

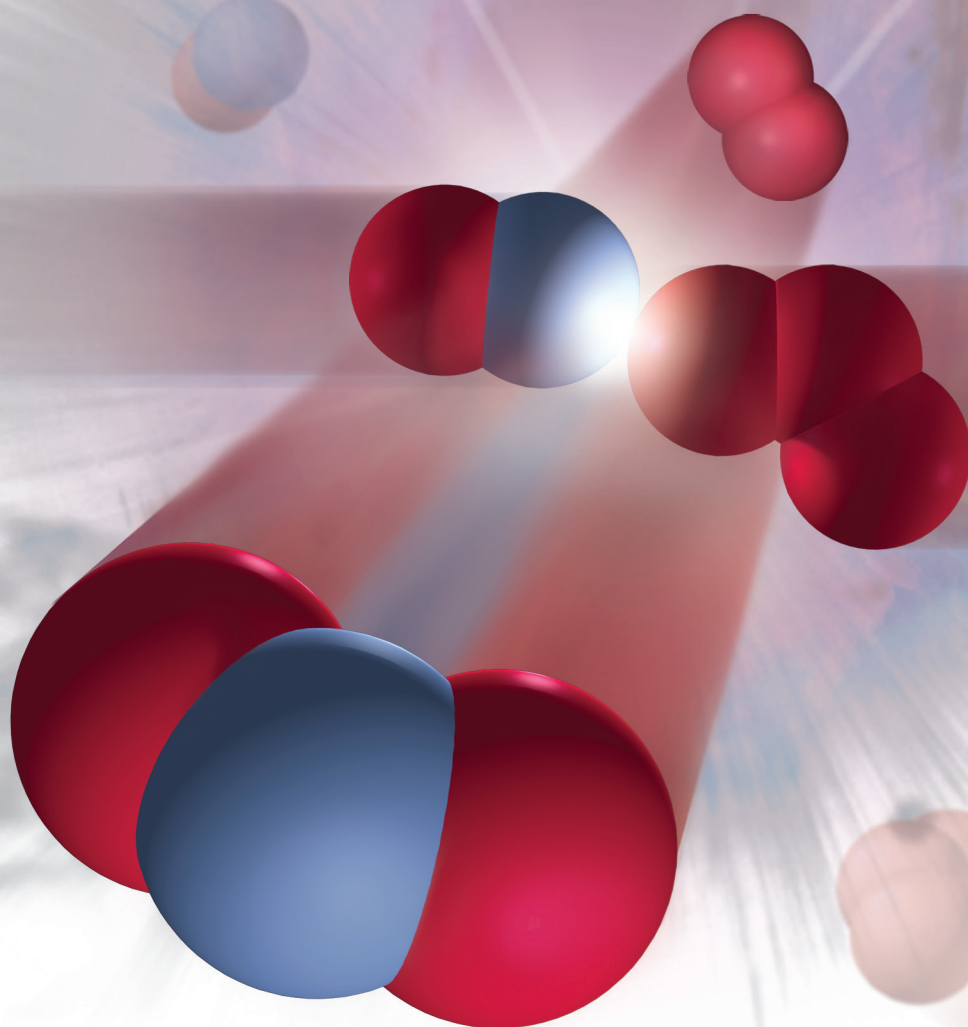
Silberberg & Amateis

7e

# CHEMISTRY

The Molecular Nature of Matter and Change

Silberberg  Amateis



7e

# CHEMISTRY

The Molecular Nature of Matter and Change

Mc  
Graw  
Hill  
Education



CHEMISTRY: THE MOLECULAR NATURE OF MATTER AND CHANGE, SEVENTH EDITION

Published by McGraw-Hill Education, 2 Penn Plaza, New York, NY 10121. Copyright © 2015 by McGraw-Hill Education. All rights reserved. Printed in the United States of America. Previous editions © 2012, 2009, and 2006. No part of this publication may be reproduced or distributed in any form or by any means, or stored in a database or retrieval system, without the prior written consent of McGraw-Hill Education, including, but not limited to, in any network or other electronic storage or transmission, or broadcast for distance learning.

Some ancillaries, including electronic and print components, may not be available to customers outside the United States.

This book is printed on acid-free paper.

1 2 3 4 5 6 7 8 9 0 DOW/DOW 1 0 9 8 7 6 5 4

ISBN 978-0-07-351117-7

MHID 0-07-351117-X

Senior Vice President, Products & Markets: *Kurt L. Strand*  
Vice President, General Manager, Products & Markets: *Marty Lange*  
Vice President, Content Production & Technology Services: *Kimberly Meriwether David*  
Managing Director: *Thomas Timp*  
Executive Brand Manager: *David Spurgeon, Ph.D.*  
Director of Development: *Rose Koos*  
Senior Development Editor: *Lora Neyens*  
Executive Marketing Manager: *Tamara L. Hodge*  
Director of Digital Content: *Shirley Hino, Ph.D.*  
Director, Content Production: *Terri Schiesl*  
Content Project Manager (print): *Peggy Selle*  
Content Project Manager (media): *Laura Bies*  
Senior Buyer: *Sandy Ludovissy*  
Senior Designer: *David W. Hash*  
Cover Image: *@Victor Habbick Visions/Getty Images, Lachina Publishing Services*  
Senior Content Licensing Specialist: *Lori Hancock*  
Compositor: *Lachina Publishing Services*  
Typeface: *10/12 Times*  
Printer: *R. R. Donnelley*

All credits appearing on page or at the end of the book are considered to be an extension of the copyright page.

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

Silberberg, Martin S. (Martin Stuart), 1945-

Chemistry : the molecular nature of matter and change / Martin S. Silberberg, Patricia Amateis, Virginia Polytechnic. – Seventh edition.

pages cm

Includes index.

ISBN 978-0-07-351117-7 — ISBN 0-07-351117-X (hard copy : alk. paper) 1. Chemistry—Textbooks.

I. Amateis, Patricia. II. Title.

QD33.2.S55 2015

540-dc23

2013033592

The Internet addresses listed in the text were accurate at the time of publication. The inclusion of a website does not indicate an endorsement by the authors or McGraw-Hill Education, and McGraw-Hill Education does not guarantee the accuracy of the information presented at these sites.

To Ruth and Daniel, with all my love and gratitude.

MSS

To Ralph, Eric, Samantha, and Lindsay:  
you bring me much joy.

PGA

# BRIEF CONTENTS

Preface xx

Acknowledgments xxxi

- 1** *Keys to the Study of Chemistry* 2
- 2** *The Components of Matter* 40
- 3** *Stoichiometry of Formulas and Equations* 90
- 4** *Three Major Classes of Chemical Reactions* 138
- 5** *Gases and the Kinetic-Molecular Theory* 198
- 6** *Thermochemistry: Energy Flow and Chemical Change* 250
- 7** *Quantum Theory and Atomic Structure* 286
- 8** *Electron Configuration and Chemical Periodicity* 322
- 9** *Models of Chemical Bonding* 358
- 10** *The Shapes of Molecules* 394
- 11** *Theories of Covalent Bonding* 428
- 12** *Intermolecular Forces: Liquids, Solids, and Phase Changes* 454
- 13** *The Properties of Mixtures: Solutions and Colloids* 516
- 14** *Periodic Patterns in the Main-Group Elements* 568
- 15** *Organic Compounds and the Atomic Properties of Carbon* 616
- 16** *Kinetics: Rates and Mechanisms of Chemical Reactions* 674
- 17** *Equilibrium: The Extent of Chemical Reactions* 730
- 18** *Acid-Base Equilibria* 776
- 19** *Ionic Equilibria in Aqueous Systems* 826
- 20** *Thermodynamics: Entropy, Free Energy, and the Direction of Chemical Reactions* 876
- 21** *Electrochemistry: Chemical Change and Electrical Work* 918
- 22** *The Elements in Nature and Industry* 976
- 23** *Transition Elements and Their Coordination Compounds* 1016
- 24** *Nuclear Reactions and Their Applications* 1052

---

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

**Appendix B** Standard Thermodynamic Values for Selected Substances A-5

**Appendix C** Equilibrium Constants for Selected Substances A-8

**Appendix D** Standard Electrode (Half-Cell) Potentials A-14

**Appendix E** Answers to Selected Problems A-15

Glossary G-1

Credits C-1

Index I-1



## CHAPTER

# 1

## Keys to the Study of Chemistry 2

- 1.1 Some Fundamental Definitions 4**
    - The States of Matter 4
    - The Properties of Matter and Its Changes 5
    - The Central Theme in Chemistry 8
    - The Importance of Energy in the Study of Matter 8
  - 1.2 Chemical Arts and the Origins of Modern Chemistry 10**
    - Prechemical Traditions 10
    - The Phlogiston Fiasco and the Impact of Lavoisier 11
  - 1.3 The Scientific Approach: Developing a Model 12**
  - 1.4 Measurement and Chemical Problem Solving 13**
    - General Features of SI Units 13
    - Some Important SI Units in Chemistry 14
    - Units and Conversion Factors in Calculations 17
    - A Systematic Approach to Solving Chemistry Problems 19
    - Temperature Scales 24
    - Extensive and Intensive Properties 26
  - 1.5 Uncertainty in Measurement: Significant Figures 27**
    - Determining Which Digits Are Significant 28
    - Significant Figures: Calculations and Rounding Off 29
    - Precision, Accuracy, and Instrument Calibration 31
- CHAPTER REVIEW GUIDE 32  
PROBLEMS 35

## CHAPTER

# 2

## The Components of Matter 40

- 2.1 Elements, Compounds, and Mixtures: An Atomic Overview 42**
  - 2.2 The Observations That Led to an Atomic View of Matter 44**
    - Mass Conservation 44
    - Definite Composition 45
    - Multiple Proportions 46
  - 2.3 Dalton's Atomic Theory 47**
    - Postulates of the Atomic Theory 47
    - How the Theory Explains the Mass Laws 48
  - 2.4 The Observations That Led to the Nuclear Atom Model 49**
    - Discovery of the Electron and Its Properties 49
    - Discovery of the Atomic Nucleus 51
  - 2.5 The Atomic Theory Today 52**
    - Structure of the Atom 53
    - Atomic Number, Mass Number, and Atomic Symbol 53
    - Isotopes 54
    - Atomic Masses of the Elements 55
  - TOOLS OF THE LABORATORY: MASS SPECTROMETRY 57**
  - 2.6 Elements: A First Look at the Periodic Table 58**
  - 2.7 Compounds: Introduction to Bonding 61**
    - The Formation of Ionic Compounds 61
    - The Formation of Covalent Substances 63
  - 2.8 Compounds: Formulas, Names, and Masses 65**
    - Binary Ionic Compounds 65
    - Compounds That Contain Polyatomic Ions 68
    - Acid Names from Anion Names 70
    - Binary Covalent Compounds 71
  - The Simplest Organic Compounds: Straight-Chain Alkanes 72
  - Molecular Masses from Chemical Formulas 72
  - Representing Molecules with Formulas and Models 75
  - 2.9 Mixtures: Classification and Separation 77**
    - An Overview of the Components of Matter 77
  - TOOLS OF THE LABORATORY: BASIC SEPARATION TECHNIQUES 79**
- CHAPTER REVIEW GUIDE 80  
PROBLEMS 82



## CHAPTER 3 Stoichiometry of Formulas and Equations 90

- 3.1 The Mole 91**  
 Defining the Mole 92  
 Determining Molar Mass 92  
 Converting Between Amount, Mass, and Number of Chemical Entities 93  
 The Importance of Mass Percent 98
- 3.2 Determining the Formula of an Unknown Compound 100**  
 Empirical Formulas 101  
 Molecular Formulas 102
- 3.3 Writing and Balancing Chemical Equations 107**
- 3.4 Calculating Quantities of Reactant and Product 112**  
 Stoichiometrically Equivalent Molar Ratios from the Balanced Equation 112  
 Reactions That Occur in a Sequence 116
- Chemical Formulas and Molecular Structures; Isomers 105
- Reactions That Involve a Limiting Reactant 117  
 Theoretical, Actual, and Percent Reaction Yields 123
- CHAPTER REVIEW GUIDE 125  
 PROBLEMS 130

## CHAPTER 4 Three Major Classes of Chemical Reactions 138

- 4.1 Solution Concentration and the Role of Water as a Solvent 139**  
 The Polar Nature of Water 140  
 Ionic Compounds in Water 140  
 Covalent Compounds in Water 143  
 Expressing Concentration in Terms of Molarity 144  
 Amount-Mass-Number Conversions Involving Solutions 144  
 Preparing and Diluting Molar Solutions 146
- 4.2 Writing Equations for Aqueous Ionic Reactions 149**
- 4.3 Precipitation Reactions 151**  
 The Key Event: Formation of a Solid from Dissolved Ions 151
- 4.4 Acid-Base Reactions 158**  
 The Key Event: Formation of  $\text{H}_2\text{O}$  from  $\text{H}^+$  and  $\text{OH}^-$  161  
 Proton Transfer in Acid-Base Reactions 162  
 Stoichiometry of Acid-Base Reactions: Acid-Base Titrations 165
- 4.5 Oxidation-Reduction (Redox) Reactions 168**  
 The Key Event: Movement of Electrons Between Reactants 168  
 Some Essential Redox Terminology 169
- Predicting Whether a Precipitate Will Form 151  
 Stoichiometry of Precipitation Reactions 155
- Using Oxidation Numbers to Monitor Electron Charge 170  
 Stoichiometry of Redox Reactions: Redox Titrations 172
- 4.6 Elements in Redox Reactions 174**  
 Combination Redox Reactions 174  
 Decomposition Redox Reactions 176  
 Displacement Redox Reactions and Activity Series 177  
 Combustion Reactions 180
- 4.7 The Reversibility of Reactions and the Equilibrium State 181**
- CHAPTER REVIEW GUIDE 184  
 PROBLEMS 189

