
Basic Chemistry Concepts and Exercises



John Kenkel



CRC Press
Taylor & Francis Group

Basic Chemistry

Concepts and Exercises

Basic Chemistry Concepts and Exercises

John Kenkel



CRC Press

Taylor & Francis Group

Boca Raton London New York

CRC Press is an imprint of the
Taylor & Francis Group, an **informa** business

CRC Press
Taylor & Francis Group
6000 Broken Sound Parkway NW, Suite 300
Boca Raton, FL 33487-2742

© 2011 by Taylor and Francis Group, LLC
CRC Press is an imprint of Taylor & Francis Group, an Informa business

No claim to original U.S. Government works

Printed in the United States of America on acid-free paper
10 9 8 7 6 5 4 3 2 1

International Standard Book Number-13: 978-1-4398-1338-6 (Ebook-PDF)

This book contains information obtained from authentic and highly regarded sources. Reasonable efforts have been made to publish reliable data and information, but the author and publisher cannot assume responsibility for the validity of all materials or the consequences of their use. The authors and publishers have attempted to trace the copyright holders of all material reproduced in this publication and apologize to copyright holders if permission to publish in this form has not been obtained. If any copyright material has not been acknowledged please write and let us know so we may rectify in any future reprint.

Except as permitted under U.S. Copyright Law, no part of this book may be reprinted, reproduced, transmitted, or utilized in any form by any electronic, mechanical, or other means, now known or hereafter invented, including photocopying, microfilming, and recording, or in any information storage or retrieval system, without written permission from the publishers.

For permission to photocopy or use material electronically from this work, please access www.copyright.com (<http://www.copyright.com/>) or contact the Copyright Clearance Center, Inc. (CCC), 222 Rosewood Drive, Danvers, MA 01923, 978-750-8400. CCC is a not-for-profit organization that provides licenses and registration for a variety of users. For organizations that have been granted a photocopy license by the CCC, a separate system of payment has been arranged.

Trademark Notice: Product or corporate names may be trademarks or registered trademarks, and are used only for identification and explanation without intent to infringe.

Visit the Taylor & Francis Web site at
<http://www.taylorandfrancis.com>

and the CRC Press Web site at
<http://www.crcpress.com>

Contents

Preface to Faculty.	xiii
Preface to Students.	xv
Acknowledgments	xvii
Author	xix
Chapter 1 Properties and Structure of Matter.	1
1.1 What Chemists Do	2
1.2 Properties and Change	2
1.3 Classification of Matter.	4
1.3.1 Introduction	4
1.3.2 Elements and Compounds	5
1.3.3 Homogeneous vs. Heterogeneous	8
1.4 Structure and Composition of Material Substances	9
1.4.1 Structure and Composition of Elements	9
1.4.2 Structure and Composition of Compounds	10
1.5 A Quick Look inside the Atom.	13
1.6 Number of Subatomic Particles	15
1.6.1 Protons and Electrons	15
1.6.2 Neutrons and Isotopes	16
1.6.3 Atomic Weight.	18
1.6.4 Important Isotopes.	19
1.7 Homework Exercises.	20

Chapter 2	Significant Figures, the Metric System, and Dimensional Analysis.	27
2.1	Introduction.	28
2.2	Reading a Measuring Device	28
2.3	Significant Figures	30
2.3.1	Rounding	32
2.3.2	Scientific Notation	33
2.3.3	Rules for Significant Figures	34
2.4	Dimensional Analysis	34
2.5	Metric System	38
2.5.1	Length	39
2.5.2	Mass (Weight)	40
2.5.3	Volume	40
2.5.4	Examples of Conversion Problems	42
2.6	Density	45
2.6.1	Introduction	45
2.6.2	Measuring the Density of Solids That Have Regular Dimensions.	47
2.6.3	Measuring the Density of Irregularly Shaped Solids	48
2.6.4	Measuring the Density of Liquids	49
2.6.5	Using Density as a Conversion Factor	49
2.6.6	Density of Water	51
2.7	Temperature	51
2.7.1	Introduction	51
2.7.2	Temperature Scales	52
2.8	Homework Exercises.	53
Chapter 3	Names and Formulas of Compounds	63
3.1	Subclassifications of Elements	64
3.2	Formation of Ions	66
3.3	Formulas of Ionic Compounds	69
3.4	Naming Ionic Compounds	73
3.5	Other Metal Ions	75
3.6	Other Negative Ions.	76
3.7	Binary Covalent Compounds	77
3.8	Identifying and Naming Acids	78

3.9	Bases and Salts	80
3.10	Homework Exercises.	81
Chapter 4 Atomic Structure		87
4.1	Introduction.	88
4.2	History of Theories and Experiments.	89
4.3	Modern Theory of Light	92
4.3.1	What Is Energy?	93
4.3.2	What Is Light?	94
4.4	Bohr Model	97
4.4.1	Introduction	97
4.4.2	Bohr Model and Light	98
4.5	Schrodinger Model	99
4.6	Orbital Diagrams	104
4.7	Electron Configuration	107
4.8	Homework Exercises.	113
Chapter 5 Periodic Table.		119
5.1	Introduction.	120
5.2	Mendeleev and Meyer	120
5.3	Descriptive Chemistry of Selected Elements	121
5.3.1	Alkali Metals	122
5.3.2	Alkaline Earth Metals	123
5.3.3	Halogens.	124
5.3.4	Hydrogen and Helium	125
5.3.5	Carbon, Nitrogen, Oxygen, Phosphorus, and Sulfur	125
5.3.6	Aluminum and Silicon.	126
5.3.7	Selected Top-Row Transition Elements: Chromium, Manganese, Iron, Cobalt, Nickel, Copper, and Zinc	127
5.4	Periodicity and Trends: Atomic Size, Ionization Energy, and Electron Affinity	128
5.4.1	Atomic Size	130
5.4.2	Ionization Energy and Electron Affinity	131
5.5	Homework Exercises.	135

Chapter 6	Molecular Structure	139
6.1	Chemical Bonding	140
6.2	Outermost Electrons	141
6.3	Simple Ionic Compounds	142
6.4	Simple Covalent Compounds: Slot Filling	143
6.5	More Complicated Covalent Compounds	146
6.6	Coordinate Covalent Bonds	152
6.7	More Complicated Ionic Compounds	153
6.8	Structure of Ionic Compounds	155
6.9	Polar Bonds and Molecules	156
6.9.1	Electronegativity and Bond Polarity	156
6.9.2	Polarity of Molecules	158
6.9.3	Water and the Hydrogen Bond	159
6.9.4	Hydrates	161
6.10	Valence-Shell Electron-Pair Repulsion Theory	162
6.11	Molecular Orbitals	165
6.12	Homework Exercises	168
Chapter 7	Atomic Weight, Formula Weight, and Moles	173
7.1	Introduction	174
7.2	Atomic Weight Revisited	174
7.3	Molecular Weight and Formula Weight	175
7.4	Mole	177
7.4.1	Atomic and Formula Weight: Number of Grams per Mole of an Element or a Compound	177
7.4.2	Avogadro's Number: Number of Particles per Mole	179
7.5	Percent Composition of Compounds	179
7.5.1	Percent Composition from Formulas and Atomic and Formula Weights	180
7.6	Derivation of Empirical Formulas	182
7.6.1	Empirical Formulas from Percent Composition	183
7.6.2	Empirical Formulas vs. Molecular Formulas	185
7.7	Homework Exercises	186

Chapter 8	Chemical Equations and Stoichiometry	191
8.1	Introduction to Chemical Equations	192
8.2	Balancing Chemical Equations	193
8.2.1	Quantitative Information from Chemical Equations	193
8.2.2	Law of Conservation of Mass	194
8.2.3	Process of Balancing Equations	195
8.3	Optional Symbols	199
8.4	Calculations Based on Equations	201
8.4.1	Mole-to-Mole Conversions	202
8.4.2	Mole-to-Gram and Gram-to-Mole Conversions	203
8.4.3	Gram-to-Gram Conversions	204
8.5	Limiting Reactant	205
8.6	Theoretical Yield, Actual Yield, and Percent Yield	207
8.7	Homework Exercises	208
Chapter 9	Gases and the Gas Laws	213
9.1	Introduction	214
9.2	Properties of Gases	214
9.3	Kinetic–Molecular Theory of Gases	218
9.4	Gas Pressure	220
9.5	Boyle’s Law	224
9.6	Charles’ Law	227
9.7	Combined Gas Law	230
9.8	Standard Temperature and Pressure	231
9.9	Correction of a Gas Volume to STP	231
9.10	Ideal Gas Law	232
9.11	Avogadro’s Law	234
9.12	Dalton’s Law	235
9.13	Vapor Pressure	236
9.14	Homework Exercises	239
Chapter 10	Solutions	247
10.1	Introduction	248
10.2	Terminology	248

10.3	Expressing Concentration	249
10.3.1	Percent	250
10.3.2	Molarity	251
10.4	Preparing Solutions.	252
10.4.1	Dilution	252
10.4.2	Liquid Solute and Volume Percent.	256
10.4.3	Solid Solute and Weight-to-Volume Percent	257
10.4.4	Solid Solute and Molarity	258
10.4.5	Summary of Calculations	259
10.5	More Terminology	260
10.6	Increasing Dissolution Rate	263
10.6.1	Particle Size	263
10.6.2	Agitation	264
10.6.3	Temperature	264
10.7	Mechanism of Dissolving	264
10.7.1	Ionic Solutes in Water	264
10.7.2	Molecular Solutes in Water	266
10.7.3	Nonaqueous Solvents	268
10.8	Homework Exercises.	270
Chapter 11 Chemical Equilibrium		279
11.1	Introduction.	280
11.2	Chemical Equilibrium.	282
11.3	Important Examples of Chemical Equilibrium	283
11.4	Equilibrium Constant	284
11.5	Position of the Equilibrium.	287
11.6	Magnitude of the Equilibrium Constant	288
11.7	Le Chatelier's Principle.	290
11.8	Effect of Concentration Change	291
11.9	Effect of Temperature Change	293
11.10	Effect of Pressure Change.	294
11.11	Effect of Catalysts.	296
11.12	Homework Exercises.	296
Chapter 12 Acids, Bases, and pH		303
12.1	Introduction.	304
12.2	Formulas and Strengths of Acids and Bases.	304

12.3	General Properties of Acids and Bases.	308
12.4	Acidity Level: pH	310
12.4.1	Water Ionization and the Ion Product Constant.	310
12.4.2	pH.	311
12.4.3	Logarithms.	312
12.4.4	Significant Figures Rule.	315
12.4.5	pOH	315
12.4.6	Relationship between pH and pOH	316
12.4.7	Calculations Summary.	316
12.5	Measurement of pH.	318
12.6	Theories of Acids and Bases.	319
12.6.1	Arrhenius Theory	319
12.6.2	Bronsted–Lowry Theory	319
12.7	Reactions Involving Acids and Bases	321
12.7.1	Neutralization	321
12.7.2	Reaction of Acids with Metals.	322
12.7.3	Reaction of Metal and Nonmetal Oxides with Water	323
12.7.4	Reaction of Metal Oxides with Acids	324
12.8	“Add Acid to Water” Rule.	324
12.9	Homework Exercises.	324
Chapter 13 Oxidation–Reduction Reactions.		329
13.1	Oxidation–Reduction Reactions	330
13.2	Ion–Electron Method for Balancing Equations	336
13.3	Homework Exercises.	338
Chapter 14 Organic Chemicals		343
14.1	Introduction.	344
14.2	Chemistry of Carbon and Its Compounds	345
14.3	Hybridization of Orbitals	345
14.4	Geometry around Bonded Carbon Atoms	346
14.5	Physical Properties	348
14.6	Hydrocarbons	350
14.6.1	Alkanes	351
14.6.2	Isomers.	352
14.6.3	Alkenes	356

14.6.4	Alkynes	357
14.6.5	Ring Structures	358
14.7	Organic Compounds Containing Oxygen.	362
14.7.1	Alcohols	363
14.7.2	Ethers	364
14.7.3	Aldehydes	365
14.7.4	Ketones	366
14.7.5	Carboxylic Acids	367
14.7.6	Esters	368
14.8	Summaries of Hydrocarbon and Oxygen Classifications	370
14.9	Polymers	370
14.9.1	Addition Polymers	371
14.10	Homework Exercises.	376
Appendix: Answers to Homework Exercises.		383
Index		435

Preface to Faculty

I have evolved. I have evolved as a person. I have evolved as a chemistry instructor. I have evolved as an author. These statements are probably true of any chemistry instructor who has spent some 30- to 40-odd years in a teaching career as I have. Back in 1977, I was extremely fortunate to have Southeast Community College in Lincoln, Nebraska, rescue me from chemistry oblivion and offer me the job that I have loved for so long. It has been my niche, you might say, and I appreciate the powers that be at SCC and the Supreme Powers that be to have given me this opportunity so many years ago.

My current students are seeing and using the products of my evolution. I stand in front of them every day and preach my love of this fascinating natural science. Teaching is learning, and believe me when I say that the rush of adrenaline I get when I walk into my lab and classroom every day to teach and learn never gets old.

The textbook you are holding in your hands is the most recent tangible evidence of my evolution, my teaching, and my learning. It is a textbook for a one-semester preparative chemistry course. My sense of what it takes for students to be successful in such a course has become a living, breathing document—a textbook. I am so pleased to be able to share it with you.

I learned as far back as my own days as a college student that a teacher who walks into a classroom and jumps directly into the middle of a topic with no sort of introduction or review of the previous day's work loses the student right away. As a student, I was lost on many such occasions. Believe it or not, neither teaching nor writing was on my radar screen during these times, but those experiences were the start of my evolution as a teacher and textbook author. I do not want my students to be lost like I was. I do not want them to be confused. I want them to be intrigued by what they are hearing, reading, and learning. I want them to be as fascinated by the world of chemistry as

am. So I have committed to paper my approach—the approach that I believe accomplishes these things.

In these pages you will not find a lot of extravagant graphics, colorful pictures, over-the-top statements, or confusing and distracting overkill. Just the facts are presented in clear, down-to-earth, conversational language with logically assembled graphics and plenty of homework exercises to assist in the learning process. I will not say any more here. You can browse the pages and see for yourself. I know my approach will not work for everyone. What I can say is that it works for me. I have evolved, and it does indeed work for me.

Best of luck to you in your teaching endeavors.

John Kenkel
Southeast Community College

Preface to Students

I feel your pain. I feel your pain. Although I was in college what seems a lifetime ago, I know the pain you feel when you read a textbook and sense that you are reading absolute irrelevant gibberish. I feel your pain when you tell your friends that you cannot wait for this class to be in your past.

“Okay,” you ask, “So what are you doing about it?”

Well, I feel that the textbook you are holding in your hands is what I have done about it.

Admittedly, chemistry can generate fear. That is especially true when you hear your friends talking about their experiences. If your professor has chosen this textbook for your preparative chemistry course, I feel that you are on your way to an enjoyable experience. Sure we talk about atoms, molecules, covalent bonds, and ionic bonds. We talk about chemical reactions, solution chemistry, acids and bases, and gas laws. You will indeed encounter such seemingly ridiculous terms as stoichiometry, the mole, chemical equilibrium, molarity, and oxidation and reduction. You will also encounter things so incredibly small that you will wonder how scientists can even know they exist.

“I’m already overwhelmed,” you say.

“Okay,” I say, “give me a chance.”

In Chapter 1, we encounter the small things: atoms, molecules, ions, electrons, nuclei, protons, and neutrons. We look through a microscope to see them. I know, no such microscope exists...but you will encounter the “mythical microscope.” So stop worrying already. Believe me, it will help. You will especially appreciate this microscope when you see my drawings in Chapter 10.

In Chapter 2, we encounter some crazy math. But I lay it all out for you, and it will really help when we get to the math applications in Chapters 7 through 10. In Chapter 4, we see some really fascinating stuff. Trust me. And when you get into Chapter 14, you will be excitedly looking forward to your next chemistry course.

So get ready for a wild ride. There will be no bells and whistles. Only the facts. And I think that you will appreciate it.

John Kenkel

Southeast Community College

Acknowledgments

This book has evolved from my past book projects at CRC Press, especially *Chemistry: An Industry-Based Introduction with CD-ROM* published by CRC Press in 2001 and *Analytical Chemistry for Technicians*, third edition, published by CRC Press in 2003. Some of the text, figures, and tables in these books have been used in this book.

While a book project such as this is an enjoyable and satisfying experience for me, it does not happen without the assistance, cooperation, and encouragement of others. I want to thank the acquisition editor at CRC Press, Barbara Glunn, for her continued willingness to consider my work for publication and for her commitment to this project in particular. It has been a pleasure to work with her.

I want to thank those who reviewed parts of the book and offered encouragement. These kind folks include Janet Johannessen of the County College of Morris, Bob Chaffins of Ashland Community and Technical College, Tommie Ann Broome of Mississippi Gulf Coast Community College, and Marcy Molzan of Wyeth Pharmaceuticals.

I want to thank my current and past students at Southeast Community College who patiently stood by me as I developed the many drawings over the years for my PowerPoint™ slides, many of which are now used as figures in this book. My students are my greatest resource when it comes to my improvement as a teacher and an author.

I want to thank my superiors and colleagues at Southeast Community College. I have been given generous autonomy over the years in my course and laboratory curricula, and my continued enjoyment of my job is part of the reason that I am driven to write books.

I want to acknowledge and thank my family. God has graciously blessed me with a wonderful wife and three beautiful daughters. These are the most precious gifts of all, and I am eternally grateful.

Finally, my heavenly Father has blessed me with the drive, determination, and will to do this kind of work, and I express my profound gratitude for that and for the many gifts so bestowed.

John Kenkel
Southeast Community College

Author

John Kenkel is a chemistry instructor at Southeast Community College (SCC) in Lincoln, Nebraska. Throughout his 33-year career at SCC, he has been directly involved in the education of chemistry-based laboratory technicians in a vocational program presently named Laboratory Science Technology. He has also been heavily involved in chemistry-based laboratory technician education on a national scale, having served on a number of American Chemical Society (ACS) committees, including the Committee on Technician Activities and the Coordinating Committee for the Voluntary Industry Standards project. In addition to these, he has served a five-year term on the ACS Committee on Chemistry in the Two-Year College, the committee that administers the Two-Year College Chemistry Consortium (2YC₃) conferences. He was the general chair of this committee in 1996. He later served a three-year term on this committee as the industrial sponsor chair (2006–2008).

Kenkel has authored several popular textbooks for chemistry-based technician education. Three editions of *Analytical Chemistry for Technicians* have been published; the first in 1988, the second in 1994 and the third in 2003. In addition, he has authored four other books: *Chemistry: An Industry-Based Introduction* and *Chemistry: An Industry-Based Laboratory Manual* (2000–2001); *Analytical Chemistry Refresher Manual* (1992); and *A Primer on Quality in the Analytical Laboratory* (2000). All were published through CRC Press/Lewis Publishers.

Kenkel has been the principal investigator for a series of curriculum development project grants funded by the National Science Foundation's Advanced Technological Education Program, from which three of his seven books evolved. He has also authored or coauthored four articles on the curriculum work in recent issues of the *Journal of Chemical Education* and has presented this work at more than 20 conferences since 1994.

In 1996, Kenkel won the prestigious National Responsible Care Catalyst Award for excellence in chemistry teaching sponsored by the Chemical Manufacturer's Association. He has a master's degree in chemistry from the University of Texas in Austin (1972) and a bachelor's degree in chemistry from Iowa State University (1970). His research at the University of Texas was directed by Professor Allen Bard. He was employed as a chemist from 1973 to 1977 at Rockwell International's Science Center in Thousand Oaks, California.

Chapter 1

Properties and Structure of Matter

What Chemists Do	2
Properties and Change	2
Classification of Matter	4
Structure and Composition of Material Substances.	9
A Quick Look inside the Atom	13
Number of Subatomic Particles	15
Homework Exercises	20

1.1 What Chemists Do

Chemistry can be defined as the study of the composition, structure, and properties of material substances and the changes associated with these substances.

The careers of chemists and chemical laboratory technicians consist of the study of the composition, structure, properties, and changes of material substances. The fruits of their labor are many. The following items are just a few that have resulted from chemists and laboratory technicians studying chemistry: batteries, plastics and clothing fibers, pharmaceuticals and medicines, pesticides and herbicides, clean water and air, fertilizers, soaps and shampoos, and perfumes and deodorants. How do we begin a study of chemistry? We begin by understanding what is meant by each item in the definition.

1.2 Properties and Change

Let us begin with the word *change*. A **change** occurs when a material substance undergoes a transformation in appearance or substance as a result of a process imposed upon it. When water is heated, it boils and is transformed into steam. When iron is exposed to air in the presence of moisture, it rusts. When crystals of sugar are dropped into a container of water, the sugar dissolves. When gasoline burns, it is converted into carbon dioxide and carbon monoxide. In each of these cases, a transformation of some type can be observed.

There are two types of change: physical change and chemical change. **Chemical change** involves a change of substance (e.g., conversion of iron to rust, conversion of gasoline to carbon dioxide and carbon monoxide) while **physical change** is change that is purely physical, i.e., no change of substance (e.g., evaporation of water, dissolving of sugar). A change of substance means that one or more substances have been converted into something entirely different. Clearly the iron and gasoline examples represent a change of substance while the evaporation of water (change from liquid phase to gas phase) and the dissolving of sugar (change from solid state to the dissolved state) do not.

Properties of material substances are simply characteristics by which these substances may be described or identified. For example, a property of water is that it is a liquid; a property of silver is that it is shiny; a property of iron is that

it rusts; a property of a pesticide, such as a mosquito repellent, is its disagreeable odor; a property of sugar is that it tastes sweet; a property of salt is that it dissolves freely in water; and a property of the chlorophyll found in green plants is that it is green. There are two kinds of properties: physical properties and chemical properties. **Physical properties** are physical characteristics that might be determined by simple observation or measurement and involve no chemical change. Most of the examples just listed—namely color, odor, taste, physical appearance, physical state (solid, liquid, or gas), and solubility—are physical properties. All of these may be determined by simple observation with one of the five senses, such as sight (e.g., color, sheen, physical state,) or smell (e.g., odor).

Other physical properties require measurement, and some of these require calculations. Examples are melting point, boiling point, and density. Melting point, the temperature at which a solid substance is converted to a liquid (such as 32°F [0°C] for ice to water), and boiling point, the temperature at which a liquid substance boils (such as 212°F [100°C] for water) require the measurement of temperature. Because of this, these physical properties have numbers associated with them. They are physical properties because they are characteristics by which water may be described or identified through simple observation or measurement.

Density is a physical property that indicates how heavy a substance is compared with how much space it occupies (its volume). Considering two metal blocks of equal size—one made of aluminum and one made of lead—it is easy to understand the concept. The lead block is much heavier than the aluminum block in spite of the fact that the blocks have the same volume. Lead has a greater density, or greater mass or weight, per given volume. Density is something that is measured and requires a calculation to know what it is. In a chemistry laboratory, it is easy to measure the weight of a quantity of matter, and it can also be easy to measure volume. Density is calculated by dividing the measured weight by the measured volume. The density of lead is 11.3 grams per milliliter. This is a physical property of lead, a characteristic by which it may be described and identified. The density of aluminum is 2.70 grams per milliliter, and this is similarly a physical property of aluminum.

Chemical properties are characteristics that involve chemical change. A chemical property of iron is that it rusts in the presence of air and moisture. A chemical property of hydrogen is that it will combine with oxygen to form

water. A chemical property of chlorine is that it will combine with sodium to form table salt. A chemical property of gasoline is that it burns readily and is converted into carbon dioxide and carbon monoxide.

Perhaps the most useful aspect of properties, physical or chemical, is the fact that they can be used for identification purposes. Chemists and chemical laboratory technicians are sometimes faced with the task of identifying an unknown material. All material substances possess a unique set of properties; no two substances will ever be completely identical in all respects.

1.3 Classification of Matter

1.3.1 Introduction

More than 13 million material substances exist. The study of the composition, structure, properties, and changes of this large number of substances is accomplished by creating classifications and subclassifications.

The vast majority of substances as they are found in nature are not pure. In fact, they are often extremely complex mixtures. Examples are ocean water, oil, dirt, and ores. Even air and the fresh water found in lakes and wells can actually be complex mixtures. For example, ocean water contains many, many substances besides water. A large number of industrial, household, and laboratory activities are dedicated to purifying or separating mixtures, substance by substance. Examples are industrial processes, such as the refining of oil into component chemicals, household processes such as the removal of minerals from water with the use of water softeners and distillers, and laboratory processes, such as the use of a sophisticated chromatography instrument to analyze for the quantity of a pesticide residue in river water. In many of these purification schemes, extremely pure substances often result and are available as consumer products or for further industrial or laboratory processing. Thus, while complex mixtures are commonplace in our lives, the same can be said about pure substances. Classification according to purity (i.e., mixtures *vs* pure substances) is our starting point.

All matter can be classified as either a mixture or a pure substance. All pure substances can be classified as either a compound or an element. The flow diagram in Figure 1.1 is a pictorial representation of this. We will now define these terms and give additional details and examples.

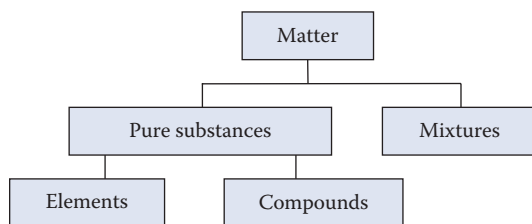


FIGURE 1.1 A classification scheme for matter based on purity.

1.3.2 Elements and Compounds

There are 109 elements in existence, and more than 13 million compounds. Many of the elements are familiar to us. Examples include various metals such as iron, copper, aluminum, and lead. Also familiar are elements that are not metals (**nonmetals**). Examples of these are helium, hydrogen, oxygen, and iodine. For convenience, all elements have been assigned **symbols**. A symbol consists of one or two letters, the first a capital letter and the second, if used, a lower case letter. These letters are derived from either the English name or the Latin name of the element. For example, *Al* is the symbol for aluminum, and *I* is the symbol for iodine. In these two examples, the symbol is either the first letter or the first two letters of the English name. The symbol *Fe* represents iron and is derived from its Latin name *ferrum*. Other examples are Cu (from *cuprum*, the Latin word for copper), Na (from *natrium*, the Latin word for sodium), and Pb (for *plumbum*, the Latin word for lead). All 109 elements are listed in a table that chemists have developed, known as the **periodic table**. The periodic table contains the names and symbols of all the elements, as well as some numeric information. This numeric information will become important to us as our study of chemistry proceeds. However, for the present discussion, let us take a look at the periodic table and especially notice the names and symbols. An example of a periodic table is presented in Figure 1.2.

The periodic table is an arrangement of small squares, each of which contain information about an element. Each square is part of a horizontal row and a vertical column. Horizontal rows are referred to as **periods** and vertical columns as **families**. The large letter or letters in the center of the squares constitute the symbol. Immediately below each symbol in the table in Figure 1.2 is the full name. In the topmost period are the squares for the elements hydrogen (symbol H) and helium (symbol He); hydrogen is on the upper left of the table

1 IA		2 IIA		Element symbol																		3 IIIB		4 IVB		5 VB		6 VIB		7 VIIB		8 VIIIB		9 IIB		10 IB		11 IIB		12 IIB		13 IIIA		14 IVA		15 VA		16 VIA		17 VIIA		18 VIIIA																																																																																																																																																																																																																																																																																																																																																																																																																																																																																																																																																																																																																																																																																															
Atomic number		Atomic weight		Element name																		Atomic number		Atomic weight		Element name		Atomic number		Atomic weight		Element name		Atomic number		Atomic weight		Element name		Atomic number		Atomic weight		Element name		Atomic number		Atomic weight		Element name		Atomic number		Atomic weight		Element name		Atomic number		Atomic weight		Element name		Atomic number		Atomic weight		Element name		Atomic number		Atomic weight		Element name		Atomic number		Atomic weight		Element name		Atomic number		Atomic weight		Element name		Atomic number		Atomic weight		Element name		Atomic number		Atomic weight		Element name		Atomic number		Atomic weight		Element name		Atomic number		Atomic weight		Element name		Atomic number		Atomic weight		Element name		Atomic number		Atomic weight		Element name		Atomic number		Atomic weight		Element name		Atomic number		Atomic weight		Element name		Atomic number		Atomic weight		Element name		Atomic number		Atomic weight		Element name		Atomic number		Atomic weight		Element name		Atomic number		Atomic weight		Element name		Atomic number		Atomic weight		Element name		Atomic number		Atomic weight		Element name		Atomic number		Atomic weight		Element name		Atomic number		Atomic weight		Element name		Atomic number		Atomic weight		Element name		Atomic number		Atomic weight		Element name		Atomic number		Atomic weight		Element name		Atomic number		Atomic weight		Element name		Atomic number		Atomic weight		Element name		Atomic number		Atomic weight		Element name		Atomic number		Atomic weight		Element name		Atomic number		Atomic weight		Element name		Atomic number		Atomic weight		Element name		Atomic number		Atomic weight		Element name		Atomic number		Atomic weight		Element name		Atomic number		Atomic weight		Element name		Atomic number		Atomic weight		Element name		Atomic number		Atomic weight		Element name		Atomic number		Atomic weight		Element name		Atomic number		Atomic weight		Element name		Atomic number		Atomic weight		Element name		Atomic number		Atomic weight		Element name		Atomic number		Atomic weight		Element name		Atomic number		Atomic weight		Element name		Atomic number		Atomic weight		Element name		Atomic number		Atomic weight		Element name		Atomic number		Atomic weight		Element name		Atomic number		Atomic weight		Element name		Atomic number		Atomic weight		Element name		Atomic number		Atomic weight		Element name		Atomic number		Atomic weight		Element name		Atomic number		Atomic weight		Element name		Atomic number		Atomic weight		Element name		Atomic number		Atomic weight		Element name		Atomic number		Atomic weight		Element name		Atomic number		Atomic weight		Element name		Atomic number		Atomic weight		Element name		Atomic number		Atomic weight		Element name		Atomic number		Atomic weight		Element name		Atomic number		Atomic weight		Element name		Atomic number		Atomic weight		Element name		Atomic number		Atomic weight		Element name		Atomic number		Atomic weight		Element name		Atomic number		Atomic weight		Element name		Atomic number		Atomic weight		Element name		Atomic number		Atomic weight		Element name		Atomic number		Atomic weight		Element name		Atomic number		Atomic weight		Element name		Atomic number		Atomic weight		Element name		Atomic number		Atomic weight		Element name		Atomic number		Atomic weight		Element name		Atomic number		Atomic weight		Element name		Atomic number		Atomic weight		Element name		Atomic number		Atomic weight		Element name		Atomic number		Atomic weight		Element name		Atomic number		Atomic weight		Element name		Atomic number		Atomic weight		Element name		Atomic number		Atomic weight		Element name		Atomic number		Atomic weight		Element name		Atomic number		Atomic weight		Element name		Atomic number		Atomic weight		Element name		Atomic number		Atomic weight		Element name		Atomic number		Atomic weight		Element name		Atomic number		Atomic weight		Element name		Atomic number		Atomic weight		Element name		Atomic number		Atomic weight		Element name		Atomic number		Atomic weight		Element name		Atomic number		Atomic weight		Element name		Atomic number		Atomic weight		Element name		Atomic number		Atomic weight		Element name		Atomic number		Atomic weight		Element name		Atomic number		Atomic weight		Element name		Atomic number		Atomic weight		Element name		Atomic number		Atomic weight		Element name		Atomic number		Atomic weight		Element name		Atomic number		Atomic weight		Element name		Atomic number		Atomic weight		Element name		Atomic number		Atomic weight		Element name		Atomic number		Atomic weight		Element name		Atomic number		Atomic weight		Element name		Atomic number		Atomic weight		Element name		Atomic number		Atomic weight		Element name		Atomic number		Atomic weight		Element name		Atomic number		Atomic weight		Element name		Atomic number		Atomic weight		Element name		Atomic number		Atomic weight		Element name		Atomic number		Atomic weight		Element name		Atomic number		Atomic weight		Element name		Atomic number		Atomic weight		Element name		Atomic number		Atomic weight		Element name		Atomic number		Atomic weight		Element name		Atomic number		Atomic weight		Element name		Atomic number		Atomic weight		Element name		Atomic number		Atomic weight			

(first column, group IA) and helium is on the upper right (last column, group VIIIA). Notice the number in the upper left corner of each square. This number is known as the **atomic number**. The atomic numbers are in numerical order, left to right, starting with hydrogen (number 1) and continuing to helium (number 2), then to the second period for lithium (number 3) and beryllium (number 4) through neon (number 10), and continuing to sodium (number 11) in the third period. The other familiar elements we mentioned can be found as follows: oxygen in the second period (number 8), aluminum in the third period (number 13), iron in the fourth period (Fe, number 26), copper in the fourth period (Cu, number 29), iodine in the fifth period (I, number 53) and lead in the sixth period (Pb, number 82).

Some elements are named after people. Examples are einsteinium, Es, number 99, (after Albert Einstein) and curium, Cm, number 96 (after Marie Curie). Others are named after places. Examples are Californium, Cf, number 98 (after California), and Europium, Eu, number 63 (after Europe). Others have rather odd names, the origins of which are not immediately obvious. Examples are praseodymium, Pr, number 59, and molybdenum, Mo, number 42.

Compounds are pure substances that are special combinations of two or more elements. The elements that are present in this combination are shown in a symbolic representation for the compound, which is known as its **chemical formula**, or simply its **formula**. The compound that is probably most familiar to everyone is water. Water is a special combination of the elements hydrogen and oxygen, which is shown in the formula for water, H_2O . Other examples of compounds and their formulas are sodium chloride (formula, NaCl), a special combination of the elements sodium (Na) and chlorine (Cl); calcium carbonate (formula, CaCO_3), a special combination of the elements calcium (Ca), carbon (C), and oxygen, (O); and potassium chromate (formula, K_2CrO_4), a special combination of the elements potassium (K), chromium (Cr), and oxygen (O).

One may ask, if compounds are combinations of two or more elements, why are they not classified as mixtures rather than compounds? The key difference between mixtures and compounds is the nature of this special combination. When two or more elements come together and form a compound, they lose their individual elemental identity and the resulting combination exhibits properties that were not exhibited before the combination. There is a complete transformation of substance, a chemical change, or a **chemical**

combination. Water is a great example. Hydrogen, a highly flammable gas, combines with oxygen, a gas which supports combustion, to produce water, a liquid that extinguishes fires. Because water does not exhibit properties of either hydrogen or oxygen, there has been a complete transformation of substance from the two elements into one, and therefore water is a compound. We can make similar statements about sodium chloride, calcium carbonate, potassium chromate, and all the other compounds that exist. The elements in a compound are **chemically combined**. The elements or substances in a mixture are not. They are simply physically mixed and retain all their properties. They can be separated by physical means, such as filtering.

The fact that there are more than 13 million compounds is intriguing. Why so many? The answer is that a very large number of different chemical combinations can result from the use of the 109 elements that exist. The composition and structure aspects of chemistry are at the heart of the concept of chemical combination. As our discussions become more and more advanced in this and later chapters, our understanding of this fascinating concept will become clearer.

1.3.3 Homogeneous vs. Heterogeneous

All pure substances, i.e., elements and compounds, are **homogeneous**. This means that their composition is identical throughout. Mixtures are not necessarily homogeneous. If the composition of a mixture is not identical throughout, then the mixture is **heterogeneous**. Examples of homogeneous mixtures are sugar water, gasoline, and stainless steel. Examples of heterogeneous mixtures are soil, a water–sand mixture, and chocolate chip ice cream. As indicated in Section 1.3.1, compounds and elements are most often found in the natural world mixed with other material substances. It is thus important for us to include mixtures in our discussion. Mixtures may be classified as either homogeneous mixtures or heterogeneous mixtures, depending on how well mixed the mixture components are. In a homogeneous mixture, the mixture components are mixed as well as they can be mixed such that you cannot tell that it is a mixture when you look at it. In a heterogeneous mixture, the mixture components are not mixed very well and you can tell that it is a mixture when you look at it. There is no chemical combination of the components of a mixture as there is with the elements that make up a compound.

1.4 Structure and Composition of Material Substances

At the end of Section 1.3.2, we indicated that composition and structure aspects of chemistry are at the heart of the concept of chemical combination. Let us now begin to explore the meaning of chemical combination by looking more closely at the structure and composition of elements and compounds.

1.4.1 Structure and Composition of Elements

Let us imagine a fictitious scenario for the analysis of the structure of material substances. Imagine that you hold in your hand a block of silver metal. Now imagine that you also have an extremely sharp cutting tool, one that is capable of cutting through metal. You proceed to cut the block of silver in half with the cutting tool. Now you have two pieces of silver in your hand, both smaller than the original. You lay one piece aside and then proceed to cut the other smaller block in half. Following this, you lay half aside and proceed to cut the other still-smaller block in half. Continuing to handle the increasingly smaller blocks and cutting them in half, you reach a point at which the block is too small to see very well. You obtain a magnifying glass and proceed to continue the halving and discarding until again you have a silver piece so small that you cannot see it adequately with the magnifying glass. So you bring out your microscope and, perhaps with a smaller cutting tool, you continue to cut and discard, cut and discard. You then reach a point at which you cannot see the speck of silver, even with the microscope set on its most powerful setting.

So, with the help of a chemist friend, you gain access to the most powerful microscope in the world, and you continue the cutting and discarding process. Finally, you get to the point at which you can absolutely go no further. At that point, you have a speck of silver barely visible through this most powerful microscope. The question in your mind is, how far would I be able to go if I could go farther? And so you start to dream and imagine that you have access to a microscope so powerful that you can see the tiniest objects that can possibly exist.

Then you resume the cutting and discarding process and you finally reach the smallest particle of silver that you can have, such that if you proceed further to divide this smallest particle, you discover that it is no longer silver. The question is, what is this smallest possible particle? The smallest particle of an element that can exist and still be that element is called an atom of that element. The **atom** is the most fundamental structural unit of all of matter, the

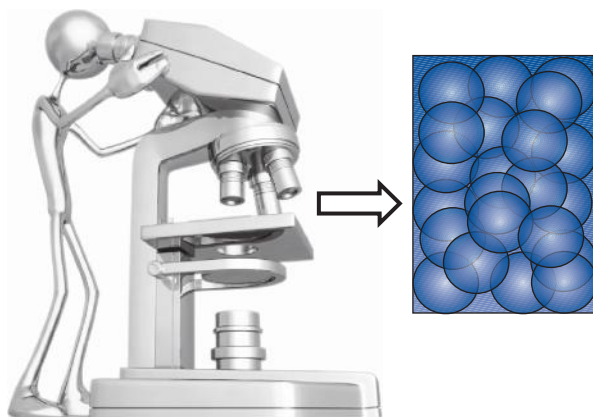


FIGURE 1.3 Looking at a sample of an element through a powerful microscope that can only be imagined, scientists believe we would see a stack of spheres that are the atoms of that element. Microscope image from Shutterstock.com.

fundamental unit of an element. Scientists most often imagine it having the shape of a sphere or ball. See Figure 1.3.

All elements are composed of atoms. The atoms of each individual element are unique to that element, which means that an atom of one element is different from an atom of any other element. However, atoms of a given element are very similar, if not identical. Chemists think of atoms as tiny, tiny spheres, or balls. One measure of this uniqueness is the size of this sphere. Hydrogen atoms are the smallest of the spheres, and atoms of francium (Fr, atomic number 87) are among the largest. Francium atoms are over nine times the size of hydrogen atoms; this is like comparing a ping pong ball to a basketball.

The utterly small sizes of all atoms, however, cannot be overemphasized. In spite of the fact that francium atoms are relatively large, it would take hundreds of thousands of them, lined up side by side, to make a line visible to the naked eye.

1.4.2 Structure and Composition of Compounds

A dividing and discarding process like that just described for the element silver may also be performed on compounds, and the result would be similar—an extremely small fundamental structural unit (visible only through our mythical microscope) that, when divided, would no longer be that compound. The situation is more complex, however, because the nature of this fundamental unit is

not the same for all compounds. Some compounds are described as **molecular** or **covalent**; others are described as **ionic**. For molecular compounds, the fundamental unit or particle is called a **molecule**. For ionic compounds, the fundamental unit is a grouping of ions that we can call the **formula unit**. An **ion** is an electrically charged atom, or grouping of atoms.

The vast majority of the 13 million compounds that exist are molecular. That fact certainly does not diminish the importance of those that are ionic, however, since there are some very common compounds in our daily lives that fall into the ionic category.

Chemical formulas, which we defined in Section 1.3.2 as symbolic designations for compounds, are used for both molecular compounds and ionic compounds. Thus we can tell what elements are chemically combined to make up a compound by looking at its formula. Examples we presented in Section 1.3.2 included H_2O (water), a combination of hydrogen (H) and oxygen (O), and NaCl (sodium chloride, or salt), a combination of sodium (Na) and chlorine (Cl). Water is an example of a molecular compound, while sodium chloride is an example of an ionic compound. Since compounds are composed of elements in chemical combination, it follows that molecules and formula units are composed of the *atoms* of the elements in chemical combination. Thus, molecules of water are composed of atoms of hydrogen in chemical combination with atoms of oxygen, and formula units of salt are composed of sodium atoms (actually ions) and chlorine atoms (ions) in chemical combination. It is the *nature* of this chemical combination that defines whether a compound is molecular or ionic. This will be detailed further as our study proceeds.

The number of atoms of each element that make up a molecule or formula unit can be seen in the formulas. The numbers written as subscripts in the formulas, such as the 2 in the formula for water, H_2O , and the 2 and the 4 in the formula for potassium chromate, K_2CrO_4 , represent the number of atoms of the elements immediately to the left of each subscript in the formula. If there is no subscript immediately to the right of an element symbol, then there is only one atom of that element in the molecule or formula unit. Thus, in a molecule of water, H_2O , there are 2 atoms of hydrogen and one atom of oxygen that are chemically combined. In a formula unit of salt, NaCl , there is one atom (ion) of each element, sodium and chlorine. In a formula unit of potassium chromate (also an ionic compound), K_2CrO_4 , there are two atoms of potassium, one atom of chromium, and four atoms of oxygen. Thus chemical formulas are reasonable representations of the molecules and formula units of which compounds are composed.

A few elements consist of two atoms that form a molecule of the element. Because of this, the symbol of these elements is actually a formula that represents this diatomic situation. These elements are hydrogen (H_2), fluorine (F_2), chlorine (Cl_2), oxygen (O_2), nitrogen (N_2), bromine (Br_2), and iodine (I_2).

Some formulas utilize parentheses. An example is the formula for calcium nitrate, $\text{Ca}(\text{NO}_3)_2$. The grouping of elements enclosed by the parentheses is an example of an ion, the nitrate ion, and it is placed in parentheses in order to maintain its identity in the formula. The 2 subscript outside the parentheses means that there are two of these ions in each formula unit of $\text{Ca}(\text{NO}_3)_2$. This formula would be equivalent to the formula CaN_2O_6 . The formula with the parentheses is preferred, however, because it gives additional information about this compound. The nitrate ion is a common grouping of atoms in which one nitrogen (N) atom is chemically bound to three oxygen atoms. Without the parentheses, this small tidbit of information would not be communicated.

We previously indicated that chemists think of atoms as tiny spheres of various sizes. Chemists think of molecules and formula units as entities that have a number of these spheres linked together or “stuck” to each other. The forces that cause them to “stick” differs, depending on whether the compound is ionic or molecular. For example, a chemist’s view of the molecules of water (H_2O) is shown in Figure 1.4. The sizes of molecules and formula units, then, are most often on the same scale as the incredibly small atoms of which they are composed.

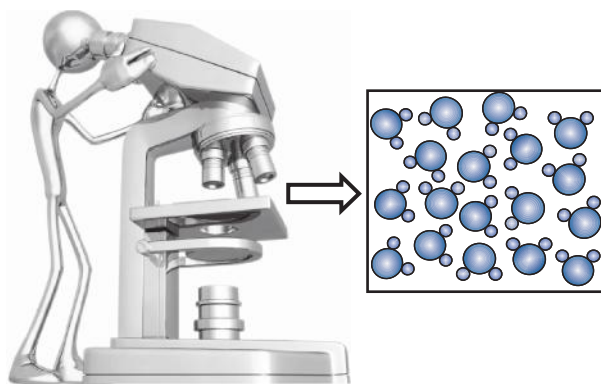


FIGURE 1.4 Looking through the mythical microscope at a sample of a molecular compound such as water, we would see the molecules of that compound. Microscope image from Shutterstock.com.

1.5 A Quick Look inside the Atom

We have stated that the atom is the smallest particle of an element that can exist and still be that element. We have also stated that if an atom is split in an attempt to make a smaller particle of an element, we would no longer have that element, but instead would have something else. We would have a collection of subatomic particles called protons, neutrons, and electrons. These are the major particles of which all atoms are composed.

Scientists discovered the existence of these particles and their properties in the late 1800s and early 1900s when some interesting experiments were performed. The modern theory describing these particles, especially the electrons, is complex and highly mathematical. In our present discussion, a very brief physical description of these particles is presented.

Let us begin with the mass of these particles. **Mass** refers to the amount of a material. This amount of a material is measured on the Earth's surface by measuring the effect of the pull of gravity on it, or in other words, its weight. References to mass in this discussion can thus be thought of as references to weight. The masses of a proton and a neutron are approximately the same. The mass of an electron is significantly less than that of a proton or neutron.

The problem with discussing the masses of these particles is that they are so tiny and weigh so little that any numerical values that we state for the mass (in grams, the metric system unit of mass or weight) would likewise be so small that they would have little meaning to us. Scientists recognized this and decided to create a new unit for mass, the **atomic mass unit**, or **amu**, and arbitrarily defined it in terms of something very small so that we could compare one particle with another. The amu is officially defined as one-twelfth the mass of a particular kind of atom of the element carbon, the so-called **carbon-12 atom**. With the amu defined this way, a proton weighs approximately 1 amu, a neutron weighs approximately 1 amu, and an electron weighs approximately 0.00055 amu.

Next, we describe the so-called electrical **charge** of these particles. The concept of electrical charge is still not well understood by scientists. However, we are familiar with the *effects* of charge. For example, lightning is the movement of charge through the atmosphere. Electrical current is the movement of charge through a conductor. The shock we feel when we touch a grounded metal object after sliding our shoes across a carpeted floor on a dry day is the movement of charge from our body to the object. This "static charge" can be a hazard in a laboratory because it can ignite flammable substances, causing fires.

In describing the charge of the subatomic particles that are part of the structure of all matter, and in order to explain their behavior, we assign each proton a charge of $+1$ and each electron a charge of -1 . They are thus equal, but opposite in their charge, which means that they attract each other, like the opposite poles of a magnet, and can neutralize each other's charge upon contact. Neutrons have no charge.

How are these particles arranged within the atom? Much of the work of the scientists in the early 1900s was focused on this question. Our modern theory of the atom states that all the protons and neutrons are located *inside* an incredibly small core situated at the center of the atom and that the electrons are located in the space *outside* this core. Obviously the core contains most of the mass of the atom, since the protons and neutrons possess most of the mass. It is a very dense agglomeration of particles, often thought of as a single particle itself. Scientists have named it the **nucleus**. To give a reasonable perspective on the nucleus in relation to the electrons, imagine that the atom is a ball with a diameter of a football field; the nucleus would then be the size of the period at the end of this sentence and the electrons would take up the remaining space. Based on this description, if we were to look through the mythical microscope at a cross section of an atom, our view would look something like that seen in Figure 1.5. See Table 1.1 for a summary of the foregoing discussion.

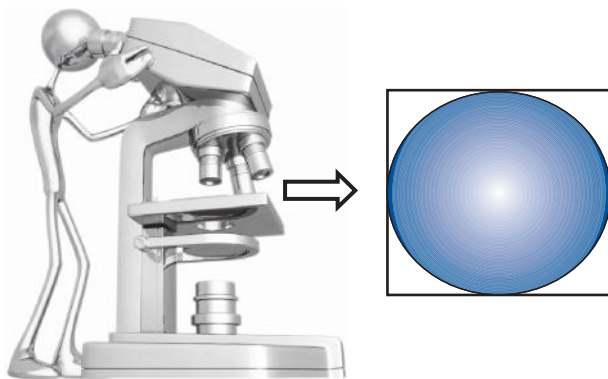


FIGURE 1.5 Looking through the mythical microscope at a cross-section of a single atom of an element, we would see that there is a tiny, dense core called the nucleus at the center where the protons and neutrons are found and space surrounding the nucleus where the electrons are found. Microscope image from Shutterstock.com.

Table 1.1 Subatomic Particles in an Atom

Particle	Mass	Charge	Location
Proton	1 amu	+1	In the nucleus
Neutron	1 amu	None	In the nucleus
Electron	0.00055 amu	-1	Outside the nucleus

1.6 Number of Subatomic Particles

One very important component of the study of the subatomic particles in atoms is the *number* of each of these particles within an atom and how one element's atoms differs from another element's atoms in terms of this number. In Section 1.4.1, we indicated that the atoms of different elements are different from each other, noting, for example, that the size of a francium atom compared with the size of a hydrogen atom is like comparing a basketball with a pingpong ball. Besides size, they also differ in terms of the number of protons, neutrons, and electrons that each has.

1.6.1 Protons and Electrons

Let us refer back to the periodic table presented in Figure 1.2. We have noted that each of the elements in this table has an identifying atomic number, which is the number located in the upper left corner of each square. For example, hydrogen is number 1, nitrogen is number 7, calcium is number 20, and gold is number 79. We can define the **atomic number** of an element as the number of protons in the nucleus of an atom of that element. All elements differ in terms of this number of protons. In fact, the atomic number, since it is unique to a given element, is a primary characteristic by which an element is identified. If you do not know the identity of an element but you know the number of protons in the nucleus (its atomic number), you can look in the periodic table to determine its identity. See Figures 1.6 and 1.7 for examples.

In an electrically neutral atom, or an atom in which the number of positive charges equals the number of negative charges, the number of protons must equal the number of electrons. Thus, for a neutral atom, the atomic number is not only the number of protons in an atom, but it is also the number of electrons. An atom may gain and lose electrons, and that is why the number of electrons is not an identifying characteristic for an element like the number of protons is. An atom cannot lose or gain protons. The particle that results from

Atomic number	17	35.453	Atomic weight
	Cl		
	Chlorine		

FIGURE 1.6 The information contained in a single element square in the periodic table (Figure 1.2) includes the element symbol and name, the atomic number, and the atomic weight. The atomic number is a whole number. The atomic weight is not a whole number.

29	63.546
Cu	
Copper	
47	107.87
Ag	
Silver	
79	196.97
Au	
Gold	

FIGURE 1.7 More examples of element squares from the periodic table. The number of protons in a nucleus is unique to each element, and so the atomic number, which is this number of protons, identifies the element. For example, if someone tells you that the atomic number of a certain element is 47, you would know that the element is silver.

the gain or loss of electrons is called an **ion**. The term *ion* was first defined in Section 1.4.2. Additional discussion will be found in Chapter 3.

1.6.2 Neutrons and Isotopes

The number of neutrons within an atom is quite variable, even for different atoms of the same element. An atom of oxygen, for example, has eight protons and eight electrons, but it can have eight, nine, or 10 neutrons. Similarly, an atom of neon has 10 protons and 10 electrons, but it can have 10, 11, or 12 neutrons. Atoms of a given element that differ because they contain a different number

of neutrons in the nucleus are called **isotopes**. The varying number of neutrons does not change the chemical identity of an element. It is oxygen or neon, for example, regardless of how many neutrons its atoms have. It also does not change the electrical neutrality because the neutron is not charged. It does, however, change the mass of the atom, since neutrons are important contributors to this mass, and it also changes a characteristic number called the mass number.

The **mass number** is the sum of the number of protons and the number of neutrons in an atom:

$$\text{Mass number} = \text{Number of protons} + \text{Number of neutrons} \quad (1.1)$$

Each isotope of a given element has the same atomic number but a different mass number.

Example 1.1

What is the mass number of an isotope of magnesium if it has 13 neutrons?

Solution 1.1

The mass number is the sum of the number of protons and the number of neutrons. The number of neutrons is given as 13. The number of protons is the atomic number of magnesium, which can be found in the periodic table. The atomic number of magnesium is 12. Therefore,

$$\text{Mass number} = 12 + 13 = 25$$

Example 1.2

Name the element described by the following: mass number = 19, number of neutrons = 10.

Solution 1.2

The atomic number (the number of protons) determines an element's identity. Since the mass number is given and the number of neutrons is given, the number of protons may be calculated:

$$\begin{aligned} \text{Number of protons} &= \text{Mass number} - \text{Number of neutrons} \\ &= 19 - 10 \\ &= 9 \end{aligned}$$

Table 1.2 Symbols for Isotopes

Element Name—Mass Number	${}^A_ZX^*$
Carbon—12	${}^{12}_6\text{C}$
Uranium—238	${}^{238}_{92}\text{U}$
Curium—239	${}^{239}_{96}\text{Cm}$
*A = mass number, Z = atomic number, X = element symbol.	

From the periodic table, we see that the element with atomic number of 9 is fluorine.

Symbols may be used to distinguish one isotope from another. One symbol that is frequently used is to write the name of the element, followed by a hyphen, and the mass number. Thus, oxygen-17 is used for the oxygen isotope with 9 neutrons, and neon-22 is used for the neon isotope with 12 neutrons. In Section 1.5, we referred to the fact that the carbon-12 isotope is used to define an amu. This would be the isotope of carbon with 6 neutrons. See the left side of Table 1.2 for more examples. The right side of this table gives yet another type of symbolism for isotopes. The superscript is the mass number and the subscript is the atomic number in this symbolism.

1.6.3 Atomic Weight

Isotopes occur in nature always in the same proportions. For example, consider boron, the element with an atomic number equal to 5. All boron atoms contain five protons, five electrons, but either five or six neutrons. Thus two isotopes of boron exist. The isotope with five neutrons constitutes approximately 20% of all boron atoms on Earth, while the isotope with six neutrons constitutes the other 80%. The source of the boron—whether it be, for example, the United States or Africa—generally makes no difference. The ratio of one isotope to the other for boron is almost always 20–80.

The mass or weight of an atom is the sum of the weights of all the particles of which it is composed: protons, neutrons, and electrons. However, since the electron's weight is negligible compared with that of the proton or the neutron, the mass or weight of an atom in amu is nearly equal to the mass number. The more neutrons there are, the heavier the atom is. Since all isotopes that occur naturally are always present in the same proportions and since attempting to

separate one isotope from another is not practical or useful in most cases, the **atomic weight** is almost always expressed as the average weight of all the isotopes that exist. This average always reflects the proportion by which these atoms are found in nature. The atomic weight of boron, for example, is 10.811, not 10 (the mass number of an atom with five neutrons), and not 11 (the mass number of an atom with six neutrons). It is also not 10.5, or the average of 10 and 11. Rather, it is 10.811, a number closer to 11 than to 10. This reflects the fact that most boron atoms on Earth (80%) have a mass number of 11, and so the average is weighted toward 11. The atomic weight for any element is never a whole number.

The atomic weight is the average weight of all the naturally occurring isotopes. *The atomic weight is found in the periodic table.* It is the number in the upper right of each square in the periodic table presented in Figure 1.2. Once again, see Figures 1.6 and 1.7 for up-close examples.

1.6.4 Important Isotopes

Probably the most important set of isotopes are those for carbon. There are three main isotopes of carbon, one with six neutrons (carbon-12, the 12 being the sum of the protons and neutrons), one with seven neutrons (carbon-13), and one with eight neutrons (carbon 14). Carbon-12 is the one which is used to define the atomic mass unit, as we have mentioned, and hence this isotope is of considerable importance. Approximately 99% of all carbon atoms are carbon-12. The carbon-14 isotope is also of considerable importance. It is used in determining the age of materials found in archeological discoveries.

The element hydrogen has three isotopes, hydrogen-1, hydrogen-2 (also known as deuterium), and hydrogen-3 (also known as tritium). Deuterium may be isolated from naturally occurring hydrogen and is used in research studies in which hydrogen atoms change positions in molecules. The movement of hydrogen atoms in chemical change processes can thus be traced.

Finally, there are a number of isotopes that are **radioactive**, such as uranium isotopes. This means that their atoms spontaneously break apart into simpler atoms with time. This breaking apart is accompanied by dangerous radiation that is given off from the materials in which these elements are found, and special precautions must be taken by technicians who work with them.

1.7 Homework Exercises

Chemistry, Properties, and Change

1. With what four aspects of matter is the science of chemistry concerned?
2. How do chemical properties of matter differ from physical properties? Give two examples of each.
3. Which of the following represent physical properties and which represent chemical properties?
 - a. Iron rusts.
 - b. Naphthalene smells like mothballs.
 - c. Potassium permanganate is purple.
 - d. The density of gold is 19.3 grams per milliliter.
 - e. Silver tarnishes.
 - f. Hydrogen reacts with oxygen to form water.
4. What is meant in chemistry by the term *physical change*? Give some examples of physical changes.
5. What is meant in chemistry by the term *chemical change*? Give some examples of chemical changes.
6. Consider the change that occurs in each of the following. Tell whether it is a physical change or a chemical change and why.
 - a. Sodium, a silvery metal, changes to a white solid in the presence of chlorine, a yellow-green gas.
 - b. A few crystals of sugar dissolve in the water in a laboratory beaker.
 - c. Alcohol, a liquid, changes to a vapor when heated.
7. Which of the following represent a chemical change and which represent a physical change?
 - a. Water is converted to steam by boiling.
 - b. Carbon, a black powder, is converted to a colorless gas when reacted with oxygen.
 - c. Salt completely dissolves in water to form a solution that looks like ordinary water.
 - d. When the chemical mercury oxide is heated in a test tube, a silvery substance condenses near the top of the test tube.
 - e. When acetic acid is cooled below 16.7°C (62°F), it is changed to a solid.

Elements, Compounds, and Mixtures

8. Tell whether each of the following is an element, a compound, or a mixture.
 - a. Water
 - b. A multicolored rock
 - c. A sample of stainless steel
 - d. CO_2
 - e. Bromine
 - f. A solution of sugar in water
 - g. Fe
 - h. Sand and water mixed together in a container
9. Fill in the blank with the appropriate letter selected from the choices that follow. Some selections may be used more than once while others not at all.

_____ Consists of two or more substances that are plainly visible.

_____ An example of this is the melting of ice.

_____ Gold.

_____ Consists of more than one substance in chemical combination.

_____ An example of this is density.

_____ An example of this is the dissolving of sugar in water.

_____ This can be found in the periodic table.

_____ Ferric chloride

_____ Consists of two or more substances that are physically mixed, but you cannot tell this by looking at it.

_____ This has a formula rather than a symbol.

 - a. Compound
 - b. Chemical property.
 - c. Homogeneous mixture
 - d. Heterogeneous mixture
 - e. Physical change
 - f. Element
 - g. Physical property
10. What is the difference between a compound and a mixture?
11. If compounds contain more than one element, why do we classify them as pure substances rather than as mixtures?

12. When two elements come together and chemically combine,
- What happens to their properties?
 - What happens to their identities?
 - What happens to their atoms?
- Give examples for each of these

Formulas, Atoms, Molecules, Formula Units, and Ions

13. What is the meaning of the subscript 2 to the right of the H and the lack of a subscript to the right of the O in the formula for water?
14. For each of the following formulas of compounds, indicate what elements (their names, not symbols) are present in the compound and also tell how many atoms of each element are present in one formula unit or molecule of the compound.
- Example: NaNO_3 : sodium, 1 atom; nitrogen, 1 atom; oxygen, 3 atoms
- Na_3PO_4
 - CCl_4
 - $\text{Al}_2(\text{SO}_4)_3$
 - $(\text{NH}_4)_2\text{CO}_3$
 - $\text{Ba}(\text{OH})_2$
15. Define each of the following terms.
- Atom
 - Molecule
 - Ion
16. State whether each of the following represents an atom, molecule, or ion.
- Se
 - CCl_4
 - HCO_3^-
 - C_6H_6
 - Pb
 - Br_2
17. What is a formula unit? How is it used in chemistry?
18. What is the difference between an atom and a molecule?

Subatomic Particles, Isotopes, and the Periodic Table

19. Define each of the following terms.

- Mass
- Atomic mass unit
- Atomic weight
- Atomic number

20. Complete the following tables.

Particle	Relative Mass	Charge	Location
Neutron			
	0.00055 amu		
			In the nucleus

Particle	Relative Mass	Charge	Location
		-1	
Neutron			
	1 amu		

21. What is an isotope? How does one isotope of an element differ from another isotope of the same element?

22. Fill in the blanks.

- The symbol for the element with atomic number 54 is _____.
- The number of electrons in a neutral atom of zinc is _____.
- The particle inside the atom that does not have a charge is the _____.
- The number of neutrons in an atom of titanium that has a mass number of 41 is _____.
- The number of protons in an atom of molybdenum is _____.
- The mass number of an atom of sodium that has 13 neutrons is _____.
- The atomic number represents the number of _____ in any atom of that element.
- The atomic number of an atom of calcium that has 18 neutrons is _____.

- i. The symbol for the element that has a mass number of 14 and a total of 8 neutrons is _____.
 - j. The particle inside the atom that has a negative charge is the _____.
23. Fill in the blanks.
- a. The symbol for the element with atomic number 47 is _____.
 - b. The number of protons in an atom of iodine is _____.
 - c. The number neutrons in an atom of chromium that has a mass number of 52 is _____.
 - d. The number of electrons in a neutral atom of calcium is _____.
 - e. The mass number of an atom of cadmium that has 47 neutrons is _____.
 - f. The atomic number of an atom of lead that has 85 neutrons is _____.
 - g. The symbol of the element that has a mass number of 27 and a total of 14 neutrons is _____.
 - h. Of atomic number, mass number, and atomic weight, the one that is not found in the periodic table is _____.
24. Fill in the blanks.
- a. The number of neutrons in an atom of bismuth that has a mass number of 210 is _____.
 - b. The atomic number of an atom of cadmium that has 56 neutrons is _____.
 - c. The symbol for the element with atomic number 35 is _____.
 - d. The symbol of the element that has a mass number of 52 and a total of 28 neutrons is _____.
 - e. The mass number of an atom of aluminum with 14 neutrons is _____.
 - f. Of atomic number, mass number and average atomic weight, the one that is *not* in the periodic table is _____.
 - g. The number of electrons in a neutral atom of an element with 13 protons and 12 neutrons. is _____.
 - h. The number of protons in an atom with atomic number 28 and a mass number of 55 is _____.
 - i. The number of protons in an atom of rubidium is _____.
25. Give the atomic number, the mass number, and the number of neutrons for each of the following:

- a. Carbon-14
 - b. Iron-58
 - c. Bismuth-209
 - d. Nitrogen-15
 - e. Chlorine-37
 - f. Sulfur-36
26. Give the atomic number, the mass number, and the number of neutrons for each of the following.
- a. $^{222}_{88}\text{Ra}$
 - b. $^{27}_{13}\text{Al}$
 - c. $^{240}_{93}\text{Np}$
 - d. $^{31}_{15}\text{P}$
 - e. $^{54}_{24}\text{Cr}$
 - f. $^{40}_{19}\text{K}$
27. An atom with a nucleus of eight protons and nine neutrons is an isotope of oxygen. Is an atom with nine protons and eight neutrons another isotope of oxygen? Explain your answer.
28. The two most common isotopes of the element chlorine are chlorine-35 and chlorine-37, but the atomic weight that is given in the periodic table for chlorine is 35.453. Explain why this atomic weight is not the same as the weight of either of these isotopes.

Chapter 2

Significant Figures, the Metric System, and Dimensional Analysis

Introduction	28
Reading A Measuring Device	28
Significant Figures	30
Dimensional Analysis	34
Metric System	38
Density	45
Temperature	51
Homework Exercises	53

2.1 Introduction

There are many, many occasions in our lives when we must make measurements on the numerous and varied systems that exist. For example, when you decide to buy a new carpet for your living room, you must measure the dimensions of the room so that you know how much carpet to buy. When a cooking recipe calls for two cups of sugar, you must be able to make this measurement so that the cookies or pies you are baking have the desired taste. When you buy an ice cream cone and it is guaranteed to have four ounces of ice cream, you would like the ice cream store clerk to have a scale so that you can be assured of getting the promised amount. When you pump air into your bicycle tire, you measure the pressure so that you know if you have pumped in the correct amount.

Chemists and chemistry laboratory technicians also make measurements, and, in fact, the accurate recording of measurements, the calculation of the results of their work, and the conclusions they draw from these measurements can be exceedingly important. The importance of the knowledge and use of measurements in a chemistry laboratory is magnified by the fact that often some very critical decisions are made based on such measurements. Examples include the measurement of an illegal drug in the urine of a race horse, the measurement of nitrate in a drinking water supply, and the measurement of the active ingredient in a pharmaceutical preparation. Often our health, our economic well-being, our integrity, indeed our very lives, depend on accurate measurements being made in a chemistry laboratory. In this chapter, the fundamentals of making and using laboratory measurements are discussed.

2.2 Reading A Measuring Device

The most fundamental rule regarding the use of any measuring device is that the device should always be read to its optimum capability with regard to precision. Precision is reflected in the number of digits in a measurement; the more digits this measurement has, the more precise it is. For example, the measurement 345.2 inches is more precise than the measurement 345 inches. If it is possible to obtain the first digit to the right of the decimal point from a device, then that digit should be recorded; if it is possible to obtain the second digit to the right of a decimal point, then that digit should be recorded, and so on. Thus, a laboratory worker should endeavor to obtain the most digits possible from whatever device is being used whether it is digital or nondigital.

This means recording all the digits shown if the device is digital and recording all digits that are known with certainty and then estimating the final digit if the device is nondigital. This statement is highlighted in Figure 2.1.

This rule is not too difficult to apply if the device is digital. Examples are shown in Figure 2.2. The rule for a nondigital device is illustrated in Figure 2.3 for a measurement one might make with an ordinary ruler. The length of the bold black line in Figure 2.3a is between 3 and 4 units. It would be expressed as “3 point something.” The 3 is known with certainty. A closer look reveals that the length is between 3.6 and 3.7, or “3 point six something.” Both the 3 and the 6 are known with certainty. Between the sixth and seventh graduation lines after the 3, there are no graduation lines. Thus, the second digit to the right of the decimal point would be the estimated digit and would perhaps be a seven. Thus, the measurement would correctly be recorded as 3.67 units.

If there were no graduation lines between the 3 and the 4, as in Figure 2.3b, then the estimated digit would be the first digit to the right of the decimal point, and the measured length would be 3.3 units. When the measurement falls exactly on a graduation line, the recorded reading must still reflect the precision of the measuring device. Thus, the correct reading in Figure 2.3c is 4.20 units, rather than 4.2 units. Similarly, the correct readings in Figure 2.3d and e are 4.0 units and 6.00 units, respectively, rather than simply 4 units or

Fundamental Rule of Measurement:

When recording a reading from a digital measuring device, record all digits shown. When recording a reading from a nondigital measuring device, write down all the digits that are known with certainty plus one digit that is estimated.

FIGURE 2.1 A fundamental rule when using any measuring device.

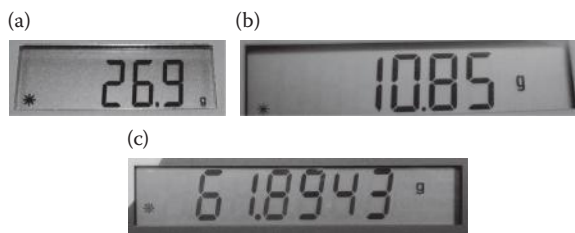


FIGURE 2.2 Example of reading on digital measuring devices. The correct readings are (a) 26.9 g, (b) 10.85 g, and (c) 61.8943 g.

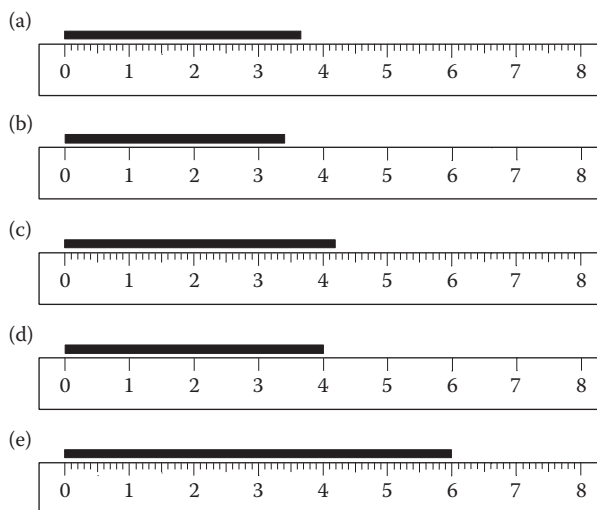


FIGURE 2.3 Several illustrations of the application of the fundamental rule of measurement for nondigital measuring devices as stated in Figure 2.1. The correct readings are (a) 3.67 units, (b) 3.3 units, (c) 4.20 units, (d) 4.0 units, and (e) 6.00 units.

6 units. Examples of other similar measuring devices with their correct readings are shown in Figures 2.4 and 2.5.

The measuring device illustrated in Figure 2.5 merits additional comment. This device is known as a **graduated cylinder** and is used for measuring volumes of liquids, such as water and water solutions. The surface of the liquid, called the **meniscus**, is curved due to the fact that the glass molecules attract water molecules with greater force than do other water molecules. The question of exactly where in the vicinity of the meniscus to take the reading becomes important. The rule is that the position of the *bottom* of the meniscus is what is read, and this is true of all volume measuring devices for liquids. Thus, the correct readings are those indicated.

2.3 Significant Figures

As stated in the last section, the number of digits in a given number indicates its precision, more digits implying greater precision. Unfortunately, there are occasions in which some of the digits in a given number are not included in the count of total digits. In other words, not all the digits found in a given

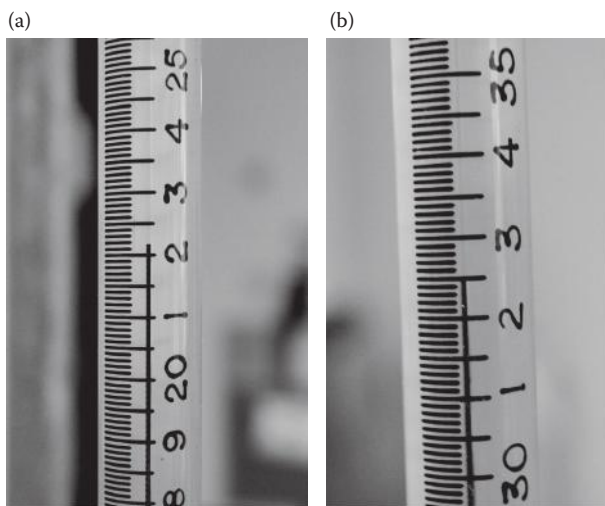


FIGURE 2.4 Additional examples of nondigital measurements, in this case classical thermometers. (a) 22.17° and (b) 32.50° .

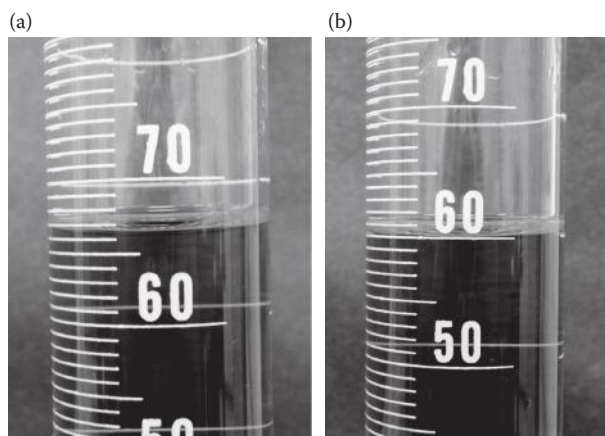


FIGURE 2.5 Additional examples of nondigital measurements, in this case, graduated cylinders. (a) 66.4 mL and (b) 60.0 mL.

number may be significant. This problem involves only zeros that may be in the number. Sometimes zeros are a part of the measurement and therefore *are* significant, but sometimes zeros are present merely to locate the decimal point and therefore are *not* significant. (In such cases, the zeros are called place-holders.) For example, if the number 430,000 is given as the population of a

city, it is not likely that all six digits reflect the true precision of the count. It is more likely that only the 4 and the 3 are significant and that the zeros are present only to locate the decimal point. The zeros in the number 0.0082 are likewise not significant because they are present only to show that the decimal point is located two places to the left of the 8. In some of these cases, the number can be expressed in a form known as **scientific notation** (described in Section 2.3.2) in order to show the significance, or lack of significance, of the placeholders.

In addition, when a number is used in a calculation and the calculator presents a long string of digits as the answer, it is important to realize that not all of these digits can be significant. A calculated answer cannot be more precise than the numbers that were used in the calculation. For example, when a calculator is used to multiply 7.51 by 149, the number found in the calculator display is 1118.99. If all the digits in this answer were counted as significant, then it would be a much more precise number than either number used in the calculation, and that is not possible. In cases such as this, a process known as rounding is used to reduce the count of digits to the proper number.

We now discuss in more detail the concepts of rounding, scientific notation, and some special rules for involving significant figures.

2.3.1 Rounding

Rounding is the removal of digits in a number when it is necessary to express the number with fewer digits. This process involves counting digits in the number, beginning from the left. When the desired number of digits has been counted, any digits further to the right must be dropped. In this process, the final counted digit may increase by one. This occurs if the set of numbers immediately to the right would indicate that the number is closer to the next higher number. For example, if 8.264 is to be rounded to 3 digits, the last digit to be counted, beginning from the left, is 6. The very next number is 4. The number 8.264 is closer to 8.26 than to 8.27, and so the correctly rounded number is 8.26. If the number to be rounded to 3 digits is 9.1793, counting from the left the last digit to be counted is the 7. Rounding to this digit gives us 9.17. However, 9.1793 is closer to 9.18 than to 9.17, and thus the correctly rounded number is 9.18.

This covers the vast majority of possibilities that may occur. The two instances that it does not cover are (1) the case in which there is just one

number to the right of the last significant digit and that number is 5, or (2) the case in which the first digit to the right is 5 and the 5 is followed by any number of zeros. For example, if the number 2.65 is to be rounded to 2 digits, we find that 2.65 is just as close to 2.6 as it is to 2.7. Another example is when 7.21500 is to be rounded to 3 digits. The number 7.21500 is just as close to 7.21 as it is to 7.22.

The rule that is most often used to solve this last problem is the even–odd rule. If the last digit to be retained is odd, then we increase it by one. If the last digit to be retained is even, then we keep it the same. Applying this rule in the these two examples, 2.65 would round to 2.6 and 7.21500 would round to 7.22. There is no theoretical principle that defines the even–odd rule. It simply means that we recognize that the original number is just as close to the higher number as it is to the lower number, and thus in half the instances we increase the last digit to be retained by one while in the other half we do not.

2.3.2 Scientific Notation

In science, we often encounter extremely large numbers and extremely small numbers. For example, in Chapter 7 we will encounter a number that has 20 zeros after the last nonzero digit before the decimal point is found. The number is

$$6022000000000000000000$$

and is known as Avogadro's number. In this type of situation, to make the number more manageable, we move the decimal point to the left such that it is immediately to the right of the nonzero first digit. Of course, each time the decimal point is moved by one digit, we are dividing the number by 10. To avoid altering the numerical value of the number, we multiply the resulting number by 10. Thus, we have the following example:

$$\begin{aligned} 723,000 &= 72,300 \times 10 \\ &= 7230 \times 100 \\ &= 723 \times 1000 \end{aligned}$$

Because $10 = 10^1$, $100 = 10^2$, and $1000 = 10^3$, the powers of 10 are used to shorten the expression of the number, as shown here

$$\begin{aligned}723,000 &= 72,300 \times 10^1 \\ &= 7,230 \times 10^2 \\ &= 723 \times 10^3\end{aligned}$$

Moving the decimal just to the right of the first digit, the correct way to express this number in scientific notation then is as follows:

$$7.23 \times 10^5$$

The exponent of the 10 in this notation is the number of places we moved the decimal point. Expressing Avogadro's number in scientific notation, we would have the following number:

$$6.022 \times 10^{23}$$

2.3.3 Rules for Significant Figures

It is important to be able to count the number of significant figures, or sig figs, in a given number. One reason that we express the answer to a calculation with the correct number of sig figs is to show how precise it is (see earlier example). Another is so that we can know how precise a measurement was when looking back some time after it was made without having to recall the nature of the measurement or the particular device that was used. Counting the sig figs in a number is not necessarily a trivial matter. The rules for counting sig figs is given in Table 2.1.

Determining the number of sig figs in the answer to a calculation requires its own set of rules. These are given in Table 2.2.

2.4 Dimensional Analysis

Have you ever calculated your car's gas mileage? If so, you know that you divide the miles driven by the gallons of gasoline used so that you have a number that is the miles per gallon for your car. This is a simple example of dimensional analysis. Knowing the units of the number that you want to calculate, you can figure out what calculation you need to do. Following is another example.

After making a measurement in a particular unit, it may be necessary to express it in some alternate unit. For example, consider homeowners who are

Table 2.1 Rules for Counting the Significant Figures in a Given Number

1. Any nonzero digit is significant. Example: 916.3 4 sig figs
2. Any zero located between 2 significant figures is significant. Example: 1208.4 5 sig figs
3. Any zero to the left of nonzero digits is not significant unless it is also covered by Rule 2. Example: 0.00345 3 sig figs
4. Any zero to the right of nonzero digits and also to the right of a decimal point is significant. Example: 34.10 4 sig figs
5. Any zero to the right of nonzero digits and to the left of a decimal point and not covered by Rule 2 may or may not be significant, depending on whether the zero is a placeholder or is actually part of the measurement. Such a number should be expressed in scientific notation to avoid any confusion. Example: 430 Unknown (4.3×10^2 or 4.30×10^2 would be better ways to express this number, depending on whether there are 2 or 3 sig figs in the number.)
<i>Note:</i> sig fig = significant figure.

installing a new carpet in their living room. They measure the length of the room and find it to be 15 feet. When they go to the carpet store they discover that they need to know the room length in yards. There are 3 feet in 1 yard. The question is, do I multiply 15 by 3 or divide 15 by 3 in order to calculate the number of yards?

While this is a very simple problem that most people may quickly respond to correctly by saying that 15 feet is 5 yards (i.e., *divide* 15 by 3), to some it may not be so simple. Chemistry students will encounter similar problems in which a more desirable method of making the decision of multiplying or dividing is important. For this reason, we use the concept of **dimensional analysis** to remove all the guesswork from this kind of problem.

At the heart of the dimensional analysis method is a technique known as the cancellation of units. The basic rule is: **when the same unit appears in both the numerator and denominator, these units “cancel.”** This cancellation of units is analogous to the cancellation of numbers when the same number appears in both the numerator and denominator.

Table 2.2 Rules for Determining the Significant Figure in Answers to Calculations

1. The correct answer to a multiplication or division problem has the same total number of sig figs as in the number with the least sig figs used in the calculation. Rounding to decrease the count of digits, or the addition of zeros to increase the count of digits, may be necessary.

Example 1: 4.3×0.882 = 3.7926 (calculator answer)
= 3.8 (answer with correct number of sig figs)

Example 2: $\frac{0.900}{0.2250}$ = 4 (calculator answer)
= 4.00 (answer with correct number of sig figs)

2. The correct answer to an addition or subtraction problem has the same number of digits to the right of the decimal point as in the number with the least such digits that is used in the calculation. Once again, rounding to decrease the count of digits, or addition of zeros to increase the count of digits, may be necessary.

Example 1: $24.992 + 3.2 = 28.192$ (calculator answer)
 $= 28.2$ (answer with correct number of sig figs)

Example 2: $772.2490 - 0.049 = 772.2$ (calculator answer)
 $= 772.200$ (answer with correct number of sig fig)

3. When several calculation steps are required, no rounding is done until the final answer is determined.

Example: $\frac{3.026 \times 4.7}{7.23}$

- = 1.9363762 (calculator answer with premature rounding)
- = 1.9 (incorrect answer due to premature rounding)
- = 1.9671093 (calculator answer correctly determined)
- = 2.0 (correct answer)

4. When both Rules 1 and 2 apply in the same calculation, follow Rules 1 and 2 in the order they are needed while also keeping Rule 3 in mind.

Example: $(3.22 - 3.034) \times 5.61$

- = 1.0659 (calculator answer with premature rounding)
- = 1.1 (incorrect answer due to premature rounding)
- = 1.04346 (calculator answer correctly determined)
- = 1.0 (correct answer)

Table 2.2 (Continued) Rules for Determining the Significant Figures in Answers to Calculations

5. Conversion factors that are exact numbers have an infinite number of sig figs.

Example: There are exactly 3 feet per yard. How many feet are there in 2.7 yards?

$$2.7 \times 3 = 8.1 \text{ (2 sig figs in correct answer, not 1)}$$

Note: sig fig = significant figure.

$$\frac{4}{5} \times \frac{21}{4} \qquad \frac{\cancel{4}}{5} \times \frac{21}{\cancel{4}}$$

For the conversion of feet to yards in the above example, we have two set-up choices, (a) and (b), as follows:

$$(a) \quad \frac{15 \text{ feet}}{1 \text{ yard}} \times \frac{3 \text{ feet}}{3 \text{ feet}} = ? \qquad (b) \quad \frac{15 \text{ feet}}{3 \text{ feet}} \times \frac{1 \text{ yard}}{3 \text{ feet}} = ?$$

In (a) we multiply by 3, while in (b) we divide by 3. The correct set-up is (b) because the unit feet appears in both the numerator and denominator (and thus cancels) and the unit yard remains in the numerator, indicating that the answer to the calculation will have yard as its unit. The answer is thus 5.0 yards.

$$\frac{15 \text{ feet}}{3 \text{ feet}} \times \frac{1 \text{ yard}}{3 \text{ feet}} = 5.0 \text{ yards}$$

It is important to remember to follow the rules of significant figures in these unit conversion problems. Notice in this case that there are *exactly* three feet per yard (infinite number of significant figures—rule 5 in Table 2.1). Thus the answer to the calculation should have the same number of significant figures as in the number 15, which is two. The correct answer is therefore 5.0 yards, not 5 yards.

An equation that gives the relationship between one unit and another is called an **equality**. The equality used in the feet-per-yard example above is shown here.

$$3 \text{ feet} = 1 \text{ yard} \qquad (2.1)$$

A **conversion factor** is one of two ratios that are derived from an equality. The following are the two possible conversion factors derived from the feet-per-yard equality.

$$\frac{1 \text{ yard}}{3 \text{ feet}} \qquad \frac{3 \text{ feet}}{1 \text{ yard}}$$

A **conversion** is a calculation in which one or more conversion factors are used to convert from one unit to another, as in the feet/yard example.

A useful approach to solving conversion problems is to write down the number, with units, that is to be converted followed by a times sign ($\times$). Leave a blank for the conversion factors (or factors) and follow this with an equals sign ($=$). To the right of the equals sign, leave a blank for the final numerical answer and follow this with the desired units. The decision of what conversion factor or factors to use then will be obvious by the fact that you can see what units are to be canceled (denominator) and what units are to be retained (numerator). The following example illustrates this procedure.

Example 2.1

How many inches are in 852 feet?

Solution 2.1

The set-up for this problem:

$$\frac{852 \text{ feet}}{\quad} \times \quad = \quad \text{inches}$$

The equality that would apply is: 1 foot = 12 inches. The two conversion factors derived from this equality are:

$$\frac{1 \text{ foot}}{12 \text{ inches}} \quad \text{and} \quad \frac{12 \text{ inches}}{1 \text{ foot}}$$

The conversion factor that would cancel the feet unit and retain the inch unit is the one on the right. Thus, we have the following:

$$\begin{aligned} \frac{852 \cancel{\text{ feet}}}{\quad} \times \frac{12 \text{ inches}}{1 \cancel{\text{ foot}}} &= 10224 \text{ inches (calculator answer)} \\ &= 1.02 \times 10^4 \text{ inches (correctly expressed answer)} \end{aligned}$$

2.5 Metric System

The measurement system that uses such familiar units as feet, yards, ounces, pounds, and gallons is known as the **English system**. Another system of measurement, the metric system, will be emphasized in this chapter and subsequent chapters. The **metric system** of measurement is used almost exclusively in all

Table 2.3 Common Prefixes Used in the Metric System and Their Meanings

Prefix	Meaning
deci-	One-tenth of
centi-	One-hundredth of
milli-	One-thousandth of
micro-	One-millionth of
kilo-	One thousand

chemistry laboratories worldwide.* The main reason for this can be summed up in one word: convenience. The entire metric system is constructed such that all units are related to other units by factors that are multiples of 10. For instance, we have the meter, which is 100 centimeters; the milligram, which is one-thousandth of a gram; and the liter, which is 1,000,000 microliters. Each measurement domain (e.g., length, mass, volume) uses prefixes to indicate what multiple of 10 is involved. Some examples of these prefixes are indicated in Table 2.3. In this chapter, we are concerned with three measurement domains: length, mass, and volume. Let us now discuss each domain and the units used in each domain.

2.5.1 Length

The base unit for length in the metric system is the **meter** (m). The meter is slightly longer than the yard in the English system and equals 39.37 inches (in). Other common metric system units of length are the centimeter (cm), the millimeter (mm), the micrometer (μm), which is also known as the micron (μ), and the kilometer (km). See Table 2.3 for the meaning of the prefixes. Equalities and conversion factors for length units compared with the meter are given in Table 2.4. One additional equality that is sometimes useful is the fact that 1 inch is equal to 2.54 cm (by definition).

* The actual system of measurement used world wide is called the International System of Units (in French, *Système International d'Unités*, or the SI system). Besides including all units that are considered metric, it has been extended to include modes of measurement that were not part of the original metric system.

Table 2.4 Units of Length with Equalities and Conversion Factors Derived from the Equalities

Meter Compared with	Equalities	Conversion Factor
Inch	39.37 in. = 1 m	$\frac{39.37 \text{ in.}}{1 \text{ m}} \quad \frac{1 \text{ m}}{39.37 \text{ in.}}$
Centimeter	100 cm = 1 m	$\frac{100 \text{ cm}}{1 \text{ m}} \quad \frac{1 \text{ m}}{100 \text{ cm}}$
Millimeter	1000 mm = 1 m	$\frac{1000 \text{ mm}}{1 \text{ m}} \quad \frac{1 \text{ m}}{1000 \text{ mm}}$
Micrometer (micron)	1,000,000 μm = 1 m	$\frac{1,000,000 \mu\text{m}}{1 \text{ m}} \quad \frac{1 \text{ m}}{1,000,000 \mu\text{m}}$
Kilometer	1000 m = 1 km	$\frac{1 \text{ km}}{1000 \text{ m}} \quad \frac{1000 \text{ m}}{1 \text{ km}}$

2.5.2 Mass (Weight)

Mass is the measure of the amount of a substance based on its weight. In other words, the measurement of the amount of a substance is done by measuring the effect of the Earth's gravitational attraction for it. This means that, technically, mass and weight are not the same. If you were to take a 180 pound man to the moon and measure his weight there, it would not be 180 pounds, yet his mass has not changed. However, weight, or the effect of the Earth's gravitational attraction for an object, is a good measure of mass because rarely would anyone measure mass anywhere except on the surface of the Earth. Thus we take mass and weight to be the same and they share the same units.

The base unit of mass in the metric system is the **gram** (g). One ounce (oz) is the same as 28.35 g. One kilogram (kg), which is 1000 g, is the same as 2.205 pounds (lb). Other common metric system units of mass are the milligram (mg) and the microgram (μg). Table 2.5 summarizes equalities and conversion factors for mass units compared with the gram.

2.5.3 Volume

Volume is the measure of the amount of space a quantity of matter occupies. The base unit for volume in the metric system is the **liter** (L). The liter is the same as 1.057 quart (qt), and 3.785 L is equal to 1 gallon (gal). Other common metric system units of volume are the milliliter (mL) and the microliter (μL).

Table 2.5 Units of Mass with Equalities and Conversion Factors Derived from the Equalities

Gram Compared with	Equalities	Conversion Factor
Ounce	$28.35 \text{ g} = 1 \text{ oz}$	$\frac{28.35 \text{ g}}{1 \text{ oz}} \quad \frac{1 \text{ g}}{28.35 \text{ g}}$
Milligram	$1000 \text{ mg} = 1 \text{ g}$	$\frac{1000 \text{ mg}}{1 \text{ g}} \quad \frac{1 \text{ g}}{1000 \text{ mg}}$
Microgram	$1,000,000 \text{ } \mu\text{g} = 1 \text{ g}$	$\frac{1,000,000 \text{ } \mu\text{g}}{1 \text{ g}} \quad \frac{1 \text{ g}}{1,000,000 \text{ } \mu\text{g}}$
Kilogram	$1000 \text{ g} = 1 \text{ kg}$	$\frac{1 \text{ kg}}{1000 \text{ g}} \quad \frac{1000 \text{ g}}{1 \text{ kg}}$

Table 2.6 Units of Volume with Equalities and Conversion Factors Derived from the Equalities

Liter Compared to	Equalities	Conversion Factor
Quarts	$1.057 \text{ qt} = 1 \text{ L}$	$\frac{1.057 \text{ qt}}{1 \text{ L}} \quad \frac{1 \text{ L}}{1.057 \text{ qt}}$
Milliliters	$1000 \text{ mL} = 1 \text{ L}$	$\frac{1000 \text{ mL}}{1 \text{ L}} \quad \frac{1 \text{ L}}{1000 \text{ mL}}$
Microliters	$1,000,000 \text{ } \mu\text{L} = 1 \text{ L}$	$\frac{1,000,000 \text{ } \mu\text{L}}{1 \text{ L}} \quad \frac{1 \text{ L}}{1,000,000 \text{ } \mu\text{L}}$

Equalities and conversion factors for volume units compared with the liter are given in Table 2.6. In addition to those units shown in Table 2.6, volumes are often also designated with cubic length units. Thus we have cubic inches, cubic feet, cubic meters, cubic centimeters, and so on. These cubic units are based on the regularly shaped object we call a cube. A cube is a solid object that has regular and equal length, width, and height dimensions. A quantity of matter that has a volume of 1 cubic foot would have a volume equal to that of a cube that is 1 foot long, 1 foot wide, and 1 foot high. If an object is a cube, or otherwise has regular length, width, and height dimensions (such as an object that is a rectangle of a certain length and width that has been moved through space to create the height dimension—that is, a rectangular prism) the volume of this object may be determined by multiplying the length, width, and height of the object together.

$$\text{volume of a cubic object} = \text{length} \times \text{width} \times \text{height} \quad (2.2)$$

The units would then be cubic length units. The abbreviations for such units are in³ for cubic inches, ft³ for cubic feet, m³ for cubic meters, and cm³ or cc for cubic centimeters. In medical applications, cc is the common unit for measuring volumes of liquid medications.

An additional convenience characteristic of the metric system is that the volume and length measurement domains are related in that one cubic centimeter, which is based on length measurement, is the same as one milliliter, which is strictly a volume unit.

$$1 \text{ cm}^3 = 1 \text{ mL} \quad (2.3)$$

This fact demonstrates that the volume domain of measurement was conceived based on the length domain, indicating that considerable thought and creativity were used in the process of setting up this system for the purpose of convenience. No similar such relationship exists within the English system because English system units were conceived in a completely arbitrary fashion. We will see examples of other metric system interdomain relationships later in this chapter.

2.5.4 Examples of Conversion Problems

Examples of English–metric and metric–metric unit conversions are given below.

Example 2.2

How many centimeters are in 5.382 m?

Solution 2.2

The set-up is shown below.

$$\frac{5.382 \text{ m}}{\quad} \times \frac{\quad}{\quad} = \quad \text{cm}$$

The required conversion factor must have cm in the numerator and m in the denominator. Thus, the correct conversion factor, from Table 2.4, is as follows:

$$\frac{100 \text{ cm}}{1 \text{ m}}$$

This gives the following result:

$$\frac{5.382 \text{ m}}{1} \times \frac{100 \text{ cm}}{\text{m}} = 538.2 \text{ cm}$$

Example 2.3

How many grams does 492.7 mg represent?

Solution 2.3

The set-up is as follows:

$$\frac{492.7 \text{ mg}}{1} \times \frac{\quad}{\quad} = \quad \text{g}$$

The required conversion has g in the numerator and mg in the denominator. From Table 2.5, we find the following:

$$\frac{1 \text{ g}}{1000 \text{ mg}}$$

Thus, we have the answer given below.

$$\frac{492.7 \text{ mg}}{1} \times \frac{1 \text{ g}}{1000 \text{ mg}} = 0.4927 \text{ g}$$

Example 2.4

How many microliters are in 0.9174 L?

Solution 2.4

The set-up is as follows.

$$\frac{0.9174 \text{ L}}{1} \times \frac{\quad}{\quad} = \quad \mu\text{L}$$

The conversion factor must have μL in the numerator and L in the denominator. This gives the following answer:

$$\frac{0.9174 \text{ L}}{1} \times \frac{1,000,000 \mu\text{L}}{1 \text{ L}} = 917400 \mu\text{L} = 9.174 \times 10^5 \mu\text{L}$$

Example 2.5

How many kilometers are represented by 8274.5 inches?

Solution 2.5

Since we have not presented a conversion factor to transform inches directly to kilometers, let us first convert to meters and then to kilometers. This will require two conversion factors. The set-up is as follows:

$$\frac{8274.5 \text{ in}}{1} \times \frac{\quad}{\quad} \times \frac{\quad}{\quad} = \quad \text{km}$$

The first conversion factor needed must convert inches to meters, while the second must convert meters to kilometers. The two conversion factors are as shown in the following:

$$\frac{1 \text{ m}}{39.37 \text{ in}} \quad \text{and} \quad \frac{1 \text{ km}}{1000 \text{ m}}$$

Thus we have the answer that follows:

$$\frac{8274.5 \text{ in}}{1} \times \frac{1 \text{ m}}{39.37 \text{ in}} \times \frac{1 \text{ km}}{1000 \text{ m}} = 0.2101727 \text{ km} = 0.2101 \text{ km}$$

Note: The number 39.37 has only four sig figs in spite of the fact that it is a conversion factor (See Rule 5 in Table 2.2). English–metric conversion factors, in general, are not exact numbers, and thus must not limit the number of sig figs in the answer to the calculation in which they are used. In this case, the answer should have four sig figs. It may be possible to find a more precise number in a handbook in order to allow the measured number, rather than the conversion factor, to limit the sig figs in the answer. One exception to this general rule is the equality $2.54 \text{ cm} = 1 \text{ in}$. This is an exact number because it defines the inch.

Example 2.6

How many cubic centimeters are in 0.339 qt?

Solution 2.6

If we are limited to the conversion factors in this text, this will require three conversion factors, one to convert from quarts to liters, one to convert from liters to milliliters and one to convert from milliliters to cubic centimeters.

$$\frac{0.339 \text{ qt}}{1} \times \frac{1 \text{ L}}{1.057 \text{ qt}} \times \frac{1000 \text{ mL}}{1 \text{ L}} \times \frac{1 \text{ cm}^3}{1 \text{ mL}} = \text{cm}^3$$

The conversion factors are as follows:

$$\frac{1 \text{ L}}{1.057 \text{ qt}} \quad \frac{1000 \text{ mL}}{1 \text{ L}} \quad \frac{1 \text{ cm}^3}{1 \text{ mL}}$$

Thus we have the answer shown as follows.

$$\frac{0.339 \text{ qt}}{1} \times \frac{1 \text{ L}}{1.057 \text{ qt}} \times \frac{1000 \text{ mL}}{1 \text{ L}} \times \frac{1 \text{ cm}^3}{1 \text{ mL}} = 320.719 \text{ cm}^3 = 321 \text{ cm}^3$$

2.6 Density

2.6.1 Introduction

In Chapter 1, we discussed physical and chemical properties of substances. We mentioned several physical properties that have numerical values associated with them and indicated that these numerical values are found by making measurements. One such property is density. Density is helpful for determining the identity of an unknown substance because it is a physical property, meaning that only one substance is likely to have a density of a particular value. Thus if that value for the unknown is determined by measurement, it can be matched to a particular known substance of that same value and identified.

Density is a measure of how heavy a given volume of a substance is. For example, if you hold in your hands a lead ball and an aluminum ball of equal size, the lead ball feels much heavier than the aluminum ball. There are more grams of lead in the lead ball than there are grams of aluminum in the aluminum ball. We say that the lead is denser, or has a greater density, than the aluminum. This means that it has a greater mass packed into the same volume. Of course, when we compare densities, the two balls must be of equal volume. An aluminum ball the size of a baseball would obviously be heavier than one the size of a small marble. This does not mean that the larger ball has the greater density. When we speak of the density of a substance, we must not only speak of its mass but also of its volume. Therefore **density** is defined as the mass of an object per unit volume. A numerical value for density of a liquid or solid may be calculated by measuring the mass and the volume of a sample and then dividing the mass by the volume.

$$\text{Density} = \frac{\text{Mass}}{\text{Volume}} \quad (2.4)$$

If D represents density, M represents mass, and V represents volume, this equation may also be written as follows:

$$D = \frac{M}{V} \quad (2.5)$$

In the metric system, the density of a liquid or a solid is most often given as the number of grams in 1 mL, or grams per milliliter. In this text, when calculating density, the mass will be expressed in grams and the volume in milliliters or cubic centimeters. The densities of some liquids and solids are given in Table 2.7.

