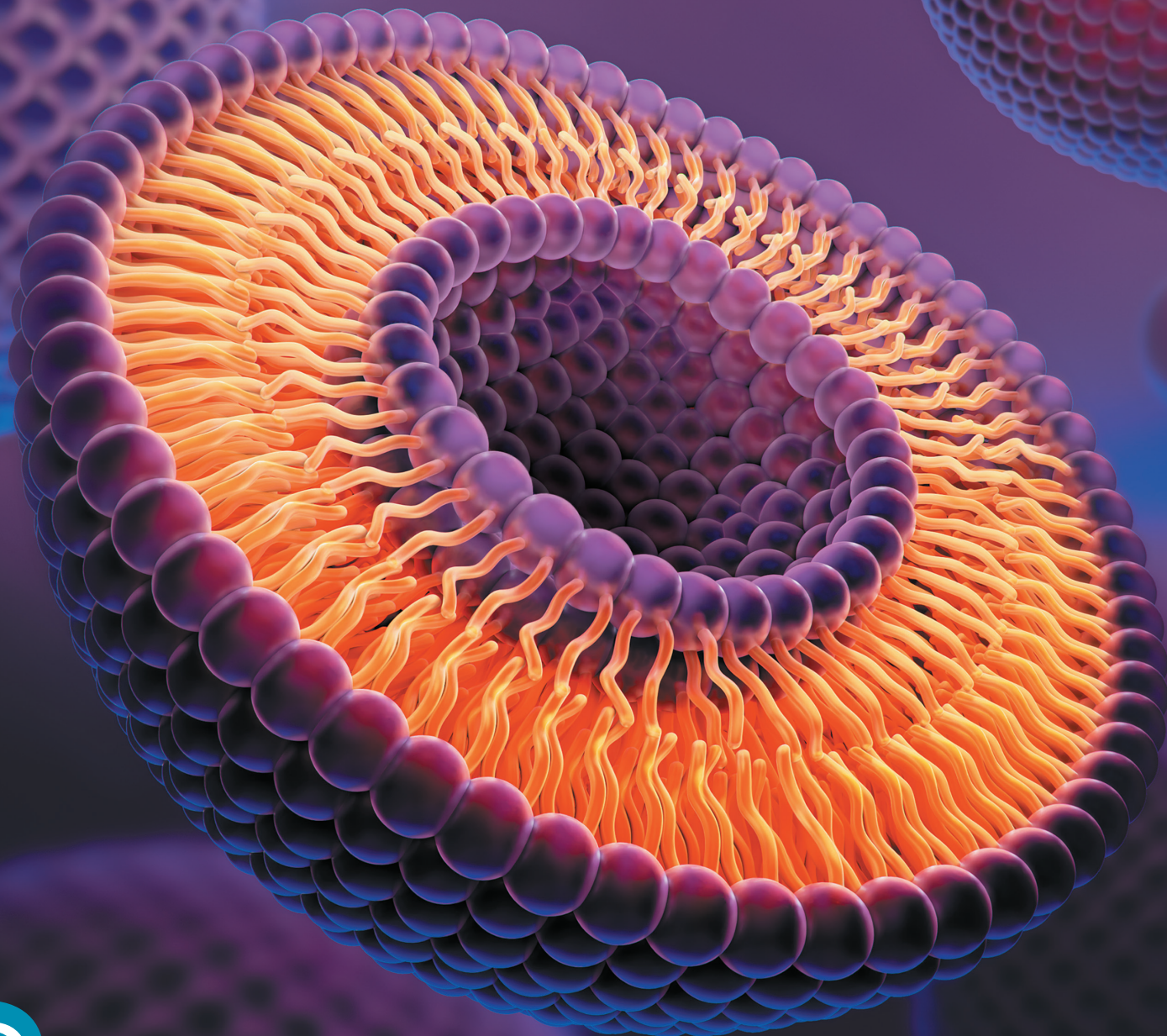
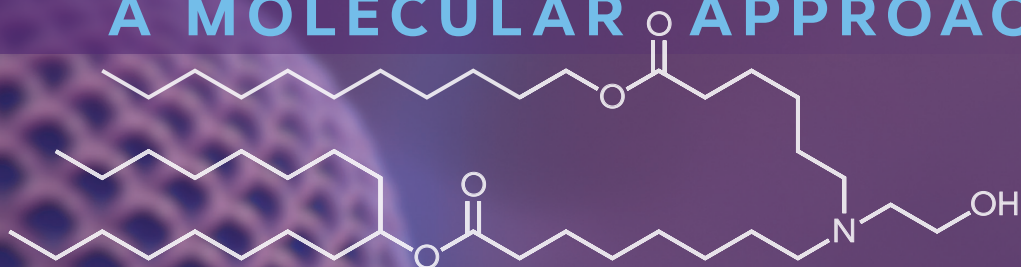


SIXTH EDITION

# CHEMISTRY

A MOLECULAR APPROACH



NIVALDO J. TRO

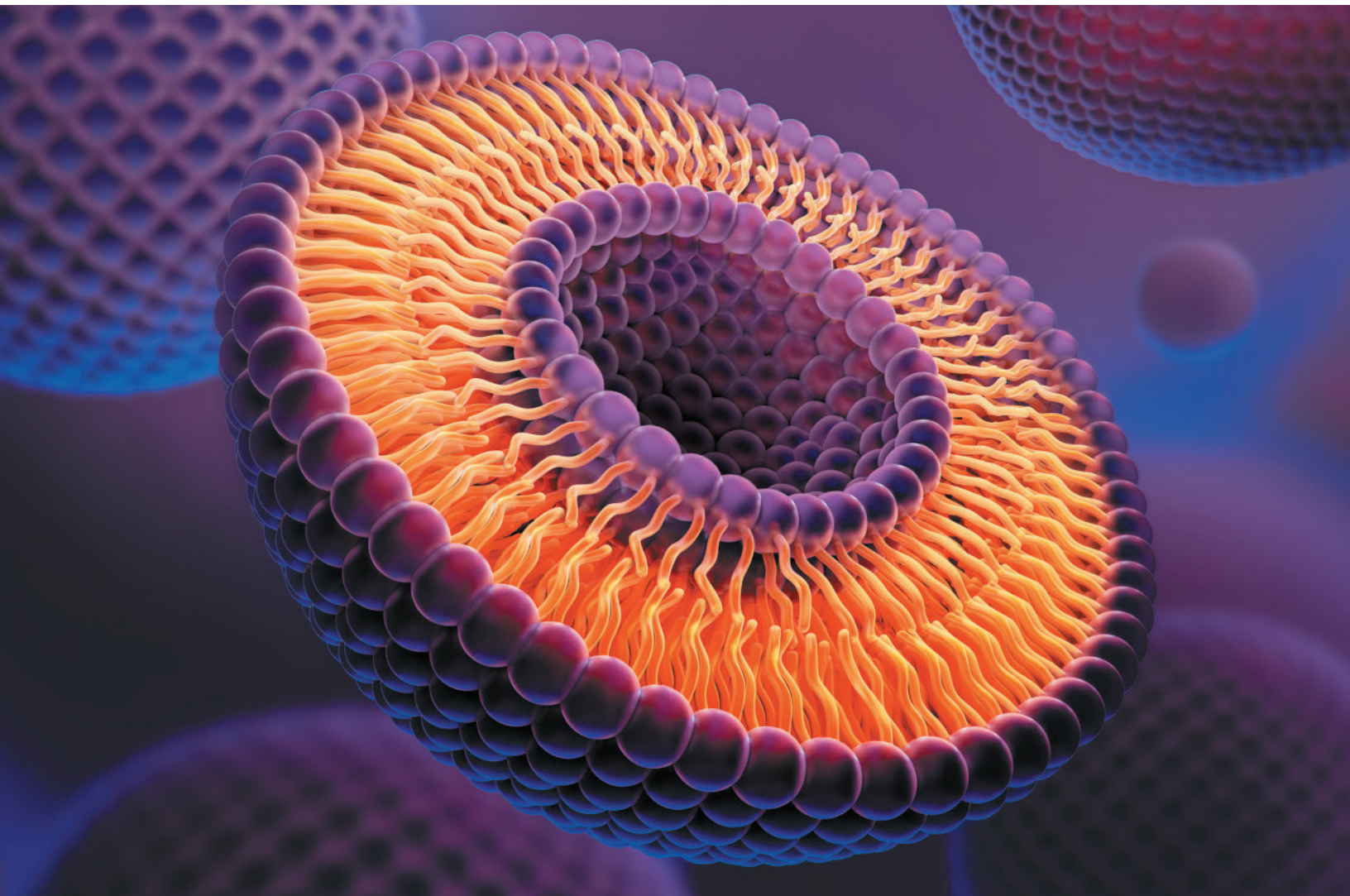
This page intentionally left blank



SIXTH EDITION

# CHEMISTRY

A MOLECULAR APPROACH



NIVALDO J. TRO



Content Development: Coleen Morrison  
Content Management: Jeanne Zalesky, Elizabeth Bell  
Content Production: Kristen Flathman, Shercian Kinoshian, Francesca Monaco, Straive, Courtney Coffman,  
Lachina Creative Inc., Tod Regan, Chloe Veylit, Jackie Jakob, Katie Foley, Jayne Sportelli, Margaret  
Trombley, Micah Brenner, Jerilyn DiCarlo, Mark Ong, Maria Guglielmo  
Product Management: Christopher Hess, Ian Desrosiers  
Product Marketing: Candice Madden, Rosemary Morton  
Rights and Permissions: Ben Ferrini, Angelica Aranas, Zosé Sotheo M. Herrero

Please contact <https://support.pearson.com/getsupport/s/> with any queries on this content

Cover Image by Lachina Creative, Inc.

Copyright © 2023, 2020, 2017 by Pearson Education, Inc. or its affiliates, 221 River Street, Hoboken, NJ 07030. All Rights Reserved. Manufactured in the United States of America. This publication is protected by copyright, and permission should be obtained from the publisher prior to any prohibited reproduction, storage in a retrieval system, or transmission in any form or by any means, electronic, mechanical, photocopying, recording, or otherwise. For information regarding permissions, request forms, and the appropriate contacts within the Pearson Education Global Rights and Permissions department, please visit [www.pearsoned.com/permissions/](http://www.pearsoned.com/permissions/).

Acknowledgments of third-party content appear on the appropriate page within the text C-1.

PEARSON and Mastering® Chemistry are exclusive trademarks owned by Pearson Education, Inc. or its affiliates in the U.S. and/or other countries.

Unless otherwise indicated herein, any third-party trademarks, logos, or icons that may appear in this work are the property of their respective owners, and any references to third-party trademarks, logos, icons, or other trade dress are for demonstrative or descriptive purposes only. Such references are not intended to imply any sponsorship, endorsement, authorization, or promotion of Pearson's products by the owners of such marks, or any relationship between the owner and Pearson Education, Inc., or its affiliates, authors, licensees, or distributors.

### **Library of Congress Cataloging-in-Publication Data**

Names: Tro, Nivaldo J., author.

Title: Chemistry: a molecular approach / Nivaldo J. Tro.

Description: Sixth edition. | Hoboken : Pearson, [2022] | Includes index.

Identifiers: LCCN 2021048850 (print) | LCCN 2021048851 (ebook) | ISBN

9780137493616 (hardcover) | ISBN 9780137832217 (ebook)

Subjects: LCSH: Chemistry, Physical and theoretical—Textbooks.

Classification: LCC QD453.3 .T759 2022 (print) | LCC QD453.3 (ebook) |

DDC 540—dc23/eng/20211110

LC record available at <https://lcn.loc.gov/2021048850>

LC ebook record available at <https://lcn.loc.gov/2021048851>

ScoutAutomatedPrintCode



Access Code Card

ISBN-10: 0-137-83196-X

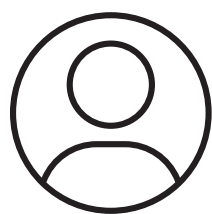
ISBN-13: 978-0-137-83196-8

Rental

ISBN-10: 0-137-49361-4

ISBN-13: 978-0-137-49361-6





# Pearson's Commitment to Diversity, Equity, and Inclusion

**Pearson is dedicated to creating bias-free content that reflects the diversity, depth, and breadth of all learners' lived experiences.**

We embrace the many dimensions of diversity, including but not limited to race, ethnicity, gender, sex, sexual orientation, socioeconomic status, ability, age, and religious or political beliefs.

Education is a powerful force for equity and change in our world. It has the potential to deliver opportunities that improve lives and enable economic mobility. As we work with authors to create content for every product and service, we acknowledge our responsibility to demonstrate inclusivity and incorporate diverse scholarship so that everyone can achieve their potential through learning. As the world's leading learning company, we have a duty to help drive change and live up to our purpose to help more people create a better life for themselves and to create a better world.

## **Our ambition is to purposefully contribute to a world where:**

- Everyone has an equitable and lifelong opportunity to succeed through learning.
- Our educational content accurately reflects the histories and lived experiences of the learners we serve.
- Our educational products and services are inclusive and represent the rich diversity of learners.
- Our educational content prompts deeper discussions with students and motivates them to expand their own learning (and worldview).

## **Accessibility**

We are also committed to providing products that are fully accessible to all learners. As per Pearson's guidelines for accessible educational Web media, we test and retest the capabilities of our products against the highest standards for every release, following the WCAG guidelines in developing new products for copyright year 2022 and beyond.



You can learn more about Pearson's commitment to accessibility at

<https://www.pearson.com/us/accessibility.html>

## **Contact Us**

While we work hard to present unbiased, fully accessible content, we want to hear from you about any concerns or needs with this Pearson product so that we can investigate and address them.



Please contact us with concerns about any potential bias at

<https://www.pearson.com/report-bias.html>



For accessibility-related issues, such as using assistive technology with Pearson products, alternative text requests, or accessibility documentation, email the Pearson Disability Support team at [disability.support@pearson.com](mailto:disability.support@pearson.com)



## About the Author

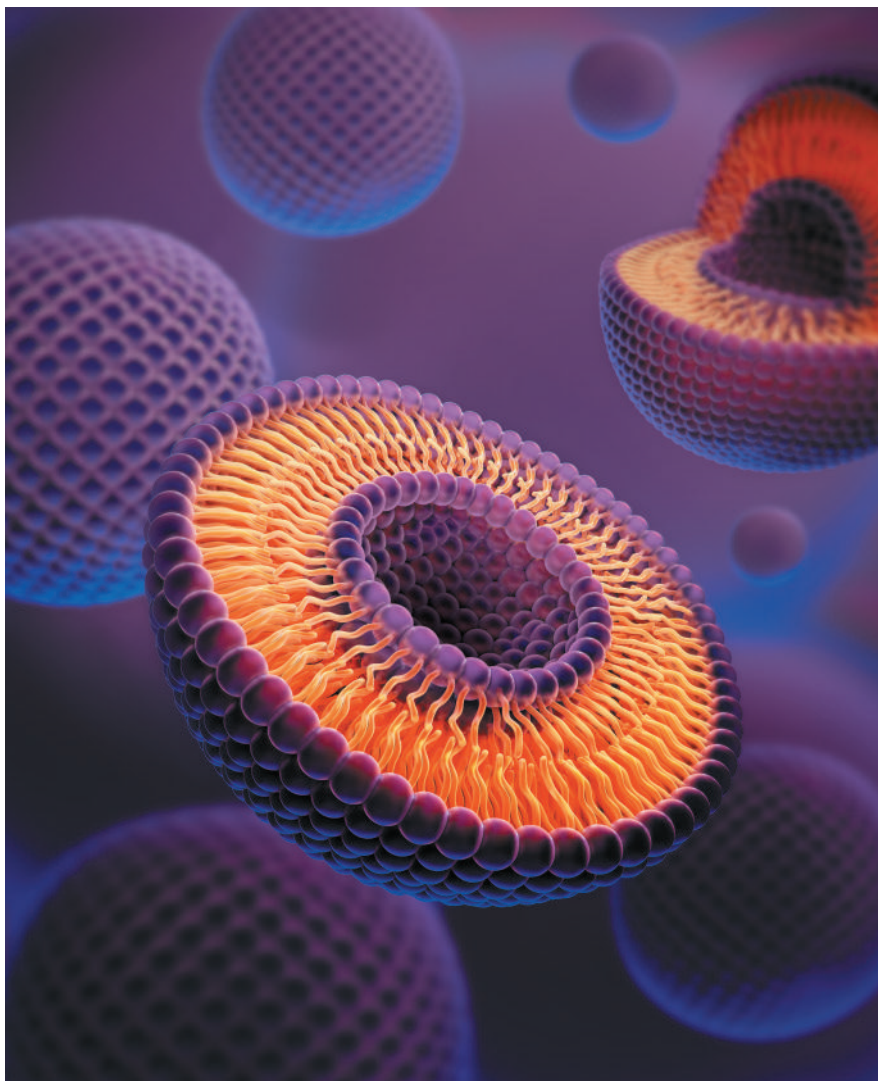


**Nivaldo Tro** has been teaching college chemistry since 1990 and is currently teaching at the College of Creative Studies at the University of California, Santa Barbara, and at Santa Barbara City College. He received his Ph.D. in chemistry from Stanford University for work on developing and using optical techniques to study the adsorption and desorption of molecules to and from surfaces in ultrahigh vacuum. He then went on to the University of California at Berkeley, where he did postdoctoral research on ultrafast reaction dynamics in solution. Professor Tro has been awarded grants from the American Chemical Society Petroleum Research Fund, the Research Corporation, and the National Science Foundation to study the dynamics of various processes occurring in thin adlayer films adsorbed on dielectric surfaces. Professor Tro lives in Santa Barbara with his wife, Ann. In his leisure time, Professor Tro enjoys cycling, surfing, and being outdoors.

*To Michael, Ali,  
Kyle, and Kaden*



## About the Cover



The cover shows a lipid nanoparticle (LNP), a fat bubble that became central to the mRNA vaccines developed against COVID-19. These vaccines use lipid nanoparticles to house mRNA and transport it to cells. The molecule shown is 8-[(2-hydroxyethyl)[6-oxo-6-(undecyloxy)hexyl]amino]-octanoic acid, 1-octylnonyl ester, one of the components of the LNPs used in the Moderna vaccine. This molecule, and others like it, have a polar end and a nonpolar end, which results in the formation of the bilayers that compose the LNP shell.

# Brief Contents

<b>1</b>	Matter, Measurement, and Problem Solving	<b>1</b>
<b>2</b>	Atoms and Elements	<b>48</b>
<b>3</b>	Molecules and Compounds	<b>90</b>
<b>4</b>	Chemical Reactions and Chemical Quantities	<b>138</b>
<b>5</b>	Introduction to Solutions and Aqueous Reactions	<b>168</b>
<b>6</b>	Gases	<b>212</b>
<b>7</b>	Thermochemistry	<b>266</b>
<b>8</b>	The Quantum-Mechanical Model of the Atom	<b>316</b>
<b>9</b>	Periodic Properties of the Elements	<b>358</b>
<b>10</b>	Chemical Bonding I: The Lewis Model	<b>400</b>
<b>11</b>	Chemical Bonding II: Molecular Shapes, Valence Bond Theory, and Molecular Orbital Theory	<b>444</b>
<b>12</b>	Liquids, Solids, and Intermolecular Forces	<b>502</b>
<b>13</b>	Solids and Modern Materials	<b>550</b>
<b>14</b>	Solutions	<b>588</b>
<b>15</b>	Chemical Kinetics	<b>642</b>
<b>16</b>	Chemical Equilibrium	<b>696</b>
<b>17</b>	Acids and Bases	<b>746</b>
<b>18</b>	Aqueous Ionic Equilibrium	<b>802</b>
<b>19</b>	Free Energy and Thermodynamics	<b>864</b>
<b>20</b>	Electrochemistry	<b>914</b>
<b>21</b>	Radioactivity and Nuclear Chemistry	<b>966</b>
<b>22</b>	Organic Chemistry	<b>1008</b>
<b>23</b>	Biochemistry	<b>1058</b>
<b>24</b>	Chemistry of the Nonmetals	<b>1092</b>
<b>25</b>	Metals and Metallurgy	<b>1130</b>
<b>26</b>	Transition Metals and Coordination Compounds	<b>1156</b>
<b>Appendix I</b>	Common Mathematical Operations in Chemistry	<b>A-1</b>
<b>Appendix II</b>	Useful Data	<b>A-5</b>
<b>Appendix III</b>	Answers to Selected Exercises	<b>A-15</b>
<b>Appendix IV</b>	Answers to In-Chapter Practice Problems	<b>A-55</b>
	<b>Glossary</b>	<b>G-1</b>
	<b>Photo and Text Credits</b>	<b>C-1</b>
	<b>Index</b>	<b>I-1</b>





## Interactive eText Media Contents

### KEY CONCEPT VIDEOS (KCVs)

- 1.1 Atoms and Molecules
- 1.3 Classifying Matter
- 1.6 Units and Significant Figures
- 1.7 Significant Figures in Calculations
- 1.8 Solving Chemical Problems
- 2.3 Atomic Theory
- 2.6 Subatomic Particles and Isotope Symbols
- 2.7 The Periodic Law and the Periodic Table
- 2.9 The Mole Concept
- 3.5 Naming Ionic Compounds
- 3.6 Naming Molecular Compounds
- 4.2 Writing and Balancing Chemical Equations
- 4.3 Reaction Stoichiometry
- 4.4 Limiting Reactant, Theoretical Yield, and Percent Yield
- 5.2 Solution Concentration
- 5.5 Reactions in Solutions
- 6.3 Simple Gas Laws and Ideal Gas Law
- 6.6 Mixtures of Gases and Partial Pressures
- 6.8 Kinetic Molecular Theory
- 7.3 The First Law of Thermodynamics
- 7.4 Heat Capacity
- 7.6 The Change in Enthalpy for a Chemical Reaction
- 7.9 Determining the Enthalpy of Reaction from Standard Enthalpies of Formation
- 8.2 The Nature of Light
- 8.4 The Wave Nature of Matter
- 8.5A Quantum Mechanics and the Atom: Orbitals and Quantum Numbers
- 8.5B Atomic Spectroscopy
- 9.3 Electron Configurations
- 9.4 Writing an Electron Configuration Based on an Element's Position on the Periodic Table
- 9.6 Periodic Trends in the Size of Atoms and Effective Nuclear Charge
- 10.5 The Lewis Model for Chemical Bonding
- 10.6 Electronegativity and Bond Polarity
- 10.7 Writing Lewis Structures for Molecular Compounds
- 10.8 Resonance and Formal Charge
- 10.9 Exceptions to the Octet Rule and Expanded Octets
- 11.2 VSEPR Theory
- 11.3 VSEPR Theory: The Effect of Lone Pairs
- 11.5 Molecular Shape and Polarity
- 11.6 Valence Bond Theory
- 11.7 Valence Bond Theory: Hybridization
- 12.3 Intermolecular Forces
- 12.5 Vaporization and Vapor Pressure
- 12.7 Heating Curve for Water
- 12.8 Phase Diagrams
- 13.3 Unit Cells: Simple Cubic, Body-Centered Cubic, and Face-Centered Cubic
- 14.4 Solution Equilibrium and the Factors Affecting Solubility
- 14.5 Solution Concentration: Molarity, Molality, Parts by Mass and Volume, Mole Fraction
- 14.6 Colligative Properties
- 15.2 The Rate of a Chemical Reaction
- 15.3 The Rate Law for a Chemical Reaction
- 15.4 The Integrated Rate Law
- 15.5 The Effect of Temperature on Reaction Rate
- 15.6 Reaction Mechanisms
- 16.3 The Equilibrium Constant
- 16.7 The Reaction Quotient
- 16.8 Finding Equilibrium Concentrations from Initial Concentrations
- 16.9 Le Châtelier's Principle
- 17.3 Definitions of Acids and Bases
- 17.4 Acid Strength and the Acid Ionization Constant
- 17.5 The pH Scale
- 17.6 Finding the  $[\text{H}_3\text{O}^+]$  and pH of Strong and Weak Acid Solutions
- 17.8 The Acid-Base Properties of Ions and Salts
- 18.2A Buffers
- 18.2B Finding pH and pH Changes in Buffer Solutions
- 18.4A The Titration of a Strong Acid with a Strong Base
- 18.4B The Titration of a Weak Acid and a Strong Base
- 19.3 Entropy and the Second Law of Thermodynamics
- 19.6 The Effect of  $\Delta H$ ,  $\Delta S$ , and  $T$  on Reaction Spontaneity
- 19.7 Standard Molar Entropies
- 20.3 Voltaic Cells
- 20.4 Standard Electrode Potentials
- 20.5 Cell Potential, Free Energy, and the Equilibrium Constant
- 21.3 Types of Radioactivity

## INTERACTIVE WORKED EXAMPLES (IWEs)

- 1.5 Determining the Number of Significant Figures in a Number
- 1.6 Significant Figures in Calculations
- 1.8 Unit Conversion
- 1.9 Unit Conversions Involving Units Raised to a Power
- 1.10 Density as a Conversion Factor
- 1.12 Problems with Equations
- 2.3 Atomic Numbers, Mass Numbers, and Isotope Symbols
- 2.5 Atomic Mass
- 2.8 The Mole Concept—Converting between Mass and Number of Atoms
- 2.9 The Mole Concept
- 3.3 Writing Formulas for Ionic Compounds
- 3.11 Using the Nomenclature Flowchart to Name Compounds
- 3.13 The Mole Concept—Converting between Mass and Number of Molecules
- 3.15 Using Mass Percent Composition as a Conversion Factor
- 3.16 Chemical Formulas as Conversion Factors
- 3.18 Obtaining an Empirical Formula from Experimental Data
- 3.21 Determining an Empirical Formula from Combustion Analysis
- 4.2 Balancing Chemical Equations
- 4.3 Balancing Chemical Equations Containing a Polyatomic Ion
- 4.4 Stoichiometry
- 4.6 Limiting Reactant and Theoretical Yield
- 5.1 Calculating Solution Concentration
- 5.2 Using Molarity in Calculations
- 5.3 Solution Dilution
- 5.4 Solution Stoichiometry
- 5.5 Predicting Whether an Ionic Compound Is Soluble
- 5.6 Writing Equations for Precipitation Reactions
- 5.9 Writing Equations for Acid–Base Reactions Involving a Strong Acid
- 5.11 Acid–Base Titration
- 5.13 Assigning Oxidation States
- 6.5 Ideal Gas Law I
- 6.7 Density
- 6.8 Molar Mass of a Gas
- 6.10 Partial Pressures and Mole Fractions
- 6.11 Collecting Gases over Water
- 6.12 Gases in Chemical Reactions
- 6.15 Graham's Law of Effusion
- 7.2 Temperature Changes and Heat Capacity
- 7.3 Thermal Energy Transfer
- 7.5 Measuring  $\Delta E_{\text{rxn}}$  in a Bomb Calorimeter
- 7.7 Stoichiometry Involving  $\Delta H$
- 7.8 Measuring  $\Delta H_{\text{rxn}}$  in a Coffee-Cup Calorimeter
- 7.9 Hess's Law
- 7.11  $\Delta H_{\text{rxn}}^\circ$  and Standard Enthalpies of Formation
- 8.2 Photon Energy
- 8.3 Wavelength, Energy, and Frequency
- 8.5 Quantum Numbers I
- 8.7 Wavelength of Light for a Transition in the Hydrogen Atom
- 9.2 Writing Orbital Diagrams
- 9.4 Writing Electron Configurations from the Periodic Table
- 9.5 Atomic Size
- 9.6 Electron Configurations and Magnetic Properties for Ions
- 9.8 First Ionization Energy
- 10.4 Writing Lewis Structures
- 10.6 Writing Lewis Structures for Polyatomic Ions
- 10.7 Writing Resonance Structures
- 10.8 Assigning Formal Charges
- 10.9 Drawing Resonance Structures and Assigning Formal Charge for Organic Compounds
- 10.10 Writing Lewis Structures for Compounds Having Expanded Octets
- 10.11 Calculating  $\Delta H_{\text{rxn}}$  from Bond Energies
- 11.1 VSEPR Theory and the Basic Shapes
- 11.2 Predicting Molecular Geometries
- 11.4 Predicting the Shape of Larger Molecules
- 11.5 Determining Whether a Molecule Is Polar
- 11.8 Hybridization and Bonding Scheme
- 11.10 Molecular Orbital Theory
- 12.1 Dipole–Dipole Forces
- 12.2 Hydrogen Bonding
- 12.3 Using the Heat of Vaporization in Calculations
- 12.5 Using the Two-Point Form of the Clausius–Clapeyron Equation to Predict the Vapor Pressure at a Given Temperature
- 13.3 Relating Unit Cell Volume, Edge Length, and Atomic Radius
- 13.4 Relating Density to Crystal Structure
- 14.2 Henry's Law
- 14.3 Using Parts by Mass in Calculations
- 14.4 Calculating Concentrations
- 14.5 Converting between Concentration Units
- 14.6 Calculating the Vapor Pressure of a Solution Containing a Nonelectrolyte and Nonvolatile Solute
- 14.9 Boiling Point Elevation
- 14.12 Calculating the Vapor Pressure of a Solution Containing an Ionic Solute
- 15.1 Expressing Reaction Rates
- 15.2 Determining the Order and Rate Constant of a Reaction
- 15.4 The First-Order Integrated Rate Law: Determining the Concentration of a Reactant at a Given Time
- 15.8 Using the Two-Point Form of the Arrhenius Equation
- 15.9 Reaction Mechanisms
- 16.1 Expressing Equilibrium Constants for Chemical Equations
- 16.3 Relating  $K_p$  and  $K_c$
- 16.5 Finding Equilibrium Constants from Experimental Concentration Measurements
- 16.7 Predicting the Direction of a Reaction by Comparing  $Q$  and  $K$