## CHAPTER 5 Gases and the Kinetic-Molecular Theory 198

- 5.1 An Overview of the Physical States of Matter 199**
- 5.2 Gas Pressure and Its Measurement 201**  
 Measuring Gas Pressure: Barometers and Manometers 202  
 Units of Pressure 202
- 5.3 The Gas Laws and Their Experimental Foundations 204**  
 The Relationship Between Volume and Pressure: Boyle's Law 205  
 The Relationship Between Volume and Temperature: Charles's Law 206  
 The Relationship Between Volume and Amount: Avogadro's Law 208  
 Gas Behavior at Standard Conditions 209
- 5.4 Rearrangements of the Ideal Gas Law 216**  
 The Density of a Gas 216  
 The Molar Mass of a Gas 218  
 The Partial Pressure of Each Gas in a Mixture of Gases 219  
 The Ideal Gas Law and Reaction Stoichiometry 222
- 5.5 The Kinetic-Molecular Theory: A Model for Gas Behavior 224**  
 How the Kinetic-Molecular Theory Explains the Gas Laws 224  
 Effusion and Diffusion 229
- The Ideal Gas Law 209  
 Solving Gas Law Problems 210
- The Chaotic World of Gases: Mean Free Path and Collision Frequency 232
- CHEMICAL CONNECTIONS TO ATMOSPHERIC SCIENCE: HOW THE GAS LAWS APPLY TO EARTH'S ATMOSPHERE 233**
- 5.6 Real Gases: Deviations from Ideal Behavior 235**  
 Effects of Extreme Conditions on Gas Behavior 235  
 The van der Waals Equation: Adjusting the Ideal Gas Law 237
- CHAPTER REVIEW GUIDE 238  
 PROBLEMS 241



## CHAPTER

## 6

## Thermochemistry: Energy Flow and Chemical Change 250

**6.1 Forms of Energy and Their Interconversion 251**

Defining the System and Its Surroundings 252  
 Energy Change ( $\Delta E$ ): Energy Transfer to or from a System 252  
 Heat and Work: Two Forms of Energy Transfer 252  
 The Law of Energy Conservation 255  
 Units of Energy 255  
 State Functions and the Path Independence of the Energy Change 256

**6.2 Enthalpy: Changes at Constant Pressure 257**

The Meaning of Enthalpy 258  
 Comparing  $\Delta E$  and  $\Delta H$  259  
 Exothermic and Endothermic Processes 259

**6.3 Calorimetry: Measuring the Heat of a Chemical or Physical Change 261**

Specific Heat Capacity 261  
 The Two Major Types of Calorimetry 262

**6.4 Stoichiometry of Thermochemical Equations 266****6.5 Hess's Law: Finding  $\Delta H$  of Any Reaction 267****6.6 Standard Enthalpies of Reaction ( $\Delta H_{\text{rxn}}^\circ$ ) 270**

Formation Equations and Their Standard Enthalpy Changes 270  
 Determining  $\Delta H_{\text{rxn}}^\circ$  from  $\Delta H_f^\circ$  Values for Reactants and Products 271

**CHEMICAL CONNECTIONS TO ENVIRONMENTAL SCIENCE: THE FUTURE OF ENERGY USE 273**

CHAPTER REVIEW GUIDE 277

PROBLEMS 280

## CHAPTER

## 7

## Quantum Theory and Atomic Structure 286

**7.1 The Nature of Light 287**

The Wave Nature of Light 288  
 The Particle Nature of Light 291

**7.2 Atomic Spectra 294**

Line Spectra and the Rydberg Equation 294  
 The Bohr Model of the Hydrogen Atom 295  
 The Energy Levels of the Hydrogen Atom 297

**TOOLS OF THE LABORATORY: SPECTROMETRY IN CHEMICAL ANALYSIS 300**

**7.3 The Wave-Particle Duality of Matter and Energy 302**

The Wave Nature of Electrons and the Particle Nature of Photons 302  
 Heisenberg's Uncertainty Principle 305

**7.4 The Quantum-Mechanical Model of the Atom 306**

The Atomic Orbital and the Probable Location of the Electron 306

Quantum Numbers of an Atomic Orbital 308  
 Quantum Numbers and Energy Levels 309  
 Shapes of Atomic Orbitals 311  
 The Special Case of Energy Levels in the Hydrogen Atom 313

CHAPTER REVIEW GUIDE 314

PROBLEMS 316

## CHAPTER

## 8

## Electron Configuration and Chemical Periodicity 322

**8.1 Characteristics of Many-Electron Atoms 324**

The Electron-Spin Quantum Number 324  
 The Exclusion Principle 325  
 Electrostatic Effects and Energy-Level Splitting 325

**8.2 The Quantum-Mechanical Model and the Periodic Table 327**

Building Up Period 1 327  
 Building Up Period 2 328  
 Building Up Period 3 330

Similar Electron Configurations Within Groups 331  
 Building Up Period 4: The First Transition Series 332  
 General Principles of Electron Configurations 333  
 Intervening Series: Transition and Inner Transition Elements 334

**8.3 Trends in Three Atomic Properties 336**

Trends in Atomic Size 336  
 Trends in Ionization Energy 339  
 Trends in Electron Affinity 343

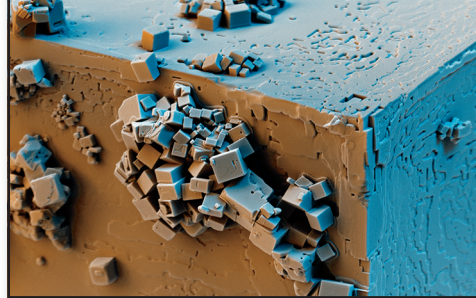
**8.4 Atomic Properties and Chemical Reactivity 344**

Trends in Metallic Behavior 344  
 Properties of Monatomic Ions 346

CHAPTER REVIEW GUIDE 352

PROBLEMS 354





## CHAPTER

## 9

## Models of Chemical Bonding 358

**9.1 Atomic Properties and Chemical Bonds 359**

The Three Ways Elements Combine 359  
Lewis Symbols and the Octet Rule 361

**9.2 The Ionic Bonding Model 362**

Why Ionic Compounds Form:  
The Importance of Lattice Energy 363  
Periodic Trends in Lattice Energy 365  
How the Model Explains the Properties of Ionic Compounds 367

**9.3 The Covalent Bonding Model 369**

The Formation of a Covalent Bond 369  
Bonding Pairs and Lone Pairs 370  
Properties of a Covalent Bond:  
Order, Energy, and Length 370

How the Model Explains the Properties of Covalent Substances 373

**TOOLS OF THE LABORATORY: INFRARED SPECTROSCOPY 374****9.4 Bond Energy and Chemical Change 376**

Changes in Bond Energy: Where Does  $\Delta H_{\text{rxn}}^{\circ}$  Come From? 376  
Using Bond Energies to Calculate  $\Delta H_{\text{rxn}}^{\circ}$  376  
Bond Strengths and the Heat Released from Fuels and Foods 379

**9.5 Between the Extremes: Electronegativity and Bond Polarity 380**

Electronegativity 380  
Bond Polarity and Partial Ionic Character 382

The Gradation in Bonding Across a Period 384

**9.6 An Introduction to Metallic Bonding 385**

The Electron-Sea Model 385  
How the Model Explains the Properties of Metals 386

CHAPTER REVIEW GUIDE 387

PROBLEMS 389

## CHAPTER

## 10

## The Shapes of Molecules 394

**10.1 Depicting Molecules and Ions with Lewis Structures 395**

Applying the Octet Rule to Write Lewis Structures 395  
Resonance: Delocalized Electron-Pair Bonding 399  
Formal Charge: Selecting the More Important Resonance Structure 401  
Lewis Structures for Exceptions to the Octet Rule 402

**10.2 Valence-Shell Electron-Pair Repulsion (VSEPR) Theory 406**

Electron-Group Arrangements and Molecular Shapes 407

The Molecular Shape with Two Electron Groups (Linear Arrangement) 407  
Molecular Shapes with Three Electron Groups (Trigonal Planar Arrangement) 408  
Molecular Shapes with Four Electron Groups (Tetrahedral Arrangement) 409  
Molecular Shapes with Five Electron Groups (Trigonal Bipyramidal Arrangement) 410  
Molecular Shapes with Six Electron Groups (Octahedral Arrangement) 411  
Using VSEPR Theory to Determine Molecular Shape 412  
Molecular Shapes with More Than One Central Atom 415

**10.3 Molecular Shape and Molecular Polarity 416**

Bond Polarity, Bond Angle, and Dipole Moment 417  
The Effect of Molecular Polarity on Behavior 419

**CHEMICAL CONNECTIONS TO SENSORY PHYSIOLOGY: MOLECULAR SHAPE, BIOLOGICAL RECEPTORS, AND THE SENSE OF SMELL 420**

CHAPTER REVIEW GUIDE 421

PROBLEMS 424



## CHAPTER

## 11

**Theories of Covalent Bonding 428****11.1 Valence Bond (VB) Theory and Orbital Hybridization 429**

The Central Themes of VB Theory 429  
Types of Hybrid Orbitals 430

**11.2 Modes of Orbital Overlap and the Types of Covalent Bonds 436**

Orbital Overlap in Single and Multiple Bonds 436  
Orbital Overlap and Rotation Within a Molecule 439

**11.3 Molecular Orbital (MO) Theory and Electron Delocalization 440**

The Central Themes of MO Theory 440  
Homonuclear Diatomic Molecules of Period 2 Elements 442  
Two Heteronuclear Diatomic Molecules: HF and NO 446  
Two Polyatomic Molecules: Benzene and Ozone 447

CHAPTER REVIEW GUIDE 448

PROBLEMS 450

## CHAPTER

## 12

**Intermolecular Forces: Liquids, Solids, and Phase Changes 454****12.1 An Overview of Physical States and Phase Changes 455****12.2 Quantitative Aspects of Phase Changes 458**

Heat Involved in Phase Changes 459  
The Equilibrium Nature of Phase Changes 462  
Phase Diagrams: Effect of Pressure and Temperature on Physical State 466

**12.3 Types of Intermolecular Forces 468**

How Close Can Molecules Approach Each Other? 468  
Ion-Dipole Forces 469  
Dipole-Dipole Forces 469  
The Hydrogen Bond 470

Polarizability and Induced Dipole Forces 471

Dispersion (London) Forces 472

**12.4 Properties of the Liquid State 474**

Surface Tension 475  
Capillarity 475  
Viscosity 476

**12.5 The Uniqueness of Water 477**

Solvent Properties of Water 477  
Thermal Properties of Water 477  
Surface Properties of Water 478  
The Unusual Density of Solid Water 478

**12.6 The Solid State: Structure, Properties, and Bonding 479**

Structural Features of Solids 479

**TOOLS OF THE LABORATORY: X-RAY DIFFRACTION ANALYSIS AND SCANNING TUNNELING MICROSCOPY 486**

Types and Properties of Crystalline Solids 487  
Amorphous Solids 490  
Bonding in Solids: Molecular Orbital Band Theory 490

**12.7 Advanced Materials 493**

Electronic Materials 493  
Liquid Crystals 495  
Ceramic Materials 498  
Polymeric Materials 500  
Nanotechnology: Designing Materials Atom by Atom 505

CHAPTER REVIEW GUIDE 507

PROBLEMS 509



## CHAPTER 13 *The Properties of Mixtures: Solutions and Colloids* 516

- 13.1 Types of Solutions: Intermolecular Forces and Solubility** 518  
 Intermolecular Forces in Solution 518  
 Liquid Solutions and the Role of Molecular Polarity 519  
 Gas Solutions and Solid Solutions 521
- 13.2 Intermolecular Forces and Biological Macromolecules** 523  
 The Structures of Proteins 523  
 Dual Polarity in Soaps, Membranes, and Antibiotics 525  
 The Structure of DNA 527
- 13.3 Why Substances Dissolve: Breaking Down the Solution Process** 528  
 The Heat of Solution and Its Components 528
- The Heat of Hydration: Dissolving Ionic Solids in Water 529  
 The Solution Process and the Change in Entropy 531
- 13.4 Solubility as an Equilibrium Process** 533  
 Effect of Temperature on Solubility 534  
 Effect of Pressure on Solubility 535
- 13.5 Concentration Terms** 536  
 Molarity and Molality 537  
 Parts of Solute by Parts of Solution 538  
 Interconverting Concentration Terms 540
- 13.6 Colligative Properties of Solutions** 541  
 Nonvolatile Nonelectrolyte Solutions 542  
 Using Colligative Properties to Find Solute Molar Mass 547
- Volatile Nonelectrolyte Solutions 548  
 Strong Electrolyte Solutions 548  
 Applications of Colligative Properties 550
- 13.7 The Structure and Properties of Colloids** 552
- CHEMICAL CONNECTIONS TO ENVIRONMENTAL ENGINEERING: SOLUTIONS AND COLLOIDS IN WATER PURIFICATION* 554
- CHAPTER REVIEW GUIDE 556  
 PROBLEMS 559

## CHAPTER 14 *Periodic Patterns in the Main-Group Elements* 568

- 14.1 Hydrogen, the Simplest Atom** 569  
 Where Hydrogen Fits in the Periodic Table 569  
 Highlights of Hydrogen Chemistry 570
- 14.2 Trends Across the Periodic Table: The Period 2 Elements** 571
- 14.3 Group 1A(1): The Alkali Metals** 574  
 Why the Alkali Metals Are Unusual Physically 574  
 Why the Alkali Metals Are So Reactive 576
- 14.4 Group 2A(2): The Alkaline Earth Metals** 576  
 How the Alkaline Earth and Alkali Metals Compare Physically 577  
 How the Alkaline Earth and Alkali Metals Compare Chemically 577  
 Diagonal Relationships: Lithium and Magnesium 577
- 14.5 Group 3A(13): The Boron Family** 579  
 How the Transition Elements Influence This Group's Properties 579  
 Features That First Appear in This Group's Chemical Properties 579
- Highlights of Boron Chemistry 581  
 Diagonal Relationships: Beryllium and Aluminum 582
- 14.6 Group 4A(14): The Carbon Family** 582  
 How Type of Bonding Affects Physical Properties 582  
 How Bonding Changes in This Group's Compounds 585  
 Highlights of Carbon Chemistry 585  
 Highlights of Silicon Chemistry 587  
 Diagonal Relationships: Boron and Silicon 588
- 14.7 Group 5A(15): The Nitrogen Family** 588  
 The Wide Range of Physical Behavior 590  
 Patterns in Chemical Behavior 590  
 Highlights of Nitrogen Chemistry 591  
 Highlights of Phosphorus Chemistry 594
- 14.8 Group 6A(16): The Oxygen Family** 596  
 How the Oxygen and Nitrogen Families Compare Physically 596  
 How the Oxygen and Nitrogen Families Compare Chemically 598  
 Highlights of Oxygen Chemistry:  
 Range of Oxide Properties 599  
 Highlights of Sulfur Chemistry 599
- 14.9 Group 7A(17): The Halogens** 601  
 Physical Behavior of the Halogens 601  
 Why the Halogens Are So Reactive 601  
 Highlights of Halogen Chemistry 603
- 14.10 Group 8A(18): The Noble Gases** 606  
 How the Noble Gases and Alkali Metals Contrast Physically 608  
 How Noble Gases Can Form Compounds 608
- CHAPTER REVIEW GUIDE 608  
 PROBLEMS 609