The volumes of liquids depend on temperature. As the temperature increases, a liquid expands, meaning that the volume becomes larger even though the mass stays the same. This means that the density decreases as the temperature increases (i.e., the same number of grams are packed into a larger volume at the higher temperature). The reverse is true when the temperature decreases; the volume becomes smaller and the density increases. For this reason, the temperature must be specified when listing the densities of liquids. You will notice that the temperature is specified as 20°C (20 degrees Celsius) in Table 2.7. This is a temperature that is about the same as the temperature of a comfortable room, often called room temperature. Concepts of temperature measurement are the subject of Section 2.7.

Table 2.7 Densities of Common Liquids and Solids at 20°C		
Substance Name	Physical State	Density (g/mL)
Lead	Solid	11.34
Aluminum	Solid	2.70
Iron	Solid	7.86
Gold	Solid	19.3
Water	Liquid	0.998
Ethyl alcohol	Liquid	0.789
Carbon tetrachloride	Liquid	1.59

Example 2.7

What is the density of a solid object if it weighs 56.22 g and occupies 6.15 mL?

Solution 2.7

$$\begin{aligned} D &= \frac{M}{V} \\ &= \frac{56.22 \text{ g}}{6.15 \text{ mL}} = 9.14 \text{ g/mL} \end{aligned}$$

2.6.2 Measuring the Density of Solids That Have Regular Dimensions

Measuring the mass of a solid substance is an easy task; we simply place the material on the laboratory device for measuring mass, a laboratory balance, and read the results, either digital or nondigital. However, measuring the volume is more challenging. If the solid object is regularly shaped (i.e., if it has a shape that has regular dimensions that can be used to calculate volume), then it is a matter of measuring the dimensions and calculating the volume. Examples of shapes that have regular dimensions that can be used to calculate volume are rectangular prisms, cylinders, and spheres. Each of these has a mathematical formula for calculating the volume once the dimensions are measured. Drawings of these objects, along with the formulas for calculating their volumes, are given in Figure 2.6.

Example 2.8

What is the density of a solid object in the shape of a rectangular prism if its mass is 78.28 g and its length, width, and height are 2.11 cm, 1.86 cm, and 5.90 cm, respectively?

Solution 2.8

$$\begin{aligned} D &= \frac{M}{V} = \frac{M}{\text{Length} \times \text{Width} \times \text{Height}} \\ &= \frac{78.28 \text{ g}}{2.11 \text{ cm} \times 1.86 \text{ cm} \times 5.90 \text{ cm}} \\ &= 3.38 \text{ g/cm}^3 \end{aligned}$$

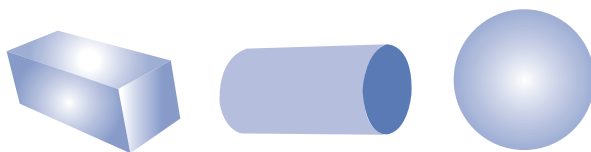


FIGURE 2.6 Left to right, a rectangular prism (Volume = length $\times$ width $\times$ height), a cylinder (Volume = $\pi r^2 h$), and a sphere (Volume = $\frac{4}{3} \pi r^3$). ($\pi = 3.1416$, r = radius, h = height).

2.6.3 Measuring the Density of Irregularly Shaped Solids

Irregularly shaped solids do not have regular dimensions and therefore cannot have their volumes calculated with a formula. Rather, we determine their volumes by water displacement (assuming that they do not dissolve in water). This means that the volume of water a solid displaces upon complete immersion is equal to the volume of the solid. In the laboratory, the process is as follows:

1. A graduated cylinder in which the solid can fit is selected.
2. A quantity of water is added (to a volume significantly less than the capacity of the cylinder but enough so that the solid can be completely immersed).
3. The volume of water in the cylinder is read.
4. The solid is immersed (being careful not to splash any water out of the cylinder and not to trap any air under or within the solid).
5. The final volume is read.

The difference between the two volumes is the volume of the solid.

Example 2.9

What is the density of an irregularly shaped piece of metal given the following data: mass = 173.03 g, water volume before immersion = 27.8 mL, water volume after immersion = 43.9 mL?

Solution 2.9

$$D = \frac{M}{V} = \frac{173.03 \text{ g}}{(43.9 - 27.8) \text{ mL}} = 10.7 \text{ g/mL}$$

2.6.4 Measuring the Density of Liquids

With solids, measuring mass is easy, but measuring the volume is a challenge requiring a calculation. With liquids, the opposite is true. It is a challenge to measure the mass, and usually requires a calculation, but measuring the volume is easy. Volumes of liquids can be conveniently measured with a graduated cylinder or some similar volume-measuring device. The mass is more difficult because the liquid must be contained while on the balance pan, and thus extra mass—the mass of the container—is included in the measurement. This procedure calls for the mass of the empty container to be determined before the liquid is introduced. After measuring the mass of the container and liquid together, the mass of the liquid is obtained by subtracting the mass of the empty container. It is convenient if the container is the same graduated cylinder in which the volume is read. Remember that volumes of liquids depend on temperature, so it is prudent to also measure the temperature of the liquid and report it along with the measured density.

Example 2.10

What is the density of a liquid if, after it is placed in a graduated cylinder that weighs 81.42 g, the liquids reaches the 48.2 mL level and the cylinder with the liquid in it weighs 133.89 g?

Solution 2.10

$$D = \frac{M}{V} = \frac{(133.89 - 81.42) \text{ g}}{48.2 \text{ mL}} = 1.09 \text{ g/mL}$$

2.6.5 Using Density as a Conversion Factor

The density of a substance is an equality from which conversion factors can be derived. For example, the density of gold (from Table 2.7) is 19.3 g/mL. In other words,

$$19.3 \text{ g of gold (Au)} = 1 \text{ mL of gold (Au)}$$

This gives us the conversion factors that can be used to convert any number of milliliters of gold to grams, or any number of grams of gold to milliliters.

$$\frac{19.3 \text{ g Au}}{1 \text{ mL Au}} \quad \text{and} \quad \frac{1 \text{ mL Au}}{19.3 \text{ g Au}}$$

In other words, density can be used to determine the volume that a certain number of grams of a substance occupies, or the mass of a certain volume of a substance. The conversion factors are as follows.

Example 2.11

What is the mass of 7.84 mL (measured at 20°C) of ethyl alcohol?

Solution 2.11

From Table 2.7, the density of ethyl alcohol at 20°C is 0.789 g/mL. The mass of 7.84 mL of ethyl alcohol is calculated as follows:

$$\frac{7.84 \text{ mL}}{1} \times \frac{0.789 \text{ g}}{1 \text{ mL}} = 6.18576 \text{ g} = 6.19 \text{ g}$$

Example 2.12

What volume will 83.224 g of iron occupy?

Solution 2.12

From Table 2.7, the density of iron is 7.86 g/mL. We calculate the volume of 83.224 g of iron as follows:

$$\frac{83.224 \text{ g}}{1} \times \frac{1 \text{ mL}}{7.86 \text{ g}} = 10.588295 \text{ mL} = 10.6 \text{ mL}$$

Sig fig note: Density is not an exact number. The three sig figs in the density Example 2.12 do limit the number of sig figs in the answer. It is possible to measure the density more precisely so that it has more sig figs or to find a more precise density in the chemical literature. In

that case, the measurement (83.224 g), and not the density, would limit the sig figs in the answer. For reference, see Rules 1 and 5 in Table 2.2.

2.6.6 Density of Water

In Section 2.5.3 we discussed a link between the length domain of measurement and the volume domain, noting that 1 cm^3 is the same as 1 mL. A link also exists between the mass and volume domains.

In Table 2.7, we see that the density of water at 20°C is equal to 0.998 g/mL . In Section 2.6.1, we note that the density of liquids changes with temperature, decreasing with increasing temperature (as the volume expands) and increasing with decreasing temperature (as the volume contracts). The density of water is equal to 1.000 g/mL at 4°C , a temperature that is near the freezing point of water. The link between the mass and volume domains then is that 1 g of water is the same as 1 mL of water at 4°C . Again, this shows the creativity and convenience that is exhibited by the metric system. Water has been used as a standard substance to create links and to set-up entire domains of measurement in the metric system.

2.7 Temperature

2.7.1 Introduction

Temperature is the measure of the hotness or coldness of a body of matter. It is *not* a measure of heat, because heat is not a measure of the hotness or coldness but of a form of energy. In this section, we will discuss the English and metric system methods of measuring temperature. Both use a device called a thermometer. Although digital thermometers and other temperature-measuring devices are in common use today, the traditional glass thermometer will be the type described here. This thermometer is typically a glass tube, closed on both ends and containing a liquid substance, which is either ethyl alcohol or mercury. The lower end is called the bulb, which contains the bulk of the liquid. It is the temperature of the matter surrounding the bulb that actually determines the reading on the thermometer.

As we learned in the last section, the volume of a liquid increases as temperature increases and decreases as temperature decreases. This fact is utilized in the design of the glass thermometer. As temperature increases, the

volume of the liquid inside increases and we observe it climbing in the tube. As temperature decreases, the volume decreases and we observe it lowering in the tube. A set of calibration marks on the tube allow us to monitor the level of the liquid in the tube, which then translates into the temperature of the matter surrounding the bulb.

2.7.2 Temperature Scales

The manner in which the positions of the graduation lines on a thermometer were determined for the English and metric systems was different. In the metric system, which measures temperature in degrees Celsius ($^{\circ}\text{C}$), water once again is used as the standard substance. The temperature at which water freezes, or the temperature of a slushy mixture of crushed ice and water, was taken to be 0°C . The temperature at which water boils (at standard atmospheric pressure) was taken to be 100°C . These two temperatures then define the magnitude of 1 degree Celsius, where each degree is one-hundredth of the length from 0 to 100 on a thermometer. This also defines the entire Celsius scale.

In the English system, we measure temperature in degrees Fahrenheit ($^{\circ}\text{F}$). The Fahrenheit degree and the Fahrenheit scale were set-up differently from those of Celsius. Water was not the standard substance for Fahrenheit. One Fahrenheit degree is very different from one Celsius degree. For example, the freezing point of water on this scale is 32° and the boiling point of water is 212° . Thus the degree interval on the Fahrenheit scale is much smaller than the degree interval on the Celsius scale.

There is another metric system scale for temperature, the Kelvin scale. By convention of the SI system of units (see footnote under Section 2.5), the Kelvin system uses neither the degree symbol ($^{\circ}$) nor the word *degree*. Thus we designate a temperature as, for example, 50 K and state it as “fifty Kelvin.” The Kelvin scale is a metric system scale, with 1 degree on the Celsius scale being equal to 1 unit on the Kelvin scale. The difference is that the Kelvin scale is an absolute scale. This means that zero on this scale is the lowest temperature that can theoretically be reached, so-called **absolute zero**. No below-zero or negative temperature readings exist on this scale. Absolute zero occurs at 273.15° below zero on the Celsius scale, or -273.15°C , and -459.67°F . See Figure 2.7.

Since 1 degree on the Celsius scale is the same as 1 unit on the Kelvin scale, the relationship between the two scales is a simple one.

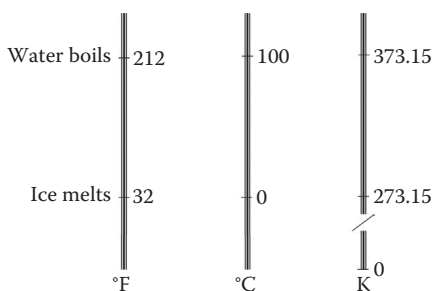


FIGURE 2.7 A comparison of three common temperature scales, Fahrenheit, Celsius, and Kelvin.

$$K = ^\circ C + 273.15 \quad (2.7)$$

or

$$^\circ C = K - 273.15 \quad (2.8)$$

On the Kelvin scale, the freezing point of water is 273.15 K and the boiling point of water is 373.15 K.

Example 2.13

What is 45.2°C expressed in Kelvin?

Solution 2.13

$$K = ^\circ C + 273.15 = 45.2 + 273.15 = 318.35 \text{ K} = 318.4 \text{ K}$$

2.8 Homework Exercises

Reading Nondigital Measuring Devices

1. Following the rules for nondigital measuring devices, give the correct reading (without units) for each of the following devices.
Figure 2.8: _____
Figure 2.9: _____
2. Again, following the rules for nondigital measuring devices, give the correct reading (without units) for each of the devices pictured in Figures 2.10 through 2.12.

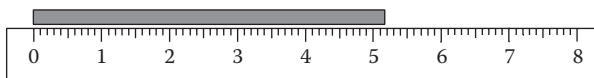
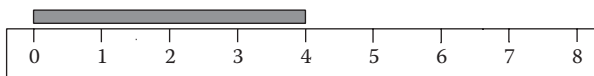
**FIGURE 2.8** The drawing for Question 1.**FIGURE 2.9** The drawing for Question 1.

Figure 2.10a: _____

Figure 2.10b: _____

Figure 2.10c: _____

Figure 2.10d: _____

Figure 2.10e: _____

Figure 2.11a: _____

Figure 2.11b: _____

Figure 2.11c: _____

Figure 2.11d: _____

Figure 2.11e: _____

Figure 2.12a: _____

Figure 2.12b: _____

Figure 2.12c: _____

Figure 2.12d: _____

Figure 2.12e: _____

Significant Figure Rules

3. How many significant figures (if known) are in each of the following numbers?
- 206
 - 210
 - 0.0540
 - 3510.0
 - 7.30×10^3
 - 0.007
 - 1830
 - 4.320
 - 690.329

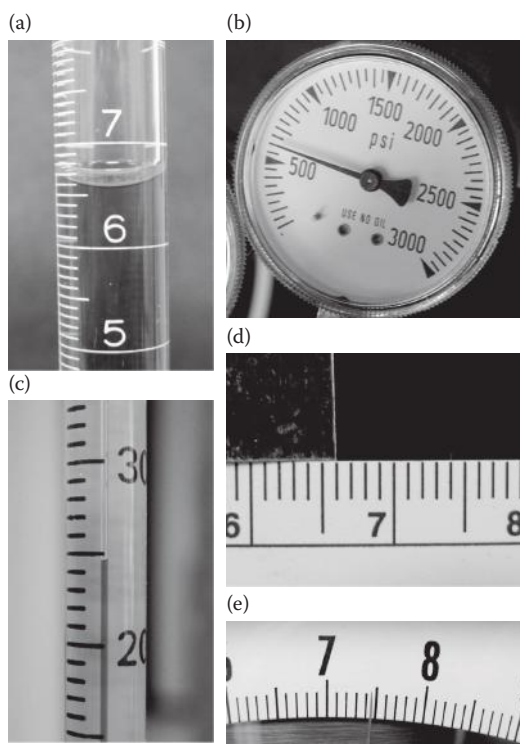


FIGURE 2.10 Photographs of the measuring devices for Question 2.

- j. 50.00
 - k. 5023.4
 - l. 0.002380
 - m. 590
 - n. 0.0345
 - o. 1.990×10^2
 - p. 607
 - q. 607.0
4. Perform the following calculations and express the answer in the correct number of significant figures.
- a. $0.00576 \div 96$
 - b. 723.5×2.19
 - c. 0.392×15.934
 - d. $98.04 + 5.6$
 - e. $28.113 - 7.1130$

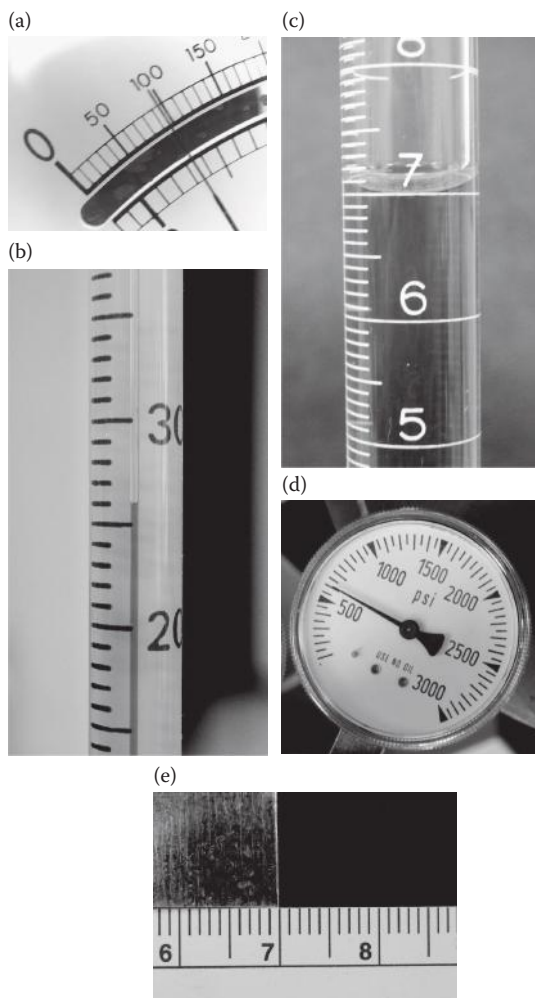


FIGURE 2.11 Photographs of the measuring devices for Question 2.

- f. $(5.330 - 5.03) \div 8.0023$
- g. $6.20 \times 10^{-3} \times 2.001 \times 10^6$
- h. 3.229×5.02
- i. $3.2929 + 8.02$
- j. 5.229×22
- k. $9.20 \times 10^3 \times 8.333 \times 10^{-9}$
- l. $8.229 \times 2.84 - 0.3890$
- m. $5.309 \times 0.88 \div 9.34384$

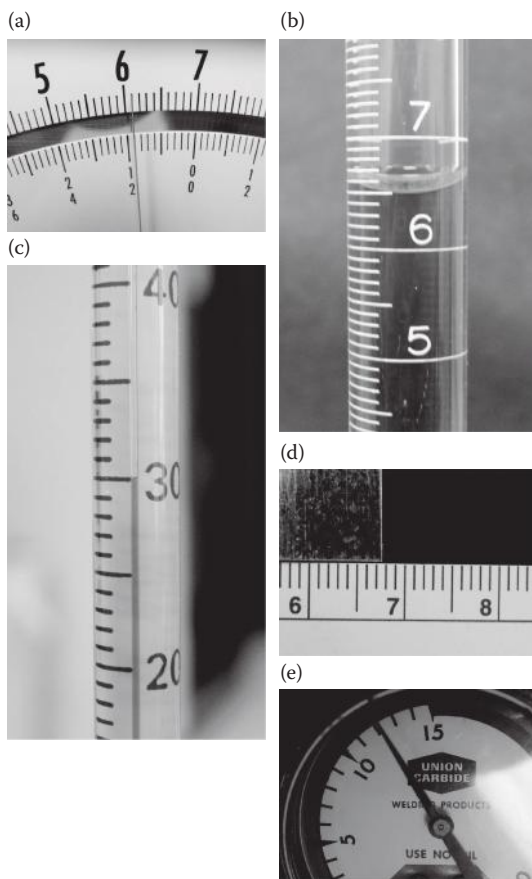


FIGURE 2.12 Photographs of the measuring devices for Question 2.

Dimensional Analysis and the Metric System

5. What is the miles per gallon in each of the following? Be sure to follow significant figure rules in expressing your answer.
 - a. You drove 348 miles on 13.5 gallons of gasoline
 - b. You drove 4.5×10^2 miles on 15.0 gallons of gasoline
 - c. You used 18.50 gallons of gasoline and drove 592.0 miles
6. If you see the following on a cereal box, how many ounces are there per gram? How many grams are there per ounce? Be sure to follow significant figure rules in expressing your answer
 - a. 13 oz (368 g)

- b. 14.0 oz (396 g)
 - c. 42 oz (1.19 kg)
7. If you see the following on a consumer product, how many milliliters are there per fluid ounce? How many fluid ounces are there per milliliter? Be sure to follow significant figure rules in expressing your answer.
- a. 16 fl.oz (473 mL)
 - b. 8 fl.oz (236 mL)
 - c. 48.0 fl.oz. (1.42 L)
8. Make the following conversions, expressing the answer with the correct number of significant figures.
- a. 2.27 kg to grams
 - b. 6224 mm to meters
 - c. 0.482 L to cubic centimeters
 - d. 76.7 cm to meters
9. Perform the calculations required in each of the following and express the answers in the correct number of significant figures.
- a. How many grams are represented by 0.338 kg?
 - b. How many liters are represent by 0.23 mL?
 - c. How many centimeters are in 5.4 m?
 - d. Convert 0.003380 kg to milligrams.
10. Fill in the blanks in the following table, expressing each to the proper number of significant figures.

	Centimeters	Meters	Kilometers
a.	49.5 cm		
b.		3.22 m	
c.			0.000048 km
d.	7.9×10^2 cm		
e.		341 m	
f.			2.303×10^{-2} km
g.	4.145×10^4 cm		
h.		7.0×10^2 m	
i.			0.018 km

11. Fill in the blanks in the following table, expressing each to the proper number of significant figures.

	Micrograms	Grams	Kilograms
a.	$6.5 \times 10^4 \mu\text{g}$		
b.		0.27 g	
c.			0.00025 kg
d.	45.21 μg		
e.		94.70 g	
f.			0.03700 kg
g.	$8.00 \times 10^6 \mu\text{g}$		
h.		0.00420 g	
i.			$6.70 \times 10^{-5} \text{ kg}$

12. Fill in the blanks in the following table, expressing each to the proper number of significant figures.

	Milliliters	Liters	Microliters
a.	65 mL		
b.		0.038 L	
c.			$9.0 \times 10^4 \mu\text{L}$
d.	$3.40 \times 10^2 \text{ mL}$		
e.		0.510 L	
f.			$1.20 \times 10^6 \mu\text{L}$
g.	8604 mL		
h.		2.000 L	
i.			1904 μL

Density

13. Calculate the density of an object in grams per milliliter (g/mL) given the following information about the object. Be sure to follow significant figure rules in expressing your answer.
- 18.9 mL of it weighs 32.85 g.
 - It weighs 56.0 g and occupies 4.00 cm^3 volume.
 - It weighs 228.6 g and occupies 38.1 mL volume.

14. Given the following information about a rectangular prism, what is its density expressed in grams per milliliter (g/mL)? Be sure to follow significant figure rules in expressing your answer.
- Mass = 1499.37 g, length = 6.05 cm, width = 5.22 cm, height = 18.93 cm.
 - Mass = 32.95 g, length = 5.6 cm, width = 2.10 cm, height = 0.495 cm.
 - Mass = 6.7×10^7 mg, length = 0.098 m, width = 3.19 cm, height = 775 cm.
15. Given the following data for an irregularly shaped solid, what is the object's density?
- Mass = 54.30 g, volume of water in a graduated cylinder before immersion of the solid = 43.6 mL, volume of water in this graduated cylinder after immersion of solid = 51.2 mL.
 - Mass = 6.22×10^2 g, volume of water in a graduated cylinder before immersion of the solid = 13.00 mL, volume of water in this graduated cylinder after immersion of solid = 81.92 mL.
 - Mass = 34.6609 g, volume of water in a graduated cylinder before immersion of the solid = 3.00 mL, volume of water in this graduated cylinder after immersion of solid = 9.46 mL.
16. In determining the density of a liquid, the following measurements are made. Calculate the density of the liquid. Be sure to follow significant figure rules in expressing your answer.
- An empty graduated cylinder weighs 28.64 g; the liquid is added so that the volume reading is 35.4 mL; after this the graduated cylinder weighs 57.56 g.
 - An empty graduated cylinder weighs 45.0 g; the liquid is added so that the volume reading is 134.5 mL; after this the graduated cylinder weighs 257.5 g.
 - An empty graduated cylinder weighs 692 g; the liquid is added so that the volume reading is 98 mL; after this the graduated cylinder weighs 947 g.
17. Convert 356 mL of a liquid to grams knowing that the density of the liquid is 1.27 g/mL.
18. What is the volume of a 33.60-g piece of gold if the density of gold is 19.3 g/mL?

19. If milk has a density of 1.030 g/cm^3 and you need 21 g of milk, how many milliliters of milk are needed?
20. What would a chunk of rock weigh in grams if it is known that it occupies 1.459 L and has a density of 8.754 g/mL ?
21. Fill in the blanks in the following table, expressing each to the proper number of significant figures.

	Density	Mass	Volume
a.	11.3 g/mL	0.7734 g	
b.	2.70 g/mL		7.002 mL
c.		34.09 g	5.2 mL
d.	7.86 g/mL	4.2309 g	
e.	11.34 g/mL		7.50 mL
f.		0.54 kg	0.27 L
g.	11.34 g/mL	68.04 g	
h.	1.59 g/mL		$6.00 \times 10^{-1} \text{ L}$
i.		$3.737 \times 10^5 \text{ mg}$	$4.382 \times 10^{-2} \text{ L}$
j.	0.789 g/mL	$8.9 \times 10^5 \text{ mg}$	

Temperature Unit Conversions

22. a. What is 23.18°C expressed in Kelvin?
b. What is 401.3 K expressed in degrees Celcius?
23. a. Convert 342 K to degrees Celcius.
b. Convert 92°C to Kelvin.
24. Fill in the blanks in the following table, expressing each to the proper number of significant figures.

	$^\circ\text{C}$	K
a.	58	
b.		334.68
c.	46.7	
d.		403.4
e.	123.45	
f.		384

Chapter 3

Names and Formulas of Compounds

Subclassifications of Elements	64
Formation of Ions	66
Formulas of Ionic Compounds	69
Naming Ionic Compounds	73
Other Metal Ions	75
Other Negative Ions	76
Binary Covalent Compounds	77
Identifying and Naming Acids	78
Bases and Salts	80
Homework Exercises	81

3.1 Subclassifications of Elements

We mentioned in Section 1.3 that chemists simplify the task of studying the enormous quantity of different material substances that exist by placing them into classifications and subclassifications. Let us now take another look at the periodic table and see what subclassifications chemists have used for the elements.

The arrangement of the elements in the periodic table may seem odd. In Chapter 4, the scientific reasons for this arrangement will become clear. However, the following discussion will make it apparent that they are grouped according to their properties. This is especially true of the vertical columns, such as Groups IA (or Group 1), IIA (or Group 2), and so on. The properties of the elements within a vertical column are so similar that vertical columns of elements have come to be known as **families**. Before we discuss some individual families, let us look at a few larger groupings.

The majority of elements exhibit the properties of **metals**. These properties are metallic sheen, electrical conductivity, malleability (ability to be flattened into foil), ductility (ability to be formed into wires), and high tensile strength (ability to withstand forces that would pull it apart). Atoms of metals tend to lose electrons to form positively charged ions. Metals are located to the left of the bold stairstep line that is shown extending diagonally across the right side of the periodic table in Figure 3.1 (see also Figure 1.2).

Two other classifications are the **nonmetals** and the **metalloids** (see Figure 3.1). Nonmetals exhibit none of the properties of metals. Many nonmetals are gases, and one (bromine) is a liquid. Those that are solids (e.g., carbon, phosphorus, and sulfur) are electrical insulators, do not have a metallic sheen, are neither ductile nor malleable, and do not have a high tensile strength. Nonmetals tend to take on electrons and form negatively charged ions. Nonmetals are located to the right of the bold diagonal line in the periodic table in Figure 3.1. Some of the elements positioned adjacent to the bold dividing line exhibit properties of both metals and nonmetals. These are the metalloids. The metalloids are boron, silicon, germanium, arsenic, antimony, tellurium, polonium, and astatine. These are shown within the dark-shaded region of Figure 3.1.

Another classification scheme divides the periodic table into three parts: the **representative** elements, the **transition elements** (also known as the transition metals), and the **inner transition elements** (also known as the inner transition metals, or rare earth elements). These divisions are shown in Figure 3.2. The

1																		18	
IA	2													13	14	15	16	17	VIIIA
	IIA													IIIA	IVA	VA	VIA	VIIA	
		3	4	5	6	7	8	9	10	11	12								
		IIIB	IVB	VB	VIB	VIIB	VIII B			IB	IIB								
													</						

FIGURE 3.1 An outline of the periodic table showing the locations of the metals (white squares), the nonmetals (light gray squares), and metalloids (dark gray squares). The stair-step dividing line referred to in the text is also shown.

inner transition elements are set apart from (below) the rest of the elements. The two horizontal rows (**periods**) that constitute this group were actually removed from the last two rows of the transition elements to conserve space. If this were not done, the table would be needlessly wide. Again, the scientific reasons for these elements being grouped together will become clear in Chapter 4. Periods are numbered from the top. We have period 1 (the top horizontal row, which consists of hydrogen and helium), period 2 (the second horizontal row from the top, which consists of lithium, beryllium, and so on) and so forth.

Finally, four of the families of elements have been given names that deserve to be mentioned here. The metals in Group IA (or Group 1) are known as the **alkali metals**. These include lithium, sodium, potassium, rubidium, cesium, and francium (Figure 3.3; also see Figure 1.2). Hydrogen, a Group IA (Group 1) element, is excluded from this list. It is a gas and considered to be a nonmetal. The scientific reasons for assigning it to this group will become clear as our discussion progresses.

The metals in Group IIA (or Group 2) are known as the **alkaline earth metals**. These include beryllium, magnesium, calcium, strontium, barium,

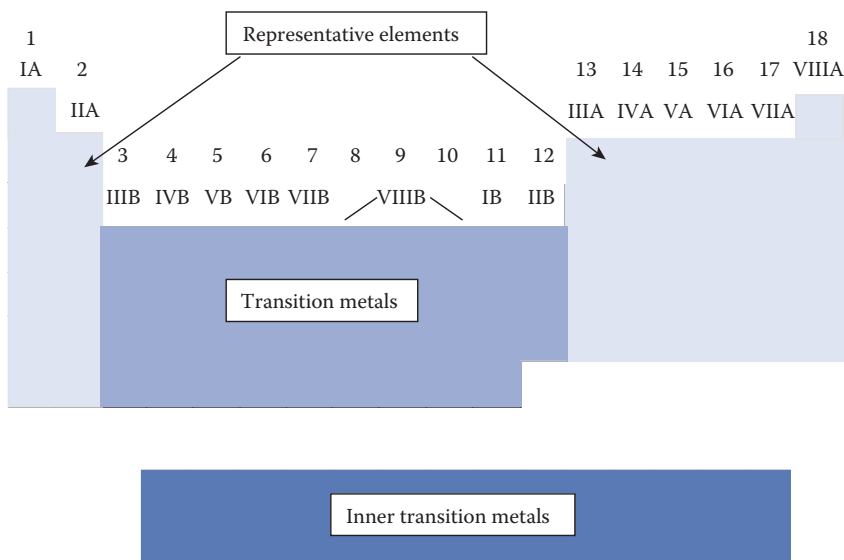


FIGURE 3.2 An outline of the periodic table showing the locations of the representative elements, the transition elements, and the inner transition elements.

and radium. The nonmetals in Group VIIA (or Group 17) are known as the **halogens**. These include fluorine, chlorine, bromine, iodine, and astatine. The elements in Group VIIIA (or Group 18) are known as the **noble gases**. These include the nonmetals helium, neon, argon, krypton, xenon, and radon. Refer to Figures 3.3 and 1.2 to clearly see the positions of these groups and these elements in the periodic table.

3.2 Formation of Ions

As indicated earlier, atoms are capable of gaining or losing electrons so that charged particles called ions form. Ions are charged because they have an imbalance of protons and electrons, which are oppositely charged particles. It is possible for single atoms to develop a charge, but it is also possible for groups of atoms to collectively develop a charge. When single atoms gain or lose electrons, a **monatomic ion** forms. When a collective group of atoms has a charge, it is called a **polyatomic ion**. The prefix *mono-* refers to one atom, and the prefix *poly-* refers to many (two or more) atoms. The number of atoms that a polyatomic ion has is variable. Some examples are discussed here.

Alkali metals												Noble gases						
1	2											13	14	15	16	17	18	
IA	IIA											IIIA	IVA	VA	VIA	VIIA	VIIIA	
		3	4	5	6	7	8	9	10	11	12							
		IIIB	IVB	VB	VIB	VII B	VIII B		IB	IIB								

FIGURE 3.3 An outline of the periodic table showing the locations of the alkali metals, the alkaline earth metals, the halogens, and the noble gases.

The gain or loss of electrons by an atom is not a random event. An atom gains or loses one or more electrons to achieve a more stable state—a state in which the atom possesses a certain number of electrons. The most stable of states is exemplified by the noble gases. Thus the number of electrons representing stable states is the number of electrons found in the atoms of noble gases, for example, 2 (helium), 10 (neon), and 18 (argon). Atoms of elements that have an electron count close to those of the noble gases will gain or lose one or more electrons (and become a monatomic ion) in order to have the same count as the nearest noble gas. Because a metal is located on the left side of the periodic table, it tends to lose one or more electrons to obtain an electron count equal to the noble gas that precedes it. Because a nonmetal is located on the right side of the periodic table, it tends to gain one or more electrons to obtain an electron count equal to the noble gas to its right (see Figure 3.4). Thus, given the opportunity, chlorine, for example, will gain an electron to become like argon and sodium will lose an electron to become like neon. The

1 or IA 2 or IIA		16 or VIA 17 or VIIA 18 or VIIIA		
Li	Be	O	F	He
Na	Mg			Ne
K	Ca	S	Cl	Ar
Rb	Sr			Kr
Cs	Ba	Br	I	Xe
Fr	Ra			Rn

FIGURE 3.4 A close-up look at the alkali metals (Group 1 or IA), alkaline earth metals (Group 2 or IIA), oxygen and sulfur (Group 16 or VIA), halogens (Group 17 or VIIA) and noble gases (Group 18 or VIIIA). An atom will gain electrons (halogens, oxygen, and sulfur) or lose electrons (alkali metals and alkaline earth metals) so that it has the same electron count as the nearest noble gas.

result is a chloride ion, an ion with a -1 charge, because the chlorine atom has gained an electron (symbolized as Cl^-), and a sodium ion, an ion with a $+1$ charge, because a sodium ion has lost an electron (symbolized as Na^+). All alkali metals tend to lose one electron and become a $+1$ ion (e.g., Li^+ , Na^+ , K^+), and all halogens tend to gain one electron and become a -1 ion (e.g., F^- , Cl^- , Br^-). Additionally, all alkaline earth metals (e.g., Be, Mg, Ca) tend to lose two electrons to achieve the electron count of the nearest noble gas. Thus, we have monatomic ions such as Be^{2+} , Mg^{2+} , and Ca^{2+} . Similarly, the monatomic ions O^{2-} and S^{2-} form due to the gain of two electrons by the elements oxygen and sulfur to become like their nearest noble gas. For all these elements, *the number of electrons that can be gained or lost by an atom is equal to the number of families the element is away from a noble gas.*

The number of electrons that transition metal and other metal atoms can lose when forming monatomic ions is not as easily predicted. The number

Table 3.1 The “Big Six” Polyatomic Ions

Formula	Name
NO_3^-	Nitrate
SO_4^{2-}	Sulfate
PO_4^{3-}	Phosphate
CO_3^{2-}	Carbonate
OH^-	Hydroxide
NH_4^+	Ammonium

of electrons that can be lost varies from one to six. Thus the charges on such ions vary from +1 to +6. In many cases, a given transition or inner transition metal can form two or more different monatomic ions. For example, the two monatomic ions of iron are Fe^{2+} and Fe^{3+} .

The formation of polyatomic ions is also not random, but the explanation as to why they form is complex and will not be addressed here. However, it is important to know some of the common polyatomic ions that do exist. Table 3.1 presents the formulas and names of six of the most common ones. For the purpose of future reference, let us call these the Big Six polyatomic ions. Others will be discussed later.

3.3 Formulas of Ionic Compounds

As we discussed in Chapter 1 (Section 1.4.2), all of the more than 13 million chemical compounds that exist consist of elements in chemical combination. Compounds that are covalent or molecular have molecules as their fundamental particles. Compounds that are ionic have formula units as their fundamental particles. Molecules are composed of various numbers of different atoms in covalent chemical combination; formula units are composed of various numbers of different ions.

Ionic and **covalent** are words that describe the nature of the chemical combination involved. In this chapter, we examine these concepts in more detail in order to characterize the nature of chemical combination and the reasons why elements come together to form compounds in the ways that they do. We also look at the schemes for naming common compounds and at the

symbolism used for expressing chemical change in the formation and destruction of compounds.

Positively charged ions are called **cations** (pronounced cat-eye-uns) and negatively-charged ions are called **anions** (pronounced ann-eye-uns). Cations and anions come together to form formula units because oppositely charged ions attract each other electrostatically. Cations attract anions, and vice versa, in a manner similar to the way in which opposite poles of a magnet attract each other. If a cation has a charge of +1 (as opposed to +2 or +3), it attracts and exactly neutralizes one anion that has a charge of -1 . With these two ions electrostatically “stuck” to each other, the result is a formula unit that has neither a positive charge nor a negative charge, but is neutral, or has no net charge. An ionic compound, with a set formula, has thus formed. For example, one sodium ion (Na^+) and one chloride ion (Cl^-) by virtue of their equal but opposite charges, attract each other, forming a formula unit. The result is a compound with the formula NaCl .* A similar one-to-one combination occurs between any cation with a charge of +1 and any anion with a charge of -1 . Examples include an ammonium ion or any alkali metal ion (+1) combined with a nitrate ion, a hydroxide ion, or any monatomic ion formed from a halogen atom (-1). These latter ions, called the halide ions include fluoride (F^-), chloride (Cl^-), bromide (Br^-), and iodide (I^-). We thus have compounds with such formulas as LiBr , KI , NaNO_3 , NH_4F , and NH_4OH .

One-to-one combinations also occur when the cation has a +2 charge and the anion has a -2 charge, since such a combination would also result in a neutral or uncharged unit. Examples of this include any alkaline earth metal ion (+2) combined with a sulfate ion and carbonate ions, and also an oxide ion (O^{2-}), or any other monatomic ion of the oxygen family (Group VIA or 16), such as the sulfide ion (S^{2-}). Thus we have compounds with such formulas as CaO , MgSO_4 , and BaCO_3 and corresponding formula units.

Two-to-one combinations occur when the cation has a +1 charge while the anion has a -2 charge. This is because two of the +1 charges are required to neutralize the -2 charge. Examples of this include an ammonium ion or any alkali metal ion combined with a carbonate ion, a sulfate ion, an oxide ion, or any other monatomic ion of the oxygen family. This gives us compounds with such formulas as Na_2SO_4 , K_2CO_3 , and $(\text{NH}_4)_2\text{S}$ and the corresponding formula

* The positive ion is always listed first in the formula of an ionic compound. Thus, it is NaCl , in this example, and not ClNa .

units. The reason for using parentheses in a formula now becomes a bit clearer (see Chapter 1, Section 1.4.2). A formula with parentheses highlights the polyatomic ion that is present in the formula unit.

One-to-two combinations occur when the cation has a +2 charge while the anion ion has a -1 charge. In this case, two of the -1 charges neutralize the +2 charge. Examples of this include any alkaline earth metal ion combined with a nitrate ion, a hydroxide ion, or any halide ion. Thus we have compounds with such formulas as $\text{Ca}(\text{NO}_3)_2$, $\text{Mg}(\text{OH})_2$, and BaCl_2 , and the corresponding formula units. The phosphate ion has a -3 charge. Also, an aluminum atom has three electrons beyond that of neon and therefore would form a +3 ion. The compound formed between aluminum and a phosphate ion would thus consist of a one-to-one combination, AlPO_4 , since that is what would produce a neutral unit. Similarly, aluminum combined with a -1 ion would produce a one-to-three unit, and a +1 ion combined with a phosphate ion would produce a three-to-one unit. Thus we have compounds with such formulas as AlCl_3 , $\text{Al}(\text{NO}_3)_3$, Na_3PO_4 , and $(\text{NH}_4)_3\text{PO}_4$ and the corresponding formula units.

The existence of +3 ions and -3 ions such as the Al^{3+} ion and the PO_4^{3-} ion means that combinations of these ions with -2 and +2 ions, respectively, present some difficulties. For example, what combination of Al^{3+} with the sulfate ion, SO_4^{2-} , or an alkaline earth metal ion with the phosphate ion, PO_4^{3-} , produces a neutral unit? In the case of the aluminum ion and sulfate ion, it takes two aluminum ions and three sulfate ions since the net charge from the aluminum ion is +6 and the net charge from the sulfate is -6. Thus we have the formula $\text{Al}_2(\text{SO}_4)_3$ and the corresponding formula unit. In the case of an alkaline earth metal ion and the phosphate ion, it takes three alkaline earth metal ions and two phosphate ions since the net charge from the alkaline earth metal ion is +6 and the net charge from the phosphate is -6. Thus we have the formula $\text{Mg}_3(\text{PO}_4)_2$ and the corresponding formula unit.

It may be noted here that the value of the charge on the positive ion becomes the subscript for the negative ion and that the value of the charge on the negative ion becomes the subscript for the positive ion. One technique to determine the correct formula of a compound if we know the ions involved therefore consists of transposing the value of the charge of each ion to the subscript of the other ion. Such a technique will work for all formulas except those in which the values of the charges are both equal to 2, 3, 4, and so on. An example of this exception is the formula for the compound formed between calcium and oxygen. Transposing the charges as described would result in

the formula Ca_2O_2 , when actually the correct formula is CaO , because a one-to-one ratio would produce the required neutral unit. We should also mention that it is only the *value* of the charge that is transposed and not its sign. For negative ions, such as -1 , -2 , and -3 , the numbers 1, 2, 3, and so on are transposed.

Example 3.1

What is the formula of the compound formed between barium and bromine?

Solution 3.1

Barium is an alkaline earth metal, and it loses two electrons when forming an ion. Bromine is a halogen, and it gains one electron when forming an ion. Thus we have Ba^{2+} and Br^{1-} . It takes two bromine (bromide) ions to form a neutral unit with the $+2$ barium ion, resulting in BaBr_2 . Alternatively, we could use the technique of transposing the values of the charges. In this case, the subscript of the barium becomes 1 while the subscript of the bromine becomes 2, again resulting in the formula BaBr_2 .

Example 3.2

What is the formula of the compound formed between the ammonium ion and the carbonate ion?

Solution 3.2

The ammonium ion, NH_4^+ , has a $+1$ charge, while the carbonate ion, CO_3^{2-} , has a -2 charge. Therefore it takes two ammonium ions to form a neutral unit with the carbonate ion, and the formula is $(\text{NH}_4)_2\text{CO}_3$. Using the technique of transposing charge values, the subscript for the ammonium ion becomes 2, and the subscript for the carbonate ion becomes 1. This means that $(\text{NH}_4)_2\text{CO}_3$ is the formula.

Some less common monatomic anions not mentioned in the previous discussion are hydride (H^-), carbide (C^{4-}), nitride (N^{3-}), phosphide (P^{3-}) and selenide (Se^{2-}). Like the other monatomic ions discussed, these ions result from the elements gaining electrons so as to have an electron count equal to that of

the noble gas atoms. Formulas for ionic compounds that include these anions can be determined by using the same schemes as those discussed earlier.

3.4 Naming Ionic Compounds

The naming scheme for ionic compounds is quite simple. We first state the name of the positive ion, which is the name of the metal if the positive ion is a metal ion, then state the name of the negative ion. The values of the subscripts do not enter into the naming scheme. Thus the name of the compound with the formula NaCl is sodium chloride, the name of the compound with the formula $\text{Ca}(\text{NO}_3)_2$ is calcium nitrate, and the name of the compound with the formula $(\text{NH}_4)_3\text{PO}_4$ is ammonium phosphate. Several other examples are listed in Table 3.2.

One curious point is how the names of the monatomic anions are determined. The names are obviously derived from the names of the corresponding elements. For example, the monatomic ion derived from chlorine is chloride and the monatomic ion derived from oxygen is oxide. The system for deriving these names is as follows:

Use the stem of element name, followed by *-ide*

For the halogens, the stem is the name of the element without the *-ine* ending. So for fluorine, it is *fluor-*; for chlorine, it is *chlor-*; for bromine, it is *brom-*; and for iodine it is *iod-*. This gives the terms *fluoride*, *chloride*, *bromide*, and *iodide* as the names of the monatomic anions of the halogens. For other non-metals, there is no set system for determining the stem. For oxygen, it is *ox-*, and the anion is oxide. For sulfur, carbon, hydrogen, and nitrogen, the stem is the first four letters of the element name. Thus, for sulfur, it is *sulf-* and the anion is sulfide. For carbon, it is *carb-*, and the anion is carbide. For nitrogen it is *nitr-* and the anion is nitride. For hydrogen it is *hydr-* and the anion is hydride. For

Table 3.2 Examples of Ionic Compounds

Formula	Name
MgBr_2	Magnesium bromide
K_2O	Potassium oxide
$\text{Al}_2(\text{SO}_4)_3$	Aluminum sulfate
NH_4OH	Ammonium hydroxide
CaS	Calcium sulfide

selenium, the stem is the first five letters of the name, and for phosphorus, it is the first six letters. Thus we have *selen-* as the stem for selenium and *phosph-* as the stem for phosphorus, and the anions are selenide and phosphide.

Example 3.3

What is the name of the compound with the formula MgI_2 ?

Solution 3.3

This is named by stating the name of the metal followed by the name of the anion. The name of the anion is *iod-* followed by *-ide*. The name of the compound is therefore magnesium iodide.

Example 3.4

What is the name of the compound with the formula $(\text{NH}_4)_2\text{CO}_3$?

Solution 3.4

The name of the cation is given first, followed by the name of the anion. The name of the compound is ammonium carbonate.

Example 3.5

What is the name of the compound with the formula Ca_2C ?

Solution 3.5

The name of the C^{4-} ion was given previously in this section—carbide. So the compound name is calcium carbide.

Example 3.6

What is the formula of the compound with the name potassium hydroxide?

Solution 3.6

The symbol for potassium (K) is given first, followed by the formula of the hydroxide ion (OH^-). Remember that subscripts may be involved.

The charge on a potassium ion is +1 while the charge on the hydroxide ion is -1 . One potassium ion with one hydroxide ion produces a neutral unit. Thus the formula of potassium hydroxide is KOH.

In the previous discussions, you may have noticed that the cations in ionic compounds are almost always metal ions. The only exception to this that we have seen is the ammonium ion, NH_4^+ . The reason for this, as stated previously, is that metals tend to lose electrons to become like the noble gas that precedes them. You may also have noticed that the anions in ionic compounds are always either ions of nonmetals or polyatomic ions containing only nonmetals. The reason for this, as stated in Section 3.2, is that nonmetals tend to gain electrons to become like the noble gas located to their right in the periodic table. Briefly stated, metals form cations and nonmetals form anions. While there are some important compounds in which metal atoms are a part of negatively charged polyatomic ions (as we will see), this description covers almost all ionic compounds.

What about covalent compounds? Covalent compounds are nonionic, molecular compounds. For the most part, they contain only nonmetal atoms. A huge number of such compounds exist, and the naming schemes for these are much more involved than what we have seen for ionic compounds. The schemes we will study in this chapter are limited to compounds that contain only two nonmetals and compounds that are acids. Compounds, whether ionic or covalent, that contain just two elements are called **binary compounds**. Compounds, whether ionic or covalent, that contain three elements are called **ternary compounds**. The naming scheme for binary covalent compounds is discussed in Section 3.7. The naming scheme for acids, both binary and ternary, is discussed in Section 3.8.

3.5 Other Metal Ions

Besides the alkali metals, the alkaline earth metals and aluminum, we also have the transition metals, among others. The charges that the cations of these metals have cannot be ascertained from their position in the periodic table. In addition, many of them form cations with several different charges. For example, iron can form cations with either a +3 charge or a +2 charge. The following are common metal ions:

- Cd^{2+}
- Cr^{3+} and Cr^{6+}

- Co^{2+} and Co^{3+}
- Cu^{+} and Cu^{2+}
- Fe^{2+} and Fe^{3+}
- Pb^{2+}
- Mn^{2+} and Mn^{4+}
- Ni^{2+}
- Hg^{+} and Hg^{2+}
- Ag^{+}
- Sn^{2+} and Sn^{4+}
- Ti^{4+}
- Zn^{2+}

Where there can be more than one possible charge for a given metal, Roman numerals are often used in the names in order to differentiate. For example, in this list, the names of the two iron ions are Iron(II) and Iron(III) and the names of the two copper ions are Copper(I) and Copper(II). The charges on monatomic ions are also their so-called **oxidation number**. More is presented on oxidation numbers in Chapter 13.

3.6 Other Negative Ions

Besides the monatomic negative ions and the Big Six polyatomic negative ions, we have many other negative ions. These are additional polyatomic ions of nitrogen, phosphorus, and sulfur as well as polyatomic ions of the halogens and ions formed by partially neutralizing oxyacids that have two or more hydrogens. There are also polyatomic negative ions that include metals. Here is a list of additional common polyatomic negative ions. The names of the ions are in parentheses.

- HCO_3^{-} (bicarbonate)
- HSO_3^{-} (bisulfite) and HSO_4^{-} (bisulfate)
- BrO_3^{-} (bromate)
- CrO_4^{2-} (chromate) and $\text{Cr}_2\text{O}_7^{2-}$ (dichromate)
- CN^{-} (cyanide)
- ClO^{-} (hypochlorite), ClO_3^{-} (chlorate), and ClO_4^{-} (perchlorate)
- IO_3^{-} (iodate)
- NO_2^{-} (nitrite)
- MnO_4^{-} (permanganate)

- $\text{S}_2\text{O}_3^{2-}$ (thiosulfate) and $\text{S}_2\text{O}_8^{2-}$ (peroxydisulfate)
- HPO_4^{2-} (monohydrogen phosphate) and H_2PO_4^- (dihydrogen phosphate)
- CNS^- (thiocyanate)

3.7 Binary Covalent Compounds

Learning the names of binary covalent compounds may also be troublesome because of the variety of naming schemes that exist. Covalent compounds are those that are mostly formed between two or more nonmetals. Like the compounds of the transition metals discussed earlier, nonmetals can exist in a variety of oxidation numbers. Thus naming schemes have also been devised to distinguish between two or more different compounds formed between the same two nonmetal elements. Examples are the compounds formed between carbon and oxygen (CO and CO_2).

Probably the most useful naming scheme for this class of compounds is one that uses prefixes to indicate the number of each type of atom in the formula. These prefixes and the corresponding number of atoms that they represent are given in Table 3.3. Utilizing this scheme, CO would be named carbon monoxide and CO_2 would be named carbon dioxide. Notice in these examples that the element that is shown first in the formula is also named first. The second word uses the stem of the second nonmetal and the *-ide* ending, as in the scheme for ionic binary compounds. However, a prefix is placed in front of the stem to indicate the number of atoms of the second nonmetal in the formula (*mono-* in the case of one atom and *di-* in the case of two). Actually, a prefix is also used in front of the name of the first element if there is more than one atom of this element indicated in the formula. For example, the compound N_2O is called

Table 3.3 Prefixes for Naming Binary Covalent Compounds

Number of Atoms	Prefix
1	mono-
2	di-
3	tri-
4	tetra-
5	penta-
6	hexa-

Table 3.4 Example of Naming of Binary Covalent Compounds Using Prefixes

Compound	Name
CO	Carbon monoxide
CO ₂	Carbon dioxide
N ₂ O	Dinitrogen monoxide
NO ₃	Nitrogen trioxide
CCl ₄	Carbon tetrachloride
P ₂ O ₅	Diphosphorus pentaoxide
SF ₆	Sulfur hexafluoride

dinitrogen monoxide. The general scheme then is as follows. The **first word** in the name consists of the **prefix (if more than one atom), followed by the name of the first element**. The **second word** in the name consists of the **prefix for the second element (in all cases) followed by the stem of the second element and the suffix -ide**. Examples are presented in Table 3.4.

Besides this scheme, the “-ous and -ic” scheme and the Roman numeral scheme are also used. For example, N₂O is also known as nitrous oxide and as nitrogen (I) oxide (the oxidation number of nitrogen in this compound is +1). While the Roman numeral scheme would provide a name for all possibilities, the “-ous and -ic” scheme is limited to only two compounds, the lower and higher of two oxidation numbers. Some nonmetal–nonmetal combinations can exist in multiple combining ratios. For example, there are seven oxides of nitrogen—N₂O, NO, NO₂, NO₃, N₂O₃, N₂O₄, and N₂O₅. Nitrous oxide is the name assigned to N₂O, and nitric oxide is the name assigned to NO. The rest of the oxides of nitrogen must be named by the other schemes.

3.8 Identifying and Naming Acids

One subclassification of covalent compounds is that of the **acids** (see Figure 3.5). Acids have unique properties that are related to the fact that they are a source of hydrogen ions (H⁺). Since they are a source of hydrogens, it follows that acids would have one or more hydrogens in their formulas. Also, since the hydrogens give acids their unique properties, the hydrogens are presented first in these formulas. Compounds whose formulas have one or more hydrogens

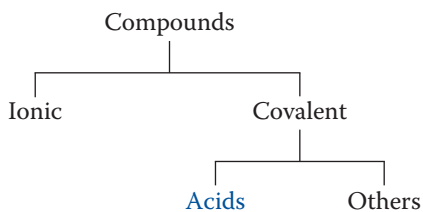


FIGURE 3.5 Acids are covalent compounds.

that are listed first are acids. Common examples are the binary acids HF, HCl, HBr, HI, and H_2S . Common examples of ternary acids are those in which the hydrogens are bonded to the familiar polyatomic ions— H_2SO_4 , HNO_3 , H_2CO_3 and H_3PO_4 . These are often called **oxyacids** because the third element of the ternary formula (following hydrogen and a nonmetal) is oxygen.

Not all compounds that have hydrogen atoms covalently bonded to other atoms are acids. That is to say, not all compounds that have hydrogens in their formulas are sources of hydrogen ions. Some compounds with covalently bonded hydrogens have a potential to release hydrogen ions and some do not. Those that do release hydrogen ions have hydrogen listed first in the formula, while those that do not release hydrogens may have hydrogen elsewhere in the formula. An example of a compound that has hydrogens, none of which can become hydrogen ions, is ammonia, which has the formula NH_3 . This formula indicates that ammonia is *not* an acid. Those hydrogen atoms that have a potential for becoming hydrogen ions are called **acidic hydrogens**. There are some compounds that have both acidic and nonacidic hydrogens; these compounds have hydrogens in two different locations in the formula. An example of such an acid that is fairly common is acetic acid, which has the formula $\text{HC}_2\text{H}_3\text{O}_2$. Examples of names of acids are given in Table 3.5.

One final comment about acids is that binary acids that are not dissolved in water are gases. These gases can be named as if hydrogen were an alkali metal, which means that they can be named like the ionic compounds in Section 3.4. Using this method, HCl would be named hydrogen chloride, HBr would be named hydrogen bromide, and H_2S would be named hydrogen sulfide. This is *not* true of *ternary* acids, however. For example, nitric acid (HNO_3) does not exist as a gas when it is not in a water solution and would *not* have hydrogen nitrate as an alternate name.

Table 3.5 Formulas and Names of Common Binary and Ternary Acids	
Acid	Name
HF	Hydrofluoric acid
HCl	Hydrochloric acid
HBr	Hydrobromic acid
HI	Hydroiodic acid
H ₂ S	Hydrosulfuric acid
H ₂ SO ₄	Sulfuric acid
HNO ₃	Nitric acid
H ₂ CO ₃	Carbonic acid
H ₃ PO ₄	Phosphoric acid

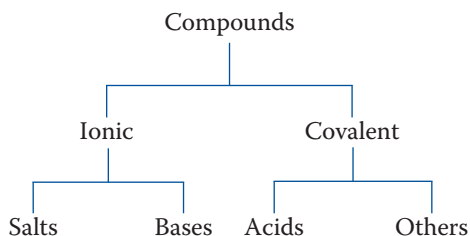


FIGURE 3.6 Salts and bases are ionic compounds.

3.9 Bases and Salts

Another subclassification of compounds that is important is that of **bases** (see Figure 3.6). Bases are the opposite of acids in the sense that they *react with* hydrogen ions rather than supply them. Examples of bases are compounds containing the hydroxide polyatomic ion OH^- . Thus, NaOH , KOH , and NH_4OH are known as bases. They react with hydrogens because the hydrogens and the hydroxides together will form water (H_2O [or HOH]).

When water forms upon contact of an acid with a base, another substance that forms is a **salt**. The word *salt* in this context does not only refer to table salt (NaCl) but rather to any ionic compound (except for the hydroxides, which are bases). A salt can be thought of as a compound that forms as a result of the replacement of the hydrogen in an acid with a metal or the ammonium ion. For example, NaCl can form from HCl , K_2SO_4 can form from H_2SO_4 , and

NH_4NO_3 can form from HNO_3 ; each of these products is a salt. This phenomenon will be discussed again in later chapters.

3.10 Homework Exercises

Subclassification of Elements

1. In the following table, place a checkmark in the box that correctly identifies the element in the left column as a metal, nonmetal, or metalloid.

	Element	Metal	Nonmetal	Metalloid
a.	Calcium			
b.	Rb			
c.	Oxygen			
d.	Si			
e.	Rhodium			
f.	Br			
g.	Arsenic			
h.	Ti			
i.	Selenium			
j.	U			

2. In the following table, place a checkmark in the box that correctly identifies the element in the left column as a representative element, a transition element, or an inner transition element.

	Element	Representative Element	Transition Element	Inner Transition Element
a.	Curium			
b.	Al			
c.	Strontium			
d.	Co			
e.	Sulfur			
f.	Mo			
g.	Californium			
h.	Au			
i.	Antimony			
j.	Li			

3. In the following table, place a checkmark in the box that correctly identifies the element in the left column as an alkali metal, an alkaline earth metal, a halogen, a noble gas, or none of these.

	Element	Alkali Metal	Alkaline Earth Metal	Halogen	Noble Gas	None of These
a.	Fluorine					
b.	B					
c.	Calcium					
d.	K					
e.	Argon					
f.	Sn					
g.	Barium					
h.	V					
i.	Iodine					
j.	Cs					

Formation of Ions

4. Define monatomic ion, polyatomic ion, cation, and anion.
5. In the following table, place a checkmark in the box that correctly identifies the ion in the left column as a monatomic ion, a polyatomic ion, a cation, or an anion. Each ion will have two boxes checked.

	Ion	Monatomic Ion	Polyatomic Ion	Cation	Anion
a.	Phosphate				
b.	I ⁻				
c.	Fluoride				
d.	NH ₄ ⁺				
e.	Oxide				
f.	Mg ²⁺				
g.	Carbonate				
h.	Na ⁺				
i.	Iodide				
j.	Al ³⁺				

6. In the following table, write down the charge of the ion (e.g., +1, -1) that the element in the left column is expected to have when that element forms a monatomic ion.

	Element	Charge
a.	Chlorine	
b.	Be	
c.	Selenium	
d.	Rb	
e.	Strontium	
f.	I	
g.	Potassium	
h.	Al	
i.	Sulfur	
j.	Li	

Formulas of Ionic Compounds

7. What are the formulas of the binary compounds formed between each of the following?
- Calcium and oxygen
 - Calcium and bromine
 - Potassium and the phosphate ion
 - Magnesium and the nitrate ion
 - Aluminum and sulfur
 - The ammonium ion and the carbonate ion
 - Potassium and oxygen
 - The ammonium ion and oxygen
 - Rubidium and the sulfate ion
 - Sodium and bromine
 - Aluminum and chlorine
 - Calcium and the nitrate ion
- m. Magnesium and the phosphate ion
- Barium and sulfur
 - Calcium and fluorine
 - Aluminum and oxygen
 - Potassium and sulfur

- r. Magnesium and oxygen
 - s. Rubidium and the carbonate ion
 - t. Lithium and the phosphate ion
 - u. Strontium and the hydroxide ion
 - v. The ammonium ion and chlorine
8. Tell which of the following chemicals are ionic and which are covalent.
 SrCO_3 , P_2O_5 , MgBr_2 , NH_4NO_3 , CCl_4 , $\text{Al}_2(\text{SO}_4)_3$, $\text{C}_3\text{H}_8\text{O}$, NCl_3 , FeBr_2 ,
 K_2O , MgCl_2 , IBr , CO_2 .
9. Identify each of the following as an acid, base, or salt.
- a. CaCl_2
 - b. H_2SO_4
 - c. KI
 - d. NH_4OH
 - e. $\text{NaC}_2\text{H}_3\text{O}_2$
 - f. $\text{Ba}(\text{OH})_2$
10. Name the following compounds.
- a. HBr
 - b. CaO
 - c. P_2O_5
 - d. KI
 - e. FeCl_2
 - f. Na_2S
 - g. SO_3
 - h. $\text{Mg}(\text{OH})_2$
 - i. CCl_4
 - j. HCl
 - k. CaBr_2
 - l. FeO
 - m. CaSO_4
 - n. FeBr_2
 - o. SF_6
 - p. HI
 - q. BaCl_2
 - r. H_2S
 - s. MgO
 - t. CuCl_2
 - u. PCl_3

11. Write the formula of the compound formed between the following.
- Potassium and the CNS^- ion.
 - The ammonium ion and the HPO_4^{2-} ion.
 - Sodium and the $\text{Cr}_2\text{O}_7^{2-}$ ion.
 - Calcium and the ClO^- ion.
 - Magnesium and the ClO_4^- ion.
 - Lead(II) and the CrO_4^{2-} ion.
 - Potassium and the CN^- ion.

Naming Simple Ionic and Covalent Compounds

12. Give the formula of each of the chemicals with the following names.
- Carbon tetrachloride
 - Copper(I) bromide
 - Ammonium chloride
 - Sodium fluoride
 - Iron(III) oxide
 - Dinitrogen trioxide
 - Tin(II) sulfate
 - Silver chloride
 - Sulfur hexafluoride
 - Hydrofluoric acid
 - Tin(IV) chloride
 - Potassium oxide
 - Sodium hydroxide
 - Titanium(IV) oxide
 - Hydrochloric acid
 - Diphosphorus pentoxide
 - Magnesium chloride
 - Nickel(II) sulfide
 - Dinitrogen pentoxide
 - Cobalt(II) fluoride
 - Aluminum bromide
13. For each name, select the correct formula from those given.
- | | |
|----------------------|--|
| a. Magnesium bromide | MgBr , MgBr_2 , or Mg_2Br |
| b. Phosphoric acid | H_3PO_4 , H_2PO_4 , H_3PO_3 , or H_2PO_3 |
| c. Iron(III) iodide | FeI_2 , FeI_3 , Fe_2I , or Fe_3I |
| d. Potassium sulfate | KSO_3 , K_2SO_3 , KSO_4 , or K_2SO_4 |

- e. Copper(I) carbonate CuCO_3 , Cu_2CO_3 , CuCO_4 , or Cu_2CO_4
 - f. Sulfur trioxide SO_3 or SO_3^{2-}
 - g. Hydrosulfuric acid HSO_4 , H_2SO_4 , HS , H_2S
 - h. Dinitrogen monoxide NO , NO_2 , N_2O , or N_2O_3
14. For each formula, select the correct name from those given.
- a. HNO_3 Hydronitric acid or nitric acid
 - b. CCl_4 Carbon chloride or carbon tetrachloride
 - c. AlBr_3 Aluminum bromide or aluminum tribromide
 - d. SnO_2 Tin oxide, tin(IV) oxide, or tin dioxide
 - e. $(\text{NH}_4)_2\text{CO}_3$ Ammonium carbonate, diammonium carbonate,
or ammonium tricarbonate

Chapter 4

Atomic Structure

Introduction	88
History of Theories and Experiments	89
Modern Theory of Light	92
Bohr Model	97
Schrodinger Model.	99
Orbital Diagrams	104
Electron Configuration.	107
Homework Exercises	113

4.1 Introduction

One of the most fascinating topics in our study of chemistry is the concept of chemical combination. What is it about chemical combination that converts a highly flammable gas (hydrogen) and a gas that supports combustion (oxygen) into a liquid that puts out fires (water)? What is it about the process of two atoms of hydrogen and one atom of oxygen chemically combining to form a molecule of water that results in the formation of a new substance that in no way resembles the substances that went into it? And that is just one example! There are innumerable examples of this phenomenon in our world. In this chapter, we begin to unravel the mystery of chemical combination by studying the structure of the smallest unit of matter, the atom, in more detail.

You will recall that we previously defined chemistry as the study of the structure, composition, and properties of matter and the changes that matter undergoes. With an examination of the phenomenon of chemical combination comes a need for the consideration of the structure and composition portions of this definition. Specifically, we need to look at the characteristics of that ultrasmall structural unit that is at the heart of the structure of all matter—elements, compounds, and mixtures alike—the atom. We presented some basic information about the atom in Chapter 1. We will now go into greater detail in describing its structure. The electron structure is what will be emphasized in this discussion. When two atoms come together and chemically combine, it is the electrons that are involved because they are outside the nucleus on the outer edges of the atoms (Figure 4.1). However, to make our discussion complete, we want to present a historical perspective, which includes how it came

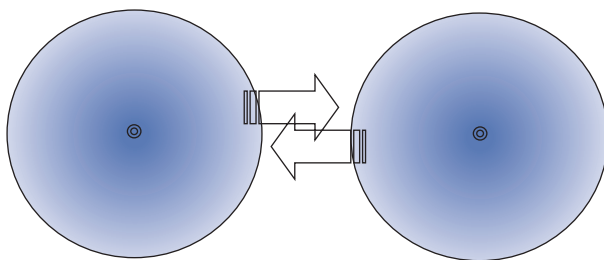


FIGURE 4.1 When two atoms come together and chemically combine, it is the electrons that are involved because they are outside the nucleus on the outer edges of the atoms.

to be known that the nucleus exists, that protons and neutrons are located in the nucleus, and that the electrons are located outside the nucleus.

4.2 History of Theories and Experiments

The structure of matter on a submicroscopic scale is something that has intrigued scientists for years. Modern theories of atomic structure had their origin in 1808 when **John Dalton**, an Englishman, published a theory based on certain facts and experimental evidence that were known at the time. Not only did he propose the existence of the atom, but he went on to precisely describe some attributes of the atom that we still hold to be true today. The listing of these attributes is called **Dalton's Atomic Theory** and consists of the following:

1. All matter is made up of tiny indivisible particles called atoms, and the atom is the smallest structural unit of an element.
2. In terms of their mass and chemical and physical properties, atoms of the same element are all very similar and atoms of different elements are very different.
3. Atoms chemically combine with each other in simple whole number ratios to form molecules.
4. In chemical reactions, atoms of elements retain their identity, but rearrange, reorganize, and reunite in the formation of other substances.

Thus important facts relating to the structure of matter have been known and accepted for some time. The mystery and intrigue since Dalton's time has mostly related to what is *inside* the atom. Indeed, if we are to understand more clearly the why's and how's of chemical combination (i.e., what exactly happens when two atoms, or an atom and a molecule, or two molecules come together and undergo chemical change), then an examination of the interior of this mysterious little particle is in order.

Two of the particles we find inside the atom, the proton and the electron, were discovered in the late 1800s and early 1900s by scientists working with instruments and devices that were fairly sophisticated for the time. The electron was discovered and characterized by **William Crookes**, **J. J. Thomson**, and **Robert Millikan** as a result of independent experimentation occurring between 1879 and 1909.* Some of the work involved the use of a device called a cathode ray tube.

* Thomson (in 1906) and Millikan (in 1923) each won the Nobel Prize in Physics.

An electric current (which we now know to consist of moving electrons) passing between two metal plates in an evacuated glass tube could be made to “bend” when a magnet is brought near. Thomson measured the ratio of the electron’s charge to its mass, and Millikan, through an experiment that has come to be known as the Millikan Oil Drop Experiment, measured the charge of an electron by measuring the velocity of a tiny electrically charged oil droplet as it moved through a varied electric field. The mass of the electron was then calculated.

After the discovery and characterization of the electron, Thomson proposed a model for the atom in which the electrons were thought of as particles embedded in a “cloud” of positive charge. This model came to be known as the **plum pudding model** because the electrons were pictured like plums embedded in plum pudding. A famous experiment performed by a team of scientists led by the Englishman **Ernest Rutherford** in 1911 disproved this model, however.*

Rutherford used a spray of positively charged particles called alpha (α) particles to probe the interior of the atom to determine the arrangement of the particles. Alpha particles are one of the three kinds of radiation that are discharged by radioactive materials. The others are beta (β) particles and gamma (γ) rays. An alpha particle consists of two neutrons and two protons in a single particle (a helium nucleus). They thus have a +2 charge. Beta particles are electrons, and gamma rays are a very high-energy form of light. Alpha particles are positively charged (and very heavy, comparatively speaking) and were used by Rutherford for his experiment.

The atom Rutherford chose to probe was a gold atom. Using an extremely thin gold foil, he set up an experiment in which the gold foil was suspended directly in the path of the moving alpha particles, as pictured in Figure 4.2,

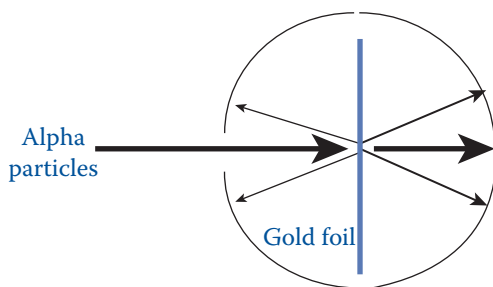


FIGURE 4.2 A drawing depicting Rutherford’s gold foil experiment.

* Rutherford won the Nobel Prize in Chemistry in 1908 for other related work.

such that the alpha particles would strike the foil at a right angle, a direct hit. In front of and around the back of the gold foil he placed detectors that were sensitive to alpha particles so that the alpha particles could be detected wherever they ended up. The results produced a model for the atom that we still subscribe to today, a century later.

Rutherford's gold foil was mostly transparent to the alpha particles—that is, the vast majority of alpha particles passed through the foil and were observed by the detectors directly behind it, apparently unaffected by the presence of the foil in their path. It was as if the foil were not even there. A small number of the particles, while not completely blocked by the foil, was found by the detectors on the back side deflected from their path a little above and a little below the main stream of particles. In addition, an extremely small number was found on the detectors on the front side, indicating that they had bounced off the foil and did not pass through it.

The model derived from these observations is called the **Rutherford model**, or the **nuclear model**, as shown by the circles in Figure 4.3. The scientists concluded that most of the mass of the atom had to be concentrated into an extremely small volume of space inside the atom, which they called the nucleus. That was the only way to explain the small number of alpha particles that were found reflected off the front side of the foil. In addition to this small volume of space where most of the mass is found, they concluded that the nuclei of neighboring atoms were surprisingly distant from one another and that the masses of any particles found there had to be extraordinarily small. This was the only way to explain finding the vast

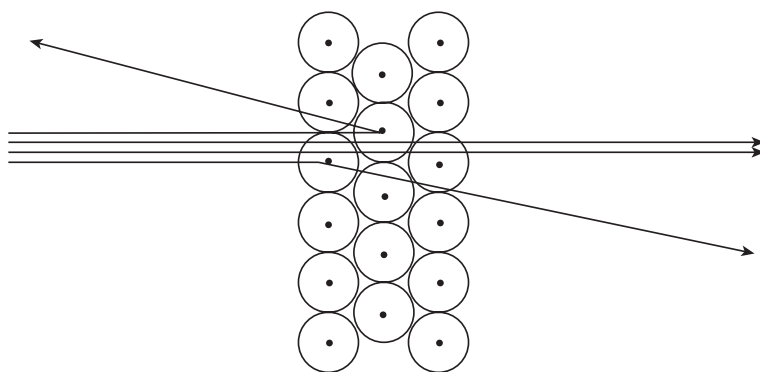


FIGURE 4.3 A close-up of the gold foil showing how the existence of the nucleus explains Rutherford's observations.

majority of the particles directly behind the foil from the point where they would have struck it.

To explain the particles that were deflected from their path and found a little above or a little below the main grouping of particles, the scientists concluded that the nucleus must have a charge that would repel the alpha particles, causing them to slightly alter their course if they came near it. Thus the nuclear model was born. In this model, the space outside and between the nuclei is viewed as being mostly empty space and is occupied by the electrons, the extremely small and light particles of negative charge. Figure 4.3 shows how there could be different paths taken by the alpha particles as they encountered, or did not encounter, the gold nuclei.

In order to explain some cathode ray phenomena observed 25 years earlier by other scientists, Rutherford proposed in 1914 that smaller positively charged particles were contained within the nucleus. He named these particles protons. Neutrons in the nucleus were proposed by the British scientist **James Chadwick** in 1932 as a result of additional experiments.*

Two other models for the atom are important. One was proposed by a Danish scientist by the name of **Neils Bohr** in 1913.† This model was called the **solar system model** because he proposed that the electrons orbit the nucleus much like the planets orbit the sun. The other was proposed by Austrian physicist **Erwin Schrodinger** in 1926.‡ The Schrodinger model is called the **quantum mechanical model** and is the model that we use today to explain and predict atomic behavior. Each of these two models is explained more fully in the sections that follow. A summary of the history of the development of atomic theory is given in Table 4.1

4.3 Modern Theory of Light

To have a clear understanding of the atomic theories espoused by Bohr and Schrodinger, and therefore a clear understanding of the current models of atomic structure, we must first briefly explore the modern theory of the nature of light, which is a form of energy.

* James Chadwick won the Nobel Prize in physics in 1935 for his discovery of the neutron.

† Neils Bohr won the Nobel Prize in physics in 1922 for his solar system model.

‡ Schrodinger won the Nobel prize in 1933 for his model.

Table 4.1 Names and Years of Noteworthy Achievements in the Development of Atomic Theory

Year	Scientist	Achievement
1808	Dalton	Atomic theory
1879–1909	Crookes, Thomson, Millikan	Discovery and characteristics
1909	Thomson	Plum pudding model
1911	Rutherford	Nuclear model
1913	Bohr	Solar system model
1926	Schrodinger	Quantum mechanical model

4.3.1 What Is Energy?

Energy is defined as the capacity for causing matter to move due to an applied force. Energy can take many forms. First, there is **mechanical energy**, also called **kinetic energy**, or the energy of motion. An example of this is the energy of a moving bowling ball as it speeds down a bowling lane. The energy it possesses is its capacity to move the pins. The kinetic energy a moving object has depends on its mass and velocity. The faster it moves and the heavier it is, the more energy it has. The mathematical relationship to describe this is:

$$\text{Kinetic energy} = \frac{1}{2} mv^2 \quad (4.1)$$

in which m is the mass and v is the velocity of the object.

There is also **potential energy**, or the energy an object has because of its position rather than its motion. An example of this is the energy of an object sitting on the edge of a table. Given a nudge, it falls from the table and its potential energy is released as kinetic energy. But as long as it sits at the edge, its energy is stored as potential energy.

This latter example illustrates an important point—one form of energy can be converted to another form and anything that can *potentially* cause an object to move is a form of energy. This means that we can add **light energy**, **electrical energy**, **chemical energy**, **heat energy**, and **nuclear energy** to the list because all have the *potential* to cause an object to move. The electrical energy possessed by a battery causes the Energizer Bunny™ to move. When the battery is put to use, the chemical energy possessed by the chemicals present inside the battery is converted to electrical energy and

ultimately to kinetic energy. Light energy from the sun can be converted to electrical energy by the solar energy cell in a solar calculator. Heat energy heats our homes in the winter. Nuclear energy possessed by certain isotopes held as the fuel in a nuclear power plant can be converted to electrical energy to drive a variety of devices in our homes. The present discussion will be limited to light energy, however. We need to have a basic sense of what light is so that we can discuss the relationship of light to the modern theory of atomic structure.

4.3.2 What Is Light?

The modern theory of light is that it has a dual nature, a rather complex and abstract notion. First, some properties of light are best explained if we think of it as a stream of particles emanating from the source of the light (e.g., light bulb or sun). These particles are called **photons**. This is the so-called **particle theory**. Other properties of light are best explained if we consider it to be a **wavelike disturbance** that has both an electrical component and a magnetic component. Hence the reason light is sometimes called electromagnetic radiation. This is the so-called **wave theory**. The wave theory is the most useful to us for our present purpose.

In the wave theory, light is thought of as a continuous series of repeating electromagnetic waves that travel through space and matter. As an analogy, you may picture a single wave as what you observe when a rock is dropped into a pond of water. A water wave is created that travels outward from where the rock contacted the water. The differences are (1) the water wave is a mechanical wave, as opposed to light, which is an electromagnetic wave and (2) the water wave is just one wave, whereas electromagnetic radiation is a continuous series of waves, repeating one after the other, flowing forth from the light source. A mechanical wave requires matter to exist.

Electromagnetic waves do not require matter to exist because they are electrical and magnetic in nature. That is why we are able to see light from the sun. The void (absence of matter) between the Earth and the sun is no problem for light as it travels to the Earth.

With electromagnetic radiation, the waves are repeating and there is a finite distance from the crest of one wave to the crest of the next wave. This distance is called the **wavelength**. Some wavelengths are extremely short while others are extremely long. A wavelength can be shorter than the radius of the smallest atom, or it can be as long as several miles. See

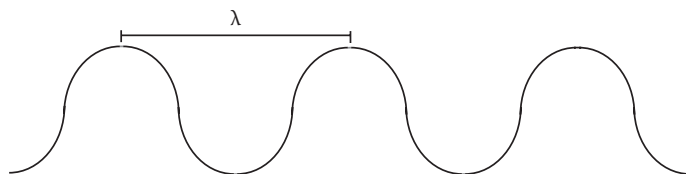


FIGURE 4.4 A drawing showing the concept of a series of repeating light waves. The wavelength (λ) is the distance from the crest of one wave to the crest of the next wave.

Figure 4.4 for the concept of wavelength. Wavelengths are expressed in units of length. Extremely short wavelengths are expressed in nanometers (1 billionth of a meter) or angstroms, which are on the order of the size of an atomic radius. Extremely long wavelengths are expressed in kilometers. The energy of light depends on this wavelength. The shorter the wavelength, the greater the energy. Thus very short wavelengths of light (angstroms) possess much energy—enough to do harm to the human body (e.g., gamma rays and X-rays). Very long wavelengths, such as radio and television waves (kilometers), are not harmful at all.

The human eye is sensitive only to a narrow range of wavelengths, the so-called visible wavelengths. Hence these are the only wavelengths we can see. Visible light wavelengths are seen as colors. While visible light, such as that from the sun or from a light bulb, is “white,” this light is actually a mixture of all the colors, or of many different wavelengths. Under certain conditions, these wavelengths can be separated from each other, such as by water droplets in the atmosphere when a rainbow is visible or by a prism, a solid piece of glass that has a triangular shape. In addition, the colors that various forms of matter appear to have are due to the fact that some wavelengths get absorbed by matter and some do not. For example, if the cover of the notebook you use for taking class notes is red, the red wavelengths are being reflected off (transmitted by) the cover while all other wavelengths are being absorbed.

The wavelength range of visible light is very narrow compared with the range that encompasses all light. And there are wider ranges of wavelengths that are much shorter—for example, gamma rays and ultraviolet waves, which are just outside the visible range on the short side. There are also wider ranges of wavelengths that are much longer—for example, radio and television wavelengths, and infrared wavelengths, which are just outside the visible range

on the long side. Blue light represents the shorter wavelengths in the visible range while red light represents the longer wavelengths. Green and yellow are in between, as in the rainbow. The wavelength range for visible light is about 350–750 nanometers (nm, $1\text{ nm} = 10^{-9}\text{ m}$). While these wavelengths are extremely short (the same as about 16,000 hydrogen atoms lined up end to end), their energies are not harmful under normal circumstances, which is fortunate because we depend on visible light to view everything we see. Blue light has the highest energy. Red light has the lowest energy. Green and yellow are in between.

As mentioned, ultraviolet wavelengths are shorter than visible wavelengths and infrared wavelengths are longer than visible wavelengths. Thus, of the three, ultraviolet has the greatest energy (e.g., causes sunburn) while infrared has the least energy. Ultraviolet, visible, and infrared wavelengths are used routinely in a chemistry laboratory for chemical analysis purposes, as we will see later in this chapter. The complete range of all light, from the short wavelength gamma rays to the long wavelength radio and television waves, is called the **electromagnetic spectrum**. Figure 4.5 summarizes the points made in this section.

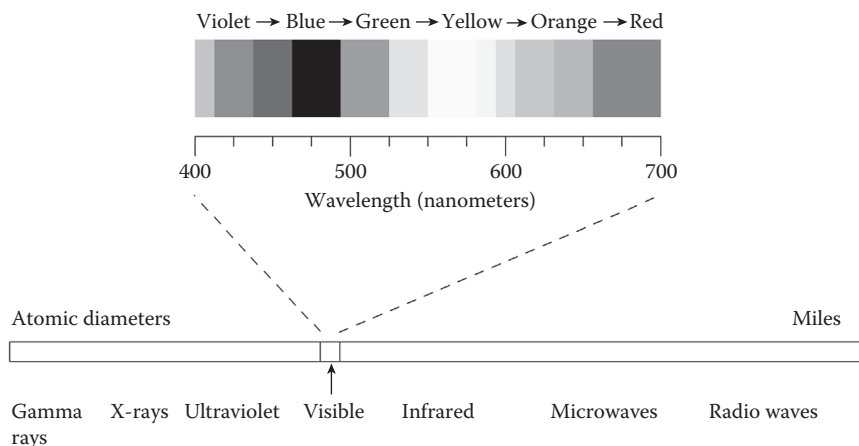


FIGURE 4.5 The electromagnetic spectrum. The visible range of wavelengths is a very narrow range in the middle of the entire electromagnetic spectrum, and the wavelengths within the visible range are the colors of the rainbow.

4.4 Bohr Model

4.4.1 Introduction

As stated previously, with the protons and neutrons in the nucleus deep inside the atom, it is logical to think that the electrons play the most important role as far as contact with other atoms. It is logical to assume that it is the outer edges of the atomic spheres, and thus the electrons, that interact when there is a close approach between the atoms, ultimately resulting in chemical combination.

In the time immediately following Rutherford's experiment, the electrons and the space they occupy still remained a mystery to scientists. They still had important questions to consider. Exactly how do we picture the electrons? If they are so very small, exactly where are they located in this vast empty space outside the nucleus? Since positive charges attract negative charges, why do these small, negatively charged particles not fall into the positively charged nucleus? Exactly what is the role of electrons when chemical combination takes place? The next step appeared to be the development of a model that would fit the known data and answer these questions. An explanation as to why the electron does not fall into the nucleus is that each electron travels in an orbit, much like the planets travel around the sun, such that they are influenced by a force directed outward (centrifugal force) that would counterbalance the force of attraction for the nucleus. This is similar to a string tied to a ball that you would twirl around with your hand. While there is an attractive force between your hand and the ball (represented by the pulling force exerted on the string), there is an equal but opposite force, due to the ball's motion, that keeps the ball from crashing into your hand. This solar system model became popular, and indeed is still popular today; the idea of electrons as particles moving around the nucleus in orbits much like planets moving around the sun is a very common representation of electrons.

According to this model, electrons, as moving particles, have kinetic energy. In order for an electron not to fall into the nucleus, it must have sufficient energy to keep it aloft. The further the electron is from the nucleus, the more energy it must have—just as the longer the string, the faster the ball must move to keep the string taut. It is clear with this model that an electron's energy is associated with its speed; the further it is from the nucleus, the faster it must move, and thus the more energy it must have. While these ideas did not answer all the questions, it was a significant step forward. There was also the

matter of light emission by chemical substances under certain conditions that Bohr wanted to address.

4.4.2 Bohr Model and Light

All matter tends to want to be in its lowest possible energy state. When an object on the edge of a table is gently pushed off the edge, it falls to the floor; it loses energy. The Bohr theory states that when an electron is in an orbit of high energy, it will drop to an orbit of lower energy if a vacancy for it exists in the lower orbit. This explained a strange phenomenon observed at the time, the emission of light by matter.

Bohr had observed that, under certain conditions, elements can emit light. Examples of this emission that we might observe today include fireworks and colored flames. Bohr's explanation was that when an atom is energized, such as in a fireworks display or a flame, an electron gains energy and moves from an orbit of lower energy to one of higher energy. Because this new orbit is not the lowest possible energy state for the electron to be in and because there is a vacancy in a lower orbit, the electron subsequently loses this gained energy and drops back to the lower orbit. The energy that is lost is the energy of light, and the light that is created is seen as a particular color (wavelength) (Figure 4.6). What color this light is depends on exactly how much energy is lost.

Thus the acceptance of the concept of the energy of the electron, and the fact that part of this energy can be lost in the form of light, is a key to the acceptance of Bohr's model. In the language of modern chemistry, we say that the energy of the electron is quantized, which means that the Bohr orbits

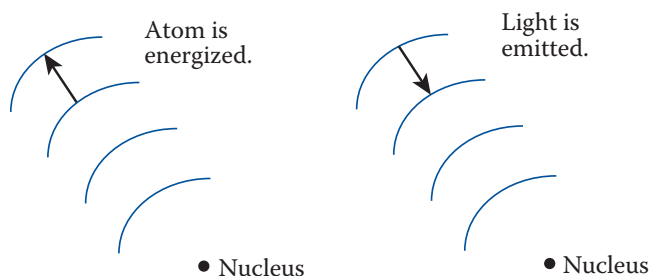


FIGURE 4.6 Bohr theorized that when electrons drop from a higher orbit to a lower orbit after being energized, light is emitted because the energy difference between one orbit and another is on the order of light energy.

represent levels of energy between which the electron can move and that the electron can only have those energies that are represented by these levels. In other words, the electron cannot be found anywhere but in an orbit, meaning it cannot be positioned between orbits. This idea can be compared to climbing a ladder. A person can only stand on one of the rungs. He cannot stand in the space between the rungs.

The Bohr model was important because, while not the correct model, it led to the modern model of the atom, the model that is widely accepted today as the true model of atomic structure.

4.5 Schrodinger Model

The modern theory of the atom was initially introduced by Erwin Schrodinger in 1926. It is popularly known as the **quantum-mechanical model**. This model is based on some very complex mathematics, with which we will not concern ourselves here. The essence of the model is that electrons exist in principal (or main) energy levels, in energy sublevels within these principal levels, and in regions of space called orbitals within the sublevels. The electrons are also thought of as having a particular spin direction.

Principal energy levels for electrons are characterized by a number assigned to that level. This number is symbolized by the letter n . The principal level closest to the nucleus is the first level and is referred to as the $n = 1$ principal level. Proceeding out from the nucleus, the next principal level is the $n = 2$ level, the next is the $n = 3$ level, and so on. Figure 4.7 presents a simplified view of the first few principal levels.

The values that n can take on (i.e., 1, 2, 3, 4, 5, and so on) correspond to the first, second, third, fourth, fifth, and so on levels. The larger the number of the level is, the further that level is from the nucleus. Theoretically, the number of such levels is limitless. Practically, however, no atom that we currently know of has more than seven principal levels. Thus the configuration of electrons of any element in the periodic table can be described by using just seven principal levels. Since there are more than 100 elements in the periodic table, it follows that the average number of electrons that can be found in each level is relatively large. In addition, the number of electrons possible in each principal level varies with the level number. A maximum of only 2 electrons can be found in the $n = 1$ level, 8 in the $n = 2$ level, 18 in the $n = 3$ level, 32 in the

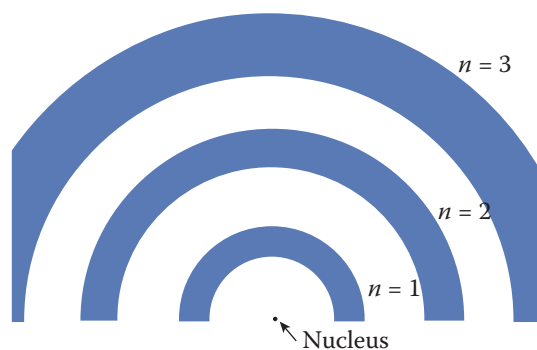


FIGURE 4.7 A drawing depicting the first three energy levels around a nucleus.

$n = 4$ level, etc. The following is a simple equation that can be used to compute the maximum number of electrons that a principal level, n , can hold:

$$\text{Maximum number of electrons possible in level } n = 2n^2 \quad (4.2)$$

It is not sufficient to simply use the n value to describe the configuration of the electron structure around the nucleus since there are so many electrons. The electrons within each principal level are not equal in energy. In fact, the modern theory says that each is unique. To describe each one, we must explore the concepts of energy sublevels within the principal levels, orbitals within the sublevels, and electron spin direction.

The different sublevels that exist are designated by the four letters s , p , d , and f . In other words, there are four possibilities for different sublevels within the principal energy levels for electrons. Not all sublevels are found in every principal level. For example, in the $n = 1$ level, there is only an s sublevel, since that is all that is required to contain the maximum of two electrons calculated using Equation 4.2, and in the $n = 2$ level, there are only s and p sublevels. As the n value increases, more sublevels are required. This is because as n increases, more electrons are found, as is determined from calculations performed with Equation 4.2. Again, theoretically, the number of sublevels is unlimited; however, all electrons in all known elements can be accounted for by only utilizing four sublevels—the four symbolized by the letters s , p , d , and f . We often refer to these sublevels as s , p , d , or f sublevels, to the orbitals as s , p , d , or f orbitals, and to the electrons they contain as the s , p , d , or f

Table 4.2 Maximum Number of Electrons and Number of Orbitals in Sublevels *s*, *p*, *d*, and *f*

Sublevel	Maximum Number of Electrons	Number of Orbitals
<i>s</i>	2	1
<i>p</i>	6	3
<i>d</i>	10	5
<i>f</i>	14	7

electrons. The maximum number of electrons and the number of orbitals that a sublevel has depends on which sublevel it is. See Table 4.2.

Orbitals are regions of space each with a particular shape and orientation. The term *orbital* is not to be confused with the electron orbit that was introduced in the Bohr model. The quantum-mechanical model does not recognize the electron as a particle, and so there is no such thing as an orbit. Instead, the **orbital** is viewed as a region of space around the nucleus in which the entity we call the electron is found.

By specifying an orbital, we come pretty close to uniquely describing each electron in an atom. We can say that a particular electron is in a particular principal level, in a particular sublevel, and in a particular orbital. Any given orbital can only hold two electrons. In order to complete this unique description, we only need to differentiate between the two electrons in the orbital. The quantum-mechanical model states that these two electrons have opposite spin direction.

Since there can be a maximum of two electrons in any orbital, the number of orbitals contained within a principal level of an atom is half the number of electrons, or simply the square of the principal level number.

$$\text{Number of orbitals in principal level } n = n^2 \quad (4.3)$$

Compare Equation 4.3 with Equation 4.2. Notice that the word *maximum* is used in Equation 4.2 but not in Equation 4.3. This is because all possible orbitals exist in all energy levels even though some may not have any electrons in them. So the number of orbitals is a constant number, not a maximum number. Therefore some orbitals are empty; some may have one electron in them; and some may be filled with two electrons. The number of electrons, on the other hand, is limited by the element under consideration. Thus the

number of electrons in a given level varies from one element to another and the maximum number is what a level will have if all orbitals are completely filled.

It is important to consider the energy of these sublevels and orbitals compared with each other. This is so that we can know which sublevels within a principal level (or which orbitals within a sublevel) have electrons and which do not in the event that the principal level (or the sublevel) is not filled with electrons. If a principal level is not filled with electrons, the sublevels with lower energy will have electrons while those at higher energies will not. Levels with lower energy get electrons before those with higher energy. Electrons tend to seek their lowest possible energy state. The order of increasing energy of sublevels within the same principal level is as follows:



These orbitals are three-dimensional (3-D) regions of space, so they have a particular shape. The *s* orbitals are spherical (like a ball with the nucleus at the center of the ball) and the *p* orbitals are bow-tie shaped (or a figure eight that has been rotated to form a 3-D figure eight), or like a bowling pin with both ends equal in size and rounded. For *p* orbitals, the nucleus is at the center, where the two lobes meet. The sizes of these orbitals get progressively larger as the *n*-value increases (Figures 4.8 and 4.9).

There is only one *s*-orbital in any principal level. There are no *p* orbitals in the $n = 1$ principal level. In all other principal levels, there are three *p* orbitals, each perpendicular to the other two. We consider them lying along the three perpendicular axes, *x*, *y*, and *z*, and refer to them as the p_x , p_y , and p_z orbitals (Figure 4.10).

We are nearly ready to describe the electron picture within the atoms of each individual element. The only questions remaining are, in what order do

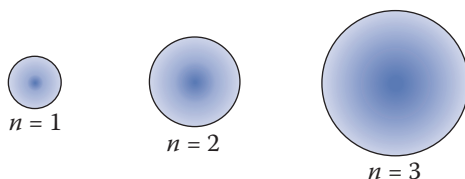


FIGURE 4.8 The *s* orbitals are spherical in shape; the further the orbitals are from the nucleus, the larger the sphere.

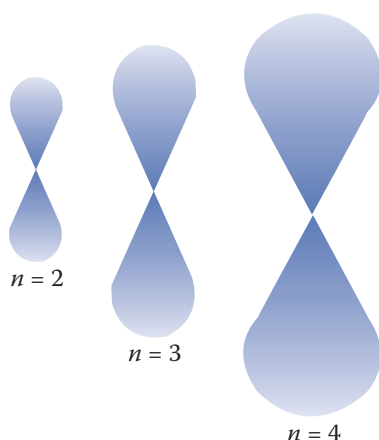


FIGURE 4.9 The p orbitals have the shape of a three-dimensional figure eight. Again, the further the orbitals are from the nucleus, the larger the figure eight.



FIGURE 4.10 The three p orbitals within a principal level are perpendicular to each other. Thus we have the p_x , the p_y , and the p_z orbitals with a given principal level. All three have the same energy.

the sublevels within a principal fill with electrons, and in what order do orbitals within the sublevels fill with electrons? The answer to the latter question is that all orbitals within a given sublevel are equal in energy. Thus deciding which orbital gets an electron is arbitrary. However, all orbitals within a given sublevel must get one electron before any get two. This is a statement of what has become known as **Hund's rule**. The application of Hund's rule is depicted in the next section in the use of orbital diagrams. When an orbital has two electrons, the electrons are said to be paired. When an orbital has just one electron, that electron is said to be unpaired.

4.6 Orbital Diagrams

One way to diagram the sublevels, orbitals, and electrons is to use small squares (boxes) to symbolize orbitals and to use arrows inside the boxes to symbolize electrons. The use of arrows for electrons is valuable because we can symbolize the fact that the electrons within a given orbital have opposite spin directions by drawing two arrows pointing in opposite directions, one up and one down. An orbitals might have no electrons, one electron, or be filled with two electrons. See Figure 4.11 for the symbolism.

The designation of the spin direction is the final step in the characterization of an individual electron. No two electrons in an atom will have the same four such designations (principal level, sublevel, orbital, and spin), and hence each electron is unique.

Let us look at some examples of the boxes-and-arrows orbital diagrams. We will begin with hydrogen and proceed through the first 10 elements in the periodic table. To label each sublevel, we will place symbols such as $1s$, $2s$, and $2p$ under the boxes. This will identify each set of orbitals as being in the $n = 1, 2, 3$, and so on principal level and in the s, p, d , or f sublevel. Remember that the electrons will go into their lowest possible energy state. The principal levels increase in energy in the order $n = 1, 2, 3$, and so on, and the sublevels within the same principal level increase in energy in the order s, p, d , and f . Only the s sublevel occurs in the $n = 1$ principal level (i.e., in the first period [horizontal row] of the periodic table). See Figure 4.12 for the orbital diagrams of the first four elements.

Helium represents a situation in which the $n = 1$ principal level is filled and the s sublevel is also filled. When we proceed to lithium and beryllium in the second period of the periodic table, the $n = 2$ principal level and the s sublevel within the $n = 2$ principal level must be used for the third and fourth electrons, as shown in Figure 4.12.

The $n = 2$ principal level has a p sublevel as well as an s sublevel, so beryllium does not represent a filled principal level. When we proceed to boron, the



FIGURE 4.11 Atomic orbitals with their electrons can be symbolized by boxes and arrows. A given orbital may have zero, one, or two electrons.

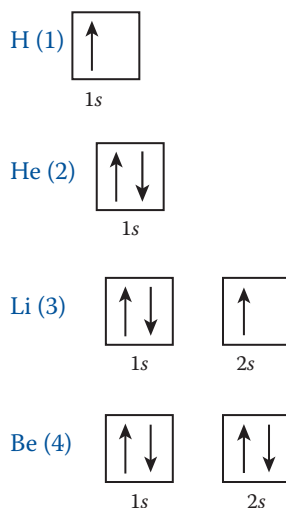


FIGURE 4.12 Orbital diagrams of the first four elements in the periodic table.

fifth electron must be shown in the $2p$ level. Remember that p sublevels have three orbitals. In this boxes-and-arrows symbolism, if a p sublevel is used, all three p orbitals are shown whether or not they have electrons. See Figure 4.13 for orbital diagrams of elements 5 through 10.

As stated previously, the three p orbitals within a principal level are all equal in energy, and since the assignment of an electron to a p orbital is arbitrary, the individual p orbitals are not given individual labels. The only distinction we have for these orbitals is that each is said to lie along an axis (x , y , or z) and that they are thus perpendicular to each other (Figure 4.10). With boxes and arrows, then, there is no distinguishing feature noted.

In the diagram for carbon and nitrogen, Hund's rule must be invoked. Thus the p orbitals each get an electron before any get two. See Figure 4.13. Following Hund's rule and filling the $2p$ level before moving on to the $n = 3$ level results in the diagrams for boron, carbon, nitrogen, oxygen, fluorine, and neon seen in Figure 4.13.