- 16.8 Finding Equilibrium Concentrations When You Know the Equilibrium Constant and All but One of the Equilibrium Concentrations of the Reactants and Products
- 16.9 Finding Equilibrium Concentrations from Initial Concentrations and the Equilibrium Constant
- 16.12 Finding Equilibrium Concentrations from Initial Concentrations in Cases with a Small Equilibrium Constant
- 16.14 The Effect of a Concentration Change on Equilibrium
- 17.1 Identifying Brønsted–Lowry Acids and Bases and Their Conjugates
- 17.3 Calculating pH from  $[H_3O^+]$  or  $[OH^-]$
- 17.5 Finding the  $[H_3O^+]$  of a Weak Acid Solution
- 17.7 Finding the pH of a Weak Acid Solution in Cases Where the  $x$  is *small* Approximation Does Not Work
- 17.8 Finding the Equilibrium Constant from pH
- 17.9 Finding the Percent Ionization of a Weak Acid
- 17.12 Finding the  $[OH^-]$  and pH of a Weak Base Solution
- 17.14 Determining the pH of a Solution Containing an Anion Acting as a Base
- 17.16 Determining the Overall Acidity or Basicity of Salt Solutions
- 18.2 Calculating the pH of a Buffer Solution as an Equilibrium Problem and with the Henderson–Hasselbalch Equation
- 18.3 Calculating the pH Change in a Buffer Solution after the Addition of a Small Amount of Strong Acid or Base
- 18.4 Using the Henderson–Hasselbalch Equation to Calculate the pH of a Buffer Solution Composed of a Weak Base and Its Conjugate Acid
- 18.6 Strong Acid–Strong Base Titration pH Curve
- 18.7 Weak Acid–Strong Base Titration pH Curve
- 18.8 Calculating Molar Solubility from  $K_{sp}$
- 18.12 Predicting Precipitation Reactions by Comparing  $Q$  and  $K_{sp}$
- 19.2 Calculating  $\Delta S$  for a Change of State
- 19.3 Calculating Entropy Changes in the Surroundings
- 19.4 Calculating Gibbs Free Energy Changes and Predicting Spontaneity from  $\Delta H$  and  $\Delta S$
- 19.5 Calculating Standard Entropy Changes ( $\Delta S_{rxn}^\circ$ )
- 19.6 Calculating the Standard Change in Free Energy for a Reaction Using  $\Delta G_{rxn}^\circ = \Delta H_{rxn}^\circ - T\Delta S_{rxn}^\circ$
- 19.10 Calculating  $\Delta G_{rxn}$  under Nonstandard Conditions
- 19.11 The Equilibrium Constant and  $\Delta G_{rxn}^\circ$
- 20.2 Half-Reaction Method of Balancing Aqueous Redox Equations in Acidic Solution
- 20.3 Balancing Redox Reactions Occurring in Basic Solution
- 20.4 Calculating Standard Potentials for Electrochemical Cells from Standard Electrode Potentials of the Half-Reactions
- 20.6 Relating  $\Delta G^\circ$  and  $E_{cell}^\circ$
- 21.1 Writing Nuclear Equations for Alpha Decay
- 21.2 Writing Nuclear Equations for Beta Decay, Positron Emission, and Electron Capture
- 21.4 Radioactive Decay Kinetics
- 21.5 Radiocarbon Dating
- 22.3 Naming Alkanes



## KEY CONCEPT INTERACTIVES (KCI's)

- 1.8A Unit Conversion
- 1.8B Problem Solving
- 2.6 Isotopes and Atomic Mass
- 2.9 The Mole Concept
- 3.5 Nomenclature
- 3.10 Determining a Chemical Formula from Experimental Data
- 4.2 Balancing Chemical Equations
- 4.4 Stoichiometry, Limiting Reactant, Excess Reactant, Theoretical Yield, and Percent Yield
- 5.4 Solubility Rules
- 5.5 Aqueous Reactions
- 6.4 Ideal Gas Law
- 6.6 Mixtures of Gases and Partial Pressures
- 7.5 Bomb Calorimetry
- 7.7 Coffee-Cup Calorimetry
- 7.9 Determining Enthalpies of Reaction from Standard Enthalpies of Formation
- 8.3 Atomic Spectroscopy
- 8.5 Quantum Numbers
- 9.4 Electron Configurations from the Periodic Table
- 9.6 Periodic Trends
- 10.7 Drawing Lewis Structures
- 10.8 Resonance and Formal Charge
- 11.2 VSEPR Theory
- 11.6 Valence Bond Theory
- 12.3 Intermolecular Forces
- 12.5 Vaporization
- 12.8 Phase Diagrams
- 13.3 Unit Cells
- 14.4 Factors Affecting Solubility
- 14.6 Colligative Properties
- 15.2 The Rate of a Chemical Reaction
- 15.3 The Rate Law
- 15.5 The Effect of Temperature on Reaction Rate
- 16.2 Dynamic Equilibrium and the Equilibrium Constant
- 16.7 The Reaction Quotient
- 16.8 Finding Equilibrium Concentrations
- 16.9 Le Châtelier's Principle
- 17.6 Finding  $[H_3O^+]$  in Strong and Weak Acid Solutions
- 17.8 The Acid–Base Properties of Ions and Salts
- 18.2 Buffers
- 18.4A The Titration of a Strong Acid with a Strong Base
- 18.4B The Titration of a Weak Acid with a Strong Base
- 19.6 Gibbs Free Energy Changes
- 19.7 Calculating the Change in Entropy for a Chemical Reaction from Standard Molar Entropies
- 20.3 Voltaic Cells
- 20.4 Standard Electrode Potentials and Cell Potential
- 21.3 Types of Radioactivity
- 21.6 Kinetics of Radioactive Decay
- 22.3 Alkanes, Alkenes, and Alkynes
- 22.8 Functional Groups
- 23.4 Proteins and Amino Acids

# Contents

PREFACE xxiii

## 1 Matter, Measurement, and Problem Solving 1

- 1.1 Atoms and Molecules 1
- 1.2 The Scientific Approach to Knowledge 3
  - THE NATURE OF SCIENCE** *Thomas S. Kuhn and Scientific Revolutions* 5
- 1.3 The Classification of Matter 5
  - The States of Matter: Solid, Liquid, and Gas 6
  - Classifying Matter by Composition: Elements, Compounds, and Mixtures 7
  - Separating Mixtures 8
- 1.4 Physical and Chemical Changes and Physical and Chemical Properties 9
- 1.5 Energy: A Fundamental Part of Physical and Chemical Change 12
- 1.6 The Units of Measurement 13
  - Standard Units 14
  - The Meter: A Measure of Length 14
  - The Kilogram: A Measure of Mass 14
  - The Second: A Measure of Time 14
  - The Kelvin: A Measure of Temperature 15
  - Prefix Multipliers 17
  - Derived Units: Volume and Density 17
  - Volume 18
  - Density 18
  - Calculating Density 19
  - CHEMISTRY AND MEDICINE** *Bone Density* 20
- 1.7 The Reliability of a Measurement 20
  - Counting Significant Figures 22
  - Exact Numbers 22
  - Significant Figures in Calculations 23
  - Precision and Accuracy 25
  - CHEMISTRY IN YOUR DAY** *Integrity in Data Gathering* 26
- 1.8 Solving Chemical Problems 26
  - Converting from One Unit to Another 26
  - General Problem-Solving Strategy 28
  - Units Raised to a Power 30
  - Order-of-Magnitude Estimations 31
  - Problems Involving an Equation 32
- 1.9 Analyzing and Interpreting Data 33
  - Identifying Patterns in Data 33
  - Interpreting Graphs 34

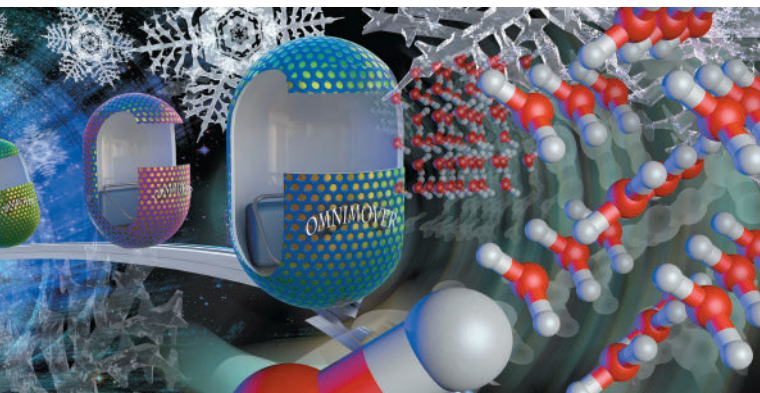
**CHAPTER IN REVIEW** Self-Assessment Quiz 36 Terms 37  
Concepts 38 Equations and Relationships 38 Learning Outcomes 38

**EXERCISES** Review Questions 39 Problems by Topic 39  
Cumulative Problems 44 Challenge Problems 45 Conceptual Problems 46 Questions for Group Work 46 Data Interpretation and Analysis 47 Answers to Conceptual Connections 47

## 2 Atoms and Elements 48



- 2.1 Brownian Motion: Atoms Confirmed 49
- 2.2 Early Ideas about the Building Blocks of Matter 51
- 2.3 Modern Atomic Theory and the Laws That Led to It 51
  - The Law of Conservation of Mass 51
  - The Law of Definite Proportions 52
  - The Law of Multiple Proportions 53
  - John Dalton and the Atomic Theory 54
  - CHEMISTRY IN YOUR DAY** *Atoms and Humans* 54
- 2.4 The Discovery of the Electron 55
  - Cathode Rays 55
  - Millikan's Oil Drop Experiment: The Charge of the Electron 56
- 2.5 The Structure of the Atom 57
- 2.6 Subatomic Particles: Protons, Neutrons, and Electrons in Atoms 59
  - Elements: Defined by Their Numbers of Protons 60
  - Isotopes: When the Number of Neutrons Varies 61
  - Ions: Losing and Gaining Electrons 63
  - CHEMISTRY IN YOUR DAY** *Where Did Elements Come From?* 64
- 2.7 Finding Patterns: The Periodic Law and the Periodic Table 65
  - Modern Periodic Table Organization 66
  - Ions and the Periodic Table 68
  - CHEMISTRY AND MEDICINE** *The Elements of Life* 69
- 2.8 Atomic Mass: The Average Mass of an Element's Atoms 69
  - Mass Spectrometry: Measuring the Mass of Atoms and Molecules 70
  - CHEMISTRY IN YOUR DAY** *Evolving Atomic Masses* 72





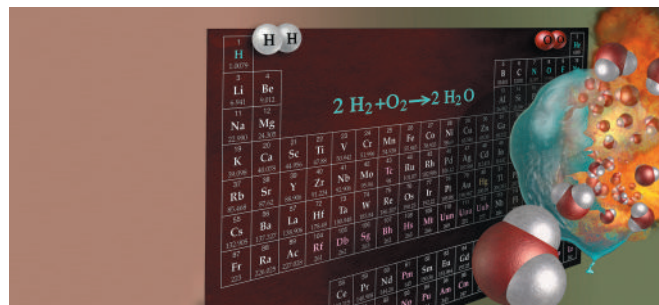
## 2.9 Molar Mass: Counting Atoms by Weighing Them 73

The Mole: A Chemist's "Dozen" 73 Converting between Number of Moles and Number of Atoms 74  
Converting between Mass and Amount (Number of Moles) 75

**CHAPTER IN REVIEW** Self-Assessment Quiz 78 Terms 79  
Concepts 80 Equations and Relationships 80 Learning Outcomes 81

**EXERCISES** Review Questions 81 Problems by Topic 82  
Cumulative Problems 85 Challenge Problems 86 Conceptual Problems 87 Questions for Group Work 88 Data Interpretation and Analysis 88 Answers to Conceptual Connections 89

## 3 Molecules and Compounds 90



### 3.1 Hydrogen, Oxygen, and Water 91

### 3.2 Chemical Bonds 93

Ionic Bonds 93 Covalent Bonds 94

### 3.3 Representing Compounds: Chemical Formulas and Molecular Models 94

Types of Chemical Formulas 94 Molecular Models 96

### 3.4 An Atomic-Level View of Elements and Compounds 96

### 3.5 Ionic Compounds: Formulas and Names 100

Writing Formulas for Ionic Compounds 100 Naming Ionic Compounds 101 Naming Binary Ionic Compounds Containing a Metal That Forms Only One Type of Cation 102 Naming Binary Ionic Compounds Containing a Metal That Forms More Than One Kind of Cation 103 Naming Ionic Compounds Containing Polyatomic Ions 104 Hydrated Ionic Compounds 105

### 3.6 Molecular Compounds: Formulas and Names 106

Naming Molecular Compounds 106 Naming Acids 107 Naming Binary Acids 108 Naming Oxyacids 108

**CHEMISTRY IN THE ENVIRONMENT** Acid Rain 108

### 3.7 Summary of Inorganic Nomenclature 109

### 3.8 Formula Mass and the Mole Concept for Compounds 111

Molar Mass of a Compound 111 Using Molar Mass to Count Molecules by Weighing 111

### 3.9 Composition of Compounds 113

Mass Percent Composition as a Conversion Factor 114  
Conversion Factors from Chemical Formulas 116

**CHEMISTRY AND MEDICINE** Methylmercury in Fish 118

## 3.10 Determining a Chemical Formula from Experimental Data 118

Determining Molecular Formulas for Compounds 120  
Combustion Analysis 121

## 3.11 Organic Compounds 123

Hydrocarbons 124 Functionalized Hydrocarbons 125

**CHAPTER IN REVIEW** Self-Assessment Quiz 127 Terms 128  
Concepts 128 Equations and Relationships 129 Learning Outcomes 129

**EXERCISES** Review Questions 129 Problems by Topic 130  
Cumulative Problems 134 Challenge Problems 135 Conceptual Problems 136 Questions for Group Work 136 Data Interpretation and Analysis 136 Answers to Conceptual Connections 137

## 4 Chemical Reactions and Chemical Quantities 138

### 4.1 Climate Change and the Combustion of Fossil Fuels 139

### 4.2 Writing and Balancing Chemical Equations 141

### 4.3 Reaction Stoichiometry: How Much Carbon Dioxide? 145

Making Pizza: The Relationships among Ingredients 145  
Making Molecules: Mole-to-Mole Conversions 146  
Making Molecules: Mass-to-Mass Conversions 146

### 4.4 Stoichiometric Relationships: Limiting Reactant, Theoretical Yield, Percent Yield, and Reactant in Excess 149

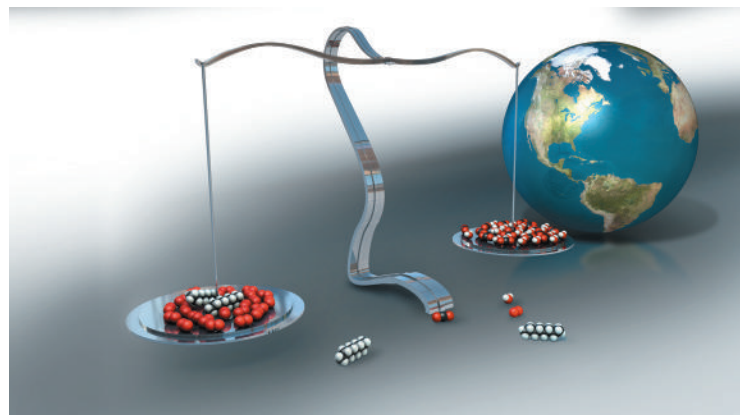
Calculating Limiting Reactant, Theoretical Yield, and Percent Yield 151 Calculating Limiting Reactant, Theoretical Yield, and Percent Yield from Initial Reactant Masses 152

### 4.5 Three Examples of Chemical Reactions: Combustion, Alkali Metals, and Halogens 155

Combustion Reactions 155 Alkali Metal Reactions 156 Halogen Reactions 156

**CHAPTER IN REVIEW** Self-Assessment Quiz 158 Terms 159  
Concepts 159 Equations and Relationships 159 Learning Outcomes 159

**EXERCISES** Review Questions 160 Problems by Topic 160  
Cumulative Problems 163 Challenge Problems 164  
Conceptual Problems 165 Questions for Group Work 165  
Data Interpretation and Analysis 166 Answers to Conceptual Connections 167



## 5 Introduction to Solutions and Aqueous Reactions 168



- 5.1 Molecular Gastronomy and the Spherified Cherry 169
- 5.2 Solution Concentration 170
  - Solution Concentration 170 Using Molarity in Calculations 172
  - Solution Dilution 173
- 5.3 Solution Stoichiometry 175
- 5.4 Types of Aqueous Solutions and Solubility 177
  - Electrolyte and Nonelectrolyte Solutions 177
  - The Solubility of Ionic Compounds 179
- 5.5 Precipitation Reactions 181
- 5.6 Representing Aqueous Reactions: Molecular, Ionic, and Net Ionic Equations 185
- 5.7 Acid–Base Reactions 187
  - Acid–Base Reactions 187
  - Acid–Base Titrations 191
- 5.8 Gas–Evolution Reactions 193
- 5.9 Oxidation–Reduction Reactions 195
  - Oxidation States 196
  - Identifying Redox Reactions 198
  - The Activity Series: Predicting Whether a Redox Reaction Is Spontaneous 200
- CHEMISTRY IN YOUR DAY** Bleached Blonde 201

**CHAPTER IN REVIEW** Self-Assessment Quiz 203 Terms 204 Concepts 204 Equations and Relationships 205 Learning Outcomes 205

**EXERCISES** Review Questions 206 Problems by Topic 206 Cumulative Problems 208 Challenge Problems 209 Conceptual Problems 210 Questions for Group Work 210 Data Interpretation and Analysis 210 Answers to Conceptual Connections 211

## 6 Gases 212

- 6.1 Supersonic Skydiving and the Risk of Decompression 213
- 6.2 Pressure: The Result of Molecular Collisions 214
  - Pressure Units 215
  - The Manometer: A Way to Measure Pressure in the Laboratory 216
- CHEMISTRY AND MEDICINE** Blood Pressure 217
- 6.3 The Simple Gas Laws: Boyle's Law, Charles's Law, and Avogadro's Law 217
  - Boyle's Law: Volume and Pressure 218
  - Charles's Law: Volume and Temperature 220
- CHEMISTRY IN YOUR DAY** Extra-Long Snorkels 221
- Avogadro's Law: Volume and Amount (in Moles) 223

- 6.4 The Ideal Gas Law 224
- 6.5 Applications of the Ideal Gas Law: Molar Volume, Density, and Molar Mass of a Gas 227
  - Molar Volume at Standard Temperature and Pressure 227
  - Density of a Gas 228
  - Molar Mass of a Gas 229
- 6.6 Mixtures of Gases and Partial Pressures 230
  - Deep-Sea Diving and Partial Pressures 233
  - Collecting Gases over Water 235
- 6.7 Gases in Chemical Reactions: Stoichiometry Revisited 237
  - Molar Volume and Stoichiometry 238
  - ANALYZING AND INTERPRETING DATA** Good News about Our Nation's Air Quality 240
- 6.8 Kinetic Molecular Theory: A Model for Gases 240
  - How Kinetic Molecular Theory Explains Pressure and the Simple Gas Laws 241
  - Kinetic Molecular Theory and the Ideal Gas Law 242
  - Temperature and Molecular Velocities 244
- 6.9 Mean Free Path, Diffusion, and Effusion of Gases 247
- 6.10 Real Gases: The Effects of Size and Intermolecular Forces 248
  - The Effect of the Finite Volume of Gas Particles 249
  - The Effect of Intermolecular Forces 250
  - Van der Waals Equation 251
  - Real Gases 251

**CHAPTER IN REVIEW** Self-Assessment Quiz 252 Terms 253 Concepts 253 Equations and Relationships 254 Learning Outcomes 254

**EXERCISES** Review Questions 255 Problems by Topic 255 Cumulative Problems 260 Challenge Problems 262 Conceptual Problems 262 Questions for Group Work 263 Data Interpretation and Analysis 263 Answers to Conceptual Connections 264



## 7 Thermochemistry 266

- 7.1 Chemical Hand Warmers 267
- 7.2 The Nature of Energy: Key Definitions 268
  - Types of Energy 268
  - Energy Conservation and Energy Transfer 269
  - Units of Energy 269
- 7.3 The First Law of Thermodynamics: There Is No Free Lunch 271
  - Internal Energy 271
  - CHEMISTRY IN YOUR DAY** Redheffer's Perpetual Motion Machine 271
  - Heat and Work 274
- 7.4 Quantifying Heat and Work 276
  - Heat 276
  - Temperature Changes and Heat Capacity 276
  - Thermal Energy Transfer 278
  - Work: Pressure–Volume Work 280



- 7.5 Measuring  $\Delta E$  for Chemical Reactions: Constant-Volume Calorimetry** 282
- 7.6 Enthalpy: The Heat Evolved in a Chemical Reaction at Constant Pressure** 285  
Exothermic and Endothermic Processes: A Molecular View 287 Stoichiometry Involving  $\Delta H$ : Thermochemical Equations 287
- 7.7 Constant-Pressure Calorimetry: Measuring  $\Delta H_{\text{rxn}}$**  289
- 7.8 Relationships Involving  $\Delta H_{\text{rxn}}$**  290
- 7.9 Determining Enthalpies of Reaction from Standard Enthalpies of Formation** 293  
Standard States and Standard Enthalpy Changes 293 Calculating the Standard Enthalpy Change for a Reaction 295
- 7.10 Energy Use and the Environment** 298  
Energy Consumption 298 Environmental Problems Associated with Fossil Fuel Use 299 Air Pollution 299 Global Climate Change 300

**CHEMISTRY IN THE ENVIRONMENT** Renewable Energy 302

**CHAPTER IN REVIEW** Self-Assessment Quiz 303 Terms 304 Concepts 304 Equations and Relationships 305 Learning Outcomes 305

**EXERCISES** Review Questions 306 Problems by Topic 306 Cumulative Problems 310 Challenge Problems 312 Conceptual Problems 312 Questions for Group Work 313 Data Interpretation and Analysis 313 Answers to Conceptual Connections 314

## **8 The Quantum-Mechanical Model of the Atom** 316

- 8.1 Schrödinger's Cat** 317
- 8.2 The Nature of Light** 318  
The Wave Nature of Light 319 The Electromagnetic Spectrum 321  
**CHEMISTRY AND MEDICINE** Radiation Treatment for Cancer 323  
Interference and Diffraction 323  
The Particle Nature of Light 324
- 8.3 Atomic Spectroscopy and the Bohr Model** 328  
**CHEMISTRY IN YOUR DAY** Atomic Spectroscopy, a Bar Code for Atoms 330
- 8.4 The Wave Nature of Matter: The de Broglie Wavelength, the Uncertainty Principle, and Indeterminacy** 331  
The de Broglie Wavelength 333 The Uncertainty Principle 334 Indeterminacy and Probability Distribution Maps 335
- 8.5 Quantum Mechanics and the Atom** 337  
Solutions to the Schrödinger Equation for the Hydrogen Atom 337 Atomic Spectroscopy Explained 340

- 8.6 The Shapes of Atomic Orbitals** 343  
s Orbitals ( $l = 0$ ) 343 p Orbitals ( $l = 1$ ) 346  
d Orbitals ( $l = 2$ ) 346 f Orbitals ( $l = 3$ ) 346 The Phase of Orbitals 347 The Shape of Atoms 348

**CHAPTER IN REVIEW** Self-Assessment Quiz 348 Terms 349 Concepts 349 Equations and Relationships 350 Learning Outcomes 350

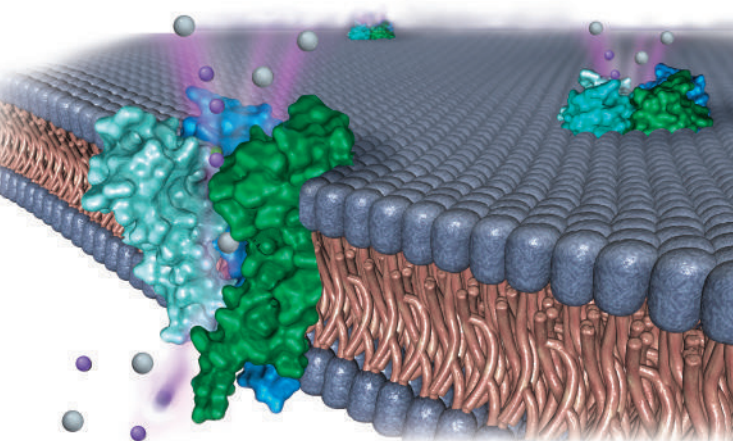
**EXERCISES** Review Questions 350 Problems by Topic 351 Cumulative Problems 353 Challenge Problems 355 Conceptual Problems 355 Questions for Group Work 356 Data Interpretation and Analysis 356 Answers to Conceptual Connections 357



## **9 Periodic Properties of the Elements** 358

- 9.1 Nerve Signal Transmission** 359
- 9.2 The Development of the Periodic Table** 360
- 9.3 Electron Configurations: How Electrons Occupy Orbitals** 361  
Electron Spin and the Pauli Exclusion Principle 362 Sublevel Energy Splitting in Multielectron Atoms 362 Coulomb's Law 363 Shielding 364 Penetration 364 Electron Spatial Distributions and Sublevel Splitting 364 Electron Configurations for Multielectron Atoms 366
- 9.4 Electron Configurations, Valence Electrons, and the Periodic Table** 369  
Orbital Blocks in the Periodic Table 370 Writing an Electron Configuration for an Element from Its Position in the Periodic Table 371 The Transition and Inner Transition Elements 372
- 9.5 The Explanatory Power of the Quantum-Mechanical Model** 373
- 9.6 Periodic Trends in the Size of Atoms and Effective Nuclear Charge** 374  
Effective Nuclear Charge 376 Atomic Radii and the Transition Elements 377
- 9.7 Ions: Electron Configurations, Magnetic Properties, Ionic Radii, and Ionization Energy** 379  
Electron Configurations and Magnetic Properties of Ions 379 Ionic Radii 381 Ionization Energy 383 Trends in First Ionization Energy 383 Exceptions to Trends in First Ionization Energy 385 Trends in Second and Successive Ionization Energies 386