## CHAPTER

## 15

## Organic Compounds and the Atomic Properties of Carbon 616

**15.1 The Special Nature of Carbon and the Characteristics of Organic Molecules 617**

The Structural Complexity of Organic Molecules 618

The Chemical Diversity of Organic Molecules 618

**15.2 The Structures and Classes of Hydrocarbons 620**

Carbon Skeletons and Hydrogen Skins 620

Alkanes: Hydrocarbons with Only Single Bonds 622

Dispersion Forces and the Physical Properties of Alkanes 625

Constitutional Isomerism 625

Chiral Molecules and Optical Isomerism 626

Alkenes: Hydrocarbons with Double Bonds 628

Restricted Rotation and Geometric (*Cis-Trans*) Isomerism 628

Alkynes: Hydrocarbons with Triple Bonds 630

Aromatic Hydrocarbons: Cyclic Molecules with Delocalized  $\pi$  Electrons 631

Variations on a Theme: Catenated Inorganic Hydrides 632

**TOOLS OF THE LABORATORY: NUCLEAR MAGNETIC RESONANCE (NMR) SPECTROSCOPY 633**

**15.3 Some Important Classes of Organic Reactions 635**

Types of Organic Reactions 635

The Redox Process in Organic Reactions 637

**15.4 Properties and Reactivities of Common Functional Groups 638**

Functional Groups with Only Single Bonds 638

Functional Groups with Double Bonds 643

Functional Groups with Both Single and Double Bonds 646

Functional Groups with Triple Bonds 650

**15.5 The Monomer-Polymer Theme I: Synthetic Macromolecules 652**

Addition Polymers 652

Condensation Polymers 653

**15.6 The Monomer-Polymer Theme II: Biological Macromolecules 654**

Sugars and Polysaccharides 654

Amino Acids and Proteins 656

Nucleotides and Nucleic Acids 658

**CHEMICAL CONNECTIONS TO GENETICS AND FORENSICS: DNA SEQUENCING AND FINGERPRINTING 663**

CHAPTER REVIEW GUIDE 665

PROBLEMS 667

## CHAPTER

## 16

## Kinetics: Rates and Mechanisms of Chemical Reactions 674

**16.1 Focusing on Reaction Rate 675****16.2 Expressing the Reaction Rate 678**

Average, Instantaneous, and Initial Reaction Rates 678

Expressing Rate in Terms of Reactant and Product Concentrations 680

**16.3 The Rate Law and Its Components 682**

Some Laboratory Methods for Determining the Initial Rate 683

Determining Reaction Orders 683

Determining the Rate Constant 688

**16.4 Integrated Rate Laws: Concentration Changes over Time 691**

Integrated Rate Laws for First-, Second-, and Zero-Order Reactions 692

Determining Reaction Orders from an Integrated Rate Law 693

Reaction Half-Life 695

**16.5 Theories of Chemical Kinetics 699**

Collision Theory: Basis of the Rate Law 699

Transition State Theory: What the Activation Energy Is Used For 703

**16.6 Reaction Mechanisms: The Steps from Reactant to Product 706**

Elementary Reactions and Molecularity 706

The Rate-Determining Step of a Reaction Mechanism 708

Correlating the Mechanism with the Rate Law 708

**16.7 Catalysis: Speeding Up a Reaction 712**

The Basis of Catalytic Action 712

Homogeneous Catalysis 713

Heterogeneous Catalysis 714

Kinetics and Function of Biological Catalysts 714

**CHEMICAL CONNECTIONS TO ATMOSPHERIC SCIENCE: DEPLETION OF EARTH'S OZONE LAYER 716**

CHAPTER REVIEW GUIDE 718

PROBLEMS 721



## CHAPTER

## 17

## Equilibrium: The Extent of Chemical Reactions 730

- 17.1 The Equilibrium State and the Equilibrium Constant 731**
- 17.2 The Reaction Quotient and the Equilibrium Constant 734**  
The Changing Value of the Reaction Quotient 734  
Writing the Reaction Quotient in Its Various Forms 735
- 17.3 Expressing Equilibria with Pressure Terms: Relation Between  $K_c$  and  $K_p$  740**
- 17.4 Comparing  $Q$  and  $K$  to Determine Reaction Direction 742**
- 17.5 How to Solve Equilibrium Problems 744**  
Using Quantities to Find the Equilibrium Constant 745  
Using the Equilibrium Constant to Find Quantities 747  
Problems Involving Mixtures of Reactants and Products 752
- 17.6 Reaction Conditions and Equilibrium: Le Châtelier's Principle 754**  
The Effect of a Change in Concentration 754  
The Effect of a Change in Pressure (Volume) 757  
The Effect of a Change in Temperature 759  
The Lack of Effect of a Catalyst 761  
Applying Le Châtelier's Principle to the Synthesis of Ammonia 763
- CHEMICAL CONNECTIONS TO CELLULAR METABOLISM: DESIGN AND CONTROL OF A METABOLIC PATHWAY 765**
- CHAPTER REVIEW GUIDE 766  
PROBLEMS 768

## CHAPTER

## 18

## Acid-Base Equilibria 776

- 18.1 Acids and Bases in Water 778**  
Release of  $H^+$  or  $OH^-$  and the Arrhenius Acid-Base Definition 778  
Variation in Acid Strength: The Acid-Dissociation Constant ( $K_a$ ) 779  
Classifying the Relative Strengths of Acids and Bases 781
- 18.2 Autoionization of Water and the pH Scale 782**  
The Equilibrium Nature of Autoionization: The Ion-Product Constant for Water ( $K_w$ ) 783  
Expressing the Hydronium Ion Concentration: The pH Scale 784
- 18.3 Proton Transfer and the Brønsted-Lowry Acid-Base Definition 787**  
Conjugate Acid-Base Pairs 788  
Relative Acid-Base Strength and the Net Direction of Reaction 789
- 18.4 Solving Problems Involving Weak-Acid Equilibria 792**  
Finding  $K_a$  Given Concentrations 793  
Finding Concentrations Given  $K_a$  794  
The Effect of Concentration on the Extent of Acid Dissociation 795  
The Behavior of Polyprotic Acids 797
- 18.5 Molecular Properties and Acid Strength 799**  
Acid Strength of Nonmetal Hydrides 799  
Acid Strength of Oxoacids 800  
Acidity of Hydrated Metal Ions 801
- 18.6 Weak Bases and Their Relation to Weak Acids 802**  
Molecules as Weak Bases: Ammonia and the Amines 802  
Anions of Weak Acids as Weak Bases 804  
The Relation Between  $K_a$  and  $K_b$  of a Conjugate Acid-Base Pair 805
- 18.7 Acid-Base Properties of Salt Solutions 807**  
Salts That Yield Neutral Solutions 807  
Salts That Yield Acidic Solutions 807  
Salts That Yield Basic Solutions 807  
Salts of Weakly Acidic Cations and Weakly Basic Anions 808  
Salts of Amphoteric Anions 809
- 18.8 Generalizing the Brønsted-Lowry Concept: The Leveling Effect 810**
- 18.9 Electron-Pair Donation and the Lewis Acid-Base Definition 811**  
Molecules as Lewis Acids 812  
Metal Cations as Lewis Acids 813  
An Overview of Acid-Base Definitions 814
- CHAPTER REVIEW GUIDE 814  
PROBLEMS 817



## CHAPTER

## 19

## Ionic Equilibria in Aqueous Systems 826

**19.1 Equilibria of Acid-Base Buffers 827**

What a Buffer Is and How It Works: The Common-Ion Effect 827  
 The Henderson-Hasselbalch Equation 832  
 Buffer Capacity and Buffer Range 833  
 Preparing a Buffer 835

**19.2 Acid-Base Titration Curves 837**

Monitoring pH with Acid-Base Indicators 837  
 Strong Acid–Strong Base Titration Curves 838  
 Weak Acid–Strong Base Titration Curves 840  
 Weak Base–Strong Acid Titration Curves 843

Titration Curves for Polyprotic Acids 844  
 Amino Acids as Biological Polyprotic Acids 845

**19.3 Equilibria of Slightly Soluble Ionic Compounds 846**

The Ion-Product Expression ( $Q_{sp}$ ) and the Solubility-Product Constant ( $K_{sp}$ ) 846  
 Calculations Involving the Solubility-Product Constant 848  
 Effect of a Common Ion on Solubility 850  
 Effect of pH on Solubility 852  
 Applying Ionic Equilibria to the Formation of a Limestone Cave 853  
 Predicting the Formation of a Precipitate:  $Q_{sp}$  vs.  $K_{sp}$  853

Separating Ions by Selective Precipitation and Simultaneous Equilibria 856

**CHEMICAL CONNECTIONS TO ENVIRONMENTAL SCIENCE: THE ACID-RAIN PROBLEM 857**

**19.4 Equilibria Involving Complex Ions 859**

Formation of Complex Ions 859  
 Complex Ions and the Solubility of Precipitates 861  
 Complex Ions of Amphoteric Hydroxides 863

CHAPTER REVIEW GUIDE 865

PROBLEMS 868

## CHAPTER

## 20

## Thermodynamics: Entropy, Free Energy, and the Direction of Chemical Reactions 876

**20.1 The Second Law of Thermodynamics: Predicting Spontaneous Change 877**

The First Law of Thermodynamics Does Not Predict Spontaneous Change 878  
 The Sign of  $\Delta H$  Does Not Predict Spontaneous Change 878  
 Freedom of Particle Motion and Dispersal of Kinetic Energy 879  
 Entropy and the Number of Microstates 880  
 Entropy and the Second Law of Thermodynamics 883  
 Standard Molar Entropies and the Third Law 883  
 Predicting Relative  $S^\circ$  of a System 883

**20.2 Calculating the Change in Entropy of a Reaction 888**

Entropy Changes in the System: Standard Entropy of Reaction ( $\Delta S_{rxn}^\circ$ ) 888  
 Entropy Changes in the Surroundings: The Other Part of the Total 889  
 The Entropy Change and the Equilibrium State 891  
 Spontaneous Exothermic and Endothermic Changes 892

**20.3 Entropy, Free Energy, and Work 893**

Free Energy Change and Reaction Spontaneity 893  
 Calculating Standard Free Energy Changes 894

The Free Energy Change and the Work a System Can Do 896

The Effect of Temperature on Reaction Spontaneity 897  
 Coupling of Reactions to Drive a Nonspontaneous Change 901

**CHEMICAL CONNECTIONS TO BIOLOGICAL ENERGETICS: THE UNIVERSAL ROLE OF ATP 902**

**20.4 Free Energy, Equilibrium, and Reaction Direction 903**

CHAPTER REVIEW GUIDE 909

PROBLEMS 912



## CHAPTER 21 *Electrochemistry: Chemical Change and Electrical Work* 918

### 21.1 Redox Reactions and Electrochemical Cells 919

A Quick Review of Oxidation-Reduction Concepts 919

Half-Reaction Method for Balancing Redox Reactions 920

An Overview of Electrochemical Cells 924

### 21.2 Voltaic Cells: Using Spontaneous Reactions to Generate Electrical Energy 925

Construction and Operation of a Voltaic Cell 926

Notation for a Voltaic Cell 928

Why Does a Voltaic Cell Work? 929

### 21.3 Cell Potential: Output of a Voltaic Cell 930

Standard Cell Potential ( $E^\circ_{\text{cell}}$ ) 930

Relative Strengths of Oxidizing and Reducing Agents 933

Using  $E^\circ_{\text{half-cell}}$  Values to Write Spontaneous Redox Reactions 934

Explaining the Activity Series of the Metals 937

### 21.4 Free Energy and Electrical Work 939

Standard Cell Potential and the Equilibrium Constant 939

The Effect of Concentration on Cell Potential 941

Following Changes in Potential During Cell Operation 943

Concentration Cells 944

### 21.5 Electrochemical Processes in Batteries 947

Primary (Nonrechargeable) Batteries 947

Secondary (Rechargeable) Batteries 949

Fuel Cells 950

### 21.6 Corrosion: An Environmental Voltaic Cell 951

The Corrosion of Iron 951

Protecting Against the Corrosion of Iron 953

### 21.7 Electrolytic Cells: Using Electrical Energy to Drive Nonspontaneous Reactions 954

Construction and Operation of an Electrolytic Cell 954

Predicting the Products of Electrolysis 955

Stoichiometry of Electrolysis: The Relation Between Amounts of Charge and Products 959

**CHEMICAL CONNECTIONS TO BIOLOGICAL ENERGETICS: CELLULAR ELECTROCHEMISTRY AND THE PRODUCTION OF ATP** 962

CHAPTER REVIEW GUIDE 964

PROBLEMS 967

## CHAPTER 22 *The Elements in Nature and Industry* 976

### 22.1 How the Elements Occur in Nature 977

Earth's Structure and the Abundance of the Elements 977

Sources of the Elements 981

### 22.2 The Cycling of Elements Through the Environment 982

The Carbon Cycle 982

The Nitrogen Cycle 984

The Phosphorus Cycle 986

### 22.3 Metallurgy: Extracting a Metal from Its Ore 988

Pretreating the Ore 989

Converting Mineral to Element 990

Refining and Alloying the Element 992

### 22.4 Tapping the Crust: Isolation and Uses of Selected Elements 994

Producing the Alkali Metals: Sodium and Potassium 994

The Indispensable Three: Iron, Copper, and Aluminum 996

Mining the Sea for Magnesium 1002

The Sources and Uses of Hydrogen 1003

### 22.5 Chemical Manufacturing: Two Case Studies 1006

Sulfuric Acid, the Most Important Chemical 1006

The Chlor-Alkali Process 1008

CHAPTER REVIEW GUIDE 1009

PROBLEMS 1010



## CHAPTER

## 23

## Transition Elements and Their Coordination Compounds 1016

**23.1 Properties of the Transition Elements 1017**

Electron Configurations of the Transition Metals and Their Ions 1018  
Atomic and Physical Properties of the Transition Elements 1020  
Chemical Properties of the Transition Elements 1022