Elements beyond neon in the third period of the periodic table involve the $n = 3$ principal level, the $3s$ and $3p$ sublevels. Examples include the sodium and sulfur diagrams shown in Figure 4.14. Notice that in all cases, the total number of arrows shown is the same as the total number of electrons in the neutral atom, which is the same as the atomic number of the element.

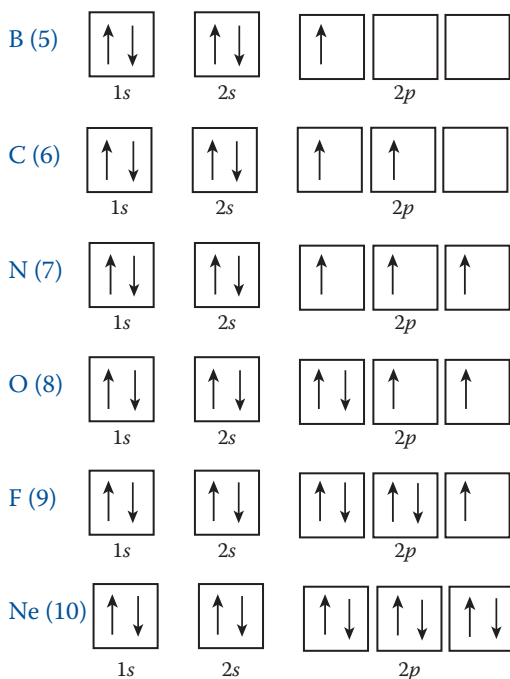


FIGURE 4.13 Orbital diagrams of the elements with atomic numbers 5 through 10. Notice that when filling a p sublevel, each p orbital gets one electron (arrow) before any get two.

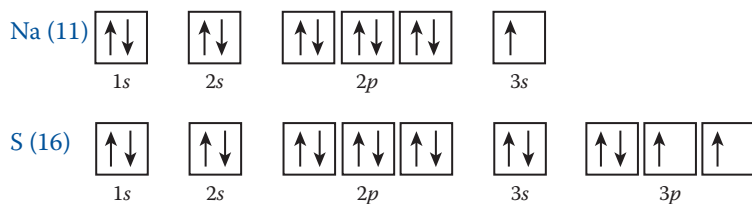


FIGURE 4.14 Orbital diagrams of two of the elements in period 3. Notice that the $n = 3$ orbitals fill with electrons in the same way that the $n = 2$ orbitals are filled.

Let us summarize these facts. Looking at Figure 4.15 and proceeding left to right across the first period in the periodic table, the $n = 1$ principal level fills with just two electrons and these two electrons go into an s orbital within an s sublevel. Proceeding across the second period, the $n = 2$ principal fills with eight electrons, two in an s orbital within an s sublevel and two in each

of three p orbitals within a p sublevel. Looking at Figure 4.16 and proceeding across the third period of the periodic table, we see that the $n = 3$ principal level begins to fill with electrons in exactly the same way that the $n = 2$ level filled in Figure 4.15. We begin to see that in the periodic table the elements are arranged according to how the electron energy levels are configured around the nucleus.

Elements beyond argon show some unexpected behavior, particularly in the order in which the sublevels are filled. The situation is, however, easily explained, as we shall see in the next section.

4.7 Electron Configuration

Besides orbital diagrams, another technique is commonly used to indicate electron structure. It is known as **electron configuration**. Electron structure by electron configuration is not as comprehensive as the electron structure by orbital diagram method as discussed in the previous section. It does not

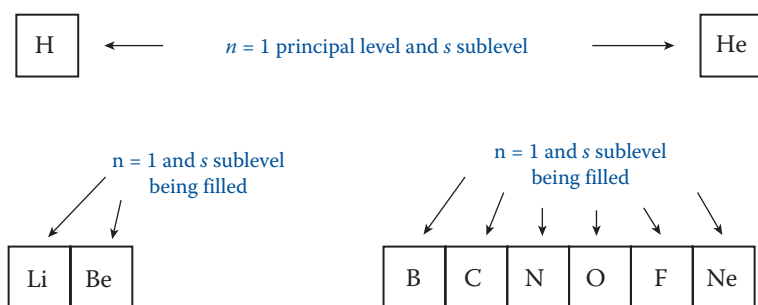


FIGURE 4.15 The first period in the periodic table is two elements wide, hydrogen and helium. An s sublevel is being filled. In the second period, both an s sublevel and a p sublevel are filled.

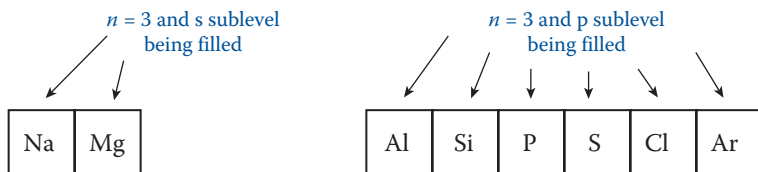


FIGURE 4.16 Period 3. It is apparent that the periodic table is arranged according to the electron arrangement within the atoms of the elements.

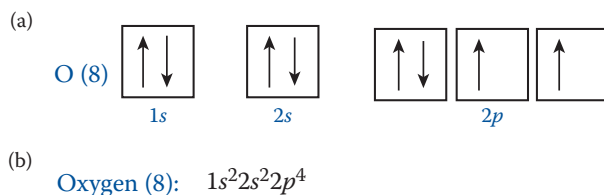


FIGURE 4.17 (a) The orbital diagram of oxygen. (b) The electron configuration of oxygen.

- H (1): $1s^1$
- Beryllium (4): $1s^2 2s^2$
- Carbon (6): $1s^2 2s^2 2p^2$
- Fluorine (9): $1s^2 2s^2 2p^5$
- Magnesium (12): $1s^2 2s^2 2p^6 3s^2$
- Sulfur (16): $1s^2 2s^2 2p^6 3s^2 3p^4$

FIGURE 4.18 Examples of electron configuration. Notice that the superscripts add up to the atomic number of the element.

designate orbital or spin, but only principal level, sublevel, and number of electrons in each sublevel. It uses the same symbolism used earlier (e.g., $1s$, $2s$, $2p$), and indicates the number of electrons in each sublevel as a superscript following the sublevel symbol (e.g., s and p). For example, Figure 4.17 shows how the orbital diagram of oxygen compares with the electron configuration of oxygen. Additional examples of electron configuration are shown in Figure 4.18.

Beyond argon, some unexpected things begin to occur. Equation 4.2 predicts that there can be up to 18 electrons in the $n = 3$ principal level, which means that while the s and p sublevels within the $n = 3$ principal level are filled when we get to argon, there is still room for 10 more electrons in this level. With potassium and beyond, we might expect that the $3d$ sublevel within the $n = 3$ principal would be next in line to receive electrons. However, that is not the case. The correct electron configuration for potassium (and also for calcium) shows that the $4s$ sublevel is used before the $3d$.

The explanation for this is simply that the $4s$ sublevel, while in a higher principal level ($n = 4$) than the $3d$ sublevel ($n = 3$), is actually lower in energy because the two principal levels overlap. The overlap is such that the $4s$ sublevel, the lowest energy sublevel in the $n = 4$ principal level, is lower in energy than the highest energy sublevel in the $n = 3$ principal level, the $3d$

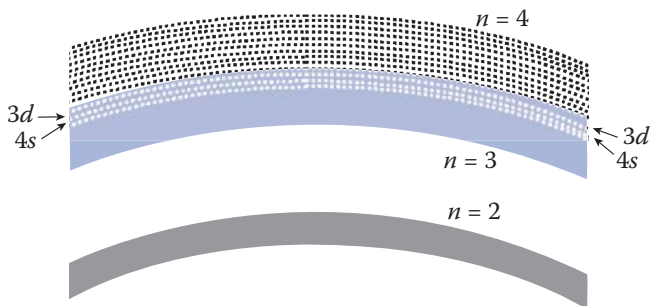


FIGURE 4.19 A drawing illustrating the overlap between the $n = 3$ and the $n = 4$ principal levels. Notice that because of the overlap, the $4s$ sublevel is actually lower in energy than the $3d$ sublevel. This is a more precise drawing than the drawing in Figure 4.7.

- Argon, Ar (18): $1s^2 2s^2 2p^6 3s^2 3p^6$
- Potassium, K (19): $1s^2 2s^2 2p^6 3s^2 3p^6 4s^1$
- Calcium, Ca (20): $1s^2 2s^2 2p^6 3s^2 3p^6 4s^2$
- Scandium, Sc (21): $1s^2 2s^2 2p^6 3s^2 3p^6 4s^2 3d^1$
- Titanium, Ti (22): $1s^2 2s^2 2p^6 3s^2 3p^6 4s^2 3d^2$

FIGURE 4.20 The $4s$ sublevel fills with electrons before the $3d$ sublevel fills.

level. Figure 4.19 should clarify how this can be. Electron configurations of the elements argon, potassium, calcium, scandium, and titanium are shown in Figure 4.20.

Overlaps that affect the order of filling are prevalent into the $n = 4$ level and beyond. See Figure 4.21. It would appear that the order of filling in these higher energy levels becomes quite complicated. It is helpful to use the mnemonic device illustrated in Figure 4.22 to remember the order of electron filling in the various sublevels. This method involves following the arrows starting with the top-most row ($1s$) and continuing downward to the lowest. The complete order of filling then is $1s \rightarrow 2s \rightarrow 2p \rightarrow 3s \rightarrow 3p \rightarrow 4s \rightarrow 3d \rightarrow 4p \rightarrow 5s \rightarrow 4d \rightarrow 5p \rightarrow 6s \rightarrow 4f \rightarrow 5d \rightarrow 6p \rightarrow 7s \rightarrow 5f \rightarrow 6d$. With this sequence, we can draw the complete electron configuration of all elements in the periodic table.

As you may imagine at this point, the periodic table itself can be used to determine electron configurations. Please refer to Figure 4.23 for this discussion. Starting in the upper left corner (with hydrogen, period 1), proceed across to the right, then move to period 2, proceed to the right, and so on. After filling

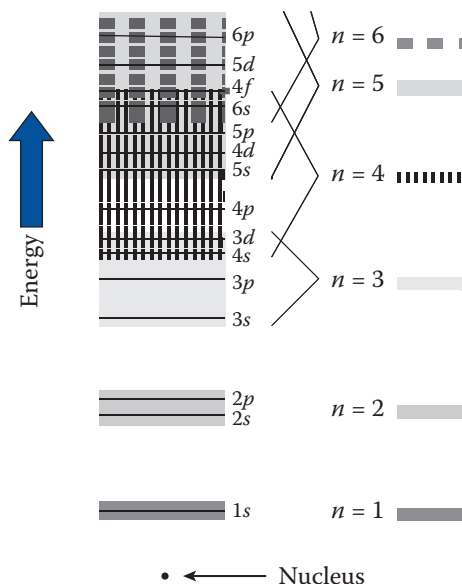


FIGURE 4.21 A diagram illustrating the complexity of the overlap that occurs between the various energy levels.

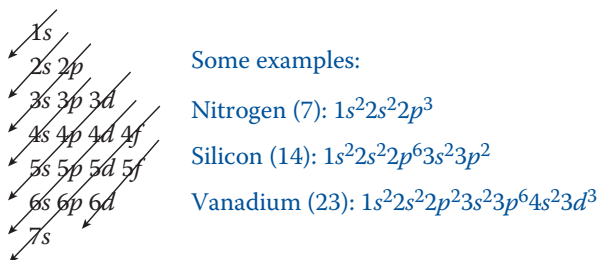


FIGURE 4.22 On the left, the mnemonic device for determining the order with which energy sublevels fill with electrons. On the right are some examples of electron configuration determined using this device.

the 6s, proceed to the 4f as indicated and then back to the 5d, and similarly from the 7s to the 5f and then to the 6d. The order of filling determined in this way is the same as determined from Figure 4.22: $1s \rightarrow 2s \rightarrow 2p \rightarrow 3s \rightarrow 3p \rightarrow 4s \rightarrow 3d \rightarrow 4p \rightarrow 5s \rightarrow 4d \rightarrow 5p \rightarrow 6s \rightarrow 4f \rightarrow 5d \rightarrow 6p \rightarrow 7s \rightarrow 5f \rightarrow 6d$.

Different blocks of elements in the periodic table have been given names based on what sublevels are being filled in these locations. We have the s-block, the p-block, the d-block, and the f-block. See Figure 4.24. For yet

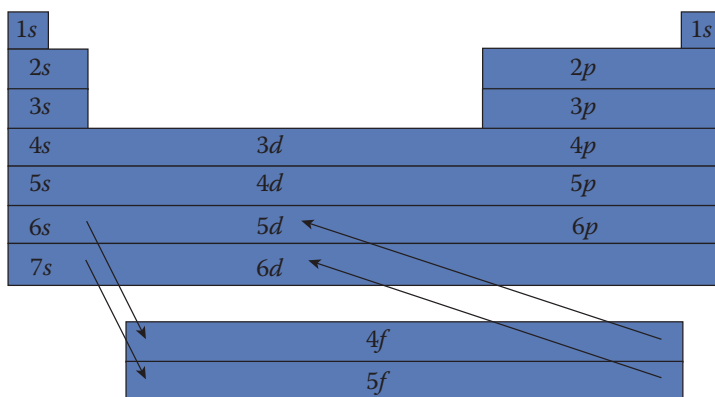


FIGURE 4.23 The order in which the energy levels fill with electrons can be deduced from the periodic table. The elements in the periodic table are arranged according to electron configuration.

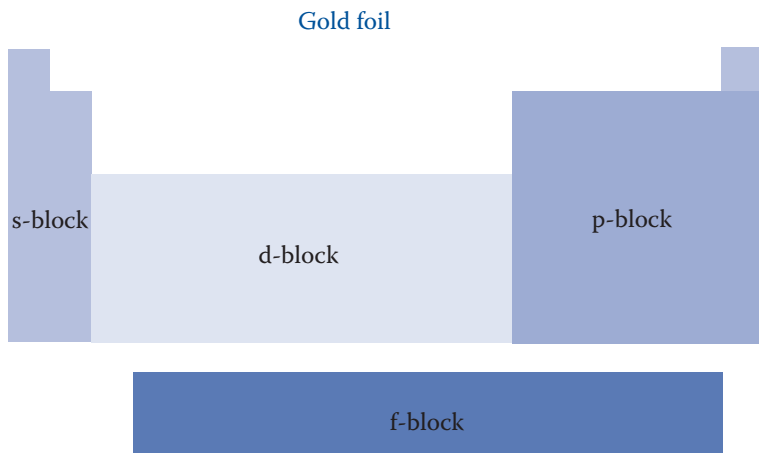


FIGURE 4.24 An outline of the periodic table showing the locations of the s-block, the p-block, the d-block, and the f-block. Note that helium in the upper right is an s-block element.

another perspective on the use of the periodic table to determine electron configuration, see Figure 4.25.

Here is one more slight complication. There are some interesting idiosyncrasies that occur *within* the *d* and *f* sublevels that warrant mentioning. The first of these occurs in the *3d* sublevel. As you proceed across period 4, in the range of the transition elements (beginning with scandium, Sc) the electron

	Period 1: 1s sublevel filled	
	Period 2: 2s and 2p sublevels filled	
	Period 3: 3s and 3p sublevels filled	
Period 4: 4s, 3d and 4p sublevels filled		
Period 5: 5s, 4d and 5p sublevels filled		
Period 6: 6s, 5d and 6p sublevels filled		
Period 7: 7s, 6d and 7p sublevels filled		
Period 6: 4f sublevel filled		
Period 7: 5f sublevel filled		

FIGURE 4.25 Another outline of the periodic table providing further information about each period in terms of what principal levels and sublevels are being filled left to right.

configurations are those shown in Table 4.3. Notice what happens when proceeding from vanadium to chromium and also from nickel to copper. For chromium, the *d* sublevel appears to “borrow” an electron from the *s* sublevel so as to leave only one electron in the *s* sublevel and to have five electrons (rather than the expected four) in the *d* sublevel. The borrowing when proceeding from nickel to copper also leaves one electron in the *s* sublevel while filling the *d* sublevel with 10 electrons.

The 4*s* and the 3*d* sublevels are extremely close in energy. It takes only a minor event to move an electron from one to the other. The minor events in these examples are the filling or half-filling of the *d* sublevel. If borrowing an electron from the *s* sublevel will result in either half-filling the *d* sublevel with five electrons (in the case of chromium) or filling the *d* sublevel with 10 electrons (in the case of copper), then the borrowing will occur as seen in Table 4.3. This kind of unexpected behavior also occurs in the other *d* sublevels and in the *f* sublevels.

One result of this latter phenomenon is the variability of ionic charge (Chapter 3) observed for the transition and inner transition elements. A good example is iron. In Chapter 3, we mentioned that iron sometimes gives up two electrons, forming the iron(II) ion, Fe^{+2} , while other times it gives up three electrons, forming the Iron(III) ion, Fe^{+3} . Iron’s electron configuration in Table 4.3 gives one possible explanation. When iron

Table 4.3 Electron Configurations of the 10 Transition Elements in Period 4 of the Periodic Table

Element (Symbol), Atomic Number	Electron Configuration
Scandium (Sc), 21	$1s^2 2s^2 2p^6 3s^2 3p^6 4s^2 3d^1$
Titanium (Ti), 22	$1s^2 2s^2 2p^6 3s^2 3p^6 4s^2 3d^2$
Vanadium (V), 23	$1s^2 2s^2 2p^6 3s^2 3p^6 4s^2 3d^3$
Chromium (Cr), 24	$1s^2 2s^2 2p^6 3s^2 3p^6 4s^1 3d^5$
Manganese (Mn), 25	$1s^2 2s^2 2p^6 3s^2 3p^6 4s^2 3d^5$
Iron (Fe), 26	$1s^2 2s^2 2p^6 3s^2 3p^6 4s^2 3d^6$
Cobalt (Co), 27	$1s^2 2s^2 2p^6 3s^2 3p^6 4s^2 3d^7$
Nickel (Ni), 28	$1s^2 2s^2 2p^6 3s^2 3p^6 4s^2 3d^8$
Copper (Cu), 29	$1s^2 2s^2 2p^6 3s^2 3p^6 4s^1 3d^{10}$
Zinc (Zn), 30	$1s^2 2s^2 2p^6 3s^2 3p^6 4s^2 3d^{10}$

forms the Fe^{+2} ion, one can imagine that it is the two $4s$ electrons that are lost, while when it forms the Fe^{+3} ion, which is a more stable ion, it is reasonable to suggest that the two $4s$ electrons are lost along with one $3d$ electron because the loss of the one $3d$ electron results in this d sublevel being half-filled.

Figure 4.21 illustrates the principal levels and sublevels through the $6p$ sublevel and represents a summary of these concepts. Notice that the overlap of principal levels is apparent, as well as the fact that some sublevels are so close in energy that movement up or down the energy scale even by the smallest increment can result in the reordering of the levels and the idiosyncrasies that we discussed.

4.8 Homework Exercises

History of Atomic Theory

- In five words or less, tell what each of the following scientists is famous for as far as atomic theories are concerned.
 - Dalton
 - Crookes
 - Thomson
 - Millikan

- e. Rutherford
 - f. Chadwick
 - g. Bohr
 - h. Schrodinger
2. Match the item on the left with the name on the right by placing the appropriate letter on the line.
- | | |
|--|-----------------|
| _____ Solar system model | (a) Schrodinger |
| _____ Nuclear model | (b) Dalton |
| _____ Original atomic theory (early 1800s) | (c) Bohr |
| _____ Quantum mechanical model | (d) Thomson |
| _____ Plum pudding model | (e) Rutherford |
| _____ Neutron | (f) Crookes |
| _____ Electron | (g) Chadwick |
| _____ Electron charge | (h) Millikan |
3. What experimental evidence did Rutherford have that led to the conclusion that the nucleus exists?

Energy, Light, and the Bohr Model

- 4. If energy is the capacity to cause matter to move due to an applied force, how can a boulder at rest on the edge of a cliff be considered to have energy? What kind of energy is it?
- 5. Name six different forms of energy. Give an example scenario for each in which each form or energy is transformed into another form.
- 6. Light is said to have a dual nature. Explain what that means.
- 7. What is different about light waves that allow them to travel through the void of space?
- 8. Define wavelength. Is a long wavelength high energy or low energy?
- 9. Which of the following are of sufficient energy to cause some kind of damage to our bodies: visible light, gamma rays, infrared light, radio waves, ultraviolet waves, microwaves, X rays? Briefly give details.
- 10. Visible light consists of the colors of the rainbow. Arrange the following in the order of increasing energy: yellow, red, green, blue, orange, violet.

11. If a wavelength of visible light is 450 nm, what color is it? Answer this same question for the following wavelengths: 600nm, 550 nm, 700 nm, 350 nm.
12. How are the different colors of light emitted by elements under certain conditions explained by the Bohr model?
13. What basic premise of the Bohr model was discounted by the modern theory?

Quantum Mechanical (Schrodinger) Model

14. What is wrong with this statement: "In the quantum mechanical model, the electron is pictured as a particle in an orbit around the nucleus much like a planet is in an orbit around the sun."
15. Fill in the blanks.
 - a. The maximum number of electrons that can be found in the $n = 3$ principal level is _____
 - b. There are _____ orbitals are in the $n = 3$ principal level.
 - c. The maximum number of electrons that can be found in a $2p$ sublevel is _____.
 - d. The maximum number of electrons that can be found in a single d orbital is _____.
16. How many orbitals are in the $n = 4$ principal level? How many are in the $n = 5$ level?
17. Sublevels within principal levels are designated by the letters s , p , d , and f . Within a given principal level, which of these has the most energy? Which of these can hold 10 electrons? Which of these has only 3 orbitals?
18. State the maximum number of electrons each of the following can hold.
 - a. A single f orbital
 - b. The $3p$ sublevel
 - c. The $n = 4$ principal level
 - d. Any s sublevel
19. Fill in the blanks with one of the following:
principal level sublevel orbital
 - a. A region of space around the nucleus in which a maximum of two electrons can be found is the _____.

- b. An energy level designated by n is the _____.
 - c. An energy level that is within a larger energy level is the _____.
20. What is the maximum number of electrons that can be found in each of the following?
- a. The $n = 3$ level
 - b. A single d orbital
 - c. A p sublevel
 - d. The $4s$ level
21. How many different orientations within an energy sublevel can a p orbital have and what are they? How do they compare in terms of energy? Where is the nucleus located?
22. What atomic orbital is spherical in shape? What is the lowest principal energy level in which this orbital is found? How many different orientations can it have?

Orbital Diagrams and Electron Configuration

23. The orbital diagrams of the first 11 elements are given in Figures 4.12, 4.13, and 4.14. If the identity of the element represented by any one of these was not known, how could you determine what element it represents?
24. How many arrows would be shown in the complete orbital diagrams of each of the following elements?
- a. Magnesium
 - b. Aluminum
 - c. Silicon
 - d. Phosphorus
 - e. Chlorine
 - f. Argon
25. How many arrows would be shown in the $4p$ sublevel of the orbital diagram of (a) gallium, (b) arsenic, and (c) krypton.
26. How many arrows would be shown in the $4d$ sublevel of the orbital diagram of (a) zirconium, (b) ruthenium, and (c) silver?
27. Fill in the blanks in the following table.

	Element	Electron Configuration
(a)		$1s^2 2s^2 2p^6 3s^2 3p^3$
(b)	Boron	
(c)		$1s^2 2s^2 2p^6 3s^2 3p^6 4s^2 3d^3$
(d)	Cl	
(e)		$1s^2 2s^2 2p^4$
(f)	Nickel	
(g)		$1s^2 2s^2 2p^6 3s^2 3p^6$
(h)	K	
(i)		$1s^2 2s^2 2p^6 3s^2 3p^6 4s^2 3d^{10} 4p^6 5s^2 4d^{10} 5p^2$
(j)	Magnesium	
(k)		$1s^2 2s^2 2p^6 3s^2 3p^6 4s^2 3d^6$
(l)	Ne	
(m)		$1s^2 2s^2 2p^6 3s^2 3p^4$
(n)	Beryllium	
(o)		$1s^2 2s^2 2p^6 3s^2 3p^6 4s^2 3d^{10} 4p^6 5s^2 4d^{10} 5p^6 6s^1$

28. In terms of energy level overlap, explain why the $4s$ level gets electrons before the $3d$ level when building electron configurations.
29. What does it mean for an energy sublevel to be half-filled? Explain what the half-filled condition has to do with the unexpected electron configuration of chromium, for example.

Chapter 5

Periodic Table

Introduction	120
Mendeleev and Meyer	120
Descriptive Chemistry of Selected Elements.	121
Periodicity and Trends: Atomic Size, Ionization Energy, and Electron Affinity.	128
Homework Exercises	135

5.1 Introduction

The periodic table is an icon that people associate with chemists and chemistry, or the work that chemists do. Indeed, a wealth of information important to chemists is available from the modern periodic table. Some of this information we have already discussed.

We introduced the periodic table in Chapter 1, mentioning first that the names and symbols of the pure substances known as elements are listed there. Such a list is important because all matter is composed of these elements either in their pure state or in physical or chemical combination. In addition, we have observed that the periodic table presents a rather odd arrangement of the elements, but the manner of this arrangement provides us with additional useful information. For example, we have noted that the elements are subdivided into groups. These groups include metals, nonmetals, and metalloids; representative, transition, and inner transition elements; alkali metals and alkaline earth metals; and halogens and noble gases (for a review of these groups, see Section 3.1). These group names have since been used in our discussions of compound formation (Chapters 3 and 4), the nomenclature of compounds (Chapter 3), and atomic structure (Chapter 4). One reason for assigning elements to groups is that all elements in a group tend to behave the same way. For instance, all alkali metals tend to form monatomic ions with a +1 charge.

In Chapter 1, we also defined two important properties of the elements, the atomic number and the atomic weight. Both of these properties are listed in the modern periodic table. The atomic weight is a property that will be especially important as we introduce some quantitative concepts of chemistry in later chapters.

Because of the importance of the periodic table to chemists, the general public's tendency to recognize the periodic table as the icon of chemists and their work is understandable.

5.2 Mendeleev and Meyer

A historical perspective on the development of the modern periodic table is both interesting and important. A Russian chemist by the name of Dimitri Mendeleev is generally credited with the development of the first periodic table, although a German chemist, Lothar Meyer, working independent of Mendeleev at approximately the same time, did much the same work. Their

work was published in 1869, long before the modern theory of the atom, and hence electron configuration, was developed. Their work involved an extensive study of the properties of the elements that were known at the time. They had observed that as one steps from one element to the next, beginning with hydrogen and increasing in atomic number, physical and chemical properties seem to repeat in a regular fashion. In other words, a pattern seems to emerge. The physical properties that they observed included such basic properties as color and density; the chemical properties they studied included the manner of reaction of an element with acid (e.g., no reaction, slow reaction, fast reaction, or violent reaction) and the formula of the compound formed between each element and chlorine. Their published results included the grouping of elements into families, each containing elements that exhibited similar properties. The result was a periodic table with periods and families that were remarkably similar to the modern periodic table.

At that time, the fact that not all the elements had yet been discovered was unknown to Mendeleev and Meyer. Mendeleev's table has only 62 elements. However, due to the regular pattern he observed, Mendeleev made some bold predictions that as yet undiscovered elements would be discovered at some point in the future and would fill certain holes he observed in his table. He even went so far as to predict specific properties for these elements. For example, the element germanium was not known at that time. There was a space in Mendeleev's table between silicon and tin in Group IVA (or Group 14). Mendeleev not only predicted the discovery of an element to fill this space, but he even went on to predict specific properties this element would have. He named this element "eka-silicon." Germanium was later discovered and was found to have properties nearly identical to what Mendeleev had predicted.

The important lesson to be learned from this story is that the electron configurations of elements are closely related to the elements' properties. The periodic table based on electron configurations is remarkably similar to the one based on properties.

5.3 Descriptive Chemistry of Selected Elements

What specifically were the properties that Mendeleev and Meyer observed leading up to the development of their periodic tables? What specific properties are we talking about when we say that elements in a given family have similar properties? What specific descriptions can be given for elements within a family

Li
Na
K
Rb
Cs
Fr

FIGURE 5.1 The alkali metals.

that would demonstrate this similarity of properties? The specific description of physical and chemical properties (and other interesting facts about elements and compounds) is called **descriptive chemistry**. We now provide a summary of the descriptive chemistries of selected families and other elements and also present at least one important fact about each of these elements.*

5.3.1 Alkali Metals

The alkali metal family is presented in Figure 5.1. Lithium, sodium, potassium, rubidium, and cesium are all highly reactive metals. They react violently with water, forming hydrogen gas and the metal hydroxide with the formula MeOH (Me symbolizing the metal), and they react readily with oxygen in the air, forming the metal oxide with the formula Me_2O , but can also form superoxides of the formula MeO_2 . These latter compounds are used in self-contained breathing devices, such as those that deep-sea divers use, to generate oxygen for breathing. Each alkali metal must be stored under mineral oil or kerosene to preserve them in the elemental state. Since they are so reactive, they are never found in nature in the free elemental state, only in compounds. Each is a silvery, soft, pliable metal that can be cut with a knife.

Lithium is used in lithium batteries because it is a light element and produces a higher battery voltage than other metals. Sodium forms a plethora of compounds used in consumer products important in our lives, including NaCl (table salt), NaHCO_3 (baking soda), and NaOH (drain cleaner, oven cleaner, soap

* Reference: Stwertka, Albert, *A Guide to the Elements*, second edition, Oxford University Press, New York, NY 2002.

Be
Mg
Ca
Sr
Ba
Ra

FIGURE 5.2 The alkaline earth metals.

making). Potassium is important for plant growth and, in the form of potassium nitrate (KNO_3) and other compounds, is a component of fertilizers. Rubidium is of little value. It is found as a trace impurity in ores of other alkali metals. The only naturally occurring isotope of cesium, cesium-133, is used as the world's official measure of time. All isotopes of francium are radioactive and unstable.

5.3.2 Alkaline Earth Metals

The alkaline earth metals are presented in Figure 5.2. Beryllium, magnesium, calcium, strontium, barium, and radium are the alkaline earth metals. They are grayish white in color and vary in hardness from being able to scratch glass (beryllium) to being as soft as lead (barium). They are chemically reactive, but less so than the alkali metals. They are, for the most part, not stable in air, reacting with oxygen to form the metal oxide with the formula MeO . In fact, magnesium, strontium, and barium will actually burn in the presence of air, with magnesium emitting an extremely bright light in the process. They also react with water to form hydrogen gas (H_2) and the metal hydroxide $[\text{Me}(\text{OH})_2]$, but not in the violent manner by which the alkali metals are known to react. Like the alkali metals, however, their reactivity is such that they are never found in the free elemental state in nature.

The precious gems of emerald and aquamarine are forms of the mineral beryl, a compound of beryllium, silicon, and oxygen. Compounds of magnesium are found in two consumer products. Magnesium hydroxide $[\text{Mg}(\text{OH})_2]$ is the active ingredient in milk of magnesia, and magnesium sulfate (MgSO_4) is the active ingredient in Epsom salt. Calcium is found in human teeth and bones, hard water, and limestone, which is an ingredient in cement. Strontium compounds are used in fireworks, emitting a red-colored light. A compound of barium (BaSO_4) is

F
Cl
Br
I
At

FIGURE 5.3 The halogens.

the white solid in the suspension we swallow so that X rays can be taken of our intestinal tract. Radium has 25 known isotopes, the most common of which is radium-226, the form of radium discovered by Marie Curie.

5.3.3 Halogens

The halogen family is presented in Figure 5.3. The halogens—fluorine, chlorine, bromine, iodine, and astatine—are a group of highly reactive nonmetals. This reactivity is the primary way in which they are similar. Fluorine (F_2) and chlorine (Cl_2) are gases, bromine (Br_2) is a liquid, and iodine (I_2) is a solid. Astatine has 20 known isotopes, and all are radioactive and highly unstable, which means that only a minute amount of this element exists. However, fluorine, chlorine, bromine, and iodine are readily available, and their properties are well known.

Fluorine is a yellowish gas and is the most reactive halogen. Its most recognized use is as sodium fluoride (NaF), which is used for tooth decay prevention in toothpastes and municipal water supplies. Chlorine is a yellow-green gas, is extremely poisonous, and has a stifling, pungent odor—the odor of common household bleach. While it is indeed useful as a bleach, the primary ingredient in household bleach is sodium hypochlorite ($NaClO$), an important compound of chlorine.

Bromine is a dark red, dense liquid with a strong peppery odor. It is highly toxic and can cause serious chemical burns. A compound of bromine, silver bromide ($AgBr$), is used in the manufacture of photographic film, as is silver iodide (AgI). Iodine is a solid existing as dark purple crystals. It is used as an antiseptic called tincture of iodine, which is a 50% solution of iodine in alcohol. A compound of iodine, potassium iodide (KI), is used as an additive to table salt to make iodized salt.

H	He

FIGURE 5.4 Period 1: hydrogen and helium.

C	N	O
	P	S

FIGURE 5.5 Some important nonmetals.

5.3.4 Hydrogen and Helium

Hydrogen and helium, the period 1 elements in the periodic table (Figure 5.4), are the lightest of all elements. Both are gases. Hydrogen (H_2) is the most abundant element in the universe, and helium is the second most abundant. The most common compound of hydrogen is water, but it is found in all organic compounds (see Chapter 14) and in acids and bases. Hydrogen gas is extremely flammable and is used as a fuel for rockets and space shuttles. A very important use of hydrogen gas is in the manufacture of ammonia (NH_3).

Helium is one of the noble gases; the others are neon, argon, krypton, xenon, and radon (Family 18, or VIIIA). These gases are described as noble or inert because they only form compounds under extraordinary conditions. They are all monatomic, which means that the fundamental unit is the atom rather than the diatomic molecule found as the fundamental unit for other gaseous elements, such as H_2 and Cl_2 . Helium is used whenever a very light unreactive gas is needed, such as in toy balloons and blimps.

5.3.5 Carbon, Nitrogen, Oxygen, Phosphorus, and Sulfur

Carbon, nitrogen, oxygen, phosphorus, and sulfur—a cluster of nonmetals in the periodic table (Figure 5.5)—are extremely important elements. Carbon is important because it is the element upon which all life is based. All organic compounds, which encompasses nearly all 13 million compounds that exist (see Chapter 14), contain carbon. Elemental carbon exists as several different allotropes, or different forms of the same element. These include diamond, graphite, and buckminsterfullerene, which is a form more recently synthesized. Inorganic compounds of carbon are also important. These include

carbon dioxide (CO_2), which in solid form is dry ice, carbon monoxide (CO), and the ionic carbonates.

Of all chemicals produced in the United States, nitrogen gas (N_2) ranks second in terms of quantity produced annually, a fact that demonstrates its importance. It is a major component of air (78%). Liquid nitrogen is used to achieve extremely cold temperatures needed for some applications, such as preserving biological systems. Compounds of nitrogen include organic compounds such as amino acids and proteins. Important inorganic compounds of nitrogen include ammonia (NH_3), nitric acid (HNO_3), and the gaseous oxides.

Oxygen is by far the most abundant element on earth, and 46% of the Earth's crust is oxygen. Oxygen gas (O_2) supports all life on our planet. Oxygen is contained in a large number of organic compounds (see Chapter 14), but it is also contained in many inorganic compounds, including water and compounds of most polyatomic ions, such as nitrates and sulfates. The most common of these is probably calcium carbonate (CaCO_3).

Phosphorus exists in two allotropic forms, white phosphorus and red phosphorus. Both are solids. The most important compound of phosphorus is phosphoric acid (H_3PO_4), which is important in fertilizer production. Phosphate compounds are used in fertilizers because compounds of phosphorus are key components of living plant cells. Phosphates have also been important ingredients of detergents because they soften hard water through the reaction of phosphate with the calcium and magnesium in the water, thus enhancing the cleaning action. However, this produces waste that has phosphates in it and such waste is not friendly to the environment, so many areas now ban the use of phosphate detergents.

Sulfur is a yellow solid that exists as several allotropes that form when sulfur is heated to different temperatures. Elemental sulfur is found in nature in underground deposits and is mined by the Frasch Process, which uses hot water to melt the sulfur so that it can be forced to the surface with compressed air. Most of the mined sulfur is used to make sulfur dioxide (SO_2). The most important compound of sulfur is sulfuric acid (H_2SO_4). This compound is made by converting SO_2 to SO_3 and then reacting the SO_3 with water. Sulfuric acid is the largest volume chemical produced in this United States.

5.3.6 Aluminum and Silicon

Aluminum and silicon are neighbors in the periodic table (Figure 5.6). Silicon is the second most abundant element in the Earth's crust, second only to

Al	Si

FIGURE 5.6 Aluminum and silicon.

Fe	Co	Ni	Cu	Zn

FIGURE 5.7 Some important first-row transition metals.

oxygen. Silicon and oxygen are found together in nature, chemically combined in compounds such as in silicon dioxide (SiO_2) in rock, sand, and soil. Elemental silicon is structurally similar to diamond, a form of carbon and an element that is in the same family as silicon. Silicon is used in computer chips, which is the reason the Silicon Valley is so named. A pure form of silicon dioxide is quartz, which has a modern use in clocks.

Aluminum is the third most abundant element in the Earth's crust. The aluminum industry is huge because aluminum metal has so many uses in consumer products. Examples are aluminum foil and beverage cans. It is manufactured by the Hall–Héroult process by electroplating it from a molten mixture of two aluminum ores, bauxite, and cryolite.

5.3.7 Selected Top-Row Transition Elements: Chromium, Manganese, Iron, Cobalt, Nickel, Copper, and Zinc

While they are not in the same family (Figure 5.7), these elements display similarities because the $3d$ sublevel is being filled and they all have the same number of electrons in the $4s$ sublevel, which can be considered the outermost level. They are all fairly unreactive metals, and their compounds are known for their striking colors.

Iron is a common, inexpensive metal has the structural integrity needed to build buildings and bridges and a host of other manufactured products. It is probably the most recognizable metal we have in terms of uses and availability. It does have undesirable corrosion properties; it rusts in the presence of air and moisture. This problem is overcome, if necessary, by combining it with other metals in a homogeneous mixture to make various steels, including stainless steel, which is typically a combination of iron, nickel, and chromium.

Compounds of iron are usually red (like rust), orange, or yellow. Iron(III) chloride, for example, is yellow.

Cobalt shares the property of magnetism with iron. Its compounds are usually either red or blue in color. Glass made with trace amounts of cobalt is blue, often called cobalt blue. It is much more rare than iron, which is why it is not as available or usable as iron.

Nickel is also rather rare. The common U.S. coin called the nickel is mostly composed of copper. It is an alloy that is about 25% nickel. An alloy of chromium and nickel, called nichrome, is a nickel–chromium alloy and is used to make heating elements and wires. Nickel is also used in nickel-cadmium (Ni–Cad) batteries, rechargeable batteries that have been used in laptop computers and other devices. Nickel compounds also display colors, typically green.

Copper is another metal recognizable to everyone. It is used to make pennies and other coins. An interesting fact about pennies is that before 1981 they were made of pure copper, but since 1981, they are composed of zinc with a copper layer deposited on the surface. Copper wire is used in electric power lines. Copper is also used in plumbing pipes. The popular alloy brass is a combination of copper and zinc, while bronze is a combination of copper and tin. Compounds of copper are typically blue or green.

Zinc is the most reactive of the transition metals, but after exposure to air it forms a coating of zinc oxide (ZnO) that protects it and makes it much less reactive. Zinc is used to coat steel to make a product called galvanized steel, which is used for such materials as household heating ducts, roofing nails, and highway guardrails. Galvanized steel does not disintegrate due to rusting and is therefore used to maintain the integrity of steel. Zinc is also used in the ordinary dry cell battery used in flashlights and other devices.

5.4 Periodicity and Trends: Atomic Size, Ionization Energy, and Electron Affinity

Have you ever wondered why the word **periodic** is used in the title of the table of elements? It is because element properties and electron configurations repeat at regular intervals as atomic number increases. In other words, they repeat *periodically* as the atomic number increases. This means that there is a regular pattern, like day and night temperatures—as we go from day to night to day to night the temperature goes down, then up, and then back down again (Figure 5.8).

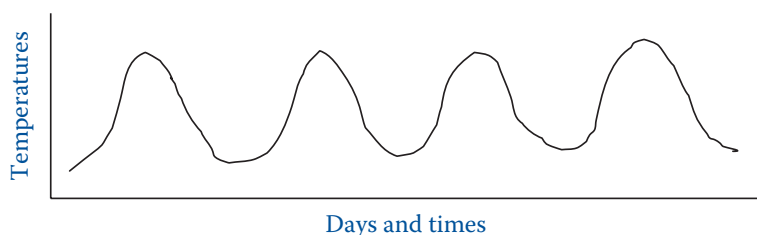


FIGURE 5.8 An illustration of a periodicity: Day–night temperature cycle.

Table 5.1 Trend in Ionic Charges

Group IA	Group IIA	Group IIIA	Group IVA	Group VA	Group VIA	Group VIIA	Group VIIIA
+1	+2	+3	+4 (−4)	−3	−2	−1	0

Note: The trend shown here is based on the need to lose and gain electrons in order to match the electron configuration of the noble gases.

From the discussions in Chapter 4, it is apparent that the elements in the periodic table are arranged according to an electron configuration pattern. Beginning at the left side in a particular period, the outermost *s* sublevel acquires one electron (alkali metal) and then a second electron (alkaline earth metal). As the right side is approached, the outermost *p* sublevel acquires six electrons, one at a time, until the right side is reached and both these *s* and *p* sublevels are filled. This pattern repeats in each and every period. In other words, the manner in which sublevels fill with electrons recurs *periodically* as the atomic number increases.

In Section 5.3, we concluded that properties are closely related to electron configuration. Element properties also recur *periodically* as the atomic number increases. Indeed, this plethora of periodic behavior is the reason that the table is called the *periodic* table, and it is also the basis for the so-called **Periodic Law**, which states that the properties of the elements are periodic functions of their atomic numbers. A good example of this is the charges that monatomic ions have. As you move across the table through the *s*-block and *p*-block, left to right, the common monatomic ion charges, as shown in Table 5.1, are +1, +2, +3, +4 (−4), −3, −2, −1, 0. Then it starts over again. See Figure 5.9. Let us examine three additional properties of elements in order to observe more evidence of this periodicity.

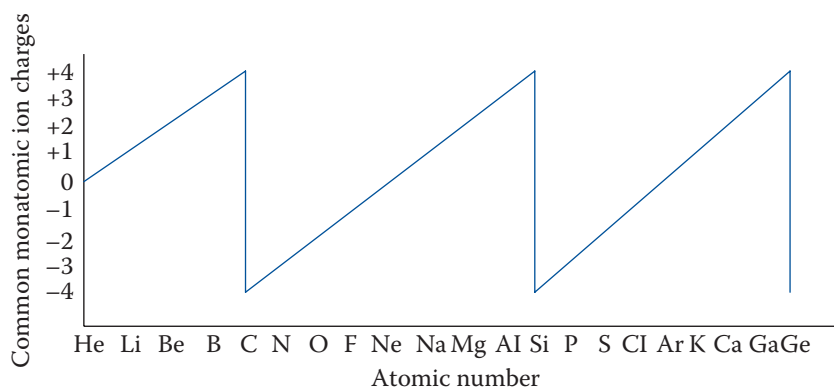


FIGURE 5.9 Another illustration of periodicity: Common monatomic ion charge vs atomic number.

5.4.1 Atomic Size

In Section 1.4.1, we defined the atom and mentioned how utterly small the atom is, indicating that it would take hundreds of thousands of even the largest of atoms lined side by side to make a spot visible to the naked eye. While atoms are extremely small particles, we would observe a significant variation in their sizes if we were to compare atoms of one element with atoms of another. In Section 1.4.1, for example, we indicated that comparing the size of a hydrogen atom (the smallest) with the size of a francium atom (the largest) would be like comparing a ping pong ball with a basketball. A francium atom is nine times the size of a hydrogen atom. Let us now examine the trends and periodicity of the sizes of the atoms of the elements.

The size of an atom is directly related (1) how many principal energy levels it has and (2) how tightly the electrons are held by the nucleus. The more principal energy levels an atom has, the more space is required to hold them and the larger the atom will be. The tighter the electrons are held, the more compact the atom becomes and the smaller the atom will be. As we move across a period from left to right, how tightly electrons are held is more important because no new principal level is required and the electrons are held more tightly due to the increased charge (more protons) in the nucleus. Thus the trend is for the atoms to decrease in size. As a new (higher) principal energy level is used at the beginning of the next period, the size becomes larger again. Thus atomic size exhibits periodic behavior: large atoms on the

left side, smaller atoms as we proceed left to right, and then larger atoms again when a new period begins.

As we move vertically down a family of elements, the number of principal energy levels becomes more important because a new principal energy level of electrons is present in each period. While the charge in the nucleus is substantially increased as we move down a family—and one may expect smaller atoms because of that—the effect of adding higher principal levels is more important.

To summarize these comments, we can draw arrows in the direction of an increase to show both the horizontal and vertical trends. A horizontal arrow pointing to the right indicates that the particular property increases as we go from left to right. If it points to the left, the property increases right to left. A vertical arrow pointing upward indicates that the property increases from bottom to top, while a vertical arrow pointing downward indicates that the property increases top to bottom. Atomic size thus exhibits the trends indicated in Figure 5.10. A periodic table picturing the relative sizes of atoms is shown in Figure 5.11.

It is apparent from Figure 5.11 that there are exceptions to the trend. For example, the size of an indium (In) atom is greater than the size of a cadmium (Cd) atom even though the general trend is in the opposite direction. Such idiosyncrasies can sometimes be traced to sublevels being filled or half-filled with electrons. For example, the cadmium-to-indium size increase is likely due to the fact that cadmium represents a filled *d* sublevel, while indium represents the beginning of the *p* sublevel and is thus a slightly larger atom.

5.4.2 Ionization Energy and Electron Affinity

Aside from having relatively large atoms, elements that have one or two electrons in their outermost level (left side of the periodic table) also lose these electrons readily. It is almost as if these electrons, which are located on the

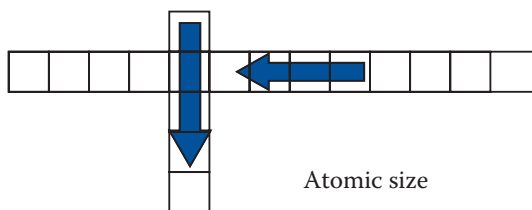


FIGURE 5.10 Atomic size increases from right to left across a period and from top to bottom in a family.

H •																			He
Li •	Be •													B •	C •	N •	O •	F •	
Na •	Mg •													Al •	Si •	P •	S •	Cl •	
K •	Ca •													Ga •	Ge •	As •	Se •	Br •	
Rb •	Sr •												Cd •	In •	Sn •	Sb •	Te •	I •	
Cs •	Ba •													Tl •	Pb •	Bi •	Po •	At •	
Fr •	Ra •																		

FIGURE 5.11 An outline of the periodic table showing the relative sizes of atoms in the s-block and p-block. Notice the trend to smaller atom as one goes left to right and from bottom to top. The relative size of the cadmium atom is also indicated in order to compare it with the indium (In) atom, as discussed in the text. (From Kenkel, J., Kelter, P., and Hage, D., *Chemistry: An Industry-Based Introduction with CD-ROM*, CRC Press, Boca Raton, FL, 2001. With Permission.)

outer fringes of these large atoms and far from the nucleus, are inviting substances that attract electrons to come by and take them away! Such atoms also have very little attraction for more electrons. Similarly, elements that have relatively small atoms and have their outermost level nearly filled (right side of the periodic table) readily take on more electrons. The outermost electrons in these atoms are also located on the outer fringes of these atoms, but, since the atoms are smaller, these electrons are located much closer to the nucleus. These atoms thus have very little tendency to give up electrons.

Elements whose atoms give up electrons readily and have a weak attraction for more electrons are located on the left side of the periodic table, while elements whose atoms do not give up electrons readily and have a strong attraction for more electrons are located on the right side of the periodic table. The energy required to remove an electron from an atom is called **ionization energy**. The attraction an atom has for more electrons is called **electron affinity**. Both increase as one moves left to right across a period, a trend opposite to atomic size (Figure 5.12). A strong attraction for electrons corresponds to

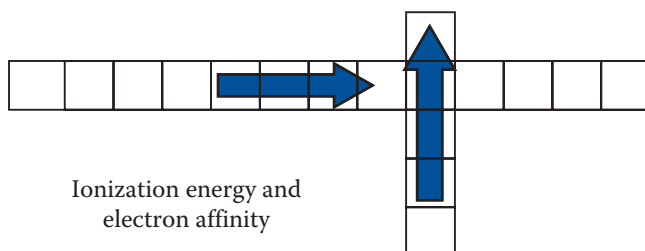


FIGURE 5.12 Ionization energy and electron affinity increase from left to right across a period and from bottom to top in a family.

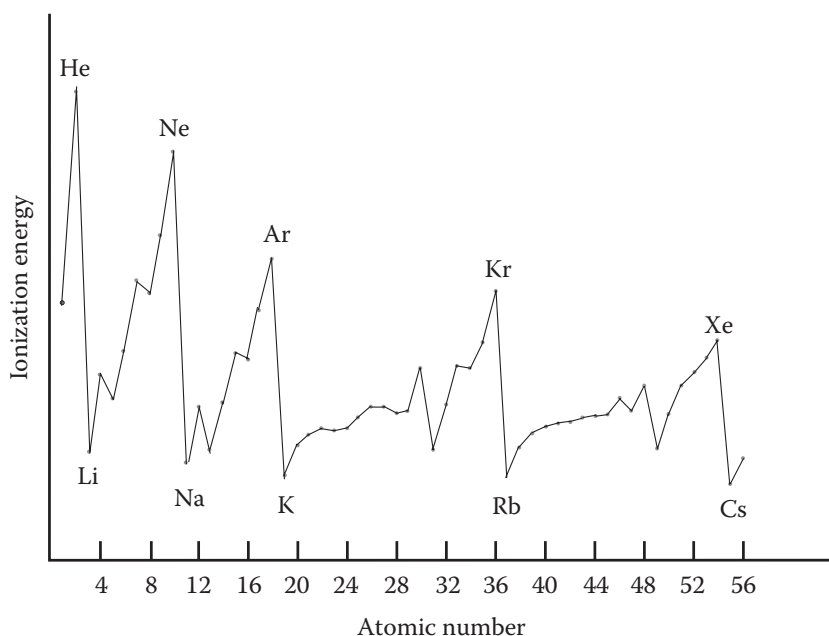


FIGURE 5.13 A graph of ionization energy versus atomic number demonstrating that ionization energy is a periodic property. (From Kenkel, J., Kelter, P., and Hage, D., *Chemistry: An Industry-Based Introduction with CD-ROM*, CRC Press, Boca Raton, FL, 2001. With Permission.)

a high ionization energy, and vice versa. When the end of a period is reached and a new period begins, both ionization energy and electron affinity sharply decrease and the increase occurring from left to right begins again. Both ionization energy and electron affinity are thus periodic properties just like atomic size. Figure 5.13, which is a graph of ionization energy versus atomic

number, presents this in graphic form where the periodic behavior is indicated by the up-and-down changes occurring as atomic number increases.

As stated above, the behavior of these properties within a family is the opposite of that observed for atomic size. Both ionization energy and electron affinity increase as we proceed from the bottom to the top within a family. The top-bottom behavior is affected by what may be called the screening effect. Atoms that have a high atomic number (near the bottom of a family) have a large number of electrons between the nucleus and the outermost level. These electrons “screen” the outer electrons from the pull of the nucleus, causing them to be loosely held and causing the atom in general to have a low ionization energy. Thus the ionization energy is high where there are fewer inner-level electrons to screen the outer-level electrons from the nucleus and low where there are more such electrons. We can make a similar statement about electron affinity. Electrons are less likely to be captured by an atom when the pull of the nucleus for such electrons is lessened by the opposing force of the inner-level electrons. They are more likely to be captured when there are fewer inner-level electrons.

Example 5.1

Which of the following has the largest atom according to periodic trends, Ca or Cl?

Solution 5.1

Atomic size increases from right to left and top to bottom in the periodic table (Figure 5.4). Calcium (Ca) is to the left and down from chlorine (Cl), so calcium has a larger atom.

Example 5.2

Which of the following has the lower electron affinity according to periodic trends, Al or Zr?

Solution 5.2

Electron affinity increases left to right and bottom to top in the periodic table (see Figure 5.7). Thus the further we go in this table to the left and down, the smaller the electron affinity will be. Zirconium (Zr) lies further

to the left and is down from aluminum (Al), so zirconium has the lower electron affinity.

Example 5.3

Arrange the following in order of decreasing ionization energy: Mn, Ba, S, Se.

Solution 5.3

Ionization energy increases from left to right and from bottom to top in the periodic table (Figure 5.7). The correct order of decreasing ionization energy would be S, Se, Mn, and Ba.

5.5 Homework Exercises

Mendeleev and Meyer

1. What is so remarkable about the work of the Russian chemist Mendeleev?
2. Mendeleev did not know anything about electron configuration, yet he was able to create a table that looked remarkably similar to the modern periodic table, which is based on electron configuration. Explain this.
3. On what did Mendeleev base the arrangement of the elements in his periodic table? On what is the arrangement of the elements in the modern periodic table based? Given that the two were based on two different things, how is it that Mendeleev's table looked so similar to the modern table?

Descriptive Chemistry

4. What is descriptive chemistry?
5. a. What might mineral oil be used for in reference to the alkali metals? Why?
b. Are any metals so soft you can cut them with a knife? If so, which?
c. Write and balance the equation for the reaction of an alkali metal of your choice with water.

6. a. Give at least one common use of each of the alkaline earth metals.
b. Write and balance the equation for the reaction of an alkaline earth metal of your choice with water.
7. Describe each of the halogens in terms of physical state and color.
8. Hydrogen and helium are both much lighter than air. Why is helium chosen over hydrogen for blimps?
9. What is an allotrope? Describe the different allotropes of carbon, phosphorus, and sulfur.
10. Give one interesting or important fact about each of the following elements: silicon, aluminum, iron, cobalt, nickel, copper, and zinc.
11. Select a specific compound of each of the following elements and give an important or interesting fact about it: sodium, magnesium, barium, fluorine, iodine, phosphorus, sulfur, silicon, and zinc.

Periodic Properties and Trends

12. When we say that a property is periodic, what does that mean?
13. Why is the table of elements called the periodic table?
14. What is the periodic law? Explain why this is important in chemistry.
15. What is the trend that is observed in atomic size as you go from left to right and from top to bottom in the periodic table?
16. What is the explanation for the observed trend in atomic size in the periodic table?
17. What is the trend that is observed in ionization energy as you go from left to right and from top to bottom in the periodic table?
18. What is the explanation for the observed trend in ionization energy in the periodic table?
19. What is the trend that is observed in electron affinity as you go from left to right and from top to bottom in the periodic table?
20. What is the explanation for the observed trend in electron affinity in the periodic table?
21. What is the screening effect? How does this affect the properties that are exhibited by atoms?
22. One reason why an atom of fluorine has a higher electron affinity compared with an atom of bromine is the screening effect. Explain.

23. Based on trends
- Which has the smallest atom, Sr or Ca?
 - Which has the lowest electron affinity, O or S?
 - Which has the highest ionization energy, Na or Mg?
 - Which has the largest atom, N or F?
 - Which has the highest electron affinity, N or As?
 - Which has the smallest ionization energy, Al or P?
24. Based on trends, write the following in order of
- Increasing atomic radius: Cr, S, Rb, K
 - Decreasing electron affinity: O, Ca, Cs, F
 - Decreasing ionization energy: Fe, Ge, Ba, F
 - Decreasing ionization energy: Hg, Fr, S, Te
 - Increasing electron affinity: Al, Tc, O, Ba
 - Decreasing ionization energy: Mo, Cs, Br, Sr
 - Increasing electron affinity: Au, S, Ba, Ir
 - Decreasing atomic size: Co, Pd, Pt, Ag
 - Increasing atomic size: Zn, O, Rb, Ag
 - Decreasing ionization energy: N, Mo, Ba
 - Increasing electron affinity: F, Sn, S, As
25. Does an element with a low ionization energy tend to form a cation or an anion? Explain.
26. In what part of the periodic table do you find elements that have relatively large atoms?
27. In what part of the periodic table do you find elements that have relatively small atoms?
28. Which has the smallest radius, K or K^+ ? Explain why.
29. Define ionization energy.

Chapter 6

Molecular Structure

Chemical Bonding	140
Outermost Electrons.	141
Simple Ionic Compounds.	142
Simple Covalent Compounds: Slot Filling	143
More Complicated Covalent Compounds	146
Coordinate Covalent Bonds	152
More Complicated Ionic Compounds	153
Structure of Ionic Compounds	155
Polar Bonds and Molecules	156
Valence-Shell Electron-Pair Repulsion Theory	162
Molecular Orbitals	165
Homework Exercises	168

6.1 Chemical Bonding

In this chapter, we present the remaining details of the topic of chemical combination, a topic that has been a point of mystery and intrigue since the opening chapter. In Chapter 1, chemical combination was given as the explanation as to why compounds, while consisting of more than one element, possess only some, or in many cases, none of the properties of those elements. We have indicated that there are forces that bind the atoms of these elements together and that these special forces, known as chemical bonding, dictate a complete transformation of properties. The result is that the compound in question possesses properties that are unique and that this new substance bears no resemblance to the elements of which it is composed. Our example in Chapter 1 was water, a liquid substance that extinguishes flames yet is composed of hydrogen, a highly flammable gaseous element, and oxygen, a gaseous element that supports combustion.

In Chapter 3, the formulas and names of compounds were introduced and, in order to establish naming schemes, a distinction was made between ionic compounds and molecular compounds. Ionic compounds were described as compounds, usually between metals and nonmetals, that are composed of formula units that include monatomic or polyatomic ions held together by electrostatic forces. Molecular, or covalent, compounds were described as compounds between two nonmetals that are not composed of ions, but of molecules—discreet units made up of atoms held together by other forces. As we have seen, metals like to give up electrons to form positively-charged ions and nonmetals like to take on electrons, producing negatively-charged ions. Ionic compounds are often considered a result of the transfer of electrons from the metal to the nonmetal. The chemical bonding that occurs in ionic compounds is called **ionic bonding**, which is the electrostatic attraction between oppositely charged ions. Molecular compounds are composed of two or more nonmetals, both of which like to take on electrons. Thus there can be no complete transfer of electrons from one element to another in these compounds. However, both of the nonmetals can achieve the same end result (i.e., a filled outermost p sub-level) by sharing two or more electrons. The chemical bonding that occurs in molecular, or covalent, compounds is called **covalent bonding**, which is the sharing of electrons between atoms.

6.2 Outermost Electrons

We have referred to the fact that electrons can be gained or lost by atoms and that ions form as a result of this gain or loss. This occurs in order for these atoms to achieve a more stable state (see Section 3.2). The electrons that are lost or gained are the so-called outermost electrons—those found in the outermost energy level, or those farthest from the nucleus. These electrons are often called valence electrons. Filling and sometimes half-filling the outermost energy levels with electrons is a stable state sought by atoms. Since all the inner levels are filled with electrons already, it stands to reason that the outermost energy level of an atom—which in most cases is not filled when the atom is in a free, uncombined state (or elemental state)—is the level that will be interacting with other atoms to satisfy the need to be filled. Because of this, a symbolism has been adopted that shows exactly what the situation is in the outermost level of atoms. This symbolism is known as the **Lewis electron-dot structure** of the atom, named after the chemist who proposed this idea, G.N. Lewis (1875–1946). His symbolism also applies to ions and molecules, as we will see in the following sections.

To draw a Lewis electron-dot structure of an atom, we only need to know the number of electrons in the outermost level of the atom and whether a given electron is alone in a given orbital or paired with another electron. The most convenient source of this information is the periodic table. As noted previously, in the first three periods of this table, the period number corresponds to the principal level being filled and the number of electrons in this level can be determined by simply counting elements starting from the left side of the table and proceeding to the right until we get to the element in question. Let us look at nitrogen as an example. Nitrogen is in period 2. Beginning with the first element in period 2, lithium, and counting to the right, we see that nitrogen is the fifth element. Thus there are five electrons to show in the Lewis structure of nitrogen.

The first two electrons in a given principal level are always in the *s* sublevel and are together (paired) in the same orbital, the spherical *s* orbital. Recalling Hund's rule from this same discussion, electrons in the *p* sublevel do not pair up until after all three orbitals get an electron. Lewis structures use the element symbol with dots around the symbol to characterize the outermost electron level of that element. If you imagine the symbol enclosed in a square, each of the four sides of the square represents an orbital. Thus one side of the square represents an *s* orbital and the other three sides represent *p* orbitals within the outermost level. See Figure 6.1. Which side we use to represent

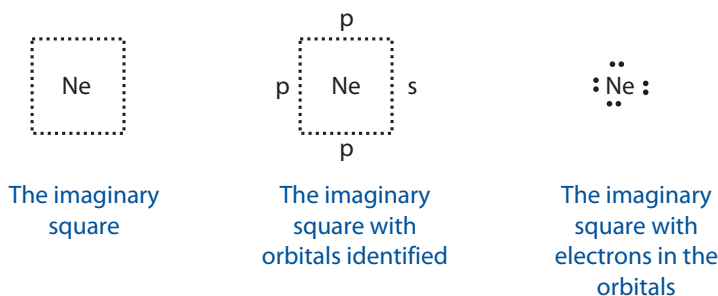


FIGURE 6.1 The concept of the Lewis electron-dot structure.

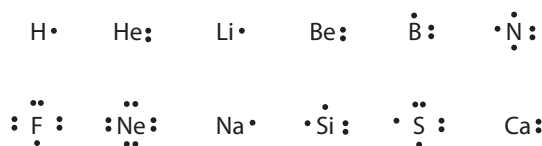


FIGURE 6.2 Lewis structures of selected elements.

each orbital is arbitrary. Figure 6.2 shows Lewis structures for some of the first 20 elements. It is important to note that elements within a given family or group (e.g., hydrogen, lithium, and sodium in Group 1) have the same Lewis structure. This should reinforce the notion that the elements are arranged in the periodic table according to their electron configurations.

6.3 Simple Ionic Compounds

Monatomic ions are easily represented in the format of Lewis electron-dot structures. First, atoms of nonmetals gain electrons when they form monatomic ions, always filling the outermost *s* and *p* sublevels in the process. The result is a total of eight electrons in these sublevels, two in each of the four orbitals. The Lewis electron-dot structure in this case shows all eight electrons, two on each of the four sides of the invisible square, plus the charge on the ion. Examples are shown in Figure 6.3. Atoms of alkali metals, alkaline earth metals, and aluminum lose all the electrons in their outermost energy level when their monatomic ions form. Thus the Lewis structures of monatomic ions of these metals simply show the positive charge that results from the loss of electrons, i.e., no dots. Monatomic ions of other metals may also be symbolized with Lewis structures as long as the charge on these ions is known. The Lewis structure of such metal ions also show no dots. Examples are shown in Figure 6.4.



FIGURE 6.3 Lewis structures of the chloride and oxide monatomic ions.



FIGURE 6.4 Lewis structures of the sodium, magnesium, and aluminum monatomic ions.

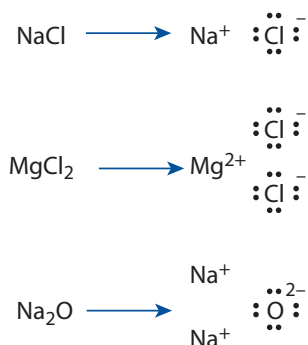


FIGURE 6.5 Lewis structures of several simple ionic compounds.

Lewis structures of ionic compounds depict the electron-dot structures of the positive ions together with the negative ions, as shown in Figure 6.5.

6.4 Simple Covalent Compounds: Slot Filling

The concept of sharing electrons in covalent bonding may also be represented by Lewis electron-dot structures. In simple cases, we recognize that the nonmetal atoms have orbital “slots” to fill, which, when filled, would give these atoms the stable noble gas configuration of eight electrons or dots. Thus Lewis electron-dot structures of some covalent compounds can be determined by a slot-filling method. In these situations, an atom has at least one orbital with only one electron, the so-called **unpaired** electrons mentioned in Section 6.2. In the process of sharing electrons, such an orbital is merged with a similar such orbital from the other atom. The result is a **molecular orbital** in which the two electrons reside. Lewis electron-dot structures show the two atoms in close proximity and horizontally adjacent to each other, with two dots representing the two shared electrons in between. An example is molecular hydrogen (H_2). Each individual

hydrogen atom in H_2 has a single electron in the outermost s sublevel. One additional electron for each would represent a filled $1s$ principal level, a stable situation that is equivalent to the stable electron configuration of helium. The Lewis electron-dot structure for H_2 shows two hydrogen atoms arranged with a pair of electrons shared between them. In Figure 6.6, two hydrogen atoms (symbolized as H) are shown as coming together to each fill an orbital slot of the other. The completed structure is shown on the right in this figure.

Lewis electron-dot structures of molecular fluorine (F_2), chlorine (Cl_2), bromine (Br_2), and iodine (I_2) are drawn in a very similar way. In these cases, one p orbital on each atom, rather than an s orbital, has a slot to fill. All other orbitals in the outermost level are already filled. The merging of two p orbitals thus occurs. This is shown for chlorine in Figure 6.7.

Simple formulas indicating combinations of unlike atoms, such as in HCl, H_2O , and NH_3 , may also be represented by Lewis electron-dot structures determined by slot filling. In these cases, the s orbital on each hydrogen merges with a p orbital on the chlorine, oxygen, or nitrogen. The slot-filling process for HCl is as shown in Figure 6.8. The process for H_2O is shown in Figure 6.9 and for NH_3 in Figure 6.10.

The covalent bonds in each of the above examples consist of a pair of shared electrons in what is called a **single bond**. It is possible and common, especially

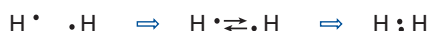


FIGURE 6.6 The process of determining the Lewis structure of hydrogen gas (H_2), showing two atoms coming together, filling each other's open orbital slot, and sharing the two electrons so that each "owns" two electrons and each has the stable configuration of helium.

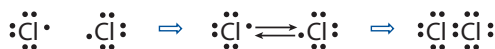


FIGURE 6.7 The process of determining the Lewis structure of chlorine gas (Cl_2) in a way similar to that of H_2 . In this case, it is a p orbital on each that has the open slot prior to bonding.

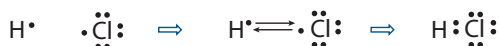


FIGURE 6.8 The slot-filling method for determining the Lewis structure of hydrogen chloride (HCl). In this case, the s orbital on the H and a p orbital on the Cl are involved.

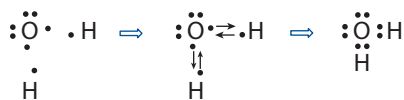


FIGURE 6.9 The slot-filling method for determining the Lewis structure of water (H_2O). The two open orbital slots on the oxygen are filled by the hydrogens.

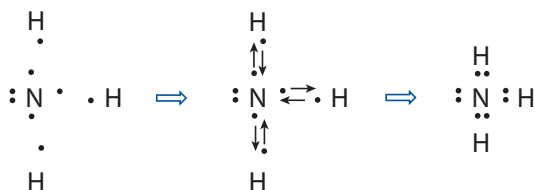


FIGURE 6.10 The slot-filling method for determining the Lewis structure of ammonia (NH_3).

among organic compounds, to have two pairs (**double bond**)—and even three pairs (**triple bond**)—of electrons shared by two atoms. To illustrate, we will consider two organic compounds, ethylene (C_2H_4) and acetylene (C_2H_2).

First, we need to point out that a hydrogen atom never forms a double bond. A hydrogen atom has only one slot to fill, as is seen in all of the foregoing examples, so it can only form one covalent bond. Therefore the double and triple bonds in ethylene and acetylene are between the two carbon atoms in each case.

Second, the carbon atoms must undergo a process called **hybridization** to form these compounds and, indeed, for any carbon compound to form, as we shall see in Chapter 14. This process is very simply shown in Figure 6.11a. On the left, we see atomic carbon's Lewis structure as we would expect to see it based on carbon's position in the periodic table. In hybridization, all of a carbon atom's outermost electrons become unpaired, as shown in the middle of Figure 6.11a. In the formation of ethylene, two hydrogen atoms then fill two of the slots, leaving two open slots as shown on the right in Figure 6.11a. The final step in the formation of ethylene is then shown in Figure 6.11b. Each carbon atom, before bonding to the other carbon atom, has two unpaired electrons and two slots to fill. Both slots are filled by the lone electrons from the other atom such that *two* molecular orbitals form. Thus a **double bond** forms. Both carbon atoms have the required eight electrons, four of which are shared in two covalent bonds rather than two electrons in just one bond.

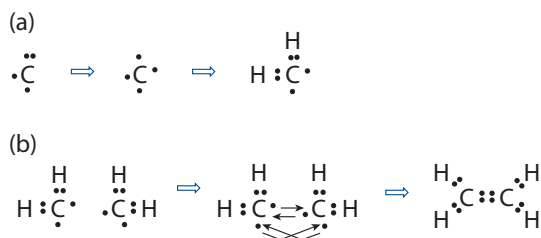


FIGURE 6.11 The slot-filling method for determining the Lewis structure of the organic compound ethylene (C_2H_4). (a) The four orbitals on each of the carbons hybridize such that all four have an open slot and all four electrons are unpaired. (b) Hydrogens then fill two of the open orbital slots while a double bond forms as shown so that the other open slots are also filled.

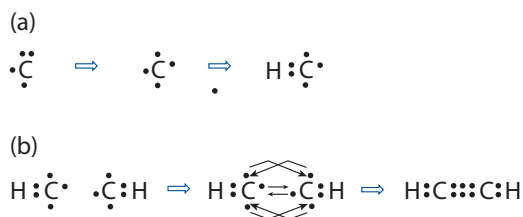


FIGURE 6.12 The formation of the triple bond in the organic compound acetylene (C_2H_2). (a) The orbitals on each carbon atom hybridize as in Figure 6.11, and (b) the other open slots are filled as shown.

Acetylene (C_2H_2) represents a similar covalent bonding example (Figure 6.12). After hybridization and after a lone hydrogen atom fills one of the slots on each carbon atom (Figure 6.12a), each carbon atom has three unpaired electrons and three slots to fill. The result is that the covalent bonding for acetylene consists of three bonds between carbons in which six electrons are being shared, a so-called triple bond. Again, both carbon atoms now have the required eight electrons—three pairs are shared between the carbon atoms and one pair is with a hydrogen atom.

6.5 More Complicated Covalent Compounds

As we progress from simple binary covalent compounds such as those discussed in Section 6.4 to other binary covalent compounds and to ternary compounds, we soon find the slot-filling procedure to be unworkable. In these examples, the atoms of interest (such as sulfur and chlorine) must have more

shared electrons than would be determined by using just slot filling. In other words, the number of such electrons in these cases cannot be determined by the slot-filling process.

To solve this problem, a set of steps is given in Table 6.1 to determine the correct Lewis electron-dot structure for these types of compounds. We now present some examples of the use of these steps.

Example 6.1

Draw the Lewis electron dot structure of carbon tetrachloride (CCl_4).

Solution 6.1

Step 1: A Carbon atom has four electrons in the outermost level, and each chlorine atom has seven electrons. The total count is $4 + 28 = 32$.

Step 2: The Carbon atom is the logical choice for the central atom. It has a lower electron affinity than does chlorine. Thus the arrangement would be as shown in Figure 6.13a.

Step 3: Giving each atom eight electrons would result in Figure 6.13b.

Step 4: The total count of electrons in step 3 is $4 \times 8 = 32$. This count matches the count in step 1. Therefore the drawing in 6.13b is the correct one.

Example 6.2

Draw the Lewis electron dot structure of chloric acid (HClO_3).

Solution 6.2

Step 1: The count is tabulated as follows:

$$\begin{array}{r}
 1(\text{H}) = 1 \times 1 = 1 \\
 1(\text{Cl}) = 1 \times 7 = 7 \\
 + \quad 3(\text{O}) = 3 \times 6 = 18 \\
 \hline
 \text{Total} = 26 \text{ electrons}
 \end{array}$$

Step 2: Since a hydrogen atom cannot be the central atom and since an oxygen atom will not bond to another oxygen atom in this case, the chlorine atom is the central atom. The oxygens are bonded

Table 6.1 Steps to Determine Lewis Electron-Dot Structure of Covalent Compounds and Polyatomic Ions That Are Too Complicated to Determine by Normal Slot-Filling

1. **Count the electrons.** Count the number of outermost electrons on each of the atoms in the formula before bonding, and add to obtain the total number. This is the number of dots that must be shown in the Lewis electron-dot structure. If the formula is that of a polyatomic ion, also add to this number the extra electrons that are represented by the charge on the ion.
2. **Arrange the atoms.** Designate one of the atoms as the central atom and place the other atoms around it in some logical fashion. If there is more than one possibility for the central atom, choose the one that has the lowest electron affinity according to periodic trends. In addition, the following points should be considered:
 - a. Hydrogen cannot be a central atom. This is logical because hydrogen becomes stable by sharing just two electrons, which means it only forms one bond with another atom. One bond means just one other atom can be attached to hydrogen. Central atoms may have from one to six atoms attached, as we shall see later. In oxyacids, each hydrogen is bonded to an oxygen.
 - b. Oxygen never bonds to another oxygen except in oxygen gas (O_2), ozone gas (O_3), and peroxide (O_2^{2-}). Except for these, oxygen will never be the central atom. Normally, an oxygen atom bonds only twice, either with two single bonds or with one double bond.
3. **Distribute the electrons.** Without regard (yet) to the count in step 1, place eight electrons (dots) around all atoms, one pair on each side of the invisible square (making only single bonds for now), except for any hydrogens, which get just two dots.
4. **Count the electrons** that were distributed in step 3 and compare this count with the count that was obtained in step 1. These counts must match. If the counts match at this point, the electron dot structure you have drawn is the correct one. If it is a polyatomic ion, place parentheses around it and write the charge to the upper right.
5. **If the counts in steps 1 and 4 do not match,** make double or triple bonds between the central atom and other logical atoms. If the count in step 4 is greater than the count in step 1, the count in step 4 must be reduced. This can be done by introducing double or triple bonding (or both). Each double bond reduces the count in step 4 by two. Each triple bond reduces the count in step 4 by four. Of course, hydrogen cannot participate in a double or triple bond.
6. **If the central atom is a phosphorus atom,** or an atom of an element to the right or below phosphorus in the periodic table, the central atom may accommodate more than eight electrons, if necessary, to make the counts in steps 1 and 4 match.

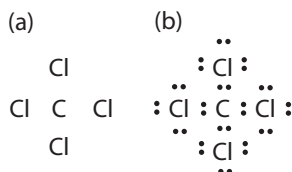


FIGURE 6.13 The process of determining the Lewis structure of carbon tetrachloride, CCl_4 , according to the steps in Table 6.1. (a) The arrangement of the atoms after completing step 2. (b) The Lewis structure after completing step 4.

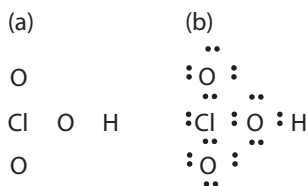


FIGURE 6.14 The process of determining the Lewis structure of chloric acid (HClO_3). (a) The arrangement of the atoms after completing step 2. (b) The Lewis structure after completing step 4.

to the chlorine atom, and the hydrogen atom is bonded to an oxygen atom. See Figure 6.14a.

Step 3: Give oxygen atoms and the chlorine atom eight electrons each. This means that the hydrogen atom has the required two, and we then have the drawing in 6.14b.

Step 4: The total count from step 3 is 26. This matches the count from step 1, so the structure drawn in step 3 (Figure 6.14b) is the correct one.

Example 6.3

Draw the Lewis electron dot structure of nitric acid (HNO_3).

Solution 6.3

Step 1:

$$\begin{array}{rcl} 1(\text{H}) & = & 1 \times 1 = 1 \\ 1(\text{N}) & = & 1 \times 5 = 5 \\ + \quad 3(\text{O}) & = & 3 \times 6 = 18 \\ \hline & & \text{Total} = 24 \text{ electrons} \end{array}$$

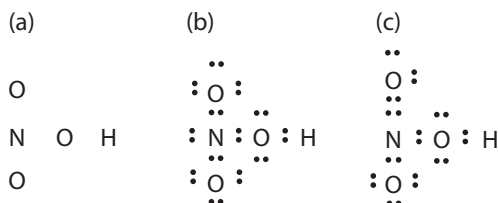


FIGURE 6.15 The process of determining the Lewis structure of nitric acid (HNO_3). (a) The arrangement of the atoms after completing step 2. (b) The Lewis structure after completing step 4. (c) In step 5, we see that a double bond is needed here since the counts compared in step 4 do not match.

Step 2: The Nitrogen atom is the only choice for the central atom, and the hydrogen atom bonds to an oxygen atom. See Figure 6.15a.

Step 3: Give the hydrogen atom two electrons and each other atom eight. This results in the drawing in Figure 6.15b.

Step 4: The total count here is 26. This does not match the count in step 1, which is 24. Therefore this cannot be the correct structure, and we need to go on to step 5.

Step 5: Let us draw a double bond between the nitrogen atom and an oxygen atom, making sure that the nitrogen atom and the oxygen atom still have eight dots each. See Figure 6.15c. Now the counts match; the drawing in Figure 6.15c is the correct one.

Note: Which oxygen atom gets the double bond is arbitrary. This bond may be shown between the nitrogen atom and any of the three oxygen atoms.

Example 6.4

Draw the Lewis electron dot structure of the carbonate ion (CO_3^{2-}).

Solution 6.4

Note: While the carbonate is an ion and is found in ionic compounds, the bonding between the carbon atom and the oxygen atoms is covalent and the Lewis structure is determined by the steps in Table 6.1. This is true of all polyatomic ions. The Lewis electron-dot structures of ionic compounds that contain polyatomic ions are discussed in Section 6.7.

Step 1:

$$\begin{array}{r}
 1(\text{C}) = 1 \times 4 = 4 \\
 3(\text{O}) = 3 \times 6 = 18 \\
 + \text{ Extra electrons} = 2 \\
 \hline
 \text{Total} = 24 \text{ electrons}
 \end{array}$$

Step 2: The Carbon atom is the choice for the central atom, so the atoms are arranged as in Figure 6.16a.

Step 3: Distribute the electrons so that each atom has eight. We now have the drawing shown in Figure 6.16b. Note the parentheses and the charge in this structure; these are necessary to show that the structure is a polyatomic ion.

Step 4: The total count in the drawing in Figure 6.16b is 26. Two electrons must be dropped for the count to equal that in step 1 (24). Therefore the drawing in Figure 6.16b is not a correct structure.

Step 5: Placing a double bond between the carbon atom and an oxygen atom produces the result in Figure 6.16c. This is the correct structure since the counts now match at 24 electrons.

Example 6.5

Draw the Lewis electron dot structure of phosphorus pentachloride (PCl_5).

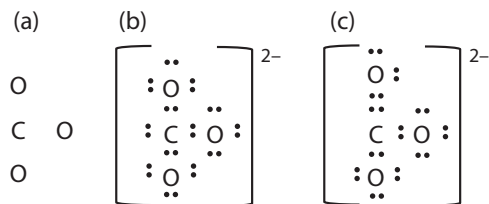


FIGURE 6.16 The process of determining the Lewis structure of the carbonate ion (CO_3^{2-}). (a) The arrangement of the atoms according to step 2. (b) The Lewis structure after completing step 4. Notice the parentheses and the charge in the upper right, since this is a polyatomic ion. (c) In step 5, we again see that a double bond is needed since the counts compared in step 4 do not match.

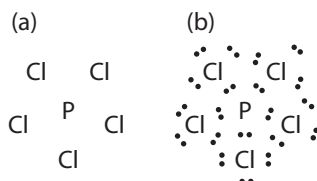


FIGURE 6.17 The process of determining the Lewis structure of phosphorus pentachloride (PCl_5). (a) The arrangement of the atoms according to step 2. (b) Step 6 is needed since the central atom must accommodate five pairs of electrons.

Solution 6.5

Step 1:

$$\begin{array}{r}
 1(\text{P}) = 1 \times 5 = 5 \\
 5(\text{Cl}) = 5 \times 7 = 35 \\
 \hline
 \text{Total} = 40 \text{ electrons}
 \end{array}$$

Step 2: The Phosphorus atom is the central atom and has five chlorine atoms around it as in Figure 6.17a.

Step 3: The Phosphorus atom can accommodate five chlorine atoms by sharing five electron pairs, or 10 electrons. Each chlorine atom will have eight electrons. This is an example of a central atom having more than eight electrons, so step 6 is used. This results in the drawing shown in Figure 6.17b.

Step 4: Forty electrons matches the count in step 1. Thus this is the correct structure.

It should be pointed out that all structures that can be drawn by slot filling may also be drawn using the steps in Table 6.1.

6.6 Coordinate Covalent Bonds

Covalent bonds form as a result of an orbital on one atom merging with an orbital on another atom such that each shares a pair of electrons with the other. With the slot-filling process, each of these orbitals has one electron before bonding occurs. However, having one electron before bonding is not a requirement of these orbitals. It is possible for one orbital to have two electrons while the other has none. A bond formed in such a manner is called a coordinate covalent bond.

An obvious example of a coordinate covalent bond is one of the N-H bonds in the ammonium ion, NH_4^+ . The ammonium ion forms when the ammonia

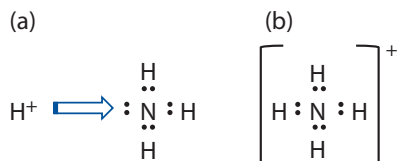


FIGURE 6.18 The formation of a coordinate covalent bond. (a) When hydrogen ions are in a solution together with ammonia molecules, the hydrogen ion moves in to share the open pair of electrons on the ammonia molecule (refer back to Figure 6.10). (b) The complete Lewis structure of the ammonium ion.

molecule, NH_3 , encounters a hydrogen ion, such as when a solution of ammonia is mixed with a solution of an acid. The hydrogen ion, which has no electrons, attaches itself to the nitrogen atom of the ammonia molecule at the site of the nonbonded pair of electrons on the nitrogen atom. This creates a total of 4 N-H bonds and a net positive charge, resulting in the ammonium ion. See Figure 6.18. Both electrons shared with the fourth hydrogen atom originated on the nitrogen atom and thus form a coordinate covalent bond.

6.7 More Complicated Ionic Compounds

We are now ready to draw Lewis electron-dot structures of ionic compounds that include at least one polyatomic ion. The polyatomic ions in such structures are drawn according to the steps in Table 6.1. The total structure is then assembled as was done for the other ionic compounds that were discussed in Section 6.3.

Example 6.6

Draw the Lewis electron-dot structure of sodium sulfate (Na_2SO_4).

Solution 6.6

Two sodium ions, Na^+ , will be shown with one sulfate ion. We draw the sulfate ion according to the procedure in Table 6.1.

Step 1:

$$\begin{array}{rcl}
 1(\text{S}) & = & 1 \times 6 = 6 \\
 4(\text{O}) & = & 4 \times 6 = 24 \\
 + \text{ Extra electrons} & = & 2 \\
 \hline
 \text{Total} & = & 32 \text{ electrons}
 \end{array}$$

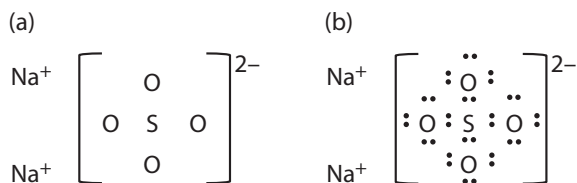


FIGURE 6.19 An example of an ionic compound sodium sulfate (Na_2SO_4) that includes a polyatomic ion, the sulfate ion SO_4^{2-} . The two sodium ions are shown and then the Lewis structure of the sulfate ion must be determined. (a) After performing step 2, including the brackets and charge. (b) The complete structure after performing step 4 for the sulfate ion.

Step 2: The sodium ions are pictured apart from the sulfate ion. For the sulfate ion, the sulfur atom is the central atom with the oxygen atoms around it. See Figure 6.19a.

Steps 3 and 4: If each oxygen atom is given eight electrons, the sulfur will have eight and the total count will be $4 \times 8 = 32$, matching the count in step 1. With the parentheses and charge in place, we have the final structure shown in Figure 6.19b.

Example 6.7

Draw the Lewis electron-dot structure of ammonium phosphate $[(\text{NH}_4)_3\text{PO}_4]$.

Solution 6.7

The three ammonium ions are drawn as noted in Section 6.6 and will be included here in steps 3 and 4. The phosphate ion must be drawn according to the procedure in Table 6.1.

Step 1:

$$\begin{array}{rcl}
 1(\text{P}) & = & 1 \times 5 = 5 \\
 4(\text{O}) & = & 4 \times 6 = 24 \\
 + \text{ Extra electrons} & = & 3 \\
 \hline
 \text{Total} & = & 32 \text{ electrons}
 \end{array}$$

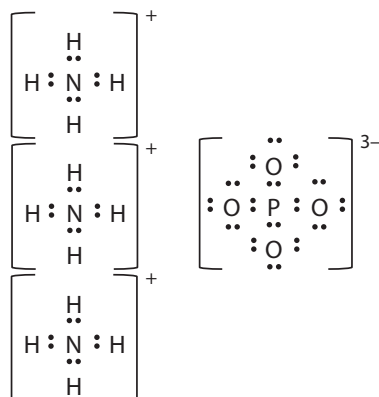


FIGURE 6.20 The Lewis structure of ammonium phosphate ($[\text{NH}_4]_3\text{PO}_4$). The Lewis structure of the phosphate was determined according to the steps in Table 6.1.

Steps 2 through 4: The phosphorus atom is the central atom, and the oxygen atoms are around it. Giving each oxygen atom eight electrons will give the phosphorus atom eight electrons and the total count will be 32, matching the count in step 1. With the ammonium ions included, the final structure is as shown in Figure 6.20.

6.8 Structure of Ionic Compounds

In Chapter 1 (Section 1.4.2), we made a distinction between the fundamental unit of an ionic compound and the fundamental unit of a covalent compound, saying that for an ionic compound the fundamental unit is the formula unit but for a covalent compound it is the molecule. Molecules are distinct individual units that exist independent of other molecules. Formula units, on the other hand, are collections of ions in the ratio indicated in the formula (and in their Lewis electron-dot structures). They do *not* exist as independent distinct units, however. The ions that make up ionic compounds, both anions and cations, exist in three-dimensional (3-D) space like a stack of marbles, with one layer placed upon another. The cations and anions interact equally with all ions in their immediate vicinity, which is more than what is shown in the formula. It

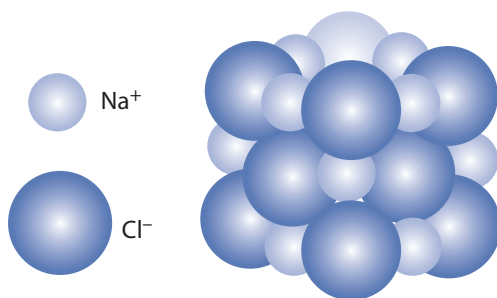


FIGURE 6.21 The structure of NaCl. Note that no single sodium ion is associated with any single chloride ion—no molecule. (From Kenkel, J., Kelter, P., and Hage, D., *Chemistry: An Industry-Based Introduction with CD-ROM*, CRC Press, Boca Raton, FL, 2001. With Permission.)

is this 3-D stack of cations and anions that make up the structure of crystals of ionic compounds. Thus, while the formula indicates the whole number ratio of cations to anions, no distinct particle can be called a molecule because there is no independent unit exemplified by the formula of the compound. The structure of NaCl is shown in Figure 6.21 as an example.

6.9 Polar Bonds and Molecules

The pair or pairs of electrons that are shared by atoms that are linked by covalent bonds occupy spaces between the two atoms that are often called molecular orbitals. A molecular orbital is the region of space in molecules that is analogous to atomic orbitals in atoms. They can be described as electron “clouds” that are located between the bonded atoms and in which the electrons reside.

6.9.1 Electronegativity and Bond Polarity

Each of the atoms taking part in a bond have an attraction for their shared electrons in a fashion similar to the attraction nonbonded atoms have for electrons. The attraction that a bonded atom has for its shared electrons is called its **electronegativity**. Linus Pauling, a prominent chemist of the twentieth century, assigned numerical values to the electronegativities of elements for the purpose of quantitatively comparing one with another. Since covalent bonds are mostly between nonmetal atoms, it is the electronegativities of nonmetals

Table 6.2 Pauling's Electronegativities of Nonmetals (Excluding Noble Gases) in Decreasing Order

Nonmetal	Electronegativity
Fluorine	4.0
Oxygen	3.5
Chlorine	3.0
Nitrogen	3.0
Bromine	2.8
Carbon	2.5
Iodine	2.5
Sulfur	2.5
Selenium	2.4
Astatine	2.2
Hydrogen	2.1
Phosphorus	2.1
Tellurium	2.1
Arsenic	2.0
Boron	2.0
Silicon	1.8

that are important for the present discussion. Pauling's electronegativities of the nonmetals (excluding the noble gases) are listed in Table 6.2.

If the electronegativities of two bonded atoms are equal, the orbital between them is symmetrically shaped and the electron charge is balanced between the two atoms. Such a bond is said to be **nonpolar**. If the electronegativities are unequal, the electrons in the bond are attracted more strongly toward the atom that is more electronegative. This means that the shared electrons are positioned more closely toward that atom. This polarizes the orbital and makes the orbital unsymmetrical because the distribution of electron charge is not balanced between the two atoms. The bond then becomes like a small magnet that is slightly negatively charged on one end and slightly positively charged on the other end. Such a bond is said to be **polar**, meaning that it has both a negative pole and a positive pole like the north and south poles of a magnet (Figure 6.22). This polarity is not to be confused with the total separation of

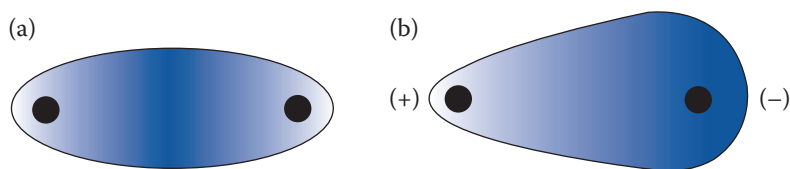


FIGURE 6.22 (a) An illustration of a nonpolar bond in which the shared electrons are distributed symmetrically in the space between the nuclei. (b) An illustration of a polar bond in which the electrons are not distributed symmetrically and slight charges appear on both ends. (From Kenkel, J., Kelter, P., and Hage, D., *Chemistry: An Industry-Based Introduction with CD-ROM*, CRC Press, Boca Raton, FL, 2001. With Permission.)

charge that is seen in ionic bonds. Polar bonds are covalent bonds that have developed a charge imbalance due to the electronegativity difference between the bonded atoms.

Covalent bonds between two atoms of the same element are nonpolar. Both atoms, being the same element, have the same electronegativity. Thus the H-H bond in H_2 is nonpolar and the Cl-Cl bond in Cl_2 is nonpolar. Also, covalent bonds between atoms that have the same electronegativity value in Table 6.2 will also be nonpolar. For example, a bond between hydrogen and phosphorus or between tellurium and boron is nonpolar.

Covalent bonds between atoms that have unequal electronegativities are polar. Examples include the H-Cl bond in HCl, the H-O bonds in water, and each of the C-H bonds in methane (CH_4).

6.9.2 Polarity of Molecules

These latter examples warrant more discussion relating to the overall polarities of entire molecules. The water molecule is extremely polar while the methane molecule is nonpolar. This is true despite the fact that both of these molecules contain polar bonds. The reason for this apparent contradiction has to do with the arrangement of the bonds within the two molecules. In methane, all bonds are C-H bonds and are arranged in a perfectly symmetrical fashion, a so-called tetrahedral arrangement. Because of this symmetry, the individual bond polarities balance out, resulting in a nonpolar molecule. In water, the H-O bonds are not symmetrically arranged and, as a result, the molecule is polar. The molecular arrangements of the methane and water molecules are given in Figure 6.23. In these ball-and-stick models, the bonded atoms are represented as balls and the bonds are represented as sticks connecting the balls.

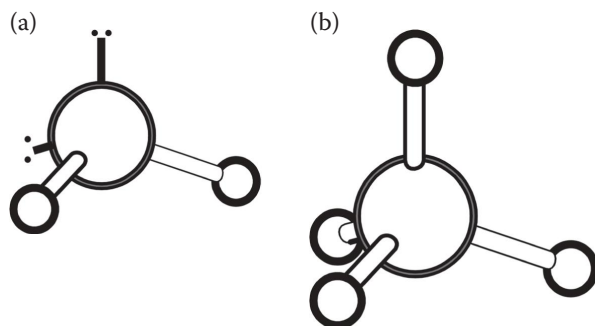


FIGURE 6.23 (a) A representation of the asymmetrical water molecule. (b) A representation of a symmetrical methane molecule. (From Kenkel, J., Kelter, P., and Hage, D., *Chemistry: An Industry-Based Introduction with CD-ROM*, CRC Press, Boca Raton, FL, 2001. With Permission.)

In Figure 6.23b, the model for CH₄ shows the carbon atom as the large ball in the center and the hydrogens as small balls connected to it. This is referred to as a tetrahedral arrangement because the four hydrogens connect to each other with imaginary lines to create a pyramid-like structure with four faces, one of which is the base of the structure. A four-faced 3-D shape of this nature is called a tetrahedron. The carbon atom is in the center of the tetrahedron (in a location where an Egyptian pharaoh's tomb may be found in a pyramid), and the hydrogen atoms are located in the corners. As such, it is quite a symmetrical structure—so symmetrical that the individual C–H bond polarities exactly balance with the others, giving a nonpolar molecule.

The general structure of the water molecule (Figure 6.23a) is also a tetrahedron. In this structure, the oxygen atom is located in the center and the two hydrogen atoms are located in two of the tetrahedral corners. The other two corners are taken up by the unshared pairs of electrons that the oxygen atom in water has. If one does not count the unshared electrons in a structure as part of the molecular geometry, the water molecule would not be referred to as tetrahedral, but as angular. This is a highly unbalanced, nonsymmetrical molecule; the individual H–O bond polarities do *not* balance, and thus the molecule is polar—*very* polar, in fact.

6.9.3 Water and the Hydrogen Bond

The polar/nonpolar nature of molecules dictate a host of important properties that compounds have. Because water molecules are polar, they possess a strong

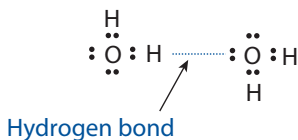


FIGURE 6.24 An illustration of the hydrogen bond that occurs between two water molecules.

attraction for each other. A container filled with water is like a container filled with small magnets. It takes much energy to separate these molecules from each other. For example, this means that water does not tend to evaporate as readily as its counterparts with nonpolar molecules, such as methane. Water's boiling point is quite high compared with the boiling point of other substances with similar size and weight. The boiling point of water is 100°C , compared to -161.5°C for methane.

The intermolecular force that holds water molecules together is known as the **hydrogen bond**. This bond is technically defined as the intermolecular force in which the hydrogen atom that is bonded to a nitrogen atom, oxygen atom, or fluorine atom in one molecule is attracted to an unshared pair of electrons on an electronegative atom in another molecule. See Figure 6.24.

Such a bond is indeed an *intermolecular* force, and *not* a covalent or ionic bond. It is also much weaker than a covalent or ionic bond, although strong enough to dictate a host of unique properties. It occurs in water but, obviously from the our definition, is certainly not limited to water.

Besides the effect on the boiling point of water, the hydrogen bond in water creates other unique properties. One is the structure and properties of ice. In liquid water, the intermolecular attraction is a reality, but, in spite of it, the water remains fluid, meaning that the molecules are still free to move and flow around one another. However, when water freezes, the molecules become immobilized and the hydrogen bonds become rigid. This creates a defined structure for ice that includes void spaces within. The presence of the void spaces means that the ice must occupy a larger volume than it did when it was a liquid. This means that its density must be lower as it converts from liquid water to ice. These facts dictate properties with which we are all familiar, including the fact that ice floats on water (such as in a glass of iced tea) and the fact that freezing water can burst plumbing pipes and water hoses.

The hydrogen bond plays an important role in other systems, too. To note one important example, hydrogen bonding is the force that holds the large DNA molecular strands together within the cells of all living things.