## 9.8 Electron Affinities and Metallic Character 387

Electron Affinity 387 Metallic Character 388

## 9.9 Periodic Trends Summary 391

**CHAPTER IN REVIEW** Self-Assessment Quiz 391 Terms 392 Concepts 392 Equations and Relationships 393 Learning Outcomes 393

**EXERCISES** Review Questions 394 Problems by Topic 395 Cumulative Problems 396 Challenge Problems 397 Conceptual Problems 398 Questions for Group Work 398 Data Interpretation and Analysis 399 Answers to Conceptual Connections 399

# 10 Chemical Bonding I: The Lewis Model 400

- 10.1 Bonding Models and AIDS Drugs 401
- 10.2 Types of Chemical Bonds 402
- 10.3 Representing Valence Electrons with Dots 404
- 10.4 Ionic Bonding: Lewis Symbols and Lattice Energies 405
  - Ionic Bonding and Electron Transfer 405 Lattice Energy: The Rest of the Story 406 The Born–Haber Cycle 406 Trends in Lattice Energies: Ion Size 409 Trends in Lattice Energies: Ion Charge 409 Ionic Bonding: Models and Reality 410
- CHEMISTRY AND MEDICINE** Ionic Compounds in Medicine 411
- 10.5 Covalent Bonding: Lewis Structures 412
  - Single Covalent Bonds 412 Double and Triple Covalent Bonds 412 Covalent Bonding: Models and Reality 413
- 10.6 Electronegativity and Bond Polarity 414
  - Electronegativity 415 Bond Polarity, Dipole Moment, and Percent Ionic Character 416
- 10.7 Lewis Structures of Molecular Compounds and Polyatomic Ions 418
  - Writing Lewis Structures for Molecular Compounds 418 Writing Lewis Structures for Polyatomic Ions 420
- 10.8 Resonance and Formal Charge 420
  - Resonance 420 Formal Charge 422

## 10.9 Exceptions to the Octet Rule: Odd-Electron Species, Incomplete Octets, and Expanded Octets 425

Odd-Electron Species 426 Incomplete Octets 426

**CHEMISTRY IN THE ENVIRONMENT** Free Radicals and the Atmospheric Vacuum Cleaner 427

Expanded Octets 428

## 10.10 Bond Energies and Bond Lengths 430

Bond Energy 430 Using Average Bond Energies to

Estimate Enthalpy Changes for Reactions 431

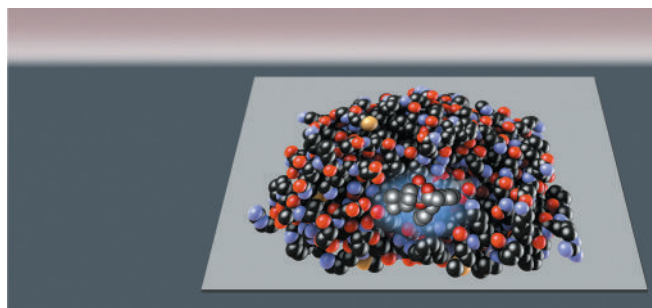
Bond Lengths 432

## 10.11 Bonding in Metals: The Electron Sea Model 433

**CHEMISTRY IN THE ENVIRONMENT** The Lewis Structure of Ozone 434

**CHAPTER IN REVIEW** Self-Assessment Quiz 435 Terms 436 Concepts 436 Equations and Relationships 437 Learning Outcomes 437

**EXERCISES** Review Questions 438 Problems by Topic 438 Cumulative Problems 440 Challenge Problems 442 Conceptual Problems 442 Questions for Group Work 442 Data Interpretation and Analysis 443 Answers to Conceptual Connections 443



# 11 Chemical Bonding II: Molecular Shapes, Valence Bond Theory, and Molecular Orbital Theory 444

## 11.1 Morphine: A Molecular Imposter 445

## 11.2 VSEPR Theory: The Five Basic Shapes 446

Two Electron Groups: Linear Geometry 447 Three Electron Groups: Trigonal Planar Geometry 447 Four Electron Groups: Tetrahedral Geometry 447 Five Electron Groups: Trigonal Bipyramidal Geometry 449 Six Electron Groups: Octahedral Geometry 449

## 11.3 VSEPR Theory: The Effect of Lone Pairs 450

Four Electron Groups with Lone Pairs 450 Five Electron Groups with Lone Pairs 452 Six Electron Groups with Lone Pairs 453

## 11.4 VSEPR Theory: Predicting Molecular Geometries 455

Representing Molecular Geometries on Paper 457 Predicting the Shapes of Larger Molecules 457



- 11.5 Molecular Shape and Polarity** 458  
 Vector Addition 460  
**CHEMISTRY IN YOUR DAY** *How Soap Works* 462
- 11.6 Valence Bond Theory: Orbital Overlap as a Chemical Bond** 463
- 11.7 Valence Bond Theory: Hybridization of Atomic Orbitals** 465  
 $sp^3$  Hybridization 466  $sp^2$  Hybridization and Double Bonds 468  
**CHEMISTRY IN YOUR DAY** *The Chemistry of Vision* 472  
 $sp$  Hybridization and Triple Bonds 472  
 $sp^3d$  and  $sp^3d^2$  Hybridization 474  
 Writing Hybridization and Bonding Schemes 475
- 11.8 Molecular Orbital Theory: Electron Delocalization** 478  
 Linear Combination of Atomic Orbitals (LCAOs) 479  
 Period Two Homonuclear Diatomic Molecules 483  
 Second-Period Heteronuclear Diatomic Molecules 488  
 Polyatomic Molecules 490
- CHAPTER IN REVIEW** Self-Assessment Quiz 491 Terms 492  
 Concepts 492 Equations and Relationships 492 Learning Outcomes 493
- EXERCISES** Review Questions 493 Problems by Topic 494  
 Cumulative Problems 497 Challenge Problems 499  
 Conceptual Problems 500 Questions for Group Work 500  
 Data Interpretation and Analysis 501 Answers to Conceptual Connections 501

## 12 Liquids, Solids, and Intermolecular Forces 502

- 12.1 Water, No Gravity** 503
- 12.2 Solids, Liquids, and Gases: A Molecular Comparison** 504  
 Differences between States of Matter 504  
 Changes between States 506
- 12.3 Intermolecular Forces: The Forces That Hold Condensed States Together** 507  
 Dispersion Force 508 Dipole–Dipole Force 510  
 Hydrogen Bonding 513 Ion–Dipole Force 515  
**CHEMISTRY AND MEDICINE** *Hydrogen Bonding in DNA* 516

- 12.4 Intermolecular Forces in Action: Surface Tension, Viscosity, and Capillary Action** 517  
 Surface Tension 517 Viscosity 519  
**CHEMISTRY IN YOUR DAY** *Viscosity and Motor Oil* 519  
 Capillary Action 519
- 12.5 Vaporization and Vapor Pressure** 520  
 The Process of Vaporization 520 The Energetics of Vaporization 522 Vapor Pressure and Dynamic Equilibrium 523 Temperature Dependence of Vapor Pressure and Boiling Point 525 The Clausius–Clapeyron Equation 526 The Critical Point: The Transition to an Unusual State of Matter 529
- 12.6 Sublimation and Fusion** 530  
 Sublimation 530 Fusion 531 Energetics of Melting and Freezing 531
- 12.7 Heating Curve for Water** 532
- 12.8 Phase Diagrams** 535  
 The Major Features of a Phase Diagram 535 Navigation within a Phase Diagram 536 The Phase Diagrams of Other Substances 537
- 12.9 Water: An Extraordinary Substance** 537  
**CHEMISTRY IN THE ENVIRONMENT** *Water Pollution* 539

**CHAPTER IN REVIEW** Self-Assessment Quiz 539 Terms 540  
 Concepts 540 Equations and Relationships 541 Learning Outcomes 541

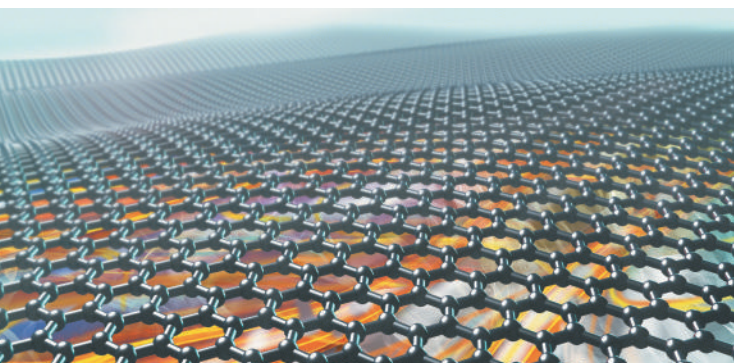
**EXERCISES** Review Questions 542 Problems by Topic 542  
 Cumulative Problems 545 Challenge Problems 546  
 Conceptual Problems 547 Questions for Group Work 547  
 Data Interpretation and Analysis 548 Answers to Conceptual Connections 549



## 13 Solids and Modern Materials 550

- 13.1 Friday Night Experiments: The Discovery of Graphene** 551
- 13.2 X-Ray Crystallography** 552
- 13.3 Unit Cells and Basic Structures** 555  
 Cubic Unit Cells 555 Closest-Packed Structures 561





### 13.4 The Fundamental Types of Crystalline Solids 562

Molecular Solids 563

**CHEMISTRY IN YOUR DAY** *Chocolate, An Edible Material* 564

Ionic Solids 565 Atomic Solids 565

### 13.5 The Structures of Ionic Solids 566

### 13.6 Network Covalent Atomic Solids:

**Carbon and Silicates** 568

Carbon 568 Silicates 571

### 13.7 Ceramics, Cement, and Glass 571

Ceramics 571 Cement 572 Glass 573

### 13.8 Semiconductors and Band Theory 573

Molecular Orbitals and Energy Bands 573 Doping: Controlling the Conductivity of Semiconductors 575

### 13.9 Polymers and Plastics 575

**CHEMISTRY IN YOUR DAY** *Kevlar* 578

**CHAPTER IN REVIEW** Self-Assessment Quiz 579 Terms 580 Concepts 580 Equations and Relationships 581 Learning Outcomes 581

**EXERCISES** Review Questions 581 Problems by Topic 582 Cumulative Problems 585 Challenge Problems 586 Conceptual Problems 586 Questions for Group Work 586 Data Interpretation and Analysis 587 Answers to Conceptual Connections 587

## 14 Solutions 588

### 14.1 Thirsty Solutions: Why You Shouldn't Drink Seawater 589

### 14.2 Types of Solutions and Solubility 591

Nature's Tendency toward Mixing: Entropy 592 The Effect of Intermolecular Forces 592

### 14.3 Energetics of Solution Formation 596

Energy Changes in Solution Formation 596 Aqueous Solutions and Heats of Hydration 598

### 14.4 Solution Equilibrium and Factors Affecting Solubility 599

The Temperature Dependence of the Solubility of Solids 601 Factors Affecting the Solubility of Gases in Water 601

### 14.5 Expressing Solution Concentration 604

**CHEMISTRY IN THE ENVIRONMENT** *Lake Nyos* 604

Molarity 605 Molality 606 Parts by Mass and Parts by Volume 606 Using Parts by Mass (or Parts by Volume) in Calculations 607 Mole Fraction and Mole Percent 608

**CHEMISTRY IN THE ENVIRONMENT** *The Dirty Dozen* 608

### 14.6 Colligative Properties: Vapor Pressure Lowering, Freezing Point Depression, Boiling Point Elevation, and Osmotic Pressure 611

Vapor Pressure Lowering 612 Vapor Pressures of Solutions Containing a Volatile (Nonelectrolyte) Solute 615 Freezing Point Depression and Boiling Point Elevation 618

**CHEMISTRY IN YOUR DAY** *Antifreeze in Frogs* 621 Osmotic Pressure 621

### 14.7 Colligative Properties of Strong Electrolyte Solutions 623

Strong Electrolytes and Vapor Pressure 624 Colligative Properties and Medical Solutions 625

### 14.8 Colloids 626

**CHAPTER IN REVIEW** Self-Assessment Quiz 629 Terms 630 Concepts 630 Equations and Relationships 631 Learning Outcomes 631

**EXERCISES** Review Questions 632 Problems by Topic 633 Cumulative Problems 637 Challenge Problems 638 Conceptual Problems 639 Questions for Group Work 639 Data Interpretation and Analysis 640 Answers to Conceptual Connections 641





## 15 Chemical Kinetics 642

- 15.1 Catching Lizards 643
- 15.2 The Rate of a Chemical Reaction 644
  - Definition of Reaction Rate 644 Measuring Reaction Rates 648
- 15.3 The Rate Law: The Effect of Concentration on Reaction Rate 649
  - The Three Common Reaction Orders ( $n = 0, 1$ , and  $2$ ) 649
  - Determining the Order of a Reaction 650 Reaction Order for Multiple Reactants 652
- 15.4 The Integrated Rate Law: The Dependence of Concentration on Time 654
  - The Integrated Rate Law 654 The Half-Life of a Reaction 658
- 15.5 The Effect of Temperature on Reaction Rate 662
  - The Arrhenius Equation 662 The Activation Energy, Frequency Factor, and Exponential Factor 663
  - Arrhenius Plots: Experimental Measurements of the Frequency Factor and the Activation Energy 664
  - The Collision Model: A Closer Look at the Frequency Factor 667
- 15.6 Reaction Mechanisms 668
  - Rate Laws for Elementary Steps 669 Rate-Determining Steps and Overall Reaction Rate Laws 670 Mechanisms with a Fast Initial Step 671
- 15.7 Catalysis 673
  - Homogeneous and Heterogeneous Catalysis 675 Enzymes: Biological Catalysts 676
  - CHEMISTRY AND MEDICINE** Enzyme Catalysis and the Role of Chymotrypsin in Digestion 678

**CHAPTER IN REVIEW** Self-Assessment Quiz 679 Terms 681 Concepts 681 Equations and Relationships 682 Learning Outcomes 682

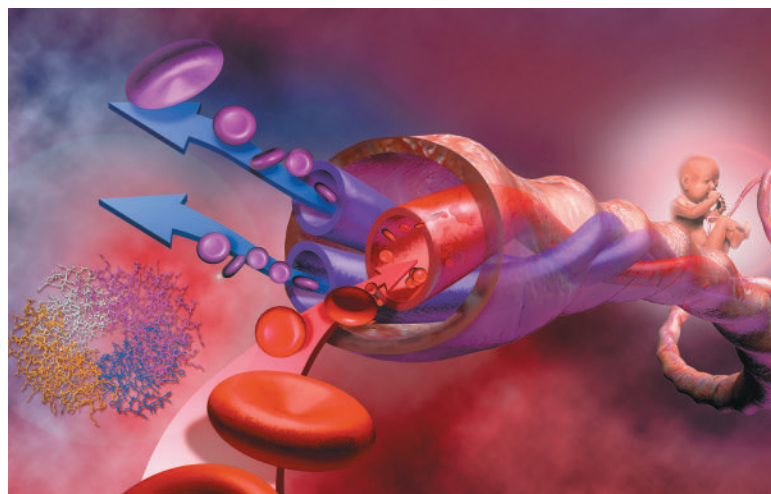
**EXERCISES** Review Questions 682 Problems by Topic 683 Cumulative Problems 690 Challenge Problems 692 Conceptual Problems 693 Questions for Group Work 694 Data Interpretation and Analysis 694 Answers to Conceptual Connections 695

## 16 Chemical Equilibrium 696

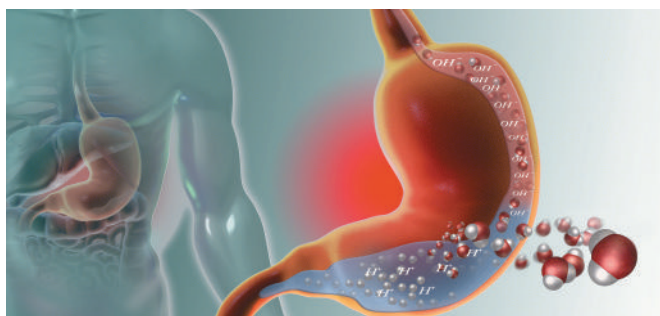
- 16.1 Fetal Hemoglobin and Equilibrium 697
- 16.2 The Concept of Dynamic Equilibrium 699
- 16.3 The Equilibrium Constant ( $K$ ) 702
  - Expressing Equilibrium Constants for Chemical Reactions 702 The Significance of the Equilibrium Constant 703
  - CHEMISTRY AND MEDICINE** Life and Equilibrium 704 Relationships between the Equilibrium Constant and the Chemical Equation 705
- 16.4 Expressing the Equilibrium Constant in Terms of Pressure 706
  - Relationship between  $K_p$  and  $K_c$  707 Units of  $K$  708
- 16.5 Heterogeneous Equilibria: Reactions Involving Solids and Liquids 709
- 16.6 Calculating the Equilibrium Constant from Measured Equilibrium Concentrations 710
- 16.7 The Reaction Quotient: Predicting the Direction of Change 713
- 16.8 Finding Equilibrium Concentrations 715
  - Finding Equilibrium Concentrations from the Equilibrium Constant and All but One of the Equilibrium Concentrations of the Reactants and Products 716
  - Finding Equilibrium Concentrations from the Equilibrium Constant and Initial Concentrations or Pressures 717
  - Simplifying Approximations in Working Equilibrium Problems 721
- 16.9 Le Châtelier's Principle: How a System at Equilibrium Responds to Disturbances 725
  - The Effect of a Concentration Change on Equilibrium 726
  - The Effect of a Volume (or Pressure) Change on Equilibrium 728 The Effect of a Temperature Change on Equilibrium 730

**CHAPTER IN REVIEW** Self-Assessment Quiz 732 Terms 733 Concepts 733 Equations and Relationships 734 Learning Outcomes 734

**EXERCISES** Review Questions 735 Problems by Topic 736 Cumulative Problems 740 Challenge Problems 742 Conceptual Problems 742 Questions for Group Work 743 Data Interpretation and Analysis 743 Answers to Conceptual Connections 745





**17 Acids and Bases** 746

- 17.1 Heartburn** 747
- 17.2 The Nature of Acids and Bases** 748
- 17.3 Definitions of Acids and Bases** 750  
The Arrhenius Definition 750 The Brønsted–Lowry Definition 751
- 17.4 Acid Strength and the Acid Ionization Constant ( $K_a$ )** 753  
Strong Acids 753 Weak Acids 754 The Acid Ionization Constant ( $K_a$ ) 755
- 17.5 Autoionization of Water and pH** 756  
The pH Scale: A Way to Quantify Acidity and Basicity 758 pOH and Other p Scales 759  
**CHEMISTRY AND MEDICINE** Ulcers 760
- 17.6 Finding the  $[H_3O^+]$  and pH of Strong and Weak Acid Solutions** 761  
Strong Acids 761 Weak Acids 761 Percent Ionization of a Weak Acid 766 Mixtures of Acids 767
- 17.7 Base Solutions** 770  
Strong Bases 770 Weak Bases 770 Finding the  $[OH^-]$  and pH of Basic Solutions 772  
**CHEMISTRY AND MEDICINE** What's in My Antacid? 774
- 17.8 The Acid–Base Properties of Ions and Salts** 774  
Anions as Weak Bases 775 Cations as Weak Acids 778 Classifying Salt Solutions as Acidic, Basic, or Neutral 779
- 17.9 Polyprotic Acids** 781  
Finding the pH of Polyprotic Acid Solutions 782 Finding the Concentration of the Anions for a Weak Diprotic Acid Solution 784
- 17.10 Acid Strength and Molecular Structure** 786  
Binary Acids 786 Oxyacids 787
- 17.11 Lewis Acids and Bases** 788  
Molecules That Act as Lewis Acids 788 Cations That Act as Lewis Acids 789

**CHAPTER IN REVIEW** Self-Assessment Quiz 790 Terms 791 Concepts 791 Equations and Relationships 792 Learning Outcomes 792

**EXERCISES** Review Questions 793 Problems by Topic 793 Cumulative Problems 797 Challenge Problems 799 Conceptual Problems 799 Questions for Group Work 799 Data Interpretation and Analysis 800 Answers to Conceptual Connections 801

**18 Aqueous Ionic Equilibrium** 802

- 18.1 The Danger of Antifreeze** 803
- 18.2 Buffers: Solutions That Resist pH Change** 804  
Calculating the pH of a Buffer Solution 806 The Henderson–Hasselbalch Equation 807 Calculating pH Changes in a Buffer Solution 810 The Stoichiometry Calculation 810 The Equilibrium Calculation 810 Buffers Containing a Base and Its Conjugate Acid 814

- 18.3 Buffer Effectiveness: Buffer Range and Buffer Capacity** 815  
Relative Amounts of Acid and Base 815 Absolute Concentrations of the Acid and Conjugate Base 816 Buffer Range 817

**CHEMISTRY AND MEDICINE** Buffer Effectiveness in Human Blood 818  
Buffer Capacity 818

- 18.4 Titrations and pH Curves** 819  
The Titration of a Strong Acid with a Strong Base 820 The Titration of a Weak Acid with a Strong Base 824 The Titration of a Weak Base with a Strong Acid 829 The Titration of a Polyprotic Acid 830 Indicators: pH-Dependent Colors 830

- 18.5 Solubility Equilibria and the Solubility Product Constant** 833

$K_{sp}$  and Molar Solubility 833

**CHEMISTRY IN YOUR DAY** Hard Water 835

$K_{sp}$  and Relative Solubility 836

The Effect of a Common Ion on Solubility 836

The Effect of pH on Solubility 838

- 18.6 Precipitation** 839  
Selective Precipitation 840

- 18.7 Qualitative Chemical Analysis** 842  
Group 1: Insoluble Chlorides 843 Group 2: Acid-Insoluble Sulfides 843 Group 3: Base-Insoluble Sulfides and Hydroxides 844 Group 4: Insoluble Phosphates 844 Group 5: Alkali Metals and  $NH_4^+$  844

- 18.8 Complex Ion Equilibria** 845  
The Effect of Complex Ion Equilibria on Solubility 847 The Solubility of Amphoteric Metal Hydroxides 848

**CHAPTER IN REVIEW** Self-Assessment Quiz 849 Terms 850 Concepts 850 Equations and Relationships 851 Learning Outcomes 851

**EXERCISES** Review Questions 852 Problems by Topic 853 Cumulative Problems 859 Challenge Problems 860 Conceptual Problems 861 Questions for Group Work 861 Data Interpretation and Analysis 862 Answers to Conceptual Connections 863



## 19 Free Energy and Thermodynamics 864



- 19.1 Cold Coffee and Dead Universes 865
- 19.2 Spontaneous and Nonspontaneous Processes 866
- 19.3 Entropy and the Second Law of Thermodynamics 868
  - Entropy 870 The Entropy Change upon the Expansion of an Ideal Gas 872
- 19.4 Entropy Changes Associated with State Changes 874
  - Entropy and State Change: The Concept 875 Entropy and State Changes: The Calculation 876
- 19.5 Heat Transfer and Changes in the Entropy of the Surroundings 878
  - The Temperature Dependence of  $\Delta S_{\text{surr}}$  879 Quantifying Entropy Changes in the Surroundings 879
- 19.6 Gibbs Free Energy 881
  - The Effect of  $\Delta H$ ,  $\Delta S$ , and  $T$  on Spontaneity 882
- 19.7 Entropy Changes in Chemical Reactions: Calculating  $\Delta S_{\text{rxn}}^{\circ}$  885
  - Defining Standard States and Standard Entropy Changes 885 Standard Molar Entropies ( $S^{\circ}$ ) and the Third Law of Thermodynamics 885 Calculating the Standard Entropy Change ( $\Delta S_{\text{rxn}}^{\circ}$ ) for a Reaction 889
- 19.8 Free Energy Changes in Chemical Reactions: Calculating  $\Delta G_{\text{rxn}}^{\circ}$  889
  - Calculating Standard Free Energy Changes with  $\Delta G_{\text{rxn}}^{\circ} = \Delta H_{\text{rxn}}^{\circ} - T\Delta S_{\text{rxn}}^{\circ}$  890 Calculating  $\Delta G_{\text{rxn}}^{\circ}$  with Tabulated Values of Free Energies of Formation 891
  - CHEMISTRY IN YOUR DAY** Making a Nonspontaneous Process Spontaneous 893
  - Calculating  $\Delta G_{\text{rxn}}^{\circ}$  for a Stepwise Reaction from the Changes in Free Energy for Each of the Steps 893
  - Why Free Energy Is “Free” 894
- 19.9 Free Energy Changes for Nonstandard States: The Relationship between  $\Delta G_{\text{rxn}}^{\circ}$  and  $\Delta G_{\text{rxn}}$  896
  - Standard versus Nonstandard States 896 The Free Energy Change of a Reaction under Nonstandard Conditions 896 Standard Conditions 896
  - Equilibrium Conditions 897 Other Nonstandard Conditions 898
- 19.10 Free Energy and Equilibrium: Relating  $\Delta G_{\text{rxn}}^{\circ}$  to the Equilibrium Constant ( $K$ ) 899
  - The Relationship between  $\Delta G_{\text{rxn}}^{\circ}$  and  $K$  899 The Temperature Dependence of the Equilibrium Constant 901

**CHAPTER IN REVIEW** Self-Assessment Quiz 902 Terms 903 Concepts 903 Equations and Relationships 904 Learning Outcomes 904

**EXERCISES** Review Questions 905 Problems by Topic 906 Cumulative Problems 910 Challenge Problems 911 Conceptual Problems 912 Questions for Group Work 912 Data Interpretation and Analysis 913 Answers to Conceptual Connections 913

## 20 Electrochemistry 914

- 20.1 Lightning and Batteries 915
- 20.2 Balancing Oxidation–Reduction Equations 916
- 20.3 Voltaic (or Galvanic) Cells: Generating Electricity from Spontaneous Chemical Reactions 919
  - The Voltaic Cell 920 Current and Potential Difference 921 Anode, Cathode, and Salt Bridge 922 Electrochemical Cell Notation 923
- 20.4 Standard Electrode Potentials 923
  - Predicting the Spontaneous Direction of an Oxidation–Reduction Reaction 928 Predicting Whether a Metal Will Dissolve in Acid 931
- 20.5 Cell Potential, Free Energy, and the Equilibrium Constant 931
  - The Relationship between  $\Delta G^{\circ}$  and  $E_{\text{cell}}^{\circ}$  932 The Relationship between  $E_{\text{cell}}^{\circ}$  and  $K$  934
- 20.6 Cell Potential and Concentration 935
  - Cell Potential under Nonstandard Conditions: The Nernst Equation 935 Concentration Cells 938
  - CHEMISTRY AND MEDICINE** Concentration Cells in Human Nerve Cells 940
- 20.7 Batteries: Using Chemistry to Generate Electricity 940
  - Dry-Cell Batteries 940 Lead–Acid Storage Batteries 941 Other Rechargeable Batteries 942 Fuel Cells 943
  - CHEMISTRY IN YOUR DAY** The Fuel Cell Breathalyzer 944
- 20.8 Electrolysis: Driving Nonspontaneous Chemical Reactions with Electricity 944
  - Predicting the Products of Electrolysis 947 Stoichiometry of Electrolysis 950
- 20.9 Corrosion: Undesirable Redox Reactions 951
  - Corrosion of Iron 952 Preventing the Corrosion of Iron 953