**23.2 The Inner Transition Elements 1024**

The Lanthanides 1024  
The Actinides 1025

**23.3 Coordination Compounds 1026**

Complex Ions: Coordination Numbers, Geometries, and Ligands 1026  
Formulas and Names of Coordination Compounds 1028  
Isomerism in Coordination Compounds 1031

**23.4 Theoretical Basis for the Bonding and Properties of Complexes 1034**

Applying Valence Bond Theory to Complex Ions 1034  
Crystal Field Theory 1036

**CHEMICAL CONNECTIONS TO NUTRITIONAL SCIENCE: TRANSITION METALS AS ESSENTIAL DIETARY TRACE ELEMENTS 1043**

CHAPTER REVIEW GUIDE 1045  
PROBLEMS 1046

## CHAPTER

## 24

## Nuclear Reactions and Their Applications 1052

**24.1 Radioactive Decay and Nuclear Stability 1053**

Comparing Chemical and Nuclear Change 1054  
The Components of the Nucleus: Terms and Notation 1054  
The Discovery of Radioactivity and the Types of Emissions 1055  
Modes of Radioactive Decay; Balancing Nuclear Equations 1055  
Nuclear Stability and the Mode of Decay 1058

**24.2 The Kinetics of Radioactive Decay 1062**

Detection and Measurement of Radioactivity 1063  
The Rate of Radioactive Decay 1064  
Radioisotopic Dating 1067

**24.3 Nuclear Transmutation: Induced Changes in Nuclei 1069**

Early Transmutation Experiments; Nuclear Shorthand Notation 1070  
Particle Accelerators and the Transuranium Elements 1070

**24.4 Effects of Nuclear Radiation on Matter 1072**

Effects of Ionizing Radiation on Living Tissue 1073  
Background Sources of Ionizing Radiation 1075  
Assessing the Risk from Ionizing Radiation 1076

**24.5 Applications of Radioisotopes 1077**

Radioactive Tracers 1077  
Additional Applications of Ionizing Radiation 1080

**24.6 The Interconversion of Mass and Energy 1081**

The Mass Difference Between a Nucleus and Its Nucleons 1081  
Nuclear Binding Energy and Binding Energy per Nucleon 1082

**24.7 Applications of Fission and Fusion 1084**

The Process of Nuclear Fission 1084  
The Promise of Nuclear Fusion 1088

**CHEMICAL CONNECTIONS TO COSMOLOGY: ORIGIN OF THE ELEMENTS IN THE STARS 1090**

CHAPTER REVIEW GUIDE 1092  
PROBLEMS 1094

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

**Appendix B** Standard Thermodynamic Values for Selected Substances A-5

**Appendix C** Equilibrium Constants for Selected Substances A-8

**Appendix D** Standard Electrode (Half-Cell) Potentials A-14

**Appendix E** Answers to Selected Problems A-15

**Glossary** G-1

**Credits** C-1

**Index** I-1



## LIST OF SAMPLE PROBLEMS *(Molecular-scene problems are shown in color.)*

### Chapter 1

- 1.1 Visualizing Change on the Atomic Scale 6
- 1.2 Distinguishing Between Physical and Chemical Change 7
- 1.3 Converting Units of Length 20
- 1.4 Converting Units of Volume 21
- 1.5 Converting Units of Mass 21
- 1.6 Converting Units Raised to a Power 22
- 1.7 Calculating Density from Mass and Volume 24
- 1.8 Converting Units of Temperature 26
- 1.9 Determining the Number of Significant Figures 28
- 1.10 Significant Figures and Rounding 30

### Chapter 2

- 2.1 Distinguishing Elements, Compounds, and Mixtures at the Atomic Scale 43
- 2.2 Calculating the Mass of an Element in a Compound 46
- 2.3 Visualizing the Mass Laws 48
- 2.4 Determining the Number of Subatomic Particles in the Isotopes of an Element 54
- 2.5 Calculating the Atomic Mass of an Element 55
- 2.6 Identifying an Element from Its *Z* Value 60
- 2.7 Predicting the Ion an Element Forms 63
- 2.8 Naming Binary Ionic Compounds 66
- 2.9 Determining Formulas of Binary Ionic Compounds 67
- 2.10 Determining Names and Formulas of Ionic Compounds of Metals That Form More Than One Ion 68
- 2.11 Determining Names and Formulas of Ionic Compounds Containing Polyatomic Ions 69
- 2.12 Recognizing Incorrect Names and Formulas of Ionic Compounds 69
- 2.13 Determining Names and Formulas of Anions and Acids 70
- 2.14 Determining Names and Formulas of Binary Covalent Compounds 71
- 2.15 Recognizing Incorrect Names and Formulas of Binary Covalent Compounds 72
- 2.16 Calculating the Molecular Mass of a Compound 73
- 2.17 Using Molecular Depictions to Determine Formula, Name, and Mass 74

### Chapter 3

- 3.1 Converting Between Mass and Amount of an Element 94
- 3.2 Converting Between Number of Entities and Amount of an Element 95
- 3.3 Converting Between Number of Entities and Mass of an Element 95
- 3.4 Converting Between Number of Entities and Mass of a Compound I 96
- 3.5 Converting Between Number of Entities and Mass of a Compound II 97
- 3.6 Calculating the Mass Percent of Each Element in a Compound from the Formula 98
- 3.7 Calculating the Mass of an Element in a Compound 100
- 3.8 Determining an Empirical Formula from Amounts of Elements 101
- 3.9 Determining an Empirical Formula from Masses of Elements 102
- 3.10 Determining a Molecular Formula from Elemental Analysis and Molar Mass 103
- 3.11 Determining a Molecular Formula from Combustion Analysis 104
- 3.12 Balancing Chemical Equations 110
- 3.13 Balancing an Equation from a Molecular Scene 111
- 3.14 Calculating Quantities of Reactants and Products: Amount (mol) to Amount (mol) 114

- 3.15 Calculating Quantities of Reactants and Products: Amount (mol) to Mass (g) 114
- 3.16 Calculating Quantities of Reactants and Products: Mass to Mass 115
- 3.17 Writing an Overall Equation for a Reaction Sequence 116
- 3.18 Using Molecular Depictions in a Limiting-Reactant Problem 119
- 3.19 Calculating Quantities in a Limiting-Reactant Problem: Amount to Amount 120
- 3.20 Calculating Quantities in a Limiting-Reactant Problem: Mass to Mass 121
- 3.21 Calculating Percent Yield 123

### Chapter 4

- 4.1 Using Molecular Scenes to Depict an Ionic Compound in Aqueous Solution 142
- 4.2 Determining Amount (mol) of Ions in Solution 143
- 4.3 Calculating the Molarity of a Solution 144
- 4.4 Calculating Mass of Solute in a Given Volume of Solution 145
- 4.5 Determining Amount (mol) of Ions in a Solution 145
- 4.6 Preparing a Dilute Solution from a Concentrated Solution 147
- 4.7 Visualizing Changes in Concentration 148
- 4.8 Predicting Whether a Precipitation Reaction Occurs: Writing Ionic Equations 153
- 4.9 Using Molecular Depictions in Precipitation Reactions 154
- 4.10 Calculating Amounts of Reactants and Products in a Precipitation Reaction 156
- 4.11 Solving a Limiting-Reactant Problem for a Precipitation Reaction 157
- 4.12 Determining the Number of  $H^+$  (or  $OH^-$ ) Ions in Solution 160
- 4.13 Writing Ionic Equations for Acid-Base Reactions 161
- 4.14 Writing Proton-Transfer Equations for Acid-Base Reactions 165
- 4.15 Calculating the Amounts of Reactants and Products in an Acid-Base Reaction 165
- 4.16 Finding the Concentration of an Acid from a Titration 167
- 4.17 Determining the Oxidation Number of Each Element in a Compound (or Ion) 170
- 4.18 Identifying Redox Reactions and Oxidizing and Reducing Agents 171
- 4.19 Finding the Amount of Reducing Agent by Titration 173
- 4.20 Identifying the Type of Redox Reaction 180

### Chapter 5

- 5.1 Converting Units of Pressure 203
- 5.2 Applying the Volume-Pressure Relationship 211
- 5.3 Applying the Volume-Temperature and Pressure-Temperature Relationships 211
- 5.4 Applying the Volume-Amount and Pressure-Amount Relationships 212
- 5.5 Applying the Volume-Pressure-Temperature Relationship 213
- 5.6 Solving for an Unknown Gas Variable at Fixed Conditions 214
- 5.7 Using Gas Laws to Determine a Balanced Equation 214
- 5.8 Calculating Gas Density 216
- 5.9 Finding the Molar Mass of a Volatile Liquid 218
- 5.10 Applying Dalton's Law of Partial Pressures 220
- 5.11 Calculating the Amount of Gas Collected over Water 221
- 5.12 Using Gas Variables to Find Amounts of Reactants or Products I 222
- 5.13 Using Gas Variables to Find Amounts of Reactants or Products II 223
- 5.14 Applying Graham's Law of Effusion 230

## Chapter 6

- 6.1 Determining the Change in Internal Energy of a System 256
- 6.2 Calculating Pressure-Volume Work Done by or on a System 258
- 6.3 Drawing Enthalpy Diagrams and Determining the Sign of  $\Delta H$  260
- 6.4 Finding the Quantity of Heat from a Temperature Change 262
- 6.5 Determining the Specific Heat Capacity of a Solid 263
- 6.6 Determining the Enthalpy Change of an Aqueous Reaction 263
- 6.7 Calculating the Heat of a Combustion Reaction 265
- 6.8 Using the Enthalpy Change of a Reaction ( $\Delta H$ ) to Find the Amount of a Substance 267
- 6.9 Using Hess's Law to Calculate an Unknown  $\Delta H$  269
- 6.10 Writing Formation Equations 270
- 6.11 Calculating  $\Delta H_{\text{rxn}}^{\circ}$  from  $\Delta H_f^{\circ}$  Values 272

## Chapter 7

- 7.1 Interconverting Wavelength and Frequency 289
- 7.2 Interconverting Energy, Wavelength, and Frequency 293
- 7.3 Determining  $\Delta E$  and  $\lambda$  of an Electron Transition 298
- 7.4 Calculating the de Broglie Wavelength of an Electron 303
- 7.5 Applying the Uncertainty Principle 305
- 7.6 Determining Quantum Numbers for an Energy Level 309
- 7.7 Determining Sublevel Names and Orbital Quantum Numbers 310
- 7.8 Identifying Incorrect Quantum Numbers 310

## Chapter 8

- 8.1 Correlating Quantum Numbers and Orbital Diagrams 329
- 8.2 Determining Electron Configurations 335
- 8.3 Ranking Elements by Atomic Size 338
- 8.4 Ranking Elements by First Ionization Energy 341
- 8.5 Identifying an Element from Its Ionization Energies 342
- 8.6 Writing Electron Configurations of Main-Group Ions 347
- 8.7 Writing Electron Configurations and Predicting Magnetic Behavior of Transition Metal Ions 350
- 8.8 Ranking Ions by Size 351

## Chapter 9

- 9.1 Depicting Ion Formation 362
- 9.2 Predicting Relative Lattice Energy from Ionic Properties 366
- 9.3 Comparing Bond Length and Bond Strength 372
- 9.4 Using Bond Energies to Calculate  $\Delta H_{\text{rxn}}^{\circ}$  378
- 9.5 Determining Bond Polarity from EN Values 383

## Chapter 10

- 10.1 Writing Lewis Structures for Species with One Central Atom 397
- 10.2 Writing Lewis Structures for Molecules with More Than One Central Atom 398
- 10.3 Writing Lewis Structures for Molecules with Multiple Bonds 398
- 10.4 Writing Resonance Structures 400
- 10.5 Writing Lewis Structures for Octet-Rule Exceptions 405
- 10.6 Examining Shapes with Two, Three, or Four Electron Groups 414
- 10.7 Examining Shapes with Five or Six Electron Groups 414
- 10.8 Predicting Molecular Shapes with More Than One Central Atom 416
- 10.9 Predicting the Polarity of Molecules 418

## Chapter 11

- 11.1 Postulating Hybrid Orbitals in a Molecule 435
- 11.2 Describing the Types of Bonds in Molecules 438
- 11.3 Predicting Stability of Species Using MO Diagrams 442
- 11.4 Using MO Theory to Explain Bond Properties 446

## Chapter 12

- 12.1 Finding the Heat of a Phase Change Depicted by Molecular Scenes 461
- 12.2 Applying the Clausius-Clapeyron Equation 464
- 12.3 Using a Phase Diagram to Predict Phase Changes 467
- 12.4 Drawing Hydrogen Bonds Between Molecules of a Substance 471
- 12.5 Identifying the Types of Intermolecular Forces 473
- 12.6 Determining the Number of Particles per Unit Cell and the Coordination Number 480
- 12.7 Determining Atomic Radius 484
- 12.8 Determining Atomic Radius from the Unit Cell 485

## Chapter 13

- 13.1 Predicting Relative Solubilities 521
- 13.2 Calculating an Aqueous Ionic Heat of Solution 531
- 13.3 Using Henry's Law to Calculate Gas Solubility 536
- 13.4 Calculating Molality 537
- 13.5 Expressing Concentrations in Parts by Mass, Parts by Volume, and Mole Fraction 539
- 13.6 Interconverting Concentration Terms 540
- 13.7 Using Raoult's Law to Find  $\Delta P$  543
- 13.8 Determining Boiling and Freezing Points of a Solution 545
- 13.9 Determining Molar Mass from Osmotic Pressure 547
- 13.10 Depicting Strong Electrolyte Solutions 549

## Chapter 15

- 15.1 Drawing Hydrocarbons 621
- 15.2 Naming Alkanes, Alkenes, and Alkynes 630
- 15.3 Recognizing the Type of Organic Reaction 636
- 15.4 Predicting the Reactions of Alcohols, Alkyl Halides, and Amines 642
- 15.5 Predicting the Steps in a Reaction Sequence 645
- 15.6 Predicting Reactions of the Carboxylic Acid Family 649
- 15.7 Recognizing Functional Groups 651