6.9.4 Hydrates

There is one additional property of the polar water molecule that is important to consider. This is its tendency to become trapped within the marble-stack crystal structure of ionic compounds when these compounds crystallize from a water solution. Ionic compounds that have water molecules trapped in their crystal structure are called hydrates and are described as being hydrated. Hydrates are very common compounds. If we think of the cations and anions as having a spherical shape (recall the stack-of-marbles discussion), it is easy to see that there are, in fact, void spaces in which small particles, such as water molecules, can fit. Due to the fact that water molecules are polar, they tend to cling to the cations and anions upon crystallization and then subsequently become trapped within the crystal structure. It is interesting that only a certain number of these molecules can fit there and that this number depends on the sizes of the cations and anions, which in turn depends on the identities of these ions and the way in which they fit together. For example, for every formula unit of CaCl_2 , six water molecules become trapped; for every formula unit of BaCl_2 , two water molecules become trapped; and for every formula unit of MgSO_4 , seven water molecules become trapped. This trapped water is not considered an impurity but part of the structure of the crystal, and thus it must be shown in the chemical formula of these compounds. The manner in which it is shown in the formula may be considered a bit unusual. Following the formula of the compound, a dot is written, followed by a number indicating the number of water molecules per formula unit and the water formula. The names of hydrates use the name of the compound that is represented by the ionic formula, and this compound is followed by a word that adds a prefix specifying the number of water molecules to the word *hydrate*. For example, we have the following formulas and names for hydrates:

$\text{CaCl}_2 \bullet 6\text{H}_2\text{O}$ Calcium chloride hexahydrate

$\text{BaCl}_2 \bullet 2\text{H}_2\text{O}$ Barium chloride dihydrate

$\text{MgSO}_4 \bullet 7\text{H}_2\text{O}$ Magnesium sulfate heptahydrate

The prefixes used in the names of these hydrates are *mono-* (one), *di-* (two), *tri-* (three), *tetra-* (four), *penta-* (five), *hexa-* (six), *hepta-* (seven) etc.

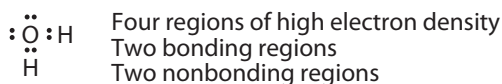
6.10 Valence-Shell Electron-Pair Repulsion Theory

Molecules of different covalent substances possess a wide variety of geometric shapes and structures that dictate whether they are symmetrical or non-symmetrical and whether they are polar or nonpolar, and thus what properties they will have.

The theoretical basis for a molecule possessing a particular geometric shape is the concept that electron pairs, whether they are part of a covalent bond (as in a bonding pair) or not (as in a nonbonding pair), will repel each other. In methane, water, and all other molecules, this repulsion means that the electron pairs will get as far away from each other as they can get. The general name for this theory is the valence-shell electron-pair repulsion (VSEPR) theory.

The determination of a molecule's structure or geometry is accomplished by first looking at the Lewis electron-dot structure and counting the number of regions of high electron density around the central atom. A region of high electron density can mean a pair of nonbonding electrons, a single bond, a double bond, or a triple bond. For both methane and water, for example, there are four such regions. See Figure 6.25. The number of such regions dictates one of five possible general shapes. These are shown in a vertical column along the left margin in Figure 6.26. These shapes are known as linear,

Water:



Methane:

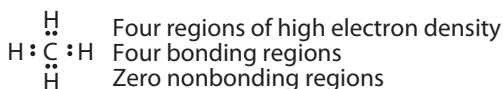


FIGURE 6.25 Two examples of molecules in which we count (1) the total number of regions of high electron density, (2) the number of these regions that are bonding, and (3) the number that are nonbonding as a prelude for using Figure 6.26 to determine the structure.

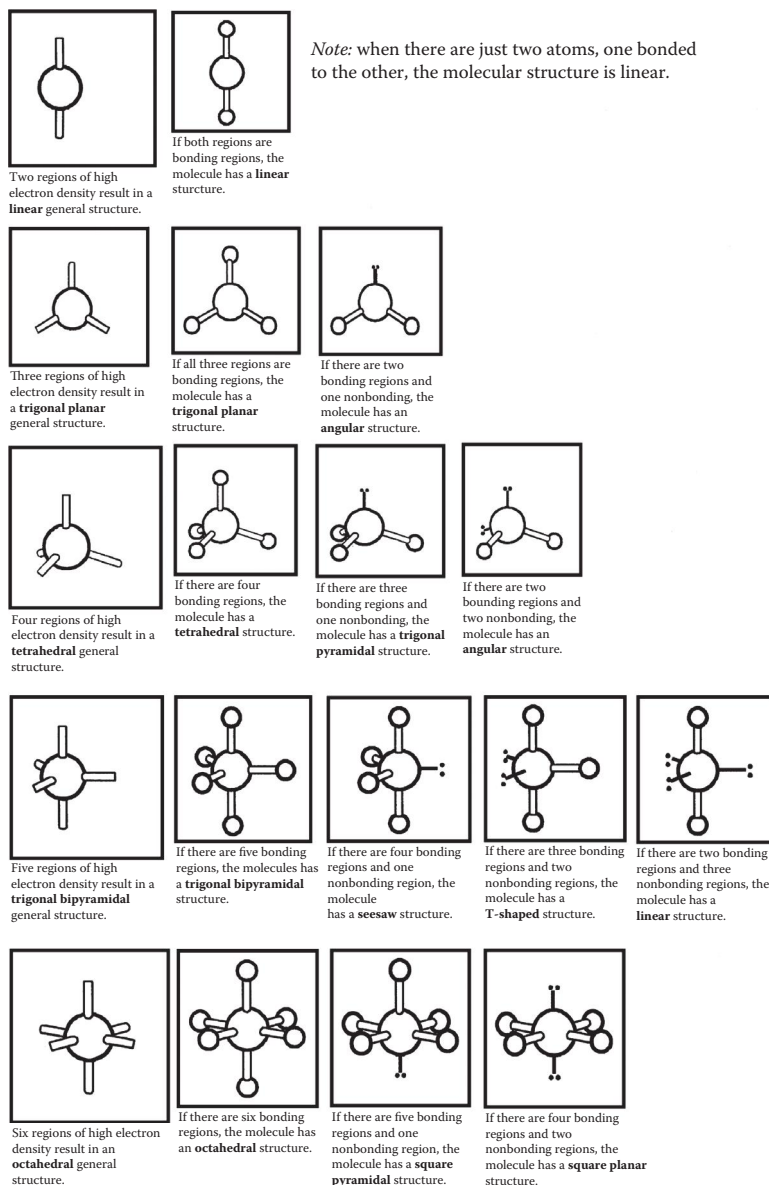


FIGURE 6.26 The various structures that result from the VSEPR theory. General structures are on the left and the possibilities for the actual molecular structures are on the right in each row. (From Kenkel, J., Kelter, P., and Hage, D., *Chemistry: An Industry-Based Introduction with CD-ROM*, CRC Press, Boca Raton, FL, 2001. With Permission.)

trigonal planar, tetrahedral, trigonal bipyramidal, and octahedral. Notice that with both water and methane the general structure is tetrahedral.

The actual molecular structure is determined by looking at the possibilities to the right of the general structure in Figure 6.26. It is apparent from this that methane is tetrahedral and that water is angular. Refer to Figure 6.25 to see the logic by which this was determined. All molecular structures can be determined by this scheme.

Example 6.8

What is the molecular structure of carbon tetrachloride (CCl_4)? (Refer back to Example 6.1 and Figure 6.13b for the Lewis electron-dot structure.)

Solution 6.8

Referring back to the Lewis structure in Example 6.1, we see that there are four regions of high electron density around the central atom, Cl. Looking at Figure 6.26, we see that this would dictate a tetrahedral general structure. Since all four regions are bonding regions, we see also from Figure 6.26 that the molecular structure is also tetrahedral.

Example 6.9

What is the molecular structure of chloric acid (HClO_3)? (Refer back to Example 6.2 and Figure 6.14b for the Lewis electron-dot structure.)

Solution 6.9

Referring back to the Lewis structure in Example 6.2, we see that there are four regions of high electron density around the central atom, Cl. From Figure 6.26, we see that chloric acid would have a tetrahedral general structure. However, three of these regions are bonding regions and one region is a nonbonding region. Looking again at Figure 6.26, we see that the molecular structure would be trigonal pyramidal.

Example 6.10

What is the molecular structure of nitric acid (HNO_3)? (Refer back to Example 6.3 and Figure 6.15c for the Lewis electron-dot structure.)

Solution 6.10

Referring back to the Lewis structure of nitric acid in Example 6.3, we see that there are three regions of high electron density around the central atom, N. From Figure 6.26, we see that nitric acid would have a trigonal planar general structure. All three regions are bonding regions. Looking again at Figure 6.26, we see that the molecular structure of nitric acid is also trigonal planar.

Example 6.11

What is the molecular structure of phosphorus pentachloride (PCl_5). (See Example 6.5 and Figure 6.17b for the Lewis electron-dot structure.)

Solution 6.11

Referring back to the Lewis structure of PCl_5 in Example 6.5, we see that there are five regions of high electron density around the central atom, P. From Figure 6.26, we see that PCl_5 would have a trigonal bipyramidal structure. All five regions are bonding regions. Looking again at Figure 6.26, we see that the molecular structure of PCl_5 is also trigonal bipyramidal.

6.11 Molecular Orbitals

What is the nature of the molecular orbitals, or the regions of space in which shared electrons reside in covalent bonding? One word that has been used to describe these regions is **overlap**, meaning that an atomic orbital on one atom overlaps, or occupies the same space, as an atomic orbital on another atom, which results in a merging of two spatial regions. There are two kinds of overlap, the end-to-end overlap and the side-to-side overlap. End-to-end overlap can occur between any two kinds of atomic orbitals, and these orbitals can each be the same kind of orbital or they can be different. Figure 6.27 gives representations of end-to-end overlap between two s orbitals, one s and one p orbital, and two p orbitals. The pair of electrons shared when two hydrogen atoms come together to form hydrogen gas (H_2) occupy two s orbitals before the overlap (the unpaired outermost electrons occupy s orbitals in hydrogen atoms). The top example in Figure 6.27 represents hydrogen. The pair of electrons that are shared when two chlorine atoms, for example, come together to

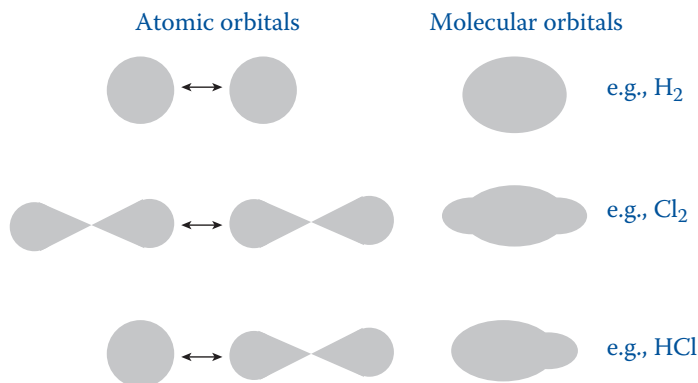


FIGURE 6.27 End-to-end overlapping of various orbitals. Top: two s orbitals as in hydrogen gas, (H_2) (see also Figure 6.6). Center: Two p orbitals as in chlorine gas (Cl_2) (see also Figure 6.7). Bottom: One s orbital and one p orbital as in hydrogen chloride (HCl) (see also Figure 6.8). Each is an illustration of a single bond (i.e., a σ -bond).

form chlorine gas (Cl_2) occupy two p orbitals before the overlap (the unpaired outermost electrons occupy p orbitals in chlorine atoms). The center example in Figure 6.27 represents chlorine. When one hydrogen atom and one chlorine atom, for example, come together to form hydrogen chloride (HCl), the unpaired electrons occupy one s orbital (the hydrogen) and one p orbital (the chlorine) before the overlap. The bottom example in Figure 6.27 represents hydrogen chloride (HCl). Orbitals that are hybrids of s , p , d , or f (recall the discussion accompanying Figures 6.11 and 6.12, for example) may also overlap end to end. The end-to-end overlap described here is called sigma bonding. A **sigma bond** (or, using the Greek letter σ for sigma, a σ bond) is the end-to-end overlap between *any two orbitals*.

Side-to-side overlap can only occur between p orbitals. Figure 6.28 presents a drawing of two p orbitals coming together and overlapping side to side. The side-to-side overlap between p orbitals is called pi bonding. A **pi bond** (or using the Greek letter π for pi, a π bond), is then defined as the side-to-side overlap between p orbitals. In a double bond, one of the bonds is a σ bond and one is a π bond. In a triple bond, one of the bonds is a σ bond and two are π bonds. In Figure 6.29, two p orbitals are coming together to form a σ bond and two other p orbitals are coming together to form the π bond.

Table 6.3 summarizes the facts concerning σ and π bonds.

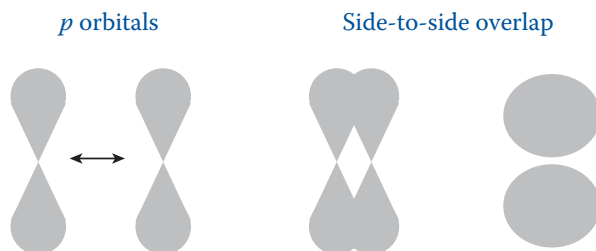


FIGURE 6.28 The side-to-side overlap of p orbitals that constitutes a π bond, the second bond in a double bond and the second and third bonds in a triple bond. See also Figures 6.11 and 6.12.



FIGURE 6.29 A representation of a double bond (i.e., both a σ bond and a π bond together, such as in ethylene [Figure 6.11]).

Table 6.3 Summary of Facts Concerning σ and π Bonds

σ -Bond

1. A σ bond consists of the end-to-end overlap between any two kinds of orbitals.
2. Whenever there is just one covalent bond between atoms, it is a σ bond.
3. A σ bond is always one of the bonds in a double bond and one of the bonds in a triple bond.

π Bond

1. A π bond consists of the side-to-side overlap between two p -orbitals. No other kind of orbital can participate in a π bond.
2. A π bond is always accompanied by a σ bond. Thus π bonds only occur in double or triple bonds.
3. In a double bond, there is one π bond. In a triple bond, there are two π bonds.

6.12 Homework Exercises

Ionic and Covalent Bonding

1. Define (a) ionic bond and (b) covalent bond.
2. Is KBr an ionic compound or molecular compound? Is HBr an ionic compound or a covalent compound? Explain your answers.
3. Why do we focus on the outermost electrons in our explanation of why and how ionic and covalent bonds form?

Lewis Electron Dot Structures

4. Draw the Lewis electron-dot structure for each of the elements that have atomic number between 1 and 20 and are missing from Figure 6.2.
5. Draw the Lewis electron-dot structure of each of the following.
 - a. K_2O
 - b. H_2CO_3
 - c. NO_3^-
 - d. AlCl_3
 - e. Na_2SO_3
 - f. XeF_4
6. Draw the Lewis electron-dot structure of each of the following.
 - a. IO_2^-
 - b. Br_2
 - c. BaCl_2
 - d. H_2S
 - e. IF_5
 - f. CN^-
 - g. H_2PO_4^-
7. Draw the Lewis electron-dot structure of each of the following.
 - a. BrF_3
 - b. CO
 - c. PH_3
 - d. SO_4^{2-}
 - e. KBrO_2
 - f. MgO
 - g. SO_3

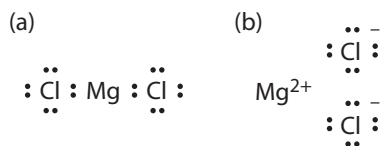


FIGURE 6.30 The Lewis structures.

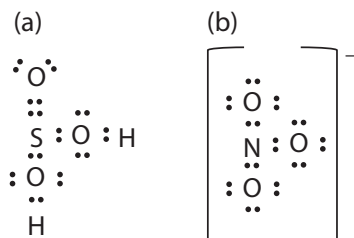


FIGURE 6.31 The Lewis structures.

8. Which of the Lewis electron-dot structures found in Figure 6.30 is the correct one for the compound magnesium chloride (MgCl_2)? Explain your answer.
9. Explain why the Lewis electron-dot structures found in Figure 6.31 are incorrect.
10. One of the N-H bonds in the ammonium ion, NH_4^+ , is a coordinate covalent bond. Explain why.
11. Explain why a molecule for an ionic compound cannot exist.

Polar Bonds and Molecules

12. Differentiate between electron affinity and electronegativity.
13. What does electronegativity have to do with bond polarity?
14. Is a polar bond an ionic bond or a covalent bond? Give an example of such a bond.
15. What is a polar molecule? Give an example.
16. Both compound A and compound B consist of molecules that have polar bonds. However, these molecules are polar in the case of compound A and nonpolar in the case of compound B. Explain this.

17. Tell whether each of the following bond is an ionic bond, a nonpolar covalent bond, or a polar covalent bond, and explain your answer. You may refer back to Table 6.2 for assistance.
- H–O
 - Br–Br
 - Ca–O
 - H–P
18. Carbon tetrachloride (CCl_4) does not mix with water, which indicates that it is nonpolar, and yet it has polar bonds in its structure. Explain this.
19. Consider a molecule that has polar covalent bonds in it. What is necessary for the molecule itself to be nonpolar?

Hydrogen Bonding

20. What is meant by the term *intermolecular force*? Give one specific example of such a force.
21. What is a hydrogen bond?
22. Identify at least two properties of water that are a direct result of the presence of the hydrogen bond.
23. Which compound would have a higher boiling point, CH_3Cl or CCl_4 ? What factors enter into your answer?

Hydrates

24. What is a hydrate? Write the formulas of at least two examples.
25. Explain the use of the dot in formulas of hydrates.

VSEPR Theory

26. What is the VSEPR theory?
27. How is the VSEPR theory used to help us understand the behavior and structure of molecules and ions?
28. Examine the ball-and-stick models for water and methane as drawn in Figure 6.23. In terms of the VSEPR theory, how are these models the same and how are they different?
29. Tell what the molecular structure is for each of the compounds in Figure 6.32. Possible answers include linear, trigonal planar, angular,

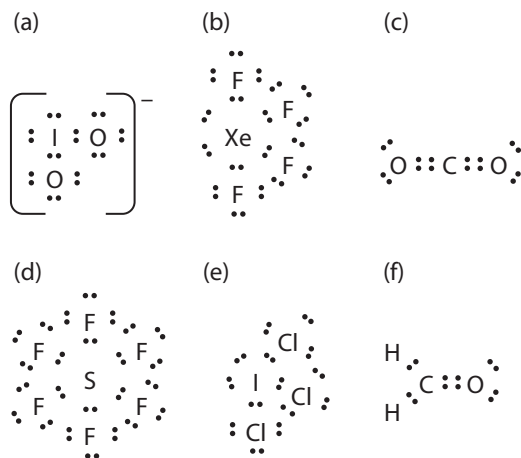


FIGURE 6.32 The Lewis structures.

tetrahedral, trigonal pyramidal, trigonal bipyramidal, seesaw, T-shaped, octahedral, square pyramidal, and square planar.

30. Ask your instructor to select certain Lewis electron-dot structures from exercises 5 through 7. Determine the molecular structure of each.

Sigma and Pi Bonding

31. What is meant by the term *molecular orbital*?
32. What is the difference between a sigma bond (σ bond) and a pi bond (π bond)? Give an example of an element, compound or ion that has a π bond. Give an example of an element, compound, or ion that has two π bonds.
33. Draw the shapes of the molecular orbitals that will be produced by each of the following combinations of orbitals.
- Two s orbitals combined end to end
 - Two p orbitals combined end to end
 - One s and one p orbital combined end to end
 - Two p orbitals combined side to side
34. Which of the situations (a–d) in Exercise 33 produce σ bonds? Which of the situations in Exercise 15 produce π bonds?

35. Consider the following: (a) Cl-Cl (b) $\text{H}_2\text{C}=\text{CH}_2$ (c) $\text{HC}\equiv\text{CH}$

Fill in the blanks with the word *all*, the word *none*, or one of the letters *a*, *b*, or *c* to correctly answer the question regarding the bonds between the single bond in (a) and the double, and triple bonds in (b) and

- a. Which of these has only one σ bond? _____
 - b. Which of these has only one π bond? _____
 - c. Which of these has two π bonds? _____
 - d. Which of these has two σ bonds? _____
 - e. Which of these has no σ bond? _____
 - f. Which of these has no π bonds? _____
 - g. Which of these has three σ bonds? _____
36. Define π bond with the help of drawings of the orbitals before and after a bond forms.
37. How many σ bonds and how many π bonds are there in a triple bond?
 σ bonds _____ π bonds _____
38. a. What kind of atomic orbital (or orbitals) are involved in a σ bond?
Is the overlap in a σ bond end-to-end or side-to-side?
- b. What kind of atomic orbital (or orbitals) are involved in a π -bond?
Is the overlap in a π bond end-to-end or side-to-side?
39. Does a double bond consist of two σ bonds, two π bonds, or one σ bond and one π bond? Explain your answer.

Chapter 7

Atomic Weight, Formula Weight, and Moles

Introduction	174
Atomic Weight Revisited	174
Molecular Weight and Formula Weight	175
Mole	177
Percent Composition of Compounds	179
Derivation of Empirical Formulas	182
Homework Exercises	186

7.1 Introduction

We defined chemistry in Chapter 1 as the study of the composition, structure, and properties of matter and the changes that matter undergoes. As our study of chemistry has unfolded in subsequent chapters, we have learned that a great deal of symbolism (e.g., element symbols and formulas of compounds) is used to describe chemical composition, structure, properties, and changes. We have also learned that some aspects of chemistry, particularly aspects of structure, can be very abstract, requiring us to think in terms of the extremely small invisible particles called atoms, molecules, and ions, and the even-smaller subatomic particles called protons, neutrons, and electrons. The special language of chemists and chemistry technicians, however, is certainly not limited to the abstract or to very small quantities. In this chapter, we move from the realm of the abstract and unseen to the realm of observable and measurable realities. It is time to apply our symbolism and our knowledge to quantities of matter that are observable and measurable in the laboratory and to the very large quantities of matter that are generated and consumed every day in the chemical industry. The expression, measurement, and calculation of quantities of matter in compounds and in reactions is called **stoichiometry**.

7.2 Atomic Weight Revisited

Let us begin with a brief review of the facts concerning atomic weight as discussed in Chapter 1. These facts are as follows:

1. A number called the atomic weight is listed for each element in the periodic table. In most tables, it is the number found near the bottom of each square in the periodic table below the symbol for the element (see Figure 1.2).
2. The atomic weight of any element is the average weight of all the weights of all the atoms that exist naturally for that element. The weight of any one atom is the sum of the masses, or weights, of all the particles found in the atom, which are the protons, neutrons, and electrons.
3. The unit associated with these masses is the atomic mass unit (amu). One amu is equal to approximately 1.66×10^{-24} grams (g).

4. The mass of a single proton and the mass of a single neutron are each approximately 1 amu, and the mass of an electron is approximately 0.00055 amu.
5. Weights of the atoms of a given element vary due to the existence of isotopes, which are atoms of a given element that differ due to differing numbers of neutrons in the nucleus.

A beginning chemistry student may conclude that an atomic weight should be approximately a whole number equal to the mass number, the sum of the protons and neutrons in an atom, since each of these weighs approximately 1 amu and since the weight of an electron is approximately 0 amu. Points 2 and 5 dispel this theory, however. A casual look at the periodic table reveals that atomic weights are anything but whole numbers.

We have used the word *approximately* in our references to the weights of protons and neutrons with respect to the amu. The question may arise as to what standard *is* used to define the amu. The answer is that the amu is defined as one-twelfth the mass of the isotope of carbon that has six protons and six neutrons, the so-called carbon-12 isotope. From this it is easy to see why the masses of the proton and the neutron are *approximately* equal to 1 amu since it is carbon-12's six electrons that create the discrepancy.

7.3 Molecular Weight and Formula Weight

While we have dealt with the atomic weights of *elements* here and in Chapter 1, we have not dealt with the weights of molecules and formula units, which are the fundamental particles of *compounds*. First, by way of review, it is important to distinguish between molecular (or covalent) compounds and ionic compounds. Molecular compounds are those that consist of distinct individual particles called molecules. Molecules interact with one another in a sample of a molecular compound, but only in a rather mild way. Ionic compounds are those that consist of indistinct particles, which are most appropriately referred to as formula units, because a formula represents the simplest ratio between the cations and anions that constitute the compound (and not a distinct particle). Ionic compounds, on the atomic scale, consist of a three-dimensional array of cations and anions, all of which strongly interact with each other. Thus there is no such particle as a molecule in these cases. Figure 7.1 depicts the difference between these two kinds of compounds.

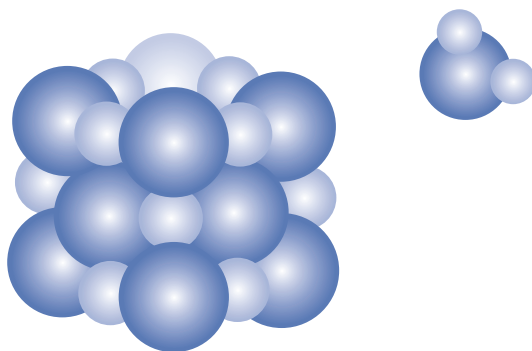


FIGURE 7.1 Left, an example of a three-dimensional array of ions (no molecule). Right, an example of a molecule, the fundamental unit of a molecular compound. (From Kenkel, J., Kelter, P., and Hage, D., *Chemistry: An Industry-Based Introduction with CD-ROM*, CRC Press, Boca Raton, FL, 2001. With Permission.)

Molecular weight, or the weight of a single molecule in amu, is thus strictly applicable to molecular compounds. A more general term would be **formula weight**, or the weight of one formula unit. This latter term is descriptive of both molecular compounds and ionic compounds. In this text, the term *formula weight* will be used to describe the weight of the fundamental unit of all compounds.

The formula weight is determined by simply adding together the atomic weights of all the elements in the formula. Some examples are given here.

Example 7.1

What is the formula weight of ethyl alcohol ($\text{C}_2\text{H}_6\text{O}$)?

Solution 7.1

$$\begin{array}{rcl}
 2(\text{C}) & = & 2(12.011) = 24.022 \text{ amu} \\
 6(\text{H}) & = & 6(1.008) = 6.048 \text{ amu} \\
 + \quad 1(\text{O}) & = & 1(15.999) = 15.999 \text{ amu} \\
 \hline
 \text{Formula weight} & = & 46.069 \text{ amu}
 \end{array}$$

Example 7.2

What is the formula weight of calcium nitrate [$\text{Ca}(\text{NO}_3)_2$]?

Solution 7.2

$$\begin{array}{rcl}
 1(\text{Ca}) & = & 1(40.078) = 40.078 \text{ amu} \\
 2(\text{N}) & = & 2(14.007) = 28.014 \text{ amu} \\
 + \quad 6(\text{O}) & = & 6(15.999) = 95.994 \text{ amu} \\
 \hline
 \text{Formula weight} & = & 164.086 \text{ amu}
 \end{array}$$

7.4 Mole

How do we solve the problem of these fundamental units of matter being too small to measure in the laboratory, as suggested in Section 7.1? Scientists have defined a quantity of matter called the **mole** as a substitute entity for the atom, molecule, or formula unit. The mole *is* measurable and observable in the laboratory. It is simply the atomic or formula weight expressed in grams rather than amu. In other words, for elements the atomic weight as listed in the periodic table is not only the average number of amu per atom, it is also the average number of grams per mole of the element. For compounds, the formula weight as calculated from the atomic weights in the table is not only the average number of amu per molecule or formula unit, but it is also the average number of grams per mole of the compound. Examples of 1 mole of various elements and compounds are presented in Figure 7.2.

7.4.1 Atomic and Formula Weight: Number of Grams per Mole of an Element or a Compound

One result of this apparently simple concept is that the atomic weight can be used to interconvert grams and moles. Since the units of atomic weight, as defined earlier, include both grams and moles, it can be used as a conversion factor to convert grams to moles, and vice versa. The dimensional analysis would appear as follows for the conversion from grams to moles:

$$\frac{\cancel{\text{Moles}}}{\cancel{\text{Mole}}} \times \frac{\text{Grams}}{\cancel{\text{Mole}}} = \text{Grams}$$

and as follows for the conversion from moles to grams:

$$\frac{\cancel{\text{Grams}}}{\cancel{\text{Gram}}} \times \frac{\text{Mole}}{\cancel{\text{Gram}}} = \text{Moles}$$

Some examples follow.



FIGURE 7.2 One mole of several elements and compounds. Clockwise, from left, 32.066 g of sulfur, 249.686 g of copper sulfate pentahydrate, 180.16 g of sugar, 58.443 g of sodium chloride, 26.982 g of aluminum, and 18.015 g of water.

Example 7.3

How many moles does 17.3 grams of silver represent?

Solution 7.3

From the periodic table, the atomic weight of silver is 107.87 g/mol. So the problem is solved as follows (mol being the official abbreviation for moles).

$$\frac{17.3 \text{ g}}{107.87 \text{ g}} \times \frac{1 \text{ mol}}{1} = 0.160 \text{ mol}$$

Example 7.4

What is the mass in grams of 569.2 moles of water?

Solution 7.4

The formula weight of water, H_2O , is needed.

$$\begin{array}{rcl} 2(\text{H}) & = & 2(1.008) = 2.016 \text{ g} \\ + & & 1(\text{O}) = 1(15.999) = 15.999 \text{ g} \\ \hline \text{Formula weight} & & = 18.015 \text{ g/mol} \end{array}$$

Thus we have the following:

$$\frac{569.2 \cancel{\text{mol}}}{1 \cancel{\text{mol}}} \times \frac{18.015 \text{ g}}{1 \cancel{\text{mol}}} = 1.025 \times 10^4 \text{ g}$$

7.4.2 Avogadro's Number: Number of Particles per Mole

The mole is a concept of convenience similar to what the dozen is for eggs and donuts. It is much more convenient for a baker to think in terms of a dozen donuts and for the egg producer to think in terms of a dozen eggs rather than just one donut or one egg. The difference here is that the number of atoms, molecules, or formula units that make up one mole is an enormous number, not something easy to visualize like 12 is. The number of atoms or molecules required for 107.87 g of silver or 18.015 g of water (the atomic or formula weights as in the earlier examples), or for one mole of anything, is 6.022×10^{23} . To illustrate the size of this number, one author* has described the mole in terms of garden peas. One mole of garden peas would be a pile so large that it would cover the entire United States to a depth of 6 km. This number has come to be known as Avogadro's number, named after Amadeo Avogadro (1776–1856), an Italian physicist.

7.5 Percent Composition of Compounds

A certain percentage of the mass of a compound is attributable to the masses of the elements of which it is composed. For example, a certain percentage of the mass of a sample of ammonia (NH_3) is attributable to the mass of the nitrogen that it contains. A certain percentage of a sample of isopropyl alcohol ($\text{C}_3\text{H}_8\text{O}$) is attributable to the mass of the oxygen that it contains. We can also say that a certain percentage of the mass of a compound is attributable to the mass of a grouping of elements that are found in the formula, such as a polyatomic ion, or water of hydration. Thus a certain percentage of the mass of a sample of ammonium chloride (NH_4Cl) is due to the ammonium ion, and

* Ralph H. Petrucci, *General Chemistry Principles and Modern Applications*, 4th Edition, New York, NY: Macmillan Publishing, 1985.

a certain percentage of the mass of a sample of copper sulfate penthydrate ($\text{CuSO}_4 \cdot 5\text{H}_2\text{O}$) is due to water.

The calculation of the percent of an element in a compound is derived from the general definition of percent.

$$\text{Percent} = \frac{\text{Part}}{\text{Whole}} \times 100 \quad (7.1)$$

If the part is a weight and the whole is a weight, then we have the following:

$$\text{Percent} = \frac{\text{Weight of part}}{\text{Weight of whole}} \quad (7.2)$$

Applying this to the percent of an element in a compound, we have the following:

$$\text{Percent} = \frac{\text{Weight of element}}{\text{Weight of compound}} \times 100 \quad (7.3)$$

7.5.1 Percent Composition from Formulas and Atomic and Formula Weights

The percent of elements or groupings of elements in compounds can be determined directly from the formulas of the compounds and the atomic and formula weights. This is illustrated in the examples provided next.

Example 7.5

What is the percent of oxygen in ethyl acetate ($\text{C}_4\text{H}_8\text{O}_2$)?

Solution 7.5

First the formula weight is needed.

$$\begin{aligned} 4(\text{C}) &= 4(12.011) = 48.044\text{g} \\ 8(\text{H}) &= 8(1.008) = 8.064\text{g} \\ + 2(\text{O}) &= 2(15.999) = 31.998\text{g} \\ \hline \text{Formula weight} &= 88.106\text{g/mol} \end{aligned}$$

We then use the equation for percent (Equation 7.3) and plug in the weights from the previous calculation:

$$\begin{aligned}\text{Percent of oxygen} &= \frac{\text{Weight of oxygen}}{\text{Weight of compound}} \times 100 \\ &= \frac{31.998}{88.105} \times 100 = 36.318 \%\end{aligned}$$

Using the method from Example 7.5, the percent of any and all elements in a compound may be determined and collectively reported. By definition, this collective report is defined as the **percent composition** of a compound.

Example 7.6

What is the percent composition of sodium carbonate (Na_2CO_3)?

Solution 7.6

Percents of all three elements in the formula—Na, C, and O—are determined and reported just as the individual percent was in Example 7.5.

$$\begin{aligned}2(\text{Na}) &= 2(22.990) = 45.980\text{g} \\ 1(\text{C}) &= 1(12.011) = 12.011\text{g} \\ + 3(\text{O}) &= 3(15.999) = 47.997\text{g} \\ \hline \text{Formula weight} &= 105.988\text{g/mol} \\ \hline \%\text{Na} &= \frac{45.980}{105.988} \times 100 = 43.382 \% \\ \%\text{C} &= \frac{12.011}{105.988} \times 100 = 11.332 \% \\ \%\text{O} &= \frac{47.997}{105.988} \times 100 = 45.285 \%\end{aligned}$$

The individual percents determined by the method of Example 7.6 should add to 100%. Occasionally, due to effects of the rounding of the atomic weights, they may not add to exactly 100%.

In a way similar to Example 7.5, the percent of a grouping of elements may be determined by this method.

Example 7.7

What is the percent of water in $\text{CuSO}_4 \cdot 5\text{H}_2\text{O}$?

Solution 7.7

First the weight of the five water units:

$$\begin{array}{r} 10(\text{H}) = 10(1.008) = 10.080 \text{ g} \\ + 5(\text{O}) = 5(15.999) = 79.995 \text{ g} \\ \hline \text{Total weight of water units} = 90.075 \text{ g/mol} \end{array}$$

Then the formula weight of the compound is calculated.

$$\begin{array}{r} 1(\text{Cu}) = 1(63.546) = 63.546 \text{ g} \\ 1(\text{S}) = 1(32.066) = 32.066 \text{ g} \\ 4(\text{O}) = 4(15.999) = 63.996 \text{ g} \\ + 5(\text{H}_2\text{O}) = 90.075 \text{ g (from above)} \\ \hline \text{Formula weight} = 249.683 \text{ g/mol} \end{array}$$

Finally, the percent of water is calculated.

$$\% \text{H}_2\text{O} = \frac{90.055}{249.683} \times 100 = 36.076 \%$$

7.6 Derivation of Empirical Formulas

While it is clear that percent composition can be derived from formulas, the reverse is also true. Empirical formulas, or formulas in which the subscripts are in their lowest terms (e.g., acetylene $[\text{C}_2\text{H}_2]$ has an empirical formula of CH), can be derived from percent composition data. Knowing what elements are in the compound and having some data about the relative amounts of these elements in the compound, we can determine the subscripts and the empirical formula.

As we learned in Chapter 1, the subscripts represent the number of atoms of each element in the formula. In the formula H_2O , for example, we know that

the 2 subscript means that there are two atoms of hydrogen in a molecule of water. We also know that the lack of a subscript with the oxygen means that there is one atom of oxygen in one molecule of water. If we consider moles instead of atoms or molecules, we can say that the 2 subscript means that there are two *moles* of hydrogen in 1 *mole* of water, and the lack of the subscript means that there is one *mole* of oxygen in 1 *mole* of water. If we are able to translate the available information about a compound into *moles* of the elements involved, we can come up with the empirical formula.

7.6.1 Empirical Formulas from Percent Composition

Percent composition represents the number of grams of each element in every 100 g of the compound. Grams can be converted to moles by dividing by the atomic weight. Subscripts can then be determined by comparing the number of moles of one element relative to another. If the number of moles of one element is twice the number of moles of another, then the subscript of the first element will be twice the other. The formal step-by-step procedure for converting percent composition data into subscripts is given in Table 7.1.

Table 7.1 Steps to Determine the Formula of a Compound When Given Percent Composition Data

1. Assume that we have 100 g of the compound. The percent composition data then translates to grams of each element.
2. Convert the number of grams of each element to moles by dividing each by the atomic weight of the element.
3. Divide each number of moles by the one that is the smallest number. If the result for each is a whole number, then these whole numbers are the subscripts in the formula.
4. If the result for each in step 3 is not a whole number, multiply each by 2. If the result for each is now a whole number, then these whole numbers are the subscripts in the formula.
5. If the result for each in step 4 is not a whole number, multiply each result from step 3 by 3, then by 4, then by 5, and so on until whole numbers are obtained for the subscripts.

Note: This is one time that the rules for significant figures are unimportant since the subscripts always consist of one digit regardless of the number of significant figures in the data. It is often acceptable, then, to judge a “near whole number” as a whole number to avoid ridiculously large numbers as subscripts.

Example 7.8

What is the empirical formula of the compound that is 56.6% potassium, 8.7% carbon, and 34.7% oxygen?

Solution 7.8

The steps that follow are the steps from Table 7.1.

Step 1: The percent figures translate directly to grams. So we have 56.6 g of potassium, 8.7 g of carbon, and 34.7 g of oxygen.

Step 2: These grams are now converted to moles.

$$\frac{56.6 \text{ gK}}{39.098 \text{ gK}} \times \frac{1 \text{ mol K}}{1} = 1.45 \text{ mol K}$$

$$\frac{8.7 \text{ gC}}{12.011 \text{ gC}} \times \frac{1 \text{ mol C}}{1} = 0.72 \text{ mol C}$$

$$\frac{34.7 \text{ gO}}{15.999 \text{ gO}} \times \frac{1 \text{ mol O}}{1} = 2.17 \text{ mol O}$$

Step 3: Divide each number of moles by the smallest number.

$$\text{K: } \frac{1.45}{0.72} = 2.01 \approx 2$$

$$\text{C: } \frac{0.72}{0.72} = 1.00 = 1$$

$$\text{O: } \frac{2.17}{0.72} = 3.01 \approx 3$$

Therefore the empirical formula is K_2CO_3 .

Example 7.9

What is the empirical formula of a compound that is 43.3% phosphorus and 56.7% oxygen?

Solution 7.9

The steps that follow are the steps from Table 7.1.

Step 1: The percentages translate to grams: 43.3 g of phosphorus and 56.7 g of oxygen.

Step 2: Converting these grams to moles, we have the following:

$$\frac{43.3 \text{ gP}}{30.974 \text{ gP}} \times \frac{1 \text{ mol P}}{1} = 1.40 \text{ mol P}$$

$$\frac{56.7 \text{ gO}}{15.999 \text{ gO}} \times \frac{1 \text{ mol O}}{1} = 3.54 \text{ mol O}$$

Step 3: Divide each number of moles by the smallest number.

$$\text{P: } \frac{1.40}{1.40} = 1.00 = 1$$

$$\text{O: } \frac{3.54}{1.40} = 2.53 \approx 2.5$$

Step 4 is needed since 2.5 is not close to a whole number.

Step 4: Multiplying each by 2 gives whole numbers.

$$\text{P: } 2 \times 1 = 2$$

$$\text{O: } 2 \times 2.5 = 5$$

Therefore the empirical formula is P_2O_5 .

7.6.2 Empirical Formulas vs. Molecular Formulas

The subscripts in a molecular formula are a whole number multiple of the subscripts in the empirical formula. For example, the empirical formula of benzene is CH while the molecular formula is C_6H_6 . The whole number multiple in this case is 6. The formula weight is the same whole number multiple of the empirical weight (the formula weight calculated from the empirical formula). In the case of benzene, the empirical weight is 13.0 g/mol and the molecular formula is 6×13.0 or 78.0 g/mol. If the empirical formula is given and the molecular formula is desired, the formula weight must be known. Then the whole number multiple can be calculated and the subscripts in the molecular formula determined.

Example 7.10

The empirical formula of borazole is BNH_2 . If the formula weight of borazole is 80.4 g/mol, what is the molecular formula?

Solution 7.10

First, calculate the empirical weight.

$$\begin{aligned} 1(\text{B}) &= 1(10.811) = 10.811\text{g} \\ 1(\text{N}) &= 1(14.007) = 14.007\text{g} \\ + 2(\text{H}) &= 2(1.008) = 2.016\text{g} \\ \hline \text{Formula weight} &= 26.834\text{g/mol} \end{aligned}$$

The whole number multiple is calculated as follows:

$$\frac{80.4}{26.834} = 3.00 \approx 3$$

The molecular formula is therefore $\text{B}_3\text{N}_3\text{H}_6$.

7.7 Homework Exercises**Formula Weight**

1. Define stoichiometry.
2. One amu is defined as one-twelfth the weight of the carbon-12 isotope, and there are six protons and six neutrons in this isotope. With being the case, why is the weight of one proton and the weight of one neutron each not equal to 1 amu?
3. Define the following terms, as pertaining to chemical substances.
 - a. Molecular weight
 - b. Formula weight
 - c. Mole
4. Why is it better to refer to the formula weight, rather than the molecular weight, of compounds?
5. Calculate the formula weight of each of the following.
 - a. KCl
 - b. $\text{Al}_2(\text{SO}_4)_3$
 - c. $\text{Mg}(\text{NO}_2)_2$

6. Complete the following table.

	Chemical Formula	Formula Weight
(a)	CO_2	
(b)	NH_3	
(c)	K_2SO_4	
(d)	Na_3PO_4	
(e)	$\text{C}_4\text{H}_{10}\text{O}$	
(f)	$\text{C}_2\text{H}_3\text{N}$	
(g)	$\text{Mg}(\text{OH})_2$	
(h)	$\text{Al}_2(\text{CO}_3)_3$	
(i)	$\text{FeSO}_4 \cdot 7\text{H}_2\text{O}$	
(j)	$\text{CoCl}_2 \cdot 6\text{H}_2\text{O}$	

Mole/Gram Conversions

7. Convert the moles of each to grams.

- 82 mol of sulfur
- 7.3 mol of silver
- 0.1883 mol of NaBr
- 34.1 mol of $\text{BaCl}_2 \cdot 2\text{H}_2\text{O}$

8. Convert the grams of each to moles.

- 0.339 g of phosphorus
- 0.04451 g of antimony
- 0.9338 g of KCl
- 7.22 g of $\text{Ca}(\text{NO}_3)_2$

9. Complete the following table.

	Chemical	Grams	Moles
(a)	Au	2.34 g	
(b)	Sb		0.4490 mol
(c)	Cr	0.034 g	
(d)	Ar		129.0 mol
(e)	CH_4	0.3148 g	
(f)	$\text{C}_3\text{H}_6\text{O}_2$		17.0 mol
(g)	$\text{Al}(\text{OH})_3$	5.63×10^4 g	
(h)	$\text{BaCl}_2 \cdot 2\text{H}_2\text{O}$		2.110×10^3 mol

10. Which has more molecules? (Hint: the one that is more moles has more molecules.)
- 0.339 g of CO_2 or 0.108 mol of CO_2
 - 0.116 g of NH_3 or 4.5 mol of NH_3

Percent Composition

- What is the percent of sulfur in K_2SO_4 ?
- What is the percent of nitrogen in (a) KNO_3 and (b) $\text{Ca}(\text{NO}_3)_2$?
- Calculate the percent of nitrogen in each of the compounds in the following table.

	Chemical Formula	Nitrogen (%)
(a)	NH_3	
(b)	NO	
(c)	N_2O	
(d)	NO_2	
(e)	NO_3	
(f)	N_2O_3	
(g)	N_2O_4	
(h)	N_2O_5	

- What is the percent composition of the following?
 - CCl_4
 - $(\text{NH}_4)_3\text{N}$
 - $\text{C}_3\text{H}_6\text{O}_2$
 - $\text{HC}_2\text{H}_3\text{O}_2$
- What is the percent of water in Epsom salt ($\text{MgSO}_4 \cdot 7\text{H}_2\text{O}$)?
- Two-thirds of the atoms in water are hydrogen atoms, yet only 11% of the weight of water is due to hydrogen. Explain.

Calculation of Empirical Formula and the True Formula

- Determine the empirical formula from the information given in each of the following.
 - 20.16 %Al, 79.84 %Cl
 - 65.3 %Sr, 10.6 %Si, 24.1 %O

- c. 58.8 %C, 9.9 %H, 31.3 %O
 - d. 72.4 %Fe, 27.6 %O
 - e. 74.96 %Al, 25.04 %C
 - f. 40.9 %C, 4.58 %H, 54.5 %O
 - g. 29.0 %Na, 40.5 %S, 30.4 %O
 - h. 21.85 %Mg, 27.85 %P, 50.31 %O
18. Write the true formula of a compound with an empirical formula as follows.
- a. $\text{C}_2\text{H}_6\text{O}$ if the true formula weight is 92.0 g/mol
 - b. $\text{C}_2\text{H}_4\text{O}$ if the true formula weight is 132.0 g/mol
 - c. $\text{C}_4\text{H}_4\text{S}$ if the true formula weight is 168.0 g/mol
 - d. $\text{C}_3\text{H}_2\text{Cl}$ if the true formula weight is 147.0 g/mol
19. The percent composition of a certain compound is 54.5% carbon, 9.1% hydrogen, and 36.4% oxygen. The formula weight of this compound is 88.0 g/mol. Calculate both the empirical formula and molecular formula of this compound.

Chapter 8

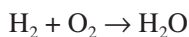
Chemical Equations and Stoichiometry

Introduction to Chemical Equations	192
Balancing Chemical Equations	193
Optional Symbols.	199
Calculations Based on Equations.	201
Limiting Reactant.	205
Theoretical Yield, Actual Yield, and Percent Yield	207
Homework Exercises	208

8.1 Introduction to Chemical Equations

Symbols and names of elements, as well as formulas and names of compounds, are fundamental to the description and expression of chemical principles and processes. To describe the formation of water from the chemical combination of hydrogen and oxygen would be difficult to do without having the name *water* assigned to the clear, colorless, and odorless liquid that is involved in this reaction, or without having the names *hydrogen* and *oxygen* assigned to the colorless, odorless, and invisible gases that are involved. Additionally, knowing that the formulas of hydrogen and oxygen are H_2 and O_2 , respectively, and knowing that the formula of water is H_2O reduces the time and space needed to describe the process. In a laboratory notebook, a chemist might state that H_2 and O_2 combine to form H_2O , rather than writing the actual names of these substances. It is very important to have symbols and formulas available for the substances involved in chemical processes, in addition to the names. These symbols and formulas provide a type of chemical shorthand that chemists have found invaluable because of their convenience. Even if you have found the task of learning symbols, formulas, and names to be laborious, ultimately this work reaps the important reward of convenience.

The use of symbols and formulas is especially helpful in the expression of chemical change. Another term for chemical change, or the transformation of one or more substances with known properties into one or more other substances with completely different properties, is chemical reaction. The chemical combination of hydrogen and oxygen to form water is an example of a chemical reaction. We say that hydrogen and oxygen react to form water, or that hydrogen reacts with oxygen to form water. A statement of a chemical reaction that uses the convenient chemical shorthand of symbols and formulas is known as a chemical equation. In a chemical equation, the symbols and formulas of the substances present before the reaction takes place are written to the left and separated by a plus (+) sign. An arrow pointing to the right signifies a chemical change. Symbols and formulas of the elements or compounds that form as a result of the reaction appear on the right. These are also separated by + signs. Thus, for the water formation reaction, we would write:



The substance or substances to the left of the arrow, the H_2 and the O_2 in the example, are called the reactants. The substance or substances to the right of

the arrow, water in this example, are called the products of the reaction. Thus, we have reactants and products separated by the arrow.

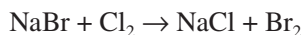
Reactants $\rightarrow$ Products

The information in a chemical equation is verbalized by using the words *plus* or *and* or *reacts with* for the + signs on the left; the words *yields* or *reacts to produce* or *to form* for the arrow; and the word *plus* or *and* for the + signs on the right. Thus the water formation reaction might be verbalized as follows: “Hydrogen reacts with oxygen to form water” or “Hydrogen and oxygen react to produce water” or “Hydrogen plus oxygen yields water.”

Examples in which there is only one reactant but two or more products, or two or more reactants with two or more products, are verbalized similarly. For example, the equation



may be verbalized as “Calcium carbonate yields calcium oxide plus carbon dioxide” and the equation

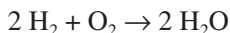


may be verbalized as “Sodium bromide reacts with chlorine to produce sodium chloride and bromine.”

8.2 Balancing Chemical Equations

8.2.1 Quantitative Information from Chemical Equations

Chemical equations also can provide quantitative information, which means that it is possible to tell how much of a reactant or a product is involved. To do that, we assume that each symbol and formula present in the equation represents exactly one atom, one molecule, or one formula unit of the element or compound. It is possible to then indicate two or more atoms, molecules, or formula units by placing coefficients in front of each of the symbols or formulas, such that the number of total atoms of each element is the same on both sides. Such an equation is said to be balanced. Using our water formation example, the following represents a balanced equation:



We could now verbalize this equation as follows: “Two molecules of hydrogen and one molecule of oxygen react to produce two molecules of water.”

Since two molecules of hydrogen gas (H_2) are represented on the left side, there are a total of four hydrogen atoms represented on the left (two molecules of H_2 with two hydrogen atoms each). Since there are two molecules of water represented on the right side, there are also four atoms of hydrogen on the right (two hydrogen atoms in each of two molecules of water). Also, since there is one molecule of oxygen on the left side (no coefficient means one), there are two atoms of oxygen on the left (two atoms per molecule); and since there are two molecules of water on the right side, there are also two atoms of oxygen represented on the right (one atom per molecule). Thus the equation is said to be balanced. The equation now indicates quantities of chemicals in terms of the number of molecules that are involved.

All chemical equations can be balanced. In Section 8.2.3, we will discuss the process of how simple chemical equations may be balanced. First, we need to understand why we must balance these equations.

8.2.2 Law of Conservation of Mass

One of the most basic of all chemical principles is the law of conservation of mass. For the purpose of this discussion, consider mass to be the same as weight. This law states that in a chemical reaction, mass can be neither created nor destroyed. This means that the total mass of the products formed in a chemical reaction must equal the total mass of the reactants that were used. It is a law that is easily proven in the laboratory. If an enclosed container is used so that no material can escape (or be added) during the reaction, and if the reaction is not initiated until the mass of the container with reactants is measured, you should find that when the reaction is initiated and allowed to proceed, the mass of the container does not change.

A balanced chemical equation is a demonstration of the law of conservation of mass. In a typical chemical reaction, atoms lose their identity because they enter into chemical combination, are released from chemical combination, or enter into a different mode of chemical combination. However, they do not lose their identity in terms of the number of protons in the nucleus.* Thus, for example, an atom of iron is an atom of iron in all instances, whether

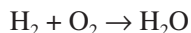
* The exception to this are nuclear reactions, or reactions in which the nuclei of atoms are either split or fused. Nuclei and atoms of other elements are thus formed.

it is chemically combined or not. Recognizing this, we can restate the law of conservation of mass to say that in a chemical reaction, the total number of atoms of each element present in the reaction container must be the same before the reaction begins and after the reaction is complete. The act of balancing a chemical equation, then, puts this important law into practice. The total number of atoms of each element on the left side of the equation must equal the total number of atoms of each element on the right side.

8.2.3 Process of Balancing Equations

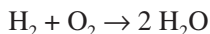
Many equations can be balanced by inspection. This is a trial-and-error method by which we try various coefficients on both sides, going back and forth, from the left side, to the right side, to the left side again, and so on until the equation is balanced. This process will be described here.

Equations are balanced by placing coefficients in front of the individual symbols and formulas. The equation representing the water-formation reaction

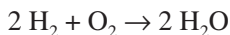


is balanced as shown below.

Starting with the first element encountered as you look at the reaction equation from left to right, we see that the hydrogen atoms are already balanced, with two on the left and two on the right. Proceeding through the equation from left to right, we next encounter oxygen. We see that we need more oxygen on the right and that a 2 in front of the H_2O would balance the oxygens.



While the oxygens are now balanced, the hydrogens have become unbalanced. However, we see that placing a 2 in front of the H_2 would balance the hydrogens and not affect the oxygens.



The equation is now balanced.

As a general rule, the coefficients used should be whole numbers and not fractions. For example, this equation is balanced if the fraction $\frac{1}{2}$ were as the coefficient for O_2 :



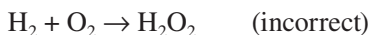
However, fractions in general should be avoided. Also, the coefficients used should be the smallest possible numbers. For instance, this same equation is balanced as follows:



but the coefficients can be smaller and still give a balanced equation.

There are two very important rules that we must introduce here. These are (1) equations cannot be balanced by changing the subscripts in the formulas, and (2) equations cannot be balanced by placing coefficients in the middle of formulas.

Changing the subscripts in the formulas is incorrect because this changes the identity of the chemical that is involved in the reaction. For example, the water formation reaction can be balanced by placing a 2 as a subscript to the O on the right.



Although this equation is balanced, the result is incorrect because hydrogen peroxide (H_2O_2) would be listed as the product instead of water, which makes this an equation representing a different reaction from the original water formation reaction. We have changed the equation; we have not balanced it.

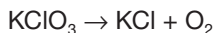
In the case of the second rule, we might be tempted to balance the equation by placing a 2 between the H and the O in the formula for water. Thus we would have the following:



This is incorrect because the formula on the right is not a legitimate formula. The examples of the trial-and-error inspection process that follow provide some hints that will be helpful in certain cases.

Example 8.1

Balance the following equation.



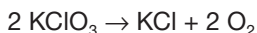
Solution 8.1

Proceeding from left to right, we notice that the potassiums and the chlorines are already balanced with one of each on each side of

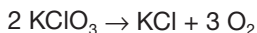
the equation. The oxygens, however, are not balanced. We need more oxygens on the right. Let us place a 2 in front of the O_2 .



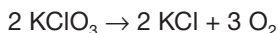
This gives us four oxygens on the right but only three on the left. Let us place a 2 in front of the $KClO_3$.



This gives us six oxygens on the left, but only four on the right. We need more oxygens on the right again. If we place a 3 in front of the O_2 , we will have six oxygens on the right and this matches the six on the left.



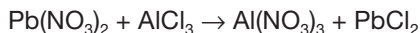
However, the potassium and the chlorine are now not balanced. Placing a 2 in front of the KCl balances the potassium and the chlorine and does not change the balance of the oxygens.



The equation is now balanced.

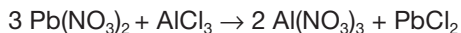
Example 8.2

Balance the following equation:



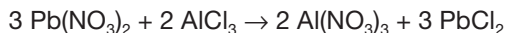
Solution 8.2

The nitrate polyatomic ion (NO_3^-) is seen on both sides of this equation. One helpful hint when you have the same polyatomic ion on both sides is to balance this ion as a group rather than doing the component elements individually. Keeping this in mind, let us balance the equation, starting with the first element on the left and proceeding to the right, as before. There is one lead on both sides, so they are already balanced, with one on each side. Next is the nitrate ion. There are two on the left but three on the right. If we place a 3 in front of the $Pb(NO_3)_2$ and a 2 in front of the $Al(NO_3)_3$, the nitrate ion will be balanced, with six on each side.



Proceeding to the aluminum, we need a 2 in front of the $AlCl_3$ to balance the two aluminums on the right. This gives us six chlorines on the left;

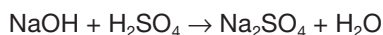
and these are balanced by placing a 3 in front of the PbCl_2 on the right.



Rechecking the lead, we see that these are now balanced, as are all the other elements in the equation.

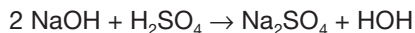
Example 8.3

Balance the following equation:

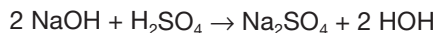


Solution 8.3

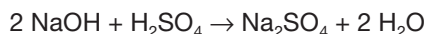
As in Example 8.2, this example has a polyatomic ion on both sides, in this case the sulfate ion, SO_4^{2-} . We will proceed to balance the sulfate as a group. This is also an example of the reaction of an acid with a base, with NaOH being the base (recognized as a base because of the hydroxide polyatomic ion) and H_2SO_4 being the acid (recognized as an acid because there is an H given first in its formula). In these so-called neutralization reactions, water is always a product. One trick that helps in balancing this equation is to consider the H_2O as HOH so that the hydroxide can be balanced as a group. Proceeding now to balance the equation and starting with sodium, we see that we need a 2 in front of NaOH so that there are two sodiums on each side.



Proceeding to the right, we balance the hydroxide groups next. We need a 2 in front of the HOH so that we will have 2 hydroxides on both sides.



Continuing to the right, we see that the hydrogens that are not part of the hydroxide groups are already balanced, with two on the left (in the H_2SO_4) and two on the right (the first H in HOH). The sulfates are also already balanced, with one on each side. All elements are thus balanced. It is now desirable to convert the HOH back to H_2O . The final balanced equation then is as shown below



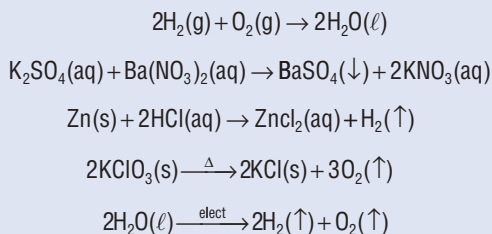
Verbalizing the equation above, we would say, “Two formula units of NaOH react with one molecule of sulfuric acid to produce one formula unit of sodium sulfate and two molecules of water.” In terms of observable and measurable realities using the mole, we would say, “Two moles of sodium hydroxide react with one mole of sulfuric acid to produce one mole of sodium sulfate and two moles of water.”

8.3 Optional Symbols

Besides identifying what substances are the reactants and products of a chemical reaction and how many atoms, molecules, and formula units are involved, chemical equations can provide additional information if the writer so chooses. Symbols may be used to provide some qualitative information about the substances involved or about the reaction in general. However, such symbols are usually optional. Symbols giving specific information about the reactants or products are placed in parentheses immediately to the right of the chemical symbol or formula in the equation. Symbols giving information about the reaction in general are usually placed above the arrow.

Three examples of symbols that are placed over the arrows are the symbol for heat (Δ), the symbol or formula for a catalyst (such as Pt for platinum), and the abbreviation *Elect* to indicate that electrolysis is required. A catalyst is a material that helps speed up a reaction but that is not chemically altered in the process. Electrolysis is a process whereby a chemical reaction is caused by the passage of electrical current through a solution. Some examples of equations in which these optional symbols are used are presented in Table 8.1.

Table 8.1 Examples of Chemical Equations (Including Special Symbols Discussed in the Text)



The symbols (s) (solid), (ℓ) (liquid), and (g) (gas) indicate the physical state of a substance. The symbol (aq) indicates that a substance is dissolved in water and is an abbreviation of the word *aqueous*, which means dissolved in water. The word *aqueous* is derived from the Latin word *aqua*, which means water. To indicate that a product of the reaction is a gas that forms in a reaction medium that is not otherwise gaseous, we use an up arrow (↑). This symbol refers to the fact that such a gas often escapes the reaction medium and is lost to the atmosphere. To indicate that a product of the reaction is a solid material that forms under the surface of a liquid (a precipitate), we use a down arrow (↓). This symbol refers to the fact that such a solid often settles to the bottom of the container. Figures 8.1 and 8.2 show drawings of a gaseous product and a precipitate as described here and should help clarify these two symbols.

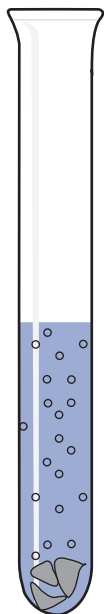


FIGURE 8.1 Hydrogen gas bubbles up and out of the solution when zinc metal is in contact with a solution of hydrochloric acid. The (↑) symbol would be used with the symbol for hydrogen in the equation for this reaction. (From Kenkel, J., Kelter, P., and Hage, D., *Chemistry: An Industry-Based Introduction with CD-ROM*, CRC Press, Boca Raton, FL, 2001. With Permission.)

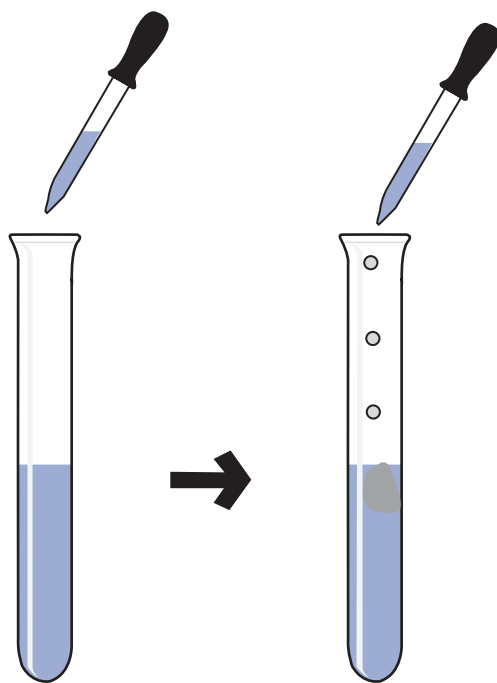
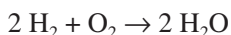


FIGURE 8.2 A solid material, or precipitate, can form when an ingredient of one solution reacts with an ingredient in another. The (↓) symbol is placed to the right of the formula for the precipitate in the equation for this reaction. (From Kenkel, J., Kelter, P., and Hage, D., *Chemistry: An Industry-Based Introduction with CD-ROM*, CRC Press, Boca Raton, FL, 2001. With Permission.)

8.4 Calculations Based on Equations

How can the observable and measurable quantities of matter expressed in chemical equations be calculated? The balancing coefficients play the central role. Let us look at a simple example to illustrate the point, comparing the moles of oxygen that will react with a given number of moles of hydrogen.

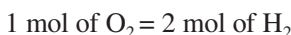


This equation states that 2 mol of H_2 react with 1 mol of O_2 . It is a two-to-one (2:1) ratio. There are twice as many moles of hydrogen involved as there are moles of oxygen. If 4 mol of hydrogen were to react, 2 mol of oxygen would be involved. If 10 mol of hydrogen were used, 5 mol of oxygen would

react. If 0.924 mol of hydrogen were involved, 0.462 mol of oxygen would be needed. It is always a 2:1 ratio of moles of hydrogen to moles of oxygen. Thus the ratio of moles of one chemical to the moles of another is involved, and this ratio is determined from the coefficients in the balanced equation. Grams may also be involved, as we will see later.

8.4.1 Mole-to-Mole Conversions

The concept of dimensional analysis may be applied to the above exercise and is very useful in avoiding mistakes, particularly if the ratio is not as straightforward as 2:1. Using the hydrogen-oxygen reaction as an example, the equality from which a conversion factor is derived is the following:



The conversion factors are as follows:

$$\frac{1 \text{ mol O}_2}{2 \text{ mol H}_2} \quad \text{and} \quad \frac{2 \text{ mol H}_2}{1 \text{ mol O}_2}$$

Such conversion factors can be used to convert moles of one chemical in an equation to moles of another.

Example 8.4

How many moles of Fe_3O_4 can be made from 0.559 moles of Fe according to the following reaction?



Solution 8.4

There is 1 mol of Fe_3O_4 involved for every 3 mol of Fe. We want the moles of Fe to cancel and the units of the answer to be moles of Fe_3O_4 .

$$0.559 \cancel{\text{ mol Fe}} \times \frac{1 \text{ mol Fe}_3\text{O}_4}{3 \cancel{\text{ mol Fe}}} = 0.186 \text{ mol Fe}_3\text{O}_4$$

Example 8.13 is an example of a mole-to-mole conversion—one in which moles of one chemical involved in a chemical reaction are converted to moles of another. The conversion factor is always the mole ratio, as determined from the balanced equation.

It should also be pointed out that the chemical for which an amount is given can be a reactant or a product and that the chemical for which an amount is sought can also be either a reactant or a product. In the illustration of hydrogen, oxygen, and water, two reactants were involved, while in Example 8.13 one was a reactant and the other was a product. In any case, the mole ratio is always set up as follows:

$$\frac{\text{Sought}}{\text{Given}}$$

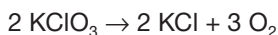
so that the moles unit of the given chemical cancels and the moles unit of the chemical being calculated remains.

8.4.2 Mole-to-Gram and Gram-to-Mole Conversions

We have learned that grams of an element or compound can be easily converted to moles and that moles can be easily converted to grams. The atomic or formula weight is the conversion factor in either case. Because of this, we can expand our discussion of calculations based on equations to include grams of the chemical sought or given. In other words, given the grams of one chemical involved in a chemical reaction, we can calculate the moles of another chemical, or given the moles of one chemical involved in a chemical reaction, we can calculate the grams of another. We thus have gram-to-mole and mole-to-gram conversions, each requiring two conversions, or steps.

Example 8.5

How many moles of O_2 can be expected to form from 0.528 grams of KClO_3 according to the following reaction?



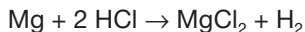
Solution 8.5

This is a gram-to-mole conversion. Grams of KClO_3 must be converted to moles before we can use the mole ratio from the equation to get moles of oxygen. The formula weight of KClO_3 is 122.548 g/mol.

$$0.528 \text{ g } \text{KClO}_3 \times \frac{1 \text{ mol } \text{KClO}_3}{122.548 \text{ g } \text{KClO}_3} \times \frac{3 \text{ mol } \text{O}_2}{2 \text{ mol } \text{KClO}_3} = 0.00646 \text{ mol } \text{O}_2$$

Example 8.6

How many grams of Mg are needed to react with 1.392 moles of HCl according to the following reaction?

**Solution 8.6**

This is a mole-to-gram conversion. Moles of HCl are converted to moles of Mg using the mole ratio from the equation. The moles of Mg are then converted to grams using the atomic weight of magnesium.

$$1.392 \cancel{\text{ mol HCl}} \times \frac{1 \cancel{\text{ mol Mg}}}{2 \cancel{\text{ mol HCl}}} \times \frac{24.305 \text{ g Mg}}{1 \cancel{\text{ mol Mg}}} = 16.92 \text{ g Mg}$$

8.4.3 Gram-to-Gram Conversions

Finally, a calculation involving equations may be a gram-to-gram conversion. That is, given grams of one chemical involved in a chemical reaction, we can determine how many grams of another are involved. Such a calculation requires three steps: (1) conversion of grams to moles, (2) conversion of moles of one chemical to moles of the other, and (3) conversion of these moles to grams. A graphical representation of reaction stoichiometry calculations is given in Figure 8.3.

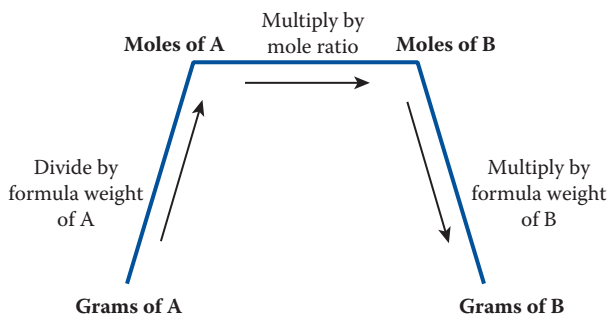
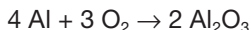


FIGURE 8.3 A graphical depiction, also called a mole map, showing the operations required for mole-to-mole, gram-to-mole, mole-to-gram, and gram-to-gram stoichiometry conversions as discussed in the text. (From Kenkel, J., Kelter, P., and Hage, D., *Chemistry: An Industry-Based Introduction with CD-ROM*, CRC Press, Boca Raton, FL, 2001. With Permission.)

Example 8.7

How many grams of Al_2O_3 can be obtained from 0.500 grams of Al according to the following reaction?

**Solution 8.7**

This is a gram-to-gram conversion: Given 0.500 grams of Al, how many grams of Al_2O_3 can form? Thus the three steps are needed. The formula weight of Al_2O_3 is 101.961 g/mol.

$$0.500 \text{ g Al} \times \frac{1 \text{ mol Al}}{26.982 \text{ g Al}} \times \frac{2 \text{ mol Al}_2\text{O}_3}{4 \text{ mol Al}} \times \frac{101.961 \text{ g Al}_2\text{O}_3}{1 \text{ mol Al}_2\text{O}_3} = 0.945 \text{ g Al}_2\text{O}_3$$

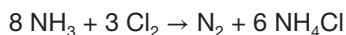
8.5 Limiting Reactant

The quantity of product obtained in a chemical reaction is called the **yield**. As we have seen, the yield depends on what quantities of reactants are used. Using the methods discussed in Section 8.4 and given the moles or grams of a reactant, the moles or grams of a product (the yield) can be calculated. But what if there are two reactants and the moles or grams of both are given? What will be the yield in this case? The yield that results when quantities of two reactants are given depends on which reactant is the limiting reactant. The limiting reactant is the reactant that limits the amount of product formed due to the fact that it is consumed first in the process. A simple analogy here will clarify the concept. Suppose you were preparing holiday greeting cards to mail to your family and friends. Each mailing requires an envelope (with a card) and a stamp. You begin with a book of stamps and a stack of envelopes with cards inside. The yield is the number of mailings that you are able to complete. You find that one book of stamps is not enough and you run out. The entire process stops until you are able to buy more stamps. The number of stamps limits the yield. The stamps are gone, but cards and envelopes are left over. The stamps represent the limiting reactant in a reaction. The completed mailings represent the product, the left over cards and envelopes represent the other reactant, which is present in excess.

In a chemical reaction in which the amounts of two reactants are given, the yield will depend on the amount of the limiting reactant. The reactant that gives less product is the **limiting reactant**. To solve a limiting reactant problem (when the amounts of two reactants are given and you are asked to determine the limiting reactant or the yield), calculate the amounts of products possible from both reactants individually. The reactant that gives less product is the limiting reactant, the lesser amount of product is the yield, and the other reactant is said to be present in excess.

Example 8.8

What is the yield of NH_4Cl (in grams) if 0.238 mol of Cl_2 and 1.627 g of NH_3 are available for the following reaction? Which reactant is the limiting reactant?



Solution 8.8

The amount of NH_4Cl possible from the NH_3 given and the amount of NH_4Cl possible from the amount of Cl_2 given must both be calculated. The reactant that gives less NH_4Cl is the limiting reactant, and the lesser amount of NH_4Cl is the yield. The formula weight of NH_3 is 17.031 g/mol, the formula weight of Cl_2 is 70.906 g/mol, and the formula weight of NH_4Cl is 53.492 g/mol.

$$0.238 \cancel{\text{mol Cl}_2} \times \frac{6 \cancel{\text{mol NH}_4\text{Cl}}}{3 \cancel{\text{mol Cl}_2}} \times \frac{53.492 \text{ g NH}_4\text{Cl}}{1 \cancel{\text{mol NH}_4\text{Cl}}} = 25.5 \text{ g NH}_4\text{Cl}$$

$$\begin{aligned} 1.627 \cancel{\text{g NH}_3} \times \frac{1 \cancel{\text{mol NH}_3}}{17.031 \cancel{\text{g NH}_3}} \times \frac{6 \cancel{\text{mol NH}_4\text{Cl}}}{8 \cancel{\text{mol NH}_3}} \times \frac{53.492 \text{ g NH}_4\text{Cl}}{1 \cancel{\text{mol NH}_4\text{Cl}}} \\ = 3.833 \text{ g NH}_4\text{Cl} \end{aligned}$$

The NH_3 is the limiting reactant because less product is obtained from it. The yield for the reaction is 3.833 g because it is the lesser of the two amounts.

8.6 Theoretical Yield, Actual Yield, and Percent Yield

Often the yield realized from a given reaction is less than the yield calculated from the data. Some possible reasons for this include (1) sloppy lab technique, such that some reactant does not get into the reaction or some product does not get measured, (2) side reactions, so that other products form that are not part of the reaction on which the calculation was based, and (3) reversible reactions, in which the given reaction does not go to completion and some reactant material remains.

In these cases, it may be customary to calculate a percent yield, which is a number that indicates how successful the process was. To calculate the percent yield, the **actual yield** (i.e., the yield that was physically obtained in the reaction) and the **theoretical yield** (i.e., the yield that was expected from the calculations) must be known or calculated. The **percent yield** is calculated as follows:

$$\text{Percent yield} = \frac{\text{Actual yield}}{\text{Theoretical}} \times 100$$

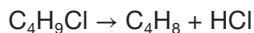
The actual yield and theoretical yield must have the same units (mole or grams). The calculated percent yield should always be a number equal to 100% or less than 100%.

Example 8.9

The following reaction can be used to synthesize t-butyl alcohol ($\text{C}_4\text{H}_9\text{OH}$):



However, the following is a side reaction that also occurs, so the percent yield of the $\text{C}_4\text{H}_9\text{OH}$ will be less than 100%:



If 2.19 g of $\text{C}_4\text{H}_9\text{Cl}$ are used and 0.0211 mol of $\text{C}_4\text{H}_9\text{OH}$ actually form, what is the percent yield? The formula weight of $\text{C}_4\text{H}_9\text{Cl}$ is 92.569 g/mol.

Solution 8.9

We can calculate the theoretical yield in moles, since the actual yield is in moles.

$$2.19 \text{ g } \text{C}_4\text{H}_9\text{Cl} \times \frac{1 \text{ mol } \text{C}_4\text{H}_9\text{Cl}}{92.569 \text{ g } \text{C}_4\text{H}_9\text{Cl}} \times \frac{1 \text{ mol } \text{C}_4\text{H}_9\text{OH}}{1 \text{ mol } \text{C}_4\text{H}_9\text{Cl}}$$

$$= 0.023658 \text{ mol } \text{C}_4\text{H}_9\text{OH}$$

Then the percent yield is as follows:

$$\text{Percent yield} = \frac{\text{Actual yield}}{\text{Theoretical}} \times 100$$

$$\text{Percent yield} = \frac{0.0211 \text{ mol } \text{C}_4\text{H}_9\text{OH}}{0.023658 \text{ mol } \text{C}_4\text{H}_9\text{OH}} \times 100 = 89.2\%$$

8.7 Homework Exercises

Reactants and Products

- In the following chemical equations, tell what chemicals are the reactants and what chemicals are the products.
 - $\text{BaCl}_2 + \text{K}_2\text{SO}_4 \rightarrow \text{BaSO}_4 + 2 \text{KCl}$
 - $2 \text{KOH} + \text{H}_2\text{C}_2\text{O}_4 \rightarrow \text{K}_2\text{C}_2\text{O}_4 + 2 \text{H}_2\text{O}$
 - $\text{Fe} + \text{O}_2 \rightarrow \text{Fe}_2\text{O}_3$
- State in words what is happening in the reactions in Exercise 1.

Balancing Equations

- What is the law of conservation of mass?
- Balance the following equations.
 - $\text{Fe} + \text{O}_2 \rightarrow \text{FeO}$
 - $\text{P} + \text{O}_2 \rightarrow \text{P}_2\text{O}_5$
 - $\text{SO}_2 + \text{O}_2 \rightarrow \text{SO}_3$
 - $\text{H}_3\text{PO}_4 + \text{NaOH} \rightarrow \text{Na}_3\text{PO}_4 + \text{H}_2\text{O}$
 - $\text{BiCl}_3 + \text{H}_2\text{O} \rightarrow \text{BiOCl} + \text{HCl}$
 - $\text{S} + \text{O}_2 \rightarrow \text{SO}_3$
 - $\text{FeO} + \text{O}_2 \rightarrow \text{Fe}_2\text{O}_3$
 - $\text{Ag} + \text{H}_2\text{S} + \text{O}_2 \rightarrow \text{Ag}_2\text{S} + \text{H}_2\text{O}$
 - $\text{NO}_2 + \text{H}_2\text{O} \rightarrow \text{HNO}_3 + \text{NO}$
 - $\text{HgO} \rightarrow \text{Hg} + \text{O}_2$
 - $\text{Mg} + \text{H}_2\text{O} \rightarrow \text{H}_2 + \text{Mg}(\text{OH})_2$

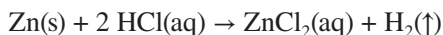
- l. $\text{N}_2 + \text{O}_2 \rightarrow \text{N}_2\text{O}$
 - m. $\text{PCl}_5 + \text{H}_2\text{O} \rightarrow \text{POCl}_3 + \text{HCl}$
 - n. $\text{HCl} + \text{Ba}(\text{OH})_2 \rightarrow \text{BaCl}_2 + \text{H}_2\text{O}$
5. State which of the reactions represented by the equations in exercise 4 are acid–base reactions, and explain why.

Optional Symbols

6. State the meaning of each of the following chemical symbols.
- a. (s)
 - b. (g)
 - c. (ℓ)
 - d. (↑)
 - e. (↓)
 - f. (Δ)
7. Consider the following chemical equation:



- a. Write the formula of a chemical that is a reactant.
 - b. Write the formula of a chemical that is a liquid.
 - c. Write the formula of a product that is a gas.
8. A chemist observes the following reaction in her laboratory: magnesium metal (Mg) and hydrobromic acid (HBr) are the reactants; magnesium bromide (MgBr_2) and hydrogen gas (H_2), which bubbles up out of the solution, are the products. The HBr and the MgBr_2 are in water solution. Given this information, write a balanced chemical equation incorporating all the symbolism appropriate for this reaction.
9. Consider the following chemical equation:



Explain what the symbols (s), (aq), and (↑) mean in this equation.

10. Given the following information, write a balanced chemical equation that includes all the appropriate formulas and symbols in their proper places.

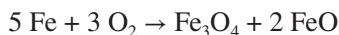
The products are PbCrO_4 and KNO_3 . The reactants are $\text{Pb}(\text{NO}_3)_2$ and K_2CrO_4 . The reaction requires heat in order to proceed. The $\text{Pb}(\text{NO}_3)_2$, KNO_3 , and K_2CrO_4 are in water solution. The PbCrO_4 is a precipitate.

Reaction Stoichiometry

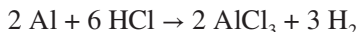
11. State with words the following equation in terms of atoms, molecules, and formula units and then restate the equation in terms of moles.



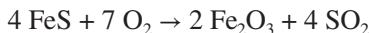
12. Given the following reaction:



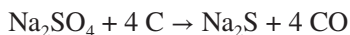
- How many moles of O_2 are needed to react with 0.00382 mol of Fe?
 - How many grams of FeO can be expected to form when 9.2 g of Fe react?
 - How many moles of Fe_3O_4 will form when 0.559 g of FeO form?
 - How many grams of FeO will form when 1.629 g of Fe react?
13. Given the following reaction:



- How many moles of Al will be used if 1.118 mol of HCl are used?
 - How many moles of H_2 form if 938.2 g of AlCl_3 form?
 - How many grams of AlCl_3 will form if 0.338 mol of HCl are used?
 - How many grams of Al are need to make 34.22 g of AlCl_3 ?
14. A chemist starts with 2.045 mol of FeS and performs the following chemical reaction:



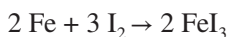
- How many moles of oxygen (O_2) will react?
 - How many grams of Fe_2O_3 should form?
15. A chemist uses 5.29 g of carbon (C) and performs the following reaction:



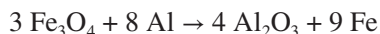
- How many moles of Na_2SO_4 are needed?
- How many grams of Na_2S can form?

Limiting Reactant

16. Given the following reaction, if 0.229 mol of iron (Fe) and 0.318 moles of I_2 are available, which would be the limiting reactant? Why?



17. 0.329 g of Fe_3O_4 and 0.812 g of Al are available for the following reaction:



What would be the limiting reactant? How many grams of iron (Fe) would be formed?

18. 0.0728 g of Cu_2O and 0.822 moles of O_2 are available for the following reaction:



What reactant would be the limiting reactant, and how many grams of CuO would form?

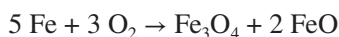
19. Consider the following balanced equation:



Complete the table.

	Moles of $\text{Al}_2(\text{CO}_3)_3$	Moles of HCl	Limiting Reactant	Moles of AlCl_3 Possible
(a)	1.0 mol	5.0 mol		
(b)	5.0 mol	1.0 mol		
(c)	0.00182 mol	0.0138 mol		
(d)	0.0838 mol	0.284 mol		

20. Consider the following reaction:



Complete the table.

	Grams of Fe	Moles of O_2	Limiting Reactant	Grams of FeO Possible
(a)	0.548 g	0.038 mol		
(b)	2.084 g	1.79 mol		
(c)	14.5 g	0.056 mol		
(d)	56.1 g	0.112 mol		

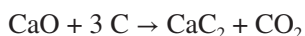
Percent Yield

21. A chemist performs the following chemical reaction starting with 0.588 g of antimony and obtains 0.643 g of Sb_4O_6 .



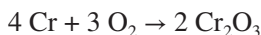
What is the percent yield of the Sb_4O_6 ?

22. A laboratory technician performs the following reaction using 4.339 g of carbon and obtains 1.993 g of CaC_2 .



What is the percent yield?

23. Starting with 9.228 g of Cr, a chemist isolates 0.0803 moles of Cr_2O_3 according to the following reaction:

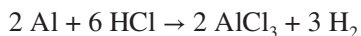


What is the percent yield?

Complete the following table.

	Amount of Cr	Actual Yield of Cr_2O_3	Theoretical Yield of Cr_2O_3	Percent Yield
(a)	0.00382 mol	0.00186 mol		
(b)	9.6 mol	705 g		
(c)	1.559 g	0.013 mol		
(d)	43.6 g	59.03 g		

24. Consider the following balanced equation:



Complete the following table.

	Amount of Al	Actual Yield of AlCl_3	Theoretical Yield of AlCl_3	Percent Yield
(a)	0.183 mol	0.179 mol		
(b)	2.004 mol	248 g		
(c)	62.4 g	2.04 mol		
(d)	9.05 g	41.2 g		

Chapter 9

Gases and the Gas Laws

Introduction	214
Properties of Gases	214
Kinetic–Molecular Theory of Gases	218
Gas Pressure	220
Boyle’s Law	224
Charles’ Law	227
Combined Gas Law	230
Standard Temperature and Pressure	231
Correction of a Gas Volume to STP	231
Ideal Gas Law	232
Avogadro’s Law	234
Dalton’s Law	235
Vapor Pressure	236
Homework Exercises	239

9.1 Introduction

Gases require special attention because they display special and unique properties— properties different from liquids and solids. In this chapter, we examine these properties and describe a theory by which they can be explained. In addition, a number of mathematical relationships, referred to as gas laws, will be developed and used to solve problems. We will also explore the stoichiometry of chemical reactions in which gases are involved.

9.2 Properties of Gases

Compared with liquids and solids, gases exhibit some very unique properties. Let us take a look at a few of these properties.

1. *Gases can be compressed.* Imagine a gas contained in a cylindrical container fitted with a movable piston, as shown on the left of Figure 9.1. With some force, we can push the piston down and

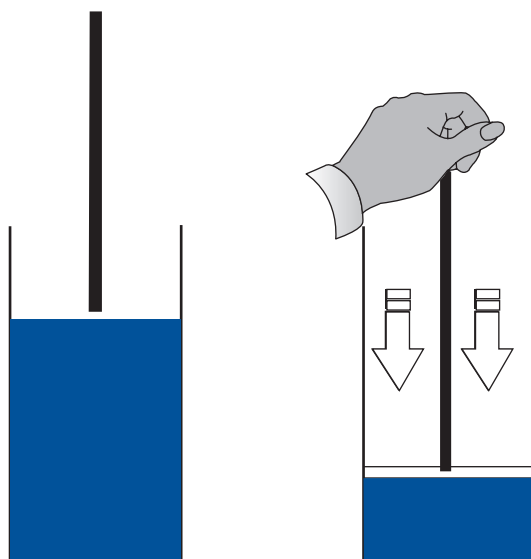


FIGURE 9.1 Unlike liquids and solids, gases can be compressed, or forced to occupy a smaller volume. (From Kenkel, J., Kelter, P., and Hage, D., *Chemistry: An Industry-Based Introduction with CD-ROM*, CRC Press, Boca Raton, FL, 2001. With Permission.)

compress the gas, or force it to occupy a smaller volume, as shown on the right in this figure. If you imagine a liquid in the container, it may be possible to push the piston down a slight distance, but nothing like when it is filled with a gas. If a solid is in this container, it would not be possible to push the piston down at all. Gases can be compressed while liquids and solids cannot.

2. *Gases exert pressure in all directions.* To maintain the piston at the position indicated on the right in Figure 9.1, the force used to push the piston down must be maintained. If we lessen this force, the piston will return to its original position. Thus there is a pressure pushing up on the piston. The contained gas is exerting a pressure, and this pressure is in all directions within the container. Liquids also exert pressure in all directions if they have been compressed as noted above, but solids (and liquids that have not been compressed) exert only a downward pressure, due to the force of gravity upon them.
3. *Gases can expand.* Imagine again the gas contained in the cylindrical container that is fitted with the piston. This time, however, imagine that we pull up on the piston rather than push down, as shown in Figure 9.2. With some force, we can pull the piston up. The gas would then fill the container and occupy a larger volume. With a liquid or solid filling the container, it would not be possible to move the piston upward. Thus gases can expand in this situation while liquids and solids cannot.
4. *Gases diffuse through one another.* When natural gas, methane mixed with an additive to give it an odor, is released into a room, the occupants of the room, no matter where they are in the room, will soon detect the odor. The methane gas and the additive diffuse in all directions through air. We have all detected both pleasant and unpleasant odors of various gaseous substances at different times. Examples include the odor of a skunk, the odor of a freshly baked apple pie, and the odor of perfume. It does not matter where we are standing relative to the source of the odor, we still smell it. We detect these odors because gaseous substances diffuse through air in all directions from the source. Indeed, *all gaseous substances diffuse through all other gaseous substances* regardless of the identity of the gases involved. Two liquids, such as alcohol and water, may diffuse through (or mix with) each other but two other liquids, such as oil and

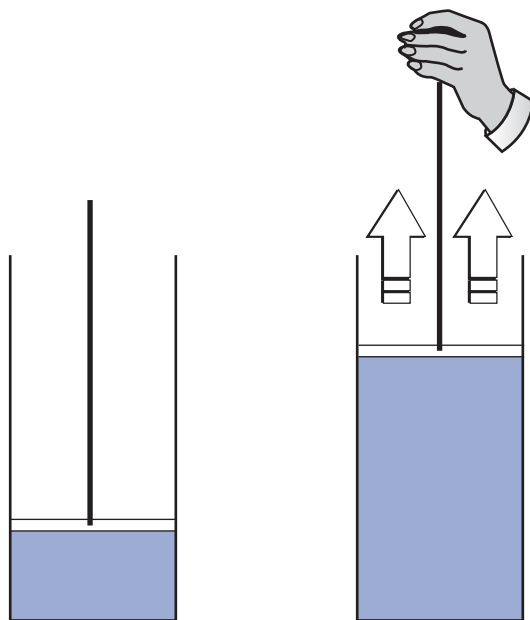


FIGURE 9.2 Gases expand and fill a container when the container is made larger. Liquids and solids do not. (From Kenkel, J., Kelter, P., and Hage, D., *Chemistry: An Industry-Based Introduction with CD-ROM*, CRC Press, Boca Raton, FL, 2001. With Permission.)

water, may not. Mixing of two solids never occurs, regardless of the identities of the solids.

5. *Gases expand when heated and contract when cooled.* If we place the cylinder with the movable piston on a hot plate, as pictured in Figure 9.3, and proceed to heat the contained gas without applying any upward or downward force on the piston, the piston moves upward as the gas expands and occupies a larger volume. Similarly, if we cool the gas by placing the cylinder in an ice bath, as pictured in Figure 9.4, the gas contracts and occupies a smaller volume. The volume of a gas depends on its temperature. Liquids behave similarly, as noted in Chapter 2 in the discussion of thermometers. The volume of a liquid depends on its temperature, and that is why thermometers can be constructed with liquids inside—the liquid level depends on the volume of the liquid, which in turn depends on the temperature. There is only a very small effect of temperature on the volume of solids.

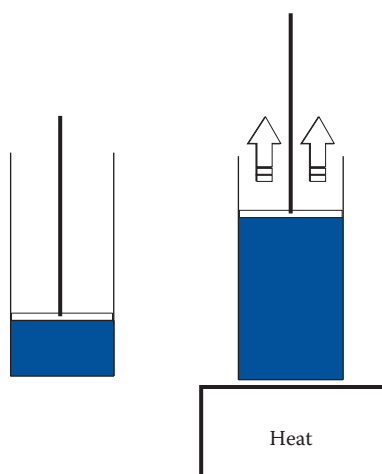


FIGURE 9.3 When a gas is heated, its volume increases. (From Kenkel, J., Kelter, P., and Hage, D., *Chemistry: An Industry-Based Introduction with CD-ROM*, CRC Press, Boca Raton, FL, 2001. With Permission.)

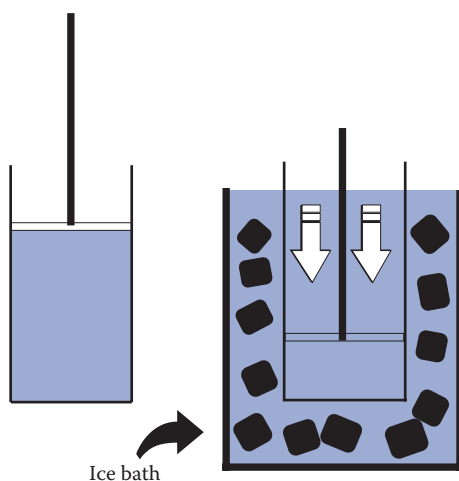


FIGURE 9.4 When a gas is cooled, its volume decreases. (From Kenkel, J., Kelter, P., and Hage, D., *Chemistry: An Industry-Based Introduction with CD-ROM*, CRC Press, Boca Raton, FL, 2001. With Permission.)

6. *If the volume of a container filled with a gas is fixed, the pressure will increase if the gas is heated and decrease if the gas is cooled.* If the piston on the left in Figure 9.3 were locked in place, the gas volume would be fixed. When heated, the piston could not move and, as a result, the gas pressure would increase. Similarly, when cooled, the piston could not move and the gas pressure would decrease. This would be similar to pushing down on the piston with increasing force (to prevent its upward movement) as the gas was heated, or pulling up on the piston with increasing force (to prevent its downward movement) as the gas was cooled. The increased downward force is due to the increased gas pressure, and the increased upward force is due to the decreased gas pressure.
7. If more gas is added to a container of fixed volume that is already filled with a gas, the pressure increases. If some gas is removed from a container of fixed volume, the pressure decreases. For example, when more air is pumped into a tire that is already filled with air, the pressure increases (as measured by the tire gauge). If some air is released from a tire filled with air, the pressure decreases.

9.3 Kinetic–Molecular Theory of Gases

Scientists attempt to explain the various properties of gases, liquids, and solids in terms of the properties of the atoms, molecules, or formula units of which they are composed. The theory used to do this is called the **kinetic-molecular theory**, or in the case of gases, the **kinetic-molecular theory of gases**. It is like looking through the mythical microscope to see what the atoms or molecules are doing. This theory has the following key points:

1. All gases are composed of atoms or molecules that are far apart from one another.
2. These atoms or molecules are in constant motion, colliding with each other and with the walls of the container.
3. Because these atoms or molecules are in constant motion, they possess the energy of motion, or **kinetic energy (K.E.)**. K.E. depends on the mass of the moving atom or molecule and the velocity with which it moves. As discussed in Chapter 4, the mathematical definition of K.E. is

$$\text{K.E.} = \frac{1}{2} mv^2$$

In this case, m is the mass of the atom or molecule and v is the velocity with which it moves.

4. The collisions are perfectly elastic, meaning that the energy of the atoms or molecules is never diminished as a result of the collisions, and so the motion never stops.
5. The atoms or molecules of all gases have the same kinetic energy at a given temperature, regardless of the identity of the gas. Since the masses of the atoms or molecules of different gases (the atomic or molecular weight) vary according to their identity, the heavier the atom or molecule is, the slower it moves.
6. Kinetic energy changes with temperature. The heat energy added to increase the temperature causes the velocity of the atom or molecule, and its kinetic energy, to increase. Conversely, the heat energy removed to decrease the temperature causes the velocity, and thus the kinetic energy, to decrease.

Let us now proceed to explain the various unique properties that gases have by using this theory and by looking through the mythical microscope.

- *Why can gases be compressed?* The atoms or molecules are far apart from one another and thus can be made to come closer together.
- *Why do gases exert a pressure in all directions?* The gas atoms or molecules are in constant motion and are constantly colliding with the walls of the container, so they exert a force on the walls. This force is observed as pressure.
- *Why can gases expand?* Gas atoms or molecules are far apart from one another and can be forced even farther apart. Since the gas molecules are in constant motion, they immediately fill all parts of the container.
- *Why do gases diffuse through one another?* Besides being far apart from one another, the atoms or molecules are in constant motion and thus move freely among other gas atoms or molecules, regardless of their identities.
- *Why do gases expand when heated and contract when cooled, and why, if the volume is fixed, does the pressure increase when the gas is heated and decrease when the gas is cooled?* The added heat energy is converted to kinetic energy, and thus the velocities of the atoms or molecules increase. If the velocities increase, they collide with the

walls of the container more often. Since gas pressure is the result of these collisions, the gas pressure increases. If the volume is not fixed, however, it will increase as a result of the increased pressure. When the gas is cooled, heat energy is removed, causing the kinetic energy to decrease, which in turn means that the velocities decrease. Lower velocities mean that the collision rate with the walls of the container decreases, and thus the pressure is lowered. If the volume is not fixed, then the volume decreases.

- *Why does pressure increase when more gas is added to a container of fixed volume and decrease when gas is removed from a container of fixed volume?* More gas atoms or molecules mean more collisions with the walls of the container, and more collisions mean increased pressure. Fewer gas atoms or molecules mean fewer collisions and thus decreased pressure.

9.4 Gas Pressure

One of the properties of gases discussed above is that of **gas pressure**. Gas pressure is something that is routinely measured in everyday life. For example, when you pump your bicycle tire or your automobile tire full of air, it is desirable to measure the pressure of the air inside the tire with a pressure gauge to determine whether you have added the correct amount of air. Also, when meteorologists measure barometric pressure, they measure the pressure exerted by the air in the vicinity of the weather station. Barometric pressure changes with changing weather conditions, so its measurement is often used to predict weather.

The pressure inside your bicycle tire is most often measured in **pounds per square inch (psi)**, the pounds of pressure exerted per square inch of the area of the inner wall of the tire. This would seem to imply that gas pressure is analogous to weight, which is also measured in pounds. The difference is that gas pressure is exerted in all directions, not just in the downward direction of gravity.

To illustrate the air pressure measured by a meteorologist, consider what happens when you fill a bottle with water by totally immersing it under the water in a sink, then pull the bottle up and out of the water bottom first, as illustrated in Figure 9.5.

Despite the fact that the water level inside the bottle is above the level of the water in the sink, the water does not flow out of the bottle. The water does

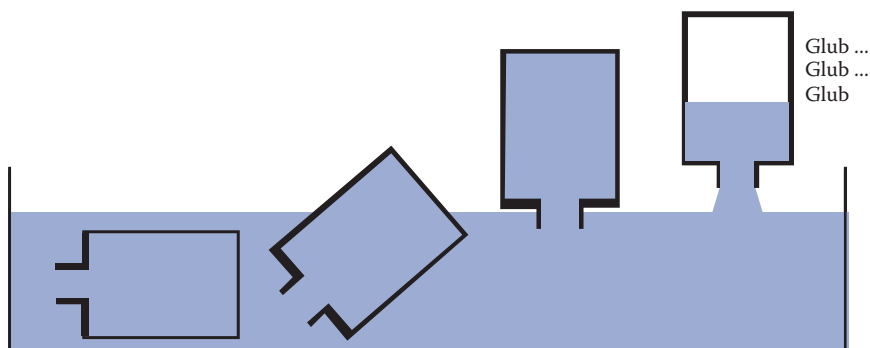


FIGURE 9.5 The water contained in a bottle immersed in water does not flow out when the bottle is lifted bottom-first from the water. Not until the mouth of the bottle is out of the water does the water flow out. See text for discussion.

not flow out until the mouth of the bottle is pulled from the water. The water is held in the bottle because of air molecules exerting pressure on the surface of the water in the sink. This pressure is atmospheric pressure, the pressure exerted by the air in the Earth's atmosphere.

The air pressure measured by a meteorologist is most often reported in inches of mercury (in. Hg). This unit of pressure is derived from the original mercury barometer invented by a seventeenth century scientist named Evangelista Torricelli. Torricelli designed an apparatus in which a long glass tube, closed at one end and filled with mercury, is placed with its open end down in a container of mercury. Just as with the bottle and water experiment, the mercury does not flow from the tube. Rather, it drops to a level the height of which represents a measure of the atmospheric pressure. Torricelli measured the height of this mercury column with a meter stick as illustrated in Figure 9.6, and this was the origin of a unit that has been widely used in chemistry applications, millimeters of mercury (mm Hg). Normal atmospheric pressure at sea level and 0°C is 760 mm Hg. As the atmospheric pressure changes from one day to the next, the height of the mercury in the tube changes. The height of the mercury is thus an effective measure of atmospheric pressure. The height of the mercury can also be measured in inches of mercury; this is the source of the unit by which meteorologists express air pressure. Besides the units of pounds per square inch, millimeters of mercury, and inches of mercury, there are

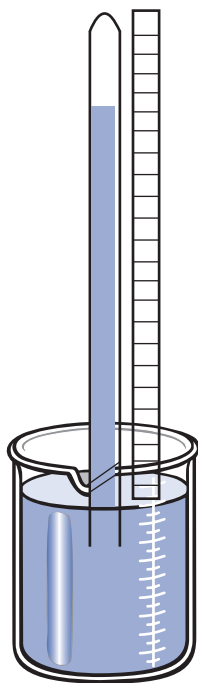


FIGURE 9.6 The Torricelli barometer with a meter stick. (From Kenkel, J., Kelter, P., and Hage, D., *Chemistry: An Industry-Based Introduction with CD-ROM*, CRC Press, Boca Raton, FL, 2001. With Permission.)