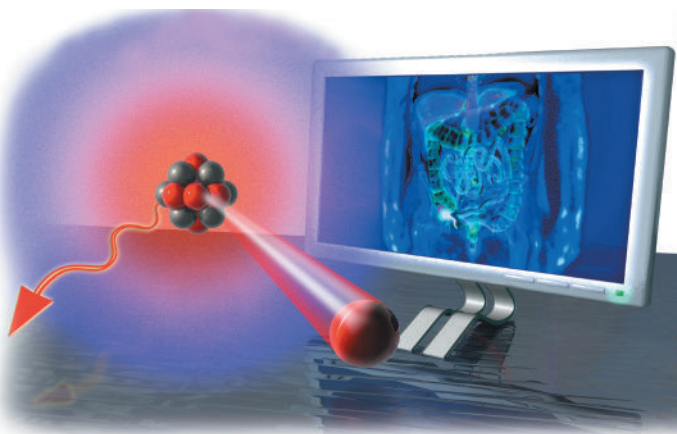
**CHAPTER IN REVIEW** Self-Assessment Quiz 954 Terms 955 Concepts 955 Equations and Relationships 956 Learning Outcomes 956

**EXERCISES** Review Questions 957 Problems by Topic 957 Cumulative Problems 961 Challenge Problems 962 Conceptual Problems 963 Questions for Group Work 963 Data Interpretation and Analysis 964 Answers to Conceptual Connections 965





## 21 Radioactivity and Nuclear Chemistry 966



- 21.1 Diagnosing Appendicitis 967
- 21.2 The Discovery of Radioactivity 968
- 21.3 Types of Radioactivity 969
  - Alpha ( $\alpha$ ) Decay 970 Beta ( $\beta$ ) Decay 971 Gamma ( $\gamma$ ) Ray Emission 972 Positron Emission 972 Electron Capture 973
- 21.4 The Valley of Stability: Predicting the Type of Radioactivity 974
  - Magic Numbers 976 Radioactive Decay Series 976
- 21.5 Detecting Radioactivity 977
- 21.6 The Kinetics of Radioactive Decay and Radiometric Dating 978
  - The Integrated Rate Law 980 Radiocarbon Dating: Using Radioactivity to Measure the Age of Fossils and Artifacts 981

**CHEMISTRY IN YOUR DAY** Radiocarbon Dating and the Shroud of Turin 983

  - Uranium/Lead Dating 983 The Age of Earth 984
- 21.7 The Discovery of Fission: The Atomic Bomb and Nuclear Power 985
  - The Manhattan Project 985 Nuclear Power: Using Fission to Generate Electricity 987 Problems with Nuclear Power 988
- 21.8 Converting Mass to Energy: Mass Defect and Nuclear Binding Energy 989
  - Mass Defect and Nuclear Binding Energy 989 The Nuclear Binding Energy Curve 991
- 21.9 Nuclear Fusion: The Power of the Sun 991
- 21.10 Nuclear Transmutation and Transuranium Elements 992
- 21.11 The Effects of Radiation on Life 994
  - Acute Radiation Damage 994 Increased Cancer Risk 994 Genetic Defects 994 Measuring Radiation Exposure and Dose 995
- 21.12 Radioactivity in Medicine and Other Applications 996
  - Diagnosis in Medicine 997 Radiotherapy in Medicine 998 Other Applications 998

**CHAPTER IN REVIEW** Self-Assessment Quiz 999 Terms 1000 Concepts 1000 Equations and Relationships 1001 Learning Outcomes 1001

**EXERCISES** Review Questions 1002 Problems by Topic 1002 Cumulative Problems 1005 Challenge Problems 1005 Conceptual Problems 1006 Questions for Group Work 1006 Data Interpretation and Analysis 1007 Answers to Conceptual Connections 1007

## 22 Organic Chemistry 1008

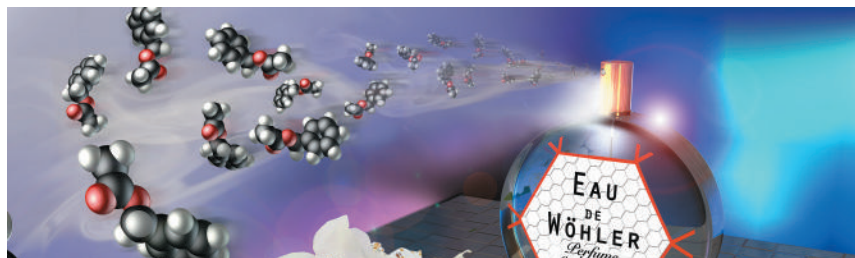
- 22.1 Fragrances and Odors 1009
- 22.2 Carbon: Why It Is Unique 1010
  - Carbon's Tendency to Form Four Covalent Bonds 1010
  - Carbon's Ability to Form Double and Triple Bonds 1010

**CHEMISTRY IN YOUR DAY** Vitalism and the Perceived Differences between Organic and Inorganic Compounds 1011

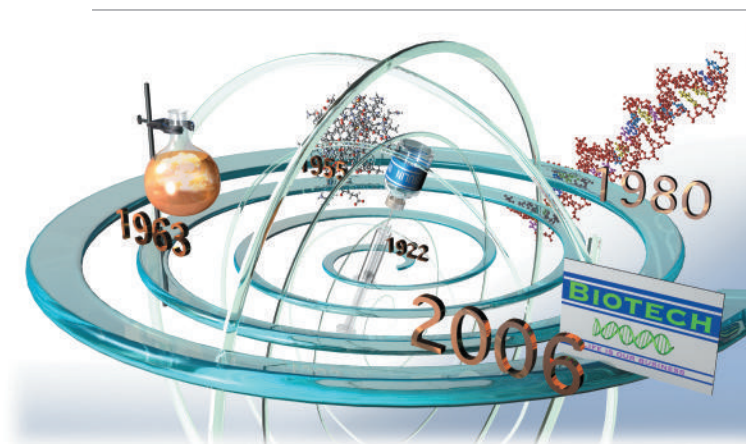
  - Carbon's Tendency to Catenate 1011
- 22.3 Hydrocarbons: Compounds Containing Only Carbon and Hydrogen 1012
  - Drawing Hydrocarbon Structures 1013 Stereoisomerism and Optical Isomerism 1015 Rotation of Polarized Light 1016 Chemical Behavior in a Chiral Environment 1017
- 22.4 Alkanes: Saturated Hydrocarbons 1018
  - Naming Alkanes 1019
- 22.5 Alkenes and Alkynes 1022
  - Naming Alkenes and Alkynes 1023 Geometric (Cis-Trans) Isomerism in Alkenes 1026
- 22.6 Hydrocarbon Reactions 1027
  - Reactions of Alkanes 1027 Reactions of Alkenes and Alkynes 1028
- 22.7 Aromatic Hydrocarbons 1030
  - Naming Aromatic Hydrocarbons 1030 Reactions of Aromatic Compounds 1032
- 22.8 Functional Groups 1033
- 22.9 Alcohols 1034
  - Naming Alcohols 1034 About Alcohols 1034 Alcohol Reactions 1034
- 22.10 Aldehydes and Ketones 1036
  - Naming Aldehydes and Ketones 1037 About Aldehydes and Ketones 1037 Aldehyde and Ketone Reactions 1038
- 22.11 Carboxylic Acids and Esters 1039
  - Naming Carboxylic Acids and Esters 1039 About Carboxylic Acids and Esters 1039 Carboxylic Acid and Ester Reactions 1040
- 22.12 Ethers 1041
  - Naming Ethers 1041 About Ethers 1042
- 22.13 Amines 1042
  - Amine Reactions 1042

**CHAPTER IN REVIEW** Self-Assessment Quiz 1043 Terms 1044 Concepts 1044 Equations and Relationships 1045 Learning Outcomes 1046

**EXERCISES** Review Questions 1047 Problems by Topic 1047 Cumulative Problems 1053 Challenge Problems 1055 Conceptual Problems 1055 Questions for Group Work 1056 Data Interpretation and Analysis 1056 Answers to Conceptual Connections 1057



## 23 Biochemistry 1058



- 23.1 Diabetes and the Synthesis of Human Insulin** 1059
- 23.2 Lipids** 1060
  - Fatty Acids 1060 Fats and Oils 1062
  - Other Lipids 1063
- 23.3 Carbohydrates** 1065
  - Simple Carbohydrates: Monosaccharides and Disaccharides 1065
  - Complex Carbohydrates 1067
- 23.4 Proteins and Amino Acids** 1068
  - Amino Acids: The Building Blocks of Proteins 1069
  - Peptide Bonding between Amino Acids 1071
- 23.5 Protein Structure** 1072
  - Primary Structure 1074
  - Secondary Structure 1074
  - Tertiary Structure 1075
  - Quaternary Structure 1076
- 23.6 Nucleic Acids: Blueprints for Proteins** 1076
  - The Basic Structure of Nucleic Acids 1076
  - The Genetic Code 1078
- 23.7 DNA Replication, the Double Helix, and Protein Synthesis** 1080
  - DNA Replication and the Double Helix 1080
  - Protein Synthesis 1081
- CHEMISTRY AND MEDICINE** *The Human Genome Project* 1082

**CHAPTER IN REVIEW** Self-Assessment Quiz 1083 Terms 1084  
Concepts 1084 Learning Outcomes 1085

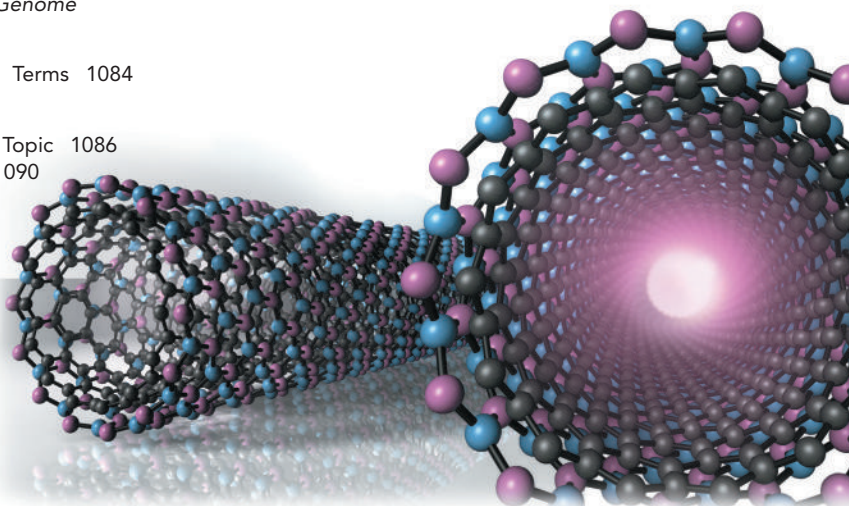
**EXERCISES** Review Questions 1086 Problems by Topic 1086  
Cumulative Problems 1089 Challenge Problems 1090  
Conceptual Problems 1090 Questions for Group Work 1090  
Data Interpretation and Analysis 1091  
Answers to Conceptual Connections 1091

## 24 Chemistry of the Nonmetals 1092

- 24.1 Insulated Nanowires** 1093
- 24.2 The Main-Group Elements: Bonding and Properties** 1094
- 24.3 Silicates: The Most Abundant Matter in Earth's Crust** 1096
  - Quartz 1096 Aluminosilicates 1097
  - Individual Silicate Units, Silicate Chains, and Silicate Sheets 1097
- 24.4 Boron and Its Remarkable Structures** 1100
  - Elemental Boron 1100
  - Boron-Halogen Compounds: Trihalides 1101
  - Boron-Oxygen Compounds 1101
  - Boron-Hydrogen Compounds: Boranes 1102
- 24.5 Carbon, Carbides, and Carbonates** 1103
  - Amorphous Carbon 1103
  - Carbides 1104
  - Carbon Oxides 1105
  - Carbonates 1106
- 24.6 Nitrogen and Phosphorus: Essential Elements for Life** 1107
  - Elemental Nitrogen and Phosphorus 1107
  - Nitrogen Compounds 1108
  - Phosphorus Compounds 1111
- 24.7 Oxygen** 1113
  - Elemental Oxygen 1113
  - Uses for Oxygen 1114
  - Oxides 1114
  - Ozone 1115
- 24.8 Sulfur: A Dangerous but Useful Element** 1115
  - Elemental Sulfur 1116
  - Hydrogen Sulfide and Metal Sulfides 1117
  - Sulfur Dioxide 1118
  - Sulfuric Acid 1118
- 24.9 Halogens: Reactive Elements with High Electronegativity** 1119
  - Elemental Fluorine and Hydrofluoric Acid 1120
  - Elemental Chlorine 1121
  - Halogen Compounds 1121

**CHAPTER IN REVIEW** Self-Assessment Quiz 1123 Terms 1124  
Concepts 1124 Learning Outcomes 1125

**EXERCISES** Review Questions 1125 Problems by Topic 1125  
Cumulative Problems 1127 Challenge Problems 1128  
Conceptual Problems 1128 Questions for Group Work 1128  
Data Interpretation and Analysis 1129  
Answers to Conceptual Connections 1129





## 25 Metals and Metallurgy 1130



### 25.1 Vanadium: A Problem and an Opportunity 1131

### 25.2 The General Properties and Natural Distribution of Metals 1132

### 25.3 Metallurgical Processes 1134

Separation 1134 Pyrometallurgy 1134  
Hydrometallurgy 1135  
Electrometallurgy 1136  
Powder Metallurgy 1137

### 25.4 Metal Structures and Alloys 1138

Alloys 1138 Substitutional Alloys 1138  
Alloys with Limited Solubility 1140  
Interstitial Alloys 1141

### 25.5 Sources, Properties, and Products of Some of the 3d Transition Metals 1143

Titanium 1143 Chromium 1144 Manganese 1145  
Cobalt 1146 Copper 1147  
Nickel 1147 Zinc 1148

**CHAPTER IN REVIEW** Self-Assessment Quiz 1148 Terms 1150  
Concepts 1150 Equations and Relationships 1150 Learning Outcomes 1151

**EXERCISES** Review Questions 1151 Problems by Topic 1151  
Cumulative Problems 1153 Challenge Problems 1153  
Conceptual Problems 1153 Questions for Group Work 1154  
Data Interpretation and Analysis 1154 Answers to Conceptual Connections 1155

## 26 Transition Metals and Coordination Compounds 1156

### 26.1 The Colors of Rubies and Emeralds 1157

### 26.2 Properties of Transition Metals 1158

Electron Configurations 1158 Atomic Size 1160  
Ionization Energy 1160 Electronegativity 1161  
Oxidation States 1161

### 26.3 Coordination Compounds 1162

Naming Coordination Compounds 1165

### 26.4 Structure and Isomerization 1167

Structural Isomerism 1167 Stereoisomerism 1169

### 26.5 Bonding in Coordination Compounds 1172

Valence Bond Theory 1172 Crystal Field Theory 1173

### 26.6 Applications of Coordination Compounds 1177

Chelating Agents 1178 Chemical Analysis 1178  
Coloring Agents 1178 Biomolecules 1178  
Hemoglobin and Cytochrome C 1179



**CHAPTER IN REVIEW** Self-Assessment Quiz 1181 Terms 1181  
Concepts 1182 Equations and Relationships 1182 Learning Outcomes 1182

**EXERCISES** Review Questions 1183 Problems by Topic 1183  
Cumulative Problems 1185 Challenge Problems 1185  
Conceptual Problems 1186 Questions for Group Work 1186  
Data Interpretation and Analysis 1186 Answers to Conceptual Connections 1187

**Appendix I** Common Mathematical Operations in Chemistry A-1

**Appendix II** Useful Data A-5

**Appendix III** Answers to Selected Exercises A-15

**Appendix IV** Answers to In-Chapter Practice Problems A-55

**Glossary** G-1

**Photo and Text Credits** C-1

**Index** I-1

## To the Student

As you begin this course, I invite you to think about your reasons for enrolling in it. Why are you taking general chemistry? More generally, why are you pursuing a college education? If you are like most college students taking general chemistry, part of your answer is probably that this course is required for your major and that you are pursuing a college education so you can get a good job some day. Although these are good reasons, I would like to suggest a better one. I think the primary reason for your education is to prepare you to *live a good life*. You should understand chemistry—not for what it can *get* you—but for what it can *do* to you. Understanding chemistry, I believe, is an important source of happiness and fulfillment. Let me explain.

Understanding chemistry helps you to live life to its fullest for two basic reasons. The first is *intrinsic*: through an understanding of chemistry, you gain a powerful appreciation for just how rich and extraordinary the world really is. The second reason is *extrinsic*: understanding chemistry makes you a more informed citizen—it allows you to engage with many of the issues of our day. In other words, understanding chemistry makes *you* a deeper and richer person and makes your country and the world a better place to live. These reasons have been the foundation of education from the very beginnings of civilization.

How does chemistry help prepare you for a rich life and conscientious citizenship? Let me explain with two examples. My first one comes from the very first page of Chapter 1 of this book. There, I ask the following question: What is the most important idea in all of scientific knowledge? My answer to that question is this: **the behavior of matter is determined by the properties of molecules and atoms**. That simple statement is the reason I love chemistry. We humans have been able to study the substances that compose the world around us and explain their behavior by reference to particles so small that they can hardly be imagined. If you have never realized the remarkable dependence of the world we *can* see on the world we *cannot*, you have missed out on a fundamental truth about our universe. To have never encountered this truth is like never having read a play by Shakespeare or seen a sculpture by Michelangelo—or, for that matter, like never having discovered that the world is round. It robs you of an amazing and unforgettable experience of the world and the human ability to understand it.

My second example demonstrates how science literacy helps you to be a better citizen. Although I am largely sympathetic to the environmental movement, a lack of science literacy within some sectors of that movement and the resulting

anti-environmental backlash create confusion that impedes real progress and opens the door to what could be misinformed policies. For example, I have heard conservative pundits say that volcanoes emit more carbon dioxide—the most significant greenhouse gas—than does petroleum combustion. I have also heard a liberal environmentalist say that we have to stop using hair spray because it is causing holes in the ozone layer that will lead to global warming. Well, the claim about volcanoes emitting more carbon dioxide than petroleum combustion can be refuted by the basic tools you will learn to use in Chapter 4 of this book. We can easily show that volcanoes emit only 1/50th as much carbon dioxide as petroleum combustion. As for hair spray depleting the ozone layer and thereby leading to global warming, the chlorofluorocarbons that deplete ozone have been banned from hair spray since 1978, and ozone depletion has nothing to do with global warming anyway. People with special interests or axes to grind can conveniently distort the truth before an ill-informed public, which is why we all need to be knowledgeable.

So this is why I think you should take this course. Not just to satisfy the requirement for your major and not just to get a good job some day, but to help you to lead a fuller life and to make the world a little better for everyone. I wish you the best as you embark on the journey to understanding the world around you at the molecular level. The rewards are well worth the effort.

## To the Professor

Thanks to all of you who adopted this book in its previous editions. You helped to make this book one of the most popular general chemistry textbooks in the world. I am grateful beyond words. Second, I have listened carefully to your feedback on the previous edition. The changes you see in this edition are the direct result of your input, as well as my own experience using the book in my general chemistry courses. If you have reviewed content or have contacted me directly, you will likely see your suggestions reflected in the changes I have made. Thank you.

This revision comes out in the midst of a pandemic and after a challenging year, for both the world and for higher education. Most of us taught remotely during the 2020–2021 academic year, lecturing into a Zoom meeting of often faceless students. The year demonstrated the power of technology for learning and also its shortcomings. I learned two important lessons: (1) Face-to-face teaching continues to be the superior mode for college instruction; and (2) technology

can deliver learning experiences that are more valuable than passive studying. This revision focuses on lesson 2.

Technology can transform passive learning into active learning, which has been repeatedly demonstrated to be more effective. One of my main goals in this revision is to continue to expand tools to engage students in active learning. Although the term *active learning* has been applied mainly to in-class learning, the principal idea—that we learn better when we are actively engaged—applies to all of learning.

For this edition, I have worked with Pearson to develop 50 *Key Concept Interactives* (KCIs). Each KCI guides the student through a key concept in that chapter. However, rather than passively reading a book, the student must continually interact with the material as they navigate through it. In the KCI on nomenclature, for example, students first learn how to categorize compounds as ionic, molecular, or acid. Before they advance, they are given a compound to categorize. If they categorize it incorrectly, they are provided with feedback to lead them in the right direction. If they categorize it correctly, they can advance. In this way, the student is completely active in the learning process, which we know from education research results in much greater retention than passive reading. In addition, each KCI has an associated follow-up question that can be assigned using Mastering Chemistry.

Other active learning tools, presented in previous editions, have been expanded in this edition. For example, the library of 3- to 6-minute *Key Concept Videos* (KCVs) now spans virtually all of the key concepts in a general chemistry course. The videos introduce a key concept and encourage active learning because they stop in the middle and pose a question that must be answered before the video continues playing. Each video also has an associated follow-up question that can be assigned using Mastering Chemistry.

Also expanded in this edition is the library of 3- to 6-minute videos called *Interactive Worked Examples* (IWEs). Each IWE video walks a student through the solution to a chemistry problem. Like the KCV, the IWE video stops in the middle and poses a question that must be answered before the video continues playing. Each video also has an associated follow-up problem that can be assigned using Mastering Chemistry. Together, the library of KCVs and IWEs now comprises approximately 250 interactive videos.

Although we have added these active learning tools to this edition and made other changes as well, the book's goal remains the same: *to present a rigorous and accessible treatment of general chemistry in the context of relevance*. Teaching general chemistry would be much easier if all of our students had exactly the same level of preparation and ability. But alas, that is not the case. My own courses are populated with students with a range of backgrounds and abilities in chemistry. The challenge of successful teaching, in my opinion, is figuring out how to instruct and challenge the best students while not losing those with lesser backgrounds and abilities. My strategy has always been to set the bar relatively high, while at the same time providing the motivation and support necessary to reach the high bar. That is exactly the philosophy of

this book. We do not have to compromise away rigor in order to make chemistry accessible to our students. In this book, I have worked hard to combine rigor with accessibility—to create a book that does not dilute the content, and yet, can be used and understood by any student willing to put in the necessary effort.

**Chemistry: A Molecular Approach is first and foremost a student-oriented book.** My main goal is to motivate students and get them to achieve at the highest possible level. As we all know, many students take general chemistry because it is a requirement; they do not see the connection between chemistry and their lives or their intended careers. *Chemistry: A Molecular Approach* strives to make those connections consistently and effectively. Unlike other books, which often teach chemistry as something that happens only in the laboratory or in industry, this book teaches chemistry in the context of relevance. It shows students *why* chemistry is important to them, to their future careers, and to their world.

**Second, Chemistry: A Molecular Approach is a pedagogically driven book.** In seeking to develop problem-solving skills, it applies a consistent approach (Sort, Strategize, Solve, and Check), usually in a two- or three-column format. In the two-column format, the left column shows the student how to analyze the problem and devise a solution strategy. It also lists the steps of the solution, explaining the rationale for each one, while the right column shows the implementation of each step. In the three-column format, the left column outlines the general procedure for solving an important category of problems that is then applied to two side-by-side examples. This strategy allows students to see both the general pattern and the slightly different ways in which the procedure may be applied in differing contexts. The aim is to help students understand both the *concept of the problem* (through the formulation of an explicit conceptual plan for each problem) and the *solution to the problem*.

**Third, Chemistry: A Molecular Approach is a visual book.** Wherever possible, I use images to deepen the student's insight into chemistry. In developing chemical principles, multipart images help show the connection between everyday processes visible to the unaided eye and what atoms and molecules are actually doing. Many of these images have three parts: macroscopic, molecular, and symbolic. This combination helps students to see the relationships between the formulas they write down on paper (symbolic), the world they see around them (macroscopic), and the atoms and molecules that compose that world (molecular). In addition, most figures are designed to teach rather than just to illustrate. They are rich with annotations and labels intended to help the student grasp the most important processes and the principles that underlie them. In this edition, the art program has been thoroughly revised in two major ways. First, navigation of the more complex figures has been reoriented to track from left to right whenever possible. Second, figure captions have been migrated into the image itself as an "author voice" that explains the image and

guides the reader through it. The resulting images are rich with information but also clear and quickly understood.

**Fourth, *Chemistry: A Molecular Approach* is a “big picture” book.** At the beginning of each chapter, a short paragraph helps students to see the key relationships between the different topics they are learning. Through a focused and concise narrative, I strive to make the basic ideas of every chapter clear to the student. Interim summaries are provided at selected spots in the narrative, making it easier to grasp (and review) the main points of important discussions. And to make sure that students never lose sight of the forest for the trees, each chapter includes several *Conceptual Connections*, which ask them to think about concepts and solve problems without doing any math. I want students to learn the concepts, not just plug numbers into equations to churn out the right answer. This philosophy is also integral to the *Key Concept Videos*, which concisely reinforce student appreciation of the core concepts in each chapter.

***Chemistry: A Molecular Approach* is lastly a book that delivers the depth of coverage faculty want.** We do not have to cut corners and water down the material in order to get our students interested. We have to meet them where they are, challenge them to the highest level of achievement, and support them with enough pedagogy to allow them to succeed.

I hope that this book supports you in your vocation of teaching students chemistry. I am increasingly convinced of the importance of our task. Please feel free to contact me with any questions or comments about the book.

Nivaldo J. Tro  
nivatro@gmail.com

## Acknowledgments

The book you hold in your hands bears my name on the cover, but I am really only one member of a large team that carefully crafted this book. Most importantly, I thank my editor, Elizabeth Bell. I have known Lizzy for many years, first as my marketing manager and now as my editor. Lizzy is a constant source of good, market-informed ideas. She is creative, energetic, and always follows through on her ideas and commitments. I am fortunate to work with such a skilled and insightful colleague. Thanks, Lizzy. Thanks also to Coleen Morrison, my new developmental editor on this project. Coleen has helped me to develop the new material in this edition. She is careful, organized, and always improves my work. Thanks, Coleen, for your attention to detail and all your hard work. Thanks also to my project manager, Shercian Kinoshian. She has managed the many details and moving parts of producing this book with care and precision. I appreciate her steady hand and hard work. Thanks also to my media editor, Jackie Jacob. Jackie and I have been working together for many years to produce innovative media pieces that are pedagogically sound and easy to use. She is simply the best in the business, and I am lucky to get to work with her. I am also grateful to Ian Desrosiers and Chloe Veylit who have helped tremendously with the development of the new Key Concept

Interactives and have stepped in at different stages to offer their talent and expertise.

I am also grateful to Jeanne Zalesky, Editor-in-Chief for Physical Sciences. She has supported me and my projects and allowed me to succeed. Thanks also to Adam Jaworski, who oversees product management in Science at Pearson. I am grateful to have his wise and steady, yet innovative, hand at the wheel, especially during the many changes that are happening within educational publishing. I am also grateful to the Lachina Creative, Inc. team and Jerilyn DiCarlo for their creativity and hard work in crafting the cover illustration and design of this text. I also thank Francesca Monaco and her coworkers at Straive. I am a picky author and Francesca is endlessly patient and a true professional. I am also greatly indebted to my copy editor, Betty Pessagno, for her dedication and professionalism over many projects.

I am also grateful to those who have supported me personally while working on this book. First on that list is my wife, Ann. Her patience and love for me are beyond description, and without her, this book would never have been written. I am also indebted to my children, Michael, Ali, Kyle, and Kaden, whose smiling faces and love of life always inspire me. I come from a large Cuban family whose closeness and support most people would envy. Thanks to my parents, Nivaldo and Sara; my siblings, Sarita, Mary, and Jorge; my siblings-in-law, Nachy, Karen, and John; and my nephews and nieces, Germain, Danny, Lisette, Sara, and Kenny. These are the people with whom I celebrate life.

I am especially grateful to Kyle Tro, who put in many hours proofreading changes in the manuscript, working problems, and organizing appendices. Kyle, you are an amazing person—it is my privilege to have you work with me on this project.

I would like to thank all of the general chemistry students who have been in my classes throughout my 32 years as a professor. You have taught me much about teaching that is now in this book, and you are the motivation that keeps me inspired to develop the best learning tools possible.

Lastly, I am indebted to the many reviewers, listed on the following pages, whose ideas are embedded throughout this book. They have corrected me, inspired me, and sharpened my thinking on how best to teach this subject we call chemistry. I deeply appreciate their commitment to this project. I am also grateful to the accuracy reviewers who tirelessly checked page proofs for correctness.