## Chapter 16

- 16.1 Expressing Rate in Terms of Changes in Concentration with Time 681
- 16.2 Determining Reaction Orders from Rate Laws 685
- 16.3 Determining Reaction Orders and Rate Constants from Rate Data 689
- 16.4 Determining Reaction Orders from Molecular Scenes 690
- 16.5 Determining the Reactant Concentration after a Given Time 692
- 16.6 Using Molecular Scenes to Find Quantities at Various Times 696
- 16.7 Determining the Half-Life of a First-Order Reaction 697
- 16.8 Determining the Energy of Activation 702
- 16.9 Drawing Reaction Energy Diagrams and Transition States 705
- 16.10 Determining Molarities and Rate Laws for Elementary Steps 707
- 16.11 Identifying Intermediates and Correlating Rate Laws and Reaction Mechanisms 710

## Chapter 17

- 17.1 Writing the Reaction Quotient from the Balanced Equation 736
- 17.2 Writing the Reaction Quotient and Finding  $K$  for an Overall Reaction 737
- 17.3 Finding the Equilibrium Constant for an Equation Multiplied by a Common Factor 739
- 17.4 Converting Between  $K_c$  and  $K_p$  741
- 17.5 Using Molecular Scenes to Determine Reaction Direction 743
- 17.6 Using Concentrations to Determine Reaction Direction 744

- 17.7 Calculating  $K_c$  from Concentration Data 746
- 17.8 Determining Equilibrium Concentrations from  $K_c$  747
- 17.9 Determining Equilibrium Concentrations from Initial Concentrations and  $K_c$  748
- 17.10 Making a Simplifying Assumption to Calculate Equilibrium Concentrations 750
- 17.11 Predicting Reaction Direction and Calculating Equilibrium Concentrations 752
- 17.12 Predicting the Effect of a Change in Concentration on the Equilibrium Position 756
- 17.13 Predicting the Effect of a Change in Volume (Pressure) on the Equilibrium Position 758
- 17.14 Predicting the Effect of a Change in Temperature on the Equilibrium Position 760
- 17.15 Determining Equilibrium Parameters from Molecular Scenes 762

## Chapter 18

- 18.1 Classifying Acid and Base Strength from the Chemical Formula 782
- 18.2 Calculating  $[\text{H}_3\text{O}^+]$  or  $[\text{OH}^-]$  in Aqueous Solution 784
- 18.3 Calculating  $[\text{H}_3\text{O}^+]$ , pH,  $[\text{OH}^-]$ , and pOH for Strong Acids and Bases 786
- 18.4 Identifying Conjugate Acid-Base Pairs 789
- 18.5 Predicting the Net Direction of an Acid-Base Reaction 790
- 18.6 Using Molecular Scenes to Predict the Net Direction of an Acid-Base Reaction 791
- 18.7 Finding  $K_a$  of a Weak Acid from the Solution pH 793
- 18.8 Determining Concentration and pH from  $K_a$  and Initial  $[\text{HA}]$  794
- 18.9 Finding the Percent Dissociation of a Weak Acid 796
- 18.10 Calculating Equilibrium Concentrations for a Polyprotic Acid 798
- 18.11 Determining pH from  $K_b$  and Initial  $[\text{B}]$  803
- 18.12 Determining the pH of a Solution of  $\text{A}^-$  805
- 18.13 Predicting Relative Acidity of Salt Solutions from Reactions of the Ions with Water 808
- 18.14 Predicting the Relative Acidity of a Salt Solution from  $K_a$  and  $K_b$  of the Ions 810
- 18.15 Identifying Lewis Acids and Bases 813

## Chapter 19

- 19.1 Calculating the Effect of Added  $\text{H}_3\text{O}^+$  or  $\text{OH}^-$  on Buffer pH 830
- 19.2 Using Molecular Scenes to Examine Buffers 834
- 19.3 Preparing a Buffer 835
- 19.4 Finding the pH During a Weak Acid–Strong Base Titration 842
- 19.5 Writing Ion-Product Expressions 847
- 19.6 Determining  $K_{sp}$  from Solubility 848
- 19.7 Determining Solubility from  $K_{sp}$  849
- 19.8 Calculating the Effect of a Common Ion on Solubility 851
- 19.9 Predicting the Effect on Solubility of Adding Strong Acid 852
- 19.10 Predicting Whether a Precipitate Will Form 854
- 19.11 Using Molecular Scenes to Predict Whether a Precipitate Will Form 854
- 19.12 Separating Ions by Selective Precipitation 856
- 19.13 Calculating the Concentration of a Complex Ion 860
- 19.14 Calculating the Effect of Complex-Ion Formation on Solubility 862

## Chapter 20

- 20.1 Predicting Relative Entropy Values 887
- 20.2 Calculating the Standard Entropy of Reaction,  $\Delta S_{\text{rxn}}^\circ$  889
- 20.3 Determining Reaction Spontaneity 890
- 20.4 Calculating  $\Delta G_{\text{rxn}}^\circ$  from Enthalpy and Entropy Values 894
- 20.5 Calculating  $\Delta G_{\text{rxn}}^\circ$  from  $\Delta G_f^\circ$  Values 896
- 20.6 Using Molecular Scenes to Determine the Signs of  $\Delta H$ ,  $\Delta S$ , and  $\Delta G$  898
- 20.7 Determining the Effect of Temperature on  $\Delta G$  899
- 20.8 Finding the Temperature at Which a Reaction Becomes Spontaneous 900
- 20.9 Exploring the Relationship Between  $\Delta G^\circ$  and  $K$  904
- 20.10 Using Molecular Scenes to Find  $\Delta G$  for a Reaction at Nonstandard Conditions 905
- 20.11 Calculating  $\Delta G$  at Nonstandard Conditions 907

## Chapter 21

- 21.1 Balancing a Redox Reaction in Basic Solution 922
- 21.2 Describing a Voltaic Cell with Diagram and Notation 928
- 21.3 Using  $E_{\text{half-cell}}^\circ$  to Find  $E_{\text{cell}}^\circ$  931
- 21.4 Calculating an Unknown  $E_{\text{half-cell}}^\circ$  from  $E_{\text{cell}}^\circ$  933
- 21.5 Writing Spontaneous Redox Reactions and Ranking Oxidizing and Reducing Agents by Strength 936
- 21.6 Calculating  $K$  and  $\Delta G^\circ$  from  $E_{\text{cell}}^\circ$  940
- 21.7 Using the Nernst Equation to Calculate  $E_{\text{cell}}$  942
- 21.8 Calculating the Potential of a Concentration Cell 945
- 21.9 Predicting the Electrolysis Products of a Molten Salt Mixture 956
- 21.10 Predicting the Electrolysis Products of Aqueous Salt Solutions 958
- 21.11 Applying the Relationship Among Current, Time, and Amount of Substance 960

## Chapter 23

- 23.1 Writing Electron Configurations of Transition Metal Atoms and Ions 1020
- 23.2 Finding the Number of Unpaired Electrons 1025
- 23.3 Finding the Coordination Number and Charge of the Central Metal Ion in a Coordination Compound 1029
- 23.4 Writing Names and Formulas of Coordination Compounds 1030
- 23.5 Determining the Type of Stereoisomerism 1034
- 23.6 Ranking Crystal Field Splitting Energies ( $\Delta$ ) for Complex Ions of a Metal 1040
- 23.7 Identifying High-Spin and Low-Spin Complex Ions 1041

## Chapter 24

- 24.1 Writing Equations for Nuclear Reactions 1058
- 24.2 Predicting Nuclear Stability 1060
- 24.3 Predicting the Mode of Nuclear Decay 1061
- 24.4 Calculating the Specific Activity and the Decay Constant of a Radionuclide 1064
- 24.5 Finding the Number of Radioactive Nuclei 1066
- 24.6 Applying Radiocarbon Dating 1069
- 24.7 Calculating the Binding Energy per Nucleon 1083

# ABOUT THE AUTHORS



**Martin S. Silberberg** received a B.S. in Chemistry from the City University of New York and a Ph.D. in Chemistry from the University of Oklahoma. He then accepted a position as research associate in analytical biochemistry at the Albert Einstein College of Medicine in New York City, where he developed methods to study neurotransmitter metabolism in Parkinson's disease and other neurological disorders. Following six years in neurochemical research, Dr. Silberberg joined the faculty of Bard College at Simon's Rock, a liberal arts college known for its excellence in teaching small classes of highly motivated students. As head of the Natural Sciences Major and Director of Premedical Studies, he taught courses in general chemistry, organic chemistry, biochemistry, and liberal-arts chemistry. The small class size and close student contact afforded him insights into how students learn chemistry, where they have difficulties, and what strategies can help them succeed. Dr. Silberberg decided to apply these insights in a broader context and established a textbook writing, editing, and consulting company. Before writing his own texts, he worked as a consulting and development editor on chemistry, biochemistry, and physics texts for several major college publishers. He resides with his wife Ruth in the Pioneer Valley near Amherst, Massachusetts, where he enjoys the rich cultural and academic life of the area and relaxes by cooking, gardening, and singing.



**Patricia G. Amateis** graduated with a B.S. in Chemistry Education from Concord University in West Virginia and a Ph.D. in Analytical Chemistry from Virginia Tech. She has been on the faculty of the Chemistry Department at Virginia Tech for 28 years, teaching General Chemistry and Analytical Chemistry. For the past 13 years, she has served as Director of General Chemistry, responsible for the oversight of both the lecture and lab portions of the large General Chemistry program. She has taught thousands of students during her career and has been awarded the University Sporn Award for Introductory Teaching, the Alumni Teaching Award, and the William E. Wine Award for a history of university teaching excellence. She and her husband live in Blacksburg, Virginia and are the parents of three adult children. In her free time, she enjoys biking, hiking, competing in the occasional sprint triathlon, and playing the double second in Panjammers, Blacksburg's steel drum band.

# PREFACE

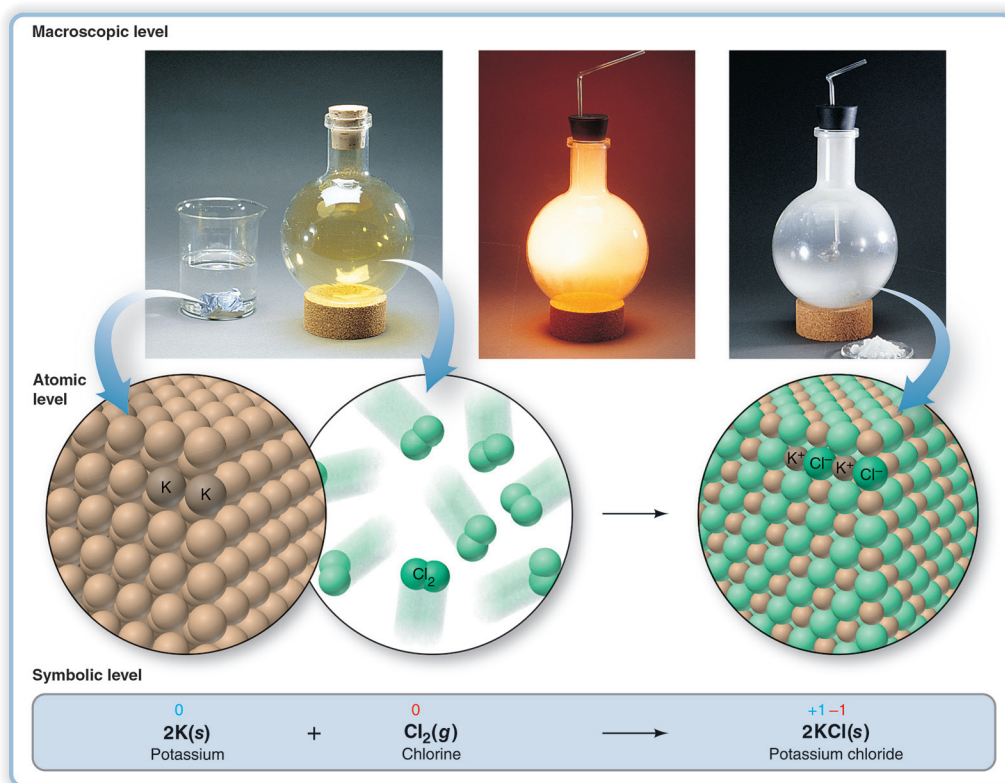
Chemistry is so crucial to an understanding of medicine and biology, environmental science, and many areas of engineering and industrial processing that it has become a requirement for an increasing number of academic majors. Furthermore, chemical principles lie at the core of some of the key societal issues we face in the 21<sup>st</sup> century—dealing with climate change, finding new energy options, and supplying nutrition and curing disease on an ever more populated planet.

## SETTING THE STANDARD FOR A CHEMISTRY TEXT

The seventh edition of *Chemistry: The Molecular Nature of Matter and Change* maintains its standard-setting position among general chemistry textbooks by evolving further to meet the needs of professor and student. The text still contains the most accurate molecular illustrations, consistent step-by-step worked problems, and an extensive collection of end-of-chapter problems. And changes throughout this edition make the text more readable and succinct, the artwork more teachable and modern, and the design more focused and inviting. The three hallmarks that have made this text a market leader are now demonstrated in its pages more clearly than ever.

## Visualizing Chemical Models—Macroscopic to Molecular

Chemistry deals with observable changes caused by unobservable atomic-scale events, requiring an appreciation of a size gap of mind-boggling proportions. One of the text's goals coincides with that of so many instructors: to help students visualize chemical events on the molecular scale. Thus, concepts are explained first at the macroscopic level and then from a molecular point of view, with pedagogic illustrations always placed next to the discussions to bring the point home for today's visually oriented students.



## Thinking Logically to Solve Problems

The problem-solving approach, based on the four-step method widely accepted by experts in chemical education, is introduced in Chapter 1 and employed *consistently* throughout the text. It encourages students to *plan* a logical approach to a problem, and only then proceed to *solve* it. Each sample problem includes a *check*, which fosters the habit of “thinking through” both the chemical and the quantitative reasonableness of the answer. Finally, for *practice* and reinforcement, each sample problem is now followed immediately by *two* similar follow-up problems. And, *Chemistry* marries problem solving to visualizing models with molecular-scene problems, which appear not only in homework sets, as in other texts, but also in the running text, where they are worked out stepwise.