Table 9.1 Relationships between Various Pressure Units

1 atm = 760 torr
1 atm = 14.696 psi
1 torr = 0.019337 psi

other important pressure units. For example, for the International System of Units (see footnote in Chapter 2, Section 2.5.), the bar and the pascal are used. Additionally, the metric system uses the units of atmosphere and torr (Table 9.1). In this text, we shall only deal further with the units of psi, atmospheres, and torr.

The torr was named in honor Torricelli. One **torr** is the same as 1 mm Hg. One **atmosphere** (atm) is defined as the pressure required to maintain a height

of exactly 760 mm in the Torricelli barometer. In other words, 760 mm Hg, or 760 torr, is exactly 1 atm. Also, 14.696 psi is 1 atm and 101.3 kilopascals (kPa) is 1 atm. These are summarized in Table 9.1 Normal atmospheric air pressure varies with altitude but will usually be reported by meteorologists in the range of 27–31 in. Hg. This corresponds to about 685–790 torr. Thus 760 torr is within the range of normal atmospheric air pressure. Later in this chapter, we will define 760 torr, or 1 atm, as standard pressure.

Example 9.1

How many torr does 734 mm Hg represent?

Solution 9.1

Since mm Hg and torr are the same unit, 734 mm Hg = 734 torr. Or, using dimensional analysis, we have the following:

$$\frac{734 \cancel{\text{ mm Hg}}}{1 \cancel{\text{ mm Hg}}} \times \frac{1 \text{ torr}}{1 \cancel{\text{ mm Hg}}} = 734 \text{ torr}$$

Example 9.2

How many psi are in 1.05 atm?

Solution 9.2

There are 14.696 psi in 1 atm.

$$\frac{1.05 \cancel{\text{ atm}}}{1 \cancel{\text{ atm}}} \times \frac{14.696 \text{ psi}}{1 \cancel{\text{ atm}}} = 15.4 \text{ psi}$$

Example 9.3

How many atmospheres does 778 torr represent?

Solution 9.3

There are 760 torr in 1 atm.

$$\frac{778 \cancel{\text{ torr}}}{760 \cancel{\text{ torr}}} \times \frac{1 \text{ atm}}{1 \cancel{\text{ torr}}} = 1.02 \text{ atm}$$

Gas pressures are measured using devices called pressure gauges. For example, the pressure of air inside a bicycle tire is measured with a tire gauge. The pressure in a compressed gas cylinder, such as the cylinder that is often used to fill balloons with helium, can be measured by a gauge that is part of the pressure regulator—the valve assembly by which the helium is delivered to the balloon. Such a gauge is used to indicate how much helium remains in the cylinder. The same kind of gauge is used on the acetylene cylinders that welders use or on the large propane tanks that are situated near houses in colder climates where propane is used to heat homes. We will not be concerned with the mechanism by which these gauges measure pressure. However, we will be concerned with the fact that such pressure can be measured, and we will use this measured pressure in calculations.

9.5 Boyle's Law

One of the properties of gases mentioned in Section 9.2 was compressibility. We said that if we push down on the piston pictured in Figure 9.1 with some force, the gas can be made to occupy a smaller volume. At any given point in this process, the force pushing down is equal to the force of the gas pushing up. In other words, the force of pushing down is equal to the pressure exerted by the gas. When the volume of the gas gets smaller by the act of compression, the force needed to push the piston down (the pressure of the gas) increases. The smaller the volume, the higher the pressure. This inverse relationship can be described mathematically. Multiplying the value of the pressure (expressed in torr, atmospheres, or any pressure unit) by the value of the volume (expressed in liters, milliliters, or any volume unit) at any point along the way in the process of pushing down the piston always gives the same number. If we symbolize the pressure as P and the volume as V , we therefore have the following relationship:

$$PV = \text{Constant} \quad (9.1)$$

If we only push the piston down slightly, such that $P = P_1$ and $V = V_1$ as in Figure 9.7, then we get the result shown in Equation 9.2.

$$P_1 V_1 = \text{Constant} \quad (9.2)$$

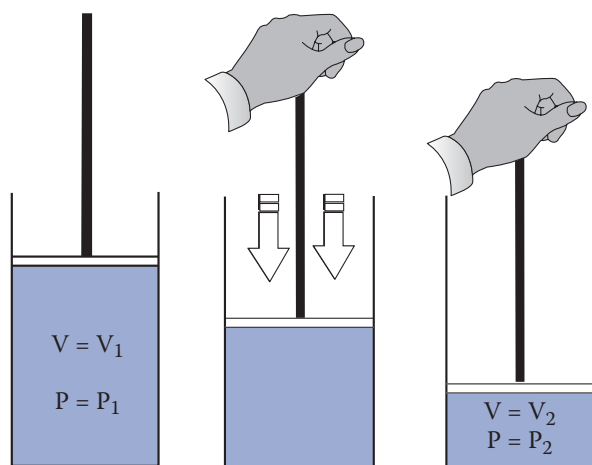


FIGURE 9.7 Pushing down the piston in the piston-cylinder arrangement changes the pressure from P_1 to P_2 and the volume from V_1 to V_2 . (From Kenkel, J., Kelter, P., and Hage, D., *Chemistry: An Industry-Based Introduction with CD-ROM*, CRC Press, Boca Raton, FL, 2001. With Permission.)

If we push the piston down further, such that $P = P_2$ and $V = V_2$, then we get the following:

$$P_2 V_2 = \text{Constant} \quad (9.3)$$

The constants indicated in Equations 9.2 and 9.3 are the same number *if the pressures and volumes were measured in the same units each time*. Thus we have the new relationship given in Equation 9.4.

$$P_1 V_1 = P_2 V_2 \quad (9.4)$$

This is a statement of what has come to be known as Boyle's law. It should be noted that Boyle's law holds true regardless of whether the pressure increases (volume becomes smaller as the piston is pushed down) or decreases (volume becomes larger as the piston is pulled up). The units of V_1 and V_2 must be the same, and the units of P_1 and P_2 must be the same. It should also be noted that the *temperature must not change* in the process of increasing or decreasing the pressure. If the temperature were to change, then the volume would be affected, as shown in Figures 9.3 and 9.4, and Boyle's law would not be valid.

Example 9.4

A gas occupies a volume of 75.3 mL at a pressure of 708 torr. What volume will this gas occupy at 758 torr if the temperature does not change?

Solution 9.4

The temperature has not changed, so Boyle's law holds true. The initial pressure, P_1 , is 708 torr; the final pressure, P_2 , is 758 torr; and the initial volume, V_1 , is 75.3 mL. We must calculate V_2 . Solving for V_2 in Equation 9.4, we have the following:

$$V_2 = V_1 \times \frac{P_1}{P_2}$$

The two pressure values have the same units, torr. The final volume will have the same units as the initial volume.

$$V_2 = 75.3 \text{ mL} \times \frac{708 \text{ torr}}{758 \text{ torr}} = 70.3 \text{ mL}$$

Example 9.5

What is the pressure of a sample of a gas held at a constant temperature when the volume is 6.02 L if the pressure was 0.893 atm when the volume was 3.41 L?

Solution 9.5

The temperature is constant, so Boyle's law holds true. The final pressure is sought, so we must rearrange Equation 9.4 to solve for P_2 .

$$P_2 = P_1 \times \frac{V_1}{V_2}$$

$$P_2 = 0.893 \text{ atm} \times \frac{3.41 \text{ L}}{6.02 \text{ L}} = 0.506 \text{ atm}$$

You may wonder why we rearrange Boyle's law to show a pressure ratio (P_1/P_2 in Example 9.4) multiplying the initial volume or a volume ratio (V_1/V_2 in Example 9.5) multiplying the initial pressure. Such ratios help us to see

whether our answers make sense. For example, if the pressure is increased as in Example 9.4, the larger pressure will be in the denominator and the pressure ratio is a fraction less than 1. This would make the new volume smaller, as we would expect from an increase in pressure. If our calculation yields a larger volume, then we know we must have reversed the pressure values in the equation. Of course, if the pressure has decreased, the opposite is true. We would expect the new volume to be larger. Such ratios also help us to see that these calculations reflect the concepts of dimensional analysis. While we have not indicated that the units cancel, obviously if the units of pressure in the P_1/P_2 ratio are the same and the units in the V_1/V_2 ratio are the same, then they would indeed cancel.

9.6 Charles' Law

The fact that gas volumes change when the temperature changes can also be treated mathematically. In this case, it is a direct relationship rather than an inverse one. As temperature increases, the volume increases. And as temperature decreases, the volume decreases. This means that if we were to *divide* the volume by the temperature we would obtain a constant.

$$\frac{V}{T} = \text{Constant} \quad (9.5)$$

Thus, at a particular temperature, T_1 , a given sample of gas would occupy a volume, V_1 , as shown on the left in Figure 9.8. Dividing this volume by this temperature would then yield a constant.

$$\frac{V_1}{T_1} = \text{Constant} \quad (9.6)$$

If the gas were heated to a new temperature, T_2 , it would expand to a new volume, V_2 , as shown on the right in Figure 9.8. Dividing this volume by this temperature would yield the same constant as before.

$$\frac{V_2}{T_2} = \text{Constant} \quad (9.7)$$

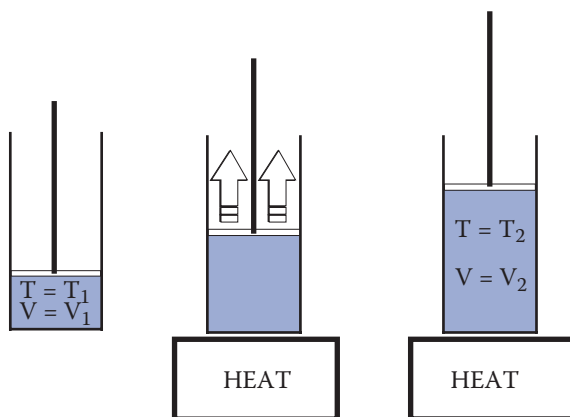


FIGURE 9.8 Heating the piston-cylinder arrangement causes the temperature to increase from T_1 to T_2 and the volume to increase from V_1 to V_2 . (From Kenkel, J., Kelter, P., and Hage, D., *Chemistry: An Industry-Based Introduction with CD-ROM*, CRC Press, Boca Raton, FL, 2001. With Permission.)

From this observation, we can obtain the following relationship:

$$\frac{V_1}{T_1} = \frac{V_2}{T_2} \quad (9.8)$$

This is a statement of what has come to be known as Charles' law. Charles' law holds true regardless of whether the temperature increases or decreases, meaning we could also have cooled the gas sample pictured in Figure 9.8 with an ice bath rather than heated it with a hot plate. The smaller volume divided by the smaller temperature would yield the same constant.

Once again, the units of volume are irrelevant as long as both V_1 and V_2 have the same units. However, both temperatures must be in Kelvin units. Thus if either temperature is given in Celsius or Fahrenheit units, they must be converted to Kelvin units before dividing. Also, Charles' law only holds true if the *pressure is held constant*. Obviously if the pressure were also to change, then the volume would change for a reason other than there being a change in temperature, and Charles' law would not hold true.

Example 9.6

A gas occupies a volume of 5.28 L at a temperature of 31°C. What volume will this gas occupy if the temperature is changed to 61°C while the pressure is held constant?

Solution 9.6

Charles' law applies because the temperature is changing at constant pressure. Solving Equation 9.8 for V_2 , we get the following:

$$V_2 = V_1 \times \frac{T_2}{T_1}$$

Plugging in the temperature and volume values and keeping in mind that the temperature units must be K, we get the answer shown below.

$$V_2 = 5.28 \text{ L} \times \frac{(61^\circ\text{C} + 273.15)\text{K}}{(31^\circ\text{C} + 273.15)\text{K}} = 5.80 \text{ L}$$

Example 9.7

What temperature would be required to change the volume of a gas from 935 mL at 316 K to 722 mL, assuming that the pressure does not change?

Solution 9.7

A change in temperature at constant pressure means that Charles' law is applicable. Rearranging Equation 9.8 for T_2 , we get the following:

$$T_2 = T_1 \times \frac{V_2}{V_1}$$

Substituting in the numerical values, we have the following:

$$T_2 = 316 \text{ K} \times \frac{722 \text{ mL}}{935 \text{ mL}} = 244 \text{ K}$$

Once again, the use of ratios (T_2/T_1 and V_2/V_1) helps us see that the answer makes sense. For example, an increase in temperature should yield a larger volume and a decrease in temperature should give a smaller volume. Knowing this, we can see if the answer is appropriately larger or smaller and conclude that the ratio we used was the correct one. Also, we can see clearly that the units of temperature or volume would cancel appropriately when we use the ratio.

9.7 Combined Gas Law

What happens if both the pressure and the temperature change? Boyle's law and Charles' law can be combined to produce a more general law that we can use to treat these situations. This law is called the combined gas law and can be represented as follows:

$$\frac{P_1 V_1}{T_1} = \frac{P_2 V_2}{T_2} \quad (9.9)$$

Example 9.8

What volume will a gas sample occupy if initially the volume is 0.227 L, and the pressure is changed from 716 torr to 734 torr, while the temperature is changed from 31°C to 16°C?

Solution 9.8

Both the pressure and the temperature are changing, so the combined gas law applies. Rearranging Equation 9.9 to solve for V_2 yields the following:

$$V_2 = V_1 \times \frac{P_1}{P_2} \times \frac{T_2}{T_1}$$

We arrive at the answer by inserting the values for the initial volume and the initial and final pressures and temperatures and remembering that the temperature units must be Kelvin.

$$V_2 = 0.227 \text{ L} \times \frac{716 \text{ torr}}{734 \text{ torr}} \times \frac{(16^\circ\text{C} + 273.15) \text{ K}}{(31^\circ\text{C} + 273.15) \text{ K}} = 0.211 \text{ L}$$

9.8 Standard Temperature and Pressure

In real-world analytical situations, pressure and temperature conditions can vary widely. Air samples for pollutant analysis, for example, can be taken on cold days or hot days, on a high mountain where the air is thin (i.e., lower pressure) or at sea level. The volume of the air sample depends on the pressure and temperature, as discovered in our discussion of Boyle's and Charles' laws. This volume, in turn, is needed in the calculation of the concentration level of the pollutant in the air, which is the critical factor in knowing whether the polluted air is hazardous to our health. The bottom line is that the volume of air that we sample needs to be corrected to standard conditions of temperature and pressure so that we may accurately compare the results of the air analysis with accepted threshold safety limits.

Another example of this problem is if the measurement of a gas volume is needed to prove a point in a research experiment. Researchers in Denver, Colorado, (which is at low pressure) working at a laboratory temperature of 25°C will report one volume for their experiment while a researcher in Los Angeles, California, (at sea level) working at a laboratory temperature of 20°C will report a different volume, simply because conditions of temperature and pressure varied. However, if their volumes are corrected to an accepted standard for temperature and pressure, their answers should be the same, all other conditions being equal.

Scientists have therefore adopted specific pressure and temperature values that represent these standard conditions. Standard pressure is 760 torr, or 1 atm. Standard temperature is 0°C, or 273.15 K. These values are referred to as **standard temperature and pressure (STP)**.

9.9 Correction of a Gas Volume to STP

The correction of a gas volume to STP conditions, as alluded to in Section 9.8, utilizes the combined gas law. The given conditions are the initial conditions (P_1 , T_1), while the STP conditions are the final conditions (P_2 , T_2).

Example 9.9

A gas occupies a volume of 27.8 L at 744 torr and 288 K. What volume will it occupy at STP?

Solution 9.9

Rearranging the combined gas law as in Example 9.8, we have the following:

$$V_2 = V_1 \times \frac{P_1}{P_2} \times \frac{T_2}{T_1}$$

STP are P_2 and T_2 .

$$V_2 = 27.8 \text{ L} \times \frac{744 \text{ torr}}{760 \text{ torr}} \times \frac{273.15 \text{ K}}{288 \text{ K}} = 25.8 \text{ L}$$

9.10 Ideal Gas Law

Earlier we discussed ways in which gas pressure can be affected in a container with a movable piston. If more gas is added to that container and no additional force is applied to the piston to keep the volume fixed (and the temperature does not change), the piston will rise (the volume will increase) to accommodate the additional gas. There is a direct relationship here between the volume of a gas and the amount of gas at constant temperature and pressure. This relationship can be expressed as follows:

$$\frac{V_1}{n_1} \times \frac{V_2}{n_2} \quad (9.10)$$

where n is the number of moles of the gas. Combining Equations 9.9 and 9.10, we get the following:

$$\frac{P_1 V_1}{n_1 T_1} = \frac{P_2 V_2}{n_2 T_2} \quad (9.11)$$

This means the following:

$$\frac{PV}{nT} = \text{Constant} \quad (9.12)$$

Pressure, volume, temperature, and the number of moles of a gas are the only parameters of a gas. If pressure, volume, temperature, or number of moles changes, one of the other parameters will also change so that the constant in Equation 9.12 does not change. For example, if pressure increases,

the volume decreases such that the constant remains the same. Thus, the constant in Equation 9.12 can be assigned a numerical value, and this value will be constant. Of course, this numerical value does depend on what units are used for the pressure and volume (the temperature units, as discussed previously, are always Kelvin). If the pressure units are atmospheres and the volume units are liters, then the value of this constant is as given below. This constant is often given the symbol R and is called the universal gas constant.

$$R = \frac{0.082057 \text{ L atm}}{\text{mol K}}, \text{ or } 0.082057 \text{ L atm mol}^{-1} \text{ K}^{-1} \quad (9.13)$$

The value of R will be as shown unless different units for pressure and volume are used. Equation 9.12 can now be written as follows:

$$\frac{PV}{nT} = R \quad (9.14)$$

Equation 9.14 can be rearranged to give the following:

$$PV = nRT \quad (9.15)$$

Equation 9.15 is known as the **ideal gas law**. Given any three of the parameters—pressure, volume, number of moles, or temperature—the fourth can be calculated using this equation. Remember that if $R = 0.082057$ is to be used, then the volume must be in liters and the pressure must be in atmospheres so that all units cancel appropriately as in Example 9.10.

Example 9.10

What volume will 0.449 mol of a gas occupy at a pressure of 1.29 atm and 293 K?

Solution 9.10

Solving Equation 9.15 for volume, we get the following:

$$V = \frac{nRV}{P}$$

After inserting the values given and the value for R , we can then calculate V as follows:

$$V = \frac{0.449 \text{ mol} \times 0.082057 \text{ L atm mol}^{-1} \text{ K}^{-1} \times 293 \text{ K}}{1.29 \text{ atm}} = 8.37 \text{ L}$$

9.11 Avogadro's Law

In all the previous gas law discussions in this chapter, no particular gas has been specified. We have always made reference to a gas or an ideal gas without saying if the gas is air, oxygen, carbon dioxide, ammonia, helium, or what. One of the tenants of the kinetic-molecular theory of gases is that all gases behave the same way in terms of compressibility, expansion, diffusion, pressure, and so on, and this ideal behavior extends to the gas laws. Thus, as long as a gas exhibits this ideal behavior,* the gas laws hold for that gas.

The ideal gas law affords us a good opportunity to summarize this in one easy statement. Using Example 9.10 as our example, we can say that 0.449 mol of helium at 1.29 atm and 293 K occupies 8.37 L. We can also say that 0.449 mol of carbon dioxide at 1.29 atm and 293 K also occupies 8.37 L. Putting it another way, a sample of helium gas at 1.29 atm and 293 K that occupies 8.37 L will consist of 0.449 mol of helium, or a sample of carbon dioxide gas at 1.29 atm and 293 K that occupies 8.37 L will consist of 0.449 mol of carbon dioxide. If we were to convert the 0.449 mol to atoms or molecules, Avogadro's number (6.022×10^{23} atoms, or molecules, per mole) would be the conversion factor regardless of the type of gas. Each gas would therefore consist of the same number of particles, whether these were atoms or molecules. Or to put this more simply, equal volumes of different gases at the same temperature and pressure consist of the same number of particles—either atoms or molecules. This is a statement of what has come to be known as **Avogadro's law**.

Avogadro's law may be expressed another way for the specific case of having exactly 1 mol of the gas held at STP. One mole of gas contains Avogadro's number of particles, so the next question is, what volume will this number of particles take up at STP? If we were to calculate this volume using the ideal gas law (see Equation 9.15), we would discover that this volume would be

* Gases deviate from ideal behavior at very high pressures and very low temperatures. Under these conditions, the gas laws would not necessarily hold.

22.4 L. In other words, 1 mol of any ideal gas occupies 22.4 L at STP. Using this statement of Avogadro's law, we can formulate an equality and two conversion factors.

$$\text{Equality: } 1 \text{ mol of gas} = 22.4 \text{ L (at STP)} \quad (9.16)$$

$$\text{Conversion factors: } \frac{22.4 \text{ L}}{1 \text{ mol}} \quad \frac{1 \text{ mol}}{22.4 \text{ L}} \quad (9.17)$$

Example 9.13

How many liters does 4.52 mol of argon gas occupy at STP?

Solution 9.13

$$4.52 \text{ mol} \times \frac{22.4 \text{ L}}{1 \text{ mol}} = 101 \text{ L}$$

We make special note of the fact that the 22.4 liters-per-mole conversion factor is good only at STP and is good only for gases. We also point out that the number 22.4 is not an exact number and will limit the number of significant figures in the answer to a calculation if it has fewer such figures than the number being converted.

9.12 Dalton's Law

We often encounter mixtures of gases in our study of chemistry and in our everyday life. The one obvious gas mixture that comes to mind is air. Air is mainly a mixture of oxygen and nitrogen. However, other gases can be present, such as carbon dioxide, since humans and animals are constantly exhaling it into their surroundings. Other examples are carbon monoxide and unburned hydrocarbons, which are components of automobile exhaust. In addition, water vapor is a component of air in virtually all locations, giving rise to what we call humidity. Gaseous mixtures exhibit all the same properties and obey all the same laws that pure gases do.

However, there is a gas law that specifically addresses gaseous mixtures. Dalton's law deals with the individual pressures that are exerted by the component gases, stating that the total pressure exerted by a gaseous mixture is

the sum of the partial pressures exerted by each of the component gases. This can be summarized as follows:

$$P_{\text{total}} = p_1 + p_2 + p_3 + p_4 \dots\dots (9.18)$$

in which P_{total} is the total pressure exerted by the mixture, and p_1 , p_2 , p_3 , and so on are the partial pressures of gases 1, 2, 3, and so on. These various gas pressures can be expressed in any pressure units. However, the total pressure and partial pressures must all have the same units.

Example 9.14

What is the partial pressure of hydrogen gas in a mixture of hydrogen, nitrogen, and ammonia gases if the total pressure exerted is 783 torr, the partial pressure of the nitrogen is 0.228 atm, and the partial pressure of the ammonia is 411 torr?

Solution 9.14

All pressures must have the same units in order to use Dalton's Law. Let us first convert the partial pressure of the nitrogen to torr.

$$0.228 \text{ atm} \times \frac{760 \text{ torr}}{1 \text{ atm}} = 173.28 \text{ torr}$$

Next, we subtract the two partial pressures given for nitrogen and ammonia from the total to determine the partial pressure of the hydrogen.

$$\begin{aligned} p_{\text{hydrogen}} &= P_{\text{total}} - p_{\text{nitrogen}} - p_{\text{ammonia}} \\ &= 783 \text{ torr} - 173.28 \text{ torr} - 411 \text{ torr} \\ &= 199 \text{ torr} \end{aligned}$$

9.13 Vapor Pressure

A useful application of Dalton's law is the case of a gas that is confined to a container in which there is also some liquid water. The water tends to evaporate

and, thus the confined gas is “contaminated” with water vapor. Perhaps the most obvious confined gas is the air that surrounds the Earth. This air is present in its container along with liquid water (e.g., oceans and lakes). The result is air that is contaminated with water vapor due to the evaporation of water from these sources, creating the condition commonly known as humid air. There is a limit to how much water vapor the air can hold before precipitation, or rain, occurs. This limit is referred to as 100% humidity, and at that point we say that the air is saturated with water vapor. Other humidity percentages, such as 45%, indicate the water vapor concentration level is at a given percentage of the saturation level (45% in this example). The actual concentration of water vapor in the air at the 100% humidity level varies with temperature.

Water—indeed, all liquids—tend to evaporate regardless of temperature. The boiling point of water does not need to be achieved for liquid water to escape the liquid phase and become a gas. We are all aware that a drop of water left on a tabletop would eventually evaporate to dryness even though we are not boiling it. Wet laundry hung on a clothes line dries because water tends to evaporate, even though it is not being boiled.

Let us examine this phenomenon in the laboratory setting. If we were to half-fill a beaker with water and set it on the laboratory bench, it would slowly evaporate, as pictured on the left of Figure 9.9. Eventually, perhaps over a period of days or weeks, the water would totally evaporate. If, however, we covered and sealed the beaker so that no water vapor could escape into the atmosphere, eventually the water vapor would saturate the air space in the beaker and create a condition of 100% humidity. At that point, the rate of evaporation equals the rate of condensation, as pictured on the right of Figure 9.9, and an equilibrium is achieved. This means that the concentration of the water vapor in the air space is constant and the liquid water would never totally evaporate.

The molecules of water in the air space exert a partial pressure like all gases in a mixture of gases. If the component of the mixture is present due to the evaporation of a liquid also present in the container, like the water vapor in this example, this partial pressure is called the vapor pressure. The formal definition of vapor pressure is the pressure exerted by the molecules of a gas in equilibrium with its liquid present in the same sealed container.

Two important points about vapor pressure are (1) all liquids have a vapor pressure and (2) vapor pressure varies with temperature. All liquids

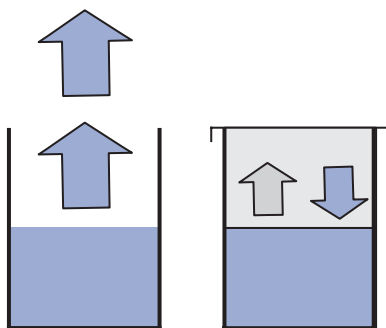


FIGURE 9.9 All the water in the beaker on the left would eventually evaporate. On the right, the water evaporates to the point of saturating the air space with water vapor and creating an equilibrium between the liquid water and the water vapor. The partial pressure of the water vapor in the air space on the right is the vapor pressure of water at the given temperature. (From Kenkel, J., Kelter, P., and Hage, D., *Chemistry: An Industry-Based Introduction with CD-ROM*, CRC Press, Boca Raton, FL, 2001. With permission.)

tend to evaporate, but some do so more than others. For example, liquid ether has a very strong tendency to evaporate, much more so than water, because ether has a very high vapor pressure compared with water. The 100% humidity level for ether would be a much higher concentration of ether vapor than water at its 100% humidity level. The concentration of the vapor, regardless of what the liquid is, depends on temperature. As temperature increases, the tendency to evaporate increases and, thus, more vapor from the liquid will result. The normal boiling point of a liquid is the temperature at which the vapor pressure equals standard atmospheric pressure (i.e., 760 torr). The dependence of the vapor pressure of water on temperature is obvious from Table 9.2. This table covers a limited range of temperatures, from 20° to 30°C.

The vapor pressure of water is another example of a property of water that is influenced by hydrogen bonding (Chapter 6). The hydrogen bonds that bind water molecules to each other are relatively strong, so water molecules require more energy than most other liquids to break free of each other and enter the gas phase. Thus the vapor pressure of water is typically much lower than other liquids at equivalent temperatures, reflecting the fact that there are fewer water molecules in the gas phase to be exerting pressure.

Table 9.2 Vapor Pressures of Water at Various Temperatures

Temperature (°C)	Vapor Pressure (torr)
20	17.5
21	18.5
22	19.8
23	21.1
24	22.4
25	23.8
26	25.2
27	26.7
28	28.3
29	30.0
30	31.8

9.14 Homework Exercises

Properties and Kinetic-Molecular Theory of Gases

1. List at least five properties that are characteristic of gases.
2. What is the kinetic-molecular theory of gases? What are the key points of this theory?
3. How does the kinetic-molecular theory of gases help explain each of the gas properties that you listed in question 1?
4. When you drive your car on the highway, the tires get hot. Explain, by the kinetic-molecular theory of gases, why the air pressure inside the tires increases.
5. If a container of ether is opened in one part of a lab, its odor will soon be noticed by people in another part of the lab. Explain this using the kinetic-molecular theory of gases.
6. When you pump air into a bicycle tire, the air pressure inside the tire increases. Explain by the kinetic-molecular theory of gases (a) how it is that you are able to add more air to a tire that seems to be filled with air and (b) why the pressure inside the tire increases.

Pressure Units

7. Fill in the blanks in the following table.

	torr	atm	psi
(a)	773		
(b)		1.429	
(c)			32
(d)	702		
(e)		0.930	
(f)			11.6
(g)	849.2		
(h)		0.7118	
(i)			18.00

8. What is 732 torr expressed (a) in psi and (b) in atmospheres?
9. The pressure of a gas sample held in a closed container is 1.29 atm. What is this pressure in torr? What is this pressure in psi?
10. What is 1.134 atm expressed in torr? What is it in psi?
11. The pressure of a sample of a gas held in a closed container is 743 torr. What is this pressure in atmospheres and in psi?

Boyle's Law

12. What is Boyle's law? What is allowed to vary and what is assumed to be constant in this law?
13. Fill in the blanks in the following table. Assume constant temperature in each case.

	Initial Pressure	Final Pressure	Initial Volume	Final Volume
(a)	728 torr	738 torr	6.38 mL	
(b)	1.192 atm	1.338 atm		16.0 mL
(c)	769.2 torr		1.66 mL	0.00180 L
(d)		1.034 atm	52.0 mL	128.3 mL
(e)	764 torr	0.961 atm	14.5 L	
(f)	714 torr	0.983 atm		1.20×10^4 mL
(g)	0.459 torr		1.28 L	0.271 L
(h)		782.4 torr	30.92 L	3.63×10^5 mL

14. The volume of a sample of a gas is 89.3 mL at 723 torr. What is the volume of this gas if the pressure is increased to 782 torr at constant temperature?
15. A sample of argon gas occupies 34.7 mL at 712 torr and 42°C. What volume will this gas occupy at the same temperature, but at 745 torr?
16. A sample of carbon dioxide occupies a volume of 55.2 mL at 777 torr. What will the pressure be if the volume is increased to 59.8 mL at constant temperature?
17. A sample of helium that occupied an unknown volume at 0.838 atm and 34°C occupies a volume of 9.0 L at 0.940 atm at the same temperature. What was the initial volume?
18. By how much must the pressure on a sample of air occupying 0.738 L at 1.20 atm and 290.3 K increase for it to occupy a volume of 0.700 L at the same temperature?

Charles' Law

19. What is Charles' law? What is allowed to vary and what is assumed to be constant in this law?
20. Fill in the blanks in the following table. Assume constant pressure in each case.

	Initial Temperature	Final Temperature	Initial Volume	Final Volume
(a)	284 K	294 K	3.0 L	
(b)	34°C	278 K		67.2 mL
(c)	78°C		5.72 L	4.220×10^3 mL
(d)		21.74°C	0.0338 L	17.9 mL
(e)	79.2°C	249.6 K	928.3 mL	
(f)	25.11°C	94.4°C		8.229 L
(g)	289.1 K		65.2 L	7.24×10^5 mL
(h)		89.24°C	9.20 mL	0.0227 L

21. If the temperature of 234.3 mL of a gas is 56.6°C, what volume will this gas occupy at 78.4°C if the pressure does not change?
22. The volume of a sample of a gas is 8.39 L at 33°C. What would be the volume if the temperature is changed to 265 K, while the pressure is constant?

23. If a gas occupies 6.50 liters at 2°C, to what temperature, in degrees Celcius, must this gas be raised in order for it to occupy 7.00 L if the pressure stays the same?
24. Consider again the gas in question 23. Suppose you want the gas to occupy 6.00 liters at the same pressure. What would the temperature need change to in order to accomplish that?
25. What increase in temperature would there need to be to change a gas volume from 34.00 mL to 35.00 mL if the initial temperature is 56°C?

Combined Gas Law

26. What is the combined gas law? What is allowed to vary and what is assumed to be constant in this law?
27. Fill in the blanks in the table below.

	Initial Pressure	Final Pressure	Initial Temperature	Final Temperature	Initial Volume	Final Volume
(a)	721 torr	741 torr	281 K	285 K	9.38 mL	
(b)	689 torr	0.793 atm	78°C		0.371 L	0.228 L
(c)	903 torr		398 K	54.3°C	0.515 L	5.3×10^2 mL
(d)	1.09 atm	0.938 atm	89.4°C	323 K		0.229 mL
(e)	1.172 atm	765 torr		34.2°C	33.4 mL	0.038 L
(f)		7.60×10^2 torr	10.0°C	273.15 K	12.0 L	16.6 L
(g)	0.838 atm	658 torr	72°C		0.4403 L	1.2295 L
(h)	0.925 atm	0.958 atm	288.0 K	298.0 K	6.3 L	
(i)	759 torr		254 K	32.9°C	2.91 mL	3.58 mL

28. What will be the volume of a gas at 14°C and 0.892 atm if it has a volume of 3.29 L at 21°C and 0.836 atm?
29. At 773 torr and 23.5°C, the volume of a gas is 0.529 L. What volume will it have if the pressure is decreased to 722 torr and the temperature is increased to 30.4°C?
30. What pressure is required to keep a gas from exceeding 4.5 L in volume if it occupied 4.0 L at 789 torr and the temperature is increased from 34°C to 44°C?
31. What temperature is required to keep a gas from dropping to less than 347 L in volume if it occupied 388 L at 24°C and the pressure is decreased from 751 torr to 738 torr.

32. Suppose the ratio of the initial temperature of a sample of gas in Kelvin to the final temperature in Kelvin is 1.273. If the gas occupies 0.0038 L at the initial temperature and 723 torr, and 0.0040 L at the final temperature, what must the final pressure be?

Correction to STP

33. What is meant by the term *standard temperature and pressure* (STP)? What are the experimental conditions that are being used when you are working at STP?
34. A chemist has performed a chemical reaction and has captured a gas that forms, noticing that 68.2 mL of the gas formed at temperature of 25°C and a pressure of 753 torr. He wishes to publish his results in a chemistry journal but realizes that he must express this volume at STP. What is the volume at STP?
35. If a gas occupies 0.334 L at 563 torr and 94.2°C, what volume will it occupy at STP?
36. If a gas occupies 89.22 mL at STP, what volume will it occupy at 785 torr and 23°C?
37. If a gas occupies 0.928 L at STP, what temperature would be required to sustain this volume when the pressure changes to 0.800 atm?
38. If a gas occupies 2.4×10^6 mL at STP, what pressure would be required to sustain this volume when the temperature changes to 215 K?

Ideal Gas Law

39. What is the ideal gas law?
40. What is the value of the R term in the ideal gas law?
41. Complete the following table.

	Pressure	Volume	Number of Moles	Temperature
(a)	1.034 atm	7.82 L	14.9 mol	
(b)	778.3 torr	0.228 L		287.5 K
(c)	693.0 torr		0.993 mol	67.2°C
(d)		46.11 mL	2.339 mol	259.3 K
(e)	0.883 atm	3.42×10^4 mL	0.668 mol	
(f)	782.3 torr	392.3 mL		34.10°C
(g)	809.12 torr		89.39 mol	293.15 K
(h)		45.1 L	6.220 mol	32.4°C

42. How many moles of gas are in a 56.3-L container at 45.1°C and 0.933 atm?
43. What volume does 0.400 mol of oxygen gas occupy at 0.933 atm and 321 K?
44. Suppose that a chemist performs a chemical reaction in which a gas is formed and finds 65.9 mL of the gas formed at 276 K and 0.987 atm. How many moles of the gas has this chemist produced?
45. What pressure, in torr, is exerted by 0.149 mol of a gas if this amount of gas occupies 93 mL at 261 K?
46. What is the temperature of a gas, in degrees Celcius, if 51.0 mol of the gas are occupying 846 L at a pressure of 0.850 atm.?

Avogadro's Law

47. What is Avogadro's law?
48. Complete the following table, assuming a gas at STP conditions.

	Volume	Moles
(a)	83.0 L	
(b)		0.0473
(c)	5.49×10^5 mL	
(d)		7.83×10^5

49. How many moles of CO₂ gas do you have if you have 0.733 L of CO₂ at STP?
50. How many liters does 0.882 mol of methane occupy at STP?
51. Is it possible to calculate, using the ideal gas law, the number of liters that 1 mol of a gas occupies at STP to five significant figures? If so, do it.

Vapor Pressure and Dalton's Law of Partial Pressure

52. What is Dalton's law?
53. Suppose that a new chemistry technician in your lab is confused and wants to know exactly what vapor pressure is. What would you tell her?
54. What is meant by the normal boiling point for a compound?

55. How does the vapor pressure of a liquid usually change with temperature?
56. A gas is collected over water at 29°C and 745 torr. The volume collected is 873.2 mL. What volume would the dry gas occupy at STP? (*Note:* The vapor pressure of water at 29°C is 30.0 torr.)
57. The vapor pressure of water at 19°C is 16.5 torr. If a sample of oxygen measuring 2.31 L is collected over water at 19°C and 729 torr, what volume will the oxygen occupy if it were dry and held at standard conditions of temperature and pressure?
58. A gas is collected over water at 24°C and 715 torr. If the volume of this “wet” gas is 9.88 L, what volume will the dry gas occupy at STP? (*Note:* The vapor pressure of water at 24°C is 22.4 torr.)

Chapter 10

Solutions

Introduction	248
Terminology	248
Expressing Concentration	249
Preparing Solutions	252
More Terminology	260
Increasing Dissolution Rate	263
Mechanism Of Dissolving	264
Homework Exercises	270

10.1 Introduction

An extraordinary amount of chemistry in our world occurs in water solution. A large percentage of the Earth's surface is covered by oceans, which are actually large bodies of water solutions. We are all very familiar with rain, which is a water solution. We are all familiar with lakes and rivers. These, too, are water solutions. All of us use water from wells and municipal water supplies for drinking, cooking, cleaning, and various other household chores. This water is also a water solution. Manufacturing industries use large amounts of water in their many processes for a variety of purposes, creating and using water solutions. Household waste and industrial waste are usually carried away by water, thus creating a water solution that can be hazardous and must be properly treated so that we can all live safe and healthy lives. Chemists and laboratory technicians use water freely in their laboratory work, sometimes isolating and analyzing for chemicals dissolved in water while at other times creating water solutions in order to facilitate their work.

In addition to water solutions, other liquid chemicals can be the basis for solution formation and use. Examples include crude oil (the source of organic chemicals), manufactured oils, many petroleum distillates, and non-water solvents used for cleaning. Gasoline, kerosene, and other fuels are solutions. In addition, organic solvents, such as acetone and methyl alcohol, are used as solvents in industrial laboratories. It follows that students of chemistry should devote much attention to this state of matter known as a solution.

10.2 Terminology

A **solution** is a homogeneous mixture, a physical combination of two or more pure substances. These substances are mixed as well as they can possibly be mixed so that we cannot tell, even with the most powerful microscope, that it is actually more than one substance. In order to tell that such a mixture is indeed a mixture and not a pure substance, one would need to be able to view the actual molecules, atoms, or ions of the pure substances involved and, in doing so, notice that there is more than one chemical species present.

Solutions, as defined above, do not have to be liquid mixtures. There are gaseous mixtures that are homogeneous and there are solid mixtures

that are homogeneous. Any homogeneous physical mixture of two or more gases is a solution. Any homogeneous physical mixture of two or more solids is a solution. An example of a gaseous solution is air, although we would need to specify a particular altitude for the composition of this solution. Air gets thinner as one climbs a mountain, for example. We would also need to specify a particular locale. The composition of the air in a room full of people exhaling carbon dioxide is different from the air in an adjoining room where there are no people. An example of a solid solution is a stainless steel alloy. Stainless steel is a homogeneous mixture of chromium, nickel, iron, and carbon. All alloys (e.g., brass, bronze) are also solid solutions.

The components of a solution are known as the solvent and the solute or solutes. There is only one solvent, but there can be one or more solutes. The **solvent** is the component that is present in the greatest amount or the component that has dissolved the other component or components in itself. The **solutes** are the components present in the lesser amount or the substances that have dissolved in the solvent. For example, if you have a large bowl of pure, distilled water and you add a few crystals of salt and a few crystals of sugar and stirred this with a spoon, you would have a solution with water as the solvent and salt and sugar as the solutes. But remember, there can be just one solute, as well as many solutes.

10.3 Expressing Concentration

The concentration of a solute in a solution (i.e., the amount of a solute dissolved in a particular amount of solvent or in a particular amount of solution) is a vitally important concept in a variety of situations in real-world chemistry. For example, nitrate in drinking water is unsafe for infants. The amount of nitrate dissolved in a given amount of drinking water, or the concentration of nitrate in drinking water, is important to know in order to determine whether that water is safe for infants. Thousands of chemists and chemistry technicians in the United States analyze solutions such as drinking water daily to determine the quantity of a solute dissolved per unit volume of solution. This concentration level of a solute in a solution is thus a very important topic for students of chemistry. In this section, various methods of expressing concentration are defined and calculations associated with the determination of solution concentration are studied.

10.3.1 Percent

The concept of percent is well known. The basic definition was reviewed in Chapter 7 and is illustrated again in Equations 10.1 through 10.3

$$\text{Percent (\%)} = \frac{\text{Part}}{\text{Whole}} \times 100 \quad (10.1)$$

For solutions, however, this concept can be confusing because the units used for the part and the whole are variable. Not only is it common for a solution concentration to be expressed as a volume percent (v/v%) when the solute is a liquid, but it is also common for the numerator units to be weight units and the denominator units to be volume units (so-called weight/volume percent [w/v%]).

$$\text{Volume percent (v/v\%)} = \frac{\text{Volume of solute used}}{\text{Total volume of solution prepared}} \times 100 \quad (10.2)$$

$$\begin{aligned} \text{Weight/volume percent (w/v\%)} \\ = \frac{\text{Weight of solute used}}{\text{Total volume of solution prepared}} \times 100 \end{aligned} \quad (10.3)$$

Note that the denominator in each case is the quantity of *solution* and not solvent. Also, the units in Equation 10.2 can be any unit of volume as long as they are the same in both the numerator and the denominator. The units in Equation 10.3 must be grams and milliliters, kilograms and liters, or milligrams and microliters.

Example 10.1

If 6.3 mL of acetone are dissolved in a water solution such that the total volume of solution is 500.0 mL, what is the concentration of the acetone expressed as a volume percent?

Solution 10.1

Equation 10.2, the definition of volume percent concentration, is used.

$$\text{Volume percent (v/v\%)} = \frac{\text{Volume of solute used}}{\text{Total volume of solution prepared}} \times 100$$

$$\text{Volume percent (v/v\%)} = \frac{6.3 \text{ mL}}{500.0 \text{ mL}} \times 100 = 1.3\%$$

Example 10.2

What is the weight/volume percent concentration of a solution if 2.48 g of solute are dissolved in 600.0 mL of solution?

Solution 10.2

Equation 10.3 defines weight/volume percent.

$$\text{Weight/Volume percent (w/v\%)} = \frac{\text{Weight of solute used}}{\text{Total volume of solution prepared}} \times 100$$

$$\text{Weight/Volume percent (w/v\%)} = \frac{2.48 \text{ g}}{600.0 \text{ mL}} \times 100 = 0.413\%$$

10.3.2 Molarity

Molarity is defined as the number of moles of solute dissolved per liter of solution and may be calculated by dividing the number of moles dissolved by the number of liters of solution.

$$\text{Molarity} = \frac{\text{Moles of solute used}}{\text{Total liters of solution prepared}} \quad (10.4)$$

So the units are moles per liter (mol/L). Solutions are referred to as being, for example, 2.0 molar, or 2.0 M. The M refers to molar, and the solution is said to have a molarity of 2.0, or 2.0 moles dissolved per liter of solution. It is important to recognize that it is the number of moles dissolved per liter of *solution* and not per liter of solvent.

Example 10.3

What is the molarity of a solution that has 4.5 mol of solute dissolved in 300.0 mL of solution?

Solution 10.3

Equation 10.4 defines molarity. The number of milliliters given must first be converted to liters, and so 0.3000 L is used in the denominator.

$$\text{Molarity} = \frac{\text{Moles of solute used}}{\text{Total liters of solution prepared}}$$

$$\text{Molarity} = \frac{4.5 \text{ mol}}{0.3000 \text{ L}} = 15 \text{ M}$$

Example 10.4

What is the molarity of a solution of NaOH that has 0.491 g dissolved in 400.0 mL of solution?

Solution 10.4

In order to use Equation 10.4, the grams of solute must be converted to moles by dividing the grams by the formula weight (FW). Once again, milliliters are converted to liters.

$$\text{Molarity} = \frac{\text{Moles of solute used}}{\text{Total liters of solution prepared}}$$

$$\text{Molarity} = \frac{\text{Grams of solute} \div \text{FW}}{\text{Total liters of solution prepared}}$$

$$\text{Molarity} = \frac{(0.491/39.997) \text{ mol}}{0.4000 \text{ L}} = 0.0307 \text{ M}$$

10.4 Preparing Solutions

A very important activity in real-world chemistry laboratories is the preparation of solutions. A variety of reasons exist to prepare a given volume of solution of a given solute at a given concentration. The given volume unit is almost always liters or milliliters, and the given concentration unit can be those defined in Section 10.3, as well as other units. We now proceed to review the specific processes and calculations involved in this activity for the percent and molarity units.

10.4.1 Dilution

One method of preparing solutions is by dilution. In other words, the solute is already in a solution but at a higher concentration than what is desired, and this more-concentrated solution must be diluted—or more solvent must be added to it—to get the desired concentration of solute. An important example of this is when making solutions of acids from the concentrated

acid solutions, such as 12 M HCl and 18 M H₂SO₄, sold by chemical supply companies. Adding a quantity of the solvent to a particular volume of the more-concentrated solution decreases the concentration to the desired value. The questions to answer here are, How much of the more-concentrated solution is to be used, and how much solvent is to be added? The volume of the more-concentrated solution to be used must be calculated, and the amount of solvent is the volume required to bring the solution volume to the desired total volume. Let us now derive a dilution equation that can be used for this calculation. We will begin by assuming that molarity is the concentration unit to be used.

When diluting a solution of a given molarity to one of lesser molarity, solvent is added. The number of moles of the solute does not change. The total number of moles of solute is the same regardless of the amount of solvent present.

$$\text{Moles of solute before dilution} = \text{Moles of solute after dilution} \quad (10.5)$$

Or using B to mean before dilution and A to mean after dilution, we have the following true statement:

$$\text{Moles}_B = \text{Moles}_A \quad (10.6)$$

The number of moles of a solute can be calculated by multiplying the molarity (moles per liter) by the volume of solution in liters.

$$\text{Moles} = \frac{\text{Moles}}{\text{Liters}} \times \cancel{\text{Liters}} \quad (10.7)$$

Thus, Equation 10.6 is transformed as follows:

$$\left(\frac{\text{Moles}}{\text{Liters}} \right)_B \times \cancel{\text{Liters}}_B = \left(\frac{\text{Moles}}{\text{Liters}} \right)_A \times \cancel{\text{Liters}}_A \quad (10.8)$$

and this translates to the following:

$$\text{Molarity}_B \times \text{Liters}_B = \text{Molarity}_A \times \text{Liters}_A \quad (10.9)$$

While it appears that the liters unit is required for proper dimensional analysis and cancellation with the liters unit in the denominator of molarity, the

milliliter unit can also be used. This is acceptable if we include the conversion factor required to convert milliliters to liters (1000 mL per 1 L).

$$\cancel{\text{mL}} \times \frac{1 \text{ L}}{1000 \cancel{\text{ mL}}} = \text{L} \quad (10.10)$$

This is the same on both sides of Equation 10.9 and thus cancels as follows:

$$M_B \times \text{mL}_B \times \frac{\cancel{1 \text{ L}}}{1000 \cancel{\text{ mL}}} = M_A \times \text{mL}_A \times \frac{\cancel{1 \text{ L}}}{1000 \cancel{\text{ mL}}} \quad (10.11)$$

$$\text{Molarity}_B \times \text{Milliliters}_B = \text{Molarity}_A \times \text{Milliliters}_A \quad (10.12)$$

The same can be said about the molarities. Any concentration unit can be used because the conversion factor to convert to molarity is the same on both sides and also cancels. Thus we can have a more general formula in which the units of concentration and volume are not specified.

$$C_B \times V_B = C_A \times V_A \quad (10.13)$$

We will refer to Equation 10.13 in future discussions as the dilution equation. In this equation, C stands for concentration and V stands for volume. Any concentration unit (e.g., molar, percent) and any volume unit (e.g., milliliter, liter) can be used, as long as we recognize that whatever unit is used, it is the same on both sides of the equation. Thus, if the volume after dilution is expressed in milliliters, then the volume before dilution will also be expressed in milliliters. If the unit of the given concentration after dilution is different from that of the given concentration before dilution, one of the two units must be converted to the other before plugging it into the equation. The form of the dilution equation needed to calculate the volume of the more-concentrated solution required to make the final solution is shown here.

$$V_B = \frac{C_A \times V_A}{C_B} \quad (10.14)$$

Example 10.5

How many milliliters of concentrated HCl (12 M) are needed to prepare 250.0 mL of a 0.15 M solution of HCl?

Solution 10.5

Using Equation 10.14, we have the following:

$$V_B = \frac{C_A \times V_A}{C_B} = \frac{0.15 \text{ M} \times 250.0 \text{ mL}}{12 \text{ M}} = 3.1 \text{ mL}$$

Note: When diluting solutions of concentrated acids, a practical rule to follow is the “add acid to water rule,” which means that some water should be placed in the container before adding the acid as a safety precaution.

The process by which solutions are prepared by dilution is important to consider. The volume of the more-concentrated solution, V_B , as calculated by the dilution equation, such as in Example 10.5, must be measured via a graduated cylinder, pipet, or similar volume transfer device and placed in a container with a graduation line corresponding to the volume of the more-diluted solution being prepared, V_A . The choice of transfer device and container depends on the precision that is required. Some pipets, such as the volumetric pipet, are high-precision transfer devices and are used for analytical work that requires more significant figures than a graduated cylinder can provide. Similarly, volumetric flasks are high-precision solution containers, while bottles, beakers, and Erlenmeyer flasks are less precise.

Having placed the more-concentrated solution in the chosen container, the solvent, usually water, is added in such a quantity that the bottom of the meniscus (the curved surface of the solvent) is resting on the calibration line. To properly judge its height of the meniscus, the meniscus must be at eye level. Once the appropriate amount of solvent is added, the solution is shaken, stirred, or swirled so as to make the solution homogeneous. Figure 10.1 illustrates this process using an Erlenmeyer flask as the container.

Example 10.6

How would you prepare 100 mL of a 5.0-ppm (parts per million) copper solution from a 1000-ppm copper solution?

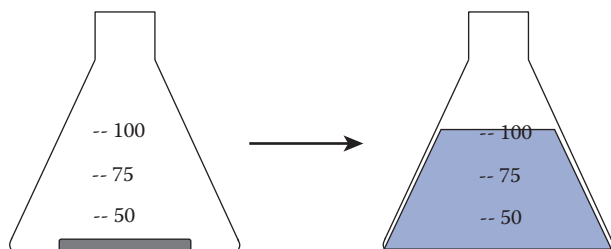


FIGURE 10.1 An illustration of the preparation of 100 mL of a solution by dilution. The flask at left has the measured volume of the undiluted solution (V_b) inside. The flask on the right contains the volume of diluted solution after swirling.

Solution 10.6

Using Equation 10.14, we can calculate how many milliliters of the 1000-ppm solution are required.

$$V_B = \frac{C_A \times V_A}{C_B} = \frac{5.0 \text{ ppm} \times 100.0 \text{ mL}}{1000 \text{ ppm}} = 0.50 \text{ mL}$$

To answer the how-would-you-prepare question, you would measure 0.50 mL of the 1000-ppm solution, place this volume in a 100-mL flask, dilute with water to 100 mL, and shake.

If high precision is important, a volumetric pipet or other high-precision device must be used for the transfer of V_B to the container and a volumetric flask must be used for the container. Also, if too much water is added by mistake, such that the solution level is above the required calibration line, the solution must be discarded and the entire process repeated. Removing solution from the container to correct such a situation is not an acceptable method for accurately preparing solutions. The reason for this is that the solution being removed likely contains some solute, and thus the concentration of the solute would change in an indeterminable way if some were removed.

10.4.2 Liquid Solute and Volume Percent

Expressing concentration as a volume percent is appropriate and common when the solute is a liquid, because (unlike the volume of a solid) the volume of a liquid is easy to measure, whether it be with a graduated cylinder, pipet, or graduated container. The calculation and procedure can actually be thought of as

identical to that for dilution. The concentration before dilution, C_B in Equations 10.13 and 10.14, is 100%. However, it is also similar to the weight percent calculation and procedure discussed earlier. The following formula is useful:

$$\text{Milliliters of solute} = (\text{Percent desired} \div 100) \times \text{Total mL desired} \quad (10.15)$$

The solute is measured (or pipetted if high precision is important) into a graduated container (or a volumetric flask if high precision is important), and the solvent is added so that the total volume is the volume of solution that is desired. Notice that the identity of the solute does not enter into the calculation or process.

Example 10.7

How would you prepare 500.0 mL of a 15% (v/v) ethanol solution in water?

Solution 10.7

$$\text{Volume to measure} = (15\% \div 100) \times 500.0 \text{ mL} = 75 \text{ mL}$$

Thus, 75 mL of ethanol are measured into the vessel to contain the solution. Water is then added to the 500.0-mL level, and the solution is stirred, shaken, or swirled.

10.4.3 Solid Solute and Weight-to-Volume Percent

The weights of solid solutes are more conveniently measured than their volumes, and the volumes of liquid solutions are more conveniently measured than their weights. These two facts likely gave rise to the method of weight/volume percent that is sometimes used for expressing solution concentration. The amount of solute to be weighed is a percentage of the total solution volume rather than the total solution weight.

$$\text{Grams of solute} = (\text{Percent desired} \div 100) \times \text{Solution volume desired} \quad (10.16)$$

Thus the solute can be weighed and placed in the container and the solvent added up to the total solution volume desired. The volume unit is milliliters, if grams is being calculated.

Example 10.8

How would you prepare 500.0 mL of a 5.0% (w/v) solution of NaCl?

Solution 10.8

$$\text{Weight to measure} = (5.0\% \div 100) \times 500.0 \text{ mL} = 25 \text{ g}$$

Thus, 25 g of NaCl are weighed, placed in the container, and dissolved. More water is then added such that the total solution volume is 500.0 mL, and the solution is shaken to make it homogeneous.

10.4.4 Solid Solute and Molarity

To prepare a given volume of a solution of a given molarity of solute when the weight of the pure, solid solute is to be measured, it is necessary to calculate the grams of solute that are required. This number of moles can be calculated from the desired molarity and the desired volume of solution.

$$\text{Moles} = \cancel{\text{Liters}} \times \frac{\text{Moles}}{\cancel{\text{Liters}}} \quad (10.17)$$

Grams can be calculated by multiplying moles by the formula weight, as we saw often in Chapters 7 and 8.

$$\text{Grams} = \cancel{\text{Moles}} \times \frac{\text{Grams}}{\cancel{\text{Mole}}} \quad (10.18)$$

Combining Equations 10.17 and 10.18 gives us the formula needed.

$$\text{Grams} = \cancel{\text{Liters}} \times \frac{\text{Moles}}{\cancel{\text{Liters}}} \times \frac{\text{Grams}}{\cancel{\text{Mole}}} \quad (10.19)$$

The liters in Equation 10.19 are the liters of solution that are desired; the moles per liter is the molarity desired; and the grams per mole is the formula weight of the solute. Therefore we have the following formula:

$$\text{Grams to weigh} = L_D \times M_D \times FW_{\text{SOL}} \quad (10.20)$$

in which L_D refers to the liters that are desired, M_D to desired molarity, and FW_{SOL} to the formula weight of solute. The grams of solute thus calculated is weighed out, placed in the container, and water is added to dissolve and dilute the solution to volume.

Example 10.9

How would you prepare 500.0 mL of a 0.20 M solution of NaOH from pure, solid NaOH?

Solution 10.9

Using Equation 10.20, we have the following:

$$\text{Grams to weigh} = L_D \times M_D \times FW_{SOL}$$

$$\text{Grams to weigh} = 0.5000 \times 0.20 \times 40.00 = 4.0 \text{ g}$$

Weigh 4.0 grams of NaOH, place this in a container with a 500-mL calibration line, add water to dissolve the solid, and then dilute to volume and shake, stir, or swirl.

If the solute is a liquid (somewhat rare), the weight calculated from Equation 10.20 can, but rather inconveniently, be measured on a balance. However, this weight can also be converted to milliliters by using the density of the liquid. In this way, the volume of the liquid, rather than its weight, can be measured, and it can be pipetted into the container.

10.4.5 Summary of Calculations

The various equations derived in this section for use in solving solution preparation problems are summarized in Table 10.1.

Table 10.1 Four Solution Preparation Possibilities and Their Relevant Equations Used for Their Calculation

Solution Preparation	Equation Number
Dilution	10.14
Liquid solute and volume percent	10.15
Solid solute and weight to volume percent	10.16
Solid solute and molarity	10.20

10.5 More Terminology

Sugar and salt are examples of compounds that dissolve in water quite readily at room temperature. There are, however, many substances that do not readily dissolve in water. Examples are sand and sawdust. There is some dependence on temperature for the ability to dissolve. Substances that dissolve readily at room temperature in a solvent are said to be **soluble** in that solvent. Substances that do not dissolve readily in a solvent are said to be **insoluble** in that solvent. Soluble and insoluble are qualitative descriptors of the behavior of compounds in the presence of a solvent at a given temperature; a compound either does dissolve or does not dissolve.

While a given solute may appear to dissolve in a given solvent and another may not, solubility is actually a relative term that depends on both quantity and temperature. For example, even though sugar is soluble in water, it is not possible to dissolve a 5-lb bag of sugar in a cup of water. All substances have an upper limit as to how much will dissolve in a given amount of solvent at a given temperature. By the same token, substances that appear to be insoluble do dissolve to a certain extent, however small that may be. In other words, all substances have a certain maximum *quantitative* solubility, such as a certain number of grams of solute per 100 mL of solvent at a given temperature, or a certain number of moles of solute per liter of solution at a given temperature. This quantitative **solubility** is defined as the maximum amount of a solute that will dissolve in a given amount of solvent at a given temperature. Table 10.2 presents the solubilities of some chemicals in water at 25°C.

Table 10.2 Common Compounds (Solids) and Their Solubilities in Water

Chemical	Solubility ^a
NaCl	35.7
Sucrose (sugar)	200
$\text{NaC}_2\text{H}_3\text{O}_2 \cdot 3\text{H}_2\text{O}$	76.2
BaSO_4	0.000222
Ca(OH)_2	0.185
^a Expressed in g/100 mL at 25°C.	

It is important to note the solubility behavior of chemicals that are gases (e.g., oxygen, nitrogen, carbon dioxide, hydrogen chloride, and ammonia). Even though such chemicals are gases, they can be quite soluble in liquid solvents. It is common practice, for example, to measure the dissolved oxygen in treated and untreated wastewater at wastewater treatment plants. The fuming properties of concentrated hydrochloric acid, or hydrogen chloride (HCl), and of concentrated ammonia (NH₃) solutions are indicative that gases can sometimes be dissolved to a very great extent. Such solubility depends on temperature, however, and the effect of temperature is the opposite of what we usually find for solids and liquids. As temperature increases, the solubility of gases decreases because there is a greater tendency for dissolved gases to escape the liquid phase and be in the gas phase.

A solution that has the maximum amount of solute represented by the solubility dissolved in it is said to be **saturated** at that temperature. The solution is then saturated with the solute so that no more can be dissolved. Such a solution may be characterized by an amount of solute in the container that is undissolved and remains undissolved indefinitely even with continuous shaking or stirring. A solution that does not have this maximum amount dissolved in it at the temperature specified by the solubility data is said to be **unsaturated**. Such a solution could be made saturated by simply dissolving more solute until the maximum amount is reached.

A solution may also be supersaturated. A **supersaturated** solution is one that has more than the maximum amount of solute specified by the solubility dissolved in it at the indicated temperature. It may seem unlikely that any solution would have more solute dissolved than is possible to be dissolved at a given temperature. Yet it is possible. The question is, how do these solutions achieve the supersaturated state? The usual technique involves temperature.

In this discussion, we have stated repeatedly that temperature is a factor. The solubility of a solute in a given solvent does indeed depend on temperature. *Usually*, as the temperature increases, the solubility increases. It is usually possible to dissolve more solute if the temperature is increased. The exact behavior of solubility versus temperature depends on what the solute and the solvent are. For example, the solubility of sodium acetate in water increases markedly as temperature is increased. However, the solubility of sodium chloride in water increases only slightly, while the solubility of sodium sulfate in water actually decreases with temperature. The sodium acetate-water behavior provides a good example of the process by which a

solution may become supersaturated. If a small amount of water (approximately 5 mL) is added to a relatively large amount of hydrated sodium acetate crystals (approximately 50 g of $\text{NaC}_2\text{H}_3\text{O}_2 \cdot 3\text{H}_2\text{O}$) in a beaker or flask, there will be very little evidence of dissolving at 25°C. However, if the container is placed on a hot plate and the temperature raised to near the boiling point of water, all the sodium acetate will dissolve. If this is followed by a cooling of the container and its contents, unless the sodium acetate crystals have a site to form and grow (something absent in a glass container with a smooth interior surface, like a beaker or flask), the solid will stay dissolved and the resultant solution is supersaturated. Supersaturation can be easily proved by providing the required surface on which the crystals can form and grow. An excellent surface is a seed crystal, a tiny solid crystal of the dissolved compound. In our example, if a tiny crystal of solid sodium acetate is added to the solution, the solution will turn completely to solid in a matter of seconds and the large amount of additional solute dissolved at the elevated temperature will precipitate or crystallize.

The terms miscible and immiscible describe the solubility behavior of two liquid substances in each other. **Miscible** describes two liquid substances that completely mix with each other in all proportions. Examples are ethyl alcohol and water, acetone and water, and toluene and benzene. **Immiscible** describes two liquids that apparently do not mix when placed in the same container, even after vigorous shaking or stirring. Examples are oil and water, carbon tetrachloride and water, heptane and water, and gasoline and water. In these cases, two distinct liquid layers form in the container, the heavier (denser) liquid sinking to the bottom and the lighter (less dense) liquid floating on top. As with solids that are insoluble in water, however, there is likely to be at least a small amount of mixing that may not be detectable by simply looking at the mixture. Thus it is also appropriate to quantify this amount by referring to one particular liquid's quantitative solubility in another if the two are obviously not miscible. For example, Table 10.3 lists the solubilities of a number of alcohols in water at 25°C.

Finally, the terms **dilute** and **concentrated** are frequently used to (rather loosely) describe the relative amount of a solute that is dissolved in a solvent. Concentrated describes a solution in which there is a relatively large amount of solute dissolved, while dilute refers to a lesser amount. These terms are used when describing solutions of common laboratory acids, such as sulfuric acid and hydrochloric acid. These acids are commonly purchased from a

Table 10.3 Some Alcohols (Liquids) and Their Solubilities in Water

Alcohol	Solubility ^a
1-Butanol	7.4
1-Pentanol	2.7
1-Hexanol	0.60
1 Heptanol	0.17
1-Octanol	0.030
^a Expressed in g/100 mL at 25°C.	

chemical supply company in the concentrated form and, as such, are known as concentrated sulfuric acid or concentrated hydrochloric acid. These concentrated solutions are nearly saturated with the solute. However, after a chemist mixes any of these with water, even if it dilutes the acid only slightly, such as in a 50–50 mixture, the resulting solution is often referred to as dilute sulfuric acid or dilute hydrochloric acid.

Again it should be noted that when making solutions of these acids, one should always add the acid to the water and not the water to the acid. This has to do with the hazards that acids present to the lab worker. See Chapter 12 for more details on this general rule.

10.6 Increasing Dissolution Rate

Sometimes, chemists and laboratory technicians attempt to dissolve a solid in water only to find out that it apparently will not dissolve even if solubility data show that it will. Such a situation is more of a problem with the rate of dissolving than with solubility. If the solute were allowed the time needed, it would dissolve. However, there are several methods for increasing the rate of dissolving.

10.6.1 Particle Size

The greater the surface area of a solid that is exposed to a solvent, the faster the solid will dissolve in the solvent. The surface area of a solid can be dramatically increased by crushing the solid into very small pieces. Thus large clumps of a solid will dissolve more slowly than powdered solid.

10.6.2 Agitation

Shaking a solution can also aid dissolving and is probably the first technique that a novice would try. If the time to dissolve is on the order of hours, however, shaking can be monotonous. In this case, a mechanical shaker can be used. The solution can also be stirred, and stirring can be accomplished with a mechanical (magnetic) stirrer with a magnetic stirring bar. If one is available, an ultrasonic cleaner bath can be very effective, often requiring a matter of minutes when a stirrer or shaker may require hours.

10.6.3 Temperature

As mentioned in Section 10.5, higher temperatures can increase solubility. Higher temperatures can also increase dissolving rate. Thus, elevating the temperature and combining this with shaking or stirring can be very effective in getting a solute dissolved quickly.

10.7 Mechanism of Dissolving

A look at what happens to molecules and ions at the surface of a solid that has been introduced into a liquid solvent, or at the interface of the two liquids when a liquid solute is introduced to a liquid solvent, helps explain why some of substances dissolve and others do not. In this section, we examine the interactions between the molecules and ions of a solute with the molecules of a solvent. We consider water as the solvent in Sections 10.7.1 and 10.7.2 and nonaqueous solvents in Section 10.7.3.

10.7.1 Ionic Solutes in Water

We first take a look at ionic solutes, such as sodium chloride (NaCl). As we stated in Chapter 6 (Figure 6.21), such compounds do not exist as molecules but rather as formula units in a three-dimensional (3-D) array of ions. That is why we refer to formula units rather than molecules and to formula weights rather than molecular weights.

When a crystal of an ionic compound is dropped into a container of water, there is an immediate attraction of the little “magnets” we call water molecules for the ionic charges in the 3-D ion array—the positive end of the water molecule being attracted to the negative ions and the negative end of the water molecule being attracted to the positive ions. As the water molecules gather around and align themselves—positive ends to negative ions and negative

ends to positive ions—they begin to tug away at the ions, attempting to pull them out of the array. In most cases, they succeed in breaking the ionic bonds and in ripping the ions from the array. The result is a solvated ion—an ion that is carried off into the water as a separate and distinct positive or negative ion, totally separated from its partner in the array and surrounded by water molecules. At some point, the entire crystal disintegrates and disappears. We say that the crystal has then dissolved. Figure 10.2 is an illustration of this process.

If the attractive strength of the water molecules is not sufficient to break the ionic bonds, then the compound does not dissolve. While most ionic compounds dissolve in water, a number of them do not. Precipitation formation in a chemical reaction is related to these concepts. If, after mixing two separate solutions together, there are two kinds of ions present that form a stronger ionic bond than the strength of the solvation bonds to water molecules, the ionic bonds will form, and an undissolved compound (or precipitate) will result.

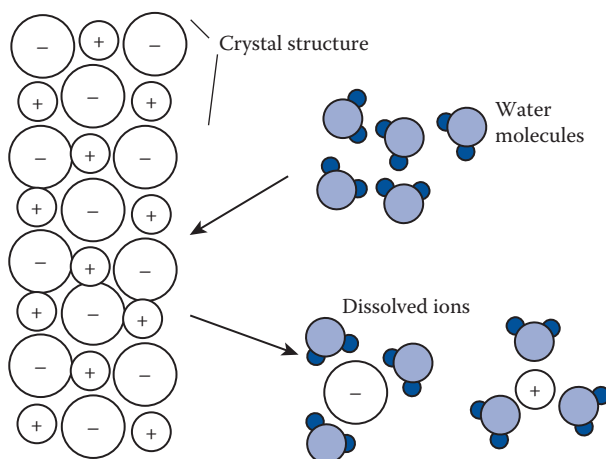


FIGURE 10.2 An illustration of an ionic compound, such as sodium chloride, dissolving in water. The water molecules, by virtue of their polar nature, pull the positive and negative ions from the ionic crystal array, and the compound dissolves. (From Kenkel, J., Kelter, P., and Hage, D., *Chemistry: An Industry-Based Introduction with CD-ROM*, CRC Press, Boca Raton, FL, 2001. With Permission.)

In general, we can make the following statements about common ionic compounds dissolving in water:

- All nitrates are soluble.
- Most acetates, sulfates, and chlorides are soluble.
- Most carbonates, hydroxides, and phosphates are insoluble.
- Most sodium, potassium, and ammonium salts are soluble.

10.7.2 Molecular Solutes in Water

Some compounds that consist of molecules (covalent compounds) are soluble in water and some are not. Some of those that are soluble form ions when they dissolve, but most do not. Let us now take a look at what happens at the molecular level in these situations.

Polar solute molecules interact with water molecules in a fashion similar to the way ions interact with water molecules. The negative end of a water molecule is attracted to the positive end of a polar solute molecule and the positive end of a water molecule is attracted to the negative end of a polar solute molecule. The net effect is a breakdown of the cohesive forces holding the solute molecules together, which causes dissolving or mixing to occur. Thus polar solutes such as sugars, low-molecular weight alcohols, and acetone dissolve readily in water. This process is depicted in Figure 10.3.

A few polar covalent solutes, namely acids, dissolve in water and also form ions in the process. The strength of the interaction with the water molecule is sufficient to break a covalent bond in these solutes and to form ions. Inorganic acids, such as HCl, H₂SO₄, and HNO₃, completely ionize and are called strong acids. This ionization can be written as follows for hydrochloric acid (HCl), for example:



In this example, no undissociated HCl molecules remain. After HCl has dissolved, all HCl molecules are present as the ionized acid, with H⁺ ions and Cl⁻ ions now present in solution.* This process is depicted in Figure 10.4.

* Technically, the hydronium ion, H₃O⁺, a combination of the hydrogen ion and a water molecule, forms instead of the hydrogen ion, H⁺.

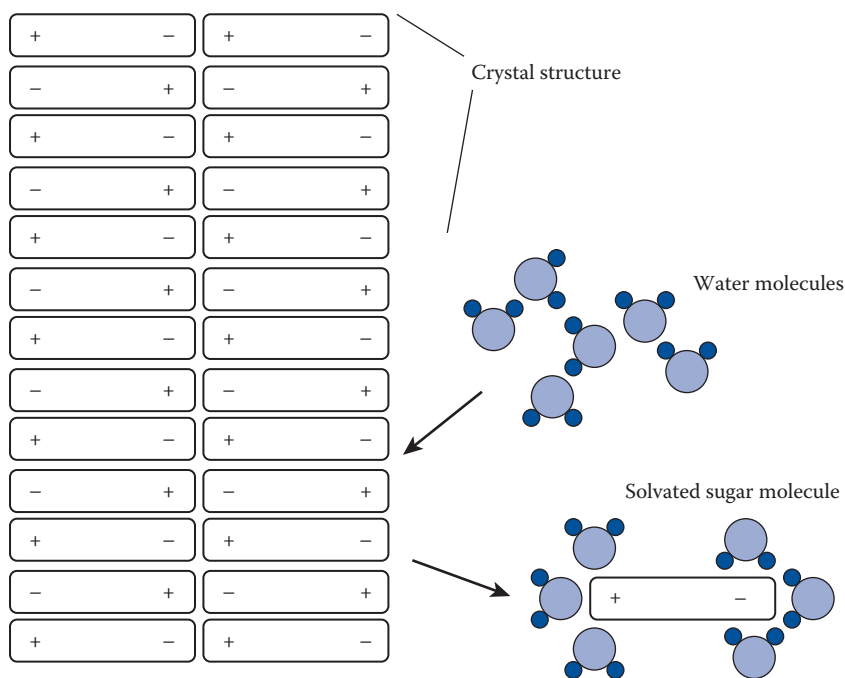


FIGURE 10.3 An illustration of a polar covalent compound, such as sugar, dissolving in water. The water molecules, by virtue of their polar nature, pull the polar sugar molecules from the crystal array, and the compound dissolves.

Strong acid case:

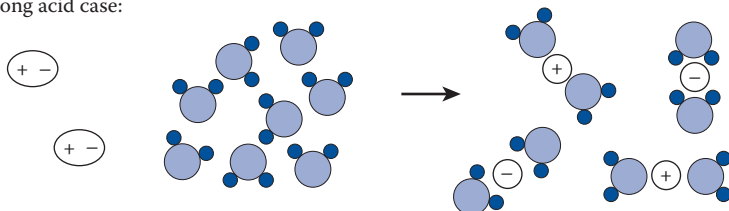


FIGURE 10.4 An illustration of a strong acid dissolving in water. The acid is a polar covalent compound, like sugar, but when it dissolves in water, it ionizes—that is, the hydrogen atoms break away from the molecule, forming hydrogen ions (+ ions) and – ions. It is a total, complete ionization.

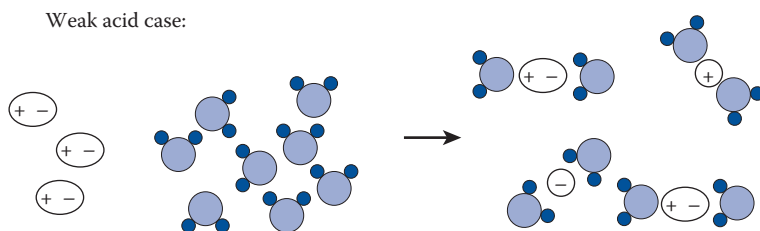


FIGURE 10.5 An illustration of a weak acid dissolving in water. The acid is a polar covalent compound, just like a strong acid. It dissolves completely in water, but only partially ionizes.

Organic acids, such as the carboxylic acids (Chapter 14), can dissolve in water but only partially ionize. The usual example of such an acid is acetic acid ($\text{HC}_2\text{H}_3\text{O}_2$). More details of this partial ionization are discussed in Chapter 11. Acids that partially ionize are called weak acids. The process of a weak acid dissolving in water is depicted in Figure 10.5.

If the molecular solute is nonpolar, there is no interaction with the polar water molecules. In addition, the water molecules are held together with the hydrogen bond, so no mixing occurs. Thus nonpolar compounds such as alkanes, alkenes, aromatic hydrocarbons, and carbon tetrachloride are immiscible with water. See Figure 10.6 for an illustration.

10.7.3 Nonaqueous Solvents

If water is not the solvent, the interaction between solvent molecules and solute ions or polar solute molecules decreases substantially and can disappear completely. This means that, while ionic compounds and polar molecular solutes may partially dissolve in nonaqueous polar solvents (such as methyl alcohol), their solubility is not as high as in water. This is illustrated in Figure 10.7.

Perhaps more interesting is the case of a nonpolar solvent with a nonpolar solute. Obviously there will be no polar–polar interaction between solvent molecules and solute molecules. One may conclude that such substances are insoluble because of this lack of interaction. However, since the molecules of both the solute and the solvent are nonpolar, there are only very weak forces, called **London forces**, that attract solute molecules to other solute molecules and solvent molecules to other solvent molecules. Thus these two mix and dissolve due to a lack of interactions. This is illustrated in Figure 10.8.

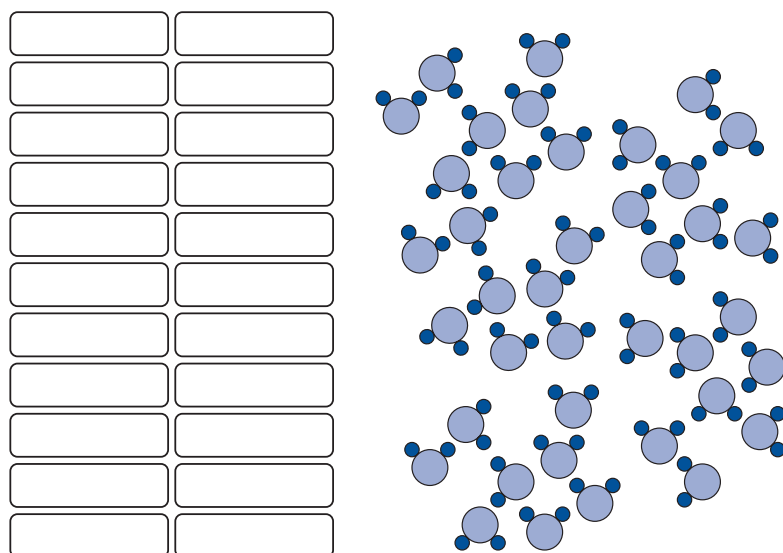


FIGURE 10.6 An illustration of a nonpolar covalent compound in contact with water. There is no interaction between the two and, in addition, the water molecules are held together by the hydrogen bond, so there is no mixing and no dissolving.

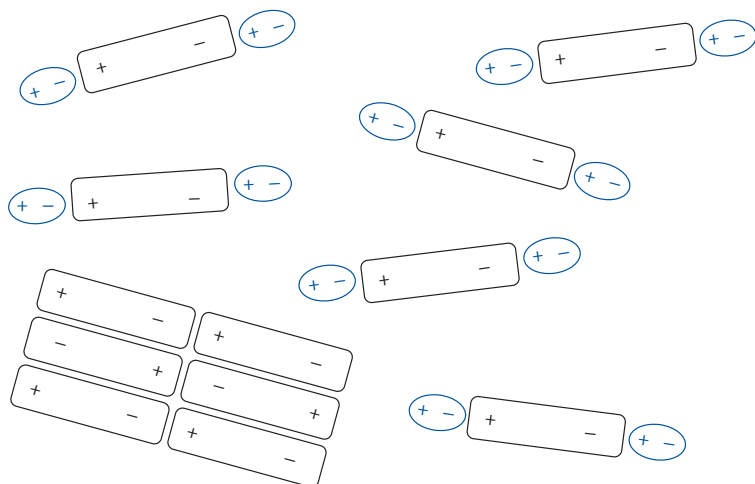


FIGURE 10.7 An illustration of a polar solvent, other than water, dissolving another polar compound such as sugar. The solvent dissolves the compound, but since it is less polar than water, the solubility will likely be lower.

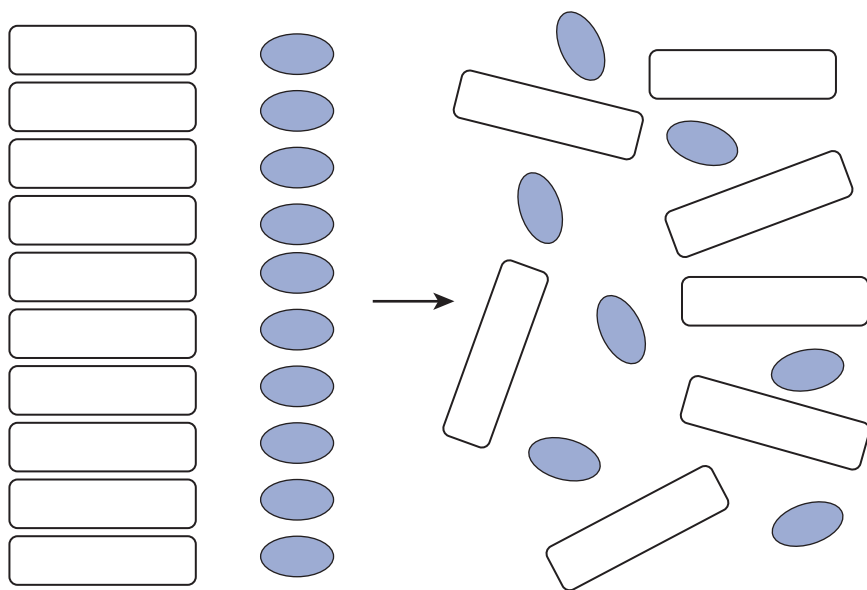


FIGURE 10.8 An illustration of two nonpolar liquids placed in the same container. While there is no polar–polar interaction among their molecules, there is also no hydrogen bonding (as there was in Figure 10.6) or other forces present that would prevent their mixing. Thus there is complete dissolving and mixing.

The old adage *like dissolves like* appears to cover all these cases. Polar (or ionic) substances dissolve in other polar substances, polar (or ionic) substances do not dissolve in nonpolar substances, and nonpolar substances dissolve in other nonpolar substances.

10.8 Homework Exercises

Basic Solution Terminology

1. Define these terms.
 - a. Solute
 - b. Solvent
 - c. Solution

2. State whether each statement is true or false, and explain your answer.
- A solution is a homogeneous mixture.
 - The solute is the solution component with the highest concentration.
 - A solution can contain more than one solute and more than one solvent.
 - The physical state of solutions can only be liquid.
 - The solute in a solution can be a solid, liquid, or gas.

Calculation of Volume Percent Concentration

3. If 33.2 mL of ethyl alcohol are dissolved in water so that the total volume of solution is 350.0 mL, what is the concentration of this solution when expressed as a volume percent?
4. Fill in the blanks in the following table.

	Volume of Solute	Volume of Solution	Volume Percent
(a)	12.0 mL	350.0 mL	
(b)	69.6 mL	200.0 mL	
(c)	56 mL	0.400 L	
(d)	0.34 L	800.0 mL	
(e)	0.838 L	3.00 L	

5. Consider a solution in which 30.0 mL of acetone are dissolved in 70.0 mL of water. Is this the same as a solution in which 30.0 mL of acetone are dissolved in 100.0 mL of a solution with water as the solvent? Explain.

Calculation of Weight/Volume Percent Concentration and Molarity

6. What is the weight/volume percent concentration of each of the following solutes in the solutions given? Explain the similarities of your answers.
- 2.337 g of K_2SO_4 dissolved in 200.0 mL of solution
 - 2.337 g of KCl dissolved in 200.0 mL of solution

7. If you prepare a solution in which 2.337 g of KCl are dissolved in 200.0 mL of water, would this solution be exactly the same as the one described in Exercise 6b? Explain.
8. Calculate the molarity of each of the following solutes in the solutions given.
 - a. 0.339 mol of a solute dissolved in 250.0 mL of solution
 - b. 0.449 g of KNO_3 dissolved in 350.0 mL of solution
 - c. 0.281 mol of NH_4OH dissolved in 500.0 mL of solution
 - d. 1.227 g of $\text{BaCl}_2 \cdot 2\text{H}_2\text{O}$ dissolved in 200.0 mL of solution
9. A chemist finds 250.0 mL of a solution for which the container label says 6.33 g of BaCl_2 are dissolved. What is the molarity of this solution?
10. Fill in the blanks in the following table.

	Grams of Solute	Volume of Solution	Molarity	w/v %
(a)	0.937 g NaOH	500.0 mL		
(b)	29 mg NaCl	600.0 mL		
(c)	0.477 g Na_2SO_4	1.20 L		
(d)	0.0083 kg CaCl_2	2500.0 mL		
(e)	549 mg K_3PO_4	0.400 L		
(f)	0.0074 kg K_2HPO_4	1.60 L		
(g)	2.59 g NaH_2PO_4	200.0 mL		
(h)	11.09 g Na_2HPO_4	3.00×10^3 mL		
(i)	123.3 mg KH_2PO_4	150.0 mL		
(j)	0.034 kg NaHCO_3	0.8000 L		

11. a. If you had a solution of KOH that is 0.450 M, what is the concentration expressed in weight/volume percent?
- b. If you had a solution of KOH that is 2.93% (weight to volume), what is the concentration in molarity?

Solution Preparation

12. How many milliliters of concentrated HCl (12.0 M) are needed to make each of the following?
 - a. 250.0 mL of 0.35 M HCl
 - b. 1.50 L of 0.60 M HCl

13. How many milliliters of concentrated (18.0 M) sulfuric acid are needed to prepare each of the following?
 - a. 500.0 mL of 0.35 M H_2SO_4
 - b. 2.50 L of 0.050 M H_2SO_4
14. How would you prepare the following?
 - a. 500.0 mL of a 0.20 M solution of nitric acid, starting with concentrated nitric acid, which is 16.0 M
 - b. 1.20 L of a 0.50 M solution of acetic acid, starting with concentrated acetic acid, which is 17 M
15. A chemist gives a laboratory technician a flask with a 2.00-L calibration line and a bottle of 12 M hydrochloric acid and asks her to prepare 2.00 L of a 1.5 M solution. If it were you, tell how you would do it.
16. Fill in the blanks in the following table.

	Concentration of the Solution to be Diluted	Volume of Diluted Solution Desired	Concentration after Dilution	Milliliters of Original Solution Needed
(a)	6.0 M NaOH	500.0 mL	0.10 M	
(b)	12 M HCl	250.0 mL	0.10 M	
(c)	1.048 M	0.200 L	0.1048 M	
(d)	5.0 M H_2SO_4	0.50 L	1.0 M	
(e)	37% HCl	1000.0 mL	5.0%	
(f)	5.56%	200.0 mL	2.0%	
(g)	96% H_2SO_4	0.250 L	15%	
(h)	25%	2.00 L	5.0%	

17. When would you use a volumetric pipet and a volumetric flask to prepare a solution by dilution, as opposed to a graduated cylinder and a beaker?
18. How many milliliters of ethyl alcohol are needed to prepare the following?
 - a. 300.0 mL of a 15.0% by volume solution
 - b. 24.0 L of a 1.50% by volume solution
19. If a person's blood alcohol level is determined to be 0.15% by volume and this person body contains 5.0 L of blood, how many liters of alcohol are in his or her blood?

20. A 15% (weight to volume) solution of $\text{Ba}(\text{OH})_2$ is needed. How many grams of solid $\text{Ba}(\text{OH})_2$ would be used if you prepared 250.0 mL of the solution?
21. How many grams of NaOH are needed to prepare 500.0 mL of a 10.0% (weight to volume) solution?
22. A 3.5% (weight to volume) NaCl solution is a solution that is corrosive to many metals and chemists often use it to simulate sea water. If a chemist asks you to make a 1.5 L of it using pure, solid NaCl , how many grams would you need?
23. How many grams of solute are required for each of the following? Since the volume of solution and concentrations are the same, why are your answers different?
 - a. 500.0 mL of a 0.25 M solution of KH_2PO_4
 - b. 500.0 mL of a 0.25 M solution of K_2HPO_4
24. Imagine yourself as a laboratory technician and your supervisor asks you to prepare 250.0 mL of a 0.60 M solution of KOH , and she hands you a container of pure, solid KOH and a piece of glassware that has a 250.0 mL calibration line. Tell how you would prepare this solution using these materials.
25. A chemist needs a 0.75 M solution of sodium carbonate (Na_2CO_3) and asks you to make it. You find that you have a bottle of pure, solid Na_2CO_3 and a solution of Na_2CO_3 that is 5.5 M in the lab. What are two methods for preparing 500.0 mL of the solution with the materials at hand and a flask with a 500.0 mL calibration line?
26. A chemist needs a 500.0 mL of a 0.35 M solution of potassium dihydrogen phosphate (KH_2PO_4). Suppose that you are assigned to prepare this solution. You have the pure solid available, but you also have a solution that is 4.0 M.
 - a. If you decide to use the pure solid, how would you prepare the solution?
 - b. If you decide to use the 4.0 M solution instead, how would you prepare the solution?
27. How would you prepare 750.0 mL of a 0.35 M solution of Na_2SO_4 using pure, solid Na_2SO_4 and a container with a 750.0 calibration line?

28. Fill in the blanks in the following table.

	Molarity or w/v % Desired	Volume Desired	Grams to be Weighed
(a)	0.874% KCl	500.0 mL	
(b)	0.874 M KCl	500.0 mL	
(c)	1.44% K_2CO_3	1.50 L	
(d)	1.44 M K_2CO_3	1.50 L	
(e)	3.0% NaCl	300.0 mL	
(f)	0.6339 M Na_2CO_3	800.0 mL	
(g)	9.11% $KHCO_3$	0.4000 L	
(h)	0.500 M NaOH	700.0 mL	
(i)	12.3% KOH	0.2500 L	
(j)	6.0 M KOH	200.0 mL	

Additional Solution Terminology

29. Define each of the following terms that relate to solutes.
 - a. Soluble
 - b. Insoluble
 - c. Solubility
30. Define each of the following terms that relate to solutions.
 - a. Saturated
 - b. Unsaturated
 - c. Supersaturated
31. Point out the differences between the following pairs of terms.
 - a. Miscible and immiscible
 - b. Dilute and concentrated
32. Tell whether each of the following is true or false and explain your answer.
 - a. If two liquids are miscible, a mixture of the two will be homogeneous.
 - b. Nonpolar substances tend to be soluble in water.
 - c. Two liquids that are not capable of mixing and forming a solution are said to be miscible.

- d. An increase in temperature increases the solubility of most solids.
 - e. A dilute solution contains more solute than does a concentrated solution.
 - f. Evidence of saturation is some undissolved solute in the container with the solution.
 - g. Since supersaturated solutions have more than the maximum amount of dissolved solute than is possible to be dissolved, one cannot make them without raising the temperature.
 - h. Carbon tetrachloride, a nonpolar liquid, is immiscible with water.
 - i. The solubility of NaCl in water is 36.0 g/100 g of water. Therefore a solution of NaCl that contains 35.2 g in 100 g of water is an unsaturated solution.
33. Answer each of the following with just one word.
- a. What word describes a solution that has some undissolved solid solute present even after shaking for a period of time?
 - b. What word describes two liquids that, after being placed in the same container and shaken, form two distinct liquid layers in the container?
 - c. What solution component is present in the lesser amount in a solution?
 - d. The maximum amount of Li_2CO_3 that can be dissolved in 100 g of water at 20°C is 1.33 g. This is known as the _____ of Li_2CO_3 .
 - e. What word describes a solution that has more solute dissolved than the maximum amount possible at a given temperature?
 - f. If you take a crystal of sodium nitrate and place it in a beaker of water and then stir, it would dissolve. We say that the sodium nitrate is _____ in water.
 - g. _____ sulfuric acid, as purchased from a chemical supply company, is 18 M.
34. A new chemistry technician who has come to work in the same lab as you says that, when she tested a compound for its solubility in water (a small crystal in 5 mL), she found it to be insoluble. She is confused, however, because a chemistry handbook says that its solubility is 0.0135 g/100 mL. How would you explain to her that while her conclusion is correct, so is the handbook?

Dissolution Rate

35. Describe three different methods of agitation that would result in a more rapid rate of dissolving a solid solute in a solvent.
36. Why is particle size a factor in dissolution rate?

Mechanism of Dissolving

37. In terms of sodium ions, chloride ions, and water molecules, describe exactly what happens when a crystal of NaCl dissolves in water.
38. Suppose that a scientist is close to a breakthrough on a new microscope through which he says he will be able to see atoms, ions, and molecules. Suppose that he wants to try it out by focusing on the process of an ionic solid dissolving in water just to see what happens to the ions and water molecules as the solid dissolves. What would you tell him he should expect to see?
39. Based on the general rules listed in Section 10.7.1 for ionic solids, which of the following would be soluble in water and which would not: NaBr, K_2SO_4 , $Fe(OH)_3$, NH_4Cl , $AgNO_3$, Ag_2CO_3 , $Ba(C_2H_3O_2)_2$, Na_3PO_4 .
40. What general class of compounds dissolve and ionize in water even though they are not ionic? Give three examples.
41. What is the difference between a weak acid and a strong acid in terms of what happens to the molecules when they dissolve in water?
42. Why does carbon tetrachloride, a nonpolar organic liquid, not dissolve in water?
43. What are London forces? How are London forces considered when discussing the solubility of a nonpolar solute in a nonpolar solvent?
44. What is meant by the phrase *like dissolves like*? Explain the chemical basis for this statement.

Chapter 11

Chemical Equilibrium

Introduction	280
Chemical Equilibrium	282
Important Examples of Chemical Equilibrium	283
Equilibrium Constant	284
Position of the Equilibrium	287
Magnitude of the Equilibrium Constant	288
Le Chatelier's Principle	290
Effect of Concentration Change	291
Effect of Temperature Change	293
Effect of Pressure Change	294
Effect of Catalysts	296
Homework Exercises	296

11.1 Introduction

Consider an experiment in which all the students in your class line up against the walls of the classroom, half on one side and half on the other. Then consider what would happen when one student on each side walks to the opposite side at exactly the same speed, passing each other in the middle. As soon as each of these students arrives at the other side, another student walks back to the opposite side, continuing at the same speed. Now imagine continuing this process indefinitely. It would be a very dynamic process—two students always moving from one side to the other in opposite directions—but there would never be any net change in the count of students on each side. If we were unable to see the students walking and were only able to count the number of students in the lines, we would never know that anything was happening. This is an example of what is known as equilibrium. **Equilibrium** is a process in which two opposing processes occur at the same time and at the same rate such that there is no net change.

The phenomenon of equilibrium occurs in chemical systems. Such systems are said to be **reversible**, which means that a process occurs in one direction, but the reverse process can also occur at the same time and at the same rate. The concept of vapor pressure (Chapter 9) is explained based on this idea. Imagine a container half filled with water. Also imagine that the container is sealed so that no water vapor escapes. The water evaporates, and this vapor diffuses into the air space above the liquid water. However, the evaporation of liquid water is reversible, and the water vapor can condense and become liquid water again. This occurs in this example because the container is sealed. The evaporation and condensation occur at the same time and at the same rate (Figure 11.1) so that there is no net change in the quantity of liquid water and no net change in the quantity of water vapor in the air space above the liquid water. The constant quantity of water vapor in the air space leads us to define the vapor pressure of a liquid as the partial pressure exerted by the vaporized liquid in the air space. This quantity is a constant at a given temperature because it represents an equilibrium situation.

Chemists utilize a double arrow symbol ($\rightleftharpoons$) to designate systems at equilibrium. For example, the equilibrium occurring in the beaker of water can be expressed as follows:



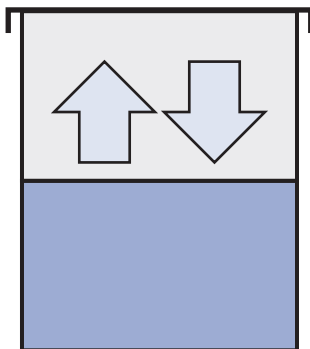


FIGURE 11.1 An example of a physical equilibrium first discussed in Chapter 9 (see also Figure 9.9) to define vapor pressure. The rate of evaporation of a liquid in a closed container is equal to the rate of condensation of the vapors. The quantity of the liquid or the vapor does not change even though both evaporation and condensation are occurring.

It is important to remember that this equilibrium only occurs when the container is sealed. If it is not sealed, the water vapor would escape into the air and would not be available for the reverse reaction to occur.

The melting of ice is also reversible. An equilibrium occurs in a beaker containing liquid water and crushed ice in a uniform slushy mixture in surroundings that are held at a constant temperature of 0°C. Under these conditions, liquid water is freezing at the same time and at the same rate as the ice is melting (Figure 11.2a). The water never completely freezes and the ice never completely melts. This equilibrium can be expressed as follows:



A third example is the behavior of a solute when given a choice of two solvents in which to dissolve. The dissolving of a solute in a solvent is reversible. Imagine a water solution of iodine (I_2), and imagine introducing a water-immiscible organic solvent, such as hexane, into the container. Because it is nonpolar, the hexane does not mix with the water and forms a separate layer of liquid in the container. After vigorous shaking, the layers separate again and the hexane layer appears pink due to iodine being dissolved in it. It is actually an equilibrium system because the iodine is moving into the hexane

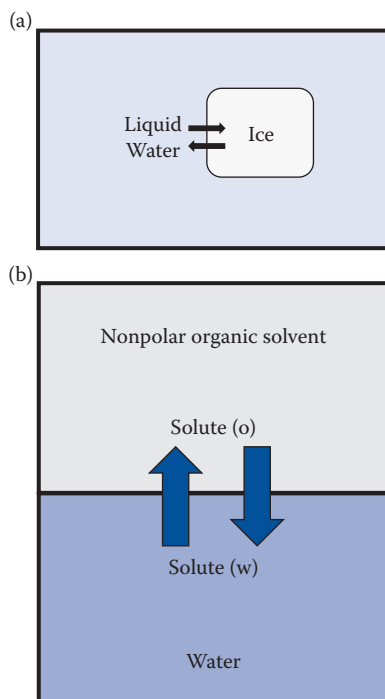


FIGURE 11.2 Illustrations of other physical equilibria. (a) The melting and freezing of water in a slushy ice–water mixture when the surroundings are at 0.0 °C. (b) The movement of a solute back and forth between immiscible layers of water and an organic solvent in an extraction experiment.

at the same time and at the same rate as it is moving back into the water layer as expressed in the following:



in which *w* refers to water and *o* refers to organic, meaning the hexane layer. See Figure 11.2b.