## Reviewers of the Sixth Edition

Paul Buckley, *Washington State University*  
Daniel Collins, *Texas A&M University*  
Gabrielle Cook, *Oklahoma State University*  
Ashley Curtiss, *Auburn University*  
Antara Dutta, *Georgia State University*  
Shashi Jain, *University of Texas Rio Grande Valley*  
Jie Jiang, *Georgia State University*  
Jennifer Kelley, *Kennesaw State University*  
Jason McAfee, *University of North Texas*  
Smita Mohanty, *Oklahoma State University*  
Joe Merola, *Virginia Tech University*



Oluwatobi Odeleye, *University of Oklahoma*  
 Tamkio Porter, *Indiana University-Purdue University Indianapolis*  
 Martina Sumner, *University of Florida*  
 Shuai Sun, *University of Kentucky*  
 Rebecca Weber, *University of North Texas*  
 Lin Zhu, *Indiana University-Purdue University Indianapolis*

## Reviewers of Previous Editions

Michael R. Abraham, *University of Oklahoma*  
 Patricia G. Amateis, *Virginia Tech*  
 Ramesh D. Arasasingham, *University of California, Irvine*  
 James A. Armstrong, *City College of San Francisco*  
 Margaret R. Asirvatham, *University of Colorado*  
 Paul Badger, *Robert Morris University*  
 Monica H. Baloga, *Florida Institute of Technology*  
 Rebecca Barlag, *Ohio University*  
 Mufeed M. Basti, *North Carolina Agricultural & Technological State University*  
 Corey Beck, *Ohio University*  
 Amy E. Beilstein, *Centre College*  
 Donald Bellew, *The University of New Mexico*  
 Maria Benavides, *University of Houston, Downtown*  
 Kyle A. Beran, *University of Texas of the Permian Basin*  
 Thomas Bertolini, *University of Southern California*  
 Christine V. Bilicki, *Pasadena City College*  
 Silas C. Blackstock, *University of Alabama*  
 Robert E. Blake, *Texas Tech University*  
 Angela E. Boerger, *Loyola University Chicago*  
 Roberto A. Bogomolni, *University of California, Santa Cruz*  
 Robert S. Boikess, *Rutgers University*  
 Paul Brandt, *North Central College*  
 Michelle M. Brooks, *College of Charleston*  
 Stacey Brydges, *University of California San Diego*  
 Gary Buckley, *Cameron University*  
 Joseph H. Bularzik, *Purdue University, Calumet*  
 Michael Burand, *Oregon State University*  
 Cindy M. Burkhardt, *Radford University*  
 Andrew E. Burns, *Kent State University at Stark*  
 Kim C. Calvo, *The University of Akron*  
 Kenneth Capps, *Central Florida Community College*  
 Stephen C. Carlson, *Lansing Community College*  
 Charles E. Carraher, *Florida Atlantic University*  
 David A. Carter, *Angelo State University*  
 Vanessa Castleberry, *Baylor University*  
 Kenneth Caswell, *University of South Florida*  
 Ferman Chavez, *Oakland University*  
 Eric G. Chesloff, *Villanova University*  
 Ted Clark, *The Ohio State University*  
 William M. Cleaver, *The University of Vermont*  
 Christopher Collison, *Rochester Institute of Technology*  
 Kim Cortes, *Kennesaw State University*  
 J. Ricky Cox, *Murray State University*  
 Charles T. Cox Jr., *Georgia Institute of Technology*  
 Samuel R. Cron, *Arkansas State University*  
 Guy Crundwell, *Central Connecticut State University*  
 David M. Sarno, *CUNY, Queensborough Community College*  
 Patrick M. Lloyd, *CUNY, Kingsborough Community College*  
 Darwin B. Dahl, *Western Kentucky University*  
 Mohammed Daoudi, *University of Central Florida*  
 Patrick Daubenmire, *Loyola University-Chicago*  
 Kate Deline, *College of San Mateo*  
 Michael Dianovsky, *South Dakota State University*  
 Robert F. Dias, *Old Dominion University*  
 Stephanie Dillon, *Florida State University*  
 Bonnie Dixon, *University of Maryland*  
 Daniel S. Domin, *Tennessee State University*  
 Ralph C. Dougherty, *Florida State University*  
 Jennifer Duis, *Northern Arizona University*  
 Alan D. Earhart, *Southeast Community College*  
 William Eck, *University of Wisconsin, Marshfield/Wood County*  
 Jack Eichler, *University of California, Riverside*  
 Robert J. Eierman, *University of Wisconsin, Eau Claire*  
 Amina K. El-Ashmawy, *Collin County Community College*

Joseph P. Ellison, *United States Military Academy at West Point*  
 Joseph M. Eridon, *Central New Mexico Community College*  
 Deborah B. Exton, *University of Oregon*  
 William A. Faber, *Grand Rapids Community College*  
 Jordan Fantini, *Denison University*  
 Michael Ferguson, *Honolulu Community College*  
 Maria C. Fermin-Ennis, *Gordon College*  
 Oscar Navarro Fernandez, *University of Hawaii at Manoa*  
 Jan Florian, *Loyola University Chicago*  
 Richard W. Frazee, *Rowan University*  
 Andrew Frazer, *University of Central Florida*  
 Andy Frazer, *University of Central Florida*  
 Candice E. Fulton, *Midwestern State University*  
 Barbara A. Gage, *Prince George's Community College*  
 Ron Garber, *California State University at Long Beach*  
 Carlos D. Garcia, *The University of Texas at San Antonio*  
 Simon Garrett, *California State University, Northridge*  
 Raymond F. Glienna, *Glendale Community College*  
 Joel Goldberg, *University of Vermont*  
 Eric S. Goll, *Brookdale Community College*  
 Robert A. Gossage, *Acadia University*  
 Pierre Y. Goueth, *Santa Monica College*  
 Thomas J. Greenbowe, *Iowa State University*  
 W. Tandy Grubbs, *Stetson University*  
 Victoria Guarisco, *Middle Georgia State University*  
 Christin Gustafson, *Illinois Central College*  
 Jerome E. Haky, *Florida Atlantic University*  
 Jason A. Halfen, *University of Wisconsin-Eau Claire*  
 Nathan Hammer, *University of Mississippi*  
 Michael D. Hampton, *University of Central Florida*  
 Tamara Hanna, *Texas Tech University*  
 Lois Hansen-Polcar, *Cuyahoga Community College-Western Campus*  
 John A. W. Harkless, *Howard University*  
 Tony Hascall, *Northern Arizona University*  
 Alton Hassell, *Baylor University*  
 Elda Hegmann, *Kent State University*  
 Monte L. Helm, *Fort Lewis College*  
 David E. Henderson, *Trinity College*  
 Susan K. Henderson, *Quinnipiac University*  
 Peter M. Hierl, *The University of Kansas*  
 Paul I. Higgs, *Barry University*  
 Melissa Hines, *Cornell University*  
 Paula Hjorth-Gustin, *San Diego Mesa College*  
 Angela Hoffman, *University of Portland*  
 Norris W. Hoffman, *University of South Alabama*  
 Tony Holland, *Wallace Community College*  
 Todd A. Hopkins, *Butler University*  
 Byron E. Howell, *Tyler Junior College*  
 Tina Huang, *University of Illinois*  
 Amy Irwin, *Monroe Community College*  
 Moheb Ishak, *St. Petersburg College, St. Petersburg*  
 Kamal Ismail, *CUNY, Bronx Community College*  
 Ralph Isovitsch, *Xavier University of Louisiana*  
 Kenneth C. Janda, *University of California, Irvine*  
 Milt Johnston, *University of South Florida*  
 Greg M. Jorgensen, *American River College*  
 Sharon K. Kapica, *County College of Morris*  
 Jason A. Kautz, *University of Nebraska-Lincoln*  
 Mark Kearley, *Florida State University*  
 Catherine A. Keenan, *Chaffey College*  
 Steven W. Keller, *University of Missouri, Columbia*  
 Resa Kelly, *San Jose State University*  
 Chulsung Kim, *Georgia Gwinnett College*  
 Ellen Kime-Hunt, *Riverside Community College, Riverside Campus*  
 Edith Preciosa Kippenhan, *University of Toledo*  
 Louis J. Kirschenbaum, *University of Rhode Island*  
 Mark Knecht, *University of Kentucky*  
 Maria Korolev, *University of Florida*  
 Bette Kreuz, *University of Michigan-Dearborn*  
 Peter J. Krieger, *Palm Beach Community College, Lake Worth*  
 Sergiy Kryatov, *Tufts University*  
 Roy A. Lacey, *State University of New York, Stony Brook*  
 Richard H. Langley, *Stephen F. Austin State University*  
 Barry Lavine, *Oklahoma State University*  
 Clifford B. Lemaster, *Boise State University*

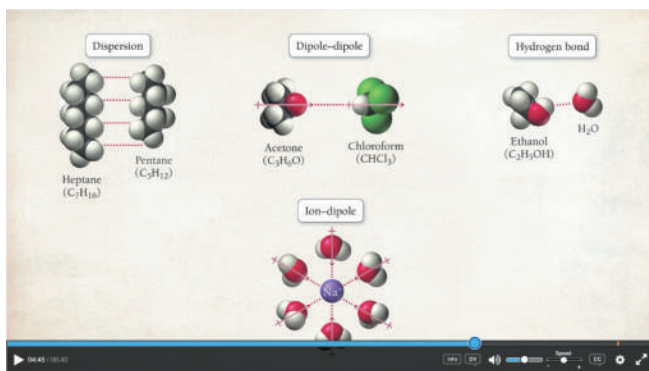
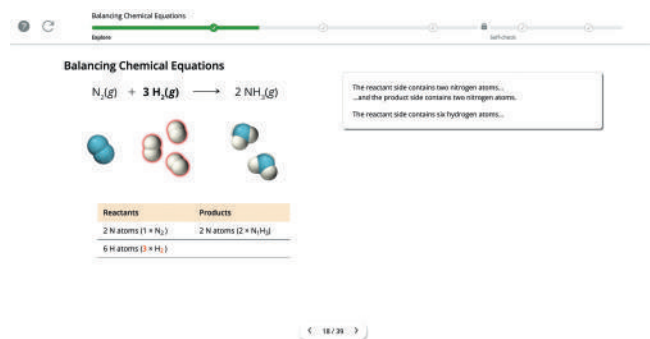
- Diana Leung, *The University of Alabama*  
 David P. Licata, *Coastline Community College*  
 Sarah Lievens, *University of California, Davis*  
 Robley Light, *Florida State University*  
 Michael E. Lipschutz, *Purdue University*  
 Adam List, *Vanderbilt University*  
 Patrick M. Lloyd, *CUNY, Kingsborough Community College*  
 Boon H. Loo, *Towson University*  
 Christopher Lovallo, *Mount Royal University*  
 James L. Mack, *Fort Valley State University*  
 Jeanette C. Madea, *Broward Community College, North*  
 Eric Malina, *University of Nebraska-Lincoln*  
 Joseph L. March, *University of Alabama, Birmingham*  
 Benjamin R. Martin, *Texas State University*  
 Lydia J. Martinez-Rivera, *University of Texas at San Antonio*  
 Marcus T. McEllistrem, *University of Wisconsin-Eau Claire*  
 Danny G. McGuire, *Cameron University*  
 Jack F. McKenna, *St. Cloud State University*  
 Charles W. McLaughlin, *University of Nebraska, Lincoln*  
 Curtis L. McLendon, *Saddleback College*  
 Lauren McMills, *Ohio University*  
 Robert C. McWilliams, *United States Military Academy*  
 Dianne Meador, *American River College*  
 Behnouth Memari, *Broward College*  
 David H. Metcalf, *University of Virginia*  
 John A. Milligan, *Los Angeles Valley College*  
 Ray Mohseni, *East Tennessee State University*  
 Alice J. Monroe, *St. Petersburg College, Clearwater*  
 Elisabeth A. Morlino, *University of the Sciences, Philadelphia*  
 Thomas Mullen, *University of North Florida*  
 Nancy Mullins, *Florida State College at Jacksonville*  
 James E. Murphy, *Santa Monica College*  
 Maria C. Nagan, *Truman State University*  
 Edward J. Neth, *University of Connecticut*  
 Heino Nitsche, *University of California at Berkeley*  
 Aric Opdahl, *University of Wisconsin La Crosse*  
 Kenneth S. Overway, *Bates College*  
 Greg Owens, *University of Utah*  
 Naresh Pandya, *University of Hawaii*  
 George Papadantonakis, *The University of Illinois at Chicago*  
 Gerard Parkin, *Columbia University*  
 Jessica Parr, *University of Southern California*  
 Yasmin Patell, *Kansas State University*  
 Pedro Patino, *University of Central Florida*  
 Tom Pentecost, *Grand Valley State University*  
 David Perdian, *Broward College*  
 Jeremy Perotti, *Nova Southeastern University*  
 Glenn A. Petrie, *Central Missouri State*  
 Norbert J. Pienta, *University of Iowa*  
 Louis H. Pignolet, *University of Minnesota*  
 Jennifer Schwartz Poehlmann, *Stanford University*  
 Jerry Poteat, *Georgia Perimeter College*  
 Daniele Ramella, *Temple University*  
 Jayashree Ranga, *Salem State University*  
 Cathrine E. Reck, *Indiana University*  
 Valerie Reeves, *University of New Brunswick*  
 Dawn J. Richardson, *Colin College*  
 Thomas G. Richmond, *University of Utah*  
 Dana L. Richter-Egger, *University of Nebraska*  
 Thomas Ridgway, *University of Cincinnati*  
 Jason Ritchie, *University of Mississippi*  
 Jil Robinson, *Indiana University*  
 Richard Rosso, *St. John's University*  
 Steven Rowley, *Middlesex County College*  
 Christopher P. Roy, *Duke University*  
 A. Timothy Royappa, *University of West Florida*  
 Stephen P. Ruis, *American River College*  
 Benjamin E. Rusiloski, *Delaware Valley College*  
 Raymond Sadeghi, *The University of Texas at San Antonio*  
 Alan E. Sadurski, *Ohio Northern University*  
 Karen Sanchez, *Florida Community College, Jacksonville*  
 Reva A. Savkar, *Northern Virginia Community College*  
 Raymond Schaak, *Penn State University*  
 Thomas W. Schleich, *University of California, Santa Cruz*  
 Rod Schoonover, *CA Polytechnic State University*  
 Mark Schraf, *West Virginia University*  
 John Selegue, *University of Kentucky*  
 Tom Selegue, *Pima Community College, West*  
 Susan Shadle, *Boise State University*  
 Kathleen Thrush Shaginaw, *Villanova University*  
 Jennifer Shanoski, *Merritt College*  
 Anju H. Sharma, *Stevens Institute of Technology*  
 Tricia Shepherd, *St. Edward's University*  
 Donald L. Siegel, *Rutgers University, New Brunswick*  
 Sarah Siegel, *Gonzaga University*  
 Mary L. Sohn, *Florida Institute of Technology*  
 Sherril A. Soman, *Grand Valley State University*  
 Sherril Soman-Williams, *Grand Valley State University*  
 Michael S. Sommer, *University of Wyoming*  
 Jie S. Song, *University of Michigan, Flint*  
 Clarissa Sorensen, *Central New Mexico Community College*  
 Mary Kay Sorenson, *University of Wisconsin, Milwaukee*  
 Allison Soult, *University of Kentucky*  
 Louise S. Sowers, *Richard Stockton College of New Jersey*  
 Stacy E. Sparks, *University of Texas, Austin*  
 Richard Spinney, *Ohio State University*  
 Anne Spuches, *East Carolina University*  
 Jeff Statler, *University of Utah*  
 William H. Steel, *York College of Pennsylvania*  
 Vinodhkumar Subramaniam, *East Carolina University*  
 Jerry Suits, *University of Northern Colorado*  
 Shuai Sun, *University of Kansas*  
 Tamar Y. Susskind, *Oakland Community College*  
 Uma Swamy, *Florida International University*  
 Ryan Sweeder, *Michigan State University*  
 Richard E. Sykora, *University of South Alabama*  
 Galina G. Talanova, *Howard University*  
 Dennis Taylor, *Clemson University*  
 Robert Craig Taylor, *Oakland University*  
 Claire A. Tessier, *University of Akron*  
 Jacquelyn Thomas, *Southwestern College*  
 Lydia Tien, *Monroe Community College*  
 Tara Todd, *Vanderbilt University*  
 David Livingstone Toppen, *California State University Northridge*  
 Marcy Towns, *Purdue University*  
 Harold Trimm, *Broome Community College*  
 Frank Tsung, *Boston College*  
 Laura VanDorn, *University of Arizona*  
 Susan Varkey, *Mount Royal College*  
 Ramaiyer Venkatraman, *Jackson State University*  
 John Vincent, *University of Alabama*  
 John B. Vincent, *University of Alabama, Tuscaloosa*  
 Kent S. Voelkner, *Lake Superior College*  
 Sheryl K. Wallace, *South Plains College*  
 Wayne E. Wesolowski, *University of Arizona*  
 Sarah E. West, *Notre Dame University*  
 John Wiginton, *University of Mississippi*  
 Kurt J. Winkelmann, *Florida Institute of Technology*  
 Gary L. Wood, *Valdosta State University*  
 Troy D. Wood, *University of Buffalo*  
 Servet M. Yatin, *Quincy College*  
 Kazushige Yokoyama, *SUNY Geneseo*  
 Lin Zhu, *IUPUI*  
 Jim Zubricky, *University of Toledo*

# What's New in This Edition?

The book has been extensively revised and contains more small changes than can be detailed here. The most significant changes to the book and its supplements are listed below:

## NEW KEY CONCEPT INTERACTIVES

50 new *Key Concept Interactives (KCIs)* have been added to the eText and are assignable in Mastering Chemistry. Each interactive guides a student through a key topic as they navigate through a series of interactive screens. As they work through the KCI, they are presented with questions that must be answered to progress. Wrong answers result in feedback to guide them toward success.



## NEW INTERACTIVE VIDEOS

15 new *Key Concept Videos (KCVs)* and 25 new *Interactive Worked Examples (IWEs)* have been added to the media package that accompanies the book. All videos are available within the eText and assignable in Mastering Chemistry. *The video library now contains approximately 250 interactive videos.* These tools are designed to help professors engage their students in active learning.

## NEW AND REVISED END-OF-CHAPTER PROBLEMS

**176 New End-of-Chapter questions** have been added throughout the book, and **378 have been revised.** Many new End-of-Chapter questions involve the interpretation of graphs and data. All new End-of-Chapter questions are assignable in Mastering Chemistry.

## NEW ONLINE PROBLEM SETS

Online problem sets are web-based, online-only problems that are algorithmically randomized. They provide answer-specific feedback and will be continually updated and expanded.

## REVISED DATA

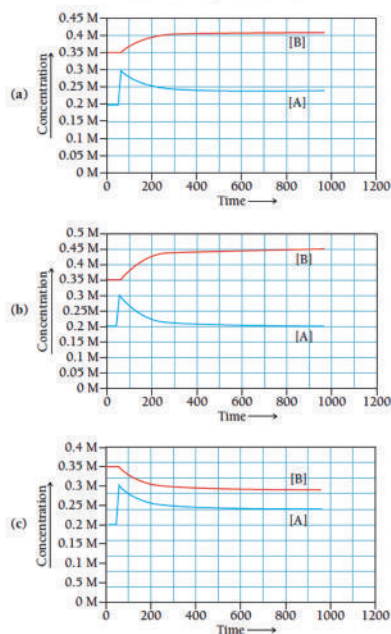
All the data throughout the book have been updated to reflect the most recent measurements available. These updates include Figure 4.2: *Carbon Dioxide Concentrations in the Atmosphere*; Figure 4.3: *Global Temperature*; the unnumbered figure in Section 7.10 of *U.S. Energy Consumption by Sector*; Figure 7.12: *Energy Consumption by Source*; Table 7.6: *Changes in National Average Pollutant Levels, 1990–2019*.

## DIVERSITY, EQUITY, AND INCLUSION REVIEW

As mentioned in the Preface, the entire book went through a detailed review to ensure the content reflects the rich diversity of our learners and is inclusive of their lived experiences.

### Le Châtelier's Principle


65. The reaction  $A(g) \rightarrow B(g)$  was at equilibrium and then disturbed by adding additional A to the reaction mixture. The reaction is then allowed to come back to equilibrium. Which graph represents the concentrations of A and B during this process?  
**MISSED THIS?** Read Section 16.9; Watch KCV 16.9





# CHEMISTRY

A MOLECULAR APPROACH



The most  
incomprehensible thing  
about the universe  
is that it is  
comprehensible.

—Albert Einstein (1879–1955)

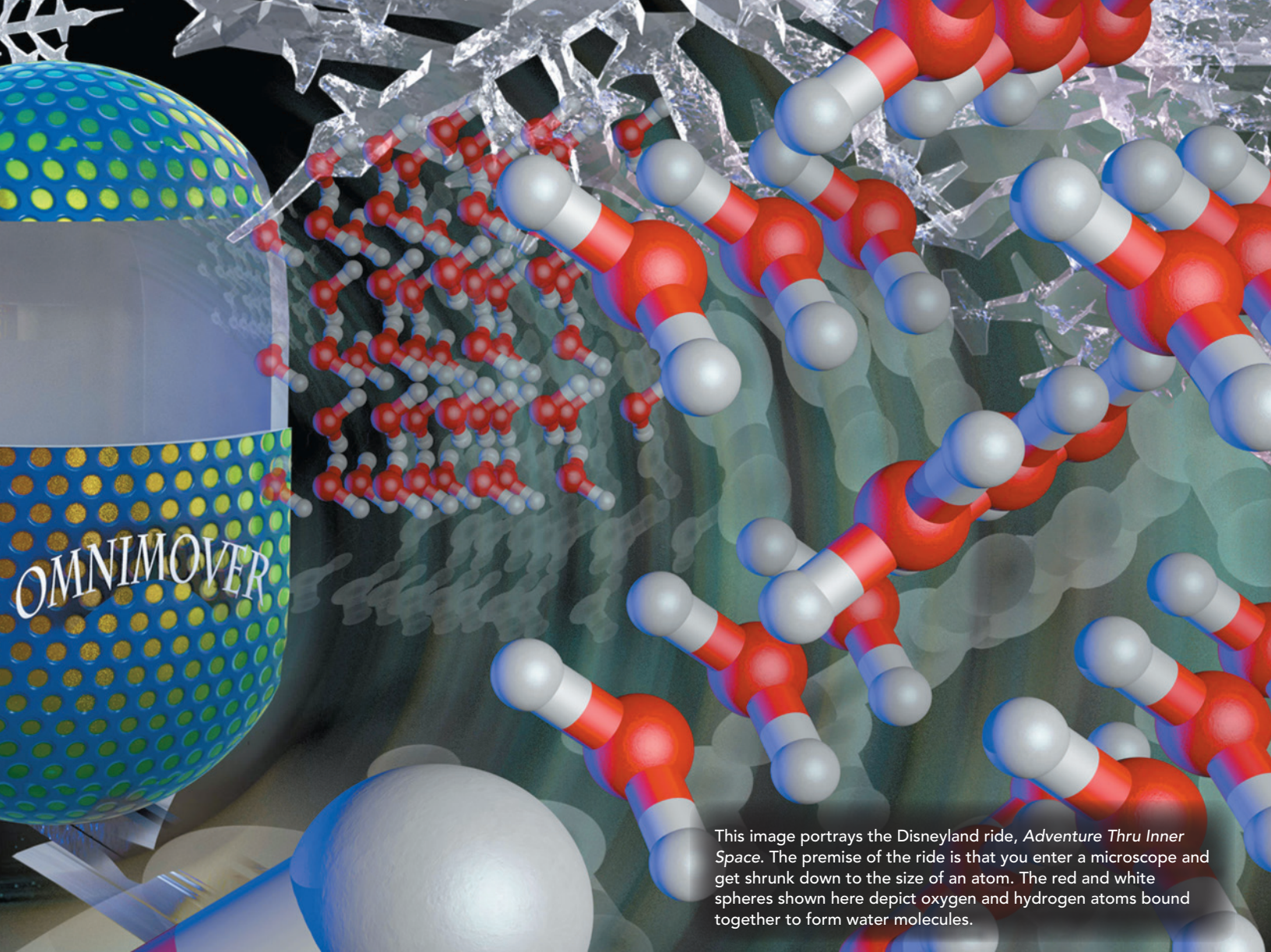
C H A P T E R

1

## Matter, Measurement, and Problem Solving

**W**hat do you think is the most important idea in all of human knowledge? This question has many possible answers—some practical, some philosophical, and some scientific. If we limit ourselves to scientific answers, mine would be this: **the properties of matter are determined by the properties of atoms and molecules.** Atoms and molecules determine how matter behaves—if they were different, matter would be different. The properties of water molecules determine how water behaves, the properties of sugar molecules determine how sugar behaves, and the properties of the molecules that compose our bodies determine how our bodies behave. The understanding of matter at the molecular level gives us unprecedented control over that matter. For example, our understanding of the details of the molecules that compose living organisms has revolutionized biology over the last 50 years.





This image portrays the Disneyland ride, *Adventure Thru Inner Space*. The premise of the ride is that you enter a microscope and get shrunk down to the size of an atom. The red and white spheres shown here depict oxygen and hydrogen atoms bound together to form water molecules.

- |   |  |
|---|--|
| <b>1.1</b> Atoms and Molecules 1  | <b>1.6</b> The Units of Measurement 13         |
| <b>1.2</b> The Scientific Approach to Knowledge 3                               | <b>1.7</b> The Reliability of a Measurement 20 |
| <b>1.3</b> The Classification of Matter 5                                       | <b>1.8</b> Solving Chemical Problems 26        |
| <b>1.4</b> Physical and Chemical Changes and Physical and Chemical Properties 9 | <b>1.9</b> Analyzing and Interpreting Data 33  |
| <b>1.5</b> Energy: A Fundamental Part of Physical and Chemical Change 12        | <b>LEARNING OUTCOMES</b> 38                    |

## 1.1 Atoms and Molecules

As I sat in the “omnimover” and listened to the narrator’s voice telling me that I was shrinking down to the size of an atom, I grew apprehensive but curious. Just minutes before, while waiting in line, I witnessed what appeared to be full-sized humans entering a microscope and emerging from the other end many times smaller. I was seven years old, and I was about to ride *Adventure Thru Inner Space*, a Disneyland ride (in Tomorrowland) that simulated what it would be like to shrink to the size of an atom. The ride



**KEY CONCEPT  
VIDEO 1.1**  
Atoms and  
Molecules

**WATCH NOW!**

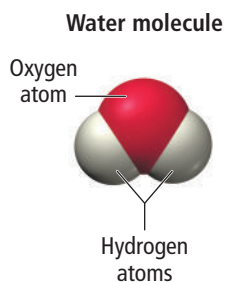




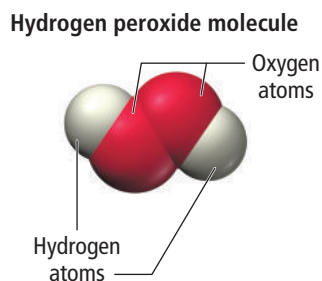
began with darkness and shaking, but then the shaking stopped and giant snowflakes appeared. The narrator explained that you were in the process of shrinking to an ever-smaller size (which explains why the snowflakes grew larger and larger). Soon, you entered the wall of the snowflake itself and began to see water molecules all around you. These also grew larger as you continued your journey into inner space and eventually ended up within the atom itself. Although this Disneyland ride bordered on being corny, and although it has since been shut down, it was my favorite ride as a young child.

That ride sparked my interest in the world of atoms and molecules, an interest that has continued and grown to this day. I am a chemist because I am obsessed with the connection between the “stuff” around us and the atoms and molecules that compose that stuff. More specifically, I love the idea that we humans have been able to figure out the connection between the *properties of the stuff* around us and the *properties of atoms and molecules*. **Atoms** are submicroscopic particles that are the fundamental building blocks of ordinary matter. Free atoms are rare in nature; instead they bind together in specific geometrical arrangements to form **molecules**. A good example of a molecule is the water molecule, which I remember so well from the Disneyland ride.