### SAMPLE PROBLEM 3.9

### Determining an Empirical Formula from Masses of Elements

**Problem** Analysis of a sample of an ionic compound yields 2.82 g of Na, 4.35 g of Cl, and 7.83 g of O. What are the empirical formula and the name of the compound?

**Plan** This problem is similar to Sample Problem 3.8, except that we are given element masses that we must convert into integer subscripts. We first divide each mass by the element’s molar mass to find the amount (mol). Then we construct a preliminary formula and convert the amounts (mol) to integers.

**Solution** Finding amount (mol) of each element:

$$\text{Amount (mol) of Na} = 2.82 \text{ g Na} \times \frac{1 \text{ mol Na}}{22.99 \text{ g Na}} = 0.123 \text{ mol Na}$$

$$\text{Amount (mol) of Cl} = 4.35 \text{ g Cl} \times \frac{1 \text{ mol Cl}}{35.45 \text{ g Cl}} = 0.123 \text{ mol Cl}$$

$$\text{Amount (mol) of O} = 7.83 \text{ g O} \times \frac{1 \text{ mol O}}{16.00 \text{ g O}} = 0.489 \text{ mol O}$$

Constructing a preliminary formula:  $\text{Na}_{0.123}\text{Cl}_{0.123}\text{O}_{0.489}$

Converting to integer subscripts (dividing all by the smallest subscript):



The empirical formula is  $\text{NaClO}_4$ ; the name is sodium perchlorate.

**Check** The numbers of moles seem correct because the masses of Na and Cl are slightly more than 0.1 of their molar masses. The mass of O is greatest and its molar mass is smallest, so it should have the greatest number of moles. The ratio of subscripts, 1/1/4, is the same as the ratio of moles, 0.123/0.123/0.489 (within rounding).

#### FOLLOW-UP PROBLEMS

**3.9A** A sample of an unknown compound is found to contain 1.23 g of H, 12.64 g of P, and 26.12 g of O. What is the empirical formula?

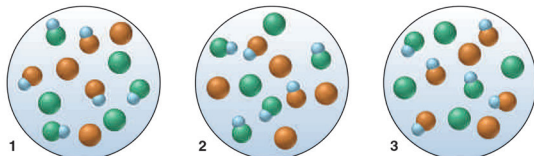
**3.9B** An unknown metal M reacts with sulfur to form a compound with the formula  $\text{M}_2\text{S}_3$ . If 3.12 g of M reacts with 2.88 g of S, what are the names of M and  $\text{M}_2\text{S}_3$ ? [*Hint*: Determine the amount (mol) of S, and use the formula to find the amount (mol) of M.]

**SOME SIMILAR PROBLEMS** 3.42(b), 3.43(b), 3.46(b), and 3.47(b)

### SAMPLE PROBLEM 18.6

### Using Molecular Scenes to Predict the Net Direction of an Acid-Base Reaction

**Problem** Given that 0.10 M HX (blue and green) has a pH of 2.88, and 0.10 M HY (blue and orange) has a pH of 3.52, which scene best represents the final mixture after equimolar solutions of HX and  $\text{Y}^-$  are mixed?



**Plan** A stronger acid and base yield a weaker acid and base, so we have to determine the relative acid strengths of HX and HY in order to choose the correct molecular scene. The concentrations of the acid solutions are equal, so we can pick the stronger acid directly from the pH values of the two acid solutions. Because the stronger acid reacts to a greater extent, fewer molecules of it will be in the scene than molecules of the weaker acid.

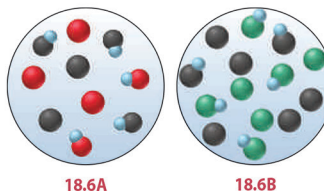
**Solution** The HX solution has a lower pH (2.88) than the HY solution (3.52), so we know right away that HX is the stronger acid and  $\text{Y}^-$  is the stronger base. Therefore, the reaction of HX and  $\text{Y}^-$  has a  $K_c > 1$ , which means the equilibrium mixture will have more HY than HX. Scene 1 has equal numbers of HX and HY, which would occur if the acids were of equal strength, and scene 2 shows fewer HY than HX, which would occur if HY were stronger. Therefore, only scene 3 is consistent with the relative acid strengths.

#### FOLLOW-UP PROBLEMS

**18.6A** The left-hand scene in the margin represents the equilibrium mixture after 0.10 M solutions of HA (blue and red) and  $\text{B}^-$  (black) react: Does this reaction have a  $K_c$  greater or less than 1? Which acid is stronger, HA or HB?

**18.6B** The right-hand scene depicts an aqueous solution of two conjugate acid-base pairs:  $\text{HC/C}^-$  and  $\text{HD/D}^-$ . HD is a stronger acid than HC. What colors represent the base  $\text{C}^-$  and the base  $\text{D}^-$ ? Does the reaction between HC and  $\text{D}^-$  have a  $K_c$  greater or less than 1?

**A SIMILAR PROBLEM** 18.39



## Applying Ideas to the Real World

As the most practical science, chemistry should have a textbook that highlights its countless applications. Moreover, today's students may enter emerging chemistry-related hybrid fields, like biomaterials science or planetary geochemistry, and the text they use should point out the relevance of chemical concepts to such related sciences. The *Chemical Connections* and *Tools of the Laboratory* boxed essays (which include problems for added relevance), the more pedagogic margin notes, and the many applications woven into the chapter content are up-to-date, student-friendly features that are directly related to the neighboring content.

### CHEMICAL CONNECTIONS TO ENVIRONMENTAL ENGINEERING

#### Solutions and Colloids in Water Purification

**Water Treatment Plants**

Most water destined for human use comes from lakes, rivers, reservoirs, or groundwater. Present in this essential resource may be soluble toxic organic compounds and high concentrations of  $\text{NO}_3^-$  and  $\text{Fe}^{3+}$ , colloidal clay and microbes, and suspended debris. Let's see how water is treated to remove these dissolved, dispersed, and suspended particles.

**Step 1. Screening and settling.** As water enters the facility, screens remove debris, and settling removes sand and other particles.

**Step 2. Coagulating.** This step and the next two remove colloids. These particles have negative surfaces that repel each other. Added aluminum sulfate (cake alum:  $\text{Al}_2(\text{SO}_4)_3$ ) or iron(III) chloride ( $\text{FeCl}_3$ ), which supply  $\text{Al}^{3+}$  or  $\text{Fe}^{3+}$  that neutralize the charges, coagulates the particles through intermolecular forces.

**Step 3. Flocculating and sedimenting.** Mixing water and flocculating agents in large basins causes a fluffy floc to form. Added cationic polymers form long-chain bridges between floc particles, which grow bigger and flow into other basins, where they form a sediment and are removed. Some plants use *dissolved air flotation* (DAF) instead: bubbles forced through the water attach to the floc, and the floating mass is skimmed.

**Step 4. Filtering.** Various filters remove remaining particles. In *slow sand filters*, the water passes through sand and/or gravel of increasing particle size. In *rapid sand filters*, the sand is backwashed with water, and the colloidal mass is removed. Membrane filters (not shown) with pore sizes of 0.1–10  $\mu\text{m}$  are thin tubes bundled into these tubes, with large central tubes and organisms rest.

**Step 5. Disinfecting.** Water sources often contain harmful microorganisms that are killed by one of three agents:

- Chlorine, as aqueous bleach ( $\text{ClO}^-$ ) or  $\text{Cl}_2$ , is most common, but carcinogenic chlorinated organic compounds can form.
- UV light emitted by high-intensity fluorescent tubes disrupts by disrupting microorganisms' DNA.
- Ozone ( $\text{O}_3$ ) gas is a powerful oxidizing agent.

Sodium fluoride (NaF) to prevent tooth decay and phosphate salts to prevent leaching of lead from pipes may then be added.

**Step 6 (not shown). Adsorbing onto granular activated carbon (GAC).** Petroleum and other organic contaminants are removed by adsorption. GAC is a highly porous agent formed by "activating" wood, coal, or coconut shells with steam: 1 kg of GAC has a surface area of 275 acres!

**Water Softening via Ion Exchange**

Water with large amounts of  $2+$  ions, such as  $\text{Ca}^{2+}$  and  $\text{Mg}^{2+}$ , is called **hard water**. Combined with fatty-acid anions in soap, these cations form solid deposits on clothes, washing machines, and sinks:

$$\text{Ca}^{2+}(\text{aq}) + 2\text{C}_{17}\text{H}_{35}\text{COONa}(\text{aq}) \longrightarrow (\text{C}_{17}\text{H}_{35}\text{COO})_2\text{Ca}(\text{s}) + 2\text{Na}^+(\text{aq})$$

insoluble deposit

When a large amount of  $\text{HCO}_3^-$  is present, the cations form *scale*, a carbonate deposit in boilers and hot-water pipes that interferes with the transfer of heat:

$$\text{Ca}^{2+}(\text{aq}) + 2\text{HCO}_3^-(\text{aq}) \longrightarrow \text{CaCO}_3(\text{s}) + \text{CO}_2(\text{g}) + \text{H}_2\text{O}(\text{l})$$

### Wastewater Treatment

Wastewater, used domestic or industrial water, is treated in several ways before being returned to a natural source:

- In **primary treatment**, the water enters a settling basin to remove particles.
- In **biological treatment**, bacteria metabolize organic compounds and are then removed by settling.
- In **advanced treatment**, a process is tailored to remove a specific pollutant. For example, ammonia, which causes excessive growth of plants and algae, is removed in two steps:
  1. **Nitrification.** Certain bacteria oxidize ammonia (electron donor) with  $\text{O}_2$  (electron acceptor) to form nitrate ion:
 
$$\text{NH}_4^+ + 2\text{O}_2 \longrightarrow \text{NO}_3^- + 2\text{H}^+ + \text{H}_2\text{O}$$
  2. **Denitrification.** Other bacteria oxidize an added compound like methanol ( $\text{CH}_3\text{OH}$ ) using the  $\text{NO}_3^-$ :
 
$$5\text{CH}_3\text{OH} + 6\text{NO}_3^- \longrightarrow 3\text{N}_2 + 5\text{CO}_2 + 7\text{H}_2\text{O} + 6\text{OH}^-$$

Thus, the process converts  $\text{NH}_3$  in wastewater to  $\text{N}_2$ , which is released to the atmosphere.

**Membrane Processes and Reverse Osmosis**

Membranes with 0.0001–0.01  $\mu\text{m}$  pores can remove unwanted ions from water. Recall that solutions of different concentrations separated by a semipermeable membrane create osmotic pressure. In **reverse osmosis**, a pressure greater than the osmotic pressure is applied to the more concentrated solution to force water back through the membrane and filter out ions. In homes, toxic *heavy-metal* ions, such as  $\text{Pb}^{2+}$ ,  $\text{Cd}^{2+}$ , and  $\text{Hg}^{2+}$ , are removed this way. On a large scale, reverse osmosis is used for **desalination**, which can convert seawater (40,000 ppm of ions) to drinking water (400 ppm) (Figure B13.3).

### TOOLS OF THE LABORATORY

#### Nuclear Magnetic Resonance (NMR) Spectroscopy

In addition to mass spectrometry (Chapter 2) and infrared (IR) spectroscopy (Chapter 9), one of the most useful tools for analyzing organic and biochemical structures is **nuclear magnetic resonance (NMR) spectroscopy**, which measures the molecular environments of certain nuclei in a molecule.

Like electrons, several types of nuclei, such as  $^{13}\text{C}$ ,  $^{19}\text{F}$ ,  $^{31}\text{P}$ , and  $^1\text{H}$ , act as if they spin in either of two directions, each of which creates a tiny magnetic field. In this discussion, we focus primarily on  $^1\text{H}$ -NMR spectroscopy, which measures changes in the nuclei of the most common isotope of hydrogen. Oriented randomly, the magnetic fields of all the  $^1\text{H}$  nuclei in a sample of compound, when placed in a strong external magnetic field ( $B_0$ ), become aligned either with the external field (parallel) or against it (antiparallel). Most nuclei adopt the parallel orientation, which is slightly lower in energy. The energy difference ( $\Delta E$ ) between the two energy states (spin states) lies in the radio-frequency (rf) region of the electromagnetic spectrum (Figure B15.1).

When an  $^1\text{H}$  (blue arrow) in the lower energy (parallel) spin state absorbs a photon in the radio-frequency region with an energy equal to  $\Delta E$ , it "flips," in a process called **resonance**, to the higher energy (antiparallel) spin state. The system then re-emits that energy, which is detected by the rf receiver of the  $^1\text{H}$ -NMR spectrometer. The  $\Delta E$  between the two states depends on the **actual magnetic field** felt by each  $^1\text{H}$  nucleus, which is affected by the tiny magnetic fields of the electrons of atoms adjacent to that nucleus. Thus, the  $\Delta E$  required for resonance of each  $^1\text{H}$  nucleus depends on its specific molecular environment—the C atoms, electronegative atoms, multiple bonds, and aromatic rings around it.  $^1\text{H}$  nuclei in different molecular environments produce different peaks in the  $^1\text{H}$ -NMR spectrum.

An  $^1\text{H}$ -NMR spectrum, which is unique for each compound, is a series of peaks that represents the resonance as a function of the changing magnetic field. The **chemical shift** of the  $^1\text{H}$  nuclei in a given environment is where a peak appears. Chemical shifts are shown relative to that of an added standard, tetramethylsilane ( $\text{CH}_4$ )<sub>4</sub>, or TMS. TMS has 12  $^1\text{H}$  nuclei bonded to four C atoms that are bonded to one Si atom in a tetrahedral arrangement, so all 12 are in identical environments and produce only one peak.

Figure B15.2 shows the  $^1\text{H}$ -NMR spectrum of acetone. The six  $^1\text{H}$  nuclei of acetone have identical environments: all six are bonded to two C atoms that are each bonded to the C atom in a C=O bond. So one peak is produced, but at a different position from the TMS peak. The spectrum of dimethoxymethane in Figure B15.3 shows two peaks in addition to the TMS peak since the  $^1\text{H}$  nuclei have two different environments. The taller peak is due to the six  $^1\text{H}$  nuclei in the two  $\text{CH}_3$  groups, and the shorter peak is due to the two  $^1\text{H}$  nuclei in the  $\text{CH}_2$  group. The area under each peak (given here in units of chart-paper spaces) is proportional to the number of  $^1\text{H}$  nuclei in a given environment. Note that the area ratio is  $20.3/6.8 \approx 3/1$ , the same as the ratio of six nuclei in the  $\text{CH}_3$  groups to two in the  $\text{CH}_2$  group. Thus, by analyzing the chemical shifts and peak areas, the chemist learns the type and number of hydrogen atoms in the compound.