11.2 Chemical Equilibrium

The two opposing processes in the definition of equilibrium may be chemical reactions. Imagine placing two chemicals that react with other, A and B, into a reaction vessel. Let us say that the products of the reaction are

C and D and that the balancing coefficients are a , b , c , and d , as indicated in Equation 11.4.



As the reaction proceeds, C and D begin to form, and their concentrations increase while the concentrations of A and B decrease. Now let us say that C and D also react with each other and that A and B are the products.



This means that the reverse of reaction shown in Equation 11.4 occurs at the same time. After a period of time, imagine that these two opposing chemical reactions occur not only at the same time but also at same rate. When this happens, chemical equilibrium occurs. **Chemical equilibrium** thus refers to two opposing chemical reactions occurring at the same time and at the same rate with no net change. Such a reaction is written with a double arrow ($\rightleftharpoons$).

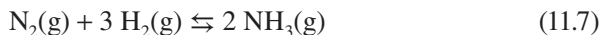


As with the vapor pressure example, equilibrium will not occur if any of the chemicals, reactants, or products escape or are removed from the container.

11.3 Important Examples of Chemical Equilibrium

Important examples of chemical equilibrium systems include (1) the Haber process for the manufacture of ammonia from hydrogen gas and nitrogen gas, (2) the ionization of weak electrolytes in water, and (3) the ionization and dissolution of ionic solids in saturated solutions.

In the Haber process, hydrogen gas and nitrogen gas are brought together at a relatively high temperature and pressure and in the presence of a catalyst in a closed container. The chemical equilibrium that occurs is expressed as follows:



Here nitrogen and hydrogen in the same container react to give ammonia, but the reverse reaction also occurs and the reaction reaches a point where the

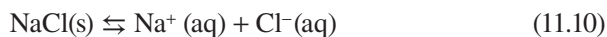
forward and reverse reactions occur at the same time and at the same rate. We will use this as an example throughout the discussion in this chapter.

Examples of weak electrolytes are the carboxylic acids, such as acetic acid ($\text{HC}_2\text{H}_3\text{O}_2$), and some bases, such as ammonium hydroxide (NH_4OH). When these chemicals are introduced into water, they partially ionize, meaning that some fraction of the molecules or formula units split apart into ions. This occurs at the same time and at the same rate as the ions are coming together to form the un-ionized molecules or formula units. They are therefore examples of chemical equilibrium. The reactions for this are written as follows:



Both of these examples of equilibrium are very important in real-world analytical laboratories. Solutions of weak electrolytes are used to control the acidity of solutions in laboratory analyses. Such solutions are known as **buffer solutions**.

An equilibrium also exists at the surface of the undissolved ionic solid in a saturated solution of the solid. An example is a saturated solution of salt (NaCl) in water (See Chapter 10, Section 10.5). The sodium ions and chloride ions break away from the ionic array at the same time and at the same rate as these same ions in solution come together to become part of the ionic array. This equilibrium can be expressed as follows:



11.4 Equilibrium Constant

In Section 11.1, we introduced the concept of equilibrium by suggesting a classroom experiment in which students in a class split into two groups and line up on two sides of the classroom, with one group on each side. We stated that as students pass from one side of the room to the other at the same time and at the same rate, there is no net change in the count of students on either side. Because there is no change in these counts, we can divide the number of students on one side by the number of students on the other side and always obtain the same number. This same concept applied to a chemical equilibrium system results in a number called the **equilibrium constant**.

In a chemical equilibrium system, the count is the molar concentration of the chemical species involved. For the purpose of symbolizing this molar concentration (and the equilibrium constant), we now introduce the symbolism of the brackets, []. The molar concentration (molarity) of a chemical species is symbolized by enclosing the formula of this species within brackets. For example, $[H^+]$ refers to the molar concentration of the hydrogen ion, $[NH_3]$ refers to the molar concentration of ammonia, and $[Cl^-]$ refers to the molar concentration of the chloride ion.

The **equilibrium constant** for a chemical equilibrium is defined as the molar concentrations of the products of the reaction, raised to the power of their respective balancing coefficients and multiplied together, divided by the molar concentrations of the reactants of the reaction, raised to the power of their balancing coefficients and multiplied together. Thus, for the general reaction



the equilibrium constant (K_{eq}) is defined as in Equation 11.12.

$$K_{eq} = \frac{[C]^c [D]^d}{[A]^a [B]^b} \quad (11.12)$$

Knowing the molar concentrations of all the chemicals involved in the equilibrium, we can calculate the number that is given by the equilibrium constant. For Equation 11.7, K_{eq} is as shown in Equation 11.13

$$K_{eq} = \frac{[NH_3]^2}{[H_2]^3 [N_2]} \quad (11.13)$$

and for Equation 11.8, K_{eq} is shown in Equation 11.14.

$$K_{eq} = \frac{[H^+] [C_2H_3O_2^-]}{[HC_2H_3O_2]} \quad (11.14)$$

and for Equation 11.9, K_{eq} is as follows:

$$K_{eq} = \frac{[NH_4^+][OH^-]}{[NH_4OH]} \quad (11.15)$$

Equations 11.8 and 11.9 represent ionization equilibria in which an uncharged chemical species (molecule or formula unit) splits apart into

ions. Equation 11.8 represents the ionization of a weak acid, acetic acid, and Equation 11.9 represents the ionization of a weak base, NH_4OH . We often refer to the equilibrium constants of ionization equilibria as **ionization constants** and symbolize them as K_i . Additionally, if the ionizing species is a weak acid, such as $\text{HC}_2\text{H}_3\text{O}_2$, we often refer to the ionization constant as the **acid dissociation constant**, K_a (a referring to acid). If the ionizing species is a weak base, such as NH_4OH , we refer to the ionization constant as K_b (b referring to base). Equations 11.14 and 11.15 are rewritten here with these more appropriate symbols.

$$K_a = \frac{[\text{H}^+][\text{C}_2\text{H}_3\text{O}_2^-]}{[\text{HC}_2\text{H}_3\text{O}_2]} \quad (11.16)$$

and

$$K_b = \frac{[\text{NH}_4^+][\text{OH}^-]}{[\text{NH}_4\text{OH}]} \quad (11.17)$$

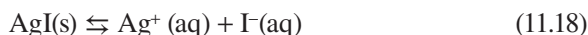
In these two ionization examples, all species in the chemical equilibrium are dissolved in water. This means that the molar concentrations needed in the equilibrium constant expression are all real numbers that can be used in the calculations. This includes the un-ionized species (e.g., $\text{HC}_2\text{H}_3\text{O}_2$ and NH_4OH), as well as the ions.

The ionization of an ionic salt, such as NaCl (see Equation 11.10), poses a special problem. The un-ionized chemical species in this equilibrium is *not* dissolved. This is an example of what is termed a **heterogeneous equilibrium** (i.e., two or more of the equilibrium participants are present in different phases [solid, liquid, gas, or dissolved]). In this case, we have a solid that is *not* dissolved while everything else *is* dissolved. Since the solid is undissolved, it does not make sense to refer to its molar concentration. An interesting and important fact about this undissolved solid is that regardless of how much is present, the concentrations of the dissolved ions are constants (at a given temperature). Thus, while the molar concentrations of the ions are real numbers, the molar concentration of the un-ionized species is a nonsensical term. In this case, a special equilibrium constant is defined which uses only the molar concentrations of the dissolved ions in its definition. This special equilibrium constant is called the **solubility product constant** (K_{sp}), which is defined as the mathematical product of the molar concentrations of the ions raised to the power of their balancing

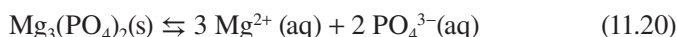
Table 11.1 Symbols Used for Various Equilibrium Constant Expressions

- K_{eq} : General symbol for all equilibrium constant expressions
- K_i : General symbol for ionization equilibrium
- K_a : Symbol for weak acid ionization equilibrium
- K_b : Symbol for weak base ionization equilibrium
- K_{sp} : Symbol for solubility product constants, which are special expressions for the ionization and solubility of ionic compounds

coefficients. It has its greatest usefulness in systems in which the ionic solid is only *slightly* soluble. Examples include silver iodide and magnesium phosphate.



$$K_{sp} = [\text{Ag}^+][\text{I}^-] \quad (11.19)$$



$$K_{sp} = [\text{Mg}^{2+}]^3 [\text{PO}_4^{3-}]^2 \quad (11.21)$$

The symbols for the equilibrium constants for the various systems are summarized in Table 11.1.

11.5 Position of the Equilibrium

The position of an equilibrium refers to how far the reaction has proceeded to the right before equilibrium is achieved. If the reaction progresses only slightly to the right before equilibrium is achieved, then we say that the equilibrium (or position) lies far to the left, meaning that the reactants predominate over the products in the reaction vessel. If the reaction progresses so that almost all of the reactants have been converted to products before equilibrium is achieved, then we say that the equilibrium (position) lies far to the right, meaning that the products predominate over the reactants. This “position” can be located at any point as the forward reaction proceeds depending on the individual reaction. Figure 11.3 is an illustration of this.

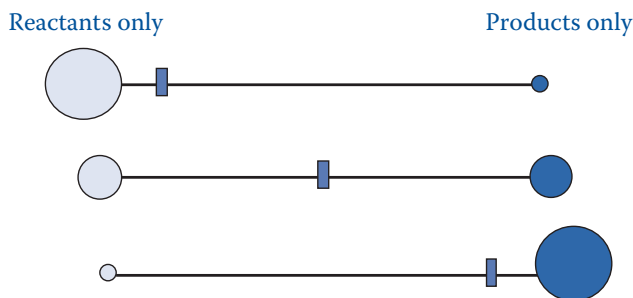


FIGURE 11.3 An illustration of what is meant by the position of a chemical equilibrium. Position refers to how far the reaction has gone to the right before equilibrium is achieved. In this figure, the small rectangles indicate the position of the equilibrium while the sizes of the circles indicate the relative quantities of reactants and products present at equilibrium.

11.6 Magnitude of the Equilibrium Constant

At this point, we want to remind you that regardless of what constant we are speaking of (the equilibrium constant, the ionization constant, or the solubility product constant) they are each a number. As indicated previously, if we know the molar concentrations of the reactants and products, this number may be calculated. We will not be concerned with such calculations at this point. However, it is useful to gain a feel for the magnitude that this number can have and of what it means when this number is large or small. Depending on the individual equilibrium process involved, this number can range from an extremely small number to an extremely large number.

An extremely small number results when the molar concentrations of the products (the numerator of the equilibrium constant expression) are very small values and the molar concentrations of the reactants (in the denominator) are very large. Dividing a small number by a large number, like when dividing the number 1 by the number 1000, results in a very small number. So an equilibrium constant is small when the concentrations of the products is small relative to the concentrations of the reactants. This occurs when the reverse reaction rate becomes equal to the forward reaction rate before the forward reaction has a chance to proceed to any appreciable extent, or when the reaction mixture consists mostly of the reactants. The position of the equilibrium would be far to the left (Figure 11.4).



FIGURE 11.4 We say that the equilibrium lies far to the left if the reactants predominate in the reaction mixture and that the equilibrium constant is therefore a small number.



FIGURE 11.5 We say that the equilibrium lies far to the right if the products predominate in the reaction mixture and that the equilibrium constant is therefore a large number.

Conversely, the equilibrium constant is large when the concentrations of the products is large compared with the concentrations of the reactants. This occurs when the reverse reaction rate becomes equal to the forward reaction rate after the reaction has consumed nearly all the reactants and is nearly complete, or when the reaction mixture consists mostly of the products. In this case, the calculation of the equilibrium constant would be like dividing a large number by a small number, meaning that the equilibrium constant would be a large number (Figure 11.5).

Weak acid and weak base ionization equilibria have extremely small ionization constants. For example, the ionization constant of acetic acid is 1.8×10^{-5} . The ionization constant of ammonium hydroxide is 1.7×10^{-5} . This means that there are very few ions present at equilibrium; most of the acid or base is present as the un-ionized species. This is the reason they are described as weak acids or bases. The position of ionization equilibrium in each case would be as illustrated in Figure 11.4.

Similarly, the solubility product constants of ionic precipitates (substances chemists usually describe as insoluble) are also extremely small. An apparently insoluble substance is never totally insoluble. In actuality, an equilibrium occurs in which there are an extremely small number of ions present. An example of this is $\text{Mg}_3(\text{PO}_4)_2$. If you were to drop a few grains of this solid into water and shake, it would appear to not dissolve. However, small

concentrations of magnesium and phosphate ions would be found in the solution. The solubility product constant of this compound is 1×10^{-25} , an extremely small number.

Finally, consider the nitrogen, hydrogen, ammonia equilibrium—the Haber process shown in Equation 11.7. This is an industrial process for the manufacture of ammonia. The equilibrium constant depends on both temperature and pressure. It can be large or small, depending on the conditions. However, to optimize the production of ammonia, chemical engineers control the conditions of temperature and pressure to give a large equilibrium constant because a large equilibrium constant means much more product (ammonia) than reactants (hydrogen and nitrogen) will be present in the reaction vessel, an important point in optimizing the production of the ammonia.

11.7 Le Chatelier's Principle

As mentioned for the Haber process, temperature and pressure can help control the position of an equilibrium, the amount of product and reactant present at equilibrium, and thus the equilibrium constant. Temperature and pressure are parameters that, when varied, create a stress on the equilibrium. Stresses on equilibrium systems cause changes in the system. The nature of these changes are described by Le Chatelier's principle.

Le Chatelier's Principle can be expressed as follows: when a stress is placed on a system at equilibrium, the equilibrium shifts to partially relieve the stress.

An equilibrium shift is defined as a spontaneous and momentary increase in the rate of an equilibrium reaction in one direction or the other, causing a temporary situation in which the reaction is not an equilibrium since the rates of the two opposing reactions are no longer equal. This phenomenon is temporary, since equilibrium is reestablished at some later moment. When the equilibrium is reestablished, the concentrations of the reactants and products are not what they were previously because the temporary rate imbalance caused a surge toward one side and away from the other. The equilibrium constant may or may not change during this process depending on what caused the shift. This will be discussed later in conjunction with the various stresses that can be placed on a chemical equilibrium and cause it to shift (Figure 11.6).

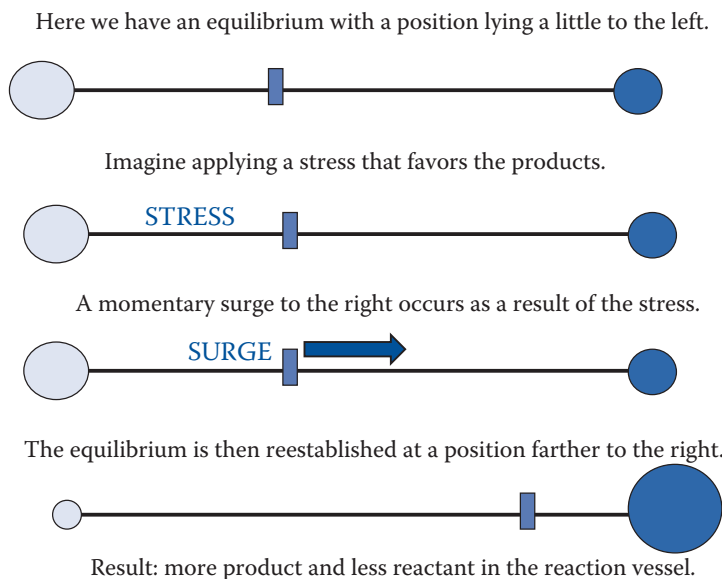
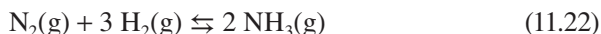


FIGURE 11.6 An illustration of Le Chatelier's principle when an applied stress favors the products. In this situation, there is a momentary surge from left to right before equilibrium is reestablished at the new position. At this new position, the amount of product is greater than it was and the amount of reactant is less.

11.8 Effect of Concentration Change

The change in concentration of a chemical participating in a chemical equilibrium is a stress on the equilibrium system. Thus such a change would cause the equilibrium to shift, as stated by Le Chatelier's principle. Relieving the stress in this case means that if the concentration increases, the equilibrium shifts so as to partially decrease it. If the concentration decreases, the equilibrium shifts so as to partially increase it. An increase in concentration may be caused simply by opening the reaction vessel and adding more of the chemical. A decrease in concentration may be caused by opening the reaction vessel and adding a chemical that would react with one of the chemicals involved in the reaction, thus removing this chemical from the system.

Let us use the Haber process for the manufacture of ammonia as an example. This reaction is shown again here.



Let us open the reaction chamber and pump in some nitrogen. This act places a concentration stress on the system consisting of an increase in the concentration of the N_2 . According to Le Chatelier's principle, the equilibrium must shift to relieve this stress. This would mean consuming some of the N_2 in order to return to equilibrium. Thus the equilibrium *shifts to the right*, as indicated here



before returning to the equilibrium state indicated in Equation 11.22.

Similarly, if we were able to somehow open the container and selectively pump out some of the hydrogen, this would also be a concentration stress. In this case, however, the equilibrium would shift to replace the hydrogen that was removed. This would mean *a shift to the left*, or toward the hydrogen in the equation so that more would be produced,



before once again returning to the equilibrium state indicated in Equation 11.22.

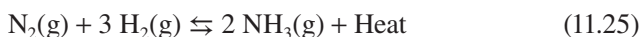
In both of these examples, the return to the original equilibrium state does *not* mean that everything is the same as before the stress was imposed. The concentrations of all the chemicals involved have changed. If there is a shift to the right, some hydrogen is consumed, and its concentration becomes lower. Also, some additional ammonia is produced, and its concentration is increased. And some nitrogen has been added and only *partially* consumed, so its concentration has increased. The concentrations of all three chemicals are different than before the stress was imposed. Two have increased while one has decreased. You may think that because these concentrations have changed, the value of the equilibrium constant for this reaction must also have changed since these concentrations are used in the calculation of K_{eq} . In reality, the value of K_{eq} does *not* change under these conditions. This indicates the true constancy of K_{eq} . In fact, we can say that in the event of a concentration stress, the equilibrium *shifts so as to maintain the constancy of the equilibrium constant*. Thus we do not look for the equilibrium constant to change; rather, we look for the shift to occur to an extent that would maintain its constancy.

11.9 Effect of Temperature Change

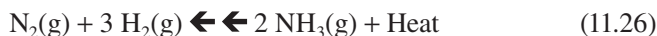
Another stress that we can impose on a chemical equilibrium system is an increase or decrease in temperature. All chemical equilibrium systems are affected by temperature changes. An increase in temperature causes a shift in one direction or the other, either to left or to the right. Also, a decrease in temperature causes a shift in one direction or the other, either to the left or to the right. Reactions that shift to the right when the temperature is increased are described as **endothermic** reactions. Reactions that shift to the left when the temperature is increased are described as **exothermic** reactions. Endothermic reactions require heat to make them proceed, while exothermic reactions release heat as the reaction proceeds.

An increase in temperature means the addition of heat to the system. A decrease in temperature means removal of heat from the system. The temperature stress may be thought of as a heat stress. One of the equilibrium reactions—either the forward reaction (the one from left to right) or the reverse reaction (the one from right to left)—is favored by the addition of heat. The situation is simplified if we simply think of heat as a reactant or product in the equation. If the forward reaction is favored by the addition of heat, then we can place the word *Heat* on the left side to indicate that the addition of heat will make the reaction proceed from left to right. If the reverse reaction is favored by the addition of heat, we can place the word *Heat* on the right side to indicate that the addition of heat will make the reaction proceed from right to left. A temperature stress then becomes like a concentration stress in which the concentration of heat either increases or decreases.

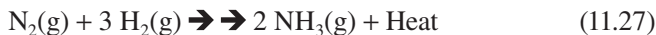
Let us look again at the Haber process. The reverse reaction of the Haber process is favored by adding heat. Thus we place the word *Heat* on the right side of Equation 11.22.



Heat is released as the reaction proceeds from left to right. The reaction is exothermic. Increasing the temperature of the reaction vessel (the addition of heat) causes a shift from right to left to partially consume the heat that was added, as indicated here.



Decreasing the temperature of the reaction vessel (the removal of heat) causes a shift from left to right to replace the heat that was removed.



In each case, the system returns to the equilibrium state of Equation 11.22.

While the word *Heat* can be added to one side or the other, depending on which reaction is favored by heat, it is important to recognize that this is only a gimmick that helps us determine what happens during the temperature stress. Of course, the word *Heat* does not appear in the equilibrium constant expression; the molar concentration of heat does not have any meaning. This means that shifts caused by temperature stresses do cause the equilibrium constant to change. The concentrations change during a temperature stress in a way in which the constancy of K_{eq} cannot be maintained. Thus all equilibrium constants are temperature dependent; they do change with temperature.

Perhaps a better way to show that heat is either consumed or generated is to indicate the enthalpy change that is involved rather than to just use the word *Heat* on one side or the other of the equation. **Enthalpy change**, symbolized as ΔH , is defined as the **heat of reaction** at constant pressure. It is a number expressed in energy units, such as joules (J) or kilojoules (kJ). If the value of ΔH is negative, the reaction is exothermic (analogous to *Heat* being placed on the right of the equation). If ΔH is positive, the reaction is endothermic (analogous to *Heat* being placed on the left of the equation). The Haber process, for example, is exothermic and can be written as follows with the ΔH symbolism:



This indicates that 92.22 kJ of heat are released when 3 mol of hydrogen react with 1 mol of nitrogen to form 2 mol of ammonia.

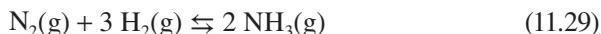
11.10 Effect of Pressure Change

The final stress we will consider is that caused by either an increase or a decrease in the pressure within the reaction chamber. This pressure stress is akin to the pressure changes discussed for systems of gases in Chapter 9. In fact, within reasonable limits, the only time a pressure change is usually a stress on an equilibrium system is when one or more gases are involved in the equilibrium. If no gases are involved, then the equilibrium does not shift in either direction.

Pressure stresses can be thought of as either pushing down or pulling up on the piston in the piston-cylinder arrangement in Chapter 9 when the equilibrium reaction is occurring within a volume enclosed by the piston and cylinder. Pushing down on the piston means a pressure increase in the system, while pulling up on the piston means a pressure decrease in the system.

If the stress is an increase in the pressure, the system responds so as to decrease it again. If the stress is a decrease in the pressure, the system responds so as to increase it again. For the system to respond by decreasing the pressure, the equilibrium must shift in the direction that would produce less gas—less gas means lower pressure. For the system to respond by increasing the pressure, the shift must be in the direction that would produce more gas—more gas means higher pressure. To predict which shift would produce more or less gas, we need to determine which side contains more or less gas. This can easily be done by looking at the equation. First, we would need to express which chemicals are gases, such as with the (g) designation. Second, we must balance the equation because the balancing coefficients for those designated with the (g) would tell us how many moles of each gas are involved. We can then determine which side of the equation consists of more or less gas.

Let us look again at the Haber process as an example.



All the reactants and products in this reaction are gases. There are four moles of gas on the left and two moles of gas on the right. If the pressure is increased, the equilibrium would shift toward the side where there is less gas. This relieves the stress by decreasing the pressure due to less gas being present. This means that the equilibrium will be a shift to the right in this case.



If the pressure is decreased, the equilibrium would shift toward the side where there is more gas. This relieves the stress by increasing the pressure due to more gas being present. This means that the equilibrium shifts to the left in this case.



In each case, the system returns to the equilibrium state of Equation 11.29.

If an equilibrium reaction has an equal numbers of moles of gas on both sides of the equation, then there would be no shift in either direction and the pressure change would have no effect.

11.11 Effect of Catalysts

A **catalyst** is a material that is present in the reaction vessel in order to increase or decrease the rate of the reaction, or the speed by which the reaction proceeds. A catalyst does not play any other role. Thus it does not appear in the equation of the reaction (except sometimes over the arrow, as was shown in Chapter 8, Section 8.3). As far as equilibria are concerned, the presence of a catalyst does not affect the position of the equilibrium at all, meaning that adding a catalyst does not impose a stress on the system and the equilibrium does not shift in either direction. It only affects the speed with which the equilibrium is reached.

11.12 Homework Exercises

Basic Concepts

- (a) Why do we speak of equilibrium when we define vapor pressure?
(b) Why do we speak of equilibrium when we describe what happens to a solute in a container with two immiscible liquid solvents?
- What does it mean when we say that a chemical reaction is reversible?
- What does a double arrow separating the reactants and products in a chemical equation mean? (To simply say “equilibrium” is not sufficient.)
- What is chemical equilibrium?
- When sulfur dioxide gas (SO_2) is mixed with oxygen gas (O_2), a reaction occurs and sulfur trioxide gas (SO_3) forms. After a certain length of time, this reaction reverses, which means that the sulfur trioxide decomposes to form oxygen gas and sulfur dioxide, and these two reactions then occur at the same time and rate. Write a single chemical equation that would describe this phenomenon.

Equilibrium Constant

- Describe what is meant by an equilibrium constant. Provide one specific example.
- Write the equilibrium constant expression for each of the following equilibria.
 - $2 \text{NO}(\text{g}) + \text{Cl}_2(\text{g}) \rightleftharpoons 2 \text{NOCl}(\text{g})$
 - $\text{N}_2(\text{g}) + \text{O}_2(\text{g}) \rightleftharpoons 2 \text{NO}(\text{g})$

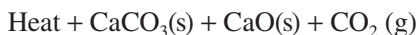
- c. $2 \text{Cl}_2(\text{g}) + 2 \text{H}_2\text{O}(\text{g}) \rightleftharpoons 4 \text{HCl}(\text{g}) + \text{O}_2(\text{g})$
 - d. $\text{H}_2\text{SO}_3(\text{aq}) \rightleftharpoons \text{H}^+(\text{aq}) + \text{HSO}_3^{2-}(\text{aq})$
 - e. $\text{H}_3\text{PO}_4(\text{aq}) \rightleftharpoons \text{H}^+(\text{aq}) + \text{H}_2\text{PO}_4^-(\text{aq})$
 - f. $2 \text{SO}_2(\text{g}) + \text{O}_2(\text{g}) \rightleftharpoons 2 \text{SO}_3(\text{g})$
8. What type of general reaction is described by an ionization constant (K_i)? Give an example.
 9. What type of general reaction is described by a solubility product constant (K_{sp})? Give an example.
 10. Write equilibrium reactions and solubility product constant expressions for the following ionic compounds dissolving in water.
 - a. Ag_2S
 - b. PbI_2
 - c. Bi_2S_3
 11. The acid HCN ionizes only very slightly when dissolved in water. Is the equilibrium constant a very small number, a number about equal to 1, or a very large number? Explain.
 12. The numerical value of the equilibrium constant for the reaction in question 7e is 0.0075. What does this say about the concentration of the H_2PO_4^- in this equilibrium compared with the concentration of the H_3PO_4 ?
 13. Consider the following equilibrium:



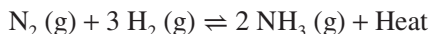
- a. Write the equilibrium constant expression for this reaction.
 - b. Which of the following symbols applies: K_i , K_a , K_b , or K_{sp} ? Explain.
 - c. If the value of the equilibrium constant is 5.8×10^{-10} , would you say that H_3BO_3 is highly ionized or only slightly ionized? Explain.
14. The equilibrium constant for the equilibrium in Exercise 7d is 1.5×10^{-2} . In this reaction, is the concentration of H^+ at equilibrium large or small? Explain.
 15. HCHO is a weak acid and a weak electrolyte.
 - a. Write an equation representing the ionization that occurs when HCHO dissolves in water.
 - b. Write the equilibrium constant expression for this ionization.
 - c. Would the K_a for this weak acid be a large number, a small number, or a number close to 1.0? Explain.

Le Chatelier's Principle

16. What is Le Chatelier's principle?
17. Why is Le Chatelier's principle important in the description of chemical reactions?
18. What is meant by an endothermic reaction?
19. How is an endothermic reaction affected by the addition of heat?
20. What is meant by an exothermic reaction?
21. How is an exothermic reaction affected by the addition of heat?
22. Given the following equilibrium, fill in the blanks with the words *increase*, *decrease*, or *no change*.

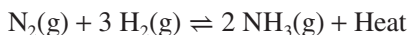


- a. If the temperature were decreased, the amount of CaO present would _____.
 - b. If the pressure were increased, the amount of CaO present would _____.
 - c. If the amount of CaCO_3 present were increased, the amount of CaO would _____.
23. Consider the following reaction:



Tell in what direction (left or right) the equilibrium will shift and what will happen to the concentration of N_2 (increase or decrease) when (a) some H_2 is added, (b) the pressure is decreased, and (c) the temperature is increased.

24. Which of the examples in Exercise 7 would be affected by a change in pressure? Which would an increase in pressure not affect?
25. Consider the following equilibrium.



In which direction will the equilibrium shift (right, left, or neither) and what will happen to the concentration of H_2 (increase, decrease, or no change) when each of the following occurs?

- a. The reaction vessel is heated.
- b. The pressure is increased.
- c. More N_2 is added to the reaction vessel.

26. Consider the following equilibrium:



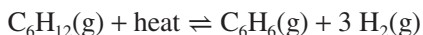
In which direction will the equilibrium shift (right, left, or neither) and what will happen to the concentration of HF (increase, decrease, or no change) when each of the following occurs?

- The concentration of F_2 is decreased.
 - The pressure is increased.
 - The reaction vessel is cooled.
27. Describe what is meant by the term *enthalpy change*?
28. Consider the following equilibrium:



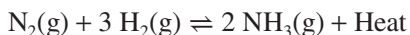
In which direction will the equilibrium shift (right, left, or neither) and what will happen to the concentration of H_2 (increase, decrease, or no change) when each of the following occurs?

- The reaction vessel is heated.
 - The pressure is decreased.
 - More O_2 is added to the reaction vessel.
29. Consider the following equilibrium:



In which direction will the equilibrium shift (right, left, or neither) and what will happen to the concentration of H_2 (increase, decrease, or no change) when each of the following occurs?

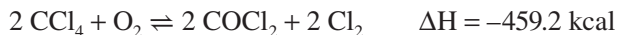
- The concentration of C_6H_6 is decreased.
 - The pressure is increased.
 - The reaction vessel is cooled.
30. Consider the following equilibrium:



In which direction will the equilibrium shift (right, left, or neither) and what will happen to the concentration of NH_3 (increase, decrease, or no change) when each of the following occurs?

- The reaction vessel is cooled.
- The pressure is decreased.
- More H_2 is added to the reaction vessel.

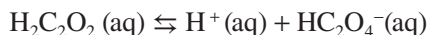
31. Consider the following equilibrium reaction:



Indicate (in the right column of the following table) what will happen to the concentration of COCl_2 (increase, decrease, no change, or not enough data given) in each of scenario given.

(a)	Pressure is decreased.	
(b)	Catalyst is added.	
(c)	Temperature is increased.	
(d)	Concentration of O_2 is decreased.	

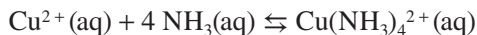
32. Consider the following equilibrium reaction:



The reaction is exothermic. Fill in the table with either *yes* or *no* to tell whether the indicated change will result in an increase in the concentration of H^+ in the solution.

(a)	Pressure is increased.	
(b)	Catalyst is added.	
(c)	Temperature is decreased.	
(d)	Concentration of $\text{C}_2\text{O}_4^{2-}$ is increased.	

33. Consider the following equilibrium reaction:

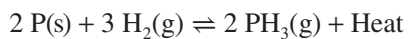


The reaction is endothermic. Indicate in the right column of the following table what direction the equilibrium will shift (left, right, neither, or not enough information given) to tell what will happen with each scenario given.

(a)	Pressure is increased.	
(b)	Catalyst is added.	
(c)	Temperature is decreased.	
(d)	Concentration of NH_3 is increased.	

34. If you were managing an industrial process in which PH_3 is made by the following reaction and you wanted to increase the output of this

product (shift the equilibrium to the right) so that you could make more money, what would you do in terms of a high or low concentration of H_2 , high or low pressure, and high or low temperature? Explain.



Chapter 12

Acids, Bases, and pH

Introduction	304
Formulas and Strengths of Acids and Bases	304
General Properties of Acids and Bases	308
Acidity Level: pH.	310
Measurement of pH	318
Theories of Acids and Bases	319
Reactions Involving Acids and Bases	321
“Add Acid to Water” Rule	324
Homework Exercises	324

12.1 Introduction

We have noted several times in previous chapters that more than 13 million chemical compounds exist. Because of this, the study of the chemistry (structure, composition, and properties) of these compounds can indeed be a formidable task. To simplify this task, the usual procedure is to divide and subdivide the compounds into various classifications and then to study individual classifications separately. This method will be especially evident in Chapter 14.

We have many different classifications of compounds. We have ionic and covalent compounds, inorganic and organic compounds, and soluble and insoluble compounds. In this chapter, we will study acids and bases. See Figure 12.1.

12.2 Formulas and Strengths of Acids and Bases

We recognize that a compound is an **acid** when it has one or more hydrogens listed first in its formula. For example, we have the well-known laboratory acids that are listed Table 12.1. In addition, there is a host of others that are less well known. Some of these are listed in Table 12.2.

The first four acids listed the Table 12.1 and all of the acids listed in Table 12.2 are **inorganic acids**, sometimes also known as **mineral acids**. Acetic acid, ascorbic acid, and acetylsalicylic acid in Table 12.1 are examples of **organic acids**. They are carboxylic acids (Chapter 14). Carboxylic acids contain the carboxyl group, which is drawn in Figure 12.2. Also in Figure 12.2 are the structure and formula of acetic acid. The hydrogen atom bonded to the oxygen in this structure is the hydrogen listed first in the formula. It is the

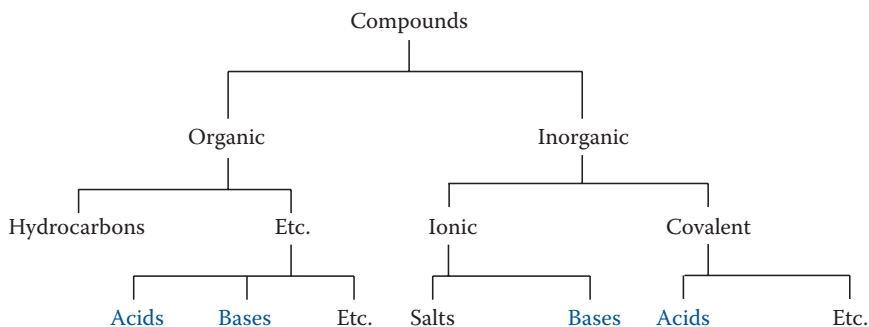


FIGURE 12.1 Acids and bases are two more classifications of compounds that we can study. They can be organic or inorganic.

Table 12.1 Formulas and Names of Well-Known Acids

Formula	Name
HCl	Hydrochloric acid
H ₂ SO ₄	Sulfuric acid
HNO ₃	Nitric acid
H ₃ PO ₄	Phosphoric acid
HC ₂ H ₃ O ₂	Acetic acid
HC ₆ H ₇ O ₆	Ascorbic acid
HC ₉ H ₇ O ₄	Acetylsalicylic acid

Table 12.2 Formulas and Names of Lesser-Known Acids

Formula	Name
HF	Hydrofluoric acid
HBr	Hydrobromic acid
HI	Hydroiodic acid
H ₂ CO ₃	Carbonic acid
HNO ₂	Nitrous acid
H ₂ SO ₃	Sulfurous acid
H ₂ S	Hydrosulfuric acid
HCN	Hydrocyanic acid

hydrogen that makes acetic acid, or any other carboxylic acid, an acid. This is the so-called acidic hydrogen in the formula. The other hydrogens are not acidic.

To be acidic, a hydrogen atom must be held relatively loosely so that it is released under certain conditions. As we saw in Chapter 10, when acids dissolve in water, the acidic hydrogen can break away from the rest of the molecule, creating a solution that has ions in it. This ionization reaction for hydrochloric acid is as follows:



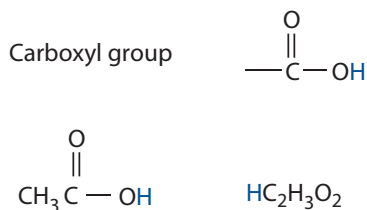
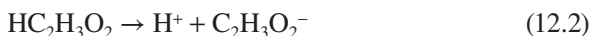


FIGURE 12.2 A structural feature common to many organic acids is the carboxyl group shown here on the upper right. The hydrogen that is released when these acids dissolve in water (the acidic hydrogen) is the hydrogen bonded to the oxygen. For example, in acetic acid, the structure and formula for which is given in the lower part of this figure, there are other hydrogens, but these are not acidic. The acidic hydrogen is written first in the formula. See Chapter 14 for more information.

These hydrogens may also be released when the acid enters into some chemical reactions, such as when it reacts with a base or with an active metal. Hydrogens that are not acidic are tightly held and are not released in these reactions. Examples of these are the middle hydrogens of the formula for acetic acid.



Acids are often classified as weak or strong. **Strong acids** are those that ionize completely when dissolved in water (Chapter 10). Strong acids are particularly corrosive and damaging to metals and organic materials, such as plant and animal tissue. Hydrochloric acid, sulfuric acid, and nitric acid are three strong acids. Special precautions are required when handling them, especially when they are at high concentration levels. **Weak acids** are those that do not ionize completely when dissolved in water. Although they can attack some metals and some organic materials and thus may require some special handling precautions, weak acids are, with some exceptions, not considered to be as dangerous as strong acids. Acetic acid is a weak acid, as are most of the other acids listed in Tables 12.1 and 12.2. Most common strong and weak acids are sold as concentrated solutions.

There are a few compounds that, while they do not have hydrogen first in their formulas, do release hydrogen ions under certain conditions. These are the salts of acids that have more than one hydrogen in their formulas and in which only some of the hydrogens have been replaced by metals.

Recall the discussion of nomenclature in Chapter 3. Examples of acids with more than one hydrogen in their formulas include sulfuric acid (H_2SO_4), phosphoric acid (H_3PO_4), and carbonic acid (H_2CO_3). Examples of salts in which only some of the hydrogens have been replaced (and therefore are capable of releasing hydrogens) are NaHSO_4 , NaH_2PO_4 , Na_2HPO_4 , and NaHCO_3 .

A comment about water (H_2O) is appropriate here. While water has hydrogen first in the formula, we do not usually think of it as an acid. A small fraction of the water molecules in a sample of pure distilled water do release hydrogen ions, however, and we will return to this phenomenon as our discussion proceeds.

A compound is a **base** when there is one or more hydroxide ions in the formula. Some of the common bases are listed in Table 12.3. The behavior of bases with their hydroxide ions is similar to the behavior of acids with their hydrogen ions. Hydroxide ions are easily released under certain conditions. With sodium hydroxide, for example, we can write



to symbolize the ionization behavior when the base dissolves in water. Similarly, hydroxide ions are released when the base reacts with an acid.

Bases are also described as strong and weak. Sodium hydroxide and potassium hydroxide fit into the category of **strong bases** and, at high concentration levels, are very corrosive and damaging to metals and plant and animal tissue. Again, special handling procedures are in order. Contrary to strong acids, these bases are normally purchased as solid chemicals.

Table 12.3 Formulas and Names of Hydroxide Bases

Formula	Name
NaOH	Sodium hydroxide
NH_4OH	Ammonium hydroxide
KOH	Potassium hydroxide
$\text{Ca}(\text{OH})_2$	Calcium hydroxide
$\text{Ba}(\text{OH})_2$	Barium hydroxide
$\text{Mg}(\text{OH})_2$	Magnesium hydroxide

Ammonium hydroxide is a good example of a **weak base**, meaning that it only ionizes slightly when dissolved in water.



It is usually purchased as a concentrated solution like the acids. This concentrated solution smells very strongly of ammonia (NH_3) because there is a second equilibrium reaction, in addition to ionization, that occurs in ammonium hydroxide–water solutions.



Thus these solutions are saturated solutions of ammonia. Because of this, solutions of ammonia are also solutions of ammonium hydroxide, and vice versa.

12.3 General Properties of Acids and Bases

The most recognizable physical property associated with acids is their sour taste. The common household product vinegar is approximately 5% acetic acid and has a sour taste. A common property of citrus fruits (e.g., oranges, grapefruit, lemons, pineapples) is their sour taste due to the presence of citric acid, which is an organic tricarboxylic acid. Its structure is given in Figure 12.3.

Bases have a bitter taste and a slippery, soapy feel. Often the terms *alkali*, *alkaline*, *lye*, and *caustic* are associated with bases. Sodium hydroxide and potassium hydroxide have, for example, been referred to as caustic soda and caustic potash, respectively. The reason is that these compounds are caustic, corrosive compounds that can chemically attack a variety of materials, including plant and animal tissue. These strong bases are used in the production of soaps.

Even to a person not educated in the properties of acids, the word *acid* conjures up thoughts of the possibility of serious chemical burns on skin. Acids in

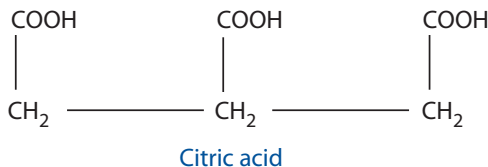


FIGURE 12.3 The structure of citric acid, the acid that gives citrus fruits their sour taste.

general are quite harmful to skin and must be handled with extreme caution. If an acid does contact skin, we often feel a burning sensation quickly, indicating that rinsing with water and that possibly some neutralizing treatment is in order. Bases can be equally harmful to the skin, and the problem is more acute in this case because we often do not feel the burning sensation as quickly. It should be emphasized that while one should never handle any chemical, liquid or solid, directly with hands and fingers, it is especially true of solid acids and bases because of the potential for chemical burns.

The chemical properties that acids and bases exhibit are mostly due to the easy release of hydrogens (in the case of acids) and hydroxides (in the case of bases). The reason we study acids and bases together is that hydrogen ions and hydroxide ions react with each other to form water.



Thus, whenever hydrogen ions are released, hydroxide ions are consumed, creating **acidic conditions** in the solution, meaning that hydrogen ions predominate over the hydroxide ions. Whenever hydroxide ions are released, hydrogen ions are consumed, creating **basic conditions** in the solution, meaning that hydroxide ions predominate over hydrogen ions.

The presence of excess H^+ or excess OH^- ions in a solution gives the solution certain colorful chemical properties. First, an excess of hydrogen ions in a solution causes litmus dye to turn pink. An excess hydroxide ions causes litmus dye to turn blue. Paper saturated with litmus dye, known as **litmus paper**, is often used as an indicator of acidic and basic conditions. There are a large number of other **indicators** that change color under various acidic and basic conditions. Phenolphthalein solutions are pink under basic conditions and colorless under acidic conditions. Methyl orange solutions are yellow under basic and slightly acidic conditions and red under very acidic conditions. It is possible to mix various indicators and thus prepare an indicator that will change colors at a number of levels of acidity and basicity. Again we can saturate paper with this indicator mixture and have what is called **pH paper**, which is paper that is saturated with a mixture of various indicators and can thus detect specific acidity and basicity levels in a solution simply by displaying different colors. Litmus paper, pH paper, and the other indicators are commonly used to determine the acidity level of a solution (i.e., whether hydrogen ions or hydroxide ions are predominating and to what extent). The acidity level is commonly expressed as the **pH** of the solution. The concept

of pH is discussed in Section 12.4, and the use of pH meters to measure it is examined in Section 12.5.

In addition to the properties of acids and bases that affect the color of other chemicals, there are other important chemical properties of these compounds. One is **neutralization**, which is the negation of the chemical properties of acids by their reaction with bases, and vice versa. In other words, it is the reaction of acids with bases. A base can be neutralized by an acid and an acid can be neutralized by a base. The reason that neutralization occurs is that hydrogen ions react with hydroxide ions to form water, as we discussed earlier. Water is one product of such a reaction, while salt is the other. This will be discussed in more detail in Section 12.7.

12.4 Acidity Level: pH

For a variety of reasons, it is very important to be able to express the acidity level of solutions. We now proceed to define the common method for doing this, the concept of pH. As we shall see, this method has applications in a chemistry lab and in our homes and environment.

12.4.1 Water Ionization and the Ion Product Constant

The fact that a small number of water molecules in pure distilled water ionize was mentioned in Section 12.2. This ionization is an equilibrium and can be written as follows:



This equilibrium lies far to the left (indicating that very little ionization has taken place). This means that the concentrations of hydrogen ions and hydroxide ions are very small. At 25°C, $[\text{H}^+] = [\text{OH}^-] = 1.00 \times 10^{-7} \text{ M}$. A special equilibrium constant, K_w , known as the ion product constant for water, is based on this ionization and is defined as follows at 25°C.

$$K_w = [\text{H}^+] [\text{OH}^-] = 1.008 \times 10^{-14} \quad (12.8)$$

Because it is an equilibrium, the addition of acids or bases to water shifts the position of this equilibrium to the left. The removal of acids or bases through neutralization reactions shifts the position of the equilibrium to the right. The addition of acids increases the $[\text{H}^+]$, and the resulting shift causes $[\text{OH}^-]$ to decrease. The addition of bases (hydroxides) increases $[\text{OH}^-]$, and the resulting shift causes

$[\text{H}^+]$ to decrease. The removal of acids decreases $[\text{H}^+]$, and the resulting shift causes $[\text{OH}^-]$ to increase. The removal of hydroxide bases decreases $[\text{OH}^-]$, and the resulting shift causes $[\text{H}^+]$ to increase. Thus, when $[\text{H}^+]$ increases, $[\text{OH}^-]$ decreases, and when $[\text{OH}^-]$ increases, $[\text{H}^+]$ decreases. However, like equilibrium constants, the multiplication product of $[\text{H}^+]$ and $[\text{OH}^-]$, or K_w , does not change. So, knowing either $[\text{H}^+]$ or $[\text{OH}^-]$, one can calculate the other from Equation 12.8. We will assume 25°C for all calculations involving K_w .

Example 12.1

What is the $[\text{OH}^-]$ in a solution in which the $[\text{H}^+]$ is $5.2 \times 10^{-3} \text{ M}$?

Solution 12.1

Rearranging Equation 12.8, we have the following:

$$[\text{OH}^-] = \frac{K_w}{[\text{H}^+]} = \frac{1.008 \times 10^{-14}}{5.2 \times 10^{-3}} = 1.9 \times 10^{-12} \text{ M}$$

12.4.2 pH

Much of the chemistry that occurs in water solution depends on what the acid content, or acidity level of the solution is. For example, a procedure for the determination of caffeine in soda pop calls for a solution that has a $[\text{H}^+]$ of $1.00 \times 10^{-3} \text{ M}$. A procedure for the extraction of potassium from soil requires a solution that has a $[\text{H}^+]$ of $1.00 \times 10^{-7} \text{ M}$. Indeed, many activities in our everyday lives depend on the acidity of solutions. Gardeners who grow rhododendrons, for example, should use soil that has a $[\text{H}^+]$ between about $1 \times 10^{-4} \text{ M}$ and $3 \times 10^{-6} \text{ M}$. The fluid in our stomachs should have $[\text{H}^+]$ of between $1 \times 10^{-1} \text{ M}$ and $1 \times 10^{-3} \text{ M}$. Tropical fish in an aquarium survive only if the $[\text{H}^+]$ is between $1 \times 10^{-6} \text{ M}$ and $1 \times 10^{-8} \text{ M}$. Fish in our lakes and rivers have been threatened by acid rain. The $[\text{H}^+]$ in lakes and rivers should be above $1 \times 10^{-5} \text{ M}$ if fish and other aquatic life are to survive. If the $[\text{H}^+]$ in our blood is not within a very narrow range (1.00×10^{-7} – $1.58 \times 10^{-8} \text{ M}$), we die.

If the $[\text{H}^+]$ in a solution is too high, a base may be added to decrease it. If the $[\text{H}^+]$ is too low, an acid may be added to increase it. The adjustment of the acidity level in real-world solutions, such as those just mentioned, has often been performed by persons with no formal training in acid–base chemistry. For example, gardeners often adjust the acidity level of soil by

applying lime (CaO) to the soil. Individuals ill with such conditions as indigestion or heartburn can adjust the $[H^+]$ of their stomach fluid by taking an antacid medication. The nonchemist is likely not to have an understanding of scientific notation nor of the often extremely small number by which $[H^+]$ is typically expressed. They may also not be comfortable with the numbers in a way that would allow them to solve problems related to acidity level.

For these reasons, and for their own convenience, scientists devised a much simpler method of expressing $[H^+]$. It is called the pH scale. The pH scale expresses the acidity level of a solution in easily understood and manageable numbers. The pH of a solution is represented by the negative logarithm of the $[H^+]$.

$$\text{pH} = -\log [H^+] \quad (12.9)$$

12.4.3 Logarithms

The **logarithm of a number** is the power to which 10 must be raised to give the number. For example, the power to which 10 must be raised to give the number 100 is 2. Thus the logarithm of 100 is 2. Similarly, the power to which 10 must be raised to give the number 1000 is 3. Thus the logarithm of 1000 is 3. See Table 12.4 for other examples.

All positive numbers have a logarithm, not just those that are multiples of 10. The power to which 10 must be raised in most cases is not a whole number like it is for the multiples of 10 in Table 12.4. For example, the logarithm of 2 is 0.30103. The logarithm of 3 is 0.4771213. The logarithm of 5.11932 is 0.7092123 (Table 12.5).

Table 12.4 Examples of Easily Determinable Logarithms	
$10^2 = 100$	$\log 100 = 2$
$10^3 = 1000$	$\log 1000 = 3$
$10^{-2} = \frac{1}{10^2} = \frac{1}{100} = 0.01$	$\log 0.01 = -2$
$10^{-3} = \frac{1}{10^3} = \frac{1}{1000} = 0.001$	$\log 0.001 = -3$

Table 12.5 Examples of Logarithms of Numbers That Are Not Multiples of 10

$10^{0.30103} = 2$	$\log 2 = 0.30103$
$10^{0.4771213} = 3$	$\log 3 = 0.4771213$
$10^{0.7092123} = 5.11932$	$\log 5.11932 = 0.7092123$

Equation 12.9 indicates that pH is the *negative* logarithm of $[H^+]$. The negative logarithm of a number is the logarithm multiplied by -1 . The $[H^+]$ is typically an extremely small positive number. All positive numbers have a logarithm no matter how small the number is. Thus even the extremely small numbers that are the hydrogen ion concentrations have logarithms (and negative logarithms), and the negative logarithms are easily understood and manageable numbers. For example, the $[H^+]$ in blood is typically 4.0×10^{-8} M, or 0.000000040 M. The negative logarithm of 0.000000040 is 7.40.

The complete scale of pH values in water ranges from 0 to 14, with $\text{pH} = 7$ being the neutral point represented by pure, distilled water or by any strong acid solution that has been exactly neutralized by a strong base. In pure, distilled water, or in a neutralized water solution, the $[H^+]$ and the $[OH^-]$ are equal to 1.0×10^{-7} M. The pH of a neutral solution, such as pure, distilled water, is thus 7.00. Any pH values between 0 and 7 indicate an acidic solution, one in which hydrogen ions predominate over hydroxide ions. Any pH values between 7 and 14 indicate a basic solution, one in which hydroxide ions predominate over hydrogen ions. We see from this that the relationship between $[H^+]$ and pH is an inverse one; as $[H^+]$ increases, pH decreases, and vice versa. A low pH indicates a high $[H^+]$ (compared with $[OH^-]$), while a high pH indicates a low $[H^+]$.

How is the logarithm of a number determined? Since every positive number has a logarithm specific to that number, tables of logarithms have been created, and logarithms can be found by looking at these tables. However, logarithms can also be determined by using the LOG key on a scientific calculator. To determine the pH of a solution given a hydrogen ion concentration that is expressed in scientific notation, you will need to learn the required keystrokes on your calculator. Here is an example for practice: What is the pH of a solution that has an $[H^+]$ of 4.59×10^{-5} M? The answer is 4.338.

It is also important to be able to calculate the $[H^+]$ given the pH. Here is an example of this for practice: If the pH of a solution is 3.51, what is the $[H^+]$? The answer is 3.1×10^{-4} M. Additional examples follow.

Example 12.2

What is the pH of a solution that has an $[H^+]$ of 2.91×10^{-4} M?

Solution 12.2

$$\begin{aligned}\text{pH} &= -\log [H^+] \\ &= -\log (2.91 \times 10^{-4}) \\ &= 3.536\end{aligned}$$

Example 12.3

What is the $[H^+]$ in a solution if the pH of the solution is 9.551?

Solution 12.3

$$\begin{aligned}\text{pH} &= -\log [H^+] \\ 9.551 &= -\log [H^+] \\ [H^+] &= 2.81 \times 10^{-10} \text{ M}\end{aligned}$$

Sometimes it may be necessary to calculate the pH of a solution when the concentration of an acid is given, but not the $[H^+]$. If the acid is a strong acid, it completely ionizes in water solution, and thus the acid concentration and the $[H^+]$ are the same because the number of hydrogen ions present equals the number of acid molecules that are dissolved.

Example 12.4

What is the pH of a 0.0283 M solution of HCl?

Solution 12.4

HCl is a strong acid that completely ionizes when it dissolves in water.



Thus the concentration of HCl (0.0283) is also the $[H^+]$.

$$\text{pH} = -\log [H^+] = -\log (0.0283) = 1.548$$

If the acid in the problem is a weak acid, the problem is more involved since the $[H^+]$ cannot be determined directly. The determination of $[H^+]$ in that case can be calculated by using the equilibrium constant expression for weak acid ionization. The pH is then calculated from the $[H^+]$ using the method shown in Solution 12.4. This type of problem will not be discussed further here.

12.4.4 Significant Figures Rule

The rule of significant figures when determining the logarithm of a number is as follows: the number of significant figures after the decimal point in the logarithm of a number is the same as the total number of significant figures in the number. You should notice that this rule was used in the examples in Section 12.4.3.

12.4.5 pOH

The hydroxide ion concentration may also be expressed in easily manageable numbers by defining the pOH in a way similar to pH.

$$\text{pOH} = -\log [OH^-] \quad (12.10)$$

Thus pOH and $[OH^-]$ may be calculated in the same way as pH and $[H^+]$ in the previous examples.

Example 12.5

What is the pOH if the $[OH^-]$ is 7.19×10^{-12} M?

Solution 12.5

$$\begin{aligned} \text{pOH} &= -\log [OH^-] \\ &= -\log (7.19 \times 10^{-12}) \\ &= 11.143 \end{aligned}$$

Example 12.6

What is the $[OH^-]$ if the pOH is 6.44?

Solution 12.6

$$\begin{aligned}\text{pOH} &= -\log [\text{OH}^-] \\ 6.44 &= -\log [\text{OH}^-] \\ [\text{OH}^-] &= 3.6 \times 10^{-7} \text{ M}\end{aligned}$$

12.4.6 Relationship between pH and pOH

Beginning with Equation 12.8, we can now derive a new equation that simplifies some of the calculations involving the water ionization process. This derivation recognizes the logarithmic identity that when

$$A = B \times C \quad (12.11)$$

the logarithm of A is the sum of the logarithms of B and C.

$$\log A = \log B + \log C \quad (12.12)$$

Also, the negative logarithm of A is the sum of the negative logarithms of B and C,

$$-\log A = -\log B - \log C \quad (12.13)$$

so that Equation 12.8 can be rewritten as follows:

$$-\log K_w = -\log [\text{H}^+] - \log [\text{OH}^-] \quad (12.14)$$

or

$$14.00 = \text{pH} + \text{pOH} \quad (12.15)$$

Example 12.7

What is the pH of a solution that has a pOH of 5.39?

Solution 12.7

$$\begin{aligned}\text{pH} + \text{pOH} &= 14.00 \\ \text{pH} &= 14.00 - \text{pOH} \\ &= 14.00 - 5.39 = 8.61\end{aligned}$$

12.4.7 Calculations Summary

We have learned four formulas in this section that can be used for the interconversion of $[\text{H}^+]$, $[\text{OH}^-]$, pH, and pOH. Given one of these four

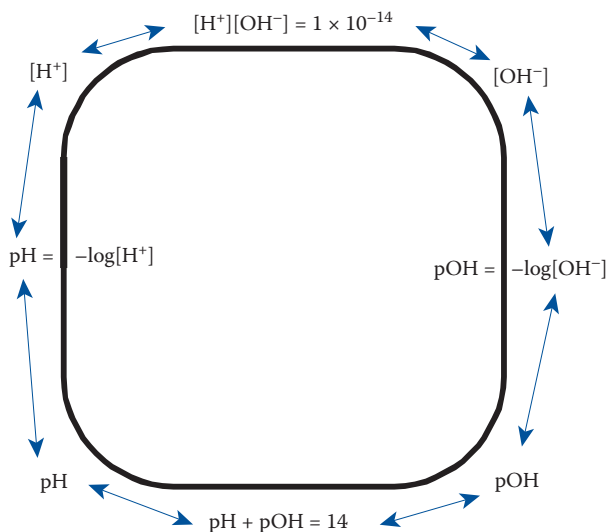


FIGURE 12.4 Different pathways for interconverting $[H^+]$, $[OH^-]$, pH, and pOH.

parameters, any of the other three may be calculated in two different ways. Figure 12.4 shows the different pathways possible. An example is worked here.

Example 12.8

What is the pH of a solution if the $[OH^-]$ is 6.3×10^{-4} M?

Solution 12.8a

One path for this is to calculate the pOH and then subtract from 14.00.

$$\begin{aligned} \text{pOH} &= -\log [OH^-] \\ &= -\log (6.3 \times 10^{-4}) = 3.20 \\ \text{pH} + \text{pOH} &= 14.00 \\ \text{pH} &= 14.00 - \text{pOH} = 10.80 \end{aligned}$$

Solution 12.8b

Another path is to calculate $[H^+]$ using the K_w expression and then take the negative logarithm.

$$[\text{H}^+][\text{OH}^-] = 1.008 \times 10^{-14}$$

$$[\text{H}^+] = \frac{K_w}{[\text{OH}^-]} = \frac{1.008 \times 10^{-14}}{6.3 \times 10^{-4}} = 1.6 \times 10^{-11} \text{ M}$$

$$\begin{aligned}\text{pH} &= -\log [\text{H}^+] \\ &= -\log (1.6 \times 10^{-11}) = 10.80\end{aligned}$$

12.5 Measurement of pH

In Section 12.3, we discussed the use of litmus paper, pH paper, and indicators as means for determining the acidity level, or pH, of a solution. While these have been widely used and are convenient and inexpensive for many applications, they do not provide the laboratory worker with the precision that is often needed. The use of pH paper is the most precise of the three, but it can only determine pH to within 1 pH unit. Often the pH of a solution needs to be measured to the first or second decimal place.

A much more accurate and precise method of measuring pH has been developed. It utilizes what is called a **pH electrode** and a **pH meter**. The pH electrode is a small glass or plastic enclosure—the size of a small test tube—that is electrically sensitive to the hydrogen ions in a solution. When this electrode is immersed into the solution to be tested, an electrical potential or voltage is produced across the glass membrane at the tip that separates the interior of the electrode from the solution. This voltage can be measured with a simple voltage measuring device, a voltmeter. The concentration of the hydrogen ions in the solution determines the voltage level that develops, and thus this voltage is proportional to $[\text{H}^+]$. Since pH is also proportional to $[\text{H}^+]$, it is possible to relate this voltage directly to the pH of the solution. The pH meter is a voltmeter that utilizes an electrical circuit that automatically converts the voltage level to pH and displays the pH on its readout. Thus, simply stated, the pH of a solution can be measured or monitored by immersing this pH probe, connected to the pH meter, into the solution of interest and reading the pH on the readout of the meter.

Since the meter is a sort of modified voltmeter, the measurement is a relative measurement. This means that the voltage developed at the tip of the probe is measured relative to some other voltage. This other voltage is the

voltage of a **reference electrode**. Modern pH probes have a reference electrode built in and are called **combination pH electrodes**. The interior design of these is beyond our scope.

12.6 Theories of Acids and Bases

12.6.1 Arrhenius Theory

The view of acids and bases as we have been discussing them thus far is known as the **Arrhenius theory**, named after Svante Arrhenius, a Swedish scientist who introduced this theory in 1884. According to this theory, an acid is defined as a substance that releases hydrogen ions when dissolved in water. A base is defined as a substance that releases hydroxide ions when dissolved in water. All Arrhenius acids have the symbol for hydrogen first in the formula. See Tables 12.1 and 12.2 for examples. All Arrhenius bases have hydroxide ions in the formula. All bases listed in Table 12.3 are Arrhenius bases.

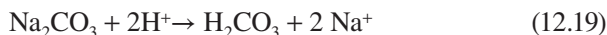
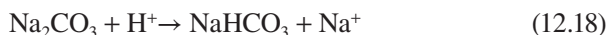
12.6.2 Bronsted–Lowry Theory

There are substances other than Arrhenius bases that turn litmus paper blue, feel slippery in water solution, or neutralize acids. In 1923, Johannes Bronsted, a Danish chemist, and Thomas Lowry, an English chemist, presented a broader definition of acids and bases. The **Bronsted–Lowry theory** states that an acid is a proton donor and a base is a proton acceptor. This definition of an acid is similar to that in the Arrhenius theory (a hydrogen ion is a proton), and thus all Arrhenius acids are also Bronsted–Lowry acids, and vice versa. However, the Bronsted–Lowry definition of a base is broader. There are substances other than Arrhenius bases that accept hydrogen ions. In other words, the hydroxide ion is a hydrogen ion acceptor, but there are other anions that also do this.

What exactly is meant by a hydrogen ion acceptor? A hydrogen ion acceptor is a chemical species to which a hydrogen ion can attach and form a covalent bond. The most important examples, in addition to the Arrhenius bases (hydroxides), are ammonia (NH_3), the carbonates, and the bicarbonates (Table 12.6). Examples of reactions in which hydrogen ions are accepted are found in Equations 12.17 through 12.21.

Table 12.6 Important Bronsted–Lowry Bases Other than Hydroxides

Name	Formula
Ammonia	NH ₃
Sodium carbonate	Na ₂ CO ₃
Potassium carbonate	K ₂ CO ₃
Calcium carbonate	CaCO ₃
Sodium bicarbonate	NaHCO ₃
Potassium bicarbonate	KHCO ₃



Notice that when hydroxides accept hydrogen ions, water is always the product. Water is not a product in the cases of ammonia, the carbonates, and the bicarbonates. However, in the case of carbonates and bicarbonates, water often does ultimately form. The carbonic acid, H₂CO₃, in Equations 12.20 and 12.21 is unstable and will spontaneously decompose into water and carbon dioxide. This reaction is often accompanied by vigorous gas (CO₂) evolution and will be discussed in more detail in Section 12.7.

The list of Bronsted–Lowry bases is actually much longer than the list that is given here. There are many anions like carbonate and bicarbonate that will form a covalent bond with hydrogen ions and form a weak acid like the carbonic acid. Thus the anion of any weak acid qualifies as a Bronsted–Lowry base. This list would include sulfide, cyanide, acetate, and anions of other carboxylic acids.

We have suggested previously that water can be thought of as an acid because in a sample of pure, distilled water, water molecules release hydrogen

ions to a small extent. According to the Bronsted–Lowry theory, water can also be thought of as a base because water molecules accept hydrogen ions to form the **hydronium ion**, H_3O^+ .



Many chemists, in fact, consider the product of acid ionization in water solution to be the hydronium ion and not the hydrogen ion.

12.7 Reactions Involving Acids and Bases

Some specific reactions of acids and bases are of some importance. In this section, we summarize them and give examples.

12.7.1 Neutralization

After acids lose or donate hydrogen ions, we say that they have been *neutralized*. Similarly, after bases accept hydrogen ions, we say that they have been neutralized. The reaction of a hydrogen ion donor with a hydrogen ion acceptor readily occurs and neutralizes both the acid and the base. The reaction of an acid with a base has thus come to be known as a neutralization reaction. The reaction is an example of a **double replacement** reaction in which the acid and the base simply “trade partners” as in the following general scheme:



Trading partners in the case of an acid and a base, in which the base is a hydroxide, results in a salt and water being formed.



Trading partners in the case of an acid and a base, in which the base is a carbonate, results in a salt and carbonic acid being formed.



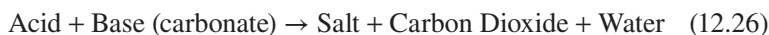
As mentioned in Section 12.6, carbonic acid (H_2CO_3) is very unstable and spontaneously decomposes to form carbon dioxide and water.



Table 12.7 Examples of Neutralization Reactions

(a)	$\text{HCl} + \text{NaOH} \rightarrow \text{NaCl} + \text{H}_2\text{O}$
(b)	$\text{H}_2\text{SO}_4 + 2\text{KOH} \rightarrow \text{K}_2\text{SO}_4 + 2\text{H}_2\text{O}$
(c)	$2\text{H}_3\text{PO}_4 + \text{Mg}(\text{OH})_2 \rightarrow \text{Mg}(\text{H}_2\text{PO}_4)_2 + 2\text{H}_2\text{O}$
(d)	$\text{H}_3\text{PO}_4 + 2\text{NaOH} \rightarrow \text{Na}_2\text{HPO}_4 + 2\text{H}_2\text{O}$
(e)	$\text{H}_3\text{PO}_4 + 3\text{KOH} \rightarrow \text{K}_3\text{PO}_4 + 3\text{H}_2\text{O}$
(f)	$2\text{HCl} + \text{CaCO}_3 \rightarrow \text{CaCl}_2 + \text{H}_2\text{CO}_3$
(g)	$\text{H}_2\text{SO}_4 + \text{Na}_2\text{CO}_3 \rightarrow \text{Na}_2\text{SO}_4 + \text{CO}_2 + \text{H}_2\text{O}$
(h)	$\text{HCl} + \text{Na}_2\text{CO}_3 \rightarrow \text{NaCl} + \text{NaHCO}_3$

Thus we can also write the following:



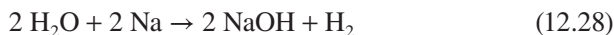
The neutralized acid may be either partially neutralized or fully neutralized. In the latter case, all acidic hydrogens shown in the formula are donated, while in the former, some but not all acidic hydrogens are donated. Thus the salt product may still show some undonated hydrogens remaining in the formula. Some examples are shown in Table 12.7.

12.7.2 Reaction of Acids with Metals

Acids react with some metals to form hydrogen gas in what is termed a **single replacement** reaction. In a single replacement reaction, one element is released from a formula and the remaining element combines with the reacting element.



Water is an acid of sufficient strength to react with the very active alkali metals, such as sodium and potassium. In fact, the reaction of these metals with water is quite violent, releasing hydrogen gas and leaving the metal hydroxide in solution. The reaction of sodium with water is as follows:



This reaction also occurs with the alkaline earth metals calcium and magnesium, although it is considerably less violent.

The less active metals, such as zinc and iron, require a stronger acid. In these cases, the products are the salt of the acid and hydrogen gas. For example, zinc reacts with hydrochloric acid as follows:



Very inactive metals, such as copper and silver, do not undergo this reaction. These metals do react with nitric acid and other oxidizing acids but give different products.

12.7.3 Reaction of Metal and Nonmetal Oxides with Water

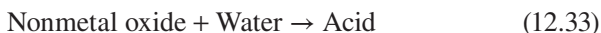
Metal oxides react with water to form the metal hydroxides, or bases.



Examples include reactions of the oxides of sodium and magnesium with water.



Nonmetal oxides, such as oxides of carbon, nitrogen, sulfur, and phosphorus, react with water to give acids.



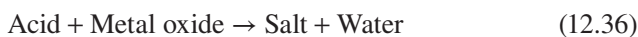
Examples include CO_2 and SO_3 .



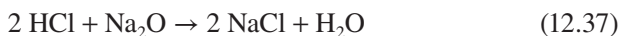
This type of reaction occurs when rainwater reacts with nonmetal oxide pollutants in the air to form what has been commonly referred to as **acid rain**. Acid rain has been a troublesome phenomenon in regions surrounding heavily industrialized areas. Vegetation can be killed; metal and granite statues are corroded; and the health of people can be adversely affected by acid rain.

12.7.4 Reaction of Metal Oxides with Acids

Since metal oxides react with water to form metal hydroxides, it is predictable that the products of the reaction of metal oxides with acids are the same as the reaction of metal hydroxides with acids. The products are a salt and water.



Let us again use Na_2O and MgO as our examples.



12.8 “Add Acid to Water” Rule

Much heat is evolved when sulfuric acid is mixed with water. This is because the ionization reaction



is highly exothermic (i.e., is accompanied by the evolution of much heat energy). Actually, the ionization reactions of all strong acids—and also of many weak acids—are exothermic. Extreme care must be exercised when any acid is mixed with water. A nontrivial question that arises is, When mixing an acid with water, should you add the acid to the water or the water to the acid? An important consideration here is that when the exothermic ionization reaction occurs, some splashing may take place due to the high heat evolution and possible boiling of the liquid at the point of contact. A serious safety hazard thus is present because the hot acid may be splashed on one’s face and clothing. To greatly minimize the problem, one should *always add the acid to the water*. The reason for this is that the splashed liquid in this scenario consists mostly of the liquid in the receiving container, and if that liquid is mostly water, then it is of course much less harmful to the face and hands

12.9 Homework Exercises

Formulas and Strengths of Acids and Bases

1. Write the names of at least four common acids and give their chemical formulas.

2. Sodium acetate ($\text{NaC}_2\text{H}_3\text{O}_2$) has several hydrogens in its formula, but it is not an acid. Explain. (It is not sufficient to say that the hydrogens are not listed first in the formula.)
3. What is the difference between strong acids and weak acids? Give examples of each.
4. Give two examples of compounds that are acids despite the fact that they do not have a hydrogen first in the formula. Explain why they are acids.
5. Name three compounds that are common laboratory bases and give their formulas.
6. Why are concentrated solutions of the weak base NH_4OH also concentrated solutions of ammonia (NH_3)? Why is NH_4OH considered a weak base?

General Properties of Acids and Bases

7. Why do citrus fruits have a sour taste?
8. What are some general properties of acids? Include both chemical and physical properties in your answer.
9. What are some general properties of bases? Include both chemical and physical properties in your answer.
10. What is an indicator? How is this used with solutions that contain acids or bases?
11. What is pH paper? How is it used with acid or base solutions?
12. What does it mean to say that H^+ ions predominate over OH^- ions? What properties does this give a solution?
13. What is meant by a neutralization reaction? Give an example.

Ion Product Constant, pH, and pOH

14. What is the definition of K_w ? Why is this important when dealing with water solutions?
15. Define pH. Why is pH important when dealing with water solutions?
16. What is the pH when $[\text{H}^+]$ has the following values?
 - a. $5.02 \times 10^{-3} \text{ M}$
 - b. $9.29 \times 10^{-5} \text{ M}$
 - c. $3.81 \times 10^{-8} \text{ M}$
 - d. $4.88 \times 10^{-11} \text{ M}$

- e. $6.06 \times 10^{-13} \text{ M}$
f. 0.00123 M
17. What is the $[\text{H}^+]$ when the pH is as follows.
- a. 4.119
b. 2.911
c. 10.449
d. 8.383
e. 6.229
18. Complete the following table.

	$[\text{H}^+]$	pH	Acidic, Basic, or Neutral?
(a)	$5.29 \times 10^{-4} \text{ M}$		
(b)		2.943	
(c)	$7.6 \times 10^{-3} \text{ M}$		
(d)		4.83	
(e)	$9.39 \times 10^{-10} \text{ M}$		
(f)		11.452	
(g)	$5.0 \times 10^{-9} \text{ M}$		
(h)		12.71	
(i)	$4.48 \times 10^{-5} \text{ M}$		
(j)		6.098	
(k)	$8.3 \times 10^{-8} \text{ M}$		
(l)		10.66	
(m)	$6.19 \times 10^{-11} \text{ M}$		
(n)		5.649	

19. Calculate the pH of the following.
- a. A solution in which $[\text{H}^+]$ is $9.42 \times 10^{-12} \text{ M}$
b. A solution in which $[\text{OH}^-]$ is $2.04 \times 10^{-3} \text{ M}$
c. A solution in which the HCl concentration is 0.077 M
20. What is the $[\text{H}^+]$ in each of the following situations?
- a. $\text{pOH} = 4.51$
b. $[\text{OH}^-] = 1.88 \times 10^{-2} \text{ M}$.
c. $\text{pH} = 7.00$
d. $\text{pH} = 3.29$

21. Complete the following table.

	[H ⁺]	[OH ⁻]	pH	pOH	Acidic, Basic, or Neutral?
(a)	$4.97 \times 10^{-5} \text{ M}$				
(b)		$7.28 \times 10^{-4} \text{ M}$			
(c)			6.234		
(d)				2.887	
(e)	$6.0 \times 10^{-11} \text{ M}$				
(f)		$3.8 \times 10^{-9} \text{ M}$			
(g)			10.35		
(h)				12.11	

22. What is meant by (a) pH meter and (b) combination pH probe?

Theories of Acids and Bases

23. What is the definition of an acid according to the Arrhenius theory? What is the definition of a base according to this theory. Give one example of an Arrhenius-type acid or base.
24. What is the definition of an acid according to the Bronsted–Lowry theory? What is the definition of a base according to this theory. Give one example of a Bronsted–Lowry acid or base.
25. Consider the following compounds:
- HNO_3
 - Na_2CO_3
 - $\text{HC}_2\text{H}_3\text{O}_2$
 - H_2SO_4
 - NH_3
 - KOH
 - CaCO_3
 - NaCl

From this list, write which are Arrhenius acids, which are Bronsted–Lowry acids, which are Arrhenius bases, and which are Bronsted–Lowry bases.

26. What is the hydronium ion? What does its formation have to do with water being a Bronsted–Lowry base?

Reactions of Acids and Bases

27. What is meant by a double replacement reaction? What is a single replacement reaction? Give one example of each.
28. Write the general reactions that occur when the following types of chemicals are combined together.
 - a. Acid + Metal
 - b. Metal oxide + Acid
 - c. Nonmetal oxide + Water
 - d. Metal oxide + Water
29. Give the products in the following equations. Write correct formulas, but do not be concerned with balancing.
 - a. $\text{K}_2\text{O} + \text{H}_2\text{O} \rightarrow$
 - b. $\text{HCl} + \text{CaCO}_3 \rightarrow$
 - c. $\text{NH}_4\text{OH} + \text{H}_2\text{SO}_4 \rightarrow$
 - d. $\text{NO}_2 + \text{H}_2\text{O} \rightarrow$
 - e. $\text{HCl} + \text{Mg} \rightarrow$
 - f. $\text{CaO} + \text{HBr} \rightarrow$
30. Fill in the product side. Write correct formulas, but do not necessarily be concerned with balancing.
 - a. $\text{H}_2\text{SO}_4 + \text{K}_2\text{CO}_3 \rightarrow$
 - b. $\text{NO}_2 + \text{H}_2\text{O} \rightarrow$
 - c. $\text{HNO}_3 + \text{Mg}(\text{OH})_2 \rightarrow$
 - d. $\text{K} + \text{H}_2\text{O} \rightarrow$
 - e. $\text{Ca} + \text{HCl} \rightarrow$
31. What would you expect to happen if you placed a piece of potassium metal in a beaker of water?
32. $\text{Mg}(\text{OH})_2$ is a base contained in Phillips' Milk of Magnesia. It reacts with the excess HCl in our stomach walls to form what products?
33. What is acid rain? How does it form? Give an example of an acid that may be dissolved in acid rain.
34. What is the "add acid to water" rule? Give a detailed description of the procedure you would use to prepare a 50–50 mixture of sulfuric acid and water.

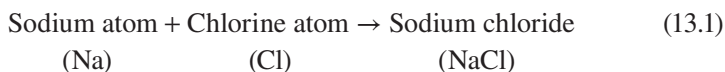
Chapter 13

Oxidation– Reduction Reactions

Oxidation–Reduction Reactions	330
Ion–Electron Method for Balancing Equations	336
Homework Exercises	338

13.1 Oxidation–Reduction Reactions

As discussed in detail in Chapter 3, many chemical species have a tendency to either give up or take on electrons. This tendency is based on the premise that a greater stability, or lower energy state, is achieved as a result of this electron donation or acceptance, such as when an atom achieves the stable noble gas electron configuration as a result of this electron transfer. An example is the reaction between a sodium atom and a chlorine atom.



A sodium atom with only one electron in its outermost energy level achieves a lower energy state when this electron is released to the chlorine atom, which also achieves a lower energy state as a result. Both atoms become ions (Na^+ and Cl^-), and each has a stable, filled outermost energy level identical with those of the noble gases—neon in the case of sodium and argon in the case of chlorine. Thus the electron transfer takes place and sodium chloride, NaCl , is formed.

Those reactions, which this sodium-chlorine case typifies, are called **oxidation–reduction reactions**. The term **oxidation** refers to the loss of electrons, while the term **reduction** refers to the gain of electrons.

It may appear strange that the term *reduction* is associated with a gaining process. Actually, the term *reduction* was coined as a result of what happens to the oxidation number of the element when the electron transfer takes place. What is meant by oxidation number?

The **oxidation number** of an element in a compound is a number assigned to that element that reflects how its electrons are involved in making up the compound. For example, if an atom of an element loses one electron and forms a +1 cation, the oxidation number of that element in a compound is +1. If an atom of an element gains one electron and forms a –1 ion, the oxidation number is –1. This is also often referred to as the **oxidation state** of an element. It is a sort of book-keeping concept that is used in some naming schemes, as well as to track electrons in redox reactions. Thus it is a concept of some importance.

While the examples of a +1 or a –1 oxidation number may appear simple enough (being simply the charge on the ion), there are many cases in which

Table 13.1 Ten Rules for Assigning Oxidation Numbers

Element	Oxidation Number
1. Any uncombined element	0
2. Combined alkali metal (group I)	+1
3. Combined alkaline earth metal (group II)	+2
4. Combined aluminum	+3
5. Halogen (group VIIA) in a binary compound with a metal or hydrogen	−1
6. Hydrogen combined with nonmetal	+1
7. Hydrogen combined with metal (hydrides)	−1
8. Combined oxygen (except peroxides)	−2
9. Oxygen in peroxides	−1
10. All others—oxidation numbers are determined from the fact that the sum of all oxidation numbers within the molecule or polyatomic ion must add up to either zero (in the case of a neutral molecule) or the charge on the ion (in the case of a polyatomic ion)	

the determination of an oxidation number is more involved, such as in the case of an element found in a molecule or polyatomic ion. For this reason, rules are used to help determine the oxidation number. These rules are listed in Table 13.1.

Let us make a few observations about these rules before we look at some examples of their use. First, in Rule 1 the uncombined element is an element that is in the free elemental state, or the state of the element when it is not combined with any other element. For most elements, this is shown by the use of the symbol of the element, as found in the periodic table. For example, the oxidation numbers of silver metal (Ag), radon gas (Rn), and mercury liquid (Hg) would be 0. However, there are some elements whose free elemental state refers to diatomic molecules, or molecules that consist of two atoms of the element that are covalently combined. This list includes hydrogen gas (H_2), fluorine gas (F_2), nitrogen gas (N_2), oxygen gas (O_2), chlorine gas (Cl_2), bromine liquid (Br_2), and iodine solid (I_2). Thus, whenever these diatomic symbols are observed, these substances are in their free elemental state and the correct oxidation number to be assigned would be 0.

Second, in Rule 5, note that the -1 oxidation number for a combined halogen atom is strictly limited to binary compounds with a metal or hydrogen. Halogens found in polyatomic ions or in ternary compounds do not have an oxidation number of -1 . Third, Rules 8 and 9 mentions peroxides. Peroxide is a polyatomic ion that consists of two oxygen atoms that have a net charge of -2 , or O_2^{2-} .

Example 13.1

What is the oxidation number of nitrogen in N_2 ?

Solution 13.1

The formula N_2 represents nitrogen in the free elemental state. Thus, according to rule 1, the oxidation number is 0.

Example 13.2

What is the oxidation number of barium in BaSO_3 ?

Solution 13.2

Barium is an alkaline earth metal, and according to Rule 4, it has an oxidation number of $+2$ when it is in the combined state.

Rule 10 is needed for such elements as Cu in CuCl . Cu is not mentioned in Rules 1 through 9. Its oxidation number is determined by knowing that it must add to the oxidation number of Cl to give zero. The oxidation number of Cl is -1 (Rule 5) and so the oxidation number of Cu is $+1$ in CuCl .

Rule 10 can be a source of confusion in the sense that often more than one atom of an element is present in the same formula. Examples are the potassium in K_2S and the oxygen in NO_3^{1-} . In these cases, the rule should be interpreted to mean that the oxidation numbers of *all* atoms must add together to either give zero or the charge on the ion. Thus, in K_2S , because there are two potassium atoms, each with an oxidation number of $+1$ (Rule 2), the total contribution from potassium is $+2$, not $+1$; thus, sulfur's oxidation number must be -2 for the sum to be zero. In NO_3^{1-} , the total contribution from oxygen is -6 (-2 for each of the three oxygen atoms in the formula—see Rule 6); thus the oxidation number for nitrogen must be $+5$ for the sum to be -1 , the charge on the ion.

Example 13.3

What is the oxidation number of sulfur in Li_2SO_4 ?

Solution 13.3

Sulfur is not covered by Rules 1 through 9, so its oxidation number must be determined by Rule 10. The sum of the oxidation numbers of all the atoms in Li_2SO_4 must be zero. Lithium is an alkali metal whose oxidation number is +1. Since there are two lithium atoms in the formula, its total charge contribution is +2. According to Rule 8, oxygen in the combined state has an oxidation number of -2 , and since there are four oxygen atoms in the Li_2SO_4 formula, the total charge contribution from oxygen is -8 . In order for all oxidation numbers to add to zero, sulfur's oxidation number must be +6, since +2, +6, and -8 add to 0.

Example 13.4

What is the oxidation number of phosphorus in the phosphate ion PO_4^{3-} ?

Solution 13.4

Phosphorus is not covered directly by Rules 1 through 9, so its oxidation number is determined by Rule 10. The sum of the oxidation numbers of all the atoms in PO_4^{3-} must be -3 . All oxygen atoms have an oxidation number of -2 (rule 8), and thus the total charge contribution from oxygen is -8 . If the sum is to be -3 , then the oxidation number of phosphorus must be +5 since +5 and -8 add to -3 .

One final point about Rule 10 involves a situation in which there is more than one element in a formula that cannot be determined by rules 1 through 9, such as when one is a transition metal and the other is a nonmetal that is part of a polyatomic ion. In this case, the charge on the polyatomic ion can be used to determine the oxidation number of the metal. The oxidation number of the nonmetal can then be determined as in the earlier examples.

Example 13.5

What are the oxidation numbers of Fe and S in FeSO_4 ?

Solution 13.5

The oxidation number of iron is +2 because the charge on the sulfate ion is -2 and they must add to 0. The oxidation number of oxygen is -2 , giving a total charge contribution of -8 for four oxygen atoms. If iron is +2, sulfur must be +6 to add to 0 (+2, +6, and -8 equal 0).

Example 13.6

What are the oxidation numbers of Sn and P in $\text{Sn}_3(\text{PO}_4)_2$?

Solution 13.6

The phosphate ion has a -3 charge and, since there are two phosphate groups in the formula, the total charge from the phosphate is -6 . This means that in order to add to zero, the total charge of Sn must be +6 and, since there are three Sn ions in the formula, each individual Sn must have an oxidation number of +2. Similarly, the P and O oxidation numbers must add to -3 , the charge on the phosphate ion. Oxygen is -2 for a total of -8 , so phosphorus must be +5 (+5 and -8 equal -3).

The usefulness of determining the oxidation number in chemistry is two-fold. First, it helps determine if there is a change in the oxidation number of a given element in a reaction. This always signals the occurrence of an oxidation–reduction reaction. Thus, it helps tell us whether a reaction is a redox reaction or some other reaction. Second, determining the oxidation number assists with determining the number of electrons involved, which aids in balancing the equation.

Example 13.7

Which of the following is a redox reaction?

- a. $\text{Cl}_2 + 2\text{NaBr} \rightarrow \text{Br}_2 + 2\text{NaCl}$
- b. $\text{BaCl}_2 + \text{K}_2\text{SO}_4 \rightarrow \text{BaSO}_4 + 2\text{KCl}$

Solution 13.7

The oxidation numbers of Cl and Br in Example 13.7a have changed. Cl has changed from 0 to -1 , while Br has changed from -1 to 0. There is no change in any oxidation number of any element in Example 13.7b. Thus Example 13.7a is redox and Example 13.7b is not.

The terms *oxidation* and *reduction* can actually be defined in two different ways—according to the gain or loss of electrons or according to the increase or decrease in oxidation number.

Oxidation: the loss of electrons or the increase in oxidation number

Reduction: the gain of electrons or the decrease in oxidation number

To say that a substance has been oxidized means that the substance has lost electrons. To say that a substance has been reduced means that the substance has gained electrons. To say that substance A has oxidized substance B means that substance A has caused substance B to lose electrons. Similarly, to say that substance A has reduced substance B means that substance A has caused substance B to gain electrons. When substance A causes substance B to be oxidized, substance A is called the **oxidizing agent**. When substance A causes substance B to be reduced, substance A is a **reducing agent**. To say that substance A causes oxidation or reduction means that substance A either removes electrons from or donates electrons to substance B. Thus oxidation is always accompanied by reduction, and vice versa. Also, every redox reaction has an oxidizing agent (which is the substance reduced) and a reducing agent (which is the substance oxidized).

Example 13.8

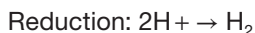
Tell what is oxidized and what is reduced and tell what the oxidizing agent and the reducing agent are in the following:



Solution 13.8



(The oxidation number of Zn has increased from 0 to +2.)



(The oxidation number of H has decreased from +1 to 0.)

Zn has been oxidized; therefore it is the reducing agent. H^+ has been reduced; therefore it is the oxidizing agent.

13.2 Ion–Electron Method For Balancing Equations

We know from our Chapter 8 discussion that a balanced chemical equation is needed for basic stoichiometry. With redox reactions, we must not only pay attention to the element balance but also to electron balance. The number of electrons given up must equal the number of electrons taken on. Our Chapter 8 technique of balancing by inspection must therefore be modified. Several special schemes have been derived for balancing redox equations. The **ion-electron method** directly and clearly takes into account the electrons that are transferred, so that is the method that we have chosen here.

This method makes use of only those species, dissolved or otherwise, that actually take part in the reaction. So-called **spectator ions**, or ions that are present but play no role in the chemistry, are not included in the balancing procedure. Solubility rules are involved here, since spectator ions result only when an ionic compound dissolves and ionizes. Also, the scheme is slightly different for acidic and basic conditions. Our purpose, however, is to discuss the procedure at a fundamental level, and thus spectator ions will be absent from all examples from the start, and acidic conditions will be the only conditions considered. The step-wise procedure we will follow is given here and an example follows.

- Step 1:** Look at the equation to be balanced and determine what is oxidized and what is reduced. This involves checking the oxidation numbers and discovering which have changed.
- Step 2:** Write a **half-reaction** for both the oxidation process and the reduction process and label as oxidation and reduction. These half-reactions show only the species being oxidized (or the species being reduced) on the left side and only the product of the oxidation (or reduction) on the right side.
- Step 3:** If oxygen appears in any formula on either side in either equation, it is balanced by writing H_2O on the opposite side. This is possible since the reaction mixture is a water solution. The hydrogen in the water is then balanced on the other side by writing H^+ , since we are dealing with acid solutions. Now balance both half-reactions for all elements by inspection.

Step 4: Balance the charges on both sides of the equation by adding the appropriate number of electrons (written as e^- in the equation) to whichever side is deficient in negative charges. The charge balancing is accomplished as if the electron is like any chemical species—place the appropriate multiplying coefficient in front of the electrons.

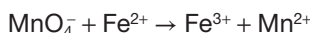
Step 5: Multiply through both equations by appropriate coefficients, so that the number of electrons involved in both half-reactions is the same. This has the effect of making the total charge loss equal to the total charge gain and thus eliminates electrons from the final balanced equation, as you will see in Step 6.

Step 6: Add the two equations together. The number of electrons, being the same on both sides, cancels out and thus does not appear in the final result. Also some H^+ and H_2O can be cancelled out if they appear on both sides at this point.

Step 7: Make a final check to see that the equation is balanced.

Example 13.9

Balance the following equation by the ion-electron method:



Solution 13.9

Step 1: Fe^{2+} oxidized from $+2 \rightarrow +3$ (loss of electrons)

MnO_4^- reduced from $+7 \rightarrow +2$ (gain of electrons)

Step 2: Oxidation: $Fe^{2+} \rightarrow Fe^{3+}$

Reduction: $MnO_4^- \rightarrow Mn^{2+}$

Step 3: Oxidation: $Fe^{2+} \rightarrow Fe^{3+}$

Reduction: $8H^+ + MnO_4^- \rightarrow Mn^{2+} + 4H_2O$

Step 4: Oxidation: $Fe^{2+} \rightarrow Fe^{3+} + 1e^-$

Reduction: $5e^- + 8H^+ + MnO_4^- \rightarrow Mn^{2+} + 4H_2O$

Step 5: Oxidation: $5(Fe^{2+} \rightarrow Fe^{3+} + 1e^-)$

Reduction: $1(5e^- + 8H^+ + MnO_4^- \rightarrow Mn^{2+} + 4H_2O)$

Step 6: Oxidation: $5(Fe^{2+} \rightarrow Fe^{3+} + 1e^-)$

Reduction: $+1(5e^- + 8H^+ + MnO_4^- \rightarrow Mn^{2+} + 4H_2O)$



Step 7: 5 Fe on each side

8 H on each side

1 Mn on each side

4 O on each side

The equation is balanced.

13.3 Homework Exercises

Oxidation Number

1. Define the term *oxidation number*.
2. What is meant by the free elemental state of an element?
3. Write the symbol for the free elemental state for each of the following elements.
 - a. Gold
 - b. Helium
 - c. Oxygen
 - d. Chlorine
4. Write the normal oxidation numbers for each of the following elements when they are combined with different elements in chemical compounds.
 - a. Oxygen
 - b. Hydrogen
 - c. Sodium
 - d. Calcium
 - e. Aluminum
 - f. Chlorine
5. What is the oxidation number for hydrogen in a hydride?
6. What is a peroxide? What is the oxidation number of oxygen in a peroxide?
7. What is the oxidation number of nitrogen in each of the following compounds or ions?
 - a. N_2O
 - b. NO_2^-
 - c. NaNO_3
 - d. N_2
 - e. Ba_3N_2
 - f. NH_4^+

8. What is the oxidation number of bromine (Br) in each of the following?
- HBrO
 - NaBr
 - BrO_3^-
 - Br_2
 - $\text{Mg}(\text{BrO}_2)_2$
 - BrO_3^-
9. What is the oxidation number of chromium (Cr) in each of the following?
- CrBr_3
 - Cr
 - CrO_3
 - CrO_4^{2-}
 - $\text{K}_2\text{Cr}_2\text{O}_7$
10. What is the oxidation number of iodine (I) in each of the following?
- HIO_4
 - CaI_2
 - I_2
 - IO_2^-
 - $\text{Mg}(\text{IO})_2$
11. What is the oxidation number of sulfur (S) in each of the following?
- SO_2
 - H_2S
 - S
 - SO_4^{2-}
 - K_2SO_3
 - K_2S
 - H_2SO_4
 - SO_3^{2-}
 - SO_2
 - SF_6
12. What is the oxidation number of phosphorus (P) in each of the following?
- P_2O_5
 - Na_3PO_3
 - H_3PO_4

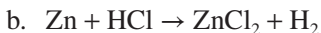
- d. PCl_3
 - e. HPO_3^{2-}
 - f. PO_4^{3-}
 - g. P^{3-}
 - h. $\text{Mg}_2\text{P}_2\text{O}_7$
 - i. P
 - j. NaH_2PO_4
13. What is the oxidation number of each of the following?
- a. Calcium in any compound?
 - b. Sodium when it is not chemically combined with anything?
 - c. Chlorine in NaCl ?
 - d. Manganese in MnO_2 ?
 - e. Oxygen in H_2O ?
 - f. Iron in FeCO_3 ?
 - g. Copper in CuNO_3 ?
 - h. Tin in $\text{Sn}(\text{SO}_4)_2$?

Terminology

14. Define oxidation, reduction, oxidation number, oxidizing agent, and reducing agent.
15. In the following redox reactions, tell what has been oxidized and what has been reduced and explain your answers.
- a. $3\text{CuO} + 2\text{NH}_3 \rightarrow 3\text{Cu} + \text{N}_2 + 3\text{H}_2\text{O}$
 - b. $\text{Cl}_2 + 2\text{KBr} \rightarrow \text{Br}_2 + 2\text{KCl}$
16. In each of the following reactions, tell what is the oxidizing agent and what is the reducing agent, and explain your answers.
- a. $\text{Mg} + 2\text{HBr} \rightarrow \text{MgBr}_2 + \text{H}_2$
 - b. $4\text{Fe} + 3\text{O}_2 \rightarrow 2\text{Fe}_2\text{O}_3$
17. Consider the following two unbalanced equations:
- a. $\text{KOH} + \text{HCl} \rightarrow \text{KCl} + \text{H}_2\text{O}$
 - b. $\text{Cu} + \text{HNO}_3 \rightarrow \text{Cu}(\text{NO}_3)_2 + \text{NO} + \text{H}_2\text{O}$

Which represents a redox reaction, (a) or (b)? In the redox reaction, what has been oxidized and what has been reduced? In the redox reaction, what is the oxidizing agent and what is the reducing agent?

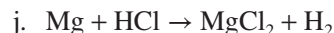
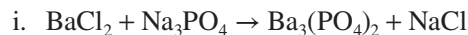
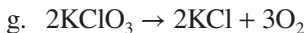
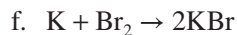
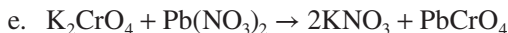
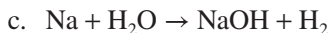
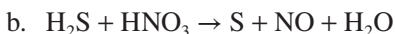
18. Consider the following two unbalanced equations:



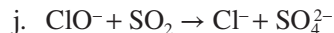
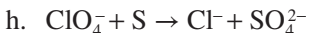
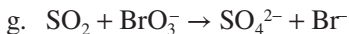
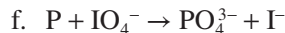
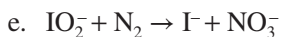
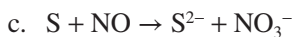
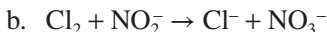
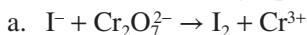
Answer the same questions as in Exercise 17.

Recognizing and Balancing Redox Reactions

19. Consider each of the following unbalanced equations and answer the question: does it represent a redox reaction? Explain your answers.



20. Balance the following equations by the ion-electron method.



Chapter 14

Organic Chemicals

Introduction	344
Chemistry of Carbon and Its Compounds	345
Hybridization of Orbitals	345
Geometry around Bonded Carbon Atoms	346
Physical Properties	348
Hydrocarbons	350
Organic Compounds Containing Oxygen	362
Summaries of Hydrocarbon and Oxygen Classifications	370
Polymers	370
Homework Exercises	376

14.1 Introduction

Organic chemistry is the chemistry of those compounds (organic chemicals) that are derived from living systems or systems that were once living. You might have heard the expression that life on Earth is carbon-based. This means that all organic chemicals have carbon in their structure and that there are relatively very few compounds of carbon that are not organic (e.g., CO_2 , CO , and metal carbonates and carbides). Therefore we can also say that **organic chemistry** is the chemistry of compounds of carbon. Organic chemicals make up the vast majority of all compounds that exist, so this is a very important branch of chemistry.

Our major source of organic compounds is petroleum. Petroleum results from the decay of living systems over time. It is interesting to imagine that the trees, grasses, animals, and other living things from years past have been folded into the Earth and have decayed to produce the pockets of crude oil under the surface of the Earth worldwide. Chemicals found in crude oil have been useful for so many of our daily interactions with matter. When most people think of petroleum, they think of gasoline. Gasoline is, of course, a very important derivative of petroleum, but it is not the only important derivative. In fact, there are literally thousands of materials with which we routinely come into contact that are derived from petroleum and are organic.

A huge industrial sector today is composed of companies that produce products from chemicals derived from petroleum. Companies operate oil refineries that refine and purify crude oil to make fuels such as gasoline and important chemicals that are used in other industrial manufacturing processes. For example, an organic chemical called ethylene is manufactured from the chemicals in petroleum, and this ethylene is then used in the manufacture of a myriad of plastics (e.g., polyethylene) and other materials that we all use. Ethylene is the highest volume organic chemical manufactured today because it is used for so many things. Plastics, yes, and other polymers as well. Near the end of this chapter will cite a number of examples of polymers and other organic chemicals that we take for granted without any thought of how these materials came to be and what role chemistry has played in their development and production.