A water molecule is composed of one oxygen atom bound to two hydrogen atoms in the shape shown at left. The exact properties of the water molecule—the atoms that compose it, the distances between those atoms, and the geometry of how the atoms are bound together—determine the properties of water. If the molecule were different, water would be different. For example, if water contained two oxygen atoms instead of just one, it would be a molecule like this:



The hydrogen peroxide we use as an antiseptic or bleaching agent is considerably diluted.



This molecule is hydrogen peroxide, which you may have encountered if you have ever bleached your hair. A hydrogen peroxide molecule is composed of *two* oxygen atoms and two hydrogen atoms. This seemingly small molecular difference results in a huge difference in the properties of water and hydrogen peroxide. Water is the familiar and stable liquid we all drink and bathe in. Hydrogen peroxide, in contrast, is an unstable liquid that, in its pure form, burns the skin on contact and is used in rocket fuel. When you pour water onto your hair, your hair simply becomes wet. However, if you put diluted hydrogen peroxide on your hair, a chemical reaction occurs that strips your hair of its color.

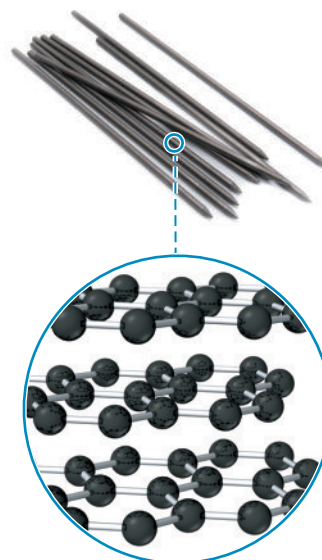
The details of how specific atoms bond to form a molecule—in a straight line, at a particular angle, in a ring, or in some other pattern—as well as the type of atoms in the molecule, determine everything about the substance that the molecule composes. If we want to understand the substances around us, we must understand the atoms and molecules that compose them—this is the central goal of chemistry. A good simple definition of **chemistry** is

**Chemistry—the science that seeks to understand the behavior of matter by studying the behavior of atoms and molecules.**

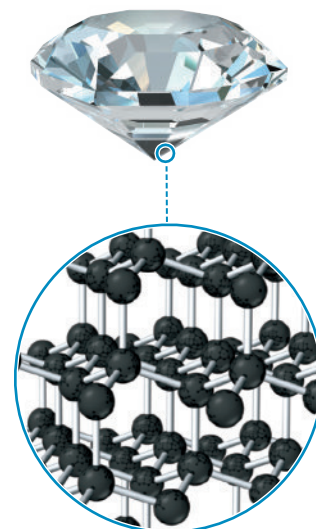
Throughout this book, we explore the connection between atoms and molecules and the matter they compose. We seek to understand how differences on the atomic or molecular level affect the properties on the macroscopic level. Before we move on, let's examine one more example that demonstrates this principle. Consider the structures of graphite and diamond.

The term *atoms* in this definition can be interpreted loosely to include atoms that have lost or gained electrons.

Graphite is the slippery black substance (often called pencil lead) that you have probably used in a mechanical pencil. Diamond is the brilliant gemstone found in jewelry. Graphite and diamond are both composed of exactly the same atoms—carbon atoms. The striking differences between the substances are a result of how those atoms are arranged. In graphite, the atoms are arranged in sheets. The atoms within each sheet are tightly bound to each other, but the sheets are *not* tightly bound to other sheets. Therefore the sheets can slide past each other, which is why the graphite in a pencil leaves a trail as you write. In diamond, by contrast, the carbon atoms are all bound together in a three-dimensional structure where layers are strongly bound to other layers, resulting in the strong, nearly unbreakable substance. This example illustrates how even the same atoms can compose vastly different substances when they are bound together in different patterns. Such is the atomic and molecular world—small differences in atoms and molecules can result in large differences in the substances that they compose.



Graphite structure



Diamond structure

## 1.2 The Scientific Approach to Knowledge

Throughout history, humans have approached knowledge about the physical world in different ways. For example, the Greek philosopher Plato (427–347 B.C.E.) thought that the best way to learn about reality was—not through the senses—but through reason. He believed that the physical world was an imperfect representation of a perfect and transcendent world (a world beyond space and time). For him, true knowledge came, not through observing the real physical world, but through reasoning and thinking about the ideal one.

The *scientific* approach to knowledge, however, is exactly the opposite of Plato's. Scientific knowledge is empirical—it is based on *observation* and *experiment*. Scientists observe and perform experiments on the physical world to learn about it. Some observations and experiments are qualitative (noting or describing how a process happens), but many are quantitative (measuring or quantifying something about the process). For example, Antoine Lavoisier (1743–1794), a French chemist who studied combustion (burning), made careful measurements of the mass of objects before and after burning them in closed containers. He noticed that there was no change in the total mass of material within the container during combustion. In doing so, Lavoisier made an important *observation* about the physical world.

Observations often lead scientists to formulate a **hypothesis**, a tentative interpretation or explanation of the observations. For example, Lavoisier explained his observations on combustion by hypothesizing that when a substance burns, it combines with a component of air. A good hypothesis is *falsifiable*, which means that it makes predictions that can be confirmed or refuted by further observations. Scientists test hypotheses by **experiments**, highly controlled procedures designed to generate observations that confirm or refute a hypothesis. The results of an experiment may support a hypothesis or prove it wrong—in which case the scientist must modify or discard the hypothesis.

In some cases, a series of similar observations leads to the development of a **scientific law**, a brief statement that summarizes past observations and predicts future ones. Lavoisier summarized his observations on combustion with the **law of conservation of mass**, which states, “In a chemical reaction, matter is neither created nor destroyed.” This statement summarized his observations on chemical reactions and predicted the outcome of future observations on reactions. Laws, like hypotheses, are also subject to experiments, which can support them or prove them wrong.

Although some Greek philosophers, such as Aristotle, did use observation to attain knowledge, they did not emphasize experiment and measurement to the extent that modern science does.



▲ French chemist Antoine Lavoisier with his wife, Marie, who helped him in his work by illustrating his experiments and translating scientific articles from English. Lavoisier, who also made significant contributions to agriculture, industry, education, and government administration, was executed during the French Revolution. (The Metropolitan Museum of Art)

Scientific laws are not *laws* in the same sense as civil or governmental laws. Nature does not follow laws in the way that we obey the laws against speeding or running a stop sign. Rather, scientific laws *describe* how nature behaves—they are generalizations about what nature does. For that reason, some people find it more appropriate to refer to them as *principles* rather than *laws*.

One or more well-established hypotheses may form the basis for a scientific **theory**. A scientific theory is a model for the way nature is and tries to explain not merely what nature does but why. As such, well-established theories are the pinnacle of scientific knowledge, often predicting behavior far beyond the observations or laws from which they were developed. A good example of a theory is the **atomic theory** proposed by English chemist John Dalton (1766–1844). Dalton explained the law of conservation of mass, as well as other laws and observations of the time, by proposing that matter is composed of small, indestructible particles called atoms. Since these particles are merely rearranged in chemical changes (and not created or destroyed), the total amount of mass remains the same. Dalton's theory is a model for the physical world—it gives us insight into how nature works and, therefore, *explains* our laws and observations.

Finally, the scientific approach returns to observation to test theories. For example, scientists can test the atomic theory by trying to isolate single atoms or by trying to image them (both of which, by the way, have already been accomplished). Theories are validated by experiments; however, theories can never be conclusively proven because some new observation or experiment always has the potential to reveal a flaw. Notice that the scientific approach to knowledge begins with observation and ends with observation. An experiment is in essence a highly controlled procedure for generating critical observations designed to test a theory or hypothesis. Each new set of observations has the potential to refine the original model. Figure 1.1▼ summarizes one way to map the scientific approach to knowledge. Scientific laws, hypotheses, and theories are all subject to continued experimentation. If a law, hypothesis, or theory is proved wrong by an experiment, it must be revised and tested with new experiments. Over time, the scientific community eliminates or corrects poor theories and laws, and valid theories and laws—those consistent with experimental results—remain.

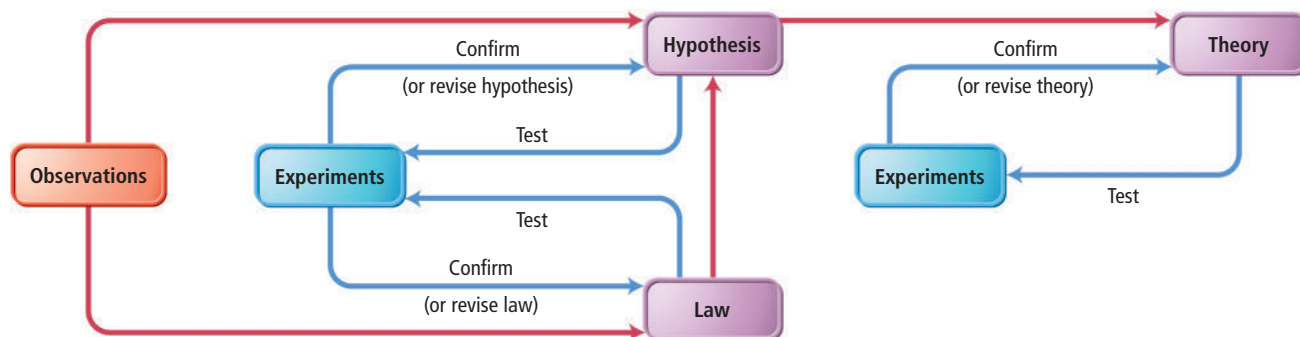
Established theories with strong experimental support are the most powerful pieces of scientific knowledge. You may have heard the phrase “That is just a theory,” as if theories are easily dismissible. Such a statement reveals a deep misunderstanding of the nature of a scientific theory. Well-established theories are as close to truth as we get in science. The idea that all matter is made of atoms is “just a theory,” but it has over 200 years of experimental evidence to support it. It is a powerful piece of scientific knowledge on which many other scientific ideas are based.

One last word about the scientific approach to knowledge: some people wrongly imagine science to be a strict set of rules and procedures that automatically leads to inarguable, objective facts. This is not the case. Even our diagram of the scientific approach to knowledge is only an idealization of real science, useful to help us see the key distinctions of science. Real science requires hard work, care, creativity, and even a bit of luck.

In Dalton's time, people thought atoms were indestructible. Today, because of nuclear reactions, we know that atoms can be broken apart into their smaller components.

▼ FIGURE 1.1 The Scientific Approach to Knowledge

### The Scientific Approach





Scientific theories do not just arise out of data—individuals of genius and creativity craft theories. A great theory is not unlike a master painting, and many see a similar kind of beauty in both. (For more on this aspect of science, see the accompanying box entitled *Thomas S. Kuhn and Scientific Revolutions*.)

**LAWS AND THEORIES** Which statement best explains the difference between a law and a theory?

- (a) A law is truth; a theory is mere speculation.
- (b) A law summarizes a series of related observations; a theory gives the underlying reasons for them.
- (c) A theory describes *what* nature does; a law describes *why* nature does it.

**1.1**  
**Cc**  
Conceptual  
Connection

ANSWER NOW!



## THE NATURE OF SCIENCE

### Thomas S. Kuhn and Scientific Revolutions

When scientists talk about science, they often talk in ways that imply that theories are “true.” Further, they talk as if they arrive at theories in logical and unbiased ways. For example, a theory central to chemistry that we have discussed in this chapter is John Dalton’s atomic theory—the idea that all matter is composed of atoms. Is this theory “true”? Was it reached in logical, unbiased ways? Will this theory still be around in 200 years?

The answers to these questions depend on how we view science and its development. One way to view science—let’s call it the *traditional view*—is as the continual accumulation of knowledge and the building of increasingly precise theories. In this view, a scientific theory is a model of the world that reflects what is *actually in* nature. New observations and experiments result in gradual adjustments to theories. Over time, theories get better, giving us a more accurate picture of the physical world.

In the twentieth century, a different view of scientific knowledge began to develop. A book by Thomas Kuhn (1922–1996), published in 1962 and entitled *The Structure of Scientific Revolutions*, challenged the traditional view. Kuhn’s ideas came from his study of the history of science, which, he argued, does not support the idea that science progresses in a smooth, cumulative way. According to Kuhn, science goes through fairly quiet periods that he called *normal science*. In these periods, scientists make their data fit the reigning theory, or paradigm. Small inconsistencies are swept aside during periods of normal science. However, when too many inconsistencies and anomalies develop, a crisis emerges. The crisis brings about a *revolution* and a new reigning theory. According to Kuhn, the new theory is usually quite different from

the old one; it not only helps us to make sense of new or anomalous information, but it also enables us to see accumulated data from the past in a dramatically new way.

Kuhn further contended that theories are held for reasons that are not always logical or unbiased, and that theories are not *true* models—in the sense of a one-to-one mapping—of the physical world. Because new theories are often so different from the ones they replace, he argued, and because old theories always make good sense to those holding them, they must not be “True” with a capital T; otherwise “truth” would be constantly changing.

Kuhn’s ideas created a controversy among scientists and science historians that continues to this day. Some, especially postmodern philosophers of science, have taken Kuhn’s ideas one step further. They argue that scientific knowledge is *completely* biased and lacks any objectivity. Most scientists, including Kuhn, would disagree. Although Kuhn pointed out that scientific knowledge has *arbitrary elements*, he also said, “*Observation . . . can and must drastically restrict the range of admissible scientific belief, else there would be no science.*” In other words, saying that science contains arbitrary elements is quite different from saying that science itself is arbitrary.

**QUESTION** In his book, Kuhn stated, “A new theory . . . is seldom or never just an increment to what is already known.” From your knowledge of the history of science, can you think of any examples that support Kuhn’s statement? Do you know of any instances in which a new theory or model was drastically different from the one it replaced?

## 1.3

## The Classification of Matter

**Matter** is anything that occupies space and has mass. Your desk, your chair, and even your body are all composed of matter. Less obviously, the air around you is also matter—it too occupies space and has mass. We call a specific instance of matter—such as air, water, or sand—a **substance**. We classify matter according to its **state** (its physical form) and its **composition** (the basic components that make it up).



WATCH NOW!

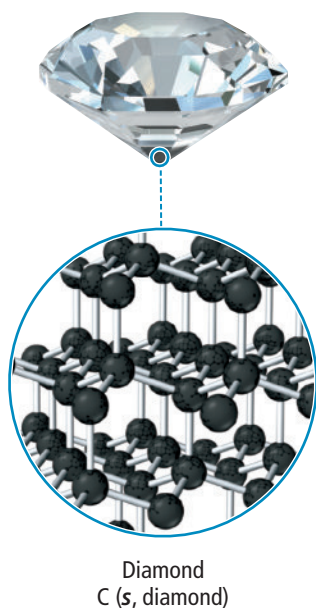
**KEY CONCEPT  
VIDEO 1.3**  
Classifying Matter



The state of matter changes from solid to liquid to gas with increasing temperature.

Glass and other amorphous solids can be thought of, from one point of view, as intermediate between solids and liquids. Their atoms are fixed in position at room temperature, but they have no long-range structure and do not have distinct melting points.

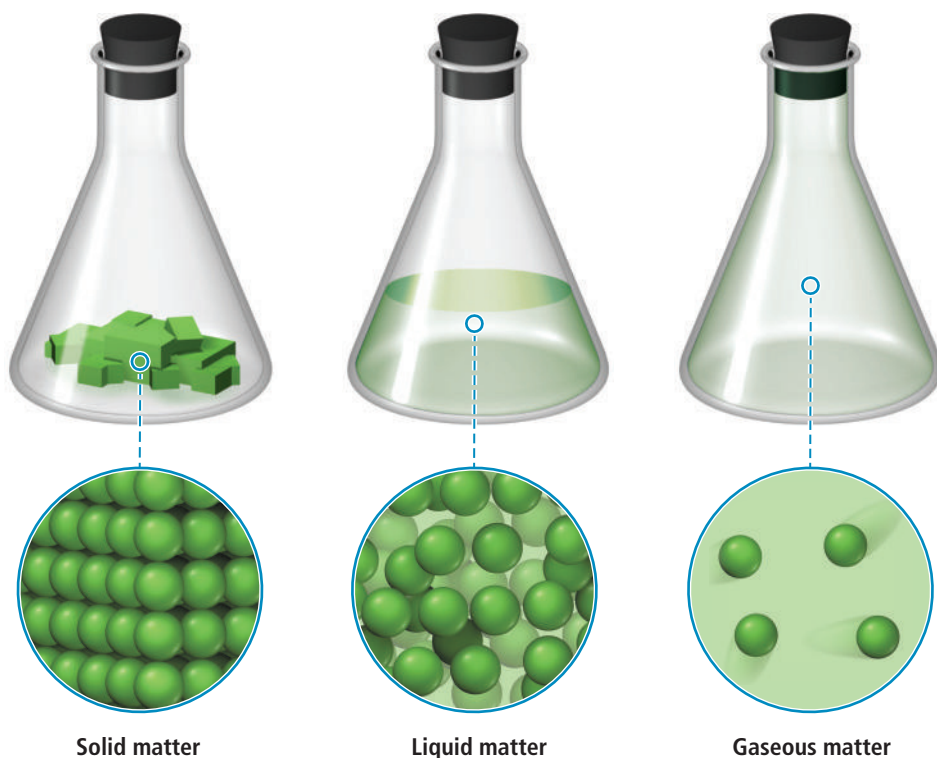
**Crystalline Solid:**  
Atoms are arranged in a regular three-dimensional pattern.



▲ **FIGURE 1.2 Crystalline Solid** Diamond (first discussed in Section 1.1) is a crystalline solid composed of carbon atoms arranged in a regular, repeating pattern.

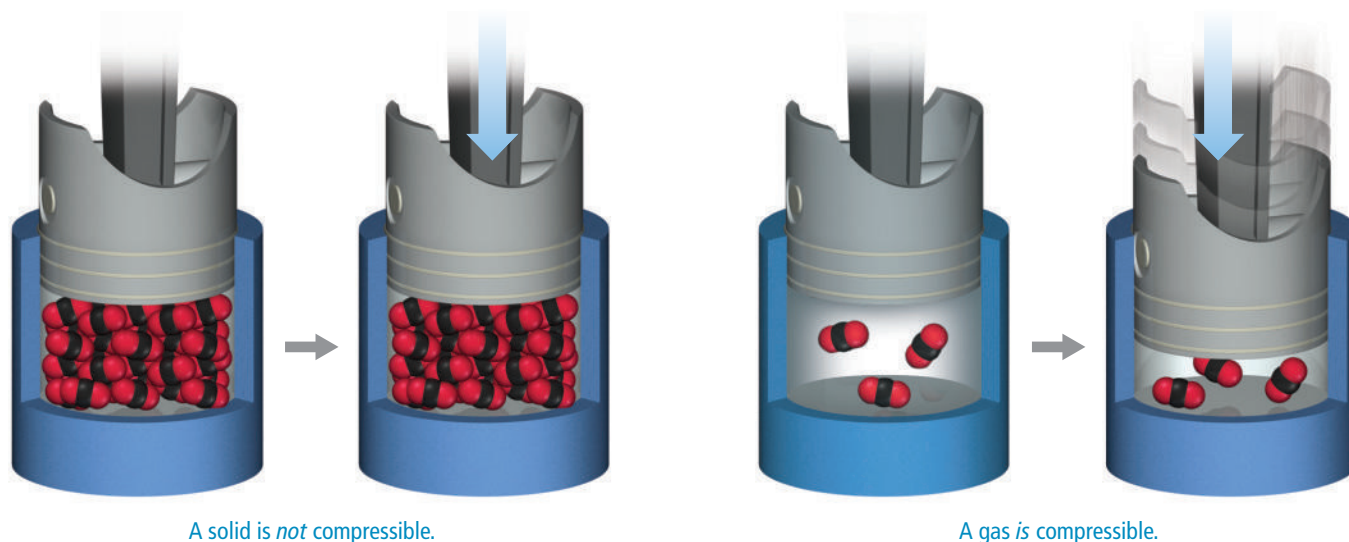
## The States of Matter: Solid, Liquid, and Gas

Matter exists in three different states: **solid**, **liquid**, and **gas**. In *solid matter*, atoms or molecules pack closely to each other in fixed locations. Although the atoms and molecules in a solid vibrate, they do not move around or past each other. Consequently, a solid has a fixed volume and rigid shape. Ice, aluminum, and diamond are examples of solids. Solid matter may be **crystalline**, in which case its atoms or molecules are in patterns with long-range, repeating order (Figure 1.2▼), or it may be **amorphous**, in which case its atoms or molecules do not have any long-range order. Table salt and diamond are examples of *crystalline* solids; the well-ordered geometric shapes of salt and diamond crystals reflect the well-ordered geometric arrangement of their atoms (although this is not the case for *all* crystalline solids). Examples of *amorphous* solids include glass and plastic. In *liquid matter*, atoms or molecules pack about as closely as they do in solid matter, but they are free to move relative to each other, giving liquids a fixed volume but not a fixed shape. Liquids assume the shape of their containers. Water, alcohol, and gasoline are all substances that are liquids at room temperature.



▲ In a solid, the atoms or molecules are fixed in place and can only vibrate. In a liquid, although the atoms or molecules are closely packed, they can move past one another, allowing the liquid to flow and assume the shape of its container. In a gas, the atoms or molecules are widely spaced, making gases compressible as well as fluid (able to flow).

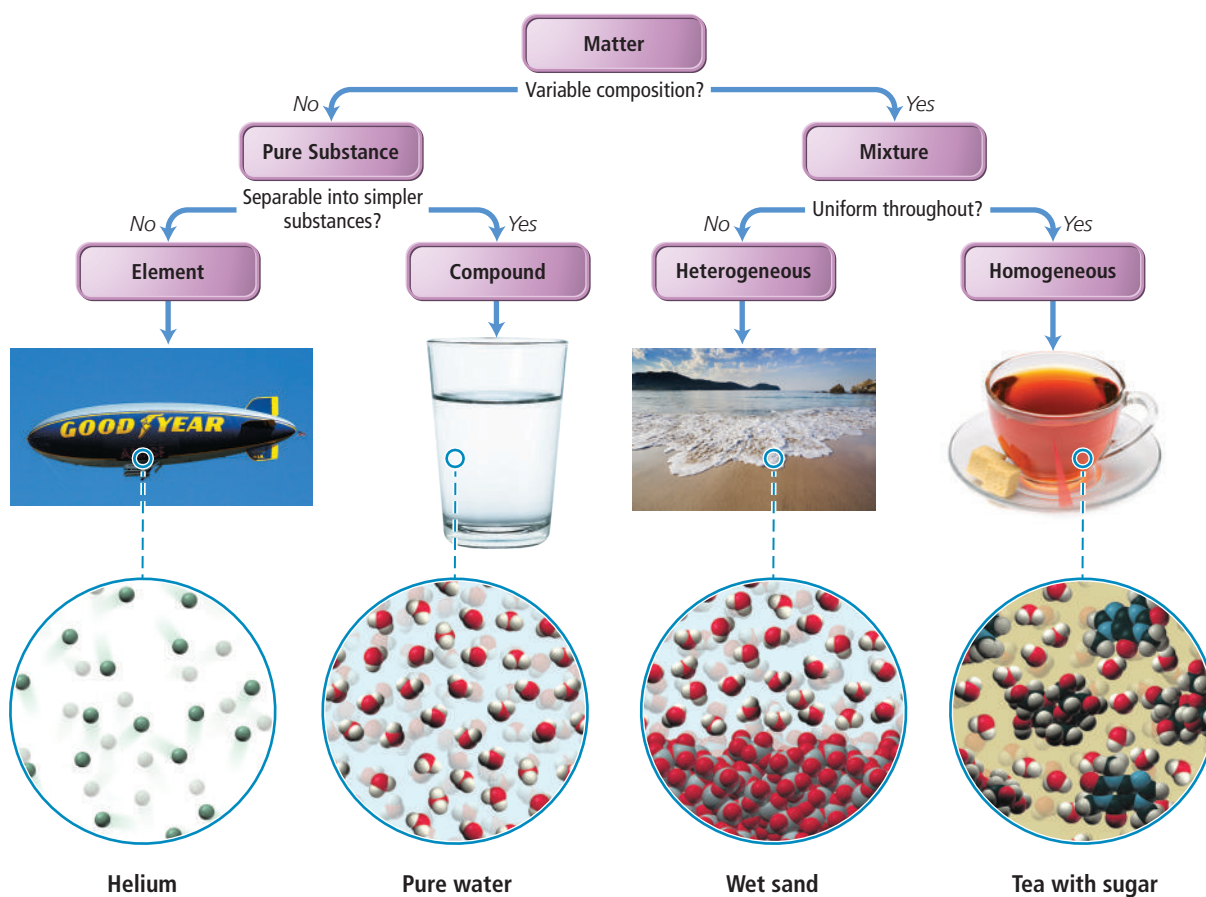
In *gaseous matter*, atoms or molecules have a lot of space between them and are free to move relative to one another, making gases *compressible* (Figure 1.3▶). When you squeeze a balloon or sit down on an air mattress, you force the atoms and molecules into a smaller space so that they are closer together. Gases always assume the shape *and* volume of their containers. Substances that are gases at room temperature include helium, nitrogen (the main component of air), and carbon dioxide.



▲ **FIGURE 1.3** The **Compressibility of Gases** Gases can be compressed—squeezed into a smaller volume—because there is so much empty space between atoms or molecules in the gaseous state.

## Classifying Matter by Composition: Elements, Compounds, and Mixtures

In addition to classifying matter according to its state, we classify it according to its composition, as shown in the following chart:





The first division in the classification of matter is between a *pure substance* and a *mixture*. A **pure substance** is made up of only one component, and its composition is invariant (it does not vary from one sample to another). The *components* of a pure substance can be individual atoms or groups of atoms joined together. For example, helium, water, and table salt (sodium chloride) are all pure substances. Each of these substances is made up of only one component: helium is made up of helium atoms, water is made up of water molecules, and sodium chloride is made up of sodium chloride units. The composition of a pure sample of any one of these substances is always exactly the same (because you can't vary the composition of a substance made up of only one component).

A **mixture**, by contrast, is composed of two or more components in proportions that can vary from one sample to another. For example, sweetened tea, composed primarily of water molecules and sugar molecules (with a few other substances mixed in), is a mixture. We can make tea slightly sweet (a small proportion of sugar to water) or very sweet (a large proportion of sugar to water) or any level of sweetness in between.

We categorize pure substances themselves into two types—*elements* and *compounds*—depending on whether or not they can be broken down (or decomposed) into simpler substances. Helium, which we just noted is a pure substance, is also a good example of an **element**, a substance that cannot be chemically broken down into simpler substances. Water, also a pure substance, is a good example of a **compound**, a substance composed of two or more elements (in this case, hydrogen and oxygen) in a fixed, definite proportion. On Earth, compounds are more common than pure elements because most elements combine with other elements to form compounds.

We also categorize mixtures into two types—heterogeneous and homogeneous—depending on how *uniformly* the substances within them mix. Wet sand is a **heterogeneous mixture**, one in which the composition varies from one region of the mixture to another. Sweetened tea is a **homogeneous mixture**, one with the same composition throughout. Homogeneous mixtures have uniform compositions because the atoms or molecules that compose them mix uniformly. Heterogeneous mixtures are made up of distinct regions because the atoms or molecules that compose them separate. Here again we see that the properties of matter are determined by the atoms or molecules that compose it.

Classifying a substance according to its composition is not always obvious and requires that we either know the true composition of the substance or are able to test it in a laboratory. For now, we focus on relatively common substances that you are likely to have encountered. Throughout this course, you will gain the knowledge to understand the composition of a larger variety of substances.