(continued)

Figure B13.2 Ion exchange to remove hard-water cations.

ions for charge balance (Figure B13.2). The hard-water cations displace the  $\text{Na}^+$  ions and bind to the anionic groups. When all resin sites are occupied, the resin is regenerated with concentrated  $\text{Na}^+$  solution that exchanges  $\text{Na}^+$  ions for bound  $\text{Ca}^{2+}$  and  $\text{Mg}^{2+}$ .

**Membrane Processes and Reverse Osmosis**

Membranes with 0.0001–0.01  $\mu\text{m}$  pores can remove unwanted ions from water. Recall that solutions of different concentrations separated by a semipermeable membrane create osmotic pressure. In **reverse osmosis**, a pressure greater than the osmotic pressure is applied to the more concentrated solution to force water back through the membrane and filter out ions. In homes, toxic *heavy-metal* ions, such as  $\text{Pb}^{2+}$ ,  $\text{Cd}^{2+}$ , and  $\text{Hg}^{2+}$ , are removed this way. On a large scale, reverse osmosis is used for **desalination**, which can convert seawater (40,000 ppm of ions) to drinking water (400 ppm) (Figure B13.3).

Figure B13.3 Reverse osmosis to remove ions. A, Part of a reverse-osmosis permeator. B, Each permeator contains a bundle of hollow fibers of semipermeable membrane. C, Pumping seawater at high pressure removes ions and pure water enters the fibers and is collected.

554

### A Purple Mule, Not a Blue Horse and a Red Donkey

A mule is a genetic mix, a hybrid, of a horse and a donkey; it is not a horse one instant and a donkey the next. Similarly, the color purple is a mix of red and blue, not red one instant and blue the next. In the same sense, a resonance hybrid is one molecular species, not one resonance form this instant and another resonance form the next. The problem is that we cannot depict the actual species, the hybrid, accurately with a single Lewis structure.

Blue horse      Red donkey      Purple mule

555

633

## Reinforcing through Review and Practice

A favorite feature, the section summaries that conclude every section restate the major ideas concisely and immediately (rather than postponing such review until the end of the chapter).

A rich catalog of study aids ends each chapter to help students review the content:

- **Learning Objectives**, with section and/or sample problem numbers, focus on the concepts to understand and the skills to master.
- **Key Terms**, boldfaced and defined within the chapter, are listed here by section (with page numbers), as well as being defined in the *Glossary*.
- **Key Equations and Relationships** are highlighted and numbered within the chapter and listed here with page numbers.
- **Brief Solutions to Follow-up Problems** triple the number of worked problems by providing multistep calculations at the end of the chapter, rather than just numerical answers at the back of the book.

### Summary of Section 9.1

- Nearly all naturally occurring substances consist of atoms or ions bonded to others. Chemical bonding allows atoms to lower their energy.
- Ionic bonding occurs when metal atoms transfer electrons to nonmetal atoms, and the resulting ions attract each other and form an ionic solid.
- Covalent bonding is most common between nonmetal atoms and usually results in individual molecules. Bonded atoms share one or more pairs of electrons that are localized between them.
- Metallic bonding occurs when many metal atoms pool their valence electrons into a delocalized electron “sea” that holds all the atoms in the sample together.
- The Lewis electron-dot symbol of a main-group atom shows valence electrons as dots surrounding the element symbol.
- The octet rule says that, when bonding, many atoms lose, gain, or share electrons to attain a filled outer level of eight (or two) electrons.

556 Chapter 13 • The Properties of Mixtures: Solutions and Colloids

### CHAPTER REVIEW GUIDE

**Learning Objectives** Relevant section (§) and/or sample problem (SP) numbers appear in parentheses.

**Understand These Concepts**

- The quantitative meaning of solubility (§13.1)
- The major types of intermolecular forces in solution and their relative strengths (§13.1)
- How the like-dissolves-like rule depends on intermolecular forces (§13.1)
- Why gases have relatively low solubilities in water (§13.1)
- General characteristics of solutions formed by various combinations of gases, liquids, and solids (§13.1)
- How intermolecular forces stabilize the structures of proteins, the cell membrane, and DNA (§13.2)
- The enthalpy components of a solution cycle and their effect on  $\Delta H_{\text{soln}}$  (§13.3)
- The dependence of  $\Delta H_{\text{soln}}$  on ionic charge density and the factors that determine whether ionic solution processes are exothermic or endothermic (§13.3)
- The meaning of entropy and how the balance between the change in enthalpy and the change in entropy governs the solution process (§13.3)
- The distinctions among saturated, unsaturated, and supersaturated solutions, and the equilibrium nature of a saturated solution (§13.4)
- The relation between temperature and the solubility of solids (§13.4)
- Why the solubility of gases in water decreases with a rise in temperature (§13.4)
- The effect of gas pressure on solubility and its quantitative expression as Henry’s law (§13.4)
- The meaning of molality, molarity, mole fraction, and parts by mass or by volume of a solution, and how to convert among them (§13.5)
- The distinction between electrolytes and nonelectrolytes in solution (§13.6)
- The four colligative properties and their dependence on number of dissolved particles (§13.6)

**Master These Skills**

- Predicting relative solubilities from intermolecular forces (SP 13.1)
- Calculating the heat of solution for an ionic compound (SP 13.2)
- Using Henry’s law to calculate the solubility of a gas (SP 13.3)
- Expressing concentration in terms of molality, parts by mass, parts by volume, and mole fraction (SPs 13.4, 13.5)
- Interconverting among the various terms for expressing concentration (SP 13.6)
- Using Raoult’s law to calculate the vapor pressure lowering of a solution (SP 13.7)
- Determining boiling point elevation and freezing point depression of a solution (SP 13.8)
- Using a colligative property to calculate the molar mass of a solute (SP 13.9)
- Calculating the composition of vapor over a solution of volatile nonelectrolyte (§13.6)
- Calculating the van’t Hoff factor (*i*) from the magnitude of a colligative property (§13.6)
- Using a solution depiction to determine colligative properties (SP 13.10)

**Key Terms** Page numbers appear in parentheses.

<b>Section 13.1</b> solute (518) solvent (518) miscible (518) solubility (5) (518) like-dissolves-like rule (518) hydration shell (518) ion-induced dipole force (518) dipole-induced dipole force (518) alloy (522)	mononucleotide (527) double helix (527)	<b>Section 13.3</b> heat of solution ( $\Delta H_{\text{soln}}$ ) (528) solution (529) hydration (529) heat of hydration ( $\Delta H_{\text{hyd}}$ ) (529) charge density (529) entropy (S) (531)	<b>Section 13.5</b> semipermeable membrane (546) osmosis (546) osmotic pressure ( $\Pi$ ) (546) fractional distillation (548) ionic atmosphere (549)	<b>Section 13.7</b> suspension (552) colloid (552) Tyndall effect (553) hard water (554) water softening (554) ion exchange (554) reverse osmosis (555) desalination (555) wastewater (555)
<b>Section 13.2</b> protein (523) amino acid (523) soap (525) lipid bilayer (526) nucleic acid (527)	<b>Section 13.4</b> saturated solution (533) unsaturated solution (533) supersaturated solution (533) Henry’s law (535)	<b>Section 13.6</b> colligative property (541) electrolyte (541) nonelectrolyte (541) vapor pressure lowering ( $\Delta P$ ) (542) Raoult’s law (542) ideal solution (542) boiling point elevation ( $\Delta T_b$ ) (543) freezing point depression ( $\Delta T_f$ ) (545)		

Chapter 13 • Chapter Review Guide 557

**Key Equations and Relationships** Page numbers appear in parentheses.

**13.1** Dividing the general heat of solution into component enthalpies (538):  

$$\Delta H_{\text{soln}} = \Delta H_{\text{solite}} + \Delta H_{\text{solvent}} + \Delta H_{\text{mix}}$$

**13.2** Dividing the heat of solution of an ionic compound in water into component enthalpies (539):  

$$\Delta H_{\text{soln}} = \Delta H_{\text{ionice}} + \Delta H_{\text{hydr}} \text{ of the ions}$$

**13.3** Relating gas solubility to its partial pressure (Henry’s law) (535):  

$$S_{\text{gas}} = k_H \times P_{\text{gas}}$$

**13.4** Defining concentration in terms of molarity (537):  

$$\text{Molarity (M)} = \frac{\text{amount (mol) of solute}}{\text{volume (L) of solution}}$$

**13.5** Defining concentration in terms of molality (537):  

$$\text{Molality (m)} = \frac{\text{amount (mol) of solute}}{\text{mass (kg) of solvent}}$$

**13.6** Defining concentration in terms of mass percent (538):  

$$\text{Mass percent [\% (w/w)]} = \frac{\text{mass of solute}}{\text{mass of solution}} \times 100$$

**13.7** Defining concentration in terms of volume percent (538):  

$$\text{Volume percent [\% (v/v)]} = \frac{\text{volume of solute}}{\text{volume of solution}} \times 100$$

**13.8** Defining concentration in terms of mole fraction (538):  

$$\text{Mole fraction (X)} = \frac{\text{amount (mol) of solute}}{\text{amount (mol) of solute} + \text{amount (mol) of solvent}}$$

**13.9** Expressing the relationship between the vapor pressure of solvent above a solution and its mole fraction in the solution (Raoult’s law) (542):  

$$P_{\text{solvent}} = X_{\text{solvent}} \times P_{\text{solvent}}^*$$

**13.10** Calculating the vapor pressure lowering due to solute (542):  

$$\Delta P = X_{\text{solute}} \times P_{\text{solvent}}^*$$

**13.11** Calculating the boiling point elevation of a solution (544):  

$$\Delta T_b = K_b m$$

**13.12** Calculating the freezing point depression of a solution (545):  

$$\Delta T_f = K_f m$$

**13.13** Calculating the osmotic pressure of a solution (546):  

$$\Pi = \frac{n_{\text{solute}}}{V_{\text{soln}}} RT = MRT$$

**BRIEF SOLUTIONS TO FOLLOW-UP PROBLEMS**

**13.1A** (a) 1,4-Butanediol is more soluble in water because it can form more H bonds.  
 (b) Chloroform is more soluble in water because of dipole-dipole forces.

**13.1B** (a) Chloroform dissolves more chloromethane due to similar dipole-dipole forces.  
 (b) Hexane dissolves more pentanol due to dispersion forces.

**13.2A** From Equation 13.2, we have  

$$\Delta H_{\text{soln}} \text{ of } \text{KNO}_3 = \Delta H_{\text{solite}} \text{ of } \text{KNO}_3 + (\Delta H_{\text{hydr}} \text{ of } \text{K}^+ + \Delta H_{\text{hydr}} \text{ of } \text{NO}_3^-)$$

$$34.89 \text{ kJ/mol} = 685 \text{ kJ/mol} + (\Delta H_{\text{hydr}} \text{ of } \text{K}^+ + \Delta H_{\text{hydr}} \text{ of } \text{NO}_3^-)$$

$$\Delta H_{\text{hydr}} \text{ of } \text{K}^+ + \Delta H_{\text{hydr}} \text{ of } \text{NO}_3^- = 34.89 \text{ kJ/mol} - 685 \text{ kJ/mol} = -650.11 \text{ kJ/mol}$$

**13.2B** From Equation 13.2, we have  

$$\Delta H_{\text{soln}} \text{ of NaCN} = \Delta H_{\text{solite}} \text{ of NaCN} + (\Delta H_{\text{hydr}} \text{ of } \text{Na}^+ + \Delta H_{\text{hydr}} \text{ of } \text{CN}^-)$$

$$1.21 \text{ kJ/mol} = 766 \text{ kJ/mol} - 410 \text{ kJ/mol} + \Delta H_{\text{hydr}} \text{ of } \text{CN}^-$$

$$\Delta H_{\text{hydr}} \text{ of } \text{CN}^- = 1.21 \text{ kJ/mol} - 766 \text{ kJ/mol} + 410 \text{ kJ/mol} = -354.79 \text{ kJ/mol}$$

**13.3A**  $S_{\text{O}_2} = (7 \times 10^{-4} \text{ mol/L} \cdot \text{atm})(0.78 \text{ atm}) = 5.46 \times 10^{-4} \text{ mol/L}$

**13.3B** In a mixture of gases, the volume percent of each gas times the total pressure equals its partial pressure (Dalton’s law, Section 5.4). From Equation 13.3, we have  

$$S_{\text{gas}} = k_H \times P_{\text{gas}} = (2.5 \times 10^{-2} \text{ mol/L} \cdot \text{atm})(0.40 \times 1.2 \text{ atm}) = 1.2 \times 10^{-2} \text{ mol/L}$$

**13.4A** Mass (g) of glucose =  $563 \text{ g ethanol} \times \frac{1 \text{ kg}}{10^3 \text{ g}} \times 2.40 \times 10^{-2} \text{ mol glucose} / 1 \text{ kg ethanol} \times 180.16 \text{ g glucose} / 1 \text{ mol glucose} = 2.43 \text{ g glucose}$

**13.4B** Convert mass (g) of  $\text{I}_2$  to amount (mol) and amount (mol) of  $(\text{CH}_3\text{CH}_2)_2\text{O}$  to mass (kg). Then divide moles of  $\text{I}_2$  by kg of  $(\text{CH}_3\text{CH}_2)_2\text{O}$ .  

$$\text{Amount (mol) of } \text{I}_2 = 15.20 \text{ g } \text{I}_2 \times \frac{1 \text{ mol } \text{I}_2}{253.8 \text{ g } \text{I}_2} = 5.989 \times 10^{-2} \text{ mol } \text{I}_2$$

$$\text{Mass (kg) of } (\text{CH}_3\text{CH}_2)_2\text{O} = 1.33 \text{ mol } (\text{CH}_3\text{CH}_2)_2\text{O} \times \frac{74.12 \text{ g } (\text{CH}_3\text{CH}_2)_2\text{O}}{1 \text{ mol } (\text{CH}_3\text{CH}_2)_2\text{O}} \times \frac{1 \text{ kg}}{10^3 \text{ g}} = 9.86 \times 10^{-2} \text{ kg } (\text{CH}_3\text{CH}_2)_2\text{O}$$

$$\text{Molality (m)} = \frac{5.989 \times 10^{-2} \text{ mol}}{9.86 \times 10^{-2} \text{ kg}} = 0.607 \text{ m}$$



Finally, an exceptionally large number of qualitative, quantitative, and molecular-scene problems end each chapter. Four types of problems are presented—three by chapter section, with comprehensive problems following:

- **Concept Review Questions** test qualitative understanding of key ideas.
- **Skill-Building Exercises** are grouped in similar pairs, with one of each pair answered in the back of the book. A group of similar exercises may begin with explicit steps and increase in difficulty, gradually weaning the student from the need for multistep directions.
- **Problems in Context** apply the skills learned in the skill-building exercises to interesting scenarios, including realistic examples dealing with industry, medicine, and the environment.
- **Comprehensive Problems**, mostly based on realistic applications, are more challenging and rely on material from any section of the current chapter or any previous chapter.