14.2 Chemistry of Carbon and Its Compounds

Why are there so many organic compounds? The main reason is the proliferation of the bonding capabilities of the element carbon. In fact, most molecular structures of such compounds have as their backbone a long string, or chain, of carbon atoms bonded together. Many even have a network of chains of carbon atoms in their molecular structure. Bonded to these backbones or networks are atoms of other elements. Hydrogen atoms are found in abundance in this arrangement. Found to a lesser extent are oxygens, nitrogens, sulfurs, and other nonmetals. Atoms of metals also find their way into these structures, but to a much lesser degree.

Molecular structures that make up living systems contain these networks of carbon chains in abundance. Modern studies of biotechnology, biochemistry, and biology have as their basis the study of the chemistry of carbon compounds. One can easily see the vast scope of organic chemistry. Not only is organic chemistry the study of the chemistry of compounds derived from living systems (from petroleum) but it is also the study of the chemistry of systems that are living, and this list includes all vegetation, all animals, and all human beings. A timely example of the latter is DNA* chemistry—an extraordinarily important topic in the modern world. But, alas, we must slow down a bit and consider the prolific behavior of that small atom we call carbon before tackling these other more engrossing topics.

14.3 Hybridization of Orbitals

The proliferation of the bonding of carbon atoms resulting in these backbones and networks is facilitated by the hybridization of carbon's atomic orbitals mentioned in Chapter 6 (Figures 6.11 and 6.12). Let us briefly review the previous discussion. With carbon's electron configuration (Figure 14.1a) and orbital diagram (Figure 14.1b), one would expect the Lewis electron-dot structure in Figure 14.1c. However, a phenomenon occurs with carbon that is relatively common among other nonmetals as well. The phenomenon is called hybridization. All four electrons become unpaired, and the orbital picture is transformed. We will not be examining the orbital picture here. The important

* Deoxyribonucleic Acid

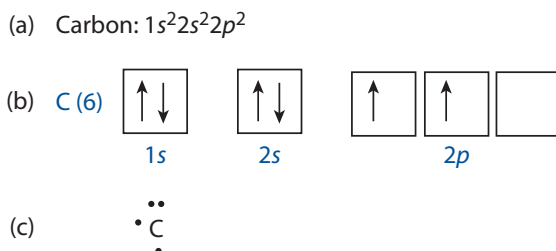


FIGURE 14.1 Electron structures of elemental carbon before hybridization. (a) Electron configuration, (b) orbital diagram, and (c) Lewis structure.



FIGURE 14.2 Lewis structures showing that all four electrons on a carbon become unpaired during hybridization.

point is that all four electrons become unpaired, as indicated in Figure 14.2, and available for bonding.

This produces a capability for covalent bonding that is unique. In all organic compounds—from the very simple to those that possess a complex network structure—carbon atoms always bond four times. While there may be single bonds, double bonds, and triple bonds involved, the total number of bonds to any carbon atom is always four—no more and no less—because of the four unpaired electrons resulting from hybridization.

One final point is that there can be p orbitals still remaining after hybridization. That is why double bonds and triple bonds are possible in organic compounds. Recall that the second bond of a double bond is a pi (π)-bond (overlap of two p orbitals) and that the second and third bonds of a triple bond are π -bonds.

14.4 Geometry around Bonded Carbon Atoms

When all four bonds to a carbon are single bonds, there are four regions of high electron density and the geometry around this carbon is tetrahedral. The molecular geometry of compounds containing just one carbon (as the central atom) is therefore tetrahedral. Examples of this are methane (CH_4) and carbon tetrachloride (CCl_4). These molecules would appear as shown with the

ball-and-stick model in Figure 14.3, with the large ball in the center representing the carbon and the four smaller balls surrounding it representing the hydrogens (in the case of methane) or chlorines (in the case of carbon tetrachloride). If the compound has two or more carbon atoms bonded together with nothing but single bonds attached (such as in many of the backbones or networks referred to earlier), then the molecular structure would consist of a tetrahedral geometry around each individual carbon. See Figure 14.4. This structure would have a carbon–carbon geometry that would be linear for two carbons. But it would begin to take on a sawtooth appearance for three or more carbons if the carbon skeleton (the structure showing only the carbon atoms and not what is bonded to them [as in Figure 14.5]) is laid out so that the bonds are as far from each other as they can be. This carbon skeleton and the sawtooth pattern are useful for visualizing the structure of most organic compounds.

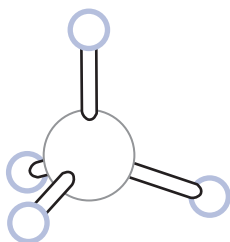


FIGURE 14.3 A drawing of the ball-and-stick model showing the tetrahedral molecular geometry exemplified by such molecules as CH_4 and CCl_4 . The sticks in such models represent single bonds (a pair of electrons being shared).

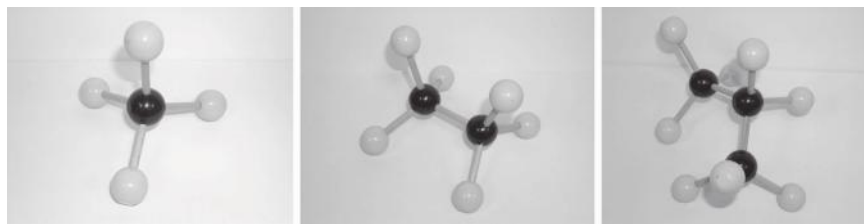


FIGURE 14.4 Photographs of three actual ball-and-stick models showing how a carbon chain forms. From left, methane, ethane, and propane. Notice the tetrahedral geometry around each of the carbons.

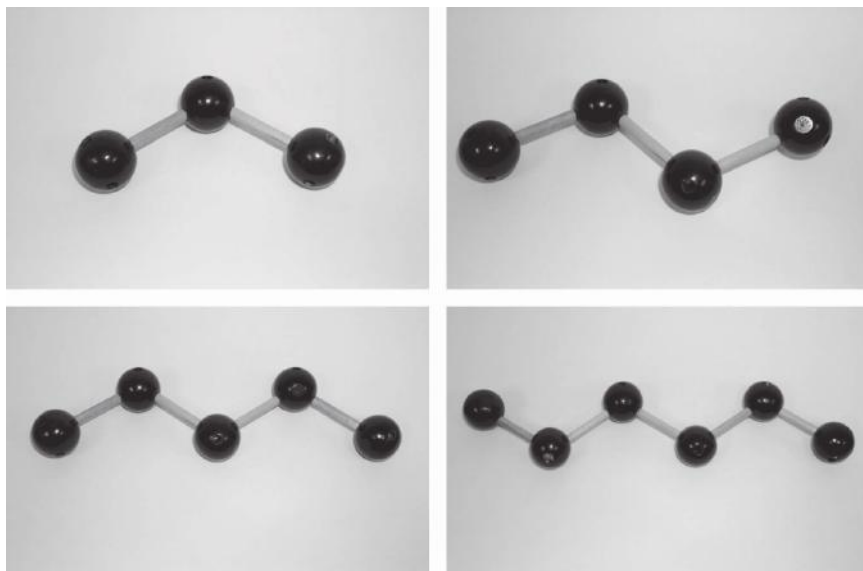


FIGURE 14.5 Ball-and-stick models (hydrogens removed) of carbon chains that are three, four, five, and six carbons long. The sawtooth appearance is due to the tetrahedral geometry around carbons where there are only single bonds attached.

When a double bond is attached to a carbon, there are three regions of high electron density and the geometry around this carbon is trigonal planar (Figure 14.6). When a triple bond is attached to the carbon, there are two regions of high electrons and the geometry around this carbon is linear (Figure 14.7). It is important to remember that the planar or linear geometry appears only around carbons to which the double or triple bonds are attached. The remainder of the structure—where there are only single bonds—has the sawtooth appearance.

14.5 Physical Properties

Crude oil, as it comes from the oil well, is a black, odorous, syrupy liquid mixture. Gasoline, a mixture of organic compounds derived from crude oil, is a yellowish, odorous, volatile, and highly flammable liquid mixture. It should not be too surprising to realize that many organic compounds are liquids at ordinary temperatures and pressures, that many have an odor, that

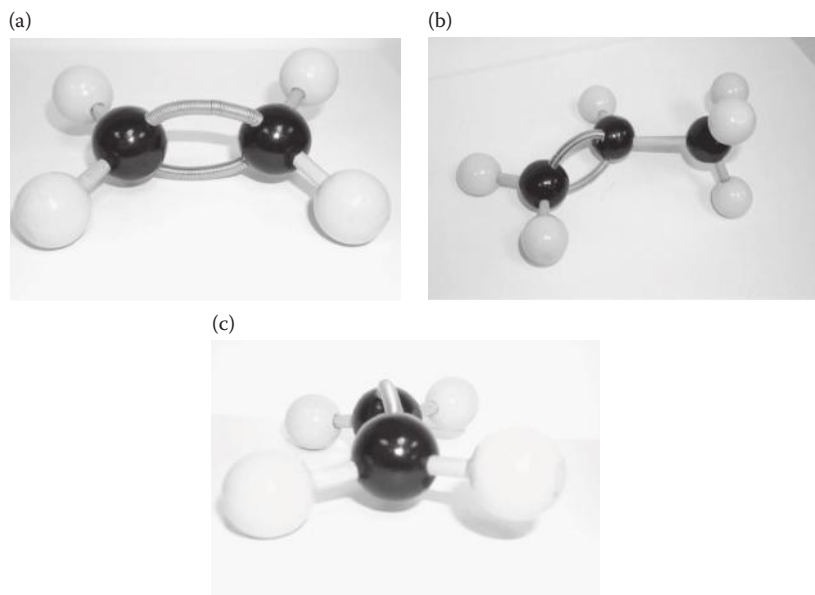


FIGURE 14.6 Ball-and-stick models of ethene (a, c) and propene (b). Flexible springs are used to show a double bond. From the angle in (c) we can see that the geometry is planar around carbons where there is a double bond attached. Notice that the third carbon on propene (b) has the tetrahedral geometry, since it has only single bonds attached.

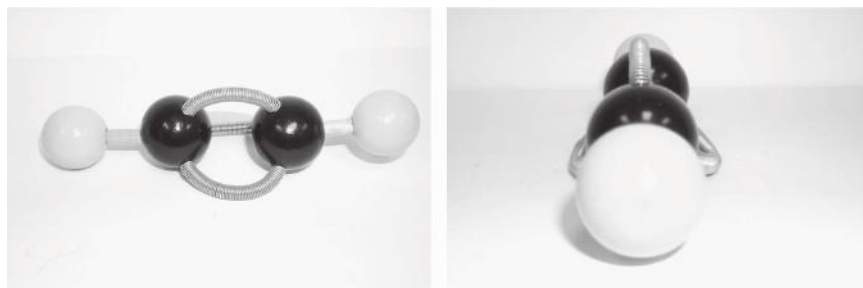


FIGURE 14.7 Ball-and-stick model of ethyne. Notice the linear geometry (right) around the carbons. This is characteristic of carbons that have a triple bond attached.

many are volatile, and that many are flammable. Compared with inorganic materials, this gives a new perspective on chemical compounds. Very few inorganic materials are liquids at ordinary temperatures and pressures, and very few have an odor. Very few are volatile, and few are flammable. The only common inorganic substance that is a liquid at ordinary temperatures and pressures is water. Of course, inorganic chemists commonly deal with water, water solutions, and other inorganic materials in their jobs. Thus various physical properties such as solubility, boiling point, freezing point, density, specific gravity, color, and refractive index are properties that are commonly observed and measured. However, with organic compounds, a broad array of other physical properties are observed for both characterization of the materials themselves and safety reasons. In addition to those physical properties just listed, others such as odor, flash point, flammability, viscosity, volatility, vapor pressure, miscibility with water, and optical rotation become important. Because of this, it is important for students of chemistry to gain first-hand experience with these compounds and with the observation and measurement of their physical properties in their laboratory work.

14.6 Hydrocarbons

The fact that there are millions of organic compounds dictates that we must find an alternative to studying them one by one. The chemical and physical properties of these compounds depend on the structural features that these compounds have. The traditional method of studying organic chemistry is to divide the huge number of compounds into classifications based on individual structural features and then to study these classifications, thus simplifying the study of both structure and properties. This will be our approach here. The structural features that place compounds into their particular classifications are called **functional groups**. The first classification we will study is the group called the hydrocarbons.

The vast majority of organic compounds have hydrogen atoms bonded to carbon atoms. Those that have only carbons and hydrogens and no other elements are called hydrocarbons. Functional groups within the broader hydrocarbon classification include the carbon–carbon single bond, the carbon–carbon double bond, and the carbon–carbon triple bond. Compounds with these functional groups fall under the general heading of aliphatic hydrocarbons. Aliphatic hydrocarbons that have nothing but single bonds between carbons are called alkanes. They are also called saturated aliphatic hydrocarbons since

they contain as many hydrogens as they can have. Aliphatic hydrocarbons that have at least one double bond between carbons are called alkenes, the carbon–carbon double bond being the functional group. They are also called olefins. Aliphatic hydrocarbons that have at least one triple bond between carbons are called alkynes, the carbon–carbon triple bond being the functional group. Compounds that are either alkenes or alkynes are unsaturated aliphatic hydrocarbons, since more hydrogens may be added by eliminating the multiple bonding between carbons. In addition to the aliphatic hydrocarbons, we also have the aromatic hydrocarbons. Compounds that are aromatic hydrocarbons possess a structural feature such as the benzene ring, which is what we will define as the aromatic hydrocarbon functional group. Figure 14.8 summarizes these facts in the form of a flow chart.

Let us now look at some examples of compounds within each of these classifications, beginning with the alkanes, since the International Union of Pure and Applied Chemistry (IUPAC) nomenclature systems for other classifications is based on the alkane group.

14.6.1 Alkanes

The simplest alkane is one with just one carbon atom. The four bonds on the one carbon are all to hydrogen atoms. The name of this compound is methane (Figure 14.9a). Methane is the major component of natural gas.

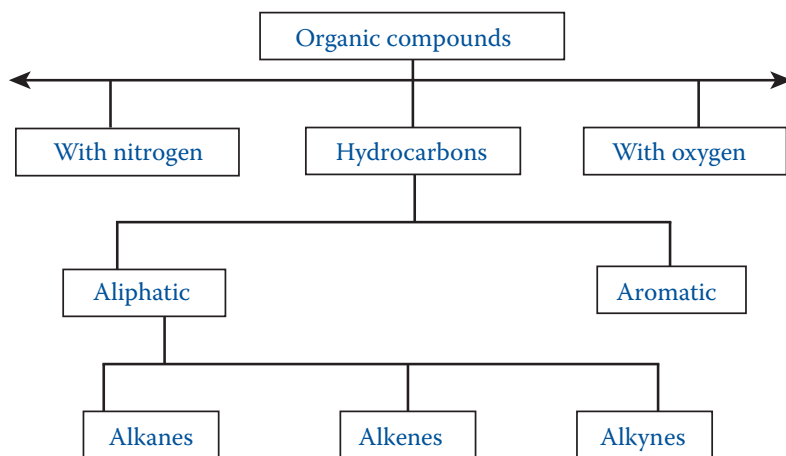


FIGURE 14.8 A flow chart showing the details of the hydrocarbon classification.

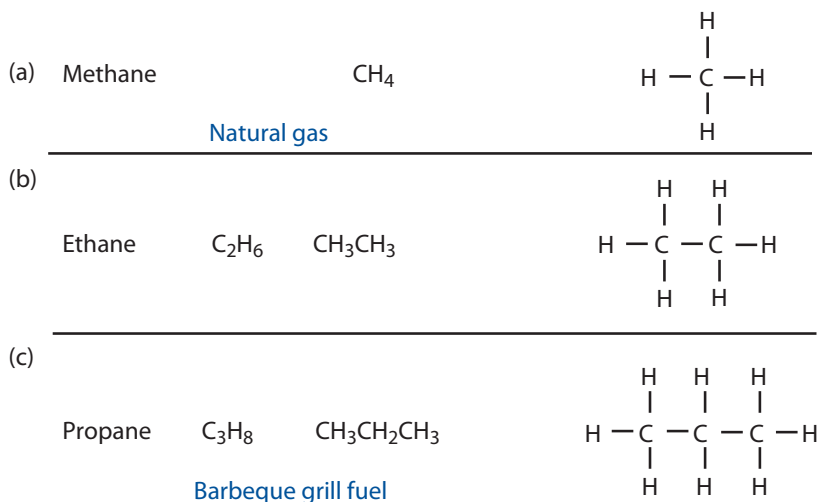


FIGURE 14.9 The names, condensed structures, and bond line structures of three simple alkanes (a) methane, (b) ethane, and (c) propane. Practical uses of methane and propane are also given.

The next simplest alkane is ethane. It has two carbon atoms and six hydrogen atoms. Three of the four bonds on each carbon are to hydrogens and one is to the other carbon (Figure 14.9b). The structures shown in the center in this figure for methane and ethane are called condensed molecular structures. In a condensed structure, bond lines are optional and are often not shown. Rather the atoms to the right of each carbon are the atoms bonded to that carbon. In addition, the carbon atom or atoms adjacent to a given carbon are bonded to it. Comparing the bond line structure on the right for ethane to the corresponding condensed structure just to its left should make this clear. Three hydrogens and the other carbon are bonded to the first carbon in the condensed structure. The other three hydrogens are bonded to the other carbon.

The alkane with three carbons is propane (Figure 14.9c). Like methane, propane is also a fuel used to heat homes, particularly in rural areas, and is also used as a fuel for outdoor barbeque grills.

14.6.2 Isomers

As we proceed from methane to ethane to propane, you may wonder why, in order to create the next alkane in the sequence, we always replace the hydrogen on the right with the “ CH_3- ” group. (a “ CH_3- ” group is called a methyl

group). It would appear that this is done to maintain a straight chain of carbons. Let us see what would happen if we replace one of the other two hydrogens in the process of progressing from ethane to propane (Figure 14.10). Figure 14.10a shows the same straight-chain structure of propane seen in Figure 14.9c. Now consider an expansion of the bonding around the second carbon (Figure 14.10b) such that instead of replacing the hydrogen directly to the right with the methyl group we instead replace one of the other two hydrogens on the second carbon. This results in either of the two structures on the right in Figure 14.10b. The fact is all three of these structures are identical. It is as if we held the first carbon stationary and rotated the rest of the structure around the bond between the first two carbons. Rotation of a single bond in this way does not create a different structure. All three structures drawn in Figure 14.10b are identical, so all three can be represented by the structure given in Figure 14.10a. The same can be said if we attempt to make a propane chain by replacing a hydrogen on the left carbon in the ethane structure. The three-carbon chain, or the propane structure, is what it is, no matter how we attempt to add the third carbon to the ethane structure.

We cannot say the same about the four-carbon alkane. The fact is we can draw two different structures (Figure 14.11). One is the four-carbon chain in which the methyl group has been attached to the end of a propane chain, and the other is the structure obtained by attaching the methyl group to the middle

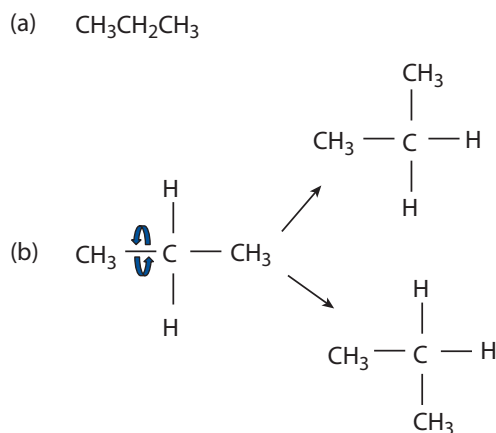


FIGURE 14.10 (a) Condensed structure of propane. (b) Rotation around the bond indicated with the semicircular arrows does not result in different structures for propane. The three structures shown here are identical.

of the propane chain. The structures are named n-butane and isobutane, as indicated in Figure 14.11. Ball-and-stick models of these two structures are given in Figure 14.12. Notice that they both have the same formula, C_4H_{10} . Despite this, they are different compounds, each with their own unique set of properties. They are what we call **isomers**—same formula but different structures. Isomers abound in organic chemistry, and as the number of carbons increases, the number of possible isomers increases as well.

In Figures 14.13 and 14.14, we can see that extending the length of the chain rather than adding the methyl group to the middle carbon does not create a different structure, no matter how the fourth carbon is oriented in space. As

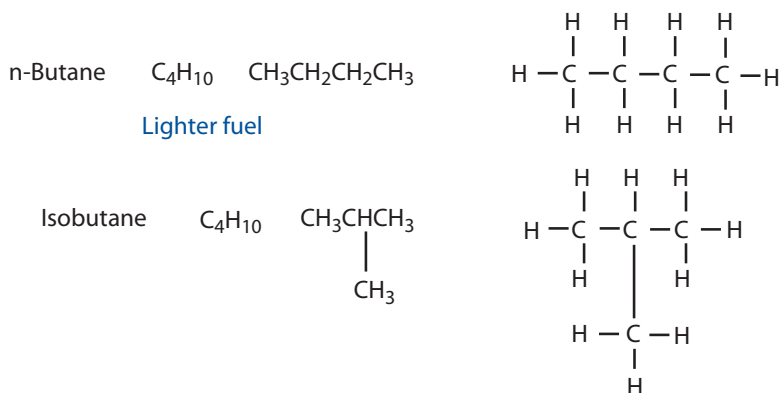


FIGURE 14.11 The names, formulas, condensed structures, and bond line structures of n-butane and isobutane. These two compounds are isomers since they have the same formula but different structures. A practical use of n-butane is also given.

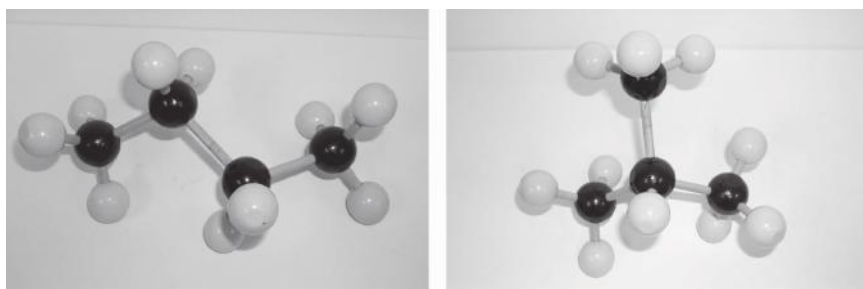


FIGURE 14.12 Ball-and-stick models of n-butane (left) and isobutane (right).

with propane, the structures in Figures 14.13 and 14.14 simply result from the rotation around a single bond, in this case, the bond between the middle two carbons. So these two structures are *not* isomers. They are identical structures and therefore the same compound.

Isobutane is an example of a branched chain structure; the methyl group attached to the center carbon of a three-carbon chain is referred to as a branch. Branches can consist any length of carbon chains, and the possibility of branching increases greatly as we proceed to longer chains of carbons—a branch can exist on any and all carbons in a chain! A very large number of different structures may be drawn as the number of carbons in an alkane gets very large. It quickly becomes easy to understand how there can be millions of compounds under the general heading of organic chemicals. And we have seen only one classification so far!

Isobutane is actually named with propane as the base name under the official naming rules adopted by the IUPAC, since the longest straight chain of carbons in isobutane is three. A detailed look at the IUPAC nomenclature

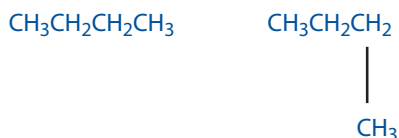


FIGURE 14.13 Two condensed structures, both representing n-butane (i.e., not isomers).

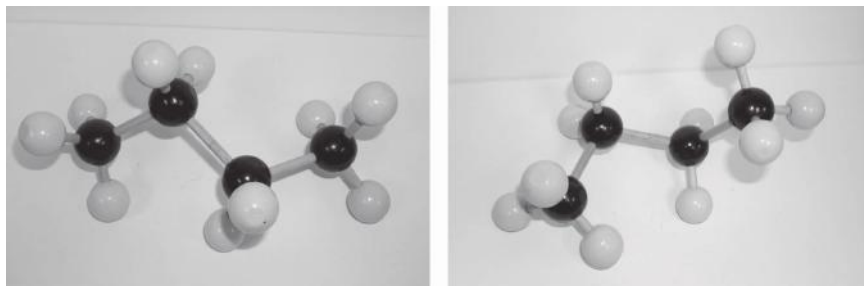


FIGURE 14.14 Ball-and-stick models of the two structures in Figure 14.13. To make the model on the right, the single bond between the middle two carbons of the model on the left has been rotated. The point is that single bond rotation does not result in an isomer.

TABLE 14.1 IUPAC Names for Alkanes and Alkyl Groups of Various Chain Lengths

Number of Carbons in Chain	Name of Alkyl Group	Name of Alkane
1	Methyl	Methane
2	Ethyl	Ethane
3	Propyl	Propane
4	Butyl	Butane
5	Pentyl	Pentane
6	Hexyl	Hexane
7	Heptyl	Heptane
8	Octyl	Octane
9	Nonyl	Nonane

rules is beyond the scope of this text. However, the base names of alkanes with chains of one to nine carbons are shown in Table 14.1. This table also lists the base names of the so-called alkyl groups for each.

All open-chain alkanes, such as those discussed in this section, fit the general formula C_nH_{2n+2} , where n is the number of carbon atoms in the molecule.

14.6.3 Alkenes

Alkenes are aliphatic hydrocarbons that have at least one double bond between carbons in the structure. The simplest alkene has two carbons. Its IUPAC name is ethene and its common name, also endorsed by the IUPAC, is ethylene (Figure 14.15). Notice that to accommodate the double bond, the carbons each have only two hydrogens. Remember that a carbon can have four bonds—no more and no less. This means that the carbons that are part of a double bond only have two remaining bonds each for use for other atoms such as hydrogen.

Notice, too, that the formal IUPAC name, ethene, is derived from the alkane with two carbons, ethane. The *a* in *-ane* changes to *e*, just as in the word *alkene*, which is derived from the word *alkane*.

The next simplest alkene (Figure 14.15) is propene, or propylene, which has three carbons. Following the same pattern as ethene, the formal IUPAC name, propene, is derived from the name of the alkane with three carbons, propane, with the *a* in *-ane* changed to *e*. The middle carbon of this three-carbon

ethane or ethylene	C_2H_4	$CH_2 = CH_2$
propene or propylene	C_3H_6	$CH_2 = CHCH_3$
1-butene	C_4H_8	$CH_2 = CHCH_2CH_3$
2-butene	C_4H_8	$CH_3CH = CHCH_3$

FIGURE 14.15 The names and condensed structures of four simple alkenes.

chain—and one of the carbons that has a double bond—has only one hydrogen attached. This is because the fourth bond is taken up by the methyl group on the right.

The double bond in this three-carbon chain could also be placed between the second and third carbons. However, this would not be a different structure, since the double bond would be between the first two carbons in a three-carbon chain either way. It does not matter from which end one counts.

The alkene with four carbons, however, is a different story. See the two butane structures in Figure 14.15. The two different structures can be drawn. They are different because of the location of the double bond. These two structures are isomers, since they have the same molecular formula (C_4H_8) but a different structure. Remember when drawing organic structures that all carbons must have exactly four bonds. Notice that a number and a hyphen to the left of the name butene indicates the carbon in the chain on which the double bond first appears (counting from the left). The lowest number possible is always used. For example, if we had counted from the right, 1-butene would have been called 3-butene, which is incorrect.

As the chain gets longer, the double bond can appear between any two carbons, a fact that implies many possible isomers. That, coupled with the possibility of branching and the possibility of more than one double bond being present in a given structure, indicates that there are a large number of alkenes.

Replacing a single bond with a double bond reduces the number of hydrogens in the structure by two. Thus, all alkenes with just one double bond fit the general formula C_nH_{2n} in which n is the number of carbons in the chain.

14.6.4 Alkynes

The alkyne classification follows the same pattern as the alkene classification. The simplest alkyne has two carbons and is named ethyne (formal IUPAC name), or acetylene (common IUPAC name). See Figure 14.16. The name acetylene

ethyne or acetylene welder's fuel	C_2H_2	$CH\equiv CH$
propyne	C_3H_4	$CH\equiv CCH_3$
1-butyne	C_4H_6	$CH\equiv CCH_2CH_3$
2-butyne	C_4H_6	$CH_3C\equiv CCH_3$

FIGURE 14.16 The names and condensed structures of four simple alkynes. A practical use of ethyne, or acetylene, is also given.

may be familiar. This compound is the gas that welders use in their torches. The IUPAC name, ethyne, is derived from the name of the base alkane, ethane, just as was done with the alkene, ethene. In this case, the *a* changes to *y*, as was done in the word *alkyne*. This is true of the formal IUPAC names of all alkynes.

Note that each carbon atom in ethyne has just one bond available for bonding to other atoms. In the case of ethyne, this one bond is to a hydrogen on each carbon.

The next simplest alkyne is propyne (Figure 14.16). Just as with propene, there are no isomers. The one bond remaining after the formation of the triple bond is taken up by a methyl group on one of the carbons and by a hydrogen on the other.

When there are four carbons, there is, once again, a pair of isomers (Figure 14.16). And when more carbons are considered—and since branching and more than one triple bond are possible in a structure—one can see that many alkynes can exist. Replacing a single bond in a structure with a triple bond reduces the number of hydrogens by four. Thus, all alkynes with just one triple bond fit the general formula C_nH_{2n-2} , in which *n* is the number of carbons in the chain.

14.6.5 Ring Structures

As stated previously, most aromatic hydrocarbons possess a unique structural feature called the benzene ring. The benzene ring is a structure that has six carbon atoms bonded together in a circle, or ring. It is not unusual to encounter other ring structures in organic chemistry; many aliphatic hydrocarbons have ring structures. An example is cyclohexane, which is a six-membered ring with a molecular formula of C_6H_{12} . See Figure 14.17 and the ball-and-stick model in Figure 14.18.

Other aliphatic rings are common, including cyclopentane (with five carbons) and cycloheptane (with seven carbons). Forming a cycloalkane ring from

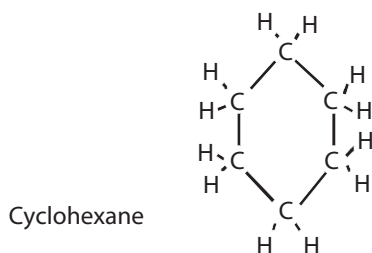


FIGURE 14.17 Bond line structure of cyclohexane.

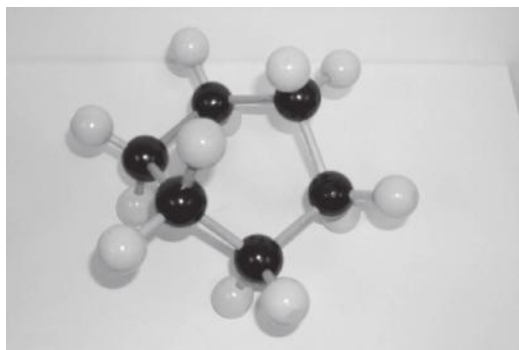


FIGURE 14.18 Ball-and-stick model of cyclohexane.

an open chain (i.e., no ring) reduces the number of hydrogens by two. Thus the general formula of a cycloalkane with just one ring and no multiple bonds is the same as that of alkenes, C_nH_{2n} . In addition, hydrogens can be removed and replaced by other structural features, such as double bonds, branches, and chains, making many different structures possible.

The cyclohexane structure is structural features drawn simply as a hexagon and the cyclopentane structure as a pentagon. In such drawings, it is understood that there is a carbon in every corner and that there are two hydrogens bonded to each carbon, even though neither the hydrogens nor the bonds are shown. Such a structure is sometimes referred to as a bond-line structure (Figure 14.19a).

It is possible to have a double bond in a ring—as with, for example, cyclopentene and cyclohexene (Figure 14.19b).

Now let us examine the aromatic hydrocarbons (Figure 14.20). Again, the benzene ring is the structural feature that identifies a hydrocarbon as aromatic (i.e., it is the aromatic hydrocarbon functional group). The benzene

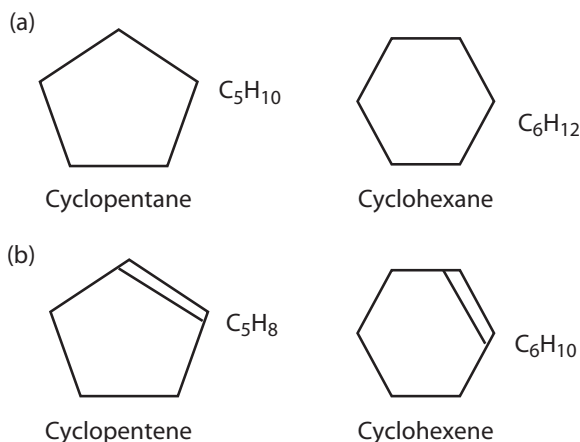


FIGURE 14.19 Bond line ring structures with the C's and H's removed. (a) The two cyclic alkanes and (b) the two cyclic alkenes discussed in the text.

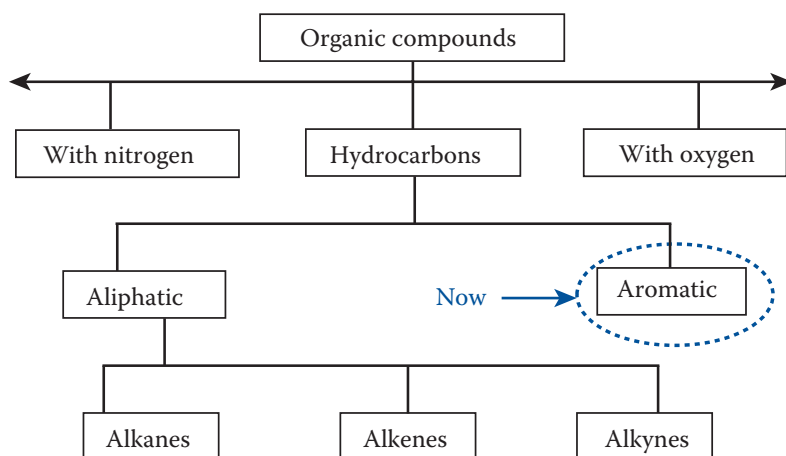


FIGURE 14.20 A flow chart of the hydrocarbon classification emphasizing the aromatic classification to be discussed next in the text.

ring is also a six-membered ring, but it is very different from cyclohexane and cyclohexene. The bonds between the carbons are not single bonds. The geometry around each carbon in the ring is identical to that around carbons to which a double bond is attached (i.e., planar). This means that each of the six carbons have a p orbital that overlaps side to side with the p orbital of the carbon next to it. Scientists originally thought that this results in alternating

double and single bonds around the ring, or π bonds on each carbon, which is the triene structure on the left in Figure 14.21. But the properties of compounds with these ringed structures indicate that there are no double bonds present at all. For example, compounds with the benzene ring in the structure do not undergo the reactions that typically characterize alkenes. Studies indicate that the π electrons are not localized between the carbons as in the triene structure but rather delocalized around the entire ring. The overlapping of the p orbitals occurs around the entire ring and creates donut-shaped molecular orbital regions above and below the ring. These electron-rich clouds are responsible for the unique properties of aromatic compounds. The most telling drawing for such a unique structure is a hexagon with a circle in the center, as shown on the right in Figure 14.21. This ring—with no substituent, chain, or branch attached to any carbon—is also an acceptable drawing for benzene, or the benzene ring, and is the one we will use in upcoming discussions in this chapter. A single hydrogen is bonded to each carbon (at each corner of the hexagon). The molecular formula of benzene is therefore C_6H_6 .

As with cyclohexane and other aliphatic hydrocarbons, the hydrogens attached to the ring of benzene can be removed and replaced with other substituents, branches, and chains. A common example of this is a hydrogen being replaced by a methyl group. The IUPAC name for this compound is methylbenzene and the common name, endorsed by the IUPAC, is toluene. See Figure 14.22 for the structures of these two simple aromatic hydrocarbons, benzene and toluene. Two views of the ball-and-stick model of benzene are shown in Figure 14.23. The view on the right uses springs to indicate double bonds and demonstrates the planar geometry exhibited around the entire ring.

Once again, we have a classification—when considering all the possibilities for substituting and branching—that makes it possible for there to be many compounds added from this group to the large total number of organic compounds that exist.

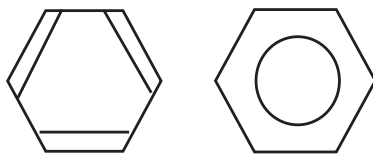


FIGURE 14.21 Representations of the compound benzene. The structure on the right is used in subsequent figures in this text to represent a benzene ring.

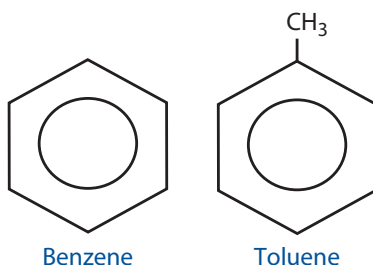


FIGURE 14.22 The bond line structures of two common aromatic hydrocarbons, benzene and toluene.

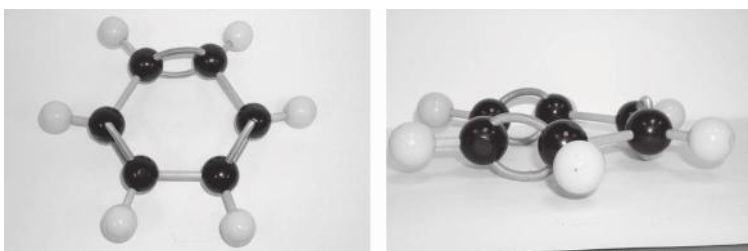


FIGURE 14.23 Two different views of the ball-and-stick model of benzene. The model on the right demonstrates that the structure is planar.

14.7 Organic Compounds Containing Oxygen

As stated previously, oxygen, nitrogen, and other elements are often a part of the structure of organic compounds. The organic compounds containing oxygen warrants division into several classifications. These include functional groups and compounds that may be familiar to you. We will take a close look at these groups (Figure 14.24).

To facilitate an organized look at these classifications, we will use an abbreviation that will allow us to focus on the functional group itself and to help represent a general picture of each classification. This abbreviation is called an R-group. An R-group represents a hydrocarbon grouping, or any organic grouping of elements. The functional group containing oxygen will be shown as attached to it. The R-group may be a simple group, such as a methyl group, but it may also represent something much more complicated. In some cases, it may represent a hydrogen. As we proceed through our study of the classifications, we will encounter some of the other relatively simple examples of

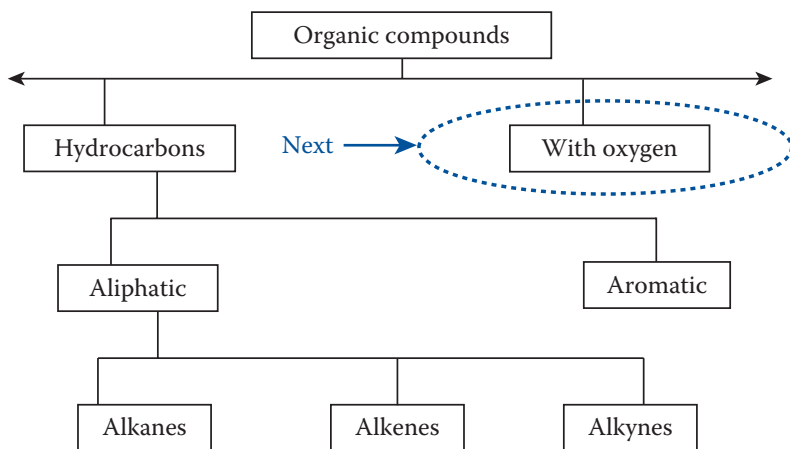


FIGURE 14.24 A flow chart emphasizing the oxygen classification to be discussed next in the text.

these alkyl groups that we listed in Table 14.1, including the ethyl and propyl groups. We will also encounter the isopropyl group, which is a three-carbon chain bonded at the middle carbon.

14.7.1 Alcohols

Let us begin with the alcohols (Figure 14.25). The functional group for alcohols is the -OH group. Using the R-group representation, the general alcohol structure would be R-OH . The common system for naming alcohols is

Name of R-group + “alcohol”

and the formal IUPAC naming system is based on the name of the alkane with the same number of carbons where the *e* at the end of the name of the alkane is dropped and replaced with *ol*. The simplest alcohols and their names are given in Figure 14.25.

Methanol is also known as wood alcohol because it can be distilled from wood. It is the major ingredient in automobile windshield washer fluid and is the automobile gasoline additive (under several different trade names) that helps car engines start in cold weather. Methanol has a low freezing point (and thus does not freeze while stored in the container that holds the windshield washer fluid under the hood of your car) and is completely miscible with water and gasoline, serving to dissolve water residues in gasoline tanks and thus assist in starting your car. It is extremely toxic if taken internally.

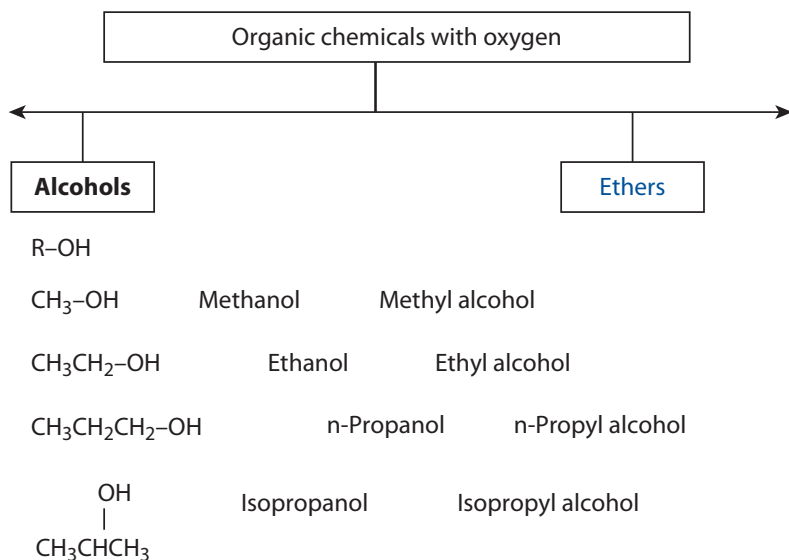


FIGURE 14.25 A flow chart emphasizing the alcohol classification and giving the condensed structures and IUPAC names (left) and common names (right) of four simple alcohols.

Ethanol is also known as grain alcohol because it can be distilled from grains, such as corn. It is the alcohol that is found in alcoholic beverages, and Ethanol is also added to some gasoline formulations.

Isopropyl alcohol is the alcohol in rubbing alcohol. It is used as an antiseptic, and in microbiology laboratories it is used to kill bacteria on benchtops and numerous materials. Isopropyl alcohol is toxic if taken internally. Notice that n-propyl alcohol and isopropyl are isomers.

As with the hydrocarbons, many more alcohols exist because of the large number of possibilities for branches and chains in the structures of the R-group.

14.7.2 Ethers

The classification known as ethers have the following general structure, or functional group: R-O-R (Figure 14.26). The common naming system for ethers simply names the two R-groups, using the prefix *di-* if both R-groups are the same, with the word *ether* following. The formal IUPAC system is slightly more complex. The simpler of the two R-groups is named with *-ane* replaced by *-oxy*, as if this R-O- is attached to a carbon of the other group; the

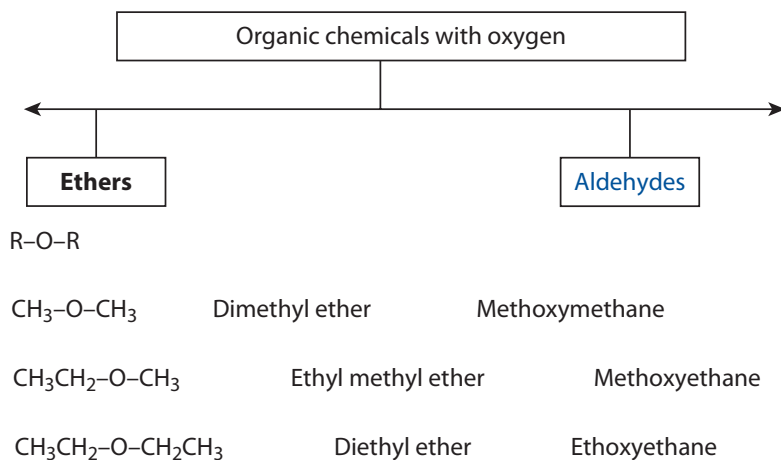


FIGURE 14.26 A flow chart emphasizing the ether classification and giving the condensed structures and common names (left) and IUPAC names (right) of three simple ethers.

name of the other group is then written as the base name. The naming of some simple ethers according to these schemes is also shown in Figure 14.26.

Diethyl ether is the ether that was formerly used in hospitals as an anesthetic. It is also known as ethyl ether and simply as ether. It has some very noteworthy uses, as well as hazards, in the laboratory. It is highly flammable, and because it evaporates readily, a dangerous situation can be created in a laboratory because a room filled with its vapor presents an explosion hazard. In addition, it decomposes readily upon standing and turns into an explosive product. Ether must be stored in metal containers in refrigerators to slow its decomposition and must be used or discarded within 9 months of its acquisition. The refrigerators in which it is stored should be explosion-proof.

Notice that dimethyl ether is an isomer of ethyl alcohol. Other ethers are also isomers of alcohols. Notice also that a large number of ethers exist, due to the large number of possible variations that can occur in the R-group.

14.7.3 Aldehydes

The four remaining classifications of oxygen-containing organic compounds have a common structural feature known as the **carbonyl group**, a carbon doubly bonded to an oxygen. The general structure, or functional group, of the aldehyde classification has one bond from the carbonyl group carbon

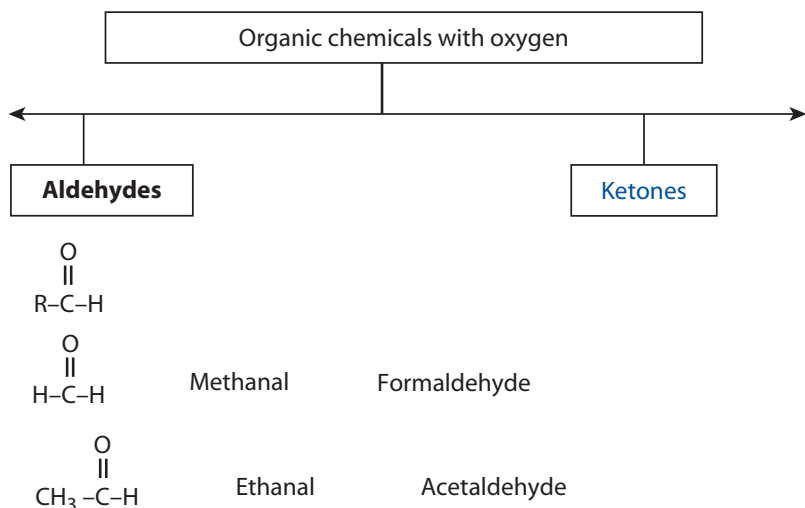


FIGURE 14.27 A flow chart emphasizing the aldehyde classification and giving the condensed structures and IUPAC names (left) and common names (right) of two simple aldehydes.

to an R-group and the other bond to a hydrogen (Figure 14.27). In the simplest aldehyde, the R-group is a hydrogen. The common name of an aldehyde uses the prefix *form-* to designate one carbon (the carbonyl carbon) and *acet-* to designate two carbons. This indicates that there is some systematic nature to the common naming system, which we shall see is used with the other three remaining oxygen-containing classifications as well. The formal IUPAC naming system for aldehydes changes the *e* on the base alkane to *al* (Figure 14.27).

Formaldehyde has been widely used as a preservative for biological specimens, although that use has been largely discontinued because it is a suspected carcinogen. It is, however, used in embalming fluids.

Again, with branches and chains creating a large number of possible variations in the R-group, a large number of aldehydes exist.

14.7.4 Ketones

Ketones are compounds that have two R-groups—which may or may not be the same—attached to the carbonyl group carbon as their functional group (Figure 14.28). The simplest ketone has both R-groups as methyl groups. The common name of this compound is acetone. Previously, we indicated

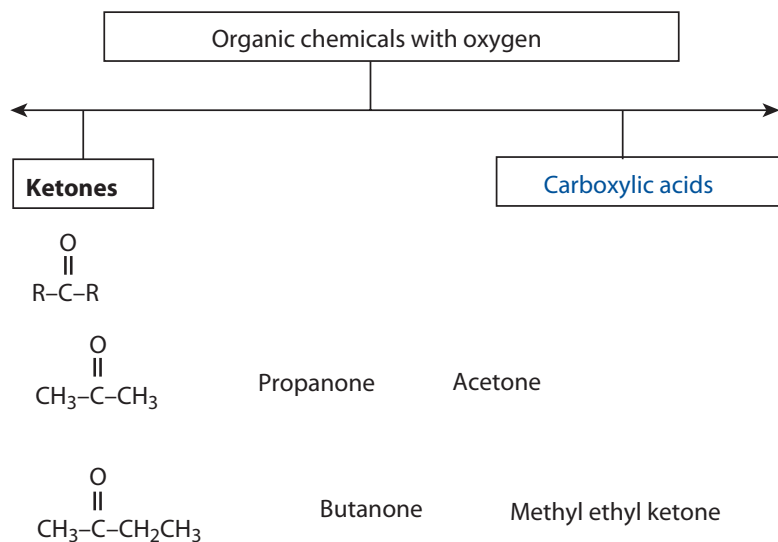


FIGURE 14.28 A flow chart emphasizing the ketone classification and giving the condensed structures and IUPAC names (left) and common names (right) of two simple ketones.

that the *acet-* prefix designates two carbons. Acetone represents an exception to that rule, since it has three carbons. This compound may also be named like the ethers: dimethyl ketone. The IUPAC name changes the *e* of the base alkane to *one*. Thus the IUPAC name for acetone is propanone (Figure 14.28). Acetone has been used as a primary ingredient in many fingernail polish removers. The four-carbon ketone in Figure 14.28 is also fairly common. This compound is commonly abbreviated as MEK. Both MEK and acetone are common laboratory solvents because they have intermediate polarity and are capable of mixing with and dissolving both polar and nonpolar substances.

14.7.5 Carboxylic Acids

The final two classifications have a structural feature in common that is derived from the carbonyl group. It is called the **carboxyl group** and has an oxygen bonded to the carbonyl carbon. The first of these two classifications is the carboxylic acid classification, in which the functional group consists of an R-group attached to the carbon of the carboxyl group and a hydrogen attached to the oxygen (Figure 14.29). These compounds are acids because

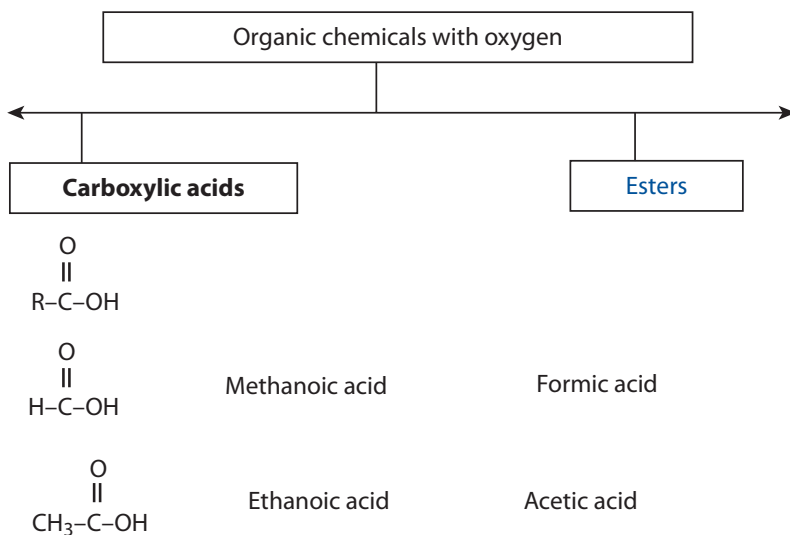


FIGURE 14.29 A flow chart emphasizing the carboxylic acid classification and giving the condensed structures and IUPAC names (left) and common names (right) of two simple carboxylic acids.

the hydrogen shown in their carboxylic acid group is acidic. This means it can be lost easily in water solution or in reaction with a base. Since these compounds are true acids, they have the H presented first in their chemical formula like all other acids. The simplest carboxylic acid also has an H for the R-group. The *form-* and *acet-* prefixes are used in the common naming system, as discussed for the previous classifications. In the formal IUPAC system, the *e* of the base alkane changes to *oic* and is followed-up by the word *acid* (Figure 14.29).

Vinegar is about 5% acetic acid, which is the ingredient that gives vinegar its sour taste. The molecular formula for formic acid is written HCHO_2 , while that for acetic acid is $\text{HC}_2\text{H}_3\text{O}_2$. Again, the R-group can be any simple or complex organic group, and thus the number of carboxylic acids in existence is large.

14.7.6 Esters

The final group of oxygen-containing organic compounds that we will list here is the esters. The general structure has a carboxyl between two R-groups (Figure 14.30). Again notice that, as with the ketones, there are two R-groups

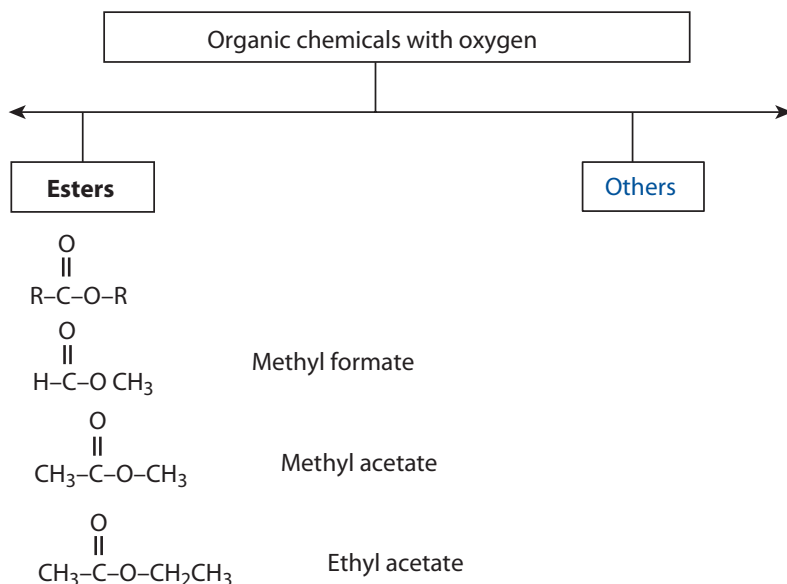


FIGURE 14.30 A flow chart emphasizing the ester classification and giving the condensed structures and common names of three simple esters.

in an ester, and they need not be the same. The simplest ester is one in which the R-group bonded to the carbon is a hydrogen and the one bonded to the oxygen is a methyl group (Figure 14.30). We also show in this figure the ester in which both R-groups are methyl groups and the ester in which the R-group bonded to the carbon is a methyl group and the R-group bonded to the oxygen is an ethyl group. The common naming system for esters uses the *form-* and *acet-* prefixes followed by *-ate*, a system similar to inorganic salts in which the *-ic* ending of an acid changes to *-ate* when naming the salt. In the IUPAC system, the *-ic* of the acid simply changes to *-ate*. Thus we have the names shown in Figure 14.30. Ethyl acetate (shown in the figure) is an ingredient in some fingernail polish removers.

Here again, the R-groups can be simple, like those shown here, or complex organic groupings. This means that there exists many compounds that can be classified as esters.

Esters are responsible for the pleasant odors and unique flavors of many fruits. For example, pentyl acetate is found in bananas, octyl acetate is found in oranges, and isobutyl formate is found in raspberries.

TABLE 14.2 Summary of Hydrocarbon and Oxygen Classifications

Name of Classification	Functional Group	Simple Example	IUPAC Name	Common Name
Alkane	C—C	$\text{CH}_3\text{—CH}_3$	Ethane	Ethane
Alkene	$\text{C}=\text{C}$	$\text{CH}_2=\text{CH}_2$	Ethene	Ethylene
Alkyne	$\text{C}\equiv\text{C}$	$\text{CH}\equiv\text{CH}$	Ethyne	Acetylene
Aromatics			Methylbenzene	Toluene
Alcohol	R—OH	$\text{CH}_3\text{CH}_2\text{—OH}$	Ethanol	Ethyl alcohol
Ether	R—O—R'	$\text{CH}_3\text{—O—CH}_3$	Methoxymethane	Dimethyl ether
Aldehyde	$\begin{array}{c} \text{O} \\ \\ \text{R—C—H} \end{array}$	$\begin{array}{c} \text{O} \\ \\ \text{CH}_3\text{—C—H} \end{array}$	Ethanal	Acetaldehyde
Ketone	$\begin{array}{c} \text{O} \\ \\ \text{R—C—R} \end{array}$	$\begin{array}{c} \text{O} \\ \\ \text{CH}_3\text{—C—CH}_3 \end{array}$	Propanone	Acetone
Carboxylic acid	$\begin{array}{c} \text{O} \\ \\ \text{R—C—OH} \end{array}$	$\begin{array}{c} \text{O} \\ \\ \text{CH}_3\text{—C—OH} \end{array}$	Ethanoic acid	Acetic acid
Ester	$\begin{array}{c} \text{O} \\ \\ \text{R—C—O—R} \end{array}$	$\begin{array}{c} \text{O} \\ \\ \text{H—C—O—CH}_3 \end{array}$	Methyl methanoate	Methyl formate

14.8 Summaries of Hydrocarbon and Oxygen Classifications

Table 14.2 summarizes the hydrocarbon and oxygen classifications and functional groups, giving one simple example for each.

14.9 Polymers

When we introduced organic chemistry in this chapter, we spoke of the broad application of the subject to our everyday lives by referring to a number of

examples of ordinary consumer products that are in fact organic compounds and mixtures of organic compounds including polymers. A polymer is an organic compound whose molecules consist of a very large number of carbon atoms and other atoms that are bonded together. As a result, these molecules have very large formula weights. In addition, polymers have repeating structural units in the chain. While the examples of organic compounds cited in other sections of this chapter have two to six carbon atoms bonded together, polymers can have hundreds and sometimes thousands of carbon atoms. The concept of the backbones or networks introduced and discussed in other sections of this chapter is especially applicable here for polymers because this is indeed is their structure—extremely long chains, branches, or chains and branches of carbons with hydrogens and functional groups bonded at regular intervals and making up a network of atoms that is the molecule. These compounds are not substances just gracing the laboratories of research scientists. They have extremely broad applications in our everyday lives.

These compounds include plastics, foams, paints, varnishes, clothing fibers, and synthetic rubbers and vinyls. In this section, we will give examples of ordinary consumer products that are polymers, discuss the structure of these molecules, and present the reactions by which they form.

14.9.1 Addition Polymers

Polymers classified as addition polymers are the most common as far as applications in our everyday lives. Molded plastic articles, such as plastic bottles and other containers, children's toys, plumbing pipes—virtually anything you can think of that is a plastic—are addition polymers. Also, we can list floor tile, some textile fibers, ropes, rubber gaskets, plastic wrap, and Styrofoam cups among the materials that are composed of addition polymers. It is difficult to imagine what our lives would be like without addition polymers. Let us take a closer look at the structure of these substances.

The word *addition* is derived from the kind of chemical reaction that is used to manufacture these polymers, namely the addition reaction of alkenes. Addition reactions of alkenes involve the adding of a reacting molecule across the double bond of the alkene, as shown in Figure 14.31. In this process, one of the two bonds in the double bond (the π bond) is broken. The carbons that were double bonded become single bonded, and the broken bond is replaced by bonds to the substituents from the reacting molecule, A and B. The molecule represented by A–B can be H–H, Cl–Cl, H–Cl, and so on, which results



FIGURE 14.31 An illustration of the addition reaction that alkenes undergo.

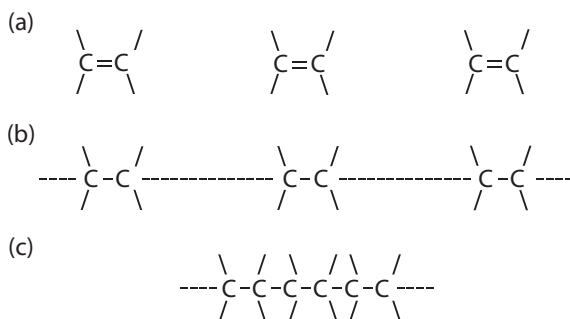
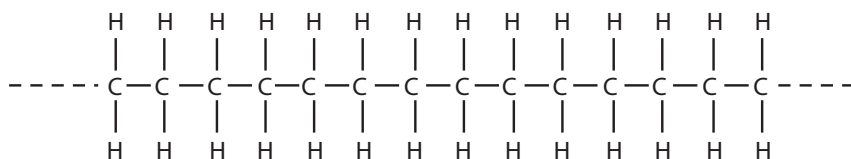
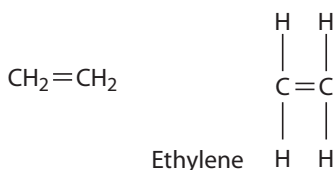


FIGURE 14.32 The formation of an addition polymer. (a) Separate alkene molecules. (b) Alkene molecules undergoing addition reactions. (c) Addition polymer formed.

in organic products that are not polymers. However, under certain conditions, the broken bond can instead be replaced by bonds to carbons from other alkene molecules that have undergone the same change, as in the sequence shown in Figure 14.32. This addition can continue in both directions, with alkene molecules adding to other alkene molecules, until a very long molecule, a polymer molecule, is formed.

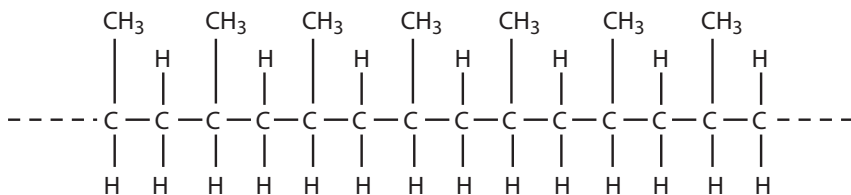
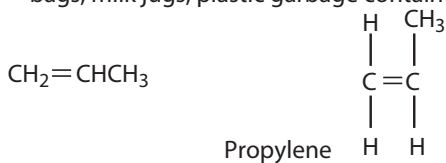
There can be many variations on this general description because the substituents bonded to the four carbons in the double bonds can be quite variable. For example, all four bonds may be to hydrogens ($\text{CH}_2 = \text{CH}_2$, ethene, or ethylene), there may be one methyl group attached ($\text{CH}_2 = \text{CHCH}_3$, propene, or propylene), there may be a chlorine attached ($\text{CH}_2 = \text{CHCl}$, vinyl chloride), and so on. Polyethylene, polypropylene, polyvinyl chloride (PVC), and many other addition polymers have been manufactured in mass quantities by this approach and used for many consumer products. Figure 14.33 lists some of the addition polymers that have been manufactured. Also listed are the individual alkene units (monomers) that are in these polymers and some of the uses of each.

(a)



Polyethylene: plastic bottles, plastic shopping bags, milk jugs, plastic garbage containers

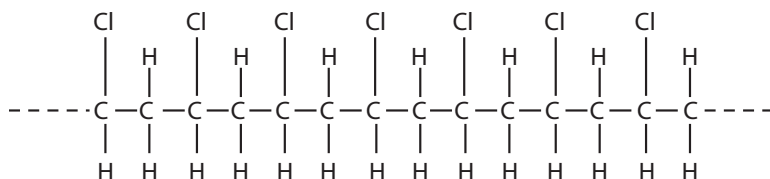
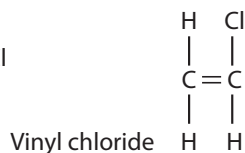
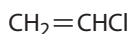
(b)



Polypropylene: plastic packaging, plastic rope, plastic chairs, plastic garbage containers, plastic rugs and mats

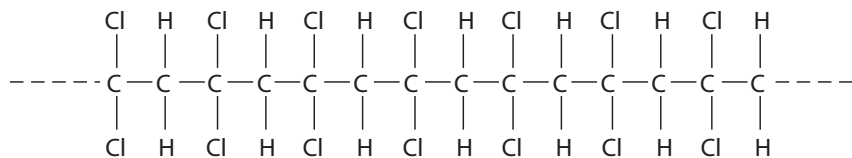
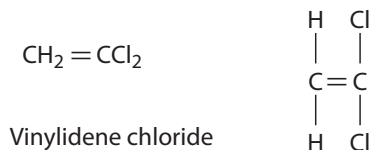
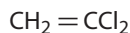
FIGURE 14.33 Examples of the formation of addition polymers. Practical uses of the polymers are also given. (a) Polyethylene forms from the polymerization of ethylene. (b) Polypropylene forms from the polymerization of propylene. (c) Polyvinyl chloride (PVC) forms from the polymerization of vinyl chloride. (d) Saran forms from the polymerization of vinylidene chloride. (e) Teflon forms from the polymerization of tetrafluoroethylene. (f) Polystyrene or styrofoam forms from the polymerization of styrene.

(c)



Polyvinyl chloride (PVC): flexible hoses, floor tile, plastic pipes, vinyl siding, rain gutters

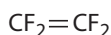
(d)



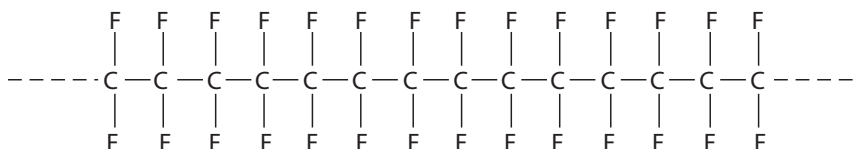
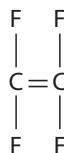
Polyvinylidene chloride, (a.k.a. Saran): saran wrap

FIGURE 14.33 (Continued) Examples of the formation of addition polymers. Practical uses of the polymers are also given. (a) Polyethylene forms from the polymerization of ethylene. (b) Polypropylene forms from the polymerization of propylene. (c) Polyvinyl chloride (PVC) forms from the polymerization of vinyl chloride. (d) Saran forms from the polymerization of vinylidene chloride. (e) Teflon forms from the polymerization of tetrafluoroethylene. (f) Polystyrene or styrofoam forms from the polymerization of styrene.

(e)

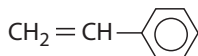


Tetrafluoroethylene

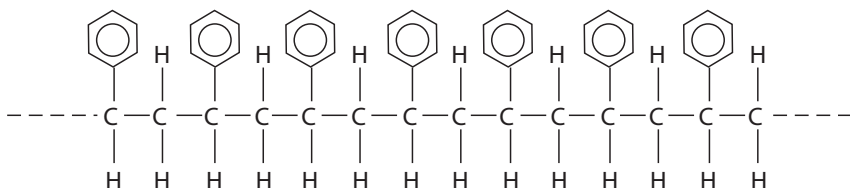
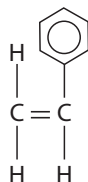


Polytetrafluoroethylene (PTFE; a.k.a. Teflon):
cookware coating, teflon tape

(f)



Styrene



Polystyrene (a.k.a. Styrofoam): Styrofoam cups,
Styrofoam insulation, Styrofoam packaging

FIGURE 14.33 (Continued) Examples of the formation of addition polymers. Practical uses of the polymers are also given. (a) Polyethylene forms from the polymerization of ethylene. (b) Polypropylene forms from the polymerization of propylene. (c) Polyvinyl chloride (PVC) forms from the polymerization of vinyl chloride. (d) Saran forms from the polymerization of vinylidene chloride. (e) Teflon forms from the polymerization of tetrafluoroethylene. (f) Polystyrene or styrofoam forms from the polymerization of styrene.

14.10 Homework Exercises

Hybridization, Geometry, and Physical Properties

1. Give two different definitions of organic chemistry. Where do most of the chemicals we call organic compounds come from?
2. What two kinds of atoms are especially abundant in the structures of organic compounds?
3. What is meant by a backbone or network of carbon atoms?
4. Define hybridization.
5. Explain why hybridization is important in explaining the bonding of carbon to other atoms.
6. Sometimes the geometry around a bonded carbon atom is tetrahedral, sometimes it is trigonal planar, and sometimes it is linear. Under what conditions is the geometry tetrahedral? Under what conditions is the geometry trigonal planar? Under what conditions is the geometry linear?
7. List four physical properties that organic compounds generally have that inorganic compounds usually do not have.
8. Many organic compounds are liquids. List some physical properties of these liquids that are important for chemists to measure.
9. If you increase the chain length of a hydrocarbon, will the boiling point increase, decrease, or stay the same? Explain your answer.

Hydrocarbons and Isomers

10. What is a functional group?
11. How are aliphatic hydrocarbons different from aromatic hydrocarbons?
12. What is the difference between a saturated hydrocarbon and an unsaturated hydrocarbon?
13. Study the structures of methane, ethane, propane, n-butane, isobutane, and other alkanes, and show that their formulas fit the general formula of alkanes, C_nH_{2n+2} .
14. Study the structures of ethene, propene, 1-butene, 2-butene, and other alkenes, as well as cyclopentane, cyclohexane, and other cycloalkanes, and show that their formulas fit the general formula C_nH_{2n} .
15. Study the structures of ethyne, propyne, butyne, and other alkynes, and show that their formulas fit the general formula for alkynes, C_nH_{2n-2} .

16. State what classification (alkane, alkene, alkyne, cycloalkane, or aromatic) can each of the following compounds belong, and draw possible structures.
- C_3H_6
 - C_4H_6
 - C_6H_6
 - C_8H_{16}
 - C_8H_{10}
17. Draw the structure of the simplest compounds that represent each of the following categories.
- Alkane
 - Alkene
 - Alkyne
 - Aromatic hydrocarbon
18. Draw condensed structures of all the isomers of the alkane with five carbons, pentane (C_5H_{12}). (Hint: there are three isomers).
19. Which of the four structures in Figure 14.34 is not an isomer of this hexane structure?

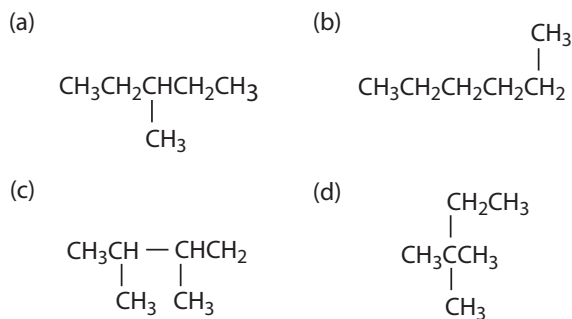
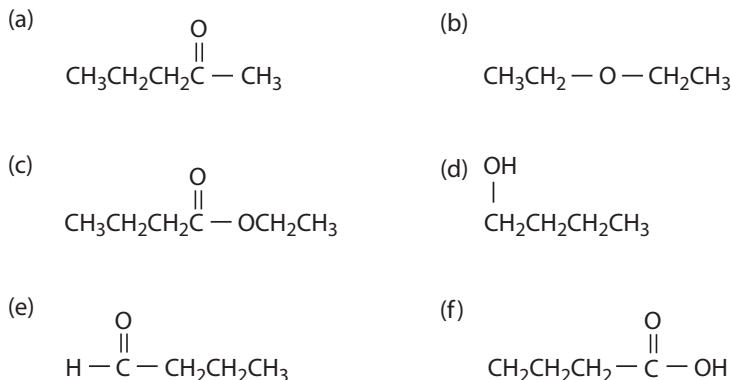
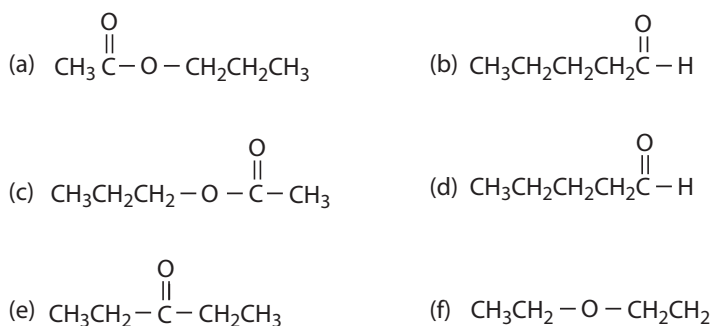


Explain why the suspected structure is not an isomer of hexane.

- What is the difference between a cyclohexane ring and a benzene ring? Explain the conventional way to draw the structures of these two compounds.
- Pick out two structures in Figure 14.35 that are isomers of each other and tell why they are isomers.
- Pick out the two structures in Figure 14.36 that are isomers of each other and tell why they are isomers.

Classification and Naming

- Draw the functional groups that are associated with each of the following oxygen-containing classifications.
 - Alcohol
 - Aldehyde
 - Carboxylic acid
 - Ester
 - Ether
 - Ketone

**FIGURE 14.34** The structures to accompany Exercise 19.**FIGURE 14.35** The structures to accompany Exercise 21.**FIGURE 14.36** The structures to accompany Exercise 22.

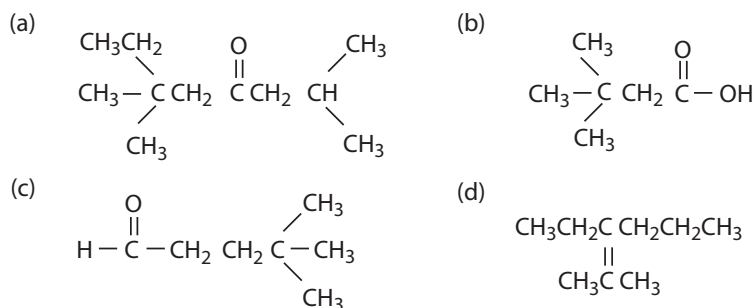


FIGURE 14.37 Four structures to accompany Exercise 24.

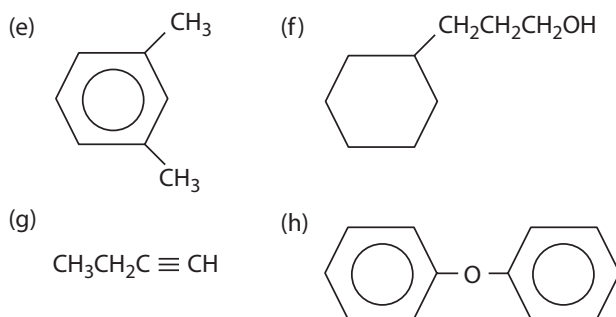


FIGURE 14.38 Four structures to accompany Exercise 24.

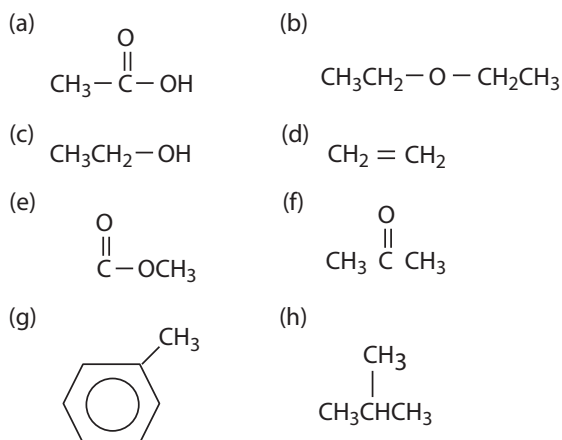
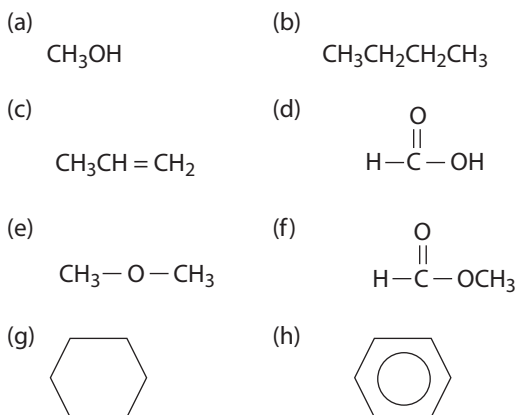


FIGURE 14.39 The structures to accompany Exercise 25.

**FIGURE 14.40** The structures to accompany Exercise 26.

24. Classify each of the structures in Figures 14.37 and 14.38 as one of the following: alkane, alkene, alkyne, aromatic hydrocarbon, alcohol, ether, aldehyde, ketone, carboxylic acid, or ester.
25. Name the compounds given in Figure 14.39.
26. Name the compounds given in Figure 14.40.
27. Draw the structures associated with the following names.
 - a. Acetaldehyde
 - b. n-Propanol
 - c. Propene
 - d. Methyl ethyl ketone
 - e. Ethane
 - f. 2-Butene
28. Draw the structure associated with the following names.
 - a. Acetylene
 - b. Ethyl methyl ether
 - c. Formaldehyde
 - d. Isopropyl alcohol
 - e. Ethyl acetate
 - f. Cyclohexene

Consumer Products

29. Match the organic chemical in the left column with the real-world use in the right column by placing the appropriate letter in the blank.

_____ Methane	a. Preservative for biological specimens

_____ Acetone	b. Alcoholic beverages
_____ Formaldehyde	c. Vinegar
_____ Acetic acid	d. Natural gas
_____ Methanol	e. Rubbing alcohol
_____ Ethanol	f. Windshield-washer fluid
_____ Isopropyl alcohol	g. Fingernail polish remover

30. Name one real-world consumer product in which each of the following is found.
- Acetylene
 - Ethyl alcohol
 - Acetone
 - Methane
 - Polyethylene
 - Methyl alcohol
 - Diethyl ether
 - Isopropyl alcohol
 - Formaldehyde
 - Acetic acid

Polymers

- What is a polymer?
- How is the addition reaction of alkenes related to the formation of addition polymers?
- Write the structure and give the name of the addition polymer that can be derived from the following alkenes.
 - $\text{CH}_2 = \text{CH}-\text{Cl}$
 - $\text{CH}_2 = \text{CH}-\text{CH}_3$
 - $\text{CF}_2 = \text{CF}_2$
- Write the structure and give the name of the alkene that is used to make the following polymers.
 - Polystyrene
 - Polyethylene
 - Saran
- Name three addition polymers, and list one corresponding consumer product that is associated with each.

Appendix: Answers to Homework Exercises

Chapter 1

1. Composition, structure, properties, and changes.
2. Chemical properties involve chemical change; physical properties do not. Example of physical properties: color, odor, texture, physical state, solubility, density, melting point, boiling point, freezing point. Examples of chemical change: irons rusts, wood burns, and so on.
3.
 - a. chemical
 - b. physical
 - c. physical
 - d. physical
 - e. chemical
 - f. chemical
4. Physical change is a change that occurs without any transformation of substance. The substance is the same substance both before and after the change. Examples include change of state (e.g., evaporation, condensation, freezing, melting) and change of texture (e.g., the crushing of a granular solid to a powdery solid).
5. Chemical change is a transformation of one substance into one or more others. Examples include the rusting of iron and the combination of hydrogen and oxygen to make water.
6.
 - a. Chemical change—there has been a transformation of substance
 - b. Physical change—sugar both before and after dissolving.
 - c. Physical change—alcohol both before and after evaporation.

7.
 - a. Physical change—water both before and after evaporation.
 - b. Chemical change—there has been a transformation of substance.
 - c. Physical change—salt both before and after dissolving.
 - d. Chemical change—there has been a transformation of substance.
 - e. Physical change—acetic acid before and after freezing.
8.
 - a. compound
 - b. heterogeneous mixture
 - c. homogeneous mixture
 - d. compound
 - e. element
 - f. homogeneous mixture
 - g. element
 - h. heterogeneous mixture
9. Blanks from top to bottom: D, E, F, A, G, E, F, A, C, A
10. In a compound, elements are chemically combined. In a mixture, the components are not chemically combined, only physically mixed.
11. Because in a compound, the elements are chemically combined, meaning that it is a different substance with properties different from the elements before they were combined. In a mixture, the components are simply physically mixed, meaning that the properties are retained.
12.
 - a. Entirely new properties.
 - b. New substance formed.
 - c. Atoms become chemically combined with other atoms, forming molecules or formula units of the new substance.
13. Two atoms of hydrogen and one atom of oxygen combined to form a molecule of water.
14.
 - a. Na_3PO_4 : Sodium, 3 atoms; phosphorus, 1 atom; oxygen, 4 atoms
 - b. CCl_4 : Carbon, 1 atom; chlorine, 4 atoms
 - c. $\text{Al}_2(\text{SO}_4)_3$: Aluminum, 2 atoms; sulfur, 3 atoms; oxygen, 12 atoms
 - d. $(\text{NH}_4)_2\text{CO}_3$: Nitrogen, 2 atoms; hydrogen, 8 atoms; carbon, 1 atom; oxygen, 3 atoms
 - e. $\text{Ba}(\text{OH})_2$: Barium, 1 atom; oxygen, 2 atoms; hydrogen, 2 atoms
15.
 - a. An atom is the smallest particle of an element that can exist and still be that element.
 - b. A molecule is the smallest particle of a molecular (or covalent) compound that can exist and still be that compound.

- c. An ion is an atom or grouping of atoms that has developed an electrical charge.
16. a. atom
b. molecule
c. ion
d. molecule
e. atom
f. molecule
17. A formula unit is the simplest particle of an ionic compound that can exist and still be the compound. The term is used to differentiate between those compounds that are ionic from those that are molecular or covalent.
18. An atom is the fundamental unit of an element. A molecule is the fundamental unit of a molecular compound. A molecule is composed of atoms in chemical combination.
19. a. Mass is the amount of material as measured by measuring its weight on the Earth's surface.
b. The atomic mass unit is one-twelfth the mass of an atom of the carbon-12 isotope.
c. The atomic weight of an element is the average weight of all atoms that exist for that element.
d. The atomic number of an element is the number of protons in the nucleus of its atom. It is the number that identifies or characterizes a given element.
20. See Table 1.1 in text.
21. An isotope of an element is an atom of that element that has a particular number of neutrons in its nucleus. One isotope of a given element differs from another in this number of neutrons.
22. a. Xe
b. 30
c. neutron
d. 19
e. 42
f. 24
g. protons
h. 20
i. C
j. electron

23. a. Ag
b. 53
c. 28
d. 20
e. 95
f. 82
g. Al
h. mass number
24. a. 127
b. 48
c. Br
d. Cr
e. 27
f. mass number
g. 13
h. 28
i. 37
25. a. Atomic number is 6; mass number is 14; number of neutrons is 8.
b. Atomic number is 26; mass number is 58; number of neutrons is 32.
c. Atomic number is 83; mass number is 209; number of neutrons is 126.
d. Atomic number is 7; mass number is 15; number of neutrons is 8.
e. Atomic number is 17; mass number is 37; number of neutrons is 20.
f. Atomic number is 16; mass number is 36; number of neutrons is 20.
26. a. Atomic number is 88; mass number is 222; number of neutrons is 134.
b. Atomic number is 13; mass number is 27; number of neutrons is 14.
c. Atomic number is 93; mass number is 240; number of neutrons is 147.
d. Atomic number is 15; mass number is 31; number of neutrons is 16.
e. Atomic number is 24; mass number is 54; number of neutrons is 30.

- f. Atomic number is 19; mass number is 40, number of neutrons is 21.
27. No. The different number of protons changes the identity of the element. An atom with nine protons is an atom of fluorine.
28. Atomic weight is the average weight of all the isotopes of an element, not the weight of just one of the atoms.

Chapter 2

1. Figure 2.8: 5.17 units; Figure 2.9: 4.0 units
2. Figure 2.10: (a) 6.58 mL, (b) 6.0×10^2 psi, (c) 24.8 degrees, (d) 6.60 units, (e) 7.46 units
- Figure 2.11: (a) 97 units, (b) 26.0 degrees, (c) 7.00 mL, (d) 6.1×10^2 psi, (e) 7.00 units
- Figure 2.12: (a) 6.10 (top scale), (b) 6.50 mL (c), 30.0 degrees, (d) 6.76 units, (e) 12.2 psi
3. a. 3
b. unknown
c. 3
d. 5
e. 3
f. 1
g. unknown
h. 4
i. 6
j. 4
k. 5
l. 4
m. unknown
n. 3
o. 4
p. 3
q. 4
4. a. 0.000060
b. 1.58×10^3
c. 6.25
d. 103.6

- e. 21.000
- f. 0.037
- g. 1.24×10^4
- h. 16.2
- i. 11.31
- j. 1.2×10^2
- k. 0.0000767
- l. 23.0
- m. 0.50
- 5. a. 25.8 mpg
- b. 3.0×10^1 mpg
- c. 32.00 mpg
- 6. a. 0.035 oz/g; 28 g/oz
- b. 0.0354 oz/g; 28.3 g/oz
- c. 0.035 oz/g; 28 g/oz
- 7. a. 3.0×10^1 mL/fl.oz; 0.034 fl.oz./mL
- b. 3×10^1 mL/fl.oz; 0.03 fl.oz./mL
- c. 29.6 mL/fl.oz; 0.0338 fl.oz./mL
- 8. a. 2.27×10^3 g
- b. 6.224 m
- c. 482 cm^3
- d. 0.767 m
- 9. a. 338 g
- b. 0.00023 L
- c. 5.4×10^2 cm
- d. 3.380×10^3 mg
- 10. a. 0.495 m 0.000495 km
- b. 322 cm 0.00322 km
- c. 4.8 cm 0.048 m
- d. 7.9 m 0.0079 km
- e. 3.41×10^4 cm 0.341 km
- f. 2303 cm 23.03 m
- g. 414.5 m 0.4145 km
- h. 7.0×10^4 cm 0.70 km
- i. 1.8×10^3 cm 18 m
- 11. a. 0.065 g 0.000065 kg
- b. 2.7×10^5 mg 0.00027 kg

- c. $2.5 \times 10^5 \mu\text{g}$ 0.25 g
d. 0.00004521 g 0.00000004521 kg
e. $9.470 \times 10^7 \mu\text{g}$ 0.09470 kg
f. $3.700 \times 10^7 \mu\text{g}$ 37.00 g
g. 8.00 g 0.00800 kg
h. $4.20 \times 10^3 \mu\text{g}$ 0.00000420 kg
i. $6.70 \times 10^4 \mu\text{g}$ 0.0670 g
12. a. 0.065 L $6.5 \times 10^4 \mu\text{L}$
b. 38 mL $3.8 \times 10^4 \mu\text{L}$
c. $9.0 \times 10^1 \text{ mL}$ 0.090 L
d. 0.340 L $3.40 \times 10^5 \mu\text{L}$
e. $5.10 \times 10^2 \text{ mL}$ $5.10 \times 10^5 \mu\text{L}$
f. $1.20 \times 10^3 \text{ mL}$ 1.20 L
g. 8.604 L $8.604 \times 10^6 \mu\text{L}$
h. 200.0 mL $2.000 \times 10^6 \mu\text{L}$
i. 1.904 mL 0.001904 L
13. a. 1.74 g/mL
b. 14.0 g/mL
c. 6.00 g/mL
14. a. 2.51 g/mL
b. 5.7 g/mL
c. 28 g/mL
15. a. 7.1 g/mL
b. 9.44 g/mL
c. 5.37 g/mL
16. a. 0.817 g/mL
b. 1.580 g/mL
c. 2.6 g/mL
17. 452 g
18. 1.74 mL
19. $2.0 \times 10^1 \text{ mL}$
20. $1.277 \times 10^4 \text{ g}$
21. a. 0.0684 mL
b. 18.9 g
c. 6.6 g/mL
d. 0.538 mL
e. 85.0 g

- f. 2.0 g/mL
 - g. 6.000 mL
 - h. 954 g
 - i. 8.528 g/mL
 - j. 1.1×10^3 mL
22. a. 296.33 K
- b. 128.2°C
23. a. 69°C
- b. 365 K
24. a. 331.K
- b. 61.53°C
- c. 319.8 K
- d. 130.2°C
- e. 396.60 K
- f. 111°C

Chapter 3

1. Metals: calcium, Rb, rhodium, Ti, and U.
Nonmetals: oxygen, Br, and selenium
Metalloids: Si and Arsenic
2. Representative elements: Al, strontium, sulfur, antimony, and Li
Transition elements: Co, Mo, and Au
Inner-transition elements: curium and californium
3. Alkali metals: K and Cs
Alkaline earth metals: calcium and barium
Halogens: fluorine and iodine
Noble gases: argon
None of these: B, Sn, and V
4. A monatomic ion is an ion that consists of just one atom. Examples: Cl^- and Ca^{2+}
A polyatomic ion is an ion that consists of more than one atom.
Examples: CO_3^{2-} and NH_4^+
A cation is a positively-charged ion. Examples: K^+ and NH_4^+
A anion is a negatively-charged ion. Examples: O^{2-} and NO_3^-
5. Monatomic ions: I^- , fluoride, oxide, Mg^{2+} , Na^+ , iodide, and Al^{3+}
Polyatomic ions: phosphate, NH_4^+ and carbonate

Cations: NH_4^+ , Mg^{2+} , Na^+ , Al^{3+}

Anions: phosphate, I^- , fluoride, oxide, carbonate, iodide

6.
 - a. -1
 - b. +2
 - c. -2
 - d. +1
 - e. +2
 - f. -1
 - g. +1
 - h. +3
 - i. -2
 - j. +1
7.
 - a. CaO
 - b. CaBr_2
 - c. K_3PO_4
 - d. $\text{Mg}(\text{NO}_3)_2$
 - e. Al_2S_3
 - f. $(\text{NH}_4)_2\text{CO}_3$
 - g. K_2O
 - h. $(\text{NH}_4)_2\text{O}$
 - i. Rb_2SO_4
 - j. NaBr
 - k. AlCl_3
 - l. $\text{Ca}(\text{NO}_3)_2$
 - m. $\text{Mg}_3(\text{PO}_4)_2$
 - n. BaS
 - o. CaF_2
 - p. Al_2O_3
 - q. K_2S
 - r. MgO
 - s. Rb_2CO_3
 - t. Li_3PO_4
 - u. $\text{Sr}(\text{OH})_2$
 - v. NH_4Cl
8. The ionic compounds are those that contain a metal ion or the ammonium ion. These are SrCO_3 , MgBr_2 , NH_4NO_3 , $\text{Al}_2(\text{SO}_4)_3$, FeBr_2 , K_2O , and MgCl_2 . The remaining ones are covalent.

9.
 - a. salt
 - b. acid
 - c. salt
 - d. base
 - e. salt
 - f. base
10.
 - a. hydrobromic acid, hydrogen bromide
 - b. calcium oxide
 - c. diphosphorus pentoxide
 - d. potassium iodide
 - e. ferrous chloride, iron(II) chloride
 - f. sodium sulfide
 - g. sulfur trioxide
 - h. magnesium hydroxide
 - i. carbon tetrachloride
 - j. hydrochloric acid, hydrogen chloride
 - k. calcium bromide
 - l. ferrous oxide, iron(II) oxide
 - m. calcium sulfate
 - n. ferrous bromide, iron(II) bromide
 - o. sulfur hexafluoride
 - p. hydroiodic acid, hydrogen iodide
 - q. barium chloride
 - r. hydrosulfuric acid, hydrogen sulfide
 - s. magnesium oxide
 - t. cupric chloride, copper(II) chloride
 - u. phosphorus trichloride
11.
 - a. KCNS
 - b. $(\text{NH}_4)_2\text{HPO}_4$
 - c. $\text{Na}_2\text{Cr}_2\text{O}_7$
 - d. $\text{Ca}(\text{ClO})_2$
 - e. $\text{Mg}(\text{ClO}_4)_2$
 - f. PbCrO_4
 - g. KCN
12.
 - a. CCl_4
 - b. CuBr
 - c. NH_4Cl

- d. NaF
- e. Fe_2O_3
- f. N_2O_3
- g. SnSO_4
- h. AgCl
- i. SF_6
- j. HF
- k. SnCl_4
- l. K_2O
- m. NaOH
- n. TiO_2
- o. HCl
- p. P_2O_5
- q. MgCl_2
- r. NiS
- s. N_2O_5
- t. CuF_2
- u. AlBr_3
- 13. a. MgBr_2
- b. H_3PO_4
- c. FeI_3
- d. K_2SO_4
- e. Cu_2CO_3
- f. SO_3
- g. H_2S
- h. N_2O
- 14. a. nitric acid
- b. carbon tetrachloride
- c. Aluminum bromide
- d. Tin (IV) oxide
- e. Ammonium carbonate

Chapter 4

- 1. a. Dalton—first comprehensive atomic theory
- b. Crookes—electron

- c. Thomson—plum pudding model
 - d. Millikan—value of electron charge
 - e. Rutherford—the nucleus exists
 - f. Chadwick—neutron
 - g. Bohr—solar system model
 - h. Schrodinger—modern atomic theory
2. Solar system model—c
 Nuclear atom—e
 Original atomic theory—b
 Quantum mechanical model—a
 Plum pudding model—d
 Neutron—g
 Electron—f
 Electron charge—h
3. A small percentage of the alpha particles were deflected from their path, and a very small percentage reflected back from the foil. A very high percentage of the alpha particles came straight through the foil. This supported the idea that most of an atom's mass and positive charge were concentrated in a small core, the nucleus.
4. The boulder has the capacity to cause matter to move because it has potential energy. Potential energy can be converted to other forms of energy, which then have this capacity.
5. The different forms of energy mentioned in the text are kinetic, potential, electrical, light, chemical, heat, and nuclear. Examples of transformation are (1) kinetic is transformed into electrical in a dam on a river, (2) potential is converted to kinetic when a boulder falls from a cliff (3) electrical is converted to light when a flashlight is switched on, (4) light is converted to electrical in a solar battery, (5) chemical is converted into electrical in a car battery, (6) heat is converted to electrical in a steam turbine, and (7) nuclear is converted to electrical in a nuclear power plant.
6. Light has properties of both particles and electromagnetic waves.
7. Light waves are electromagnetic, not mechanical, and as such, they do not require matter in order to exist.
8. Wavelength is the distance from the crest of one wave to the crest of the next wave in electromagnetic radiation. Long wavelength is low energy.

9. Gamma rays, ultraviolet waves, and X-rays. Visible and infrared light can cause eye damage if intense enough. Microwaves can indirectly cause damage through the generation of heat.
10. Red, orange, yellow, green, blue, violet.
11. From Figure 3.3: 450—blue to violet; 600—yellow to orange; 550—green to yellow; 700—red; 350—ultraviolet.
12. Different colors of light represent different energies of light. According to the Bohr model, the different electron orbits also represent different energies. Bohr theorized that electrons can gain or lose energy and move from one orbit to another. The energies that electrons can lose when jumping from a higher orbit to a lower orbit represent the different colors of light.
13. That the electron is a particle in orbit around the nucleus was discounted with the modern theory.
14. It is a different model, the obsolete “Bohr model,” not the quantum mechanical model, that describes the electrons that way.
15. a. $2n^2 =$ maximum number of electrons in principal level n . In this case, $2n^2 = 18$.
b. $n^2 =$ the number of orbitals in principal level n . In this case, $n^2 = 9$
c. There can only be 6 electrons maximum in any p sublevel.
d. There can only be 2 electrons maximum in any orbital, whether it is s , p , d , or f .
16. $n^2 =$ the number of orbitals in principal level n . If $n = 4$, 16 orbitals. If $n = 5$, 25 orbitals.
17. a. The f sublevel has the most energy.
b. The d sublevel can hold 10 electrons maximum.
c. The p sublevel has only 3 orbitals.
18. a. Any one orbital can hold only 2 electrons maximum, regardless of type.
b. The $3p$ sublevel holds 6 electrons maximum.
c. $2n^2 = 32$.
d. Any s sublevel has only 1 orbital, so just 2 electrons maximum.
19. a. orbital
b. principal level
c. sublevel
20. a. $2n^2 = 18$
b. A single orbital of any type holds a maximum of 2 electrons.

- c. 6
 - d. *s* sublevels hold 2 electrons maximum.
21. It can have three orientations, each perpendicular to the others along the *x*, *y*, and *z* axes (See Figure 4.10). They are all equal in energy. The nucleus is at the center of the figure eight.
22. The *s* orbital is spherical in shape. The lowest principal energy level in which this orbital is found is $n = 1$. It can have only one orientation.
23. Each arrow represents an electron. The total number of arrows then matches the atomic number of the element.
24. a. 12
b. 13
c. 14
d. 15
e. 17
f. 18
25. a. 1
b. 3
c. 6
26. a. 2
b. 6
c. 9
27. a. P, or phosphorus
b. $1s^2 2s^2 2p^1$
c. V, or vanadium
d. $1s^2 2s^2 2p^6 3s^2 3p^5$
e. O, or oxygen
f. $1s^2 2s^2 2p^6 3s^2 3p^6 4s^2 3d^8$
g. Ar, or argon
h. $1s^2 2s^2 2p^6 3s^2 3p^6 4s^1$
i. Sn, or tin
j. $1s^2 2s^2 2p^6 3s^2$
k. Fe, or iron
l. $1s^2 2s^2 2p^6$
m. S, or sulfur
n. $1s^2 2s^2$
o. Cs, or cesium

28. This is because the $n = 3$ and the $n = 4$ levels overlap; the lowest sublevel in the $n = 4$ level, the $4s$, actually lies below the highest sublevel in the $n = 3$ level, the $3d$.
29. A half-filled sublevel means that the sublevel has half the number of electrons that it can have and that each orbital within that sublevel has one electron. Sublevels that are half-filled are lower in energy than those that are not. To achieve the favorable half-filled condition, some elements, notably chromium, borrow electrons from other sublevels.