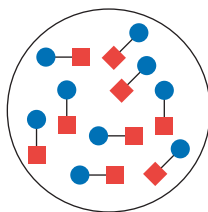
All known elements are listed in the periodic table in the inside front cover of this book.

ANSWER NOW!

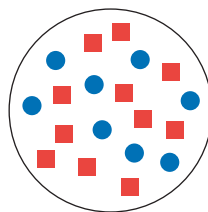


**1.2**  
**Cc**  
Conceptual  
Connection

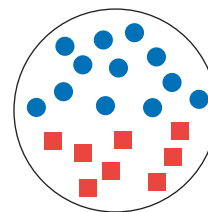
**PURE SUBSTANCES AND MIXTURES** In these images, a blue circle represents an atom of one type of element, and a red square represents an atom of a second type of element. Which image is a pure substance?



(a)



(b)



(c)

None of these  
(d)

## Separating Mixtures

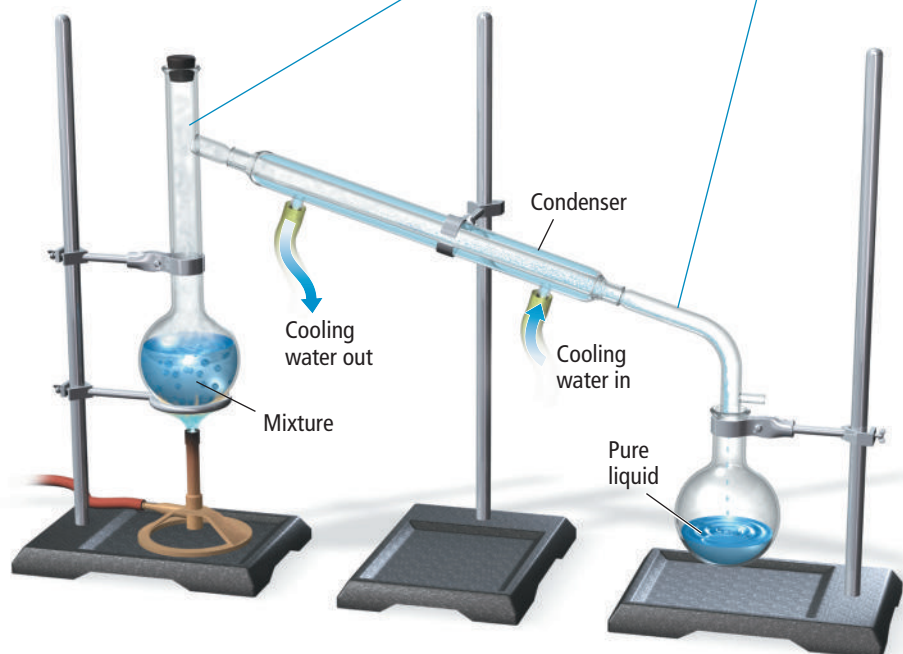
Chemists often want to separate a mixture into its components. Such separations can be easy or difficult, depending on the components in the mixture. In general, mixtures are separable because the different components have different physical or chemical properties. We can use various techniques that exploit these differences to achieve

## Distillation

When a mixture of liquids with different boiling points is heated...

... the most volatile component boils first.

The vapor is then cooled and collected as pure liquid.



▲ FIGURE 1.4 Separating Substances by Distillation

separation. For example, we can separate a mixture of sand and water by **decanting**—carefully pouring off—the water into another container. A homogeneous mixture of liquids can usually be separated by **distillation**, a process in which the mixture is heated to boil off the more **volatile** (easily vaporizable) liquid. The volatile liquid is then recondensed in a condenser and collected in a separate flask (Figure 1.4▲). If a mixture is composed of an insoluble solid and a liquid, we can separate the two by **filtration**, in which the mixture is poured through filter paper in a funnel (Figure 1.5▲).

## Filtration

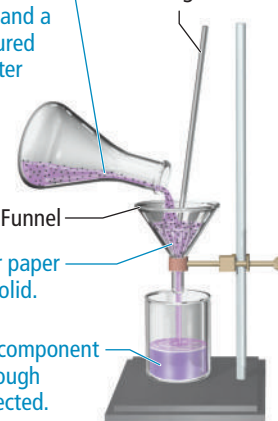
When a mixture of a liquid and a solid is poured through filter paper...

Stirring rod

Funnel

... the filter paper traps the solid.

The liquid component passes through and is collected.



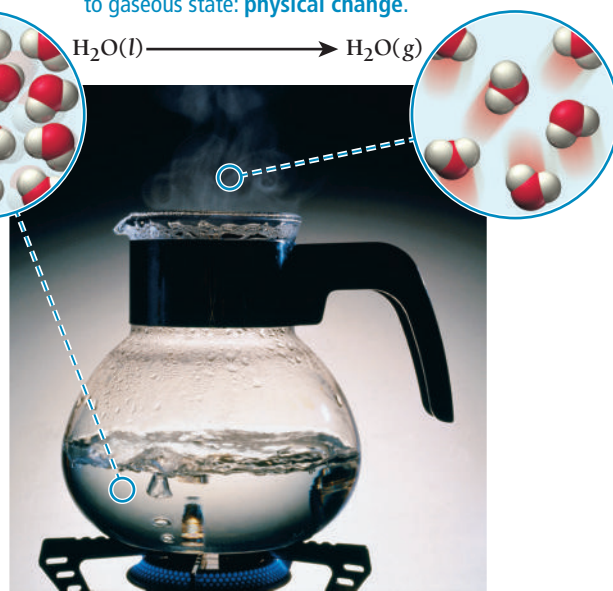
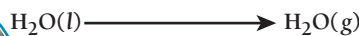
▲ FIGURE 1.5 Separating Substances by Filtration

## 1.4

## Physical and Chemical Changes and Physical and Chemical Properties

Every day we witness changes in matter: ice melts, iron rusts, gasoline burns, fruit ripens, and water evaporates. What happens to the molecules or atoms that compose these substances during such changes? The answer depends on the type of change. Changes that alter only state or appearance, but not composition, are **physical changes**. The atoms or molecules that compose a substance *do not change* their identity during a physical change. For example, when water boils, it changes its state from a liquid to a gas, but the gas remains composed of water molecules, so this is a physical change (Figure 1.6▲).

Water molecules change from liquid to gaseous state: **physical change**.



▲ FIGURE 1.6 Boiling, a Physical Change When water boils, it turns into a gas but does not alter its chemical identity—the water molecules are the same in both the liquid and gaseous states. Boiling is a physical change, and the boiling point of water is a physical property.

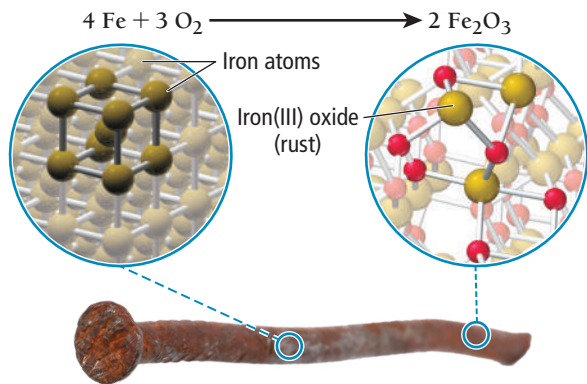
A physical change results in a different form of the same substance, while a chemical change results in a completely different substance.

In contrast, changes that alter the composition of matter are **chemical changes**. During a chemical change, atoms rearrange, transforming the original substances into different substances. For example, the rusting of iron is a chemical change. The atoms that compose iron (iron atoms) combine with oxygen molecules from air to form iron oxide, the orange substance we call rust (Figure 1.7◀). Figure 1.8▶ illustrates other examples of physical and chemical changes.

Physical and chemical changes are manifestations of physical and chemical properties. A **physical property** is a property that a substance displays without changing its composition, whereas a **chemical property** is a property that a substance displays only by changing its composition via a chemical change. The smell of gasoline is a physical property—gasoline does not change its composition when it exhibits its odor. The flammability of gasoline, in contrast, is a chemical property—gasoline does change its composition when it burns, turning into completely new substances (primarily carbon dioxide and water). Physical properties include odor, taste, color, appearance, melting point, boiling point, and density. Chemical properties include corrosiveness, flammability, acidity, toxicity, and other such characteristics.

The differences between physical and chemical changes are not always apparent. Only chemical examination can confirm whether a particular change is physical or chemical. In many cases, however, we can identify chemical and physical changes based on what we know about the changes. Changes in the state of matter, such as melting or boiling, or changes in the physical condition of matter, such as those that result from cutting or crushing, are typically physical changes. Changes involving chemical reactions—often evidenced by temperature or color changes—are chemical changes.

Iron combines with oxygen to form iron(III) oxide: **chemical change**.



▲ **FIGURE 1.7 Rusting, a Chemical Change** When iron rusts, the iron atoms combine with oxygen atoms to form a different chemical substance, the compound iron(III) oxide. Rusting is a chemical change, and the tendency of iron to rust is a chemical property. A more detailed exploration of this reaction can be found in Section 20.9.

### EXAMPLE 1.1 Physical and Chemical Changes and Properties

Determine whether each change is physical or chemical. What kind of property (chemical or physical) is demonstrated in each case?

- (a) the evaporation of rubbing alcohol
- (b) the burning of lamp oil
- (c) the bleaching of hair with hydrogen peroxide
- (d) the formation of frost on a cold night

#### SOLUTION

- (a) When rubbing alcohol evaporates, it changes from liquid to gas, but it remains alcohol—this is a physical change. The volatility (the ability to evaporate easily) of alcohol is therefore a physical property.
- (b) Lamp oil burns because it reacts with oxygen in air to form carbon dioxide and water—this is a chemical change. The flammability of lamp oil is therefore a chemical property.
- (c) Applying hydrogen peroxide to hair changes pigment molecules in hair that give it color—this is a chemical change. The susceptibility of hair to bleaching is therefore a chemical property.
- (d) Frost forms on a cold night because water vapor in air changes its state to form solid ice—this is a physical change. The temperature at which water freezes is therefore a physical property.

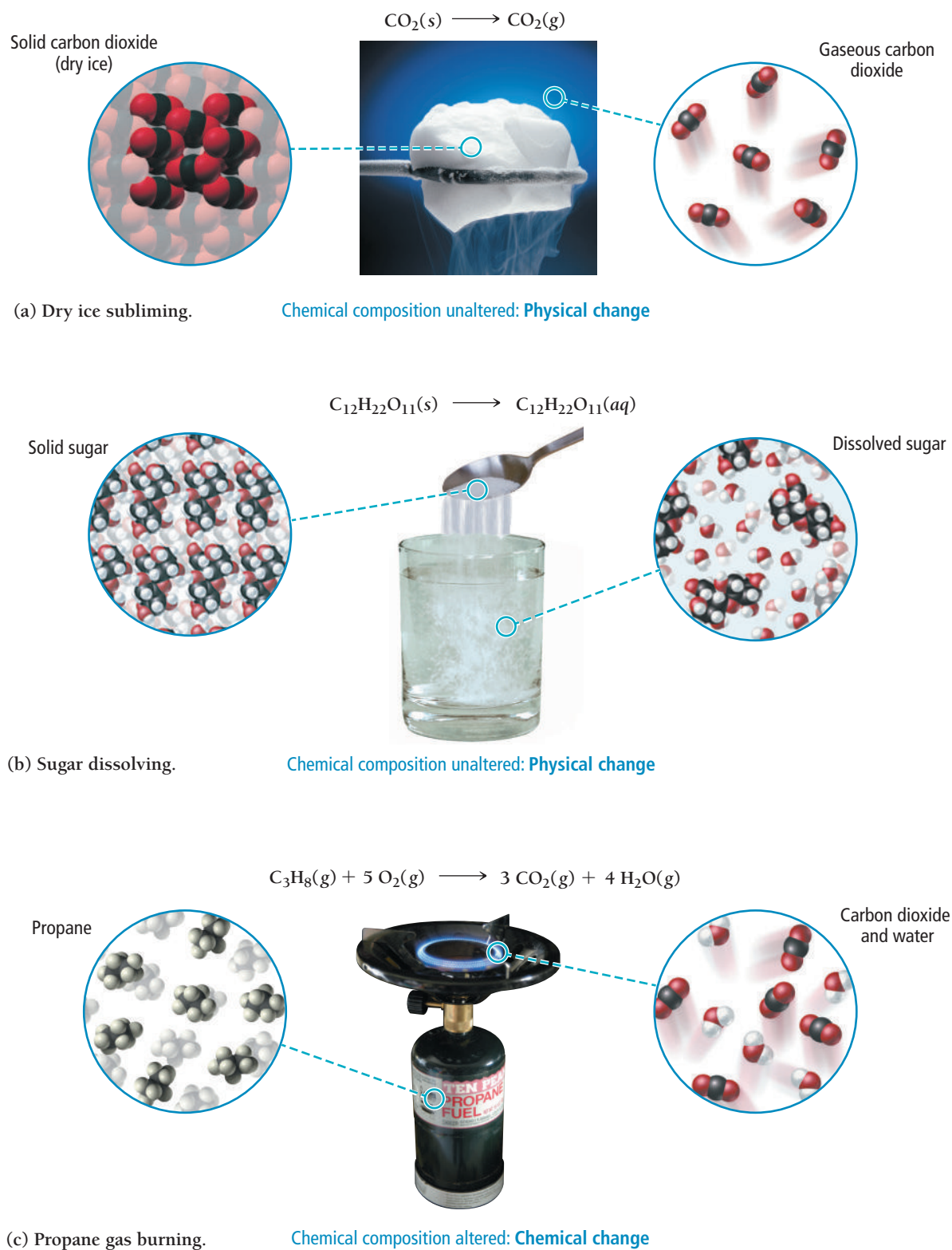
**FOR PRACTICE 1.1** Determine whether each change is physical or chemical. What kind of property (chemical or physical) is demonstrated in each case?

- (a) A copper wire is hammered flat.
- (b) A nickel dissolves in acid to form a blue-green solution.
- (c) Dry ice sublimates without melting.
- (d) A match ignites when struck on a flint.

Answers to For Practice and For More Practice problems can be found in Appendix IV.



## Physical Change versus Chemical Change



**▲ FIGURE 1.8 Physical and Chemical Changes** (a) The sublimation (the state change from a solid to a gas) of dry ice (solid  $\text{CO}_2$ ) is a physical change. (b) The dissolution of sugar is a physical change. (c) The burning of propane is a chemical change.

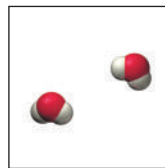
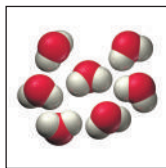
## ANSWER NOW!



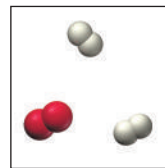
### 1.3

**Cc**  
Conceptual  
Connection

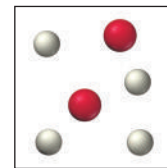
**CHEMICAL AND PHYSICAL CHANGES** The diagram on the left represents liquid water molecules in a pan. Which of the three diagrams (a, b, or c) best represents the water molecules after they have been vaporized by boiling?



(a)



(b)



(c)

## 1.5

## Energy: A Fundamental Part of Physical and Chemical Change

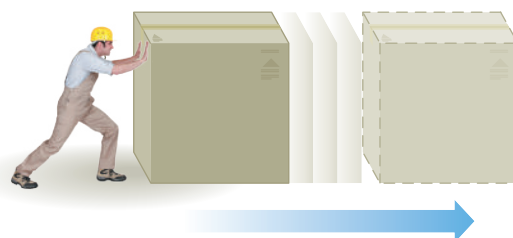
The physical and chemical changes discussed in Section 1.4 are usually accompanied by energy changes. For example, when water evaporates from your skin (a physical change), the water molecules absorb energy from your body, making you feel cooler. When you burn natural gas on the stove (a chemical change), energy is released, heating the food you are cooking. Understanding the physical and chemical changes of matter—that is, understanding chemistry—requires that you understand energy changes and energy flow.

The scientific definition of **energy** is *the capacity to do work*. **Work** is defined as the action of a force through a distance. For instance, when you push a box across the floor or pedal your bicycle across the street, you have done work.



### ▲ FIGURE 1.9 Energy

**Conversions** Gravitational potential energy is converted into kinetic energy when the weight is dropped. The kinetic energy is converted mostly to thermal energy when the weight strikes the ground.



Force acts through distance; work is done.

The **total energy** of an object is a sum of its **kinetic energy** (the energy associated with its motion) and its **potential energy** (the energy associated with its position or composition). For example, a weight held several meters above the ground has potential energy due to its position within Earth's gravitational field (Figure 1.9 ◀). If you drop the weight, it accelerates, and its potential energy is converted to kinetic energy. When the weight hits the ground, its kinetic energy is converted primarily to **thermal energy**, the energy associated with the temperature of an object. Thermal energy is actually a type of kinetic energy because it is associated with the motion of the individual atoms or molecules that make up an object. When the weight hits the ground, its kinetic energy is essentially transferred to the atoms and molecules that compose the ground, raising the temperature of the ground ever so slightly.

The first principle to note about how energy changes as the weight falls to the ground is that *energy is neither created nor destroyed*. The potential energy of the weight becomes kinetic energy as the weight accelerates toward the ground. The kinetic energy then becomes thermal energy when the weight hits the ground. The total amount of thermal energy that is released through the process is exactly equal to the difference between the initial and final potential energy of the weight. The idea that energy is neither created nor destroyed is known as the **law of conservation of energy**. Although energy can change from one type into another, and although it can flow from one object to another, the **total quantity** of energy does not change—it remains constant.

In Chapter 21 we will discuss how energy conservation is actually part of a more general law that allows for the interconvertibility of mass and energy.

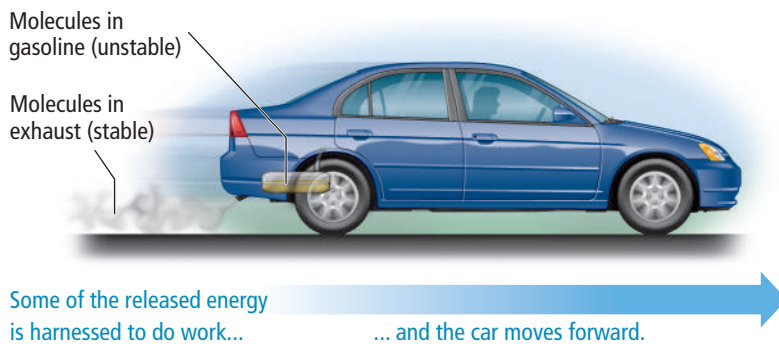
The second principle to note about the raised weight and its fall is *the tendency of systems with high potential energy to change in a way that lowers their potential energy*. For this reason, objects or systems with high potential energy tend to be *unstable*. The weight lifted several meters from the ground is unstable because it contains a significant amount of potential energy. Unless restrained, the weight will naturally fall, lowering its potential energy (due to its position in Earth's gravitational field). We can harness some of the raised weight's potential energy to do work. For example, we can attach the weight to a rope that turns a paddle wheel or spins a drill as the weight falls. After the weight falls to the ground, it contains less potential energy—it has become more *stable*.

Some chemical substances are like a raised weight. For example, the molecules that compose gasoline have a relatively high potential energy—energy is concentrated in them just as energy is concentrated in the raised weight. The molecules in the gasoline tend to undergo chemical changes (specifically combustion) that lower the molecules' potential energy. As the energy of the molecules is released, some of it can be harnessed to do work, such as moving a car forward (Figure 1.10▲). The molecules that result from the chemical change have less potential energy than the original molecules in gasoline and are more stable.

Chemical potential energy, such as the energy contained in the molecules that compose gasoline, arises primarily from electrostatic forces between the electrically charged particles (protons and electrons) that compose atoms and molecules. We will learn more about those particles, as well as the properties of electrical charge, in Chapter 2, but for now, know that molecules contain specific, usually complex, arrangements of these charged particles. Some of these arrangements—such as the one within the molecules that compose gasoline—have a much higher potential energy than others. When gasoline undergoes combustion, the arrangement of these particles changes, creating molecules with much lower potential energy and transferring a great deal of energy (mostly in the form of heat) to the surroundings.

#### Summarizing Energy:

- Energy is always conserved in a physical or chemical change; it is neither created nor destroyed.
- Systems with high potential energy tend to change in ways that lower their potential energy, transferring energy to the surroundings.



▲ **FIGURE 1.10 Using Chemical Energy to Do Work** The compounds produced when gasoline burns have less chemical potential energy than the gasoline molecules.

## ENERGY What type of energy is chemical energy?

- (a) kinetic energy
- (b) thermal energy
- (c) potential energy

**1.4**  
**Cc**  
Conceptual  
Connection

ANSWER NOW!



## 1.6

## The Units of Measurement

In 1999, NASA lost the \$125 million *Mars Climate Orbiter*. The chairman of the commission that investigated the disaster concluded, “The root cause of the loss of the spacecraft was a failed translation of English units into metric units.” As a result, the orbiter—which was supposed to monitor weather on Mars—descended too far into the Martian atmosphere and burned up. In chemistry as in space exploration, **units**—standard quantities used to specify measurements—are critical. If we get them wrong, the consequences can be disastrous.



WATCH NOW!

**KEY CONCEPT**  
**VIDEO 1.6**  
Units and  
Significant Figures

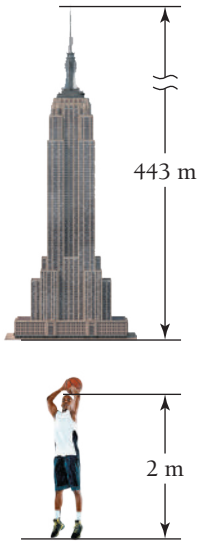






▲ The \$125 million *Mars Climate Orbiter* was lost in 1999 because two groups of engineers used different units.

The abbreviation *SI* comes from the French, *Système International d'Unités*.



▲ The Empire State Building is 443 m tall. A basketball player stands about 2 m tall.



▲ A nickel (5 cents) weighs about 5 g.

TABLE 1.1 SI Base Units		
Quantity	Unit	Symbol
Length	Meter	m
Mass	Kilogram	kg
Time	Second	s
Temperature	Kelvin	K
Amount of substance	Mole	mol
Electric current	Ampere	A
Luminous intensity	Candela	cd

The two most common unit systems are the **metric system**, used in most of the world, and the **English system**, used in the United States. Scientists use the **International System of Units (SI)**, which is based on the metric system.

### Standard Units

Table 1.1 shows the standard SI base units. In this chapter, we focus on the first four of these units: the *meter*, the standard unit of length; the *kilogram*, the standard unit of mass; the *second*, the standard unit of time; and the *kelvin*, the standard unit of temperature.

### The Meter: A Measure of Length

The **meter (m)** is slightly longer than a yard (1 yard is 36 inches, while 1 meter is 39.37 inches).



Thus, a 100-yard (yd) football field measures only 91.4 m. The meter was originally defined as 1/10,000,000 of the distance from the equator to the North Pole (through Paris). The International Bureau of Weights and Measures now defines it more precisely as the distance light travels through a vacuum in a certain period of time, 1/299,792,458 second. A tall human is about 2 m tall, and the Empire State Building stands 443 m tall (including its mast).

### The Kilogram: A Measure of Mass

The **kilogram (kg)**, defined relative to a fundamental constant called planck’s constant (*h*), is a measure of *mass*, a quantity different from *weight*. The **mass** of an object is a measure of the quantity of matter within it, while the weight of an object is a measure of the *gravitational pull* on its matter. If you could weigh yourself on the moon, for example, its weaker gravity would pull on you with less force than does Earth’s gravity, resulting in a lower weight. A 130-pound (lb) person on Earth would weigh only 21.5 lb on the moon. However, a person’s mass—the quantity of matter in his or her body—remains the same on every planet. One kilogram of mass is the equivalent of 2.205 lb of weight on Earth, so expressed in kilograms, a 130-lb person has a mass of approximately 59 kg and this book has a mass of about 2.5 kg. A second common unit of mass is the gram (g). One gram is 1/1000 kg. A nickel (5¢) has a mass of about 5 g.

### The Second: A Measure of Time

If you live in the United States, the **second (s)** is perhaps the SI unit most familiar to you. The International Bureau of Weights and Measures originally defined the second in terms of the day and the year, but a second is now defined more precisely as the duration of 9,192,631,770 periods of the radiation emitted from a certain transition in a cesium-133 atom. (We discuss transitions and the emission of radiation by atoms in Chapter 8.)

Scientists measure time on a large range of scales. The human heart beats about once every second, the age of the universe is estimated to be about  $4.32 \times 10^{17}$  s (13.7 billion years), and some molecular bonds break or form in time periods as short as  $1 \times 10^{-15}$  s.

## The Kelvin: A Measure of Temperature

The **kelvin (K)** is the SI unit of **temperature**. The temperature of a sample of matter is a measure of the average kinetic energy—the energy due to motion—of the atoms or molecules that compose the matter. The molecules in a *hot* glass of water are, on average, moving faster than the molecules in a *cold* glass of water. Temperature is a measure of this molecular motion.

Temperature also determines the direction of thermal energy transfer, what we commonly call *heat*. Thermal energy transfers from hot objects to cold ones. For example, when you touch another person's warm hand (and yours is cold), thermal energy flows *from his or her hand to yours*, making your hand feel warmer. However, if you touch an ice cube, thermal energy flows *out of your hand* to the ice, cooling your hand (and possibly melting some of the ice cube).

Figure 1.11▼ shows the three common temperature scales. The most common in the United States is the **Fahrenheit (°F) scale**, shown on the left in Figure 1.11. On the Fahrenheit scale, water freezes at 32 °F and boils at 212 °F (at sea level). Room temperature is approximately 72 °F. The Fahrenheit scale was originally determined by assigning 0 °F to the freezing point of a concentrated saltwater solution and 96 °F to normal body temperature.

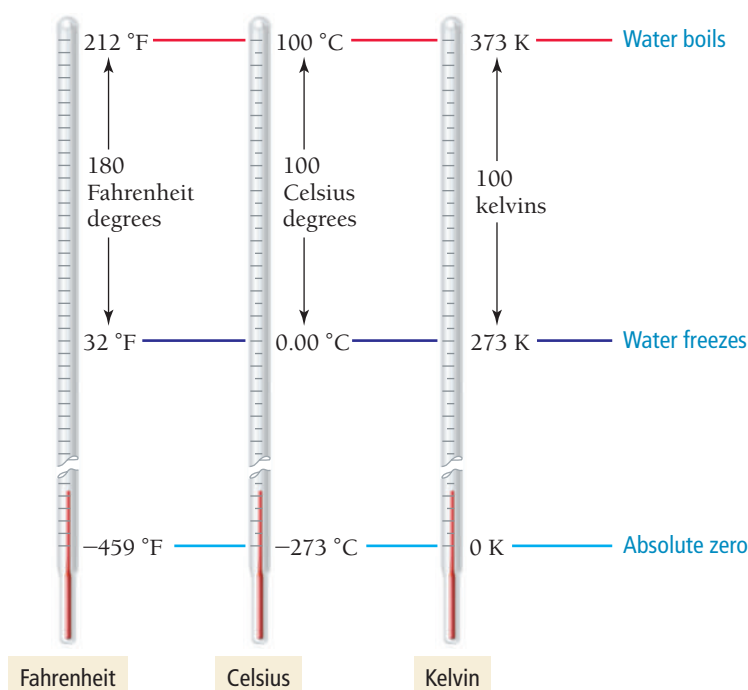
Scientists and citizens of most countries other than the United States typically use the **Celsius (°C) scale**, shown in the middle in Figure 1.11. On this scale, pure water freezes at 0 °C and boils at 100 °C (at sea level). Room temperature is approximately 22 °C. The Fahrenheit scale and the Celsius scale differ both in the size of their respective degrees and the temperature each designates as “zero.” Both the Fahrenheit and Celsius scales allow for negative temperatures.