Chapter 5

1. Mendeleev's periodic table was based on properties. However, it is remarkably similar to the modern periodic table, which is based on electron configuration. Mendeleev's work thus shows that properties have much to do with electron configuration.
2. Properties have much to do with electron configuration.
3. See answer to Exercise 1.
4. Descriptive chemistry is the study of specific properties and other important facts about individual elements and compounds.
5.
 - a. Alkali metals are highly reactive when exposed to air, water, and most other materials. Mineral oil is an exception. Because of this, these metals are often stored under mineral oil to preserve them in the free elemental state.
 - b. Yes, sodium and potassium are examples of metals that you can cut with a knife.
 - c. Potassium, for example: $2 \text{ K} + 2 \text{ H}_2\text{O} \rightarrow 2 \text{ KOH} + \text{H}_2$
6.
 - a. See Section 5.3.2
 - b. Example: $\text{Mg} + 2 \text{ H}_2\text{O} \rightarrow \text{Mg(OH)}_2 + \text{H}_2$
7. See Section 5.3.3.
8. Hydrogen is explosively flammable and is dangerous to use in blimps. Helium, being a noble gas, is not reactive.
9. Allotropes are different forms of the same element. Examples are diamond and graphite, two allotropes of carbon. See Section 5.3.5 for more discussion.
10. See Sections 5.3.6 and 5.3.7.

11. Such compounds include, for example, NaCl, $\text{Mg}(\text{OH})_2$, BaSO_4 , NaF, KI, H_3PO_4 , H_2SO_4 , SiO_2 , ZnO. See Section 5.3 for more information.
12. It repeats or recurs periodically as atomic number is increased.
13. Properties repeat or recur periodically as atomic number is increased.
14. The periodic law states that the properties of the elements are periodic functions of their atomic numbers, meaning that properties recur periodically as atomic numbers increase. This is important because it recognizes the periodicity elemental properties and can be used to predict and explain why different elements behave the way that they do.
15. From left to right across a period in the periodic table, atomic size decreases. From top to bottom with a family, atomic size increases.
16. More electrons in higher energy levels require more physical space, therefore atoms toward the bottom of a given family in the periodic are larger than atoms of elements above them. Atoms of elements on the left end of a given period hold their electrons less tightly (because they tend to be given away such that cations form), and that means they are farther from the nucleus than are the electrons of those on the right end. Farther from the nucleus means a larger diameter, a larger atom.
17. Ionization energy increases as you go from left to right within a period in the periodic table. From top to bottom within a family in the periodic table, ionization energy decreases.
18. As you go from left to right within a period, the electrons are held more tightly by the atom because you are approaching a filled p sublevel and it is a more stable situation if atoms were to acquire more electrons rather than give them up. Hence more energy is required to remove an electron. Toward the bottom within a family, electrons in the outermost energy levels are held more loosely because they are farther from the nucleus. Thus it will take less energy to remove them.
19. As you go from left to right across a period, the affinity an atom has for electrons increases. As you go from top to bottom within a family, this affinity decreases.
20. See the answer to Exercise 18. A greater ionization energy also means a higher affinity for electrons. The explanations for this are the same as for ionization energy.

21. The screening effect is the effect the electrons in the lower energy levels (nearer the interior of the atom) have on those in the outermost levels. They, in effect, screen the outermost electrons from the pull of the nucleus because they are like-charged. This results in lower ionization energy, lower electron affinity, and larger atoms near the bottom within a family.
22. Fluorine is at the top of the halogen family while bromine is nearer the bottom. Therefore fluorine will have a higher electron affinity than will bromine because fluorine has fewer electrons available in the interior levels to screen other electrons from the pull of the nucleus.
23. a. Ca
b. S
c. Na
d. N
e. N
f. Al
24. a. S, Cr, K, Rb
b. F, O, Ca, Cs
c. F, Ge, Fe, Ba
d. Fr, Hg, Te, S
e. Ba, Tc, Al, O
f. Br, Mo, Sr, Cs
g. Ba, Ir, Au, S
h. Cannot tell exactly where cobalt fits because it is both higher up and also to the left of the others. In terms of the other three, Pt, Pd, Ag
i. O, Zn, Ag, Rb
j. N, Mo, Ba
k. Sn, As, S, F
25. An element with a low ionization energy would tend to form a cation because a low ionization energy means that it does not take much energy to remove an electron.
26. Lower left
27. Upper right
28. K^+ , because the one electron that K has in its outermost energy level is loosely held, so the atom is large. However, when that electron is lost, not only is the outermost energy level the next level down, but

there are also fewer electrons than protons and all are more tightly held, making for a smaller sphere.

29. Ionization energy is the energy required to remove an electron from an atom.

Chapter 6

1. a. An ionic bond is the electrostatic attraction between oppositely charged ions.
- b. A covalent bond is a sharing of electrons between atoms.
2. KBr is an ionic compound because of the ionic bonding that it present. The important clue is the presence of a metal in the formula,

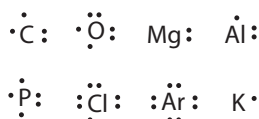


FIGURE A6.1

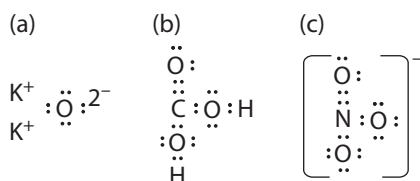


FIGURE A6.2

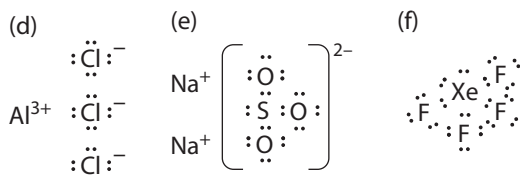


FIGURE A6.3

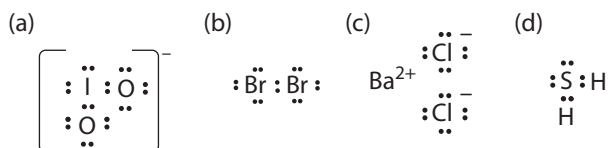


FIGURE A6.4

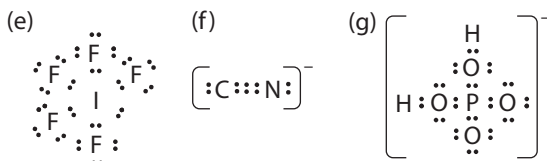


FIGURE A6.5

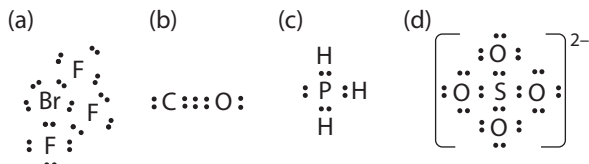


FIGURE A6.6

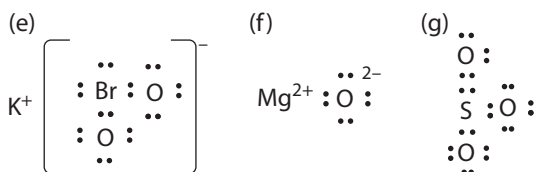


FIGURE A6.7

because metals are only rarely found in covalent bonding situations. HBr is a covalent compound since there is no metal in the formula and a pair of electrons is being shared by the H and the Br.

- When two atoms come together to form a bond, it is the outer edges of each atom that interact. Since all the inner levels are filled with electrons already, it stands to reason that the outermost energy level of an atom—which in most cases is not filled when the atom is in a free, uncombined state (or elemental state)—is the level that is involved in the bonding.
- See Figure A6.1.
- See Figures A6.2 and A6.3.
- See Figures A6.4 and A6.5.
- See Figures A6.6 and A6.7.
- The correct Lewis electron dot structure is (b). MgCl_2 is an ionic compound (metal combined with nonmetal)—one magnesium ion combined with two chloride ions as structure (b) indicates. Structure (a) would indicate that covalent bonds are present, which is inaccurate.

9. a. The electron count is wrong. Three oxygen atoms contribute 18 electrons; one sulfur contributes six electrons; and two hydrogens contribute two electrons—a total of 26 electrons. The structure shows only 24. There should be no double bonds, and there should be a nonbonded pair of electrons on the sulfur atom.
b. The nitrogen atom has only six electrons; it should have eight. There should be a double bond between the nitrogen atom and one of the oxygen atoms. One of the nonbonding pairs on the oxygen atom should be used for the second bond.
10. A hydrogen atom can only form one bond, since sharing one pair of electrons gives it the electrons needed to fill its $n = 1$ principal energy level. Thus it will always be found around the outside of a Lewis electron-dot structure.
11. Molecules are distinct units that exist independent of their surroundings. In ionic compounds, positive and negative ions exist in a three-dimensional array—positive ions surrounded by negative ions and negative ions surrounded by positive ions—such that there is no distinct unit we can identify as a molecule.
12. Electron affinity is the attraction a free atom has for electrons. Electronegativity is the attraction a bonded atom has for its shared electrons. A number is assigned to the electronegativity of an element in order to compare one element to another and determine the polarity of a bond.
13. When the electronegativity values of two nonmetal atoms are different, the bond between the two is polar, and the larger the difference, the more polar the bond.
14. A polar bond is one in which the shared electrons are pulled more toward one atom in the bond than the other, resulting in the imbalance of charge that is a polar bond. An example is the O–H bond in a water molecule.
15. A polar molecule is a molecule in which the polar bonds are not symmetrically arranged in space. An example is the water molecule because the polar O–H bonds are not symmetrically arranged in space.
16. The polar bonds in compound B are symmetrically arranged in space while the polar bonds in compound A are not.
17. a. The O–H bond is a polar covalent bond. It is covalent and not ionic because it is between two nonmetals. It is polar because the

- electronegativities of oxygen and hydrogen are different (refer to Table 6.2).
- b. The Br–Br bond is a nonpolar covalent bond. It is covalent and not ionic because it is between two nonmetals. It is nonpolar because the two atoms that are bonded are identical and thus have equal electronegativities.
 - c. The Ca–O bond is an ionic bond because a metal (calcium) is involved and metals tend to give up electrons and not share them.
 - d. The H–P bond is a nonpolar covalent bond. It is covalent and not ionic because it is between two nonmetals. It is nonpolar because the electronegativities of hydrogen and phosphorus are the same (refer to Table 6.2).
- 18. The carbon–chlorine bonds, while polar, are symmetrically arranged in space in CCl_4 (tetrahedral), so the molecule is nonpolar.
 - 19. The polar bonds must be symmetrically arranged in space.
 - 20. An intermolecular force is the force of attraction one molecule has for another. The hydrogen bond is an example.
 - 21. A hydrogen bond is an intermolecular force consisting of the attraction that a hydrogen atom bonded to an oxygen, nitrogen, or fluorine atom on one molecule has for a pair of electrons—usually on an oxygen, nitrogen, or fluorine atom—on another. See Section 6.9.3.
 - 22. Possible answers are a high boiling point, significant expansion upon freezing, ice being less dense than liquid water, and the formation of ice causing water pipes and hoses to burst.
 - 23. The boiling point of carbon tetrachloride, CCl_4 , is 78°C . The boiling point of chloroform, CHCl_3 , is 61°C . Replacing a chlorine on the nonpolar CCl_4 molecule with hydrogen to form CHCl_3 creates a polar molecule that is lighter in weight. A polar molecule would mean a higher boiling point, but a lighter weight means a lower boiling point. It is apparent that the lighter weight has a greater effect in this case.
 - 24. A hydrate is an ionic compound that has water molecules trapped within the crystal structure, a certain number of water molecules per formula unit. See Section 6.9.4 for some examples
 - 25. Since the water molecule retains its identity when combining with the ionic compound to form a hydrate, it is appropriate to show the

water molecule intact in the formula of the hydrate. The use of the dot to separate the formula of the ionic compound from the formula of water is special symbolism to display this unique situation.

26. The acronym VSEPR stands for the valence shell electron pair repulsion theory. It is a theory through which electron pair repulsions are used to predict the geometric shape of a molecule.
27. The theory is that electron pairs—or, more correctly, regions of high electron density—in the valence shell of the central atom repel each other and thereby get as far away from each other as they can get. This results in a particular geometric shape depending on how many total regions there are, how many bonding regions there are, and how many nonbonding pairs there are.
28. See Figure 6.23a and b. In terms of the total number of regions of high electron density, they are both the same,—tetrahedral (i.e., the general structure of each is tetrahedral). In terms of the specific molecular structure, for which only bonding regions are important, they are different; methane is tetrahedral and water is angular.
29.
 - a. angular
 - b. square planar
 - c. linear
 - d. octahedral
 - e. T-shaped
 - f. trigonal planar
30. Answers vary depending on which are selected.
31. A molecular orbital is the region of space occupied by the bonding electrons in a covalent bond. It is formed by the overlap of the atomic orbitals that are participating in the bond.
32. A σ bond is the end-to-end overlap between any two orbitals. A π bond is the side-to-side overlap between two p orbitals. Any compound or ion that has a double bond or triple bond has one or two π bonds, respectively. Examples of just one π bond include oxygen gas, the carbonate ion, the nitrate ion, and ethylene. Examples of two π bonds include nitrogen gas, carbon monoxide, the cyanide ion, and acetylene.
33.
 - a. See Figure 6.27
 - b. See Figure 6.27

- c. See Figure 6.27
 - d. See Figure 6.28
34. Situations (a), (b), and (c) produce a σ bond. Situation (d) produces a π bond.
35. a. All
- b. b
 - c. c
 - d. None
 - e. None
 - f. a
 - g. None
36. A π bond is the side-to-side overlap between two p orbitals. See Figure 6.28
37. A triple bond has one σ bond and two π bonds.
38. a. Any kind of orbital can be involved in a σ bond. End to end.
- b. Only p orbitals are involved in a π bond. Side to side.
39. One σ and one π . The first bond is always a σ bond. When two p orbitals that have just one electron each line up side-to-side after that, then you will also have a π bond.

Chapter 7

1. Stoichiometry is the expression, measurement, and calculation of quantities of matter in compounds and in reactions.
2. The weight of the carbon-12 isotope includes the weight of the electrons and other subatomic particles, not just the weight of the protons and neutrons.
3. a. Molecular weight is the average weight of a molecule of a covalent compound in amu. It is also the average weight of 1 mole of a covalent compound in grams.
- b. Formula weight is the average weight of a molecule or formula unit of a compound in amu per molecule or formula unit, or grams per mole. For covalent compounds, it is the same as the molecular weight.
- c. A mole is Avogadro's number of anything, just like the dozen is 12 of anything.

4. Formula weight is more general, encompassing both covalent and ionic compounds. Molecular weight is specific to covalent compounds. Formula weight acknowledges that not all compounds are composed of molecules.
5. a. 74.551 g/mol
b. 342.150 g/mol
c. 116.315 g/mol
6. a. 44.009 g/mol
b. 17.031 g/mol
c. 174.258 g/mol
d. 163.940 g/mol
e. 74.123 g/mol
f. 41.053 g/mol
g. 58.319 g/mol
h. 233.988 g/mol
i. 278.014 g/mol
j. 237.929 g/mol
7. a. 2.6×10^3 g
b. 7.9×10^2 g
c. 19.37 g
d. 8.33×10^3 g
8. a. 0.0109 mol
b. 0.0003656 mol
c. 0.01253 mol
d. 0.0440 mol
9. a. 0.0119 mol
b. 54.67 g
c. 0.00065 mol
d. 5153 g
e. 0.01962 mol
f. 1.26×10^3 g
g. 722 mol
h. 5.154×10^5 g
10. a. 0.108 mol of CO_2 has more molecules
b. 4.5 mol of NH_3 has more molecules
11. 18.401 %S

12. a. 13.854 %N
b. 17.073 %N
13. a. 82.244 %N
b. 46.681 %N
c. 63.649 %N
d. 30.447 %N
e. 22.590 %N
f. 36.855 %N
g. 30.447 %N
h. 25.937 %N
14. a. 7.8083 %C, 92.192 %Cl
b. 82.24 %N, 17.76 %H
c. 48.641 %C, 8.164 %H, 43.194 %C
d. 6.714 %H, 40.002 %C, 53.284 %O
15. 51.164 %water
16. A count percent is not the same as a weight percent unless all items weigh the same. A hydrogen atom does not weigh the same as an oxygen atom.
17. a. AlCl_3
b. Sr_2SiO_4
c. $\text{C}_5\text{H}_{10}\text{O}_2$
d. Fe_3O_4
e. Al_4C_3
f. $\text{C}_3\text{H}_4\text{O}_3$
g. $\text{Na}_2\text{S}_2\text{O}_3$
h. $\text{Mg}_2\text{P}_2\text{O}_7$
18. a. $\text{C}_4\text{H}_{12}\text{O}_2$
b. $\text{C}_6\text{H}_{12}\text{O}_3$
c. $\text{C}_8\text{H}_8\text{S}_2$
d. $\text{C}_6\text{H}_4\text{Cl}_2$
19. Empirical— $\text{C}_2\text{H}_4\text{O}$; molecular— $\text{C}_4\text{H}_8\text{O}_2$

Chapter 8

1. a. Reactants— BaCl_2 and K_2SO_4 ; products: BaSO_4 and KCl
b. Reactants— KOH and $\text{H}_2\text{C}_2\text{O}_4$; products: $\text{K}_2\text{C}_2\text{O}_4$ and H_2O
c. Reactants— Fe and O_2 ; product: Fe_2O_3

2.
 - a. One mole of barium chloride reacts with one mole of potassium sulfate to produce one mole of barium sulfate and two moles of potassium chloride.
 - b. Two moles of potassium hydroxide reacts with one mole of oxalic acid to produce one mole of potassium oxalate and two moles of water.
 - c. Four moles of iron react with three moles of oxygen to produce two moles of iron(III) oxide.
3. The law of conservation of mass is the law that states that matter can neither be created nor destroyed in a chemical reaction. The total number of atoms of each element present in the reaction container must be the same before the reaction begins and after the reaction takes place.
4.
 - a. $2 \text{ Fe} + \text{O}_2 \rightarrow 2 \text{ FeO}$
 - b. $4 \text{ P} + 5 \text{ O}_2 \rightarrow 2 \text{ P}_2\text{O}_5$
 - c. $2 \text{ SO}_2 + \text{O}_2 \rightarrow 2 \text{ SO}_3$
 - d. $\text{H}_3\text{PO}_4 + 3 \text{ NaOH} \rightarrow \text{Na}_3\text{PO}_4 + 3 \text{ H}_2\text{O}$
 - e. $\text{BiCl}_3 + \text{H}_2\text{O} \rightarrow \text{BiOCl} + 2 \text{ HCl}$
 - f. $2 \text{ S} + 3 \text{ O}_2 \rightarrow 2 \text{ SO}_3$
 - g. $4 \text{ FeO} + \text{O}_2 \rightarrow 2 \text{ Fe}_2\text{O}_3$
 - h. $4 \text{ Ag} + 2 \text{ H}_2\text{S} + \text{O}_2 \rightarrow 2 \text{ Ag}_2\text{S} + 2 \text{ H}_2\text{O}$
 - i. $3 \text{ NO}_2 + \text{H}_2\text{O} \rightarrow 2 \text{ HNO}_3 + \text{NO}$
 - j. $2 \text{ HgO} \rightarrow 2 \text{ Hg} + \text{O}_2$
 - k. $\text{Mg} + 2 \text{ H}_2\text{O} \rightarrow \text{H}_2 + \text{Mg(OH)}_2$
 - l. $2 \text{ N}_2 + \text{O}_2 \rightarrow 2 \text{ N}_2\text{O}$
 - m. $\text{PCl}_5 + \text{H}_2\text{O} \rightarrow \text{POCl}_3 + 2 \text{ HCl}$
 - n. $2 \text{ HCl} + \text{Ba(OH)}_2 \rightarrow \text{BaCl}_2 + 2 \text{ H}_2\text{O}$
5. Reactions (d) and (n) are acid–base reactions, because one reactant is an acid (formula begins with H) and one is a base (hydroxide).
6.
 - a. Solid
 - b. Gas
 - c. Liquid
 - d. A gas that is a product of a reaction and escapes the container
 - e. A precipitate
 - f. Heat added
7.
 - a. CaCO_3 and HCl are the reactants
 - b. H_2O is a liquid
 - c. CO_2 is a gas

8. $\text{Mg(s)} + \text{HBr(aq)} \rightarrow \text{MgBr}_2\text{(aq)} + \text{H}_2\text{(l)}$
9. (s) means that Zn is a solid, (aq) means that HCl and ZnCl_2 are in water solution, and (↑) means that H_2 is a gas that forms and escapes into the air.
10. $\text{Pb(NO}_3)_2\text{(aq)} + \text{K}_2\text{CrO}_4\text{(aq)} \rightarrow \text{PbCrO}_4\text{(↓)} + \text{KNO}_3\text{(aq)}$
11. One atom of sulfur reacts with three molecules of fluorine to yield one molecule of sulfur hexafluoride. One mole of sulfur reacts with three moles of fluorine to yield one mole of sulfur hexafluoride.
12. a. 0.00229 mol
b. 4.7 g
c. 0.00389 mol
d. 0.8383 g
13. a. 0.3727 mol
b. 10.55 mol
c. 15.0 g
d. 6.925 g
14. a. 3.579 mol
b. 163.3 g
15. a. 0.110 mol
b. 8.59 g
16. I_2 because it produces less product
17. Limiting reactant— Fe_3O_4 ; grams of iron—0.238 g
18. Limiting reactant— Cu_2O ; grams of CuO—0.0809 g CuO
19. a. HCl 1.7 mol AlCl_3
b. HCl 0.33 mol AlCl_3
c. $\text{Al}_2(\text{CO}_3)_3$ 0.00364 mol AlCl_3
d. HCl 0.0947 mol AlCl_3
20. a. Fe 0.282 g FeO
b. Fe 1.072 g FeO
c. O_2 2.7 g FeO
d. O_2 5.36 g FeO
21. 91.3%
22. 25.82%
23. a. 0.00191 mol 97.4%
b. 7.3×10^2 g 97%
c. 0.01499 mol 87%
d. 63.7 g 92.6%

24. a. 0.183 mol 97.8%
b. 267.2 g 92.8%
c. 2.31 mol 88.2%
d. 44.7 g 92.1%

Chapter 9

1. See Section 9.2
2. See Section 9.3
3. See Section 9.3
4. When the tires get hot, heat energy is transformed into kinetic energy and the gas molecules move faster. When they move faster, they collide with the walls of the container more often. More collisions per unit time means more pressure.
5. Ether molecules move through air molecules because of the fact that gas molecules are far apart from one another and can accommodate other moving molecules with no opposing force.
6. a. More air can be added because the air molecules already inside the tire are far apart from one another and can accommodate more molecules in the same space.
b. More gas molecules in the confined volume of the tire means more collisions with the inner walls of the tire, which means more pressure.
7. a. 1.02 atm 14.9 psi
b. 1086 torr 21.00 psi
c. 1.7×10^3 torr 2.2 atm
d. 0.924 atm 13.6 psi
e. 707 torr 13.7 psi
f. 6.00×10^2 torr 0.789 atm
g. 1.117 atm 16.42 psi
h. 541.0 torr 10.46 psi
i. 930.9 torr 1.225 atm
8. a. 14.2 psi
b. 0.963 atm
9. 9.80×10^2 torr, 19.0 psi
10. 861.8 torr, 16.67 psi
11. 0.978 atm, 14.4 psi

12. Equations 9.1 and 9.4 are expressions of Boyle's law. Temperature is held constant while pressure is varied.
13. a. 6.29 mL
b. 18.0 mL
c. 709 torr
d. 2.55 atm
e. 15.2 L
f. 1.26×10^4 mL
g. 2.17 torr
h. 9.19×10^3 torr
14. 82.6 mL
15. 33.2 mL
16. 717 torr
17. 1.0×10^1 L
18. 0.07 atm
19. Equations 9.5 and 9.8 are expressions of Charles' law. Pressure is held constant while temperature is varied.
20. a. 3.1 L
b. 74.2 mL
c. 259 K
d. 557 K
e. 657.6 mL
f. 6.678 L
g. 3.21×10^3 K
h. 147 K
21. 249.8 mL
22. 7.26 L
23. 48°C
24. 275 K or 2°C
25. 10°C
26. Equation 9.9 is an expression of the combined gas law. Both pressure and temperature are varied.
27. a. 9.26 L
b. 189 K
c. 722 torr
d. 0.221 mL
e. 41°C

- f. 1.09×10^3 torr
- g. 996 K
- h. 6.3 L
- i. 743 torr
- 28. 3.01 L
- 29. 0.580 L
- 30. 724 torr
- 31. 261 K
- 32. 5.40×10^2 torr
- 33. Standard temperature and pressure are specific temperature and pressure values that scientists use as the standard for the purpose of comparing gas volumes measured at conditions that can vary widely. Standard pressure is 760 torr (1 atm) and standard temperature is 0°C, or 273.15 K.
- 34. 61.9 mL
- 35. 0.184 L
- 36. 85.0 mL
- 37. 219 K
- 38. 0.787 atm
- 39. See Equation 9.15
- 40. $0.082057 \text{ L atm/mol K}$ or $\text{L atm mol}^{-1} \text{ K}^{-1}$
- 41. a. 6.61 K
- b. 0.00989 mol
- c. 30.4 L
- d. 1.08×10^3 atm
- e. 551 K
- f. 0.01602 mol
- g. 2.020×10^3 L
- h. 346 atm
- 42. 2.01 mol
- 43. 11.3 L
- 44. 2.87 mol
- 45. 2.6×10^4 torr
- 46. -101°C
- 47. Two statements of Avogadro's law are given in Section 9.11.
- 48. a. 3.71 mol
- b. 1.06 L

- c. 24.5 mol
- d. 1.75×10^7 L
- 49. 0.0327 mol
- 50. 19.8 L
- 51. Yes, 22.414 L
- 52. See Equation 9.18 and the accompanying discussion.
- 53. Vapor pressure is the pressure exerted by the molecules or atoms of a gas in equilibrium with its liquid state in the same sealed container.
- 54. The normal boiling point is the temperature at which the vapor pressure equals the standard atmospheric pressure, 760 torr.
- 55. The vapor pressure of a liquid usually increases with temperature.
- 56. 743 mL
- 57. 2.02 L
- 58. 8.28 L

Chapter 10

- 1. a. A solute is a solution component that is present in a lesser amount.
- b. A solvent is the solution component that is present in the greatest amount.
- c. A solution is a homogeneous mixture consisting of a solvent and one or more solutes.
- 2. a. T. This is how we defined the term in Section 10.2.
- b. F. This defines solvent, not solute.
- c. F. More than one solute, but only one solvent.
- d. F. The term solution actually applies to any homogeneous mixture.
- e. T. Any of these three can dissolve and be a solute.
- 3. 9.49%
- 4. a. 3.43%
- b. 34.8%
- c. 14%
- d. 42.5%
- e. 27.9%
- 5. Not necessarily. Liquid volumes are not necessarily additive when mixed, meaning that the two solutions described may not be the same.

6. The value of a concentration does not depend on what the chemical is. It depends only on weight and volume.
- 1.168%
 - 1.168%
7. No. The one with 200.0 mL of water likely contains more water than the one with 200.0 mL of solution. In other words, the solute probably takes up some volume, so it takes less than 200.0 mL of water to give 200.0 mL of solution.
8. a. 1.36 M
b. 0.0127 M
c. 0.562 M
d. 0.02512 M
9. 0.141 M
10. a. 0.0469 M 0.187%
b. 0.00083 M 0.0048%
c. 0.00280 M 0.0398%
d. 0.030 M 0.33%
e. 0.00647 M 0.137%
f. 0.027 M 0.462%
g. 0.108 M 1.30%
h. 0.0260 M 0.370%
i. 0.006040 M 0.08220%
j. 0.51 M 4.2%
11. a. 2.52%
b. 0.522 M
12. a. 7.3 mL
b. 75 mL
13. a. 9.7 mL
b. 6.9 mL
14. a. Dilute 6.2 mL of the 16 M HNO_3 to 500.0 mL of solution with water. Shake.
b. Dilute 35 mL of the 17 M acetic acid to 1.20 L of solution with water. Shake.
15. Dilute 0.25 L of the 12 M HCl to 2.00 L of solution with water. Shake.
16. a. 8.3 mL
b. 2.1 mL

- c. 20.0 mL
 - d. 1.0×10^2 mL
 - e. 1.4×10^2 mL
 - f. 72 mL
 - g. 39 mL
 - h. 4.0×10^2 mL
17. When a high degree of accuracy is important.
18. a. 45.0 mL
b. 3.60×10^2 mL
19. 0.0075 L
20. 38 g
21. 50.0 g
22. 52 g
23. Formula weight also enters into the calculation. See Equation 10.20.
a. 17 g
b. 22 g
24. Weigh 8.4 g of the solid into the container, add water to dissolve, and then add more water to the graduation line and shake.
25. First method: Weigh 4.0×10^1 grams of the solid into the flask, add water to dissolve, and then dilute to the mark with water and shake.
Second method: Measure 68 mL of the 5.5 M solution into the flask, add water to the mark, and shake.
26. a. Weigh 24 g of the solid into a 500-mL flask, add water to dissolve, add more water to 500.0 mL, and shake.
b. Place 44 mL of the 4.0 M solution into a 500-mL flask, add water to the 500.0 mL, and shake.
27. Weigh 37 g of the solid add water to dissolve, add more water to the line, and shake.
28. a. 4.37 g
b. 32.6 g
c. 21.6 g
d. 299 g
e. 9.0 g
f. 83.98 g
g. 36.4 g
h. 14.0 g

- i. 30.8 g
 - j. 67 g
29. a. Soluble describes a solute, a small portion of which appears to dissolve freely in a quantity of solvent.
- b. Insoluble describes a solute, a small portion of which does *not* dissolve in a quantity of solvent.
- c. Solubility has both a qualitative definition and a quantitative definition. The qualitative solubility uses the words *soluble* and *insoluble* to describe a solute. The quantitative solubility uses a certain quantity (e.g., a certain number of grams per liter) to describe how much solute dissolves in a quantity of solvent or solution.
30. a. Saturated describes a solution that has the maximum amount of solute possible dissolved at the given temperature. It is characterized by some undissolved solute present.
- b. Unsaturated describes a solution that has less than the maximum amount dissolved at the given temperature.
- c. Supersaturated describes a solution that has more than the maximum amount possible dissolved at the given temperature. This condition is achieved by heating, dissolving more, and then cooling back to the given temperature.
31. a. Miscible describes a pair liquids that appear to mix freely with each other. Immiscible describes a pair of liquids that appear not to mix freely with each other.
- b. Dilute describes a solution that does not have much solute dissolved. Concentrated refers to a solution that has a large quantity of solute dissolved. Dilute can describe a solution that has less than an amount that has been designated for a concentrated solution even though it may be a relatively large amount.
32. a. T. If they are not miscible, then the mixture would be heterogeneous
- b. F. See Figure 10.6.
- c. F. Such liquids would be immiscible.
- d. T. See Section 10.6.3.
- e. F. Concentrated solutions contains more solute.
- f. T. If a solute does not completely dissolve, we say that the solution is saturated.

- g. T. More solute can usually dissolve at a higher temperature.
 - h. T. Nonpolar liquids don't mix with polar liquids.
 - i. T. Unsaturated solutions contain an amount of solute that is less than the solubility value.
33. a. saturated
b. immiscible
c. solute
d. solubility
e. supersaturated
f. soluble
g. concentrated
34. To say that a solute is insoluble is a qualitative description that is observed in exactly the same manner by which this student observed it. So she is correct. However, all solutes dissolve to some extent, even though it may not be (cannot be) observed. The small amount that dissolves can be described quantitatively and referred to as the solubility. This is the manner by which the handbook lists the solubility, so the handbook is also correct.
35. Shaking, stirring, ultrasonic agitation
36. Large particles have less surface area in contact with the solvent, which means no contact of most of the solute with the solvent and therefore a slower rate of dissolving.
37. The negative end of polar water molecules are attracted to the positive sodium ions, and the positive end of polar water molecules are attracted to the negative chloride ions. These forces are sufficient to pull the ions from the three-dimensional array of ions and dissolve the crystal.
38. I would refer him to Figure 10.2 and tell him that there is considerable interaction between the ions in the crystal structure and the polar water molecule, causing the ions to be pulled from the crystal structure and carried off into the solution, or dissolving.
39. Soluble: NaBr , K_2SO_4 , NH_4Cl , AgNO_3 , $\text{Ba}(\text{C}_2\text{H}_3\text{O}_2)_2$, Na_3PO_4 .
Insoluble: $\text{Fe}(\text{OH})_3$, Ag_2CO_3
40. Polar covalent compounds, such as acids. Examples: hydrochloric acid, sulfuric acid, sugar.
41. A strong acid completely ionizes when dissolved in water, while a weak acid only partially ionizes.

42. Molecules of nonpolar substances do not interact with polar water molecules, so there is no force that would pull apart water molecules held together by relatively strong intermolecular forces.
43. London forces are the very weak forces by which nonpolar solute and solvent molecules interact.
44. The phrase *Like dissolves like* refers to the fact that polar substances dissolve in other polar substances and nonpolar substances dissolve in other nonpolar substances, but polar substances and nonpolar substances do not dissolve in each other

Chapter 11

1. a. Vapor pressure is the pressure exerted by a gas when it is in equilibrium with its liquid in the same closed container. Thus equilibrium is an integral part of the definition.
b. A solute moves back and forth between the two liquids at the same time and at the same rate.
2. It means that the reaction is a chemical equilibrium and the forward reaction is occurring at the same time and at the same rate as the opposing reaction.
3. The double arrow indicates that both the forward reaction and the reverse reaction are occurring at the same time and at the same rate.
4. A chemical equilibrium is a reaction in which the forward reaction is occurring at the same time and at the same rate as the opposing reaction, with no net change.
5. $2 \text{SO}_2(\text{g}) + \text{O}_2(\text{g}) \rightleftharpoons 2 \text{SO}_3(\text{g})$
6. The equilibrium constant is a constant that indicates the position of an equilibrium at a given temperature. It is the concentrations of the products of a chemical reaction raised to the power of their balancing coefficients and multiplied together divided by the concentrations of the reactants raised to the power of their coefficients and multiplied together. See Section 11.4 for examples.
7. a.
$$K = \frac{[\text{NOCl}]^2}{[\text{NO}]^2 [\text{Cl}_2]}$$

b.
$$K = \frac{[\text{NO}]^2}{[\text{N}_2][\text{O}_2]}$$

$$\text{c. } K = \frac{[\text{HCl}]^4 [\text{O}_2]}{[\text{Cl}_2]^2 [\text{H}_2\text{O}]^2}$$

$$\text{d. } K = \frac{[\text{H}^+][\text{HSO}_3^-]}{[\text{H}_2\text{SO}_3]}$$

$$\text{e. } K = \frac{[\text{H}^+][\text{H}_2\text{PO}_4^-]}{[\text{H}_3\text{PO}_4]}$$

$$\text{f. } K = \frac{[\text{SO}_3]^2}{[\text{SO}_2]^2 [\text{O}_2]}$$

8. A reaction in which a molecular or ionic species is partially converted to ions is described by an ionization constant, K_i . An example would be any weak acid or weak base ionization when dissolved in water.
9. A reaction in which an ionic compound dissolves in water and is partially converted to ions is described by a solubility product constant, K_{sp} . See Equations 11.19 and 11.21.
10. a. $\text{Ag}_2\text{S} \rightleftharpoons 2 \text{Ag}^+ + \text{S}^{2-}$ $K_{sp} = [\text{Ag}^+]^2 [\text{S}^{2-}]$
 b. $\text{PbI}_2 \rightleftharpoons \text{Pb}^{2+} + 2 \text{I}^-$ $K_{sp} = [\text{Pb}^{2+}] [\text{I}^-]^2$
 c. $\text{Bi}_2\text{S}_3 \rightleftharpoons 2 \text{Bi}^{3+} + 3 \text{S}^{2-}$ $K_{sp} = [\text{Bi}^{3+}]^2 [\text{S}^{2-}]^3$
11. It is a very small number. Since the acid ionizes only slightly, the numerator is small compared with the denominator, which makes for a very small number after division.
12. If the equilibrium constant is small, the numerator is small compared with the denominator, which means that the H_2PO_4^- concentration is small compared with the H_3PO_4 concentration.
13. a. $K_{eq} = \frac{[\text{H}^+][\text{H}_2\text{BO}_3^{2-}]}{[\text{H}_3\text{BO}_3]}$
 b. It can be K_i because it is an ionization. It can be K_a because it is a weak acid ionization.
 c. The number 5.86×10^{-10} is very small. This would indicate only slight ionization because the numerator of the K_{eq} expression, representing the products, is small compared with the denominator, which represents the reactants.
14. The equilibrium constant is small, so the numerator is small compared to the denominator. This means that the concentration of the H^+ is small at equilibrium.

15. a. $\text{HCHO} \rightleftharpoons \text{H}^+ + \text{CHO}^-$
b. $K_{\text{eq}} = \frac{[\text{H}^+][\text{CHO}^-]}{[\text{HCHO}]}$
c. It would be a small number because, with only slight ionization, the numerator will be small compared with the denominator.
16. Le Chatelier's principle is a statement that says that if a stress is placed on a system at equilibrium, the equilibrium will respond so as to relieve the stress.
17. Le Chatelier's principle is important in the description of chemical reactions because reaction conditions can represent a stress on a system and chemists must be able to predict the outcome.
18. An endothermic reaction is one in which heat is removed from the surroundings to make the reaction proceed from left to right.
19. The addition of heat would push an endothermic reaction to proceed or shift from left to right.
20. An exothermic reaction is one in which heat is added to the surroundings when the reaction proceeds from left to right.
21. The addition of heat would push an exothermic reaction to proceed or shift from right to left.
22. a. increase
b. decrease
c. increase
23. a. right, decrease
b. left, increase
c. left, increase
24. Affect: a, c, f
Not affect: b, d, e
25. a. left, increase
b. left, increase
c. right, decrease
26. a. left, decrease
b. neither, no change
c. left, decrease
27. The enthalpy change, ΔH , for a reaction is the heat that is either absorbed or given off in the reaction.

- 28. a. left, increase.
 - b. neither, no change
 - c. right, decrease
- 29. a. right, increase
 - b. left, decrease
 - c. left, decrease
- 30. a. right, increase
 - b. left, decrease
 - c. right, increase
- 31. a. increase
 - b. no change
 - c. decrease
 - d. decrease
- 32. a. no
 - b. no
 - c. yes
 - d. no
- 33. a. neither
 - b. neither
 - c. left
 - d. right
- 34. A high concentration of H_2 would be desirable because the equilibrium would shift toward the product in order to consume the added H_2 . A high pressure would be desirable because the equilibrium would shift to the right, where there is less gas. A low temperature would be desirable because the reaction is exothermic and the removal of heat would cause a shift to the right to replace this heat.

Chapter 12

- 1. See Table 12.1
- 2. The hydrogens in this formula are not acidic hydrogens. They are the hydrogens on the methyl group of the acetate ion and, as such, are not released or donated like the acidic hydrogens are.

3. A strong acid is one that completely ionizes when dissolved in water. Examples: hydrochloric acid and sulfuric acid.
A weak acid is one that is not completely ionized when dissolved in water. Examples: acetic acid and carbonic acid.
4. NaHCO_3 , NaH_2PO_4 , and other partially neutralized acids. The hydrogens in the formula are acidic hydrogens.
5. Sodium hydroxide (NaOH), potassium hydroxide (KOH), and ammonium hydroxide (NH_4OH)
6. In concentrated solutions of NH_4OH , the following equilibrium is present:

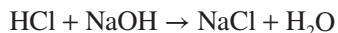


Thus, ammonia, NH_3 , is also present. In addition the following equilibrium is present:



and so the NH_4OH is considered a weak base.

7. Acids have a sour taste, and citric acid is found in citrus fruits.
8. Acids have a sour taste, attack biological materials, are corrosive, ionize to give hydrogen ions in water, and turn litmus paper red.
9. Bases have a bitter taste, attack biological materials, are caustic and corrosive, feel slippery, ionize in water to give hydroxide ions, and turn litmus paper blue.
10. An indicator is a chemical that exhibits various colors in water solutions that vary in acidity level. For example, phenolphthalein is pink in basic conditions and colorless in acidic conditions.
11. pH paper is paper that is saturated with a mixture of various indicators and can thus detect specific acidity and basicity levels by simply displaying different colors.
12. It means that there is a higher concentration of hydrogen ions compared with hydroxide ions. It gives the solution acidic properties.
13. A neutralization reaction is the reaction of an acid with a base. A common example is the reaction of HCl with NaOH :



A salt and water are the products.

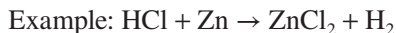
14. See Equation 12.8 for the definition. This is an important equation for water solutions because when hydrogen ions and hydroxide ions are out of balance because of acidic or basic conditions, the concentration of one may be calculated from the concentration of the other.
15. See Equation 12.9 for the definition. This is an important equation because it shows that the pH of a solution may be calculated from the hydrogen ion concentration, and vice versa.
16. a. 2.299
b. 4.032
c. 7.419
d. 10.312
e. 12.218
f. 2.910
17. a. 7.6×10^{-5} M
b. 1.23×10^{-3} M
c. 3.56×10^{-11} M
d. 4.14×10^{-13} M
e. 5.9×10^{-7} M
18. a. 3.277 acidic
b. 1.14×10^{-3} M acidic
c. 2.12 acidic
d. 1.5×10^{-3} M acidic
e. 9.027 basic
f. 3.53×10^{-12} M basic
g. 8.30 basic
h. 1.9×10^{-13} M basic
i. 4.349 acidic
j. 7.98×10^{-7} M acidic
k. 7.08 basic
l. 2.2×10^{11} M basic
m. 10.208 basic
n. 2.24×10^{-6} M acidic
19. a. 11.026
b. 11.310
c. 1.11
20. a. 3.2×10^{-10} M
b. 5.319×10^{-13} M

- c. $1.0 \times 10^{-7} \text{ M}$
 d. $5.1 \times 10^{-4} \text{ M}$
21. a. $2.01 \times 10^{-10} \text{ M}$ 4.304 9.696 acidic
 b. $1.37 \times 10^{-11} \text{ M}$ 10.862 3.138 basic
 c. $5.83 \times 10^{-7} \text{ M}$ $1.714 \times 10^{-8} \text{ M}$ 7.766 acidic
 d. $7.71 \times 10^{-12} \text{ M}$ $1.30 \times 10^{-3} \text{ M}$ 11.113 basic
 e. $1.7 \times 10^{-4} \text{ M}$ 10.22 3.78 basic
 f. $2.6 \times 10^{-6} \text{ M}$ 5.58 8.42 acidic
 g. $4.5 \times 10^{-11} \text{ M}$ $2.2 \times 10^4 \text{ M}$ 3.65 basic
 h. $1.3 \times 10^{-2} \text{ M}$ $7.8 \times 10^{-13} \text{ M}$ 1.89 acidic
22. a. A pH meter is an instrument that, in combination with a pH electrode and a reference electrode, measures and displays the pH of solutions.
 b. A combination pH probe is an electrode that develops a voltage proportional to the pH of a solution into which it is dipped and so, in combination with a pH meter, is useful for measuring pH of solutions. It consists of two concentric glass tubes, one the pH sensor and one a reference electrode.
23. According to the Arrhenius theory, an acid is a compound that donates hydrogen ions and a base is a compound that donates hydroxide ions. An example of an acid is HCl, or any compound whose formula begins with H, and an example of a base is NaOH, or any hydroxide compound.
24. According to the Bronsted–Lowry theory, an acid is a compound that donates hydrogen ions and a base is a compound that accepts hydrogen ions. An example of an acid is HCl, or any compound whose formula begins with H, and examples of bases are NaOH and Na_2CO_3 , or any hydroxide or carbonate.
25. Arrhenius acids: a, c, d
 Bronsted–Lowry acids: a, c, d
 Arrhenius base: f
 Bronsted–Lowry bases: b, e, f, g
26. The hydronium ion is H_3O^+ . It forms when water acquires a hydrogen. In acquiring a hydrogen, it behaves like a Bronsted–Lowry base, which is a substance that accepts hydrogens.

27. A double replacement reaction is one in which two reactants trade partners, meaning one element or polyatomic ion in one formula replaces an element or polyatomic ion in another, and vice versa.



A single replacement reaction is one in which an element reacts with a compound and replaces one of the elements in the formula.



28. a. $\text{Acid} + \text{Metal} \rightarrow \text{Salt} + \text{H}_2$
b. $\text{Metal oxide} + \text{Acid} \rightarrow \text{Salt} + \text{Water}$
c. $\text{Nonmetal oxide} + \text{Water} \rightarrow \text{Acid}$
d. $\text{Metal oxide} + \text{Water} \rightarrow \text{Base}$
29. a. $\text{K}_2\text{O} + \text{H}_2\text{O} \rightarrow 2 \text{KOH}$
b. $2 \text{HCl} + \text{CaCO}_3 \rightarrow \text{CaCl}_2 + \text{CO}_2 + \text{H}_2\text{O}$
c. $2 \text{NH}_4\text{OH} + \text{H}_2\text{SO}_4 \rightarrow (\text{NH}_4)_2\text{SO}_4 + 2 \text{H}_2\text{O}$
d. $\text{NO}_2 + \text{H}_2\text{O} \rightarrow \text{HNO}_3$ (not balanced)
e. $2 \text{HCl} + \text{Mg} \rightarrow \text{MgCl}_2 + \text{H}_2$
f. $\text{CaO} + 2 \text{HBr} \rightarrow \text{CaBr}_2 + \text{H}_2\text{O}$
30. a. $\text{H}_2\text{SO}_4 + \text{K}_2\text{CO}_3 \rightarrow \text{K}_2\text{SO}_4 + \text{CO}_2 + \text{H}_2\text{O}$
b. $2 \text{NO} + \text{H}_2\text{O} \rightarrow 2 \text{HNO}_3$
c. $2 \text{HNO}_3 + \text{Mg}(\text{OH})_2 \rightarrow \text{Mg}(\text{NO}_3)_2 + 2 \text{H}_2\text{O}$
d. $2 \text{K} + 2 \text{H}_2\text{O} \rightarrow 2 \text{KOH} + \text{H}_2$
e. $\text{Ca} + 2 \text{HCl} \rightarrow \text{CaCl}_2 + \text{H}_2$
31. The violent reaction represented by the equation in Exercise 29d would occur.
32. $\text{Mg}(\text{OH})_2 + 2 \text{HCl} \rightarrow \text{MgCl}_2 + 2 \text{H}_2\text{O}$
33. Acid rain is rain that has an acidic pH. It forms when rain falls through an atmosphere that contains oxides of nonmetals. An example is HNO_3 , formed when rain falls through an atmosphere that contains NO_2 .
34. See Section 12.8. In addition, one should take special precautions to cool the container.

Chapter 13

1. The oxidation number of an element is a number assigned to that element that reflects how its electrons are involved in any chemical combination. For example, it is the charge on a monatomic ion. See Section 13.1 for more discussion.
2. The free elemental state of an element is the state the element is in if it is not chemically combined with any other element.
3. For most elements, these are merely the symbols of the elements. For some nonmetals, these are the diatomic molecules of that element.
 - a. Au
 - b. He
 - c. O₂
 - d. Cl₂
4. Assuming that the word *normal* refers to what it is in most cases, we have:
 - a. -2
 - b. +1
 - c. +1
 - d. +2
 - e. +3
 - f. -1
5. -1
6. A peroxide is a compound that contains the peroxide ion, O₂⁻². The oxidation number of the oxygen is -1.
7.
 - a. +1
 - b. +3
 - c. +5
 - d. 0
 - e. -3
 - f. -3
8.
 - a. +1
 - b. -1
 - c. +5
 - d. 0

- e. +3
- f. +5
- 9. a. +3
- b. 0
- c. +6
- d. +6
- e. +6
- 10. a. +7
- b. -2
- c. 0
- d. +3
- e. +1
- 11. a. +4
- b. -2
- c. 0
- d. +6
- e. +4
- f. -2
- g. +6
- h. +4
- i. +4
- j. +6
- 12. a. +5
- b. +3
- c. +5
- d. +3
- e. +3
- f. +5
- g. -3
- h. +6
- i. 0
- j. +5
- 13. a. +2
- b. 0
- c. -1
- d. +4

- e. -2
 - f. +2
 - g. +1
 - h. +4
14. Oxidation—the loss of electrons or the increase in oxidation number.
Reduction—the gain of electrons or the decrease in oxidation number.
Oxidation number—a number indicating the state an element is in with respect to bonding.
Oxidizing agent—a chemical that causes something to be oxidized while being reduced itself.
Reducing agent—a chemical that causes something to be reduced while being oxidized itself.
15. a. CuO (or Cu) was reduced (+2 to 0) and NH₃ (or N) was oxidized (-3 to 0).
b. Cl₂ was reduced (0 to -1) and KBr (Br) was oxidized (-1 to 0)
16. a. Mg is the reducing agent because it was oxidized from 0 to +2. HBr is the oxidizing agent because H is reduced from +1 to 0.
b. Fe is the reducing agent because it was oxidized from 0 to +3. O₂ is the oxidizing agent because it was reduced from 0 to -2.
17. a. is redox. Cu is oxidized and HNO₃ (N) is reduced. Cu is the reducing agent, and HNO₃ is the oxidizing agent.
18. a. is redox. HCl (or H) is reduced and Zn is oxidized. Zn is the reducing agent and HCl is the oxidizing agent.
19. a. No, this is neutralization; there was no change in oxidation number.
b. Yes, both S and N have changed oxidation number.
c. Yes, both Na and H have changed oxidation number.
d. No, this is neutralization; there was no change in oxidation number.
e. No, there was no change in oxidation number.
f. Yes, both K and Br have changed oxidation number.
g. Yes, both Cl and O have changed oxidation number.
h. No, this is neutralization; there was no change in oxidation number.
i. No, there was no change in oxidation number.
j. Yes, Mg and H have changed oxidation number
20. a. $6 \text{ I}^- + 14 \text{ H}^+ + \text{Cr}_2\text{O}_7^{2-} \rightarrow 3 \text{ I}_2 + 2 \text{ Cr}^{3+} + 7 \text{ H}_2\text{O}$
b. $\text{Cl}_2 + \text{H}_2\text{O} + \text{NO}_2^- \rightarrow 2 \text{ Cl}^- + \text{NO}_3^- + 2 \text{ H}^+$

- c. $3 \text{ S} + 2 \text{ NO} + 4 \text{ H}_2\text{O} \rightarrow 3 \text{ S}^{2-} + 8 \text{ H}^+ + 2 \text{ NO}_3^-$
- d. $16 \text{ H}^+ + 2 \text{ MnO}_4^- + 5 \text{ H}_2\text{C}_2\text{O}_4 \rightarrow 2 \text{ Mn}^{2+} + 8 \text{ H}_2\text{O} + 10 \text{ CO}_2 + 10 \text{ H}^+$
- e. $5 \text{ IO}_2^- + 2 \text{ N}_2 + 2 \text{ H}_2\text{O} \rightarrow 5 \text{ I}^- + 4 \text{ H}^+ + 4 \text{ NO}_3^-$
- f. $2 \text{ H}_2\text{O} + 8 \text{ P} + 5 \text{ IO}_4^- \rightarrow 8 \text{ PO}_4^{3-} + 24 \text{ H}^+ + 5 \text{ I}^-$
- g. $3 \text{ H}_2\text{O} + 3 \text{ SO}_2 + \text{ BrO}_3^- \rightarrow 3 \text{ SO}_4^{2-} + 6 \text{ H}^+ + \text{ Br}^-$
- h. $3 \text{ ClO}_4^- + 4 \text{ S} + 4 \text{ H}_2\text{O} \rightarrow 3 \text{ Cl}^- + 8 \text{ H}^+ + 4 \text{ SO}_4^{2-}$
- i. $4 \text{ ClO}_3^- + 6 \text{ H}^+ + 3 \text{ N}^{3-} \rightarrow 3 \text{ H}_2\text{O} + 4 \text{ Cl}^- + 3 \text{ NO}_3^-$
- j. $\text{ClO}^- + \text{SO}_2 + \text{H}_2\text{O} \rightarrow \text{Cl}^- + 2 \text{ H}^+ + \text{SO}_4^{2-}$

Chapter 14

- Organic chemistry is the study of compounds derived from living systems, or systems that were once living. Organic chemistry is the study of the compounds of carbon. Most come from petroleum.
- Carbon atoms and hydrogens atoms.
- Carbon atoms are very prolific in their bonding capabilities and bond freely to one another, often forming long chains sometimes called backbones or a network.
- Hybridization, as defined in this module, is the rearrangement of electrons and orbitals such that carbon has four unpaired electrons, rather than two, and can therefore bond four times.
- Hybridization of a carbon atom means that the four electrons in the outermost energy level become unpaired, each in a separate atomic orbital. This results in carbon's ability to form four bonds with other atoms, including other carbon atoms, and creates the possibility of chains and networks of carbon atoms, which in turn is responsible for the huge number of organic compounds that exist.
- It is tetrahedral when there is nothing but single bonds attached. It is trigonal planar when there is a double bond attached. It is linear when there is a triple bond attached.
- The four properties mentioned in Section 14.5 are: liquids at room temperature, flammability, odorous, and volatile.
- Besides those listed in Exercise 7, other properties include specific gravity, density, viscosity, and refractive index.
- The boiling point will increase due to the increase in the weight of a molecule. More energy (higher temperature) is required to boil a liquid composed of heavier molecules.

10. A functional group is a structural feature that allows a compound to be placed into a particular classification.
11. Aromatic hydrocarbons have the benzene ring; aliphatic hydrocarbons do not.
12. A saturated hydrocarbon is one in which no more hydrogen atoms can be bonded to the carbons; it is saturated with hydrogen atoms. Saturated hydrocarbons are those with only single bonds between the carbon atoms in its structure. Hydrocarbons with double or triple bonds between carbons in the structure are not saturated with hydrogen atoms. More hydrogens can be added to the carbons involved in double or triple bonds because more hydrogens can attach to these atoms (eliminating the double or triple bond).
13. In each of the alkane molecules listed, there are $2n + 2$ hydrogen atoms bonded to n carbons. For example, in methane (CH_4) there is one carbon and $2(1) + 2$, or four, hydrogen atoms.
14. In each of the alkene or cyclic molecules listed, there are two hydrogen atoms for every carbon atom. For example, in ethene, there are two carbon atoms and four hydrogen atoms.
15. In each of the alkyne molecules listed, there are $2n - 2$ hydrogen atoms bonded to n carbon atoms. For example, in ethyne (C_2H_2), there are two carbon atoms and $2n - 2$, or two, hydrogen atoms.
16.
 - a. Alkene or cycloalkane (propylene or cyclopropane)
 - b. Alkyne (1- or 2-butyne)
 - c. Aromatic (benzene)
 - d. Alkene (any isomer of octene, or cyclooctene), cyclooctene
 - e. Probably aromatic (e.g., benzene ring with ethyl group replacing a hydrogen on the ring).
17.
 - a. The simplest alkane is methane. See Figure 14.9.
 - b. The simplest alkene is ethene, or ethylene. See Figure 14.15.
 - c. The simplest alkyne is ethyne, or acetylene. See Figure 14.16.
 - d. The simplest aromatic hydrocarbon is benzene. See Figure 14.21. These structures may also be seen in Table 14.2.
18. See Figure A14.1
19. All except (b) are isomers of the given hexane structure. Structure (b) is not an isomer because it is identical to the given structure; it has the six carbons in a chain. Twisting the chain to make it appear as shown in (b) does not make it an isomer.

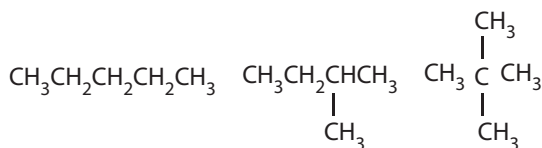


FIGURE A14.1

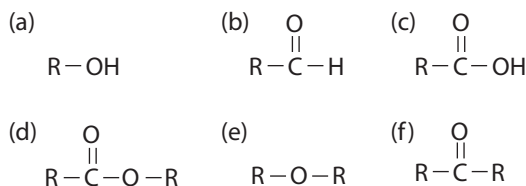
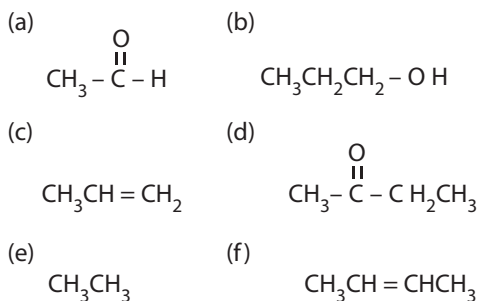
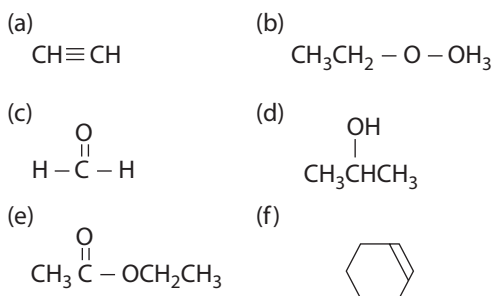


FIGURE A14.2

20. A cyclohexane ring has all single bonds between the carbons in the ring. It represents an alkane. A benzene ring has π -bonding occurring with all the carbons in the ring (besides the σ -bonding), thus creating a unique situation in which π -electrons are delocalized around the entire ring.
21. The structures (b) and (d) are isomers. They both have the same formula ($\text{C}_4\text{H}_{10}\text{O}$) but are different structure.
22. The structures (b) and (e) are isomers. They both have the same formula ($\text{C}_5\text{H}_{10}\text{O}$) but are different structures.
23. See Figure A14.2.
24.
 - a. Ketone
 - b. Carboxylic acid
 - c. Aldehyde
 - d. Alkene
 - e. Aromatic hydrocarbon
 - f. Alcohol
 - g. Alkyne
 - h. Ether
25.
 - a. Acetic acid, or ethanoic acid
 - b. Diethyl ether
 - c. Ethanol, or ethyl alcohol
 - d. Ethylene, or ethene

- e. Methyl acetate
- f. Acetone, or propanone
- g. Toluene
- h. Isobutane
- 26. a. Methanol, or methyl alcohol
- b. n-Butane
- c. Propylene, or propene
- d. Formic acid, or methanoic acid
- e. Dimethyl ether
- f. Methyl formate
- g. Cyclohexane
- h. Benzene
- 27. See Figure A14.3
- 28. See Figure A14.4
- 29. Methane, d
- Acetone, g

**FIGURE A14.3****FIGURE A14.4**

Formaldehyde, g

Acetic acid, c

Methanol, f

Ethanol, b

Isopropyl alcohol, e

30.
 - a. Acetylene—fuel for welder's torches
 - b. Ethyl alcohol—alcoholic beverages
 - c. Acetone—fingernail polish remover
 - d. Methane—natural gas
 - e. Polyethylene—plastic bottles
 - f. Methyl alcohol—windshield washer fluid
 - g. Diethyl ether—formerly used in hospitals as an anesthetic
 - h. Isopropyl alcohol—rubbing alcohol
 - i. Formaldehyde—biological specimen preserver
 - j. Acetic acid—vinegar
31. A polymer is a material consisting of very large molecules that form when small monomer units attach to one another to form a very long chain of such units.
32. Both take place by breaking one of the bonds of the double bond, creating sites on these adjacent carbons where new substituents can attach. The new substituents, in the case of polymers, are other monomer units that have undergone the same bond-breaking process.
33.
 - a. The addition polymer is polyvinyl chloride. The structure of the polymer is given in Figure 14.33c.
 - b. The addition polymer is polypropylene. The structure of the polymer is given in Figure 14.33b.
 - c. The addition polymer is Teflon. The structure of this polymer is given in Figure 14.33e.
34.
 - a. See Figure 14.33f
 - b. See Figure 14.33a
 - c. See Figure 14.33d
35. See Figure 14.33a through f

Index

A

- Absolute zero, 52
- Acetylene, triple bond, 146
- Acidic hydrogen, 79
- Acids
 - “add acid to water” rule, 324
 - Arrhenius theory, 319
 - behavior of, 307
 - Bronsted–Lowry theory, 319
 - hydronium ion, 321
 - list of bases, 320
 - chemical properties
 - hydroxide ions released, 309
 - covalent compounds, 79
 - defined, 304
 - dissociation constant, 286
 - formulas, 305
 - identifying and naming, 78
 - ion product constant, 310–311
 - neutralization, 321–322
 - pH and
 - hydrogen ion, 311–312
 - logarithm of number, 312–314
 - pOH, 315–316
 - significant figures rule, 315
 - physical properties, 308
 - rain, 323
 - reaction with metal
 - single replacement, 322
 - sodium with water, 322
 - zinc with hydrochloric acid, 323
 - reaction with metal oxide, 324
 - water ionization, 310–311
- Actual yield, 207
- Addition polymers
 - formation, 372–375
- Air composition, 249
- Air pressure, 221
- Alcohols, 363
 - IUPAC names, 364
- Alcohols and solubilities in water, 263
- Aldehydes, 365–366
- Aliphatic hydrocarbons, 351
- Alkali metal, 65
 - lithium, 122
 - potassium, 123
 - rubidium, 123
 - sodium, 122
- Alkaline earth metals
 - barium, 123–124
 - beryllium, 123
 - calcium, 123
 - list of, 123
 - magnesium, 123
 - properties of, 123
 - radium, 124
 - strontium, 123
- Alkane, 351–352
 - IUPAC names of, 356
- Alkyne
 - IUPAC names, 357–358
- Alpha particles, 90
- Aluminum
 - Hall–Héroult process, 127
 - uses, 127
- Ammonium hydroxide in water, 308
- Ammonium ion, 153
- Amu defined, 13
- Anions, 70
- Aromatic hydrocarbons, 358
- Arrhenius theory, 319
- Atmospheric pressure
 - height of mercury, 221

- metric system for units
 - atmospheres, 222–223
 - torr, 222–223
- relationship between units, 222
- Torricelli barometer for, 221–222
- Atomic size
 - and electrons in nucleus, 130
 - idiosyncrasies, 131
 - principal energy levels
 - and, 130–131
 - trends of, 131
 - exceptions, 131
 - relative sizes, 131–132
 - variations, 130
- Atoms, 9–10
 - atomic mass unit, 13
 - atomic numbers, 7
 - atomic weight, 18–19
 - of compounds, 175
 - of elements, 175
 - chemical combination, 88
 - electrons
 - arrangement within, 107
 - gain/loss by, 67
 - of elements, 67–68
 - microscopic study, 14
 - modern theory of, 14
 - nucleus, 14
 - quantum mechanical model, 92
 - Rutherford experiment
 - gold atom, 90–91
 - observations, 91
 - solar system model, 92
 - subatomic particles in, 15
 - theories of atomic structure, 89, 93
- Avogadro's number, 33–34, 179, 234–235

B

- Balancing of chemical equations
 - ion-electron method for
 - procedure steps, 336–337
 - law of conservation of mass, 194–195
 - process, 195
 - quantitative information from, 193–194
 - rules, 196
 - trial-and-error inspection process, 196
- Ball-and-stick model

- benzene, 362
 - of carbon atom, 347–348
- cyclohexane, 359
- ethene, 349
- ethyne, 349
- isobutane, 354
- n-butane, 354
- propene, 349
- Bases, 80
 - Arrhenius theory, 319
 - behavior of, 307
 - Bronsted–Lowry theory, 319
 - hydronium ion, 321
 - list of bases, 320
 - chemical properties
 - hydrogen ions released, 309
 - neutralization, 321–322
 - physical properties, 308
- Benzene, 361
 - ball-and-stick model, 362
 - bond line structure, 362
- Beta particles, 90
- Big Six polyatomic ions, 69
 - negative ions, 76
- Binary acids, 79
 - formulas and names, 80
- Binary covalent compounds, 75
 - naming scheme for, 76–78
- Bohr model, 97
 - and light, 98–99
- Boiling point, 3
- Boyle's law, 224–225
- Bronsted–Lowry theory, 319
 - hydronium ion, 321
 - list of bases, 320
- Buffer solutions, 284

C

- Calculations based on equations
 - conversions
 - factors, 202–203
 - gram-to-gram, 204
 - gram-to-mole, 203
 - mole-to-gram, 203
 - mole-to-mole, 202–203
 - observable and measurable quantities
 - of matter, 201
 - stoichiometry calculations, 204

- Cancellation of units, 35
- Carbon
 - bonding of, 345
 - compounds, 125
 - alkane, 351–352
 - alkenes, 356–357
 - alkyne, 357–358
 - isomers, 352–356
 - ring structures, 358
 - elemental, electron
 - structures, 346
 - geometry around bonded
 - atoms, 346
 - ball-and-stick model, 347–348
 - sawtooth appearance, 347
 - Lewis structures of electrons, 346
 - physical properties, 349–350
- Carbon-12 atom, 13
- Catalyst defined, 296
- Cathode ray phenomena, 92
- Cations, 70
- Celsius scale, 52
- Change
 - types
 - physical and chemical, 2
- Charles' law, 227–228
- Chemical bonding
 - covalent bonding, 140
 - molecular orbitals, 165
 - polar bonds, 158
 - unequal electronegativities, 158
 - definition of, 140
 - discreet units, 140
 - electron configurations, 142
 - ionic compounds, 140
 - molecular compounds, 140
 - transformation properties, 140
- Chemical energy, 93–94
- Chemical equations
 - balancing
 - law of conservation of
 - mass, 194–195
 - process, 195
 - quantitative information
 - from, 193–194
 - rules, 196
 - trial-and-error inspection
 - process, 96
 - examples of, 199
- Chemical equilibrium, 282
 - examples of
 - ammonium hydroxide in water, 284
 - carboxylic acids in water, 284
 - Haber process, 283–284
- Chemical formulas, 11
- Chemical properties, 3–4
- Chemical reaction
 - defined, 192
 - reactants and products, 193
 - statement of
 - chemical equation, 192
- Chemistry defined, 2
- Citric acid structure, 308
- Cobalt, 128
- Combination pH electrodes, 319
- Complex mixtures, 4
- Compounds, 5
 - chemical combination, 7–8
 - fundamental unit, 10–11
 - homogeneous and heterogeneous, 8
 - ionic, 11
 - microscopic study, 12
 - molecular/covalent, 11
 - one mole of, 178
 - percent composition of
 - atomic and formula weights, 180
 - defined, 179
 - element, 180
 - from formulas, 180
 - solubilities in water, 260
 - symbolic representation
 - chemical formula, 7
- Concentrated solutions, 262–263
- Concentration, 249
 - change effect, 291–292
 - molarity, 251
 - percent, 250
- Conversions
 - factor, 37–38
 - gram-to-gram, 204
 - gram-to-mole, 203
 - mole-to-gram, 203
 - mole-to-mole
 - conversion factors, 202–203
- Coordinate covalent bonds
 - ammonia molecule, 153
 - ion, 152

- definition of, 152
 - formation, 153
 - hydrogen ion, 153
 - nitrogen atom, 153
 - orbital, 152
 - slot-filling process, 152
 - structure of, 153
- Copper, uses, 128
- Covalent compounds
- binary compounds, 146
 - covalent bonding, 141
 - defined, 140
 - single bond, 143
 - double bond, 144
 - hydrogen atom, 145
 - ethylene formation, 145
 - hybridization, 145
 - hydrogen atom, 146
- Lewis electron-dot structures
- bromine, 143
 - chlorine, 143
 - fluorine, 143
 - iodine, 143
 - molecular hydrogen, 143
 - molecular orbitals, 143
 - form, 145
 - polyatomic ion, 146
 - slot-filling procedure, 146–147
 - slots, 143
 - ternary compounds, 146
 - triple bond
 - acetylene, 146
 - unpaired electrons, 143
- Cyclohexane
- ball-and-stick model, 359
 - bond line structure, 359
- D**
- Dalton's law, 235–236
- atomic theory, 89
- Density, 3
- as conversion factor, 49–50
 - defined, 45
 - irregularly shaped solids
 - measuring, 48
 - of liquids, 46
 - measuring, 49
 - in metric system, 46
 - of solids, 46
 - measuring of, 47
 - temperature, effect on, 46
 - of water, 51
- Descriptive chemistry of elements
- alkali metals, 122
 - alkaline earth metals
 - list of, 123
 - properties of, 123
 - aluminum and silicon, 126
 - carbon and nitrogen, 125
 - defined, 122
 - halogens, 124
 - hydrogen and helium, 125
 - oxygen and phosphorus, 125
 - sulfur, 125
 - top-row transition elements, 127
- Diethyl ether, 365
- Digital measuring devices, 29
- Dilution of solutions, 262–263
- concentrated solution, 252
 - conversion factor, 254
 - device and container, choice of, 255
 - molarity multiplying, 253
 - moles of solute
 - before and after, 253
 - units, 254
- Dimensional analysis, 34
- cancellation of units, 35
 - concept of, 35
- Dissolution rate
- agitation, 264
 - particle size, 263
 - temperatures, 264
- Dissolving mechanism
- ionic solutes in water, 264–266
 - molecular solutes in water, 266–268
 - nonaqueous solvents, 268–270
- Double bond, 144–145
- representation of, 167
- Double replacement reaction, 321–322
- E**
- Eka-silicon, 121
- Electrical charge of particles, 13
- Electrical energy, 93–94
- Electrolysis process, 199
- Electromagnetic spectrum, 96

- Electromagnetic waves, 94
- Electron affinity, 131
 - graph, 133
 - inner-level electrons, 132
 - outer-level electrons, 134
 - outermost electrons, 132
 - screening effect, 134
 - top-bottom behavior, 134
- Electronegativity
 - defined, 156–158
 - and bond polarity
 - negative pole, 157
 - nonpolar, 157
 - orbital unsymmetrical, 157
 - Pauling's electronegativities of
 - nonmetals, 157
 - polar bonds, 158
 - positive pole, 157
- Electrons
 - configuration
 - elements in periodic table and, 111
 - energy levels and, 110
 - hydrogen and helium, 107
 - mnemonic device for, 110
 - overlapping, 109
 - of oxygen, 108
 - of transition elements, 113
 - of third period, 108
 - discovery and characterization, 90
 - plum pudding model, 90
- Elemental carbon, 125–126
- Elemental sulfur, 126
- Elements
 - atomic numbers, 7
 - atoms of, 9–10
 - homogeneous and heterogeneous, 8
 - metals and nonmetals, 5
 - microscopic study, 10
 - molecule/formula unit, 11
 - names and, 7
 - one mole of, 178
 - in periodic table
 - information, 16
 - location of, 15
 - properties of metals, 64
 - subclassifications
 - alkali metals, 65
 - alkaline earth metals, 65–66
 - halogens, 66
 - inner transition elements, 64–65
 - noble gases, 66
 - nonmetals and metalloids, 64
 - representative elements, 64
 - transition elements, 64
 - symbols for, 5, 12
- Empirical formulas
 - derivation of, 182
 - and molecular formula, 185
 - from percent composition
 - step-by-step procedure for, 183
- Endothermic reactions, 293
- Energy defined, 93
- English system, 38
- Enthalpy change, 294
- Equality, 37
- Equilibrium, 280
 - chemical equilibrium, 282–284
 - constant, 284–285
 - ionization, 286
 - magnitude, 288–290
 - solubility product constant, 286–287
 - symbols used, 287
 - physical equilibrium, 281–282
 - position of, 287
 - illustration, 288
 - pressure change effect, 294–295
 - shift, 290
 - temperature change, 293–294
- Erlenmeyer flasks, 255
- Esters, 368–369
- Ethane, 352
- Ethanol, 364
- Ethene
 - ball-and-stick models, 349
- Ethers, 364–365
- Ethylene polymerization, 372–375
- Ethyne
 - ball-and-stick models, 349
- Examples
 - balancing equation, 196–198
 - conversion, 38
 - problems, 42–45
 - density, 47–51
 - diluting solutions, 254–256
 - empirical formula, 184–186
 - formula of compound, 72–73
 - formula weight, 176–177

gas pressure, 223, 226, 236
 gas volume, 229–234
 gram conversion, 205
 ion-electron method, 337–338
 Lewis electron dot structure, 147,
 149–155
 mass number, 17
 molarity, 251–252
 molecular structure, 164–165
 moles, 178–179
 conversion, 202–204
 oxidation number, 332–335
 percent composition, 180–182
 percent yield, 207–208
 periodic trends, 134–135
 pH, 311, 314–315, 317–318
 pOH, 315–316
 solutions preparation, 257–259
 temperature scales, 53
 volume percent, 250
 weight/volume percent, 251
 yield, 206
 Exothermic reactions, 293

F

Fahrenheit scale, 52
 “Fifty Kelvin,” 52
 Formaldehyde, 366
 Formula weight, 176

G

Galvanized steel, 128
 Gases
 kinetic-molecular theory
 collision in, 218
 gas composition, 218
 kinetic energy, 218–219
 motion, 219
 laws
 Boyle’s law, 224–225
 Charles’ law, 227–228
 combined gas law, 230
 piston-cylinder arrangement,
 224–225
 pressure, 220
 height of mercury, 221
 metric system for units, 222–223

 relationship between units, 222
 Torricelli barometer for, 221–222
 properties
 compressed, 214–215
 contract, 216
 diffuse through one another,
 215–216
 expand, 215–216
 in mythical microscope, 219–220
 pressure, 215, 218
 volume after cooling and heating,
 217–218
 solubility behavior, 261
 Gold foil experiment, 90–91
 Graduated cylinder, 30–31

H

Hall–Héroult process, 127
 Halogens, 66, 124
 astatine, 124
 bromine, 124
 chlorine, 124
 fluorine, 124
 group of, 124
 iodine, 124
 Heat energy, 93–94
 Heat of reaction, 294
 Helium
 uses, 125
 Heterogeneous equilibrium, 286
 High-precision transfer devices, 255
 precision, 256
 Homogeneous physical mixture, 248–249
 Hund’s rule, 103, 141
 Hybridization
 carbon compound, 145
 Hybridization of orbitals, 345–346
 Hydrates, 161
 cations and anions, 161
 crystal structure, 161
 definition of, 161
 Hydrocarbons
 classification
 flow chart of, 351, 360, 363, 365,
 368–369
 functional groups, 350
 Hydrochloric acid
 ionization reaction for, 305

Hydrogen uses, 125

Hydrogen bond

- definition of, 160
- electronegative atom, 160
- fluorine atom, 160
- intermolecular force, 160
- nitrogen atom, 160
- oxygen atom, 160
- properties of, 160

Hydroxide bases

- formulas, 307

I

Ideal gas law

- gas at constant temperature and pressure, 232
- gas constant, 233
- number of moles of gas, 232

Immiscible solutions, 262

Infrared wavelength, 96

Inner-level electrons, 134

Inner transition elements, 64–65

Inorganic acids, 304

Insoluble in solvent, 260

Ion-electron method for balancing

- equations
- procedure steps, 336–337

Ionic compounds, 153, 175

- alkali metals, 142
- alkaline earth metals, 142
- aluminum, 142
- crystal, 264–265
- formulas, 69
- illustration of, 265
- ionic bonding
 - definition of, 140
 - electrostatic forces, 140
 - monatomic ions, 140
 - polyatomic ions, 140
- monatomic ions, 142
- naming scheme for, 73–74
- positive ion, 70
- salts and bases, 80
- of sodium sulfate, 154
- solubility in water, 266
- structure
 - anions, 155
 - cations, 155

covalent compound, 155

of sodium chloride, 156

Ionic solutes in water, 264–266

Ionization constants, 286

Ionization energy, 131

- graph, 133
- inner-level electrons, 132
- outer-level electrons, 134
- outermost electrons, 132
- screening effect, 134
- top-bottom behavior, 134

Ions

- cations and anions
 - one-to-two combinations, 71
- defined, 16
- formation of, 66–69

Iron, 127

Isobutane, 355

ball-and-stick model, 354

Isopropyl alcohol, 364

Isotopes, 16

- for carbon, 19
- for hydrogen, 19
- radioactive, 19
- symbols of, 18

K

Kelvin scale, 52

Ketones, 366–367

Kinetic energy, 93

defined, 218

Kinetic-molecular theory of gases

- collision in, 218
- kinetic energy, 218–219
- motion, 219

L

Laboratory setting, 237

Law of conservation of mass, 194–195

Le Chatelier's principle, 290

Length measurement, 39–40; *see also* Metric system

Lewis electron-dot structure, 345

- of ammonia, 145
- of ammonium phosphate, 155
- of carbonate ion, 151

- of carbon tetrachloride, 149
- of chloric acid, 149
- of chlorine gas, 144
- concept of, 142
- covalent bonding, 143
- covalent compounds
 - steps in, 148
- of elements, 142
- hydrogen chloride, 144
- of hydrogen gas, 144
- ionic compounds, 143, 153
 - alkali metals, 142
 - alkaline earth metals, 142
 - aluminum, 142
 - polyatomic ion, 153
- of monatomic ions
 - chloride and oxide, 143
 - magnesium and aluminum, 143
 - sodium, 143
- of nitric acid, 150
- of nitrogen, 141
- of phosphorus pentachloride, 151
- polyatomic ions
 - steps in, 148
- slot-filling method, 143
- of sodium, 143
- of sodium sulfate, 154
- stable configuration of
 - helium, 144
- valence-shell electron, 162
- of water, 145
- Light energy, 93–94
- Light, modern theory, 92
 - energy, 93–94
 - photons, 94
 - properties of, 94
 - wave theory, 94
- Limiting reactant, 205–206
- Liquid volume measurement, 30
- Litmus paper as indicator, 309

M

- Mass
 - defined, 13
 - measurement
 - equalities and conversion
 - factors, 41
 - number, 17
- Material substances
 - properties, 2
 - physical and chemical, 3–4
 - structure and composition
 - compounds, 10–12
 - elements, 9–10
- Matter
 - classification, 4
 - based on purity, 5
 - elements and compounds, 5
- Measuring device
 - reading, 28
 - rules, 29
- Mechanical energy, 93
- Melting point, 3
- Mendeleev periodic table
 - development of, 120
 - with periods and families, 121
 - properties of elements, 121
 - space for undiscovered
 - elements, 121
- Metal ions, 75–76
- Metalloids, 64
- Metal oxides reaction with water, 323
- Methyl orange as indicator, 309
- Metric system
 - characteristic of, 42
 - of measurement, 38
 - length, 39–40
 - mass, 40
 - volume, 40–41
 - prefixes used in, 39
- Meyer periodic table
 - development of, 120
 - with periods and families, 121
 - properties of elements, 121
 - space for undiscovered elements, 121
- Millikan oil drop experiment, 90
- Mineral acids, 304
- Miscible solutions, 262
- Mnemonic device, 110
- Molarity defined, 251
- Mole
 - atomic and formula weight
 - Avogadro's number, 179
 - conversion from grams, 177
 - defined, 177
- Molecular compounds, 175
 - fundamental unit of, 176

Molecular orbitals, 143
 chlorine gas, 165–166
 hydrogen chloride, 166
 overlapping
 end-to-end, 165–166
 side-to-side, 166
 sigma bond, 166
 Molecular solutes in water, 266–268
 Molecular structure
 chemical bonding
 definition of, 140
 discreet units, 140
 extinguishes flames, 140
 ionic compounds, 140
 molecular compounds, 140
 transformation properties, 140
 complicated covalent compounds, 146
 coordinate covalent bonds, 152
 molecular orbitals, 165
 more complicated ionic compounds, 153
 outermost electrons
 definition of, 141
 uncombined state, 141
 valence electrons, 141
 polar bonds and molecules, 156
 simple covalent compounds, 143
 simple ionic compounds, 142
 monatomic ions, 142
 structure of ionic compounds, 155
 valence-shell electron, 162
 Molecule, 11, 155
 definition of, 156
 Mole map, 204
 Mole-to-mole conversions
 conversion factors, 202–203
 Monatomic ions, 66, 142
 Lewis electron-dot structures
 aluminum, 143
 chloride and oxide, 143
 sodium and magnesium, 143
 Monatomic negative ions, 76

N

Neutralization, 310, 321–322
 Neutrons
 discovery, 92
 Nickel, 128
 uses of, 128
 Nitrogen
 compounds, 126
 importance of, 126
 Noble gases, 66
 list of, 125
 stable states, 67
 Nonaqueous solvents, 268–270
 Nondigital measuring devices, 30
 graduated cylinders, 31
 thermometers, 31
 Nonmetal oxides reaction with water, 323
 Nonmetals, 64
 Nonpolar bond
 illustration, 158
 Nonpolar liquids
 illustration of, 2270
 Nuclear energy, 93–94
 Nuclear model, *see* Rutherford model
 Nuclear reactions, 194

O

Open-chain alkanes, 356
 Orbitals hybridization, 345–346
 Orbitals of electrons, 101
 diagram of
 elements in third period, 106
 first four elements, 105
 oxygen, 108
 spin direction, 104–105
 Hund's rule, 103
 overlap effect, 109
 between energy levels, 110
 p orbitals, 103
 s orbitals, 102
 Organic acids, 304
 structural feature, 306
 Organic chemistry
 carbon compounds, 346
 alkane, 351–352
 alkenes, 356–357
 alkyne, 357–358
 bonding of, 345
 chemistry, 345
 elemental, electron
 structures, 346
 geometry around bonded
 atoms, 346
 isomers, 352–356

- Lewis structures of electrons, 346
 - physical properties, 349–350
 - ring structures, 358
 - compounds with oxygen
 - alcohols, 363
 - aldehydes, 365–366
 - carboxylic acid, 367–368
 - esters, 368–369
 - ethers, 364–365
 - ketones, 366–367
 - R-group, 362
 - defined, 344
 - hydrocarbons, 350
 - IUPAC nomenclature systems, 351
 - Outer-level electrons, 134
 - Outermost electrons
 - definition of, 141
 - Hund's rule, 141
 - Lewis electron-dot structure, 141–142
 - of elements, 142
 - of nitrogen, 141
 - orbital, 141
 - paired, 141
 - spherical orbital, 141
 - symbolism, 141
 - uncombined state, 141
 - valence electrons, 141
 - Overlap
 - end-to-end
 - atomic orbitals, 165
 - structure of, 166
 - side-to-side, 165
 - double bond, 165
 - pi bonding, 165
 - structure, 165
 - structure of double bond, 167
 - triple bond, 165
 - Oxidation
 - defined, 335
 - number, 76, 330
 - combined halogen atom, 332
 - determination of, 331
 - rules for assigning, 331
 - oxidizing agent, 335
 - state, 330
 - Oxidation–reduction reactions
 - sodium and chlorine atom, 330
 - Oxyacids, 79
 - Oxygen compounds, 126
- P**
- Partial pressure, 237
 - of water vapor, 238
 - Particle theory, 94
 - Pauling's electronegativities of
 - nonmetals, 156–157
 - Percent yield, 207
 - Periodicity and trends
 - alkali metal, 129
 - alkaline earth metal, 129
 - atomic size
 - electrons in nucleus, 130–131
 - energy levels, 130–131
 - principal energy levels, 130–131
 - day–night temperature cycle, 129
 - electron configuration pattern, 129
 - ionization energy and electron
 - affinity, 131
 - graph, 133
 - inner-level electrons, 132
 - outer-level electrons, 134
 - outermost electrons, 132
 - screening effect, 134
 - top–bottom behavior, 134
 - periodic law, 129
 - Periodic law properties, 129
 - Periodic table, 6
 - alkaline earth metals, 123
 - arrangement of
 - electron configuration and, 111
 - periods and families, 5
 - carbon and nitrogen, 125
 - definition of, 5, 129
 - descriptive chemistry of elements, 121
 - elements list, 120
 - groups of, 120
 - hydrogen and helium, 125
 - importance of, 120
 - Mendeleev and Meyer, 120
 - names and symbols of elements, 5
 - outermost electrons
 - Lewis electron-dot structure, 141
 - outline of, 65, 67, 111–112
 - oxygen and phosphorus, 125
 - periodicity and trends
 - meaning of, 128
 - properties of, 120
 - sulfur, 125

- Phenolphthalein as indicator, 309
- Phosphorus
- forms of, 126
 - importance of, 126
- Photons, 94
- pH paper as indicator, 309
- pH scale, 317
- hydrogen ion, 311–312
 - logarithm of number, 312–314
 - measurement of
 - combination pH electrodes, 319
 - pH electrode, 318
 - pH meter, 318
 - reference electrode, 318–319
 - pOH, 315–316
 - significant figures rule, 315
- Physical change, 2
- Physical equilibrium, 281–282
- Physical properties, 3–4
- pi bond, 166
- side-to-side overlap of p orbitals, 167
- Plum pudding model, 90
- Polar bonds and molecules, 156
- clouds, 156
 - hydrates, 161–162
 - illustration of, 158
 - molecular orbitals, 156
 - negative pole, 157
 - Pauling's
 - electronegativity and, 156–157
 - polarity of molecules, 158–159
 - positive pole, 157
 - water and hydrogen bond, 159–160
- Polar covalent compound, 267
- illustration of, 269
- Polarity of molecules
- asymmetrical water molecule
 - structure of, 159
 - ball-and-stick models, 158
 - carbon atom, 159
 - C–H bond polarities, 159
 - C–H bonds, 158
 - four-faced 3-d shape, 159
 - hydrogen atoms, 159
 - methane molecule, 158
 - molecular arrangements, 158
 - molecular geometry, 159
 - nonpolar molecule, 159
 - nonsymmetrical molecule, 159
 - symmetrical methane molecule
 - structure of, 159
 - symmetrical structure, 159
 - tetrahedral arrangement, 158–159
 - pyramid-like structure, 159
 - tetrahedral corners, 159
 - tetrahedron, 159
 - water molecule, 158–159
- Polar water molecules
- hydrates, 160
- Polyatomic ions, 66, 179
- formation of, 69
- Polyethylene, 372–375
- Polymers, 370
- addition, 371–375
- Polypropylene, 372–375
- Polystyrene, 372–375
- Polytetrafluoroethylene (PTFE), 372–375
- Polyvinyl chloride (PVC), 372–375
- Potential energy, 93
- Precipitate defined, 201
- Precision, 30
- Pressure change effect on
- equilibrium, 294–295
- Propene
- ball-and-stick models, 349
 - condensed structure, 353
- PTFE, *see* Polytetrafluoroethylene (PTFE)
- ## Q
- Quantitative solubility, 260
- Quantum mechanical model, 92; *see also*
- Atoms
 - electrons
 - configuration, 99
 - energy state, 102
 - number of, 100
 - orbitals, 101–102
 - principal energy levels for, 99
 - sublevels, 100
 - energy levels around nucleus, 100
- ## R
- Reactions
- acetic acid
 - ionization, 306
 - with water, 284

- acids with metal
 - single replacement, 322
 - sodium with water, 322
 - zinc with hydrochloric acid, 323
 - acids with metal oxide, 324
 - addition reactions, 372
 - ammonium hydroxide
 - ionization, 308
 - with water, 284
 - carbon oxides with water, 323
 - carboxylic acids with water, 284
 - chemical
 - defined, 192
 - reactants and products, 193
 - statement of equation, 192
 - double replacement
 - reaction, 321–322
 - endothermic reactions, 293
 - exothermic reactions, 293
 - Haber process, 283–284, 294–295
 - hydrochloric acid ionization, 305
 - hydrogen-oxygen reaction, 202
 - magnesium oxides with water, 323
 - metal oxides with water, 323
 - neutralization, 322
 - nitrogen oxides with water, 323
 - nonmetal oxides with water, 323
 - nuclear, 194
 - oxidation–reduction reactions
 - sodium and chlorine atom, 330
 - phosphorus oxides with water, 323
 - reversible, 207
 - side reactions, 207
 - sodium chloride with water, 284
 - sodium oxides with water, 323
 - sodium with water, 322
 - sulfuric acid with water, 324
 - sulfur oxides with water, 323
 - water formation, 192
 - yield of, 205
 - types, 207
 - zinc with hydrochloric acid, 323
- Reduction, 330
- defined, 335
 - reducing agent, 335
- Reference electrode, 318–319
- Representative elements, 64
- Reversible equilibrium, 280–282
- Reversible reactions, 207; *see also* Yield of reaction
- Rutherford model, 90–91
- S**
- Salt, 80–81
- Saran, 372–375
- Saturated solutions, 261
- Schrödinger model, *see* Quantum mechanical model
- Scientific notation, 32
- Side reactions, 207; *see also* Yield of reaction
- Sigma bond, 166
- Significant figures, 30–31
- rounding, 32–33
 - rules for, 34
 - counting, 35
 - determining, 36–37
 - scientific notation, 33–34
- Silicon uses, 127
- Single bond, 143
- Sloppy lab technique, 207; *see also* Yield of reaction
- Slot-filling method
 - of ammonia, 145
 - ethylene, 146
 - hydrogen chloride, 144
 - of water, 145
- Solar system model, 92; *see also* Atoms
- Solubility
 - defined, 260
 - and temperature, 261
- Solubility product constant, 286
- Soluble in solvent, 260
- Solute, 249
- Solutions
 - components of
 - solvent and solute, 249
 - concentration of solute, 249
 - defined, 248
 - dilute and concentrated, 262–263
 - homogeneous, 248–249
 - miscible and immiscible, 262
 - preparation of
 - dilution, 251–254
 - liquid solute and volume percent, 256–257

solid solute and molarity, 257–259
 solid solute and weight-to-volume percent, 257
 saturated and unsaturated, 261
 solubility and, 260–261
 supersaturated, 261
 Solvent, 249
 Standard temperature and pressure (STP)
 correction of gas volume to, 231
 initial and final conditions, 231
 Static charge of particles, 13
 Stoichiometry, 174
 STP, *see* Standard temperature and pressure (STP)
 Strong acids, 306
 dissolving in water, 267
 Strong bases, 307
 Styrene polymerization, 372–375
 Styrofoam, 372–375
 Subatomic particles
 charges on, 14
 number of
 atomic weight, 18–19
 neutrons and isotopes, 16
 protons and electrons, 15
 Sulfur, 126
 properties of, 126
 uses, 126
 Sulfuric acid reaction with water, 324
 Supersaturated solutions, 261–262
 Symbols
 and physical state of substance
 aqueous, 200
 gaseous, 200
 solid, 200
 qualitative information, 199
 and symbolism, 141

T

Teflon, 372–375
 Temperature
 change effect on equilibrium, 293–294
 defined, 51
 metric system scale for, 52
 comparison, 53
 Ternary acids, 79
 formulas and names, 80
 Ternary compounds, 75

Tetrahedron, 159
 Theoretical yield, 207
 Theories
 Arrhenius theory, 319
 Bronsted–Lowry theory, 319–321
 Dalton's atomic theory, 89
 kinetic-molecular theory of gases,
 218–219
 modern theory
 of atoms, 14
 of light, 92
 particle theory, 94
 valence-shell electron-pair repulsion
 (VSEPR) theory, 162–164
 wave theory, 94
 Thermometers, 31
 Top-bottom behavior, 134
 Top-row transition elements, 127
 corrosion properties, 127
 list of, 127
 unreactive metals, 127
 Torricelli barometer for, 221–222
 Transition elements, 64
 Trends
 of atomic size, 131
 of monatomic ion charges, 129–130
 Triple bond, 145–146
 Trivial matter, 34

U

Ultraviolet wavelengths, 96
 Uncombined state, 141
 Universal gas constant, 233
 Unpaired electrons, 143
 Unsaturated solutions, 261

V

Valence electrons, 141
 Valence-shell electron-pair repulsion
 (VSEPR) theory
 angular, 164
 double bond, 162
 molecule's structure, 162
 nonbonding electrons, 162
 pair repulsion theory, 162
 polar or nonpolar, 162
 regions of water and methane, 162

- repulsion, 162
- shapes
 - linear, 162–164
 - octahedral, 162–164
 - tetrahedral, 162–164
 - trigonal bipyramidal, 162–164
 - trigonal planar, 162–164
- single bond, 162
- structure of, 163
- symmetrical/nonsymmetrical, 162
- triple bond, 162
- Vapor pressure, 236
 - evaporation of water, 237
 - humidity percentages, 237
 - of water, 238–239
- Vinegar, 368
- Visible wavelengths, 95
- Volume measurement, 40
 - of cubic object, 41–42
 - equalities and conversion factors, 41
- Volume percent, 250
- VSEPR theory, *see* Valence-shell electron-pair repulsion (VSEPR) theory

W

- Water
 - boiling point, 237
 - formation reaction, 192
 - hydrogen bond in, 159
 - boiling point, 160
 - definition of, 160
 - DNA molecular strands, 160
 - intermolecular attraction, 160
 - intermolecular force, 160

- magnets, 160
- methane, 160
- polar/nonpolar nature, 159–160
- structure of, 160
- void spaces, 160
- molecules
 - polar nature, 267
 - study by mythical microscope, 12
- structure of, 159
- wave, 94
- Wavelength
 - concept of, 95
 - defined, 94–95
 - infrared, 96
 - range of, 95–96
 - ultraviolet, 96
 - visible, 95
- Wave theory, 94
- Weak acids, 306
 - dissolving in water, 268
- Weak base, 308
- Weak electrolytes solutions, 284
- Weight/volume percent, 250

Y

- Yield of reactions, 205
 - types
 - actual, 207
 - percent, 207
 - theoretical, 207

Z

- Zinc uses, 128

Element symbol

Atomic number — 1.008 — Atomic weight

13	14	15	16	17	
IIIA	IVA	VA	VIA	VIIA	
B	C	N	O	F	He
Boron	Carbon	Nitrogen	Oxygen	Fluorine	Helium

5	6	7	8	9	10	11	12	
Al	Si	P	S	Cl	Ar			
Aluminum	Silicon	Phosphorus	Sulfur	Chlorine	Argon			
13	14	15	16	17	18	19	20	21
B	C	N	O	F	Ne			
Boron	Carbon	Nitrogen	Oxygen	Fluorine	Neon			

Element name

1	2	
H	He	
Hydrogen	Helium	

3	4	5	6	7	8	9	10	11	12
Li	Be								
Lithium	Beryllium								

11	12	13	14	15	16	17	18	19	20
Mg									
Magnesium									

Ca	Calcium	38	87.62
K	Potassium	37	85.468
Rb			
Sr			

87	Fr	Ra	La	Rf	Ha	Sg	Ns	Hs	Mt	Ds	Rg
(223)	Francium	Radium	Lawren- cium	Ruther- fordium	Hahnium	Seabor- gium	Niobium	Hassium	Mel- nerium	Darm- stadtium	Roen- tgenium
88											
89											
90											
91											
92											
93											
94											
95											
96											
97											
98											
99											
100											
101											
102											
103											
104											
105											
106											
107											
108											
109											
110											
111											
112											
113											
114											
115											
116											
117											
118											
119											
120											
121											
122											
123											
124											
125											
126											
127											
128											
129											
130											
131											
132											
133											
134											
135											
136											
137											
138											
139											
140											
141											
142											
143											
144											
145											
146											
147											
148											
149											
150											
151											
152											
153											
154											
155											
156											
157											
158											
159											
160											
161											
162											
163											
164											
165											
166											
167											
168											
169											
170											
171											
172											
173											
174											
175											
176											
177											
178											
179											
180											
181											
182											
183											
184											
185											
186											
187											
188											
189											
190											
191											
192											
193											
194											
195											
196											
197											
198											
199											
200											
201											
202											
203											
204											
205											
206											
207											
208											
209											
210											
211											
212											
213											
214											
215											
216											
217											
218											
219											
220											
221											
222											
223											
224											
225											
226											
227											
228											
229											
230											
231											
232											
233											
234											
235											
236											
237											
238											
239											
240											
241											
242											
243											
244											
245											
246											
247											
248											
249											
250											
251											
252											
253											
254											
255											
256											
257											
258											
259											
260											
261											
262											
263											
264											
265											
266											
267											
268											
269											
270											
271											
272											

Chemistry can be a daunting subject for the uninitiated and, all too often, introductory textbooks do little to make students feel at ease with the complex subject matter. **Basic Chemistry Concepts and Exercises** brings the wisdom of John Kenkel's more than 30 years of teaching experience to communicate the fundamentals of chemistry in a practical, down-to-earth manner.

Using conversational language and logically assembled graphics, the book concisely introduces each topic without overwhelming students with unnecessary detail. Example problems and end-of-chapter questions emphasize repetition of concepts, preparing students to become adept at the basics before they progress to an advanced general chemistry course. Enhanced with visualization techniques such as the first chapter's *mythical microscope*, the book clarifies challenging, abstract ideas and stimulates curiosity into what can otherwise be an overwhelming topic.

Topics discussed in this reader-friendly text include

- Properties and structure of matter
- Atoms, molecules, and compounds
- The Periodic Table
- Atomic weight, formula weights, and moles
- Gases and solutions
- Chemical equilibrium
- Acids, bases, and pH
- Organic chemicals

The appendix contains answers to the homework exercises so students can check their work and receive instant feedback as to whether they have adequately grasped the concepts before moving on to the next section. In addition, the book includes access to interactive online homework exercises from *Sapling Learning*. Designed to help students embrace chemistry not with trepidation, but with confidence, this solid preparatory text forms a firm foundation for more advanced chemistry training.



CRC Press
Taylor & Francis Group
an informa business

www.crcpress.com

6000 Broken Sound Parkway, NW
Suite 300, Boca Raton, FL 33487
270 Madison Avenue
New York, NY 10016
2 Park Square, Milton Park
Abingdon, Oxon OX14 4RN, UK

K10685

ISBN: 978-1-4398-1337-9
90000

9 781439 813379