The SI unit for temperature is the kelvin, shown in Figure 1.11. The **Kelvin scale** (sometimes also called the *absolute scale*) avoids negative temperatures by assigning 0 K to the coldest temperature possible, absolute zero. Absolute zero ( $-273.15$  °C or  $-459$  °F) is the temperature at which molecular motion virtually stops. Lower temperatures do not exist. The size of the kelvin is identical to that of the Celsius degree; the only difference is

Normal body temperature on the modern Fahrenheit scale is 98.6 °F.

Molecular motion does not *completely* stop at absolute zero because of the uncertainty principle in quantum mechanics, which we discuss in Chapter 8.

Temperature Scales

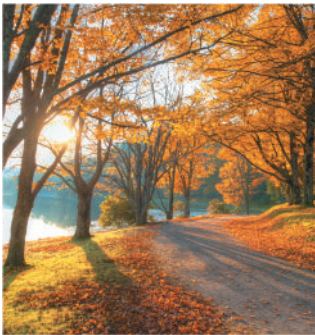


◀ **FIGURE 1.11** Comparison of the Fahrenheit, Celsius, and Kelvin Temperature Scales The Fahrenheit degree is five-ninths the size of the Celsius degree and the kelvin. The zero point of the Kelvin scale is absolute zero (the lowest possible temperature), whereas the zero point of the Celsius scale is the freezing point of water.

The Celsius Temperature Scale



0 °C – Water freezes



10 °C – Brisk fall day



22 °C – Room temperature



45 °C – Summer day in Death Valley

Note that we give Kelvin temperatures in kelvins (not “degrees Kelvin”) or K (not °K).

the temperature that each scale designates as zero. We can convert between the temperature scales with these formulas:

$$\begin{aligned} ^\circ\text{C} &= \frac{(^{\circ}\text{F} - 32)}{1.8} \\ \text{K} &= ^\circ\text{C} + 273.15 \end{aligned}$$

ANSWER NOW!



1.5  
**Cc**  
Conceptual  
Connection

**TEMPERATURE SCALES** Which temperature scale has no negative temperatures?

- (a) Kelvin
- (b) Celsius
- (c) Fahrenheit

Throughout this book, we provide examples worked out in formats that are designed to help you develop problem-solving skills. The most common format uses two columns to guide you through the worked example. The left column describes the thought processes and steps used in solving the problem, while the right column shows the implementation. Example 1.2 follows this two-column format.

EXAMPLE 1.2 Converting between Temperature Scales

A sick child has a temperature of 40.00 °C. What is the child’s temperature in (a) K and (b) °F?

SOLUTION

(a) Begin by finding the equation that relates the quantity that is given (°C) and the quantity you are trying to find (K).	$K = ^\circ\text{C} + 273.15$
Since this equation gives the temperature in K directly, substitute in the correct value for the temperature in °C and calculate the answer.	$\begin{aligned} K &= ^\circ\text{C} + 273.15 \\ K &= 40.00 + 273.15 = 313.15 \text{ K} \end{aligned}$
(b) To convert from °C to °F, first find the equation that relates these two quantities.	$^\circ\text{C} = \frac{(^{\circ}\text{F} - 32)}{1.8}$
Since this equation expresses °C in terms of °F, solve the equation for °F.	$\begin{aligned} ^\circ\text{C} &= \frac{(^{\circ}\text{F} - 32)}{1.8} \\ 1.8(^{\circ}\text{C}) &= (^{\circ}\text{F} - 32) \\ ^{\circ}\text{F} &= 1.8(^{\circ}\text{C}) + 32 \end{aligned}$
Now substitute °C into the equation and calculate the answer. <i>Note: The number of digits reported in this answer follows significant figure conventions, covered in Section 1.7.</i>	$\begin{aligned} ^{\circ}\text{F} &= 1.8(^{\circ}\text{C}) + 32 \\ ^{\circ}\text{F} &= 1.8(40.00^{\circ}\text{C}) + 32 = 104.00^{\circ}\text{F} \end{aligned}$

**FOR PRACTICE 1.2** Gallium is a solid metal at room temperature, but it will melt to a liquid in your hand. The melting point of gallium is 85.6 °F. What is this temperature on (a) the Celsius scale and (b) the Kelvin scale?



## Prefix Multipliers

Scientific notation (see Appendix IA) allows us to express very large or very small quantities in a compact manner by using exponents. For example, the diameter of a hydrogen atom is  $1.06 \times 10^{-10}$  m. The International System of Units uses the **prefix multipliers** listed in Table 1.2 with the standard units. These multipliers change the value of the unit by powers of 10 (just like an exponent does in scientific notation). For example, the kilometer has the prefix *kilo* meaning 1000 or  $10^3$ . Therefore,

$$1 \text{ kilometer} = 1000 \text{ meters} = 10^3 \text{ meters}$$

Similarly, the millimeter has the prefix *milli* meaning 0.001 or  $10^{-3}$ .

$$1 \text{ millimeter} = 0.001 \text{ meters} = 10^{-3} \text{ meters}$$

**TABLE 1.2** SI Prefix Multipliers

Prefix	Symbol	Multiplier	
exa	E	1,000,000,000,000,000,000	( $10^{18}$ )
peta	P	1,000,000,000,000,000	( $10^{15}$ )
tera	T	1,000,000,000,000	( $10^{12}$ )
giga	G	1,000,000,000	( $10^9$ )
mega	M	1,000,000	( $10^6$ )
kilo	k	1000	( $10^3$ )
deci	d	0.1	( $10^{-1}$ )
centi	c	0.01	( $10^{-2}$ )
milli	m	0.001	( $10^{-3}$ )
micro	$\mu$	0.000001	( $10^{-6}$ )
nano	n	0.000000001	( $10^{-9}$ )
pico	p	0.000000000001	( $10^{-12}$ )
femto	f	0.000000000000001	( $10^{-15}$ )
atto	a	0.000000000000000001	( $10^{-18}$ )

When reporting a measurement, choose a prefix multiplier close to the size of the quantity you are measuring. For example, to state the diameter of a hydrogen atom, which is  $1.06 \times 10^{-10}$  m, use picometers (106 pm) or nanometers (0.106 nm) rather than micrometers or millimeters. Choose the prefix multiplier that is most convenient for a particular number.

**PREFIX MULTIPLIERS** Which prefix multiplier is most appropriate for reporting a measurement of  $5.57 \times 10^{-5}$  m?

- (a) mega
- (b) milli
- (c) micro
- (d) kilo

**1.6**  
**Cc**  
Conceptual  
Connection

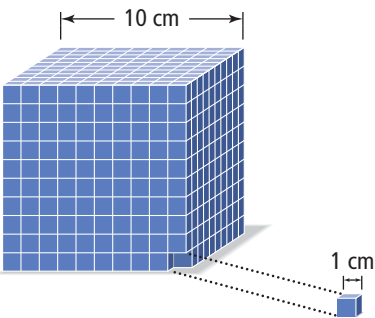
ANSWER NOW!



## Derived Units: Volume and Density

A **derived unit** is a combination of other units. For example, the SI unit for speed is meters per second (m/s), a derived unit. Notice that this unit is formed from two other SI units—meters and seconds—put together. You are probably more familiar with speed measured in miles/hour or kilometers/hour—these are also examples of derived units. Two other common derived units are those for volume (SI base unit is  $\text{m}^3$ ) and density (SI base unit is  $\text{kg}/\text{m}^3$ ).

Relationship between Length and Volume



A 10-cm cube contains 1000 1-cm cubes.

**▲ FIGURE 1.12** The Relationship between Length and Volume A cube with a 10-cm edge has a volume of (10 cm)<sup>3</sup> or 1000 cm<sup>3</sup>, and a cube with a 100-cm edge has a volume of (100 cm)<sup>3</sup> = 1,000,000 cm<sup>3</sup>.

The *m* in the equation for density is in italic type, meaning that it stands for mass rather than for meters. In general, the symbols for units such as meters (m), seconds (s), or kelvins (K) appear in regular type while those for variables such as mass (*m*), volume (*V*), and time (*t*) appear in italics.

Volume

**Volume** is a measure of space. Any unit of length, when cubed (raised to the third power), becomes a unit of volume. The cubic meter (m<sup>3</sup>), cubic centimeter (cm<sup>3</sup>), and cubic millimeter (mm<sup>3</sup>) are all units of volume. The cubic nature of volume is not always intuitive, and studies have shown that our brains are not naturally wired to process abstract concepts such as volume. For example, consider this question: how many small cubes measuring 1 cm on each side are required to construct a large cube measuring 10 cm on a side?

The answer to this question, as you can see by carefully examining the unit cube in Figure 1.12, is 1000 small cubes. When we go from a linear, one-dimensional distance to three-dimensional volume, we must raise both the linear dimension *and* its unit to the third power (not multiply by 3). Thus, the volume of a cube is equal to the length of its edge cubed.

volume of cube = (edge length)<sup>3</sup>

Other common units of volume in chemistry are the **liter (L)** and the **milliliter (mL)**. One milliliter (10<sup>-3</sup> L) is equal to 1 cm<sup>3</sup>. A gallon of gasoline contains 3.785 L. Table 1.3 lists some common units—for volume and other quantities—and their equivalents.

TABLE 1.3 Some Common Units and Their Equivalents		
Length	Mass	Volume
1 kilometer (km) = 0.6214 mile (mi)	1 kilogram (kg) = 2.205 pounds (lb)	1 liter (L) = 1000 mL = 1000 cm <sup>3</sup>
1 meter (m) = 39.37 inches (in) = 1.094 yards (yd)	1 pound (lb) = 453.59 grams (g)	1 liter (L) = 1.057 quarts (qt)
1 foot (ft) = 30.48 centimeters (cm)	1 ounce (oz) = 28.35 grams (g)	1 U.S. gallon (gal) = 3.785 liters (L)
1 inch (in) = 2.54 centimeters (cm) (exact)		

Density

An old riddle asks, “Which weighs more, a ton of bricks or a ton of feathers?” The answer is neither—they both weigh the same (1 ton). If you answered bricks, you confused weight with density. The **density (*d*)** of a substance is the ratio of its mass (*m*) to its volume (*V*).

Density =  $\frac{\text{mass}}{\text{volume}}$  or  $d = \frac{m}{V}$

Density is a characteristic physical property of substances (see Table 1.4) that depends on temperature. Density is an example of an **intensive property**, one that is *independent* of the amount of the substance. The density of aluminum, for example, is the same whether you have a gram or a kilogram. Intensive properties are often used to identify substances because these properties depend only on the type of substance, not on the amount of it. For example, from Table 1.4 you can see that pure gold has a density of 19.3 g/cm<sup>3</sup>. One way to determine whether a substance is pure gold is to determine its density and compare it to 19.3 g/cm<sup>3</sup>. Mass, in contrast, is an **extensive property**, one that depends on the amount of the substance. If you know only the mass of a sample of gold, that information alone will not allow you to identify it as gold.

The units of density are those of mass divided by volume. Although the SI-derived unit for density is kg/m<sup>3</sup>, the density of liquids and solids is most often expressed in g/cm<sup>3</sup> or g/mL. (Remember that cm<sup>3</sup> and mL are equivalent units: 1 cm<sup>3</sup> = 1 mL.) Aluminum is among the least dense metals with a density of 2.7 g/cm<sup>3</sup>, while platinum is one of the densest metals with a density of 21.4 g/cm<sup>3</sup>.

## Calculating Density

We can calculate the density of a substance by dividing the mass of a given amount of the substance by its volume. For example, suppose a small nugget we suspect to be gold has a mass of 22.5 g and a volume of 2.38 cm<sup>3</sup>. To find its density, we divide the mass by the volume:

$$d = \frac{m}{V} = \frac{22.5 \text{ g}}{2.38 \text{ cm}^3} = 9.45 \text{ g/cm}^3$$

In this case, the density reveals that the nugget is not pure gold because the density of gold is 19.3 g/cm<sup>3</sup>.

### EXAMPLE 1.3 Calculating Density

A person receives a platinum ring from their partner. Before the wedding, they notice that the ring feels a little light for its size, and so they decide to determine its density. They place the ring on a balance and find that it has a mass of 3.15 g. They then find that the ring displaces 0.233 cm<sup>3</sup> of water. Is the ring made of platinum? (Note: The volume of irregularly shaped objects is often measured by the displacement of water. To use this method, the object is placed in water and the change in volume of the water is measured. This increase in the total volume represents the volume of water *displaced* by the object and is equal to the volume of the object.)

Set up the problem by writing the important information that is *given* as well as the information that you are asked to *find*. In this case, we need to find the density of the ring and compare it to that of platinum.

*Note: In Section 1.8, we discuss this standardized way of setting up problems.*

**GIVEN:**  $m = 3.15 \text{ g}$

$V = 0.233 \text{ cm}^3$

**FIND:** Density in g/cm<sup>3</sup>

Next, write down the equation that defines density.

**EQUATION**  $d = \frac{m}{V}$

Solve the problem by substituting the correct values of mass and volume into the expression for density.

**SOLUTION**

$$d = \frac{m}{V} = \frac{3.15 \text{ g}}{0.233 \text{ cm}^3} = 13.5 \text{ g/cm}^3$$

The density of the ring is much too low to be platinum (platinum's density is 21.4 g/cm<sup>3</sup>), and the ring is therefore a fake.

**FOR PRACTICE 1.3** The person in Example 1.3 is shocked that the ring is fake and returns it. They buy a new ring that has a mass of 4.53 g and a volume of 0.212 cm<sup>3</sup>. Is the new ring genuine?

**FOR MORE PRACTICE 1.3** A metal cube has an edge that is 11.4 mm long and a mass of 6.67 g. Calculate the density of the metal and use Table 1.4 to determine the likely identity of the metal.

**TABLE 1.4**

**The Density of Some Common Substances at 20 °C**

Substance	Density (g/cm <sup>3</sup> )
Charcoal (from oak)	0.57
Ethanol	0.789
Ice	0.917 (at 0 °C)
Water	1.00 (at 4 °C)
Sugar (sucrose)	1.58
Table salt (sodium chloride)	2.16
Glass	2.6
Aluminum	2.70
Titanium	4.51
Iron	7.86
Copper	8.96
Lead	11.4
Mercury	13.55
Gold	19.3
Platinum	21.4

**DENSITY** The density of copper decreases as temperature increases (as does the density of most substances). Which change occurs in a sample of copper when it is warmed from room temperature to 95 °C?

- (a) The sample becomes lighter.
- (b) The sample becomes heavier.
- (c) The sample expands.
- (d) The sample contracts.

**1.7**  
**Cc**  
Conceptual  
Connection

**ANSWER NOW!**

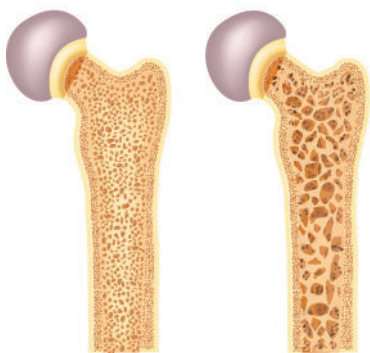






## CHEMISTRY AND MEDICINE

## Bone Density



**O**steoporosis—which means *porous bone*—is a condition in which bone density becomes low. The healthy bones of a young adult have a density of about  $1.0 \text{ g/cm}^3$ . Patients suffering from osteoporosis, however, can have bone densities as low as  $0.22 \text{ g/cm}^3$ . These low densities indicate that the bones have deteriorated and weakened, resulting in increased susceptibility to fractures, especially hip fractures. Patients suffering from osteoporosis can also experience height loss and disfigurement such as dowager's hump, a condition in which the patient becomes hunched over due to compression of the vertebrae. Osteoporosis is most common in postmenopausal women, but it can also occur in people who have certain diseases, such as insulin-dependent diabetes, or who take certain medications, such as prednisone. Osteoporosis is usually diagnosed and monitored with hip X-rays. Low-density bones absorb less of the X-rays than do high-density bones, producing characteristic differences in the X-ray image. Treatments for osteoporosis include calcium and vitamin D supplements, drugs that prevent bone weakening, exercise and strength training, and, in extreme cases, hip-replacement surgery.

**QUESTION** Suppose you find a large animal bone in the woods, too large to fit in a beaker or flask. How might you approximate its density?

◀ Top: Severe osteoporosis can necessitate surgery to implant an artificial hip joint, seen in this X-ray image. Bottom: Views of the bone matrix in a normal bone (left) and one weakened by osteoporosis (right).

## 1.7 The Reliability of a Measurement

Carbon monoxide is a colorless gas emitted by motor vehicles and found in polluted air. The table shown here lists carbon monoxide concentrations in Los Angeles County as reported by the U.S. Environmental Protection Agency (EPA) over the period 1990–2017:

Year	Carbon Monoxide Concentration (ppm)*
1990	9.9
2000	7.4
2010	1.9
2020	2.2

\*Second maximum, 8-hour average; ppm = parts per million (Pasadena Site 06-037-2005)

The first thing you should notice about these values is that they decrease over time. For this decrease, we can thank the Clean Air Act and its amendments, which have resulted in more efficient engines and specially blended fuels and consequently in cleaner air in all major U.S. cities over the last 30 years. The second thing you might notice is the number of digits to which the measurements are reported. The number of digits in a reported measurement indicates the certainty associated with that measurement. A less certain measurement of carbon monoxide levels might be reported as follows:

Year	Carbon Monoxide Concentration (ppm)
1990	10
2000	7
2010	2
2020	2

Notice that the first set of data is reported to the nearest 0.1 ppm, while the second set is reported to the nearest 1 ppm. Scientists report measured quantities in an agreed-upon standard way. The number of reported digits reflects the certainty in the measurement: more digits, more certainty; fewer digits, less certainty. Numbers are usually written so that the uncertainty is in the last reported digit. (We assume that uncertainty to be  $\pm 1$  in the last digit unless otherwise indicated.) By reporting the 2010 carbon monoxide concentration as 1.9 ppm, the scientists mean  $1.9 \pm 0.1$  ppm. The carbon monoxide concentration is between 1.8 and 2.0 ppm—it might be 2.0 ppm, for example, but it could not be 3.0 ppm. In contrast, if the reported value was 2 ppm (as in the second set of measurements), this would mean  $2 \pm 1$  ppm, or between 1 and 3 ppm. In general,

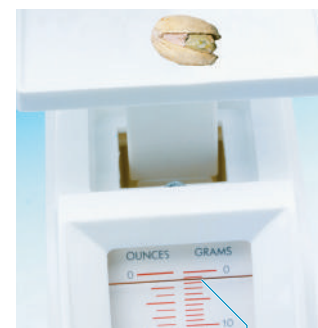
**Scientific measurements are reported so that every digit is certain except the last, which is estimated.**

For example, consider the following reported number:  
The first three digits are certain; the last digit is estimated.

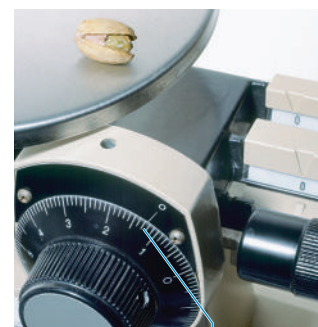
5.213  
↑ certain     ↑ estimated

The number of digits reported in a measurement depends on the measuring device. Consider weighing a pistachio nut on two different balances (Figure 1.13►). The balance on the top has marks every 1 g, while the balance on the bottom has marks every 0.1 g. For the balance on the top, we mentally divide the space between the 1- and 2-g marks into ten equal spaces and estimate that the pointer is at about 1.2 g. We then write the measurement as 1.2 g, indicating that we are sure of the “1” but we have estimated the “.2.” The balance on the bottom, with marks every *tenth* of a gram, requires that we write the result with more digits. The pointer is between the 1.2-g mark and the 1.3-g mark. We again divide the space between the two marks into ten equal spaces and estimate the third digit. For the figure shown, we report 1.27 g.

## Estimation in Weighing



(a) Markings every 1 g  
Estimated reading 1.2 g



(b) Markings every 0.1 g  
Estimated reading 1.27 g

**▲ FIGURE 1.13 Estimation in Weighing** (a) This scale has markings every 1 g, so we mentally divide the space into ten equal spaces to estimate the last digit. This reading is 1.2 g. (b) Because this balance has markings every 0.1 g, we estimate to the hundredths place. This reading is 1.27 g.

**WATCH NOW!**

### INTERACTIVE WORKED EXAMPLE 1.4

## EXAMPLE 1.4 Reporting the Correct Number of Digits

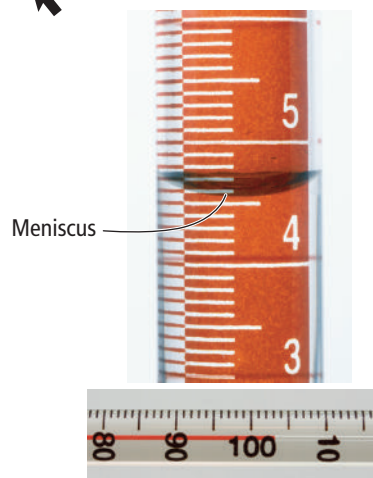
The graduated cylinder shown to the right has markings every 0.1 mL. Report the volume (which is read at the bottom of the meniscus) to the correct number of digits. (Note: The meniscus is the crescent-shaped surface at the top of a column of liquid.)

### SOLUTION

Since the bottom of the meniscus is between the 4.5-mL and 4.6-mL markings, mentally divide the space between the markings into ten equal spaces and estimate the next digit. In this case, you should report the result as 4.57 mL.

What if you estimated a little differently and wrote 4.56 mL? In general, a one-unit difference in the last digit is acceptable because the last digit is estimated and different people might estimate it slightly differently. However, if you wrote 4.63 mL, you would have misreported the measurement.

**FOR PRACTICE 1.4** Record the temperature on the thermometer shown at the right to the correct number of digits.



## Counting Significant Figures

The precision of a measurement—which depends on the instrument used to make the measurement—must be preserved, not only when recording the measurement, but also when performing calculations that use the measurement. We can accomplish the preservation of this precision by using *significant figures*. In any reported measurement, the non-place-holding digits—those that are not simply marking the decimal place—are called **significant figures** (or **significant digits**). *The greater the number of significant figures, the greater the certainty of the measurement.* For example, the number 23.5 has three significant figures while the number 23.56 has four. To determine the number of significant figures in a number containing zeroes, we distinguish between zeroes that are significant and those that simply mark the decimal place. For example, in the number 0.0008, the leading zeroes (zeroes to the left of the first nonzero digit) mark the decimal place but *do not* add to the certainty of the measurement and are therefore not significant; this number has only one significant figure. In contrast, the trailing zeroes (zeroes at the end of a number) in the number 0.000800 *do add* to the certainty of the measurement and are therefore counted as significant; this number has three significant figures.

### HOW TO: Determine the Number of Significant Figures in a Given Value

#### Significant Figure Rules

1. All nonzero digits are significant.
2. Interior zeroes (zeroes between two nonzero digits) are significant.
3. Leading zeroes (zeroes to the left of the first nonzero digit) are not significant. They only serve to locate the decimal point.
4. Trailing zeroes (zeroes at the end of a number) are categorized as follows:
  - Trailing zeroes after a decimal point are always significant.
  - Trailing zeroes before a decimal point (and after a nonzero number) are always significant.
  - Trailing zeroes before an *implied* decimal point are ambiguous and should be avoided by using scientific notation.
  - Some textbooks put a decimal point after one or more trailing zeroes if the zeroes are to be considered significant. We avoid that practice in this book, but you should be aware of it.

#### Examples

28.03	0.0540
408	7.0301
0.0032	0.00006
45.000	3.5600
140.00	2500.55
1200	ambiguous
$1.2 \times 10^3$	2 significant figures
$1.20 \times 10^3$	3 significant figures
$1.200 \times 10^3$	4 significant figures
1200.	4 significant figures (common in some textbooks)

## Exact Numbers

**Exact numbers** have no uncertainty and thus do not limit the number of significant figures in any calculation. We can regard an exact number as having an unlimited number of significant figures. Exact numbers originate from three sources:

- From the accurate counting of discrete objects. For example, 3 atoms means 3.00000 . . . atoms.
- From defined quantities, such as the number of centimeters in 1 m. Because 100 cm is defined as 1 m,
 
$$100 \text{ cm} = 1 \text{ m means } 100.00000 \dots \text{ cm} = 1.0000000 \dots \text{ m}$$
- From integral numbers that are part of an equation. For example, in the equation  $\text{radius} = \frac{\text{diameter}}{2}$ , the number 2 is exact and therefore has an unlimited number of significant figures.



WATCH NOW!

## INTERACTIVE WORKED EXAMPLE 1.5

**EXAMPLE 1.5** Determining the Number of Significant Figures in a Number

How many significant figures are in each number?

- (a) 0.04450 m    (b) 5.0003 km    (c) 10 dm = 1 m    (d)  $1.000 \times 10^5$  s    (e) 0.00002 mm    (f) 10,000 m

**SOLUTION**

(a) 0.04450 m	<i>Four significant figures.</i> The two 4s and the 5 are significant (Rule 1). The trailing zero is after a decimal point and is therefore significant (Rule 4). The leading zeroes only mark the decimal place and are therefore not significant (Rule 3).
(b) 5.0003 km	<i>Five significant figures.</i> The 5 and 3 are significant (Rule 1), as are the three interior zeroes (Rule 2).
(c) 10 dm = 1 m	<i>Unlimited significant figures.</i> Defined quantities have an unlimited number of significant figures.
(d) $1.000 \times 10^5$ s	<i>Four significant figures.</i> The 1 is significant (Rule 1). The trailing zeroes are after a decimal point and therefore significant (Rule 4).
(e) 0.00002 mm	<i>One significant figure.</i> The 2 is significant (Rule 1). The leading zeroes only mark the decimal place and are therefore not significant (Rule 3).
(f) 10,000 m	<i>Ambiguous.</i> The 1 is significant (Rule 1), but the trailing zeroes occur before an implied decimal point and are therefore ambiguous (Rule 4). Without more information, you would assume one significant figure. It is better to write this as $1 \times 10^5$ to indicate one significant figure or as $1.0000 \times 10^5$ to indicate five (Rule 4).

**FOR PRACTICE 1.5** How many significant figures are in each number?

- (a) 554 km    (b) 7 pennies    (c)  $1.01 \times 10^5$  m    (d) 0.00099 s    (e) 1.4500 km    (f) 21,000 m

**Significant Figures in Calculations**

When we use measured quantities in calculations, the results of the calculation must reflect the precision of the measured quantities. We should not lose or gain precision during mathematical operations. Follow these rules when carrying significant figures through calculations.



WATCH NOW!

**KEY CONCEPT**  
**VIDEO 1.7**  
Significant Figures  
in Calculations

**HOW TO:** Determine Significant Figures in Calculated Quantities**Rules for Calculations**

1. In multiplication or division, the result carries the same number of significant figures as the factor with the fewest significant figures.

**Examples**

$$\begin{array}{ccccccc} 1.052 & \times & 12.054 & \times & 0.53 & = & 6.7208 \\ (4 \text{ sig. figures}) & & (5 \text{ sig. figures}) & & (2 \text{ sig. figures}) & & (2 \text{ sig. figures}) \end{array}$$

$$\begin{array}{ccccccc} 2.0035 & \div & 3.20 & = & 0.626094 & = & 0.626 \\ (5 \text{ sig. figures}) & & (3 \text{ sig. figures}) & & & & (3 \text{ sig. figures}) \end{array}$$

2. In addition or subtraction, the result carries the same number of decimal places as the quantity with the fewest decimal places.

$$\begin{array}{r} 2.345 \\ 0.07 \\ \hline 2.9975 \\ \underline{5.4125} = 5.41 \end{array} \qquad \begin{array}{r} 5.9 \\ -0.221 \\ \hline 5.679 = 5.7 \end{array}$$

In addition and subtraction, it is helpful to draw a line next to the number with the fewest decimal places. This line determines the number of decimal places in the answer.

—Continued on the next page