## PROBLEMS

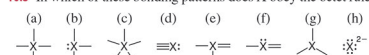
Problems with **colored** numbers are answered in Appendix E and worked in detail in the Student Solutions Manual. Problem sections match those in the text and give the numbers of relevant sample problems. Most offer Concept Review Questions, Skill-Building Exercises (grouped in pairs covering the same concept), and Problems in Context. The Comprehensive Problems are based on material from any section or previous chapter.

### Depicting Molecules and Ions with Lewis Structures

(Sample Problems 10.1 to 10.5)

#### Concept Review Questions

- 10.1** Which of these atoms *cannot* serve as a central atom in a Lewis structure: (a) O; (b) He; (c) F; (d) H; (e) P? Explain.
- 10.2** When is a resonance hybrid needed to adequately depict the bonding in a molecule? Using  $\text{NO}_2$  as an example, explain how a resonance hybrid is consistent with the actual bond length, bond strength, and bond order.
- 10.3** In which of these bonding patterns does X obey the octet rule?

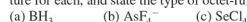


**10.4** What is required for an atom to expand its valence shell? Which of the following atoms can expand its valence shell: F, S, H, Al, Se, Cl?

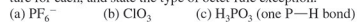
#### Skill-Building Exercises (grouped in similar pairs)

- 10.5** Draw a Lewis structure for (a)  $\text{SiF}_2$ ; (b)  $\text{SeCl}_2$ ; (c)  $\text{COF}_2$  (C is central).
- 10.6** Draw a Lewis structure for (a)  $\text{PH}_4^+$ ; (b)  $\text{C}_2\text{F}_4$ ; (c)  $\text{SbH}_3$ .
- 10.7** Draw a Lewis structure for (a)  $\text{PF}_3$ ; (b)  $\text{H}_2\text{CO}_3$  (both H atoms are attached to O atoms); (c)  $\text{CS}_2$ .
- 10.8** Draw a Lewis structure for (a)  $\text{CH}_2\text{S}$ ; (b)  $\text{S}_2\text{Cl}_2$ ; (c)  $\text{CHCl}_3$ .
- 10.9** Draw Lewis structures of all the important resonance forms of (a)  $\text{NO}_2^+$ ; (b)  $\text{NO}_2\text{F}$  (N is central).
- 10.10** Draw Lewis structures of all the important resonance forms of (a)  $\text{HNO}_3$  ( $\text{HONO}_2$ ); (b)  $\text{HAsO}_4^{2-}$  ( $\text{HOAsO}_3^{2-}$ ).
- 10.11** Draw Lewis structures of all the important resonance forms of (a)  $\text{N}_3^-$ ; (b)  $\text{NO}_2^-$ .
- 10.12** Draw Lewis structures of all the important resonance forms of (a)  $\text{HCO}_2^-$  (H is attached to C); (b)  $\text{HBrO}_3$  ( $\text{HOBrO}_2$ ).
- 10.13** Draw the Lewis structure with lowest formal charges, and determine the charge of each atom in (a)  $\text{IF}_3$ ; (b)  $\text{AlH}_4^-$ .
- 10.14** Draw the Lewis structure with lowest formal charges, and determine the charge of each atom in (a)  $\text{OCS}$ ; (b)  $\text{NO}$ .
- 10.15** Draw the Lewis structure with lowest formal charges, and determine the charge of each atom in (a)  $\text{CN}^-$ ; (b)  $\text{ClO}^-$ .
- 10.16** Draw the Lewis structure with lowest formal charges, and determine the charge of each atom in (a)  $\text{BF}_4^-$ ; (b)  $\text{ClNO}$ .

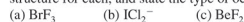
**10.19** These species do not obey the octet rule. Draw a Lewis structure for each, and state the type of octet-rule exception:



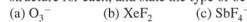
**10.20** These species do not obey the octet rule. Draw a Lewis structure for each, and state the type of octet-rule exception:



**10.21** These species do not obey the octet rule. Draw a Lewis structure for each, and state the type of octet-rule exception:



**10.22** These species do not obey the octet rule. Draw a Lewis structure for each, and state the type of octet-rule exception:



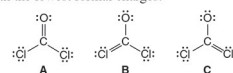
#### Problems in Context

**10.23** Molten beryllium chloride reacts with chloride ion from molten  $\text{NaCl}$  to form the  $\text{BeCl}_4^{2-}$  ion, in which the Be atom attains an octet. Show the net ionic reaction with Lewis structures.

**10.24** Despite many attempts, the perbromate ion ( $\text{BrO}_4^-$ ) was not prepared in the laboratory until about 1970. (In fact, articles were published explaining theoretically why it could never be prepared!) Draw a Lewis structure for  $\text{BrO}_4^-$  in which all atoms have lowest formal charges.

**10.25** Cryolite ( $\text{Na}_3\text{AlF}_6$ ) is an indispensable component in the electrochemical production of aluminum. Draw a Lewis structure for the  $\text{AlF}_6^{3-}$  ion.

**10.26** Phosgene is a colorless, highly toxic gas that was employed against troops in World War I and is used today as a key reactant in organic syntheses. From the following resonance structures, select the one with the lowest formal charges:

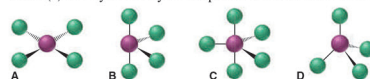


#### Valence-Shell Electron-Pair Repulsion (VSEPR) Theory

(Sample Problems 10.6 to 10.8)

#### Concept Review Questions

- 10.27** If you know the formula of a molecule or ion, what is the first step in predicting its shape?
- 10.28** In what situation is the name of the molecular shape the same as the name of the electron-group arrangement?
- 10.29** Which of the following numbers of electron groups can give rise to a bent (V-shaped) molecule: two, three, four, five, six? Draw an example for each case, showing the shape classification ( $\text{AX}_m\text{E}_n$ ) and the ideal bond angle.
- 10.30** Name all the molecular shapes that have a tetrahedral electron-group arrangement.
- 10.31** Consider the following molecular shapes. (a) Which has the most electron pairs (both shared and unshared) around the central atom? (b) Which has the most unshared pairs around the central atom? (c) Do any have only shared pairs around the central atom?



of each ion with  
es, and give oxi-  
rm of each ion  
he charges, and  
(b)  $\text{ClO}_2^-$ .

### Comprehensive Problems

**2.119** Helium is the lightest noble gas and the second most abundant element (after hydrogen) in the universe.

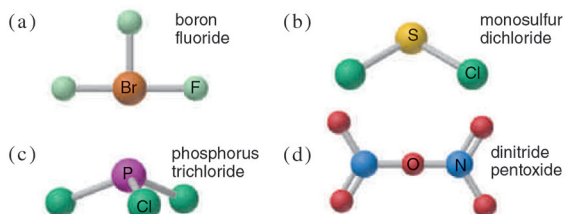
(a) The radius of a helium atom is  $3.1 \times 10^{-11}$  m; the radius of its nucleus is  $2.5 \times 10^{-15}$  m. What fraction of the spherical atomic volume is occupied by the nucleus ( $V$  of a sphere =  $\frac{4}{3}\pi r^3$ )?

(b) The mass of a helium-4 atom is  $6.64648 \times 10^{-24}$  g, and each of its two electrons has a mass of  $9.10939 \times 10^{-28}$  g. What fraction of this atom's mass is contributed by its nucleus?

**2.120** From the following ions (with their radii in pm), choose the pair that forms the strongest ionic bond and the pair that forms the weakest:

Ion:	$\text{Mg}^{2+}$	$\text{K}^+$	$\text{Rb}^+$	$\text{Ba}^{2+}$	$\text{Cl}^-$	$\text{O}^{2-}$	$\text{I}^-$
Radius:	72	138	152	135	181	140	220

**2.121** Give the molecular mass of each compound depicted below, and provide a correct name for any that are named incorrectly.



## OPTIMIZING THE TEXT

The modern chemistry student's learning experience is changing dramatically. To address the changes that students face, a modern text partnered with a suite of robust electronic tools must continue to evolve. With each edition, students and instructors alike have been involved in refining this text. From one-on-one interviews, focus groups, and symposia, as well as extensive chapter reviews and class tests of the previous edition, we learned that everyone praises the pioneering molecular art, the stepwise problem-solving approach, the abundant mix of qualitative, quantitative, and applied end-of-chapter problems, and the rigorous *and* student-friendly coverage of mainstream topics.

### Global Changes to Every Chapter

Our revision for this edition has gone further than ever to optimize the text. We completed the distillation of the text and the annotation of the illustrations so appreciated in the previous edition, and created an inviting, easy-to-follow page design. But, when professors *and* students were asked what they wanted most in a new edition, the answer inevitably cited more good problems, so that became a major focus. We are delighted to introduce the seventh edition of *Chemistry: The Molecular Nature of Matter and Change*, which takes several major steps closer to perfecting its content and approach.

**Learning ideas through focused writing and content presentation.** Once again, every discussion has been revised to optimize clarity, readability, and the conciseness and directness of the presentation. The use of additional subheads, numbered (and titled) paragraphs, and bulleted (and titled) lists that was introduced in the sixth edition has been completed in the seventh.

**Applying ideas with an enhanced problem-solving approach.** The much admired—and copied—four-part problem-solving format (plan, solution, check, follow-up) is still used throughout, in both data-based *and* molecular-scene *Sample Problems*. But, many changes make the seventh edition a problem-solving “powerhouse.” Each sample problem now includes *two Follow-up Problems*: as professors requested, the first closely matches the worked-out problem, so the student gains confidence, while the second varies a bit to test comprehension. As always, a thoroughly worked-out *Brief Solution* for each follow-up problem appears at the end of the chapter (rather than providing just a numerical answer in a distant end-of-book appendix, as is typical). As students requested, for more practice, each sample problem now lists *Some Similar Problems* within the end-of-chapter problem set. Moreover, some remaining

pedagogic gaps have been filled with 22 *new* sample problems: thus, the seventh edition has over 250 sample problems and over 500 follow-up problems.

**Re-learning ideas with annotated illustrations.** The innovative three-level figures and other art that raised the bar for molecular visualization in chemistry textbooks is still present. Many existing figures have been revised and several new ones added to create an even better teaching tool. And, continuing the innovation of last edition, wherever appropriate, figure legends have been turned into simple captions and their content into clarifying annotations within the figures themselves.

**Easier studying with attractive and functional page design and layout.** A more open layout with three-dimensional accents engages students by focusing attention on section heads, sample problems, tables, and other important features, while maintaining clean margins. The orderly presentation fosters easier viewing and studying and allows the intimate placement of figures and tables in relation to the discussion.

**Mastering the content with abundant end-of-chapter problem sets.** With the more open design for improved readability, traditional and molecular-scene problems updated and revised, and many new problems added, these problem sets are more extensive than in most other texts. They provide students and teachers with abundant choices in a wide range of difficulty and real-life scenarios.

### Content Changes to Individual Chapters

The major revision that created this seventh edition of *Chemistry* included key improvements to nearly every chapter:

- **Chapter 1** has been rearranged so that units are now discussed *before* the problem-solving approach and unit conversions are introduced. The chapter includes a *new sample problem* on converting units raised to a power.
- **Chapter 2** now clarifies the meaning of mass fraction and introduces IUPAC's new recommendations on atomic-mass ranges. It presents the periodic table updated with the latest atomic masses and includes a *new sample problem* on identifying an element from its *Z* value.
- **Chapter 3** includes more information in its road maps, and road maps in later chapters also have these enhancements. Sample problems on calculating mass percent of an element and the mass of an element have been carefully revised. The section on calculations for reactions in solution has been moved to Chapter 4.
- **Chapter 4** introduces many types of reactions in solution, so it was thoroughly revised to, more logically, include the

material on molarity and solution stoichiometry previously in Chapter 3. The table on solubility rules was revised, and a new table focuses on the reactions of strong versus weak acids in strong base. *Two new sample problems* show how to determine the amount of ions in solution and the amounts of reactants and products in a precipitation reaction. And two sample problems from the last edition were combined to show the logical relation of their material. Balancing redox reactions by the oxidation-number method was deleted, but balancing them by the half-reaction method is covered in the electrochemistry chapter (Chapter 21).

- **Chapter 5** includes *two new sample problems* that apply various combinations of the individual gas laws, and it provides a derivation of Graham's law.
- **Chapter 6** includes *a new sample problem* on calculating  $PV$  work.
- **Chapter 7** incorporates carefully revised atomic and continuous spectra throughout.
- **Chapter 8** contains a new table on changes in  $Z_{\text{eff}}$  within the sublevels of an atom.
- **Chapter 9** includes *a new sample problem* on predicting relative lattice energy and a revised figure on the properties of the covalent bond.
- **Chapter 10** has pedagogic improvements to several figures and expanded coverage in two sample problems.
- **Chapter 12** includes *two new sample problems*, one on the use of phase diagrams to predict phase changes and the other on determining the number of particles in a unit cell and the coordination number.
- **Chapter 13** presents *a new sample problem* on calculating an aqueous ionic heat of solution.
- **Chapter 14** incorporates in its unique Family Portraits all the updated atomic masses and newly synthesized elements that fill out the periodic table.
- **Chapter 16** includes revisions to two important figures and *two new sample problems*, one on rate laws and the molecularity of elementary steps and the other on intermediates and the correlation of rate laws and mechanism.
- **Chapter 17** contains *a new sample problem* on writing a reaction quotient from the balanced equation, and two important figures have been revised.
- **Chapter 18** has undergone major improvements. It has *two new sample problems*, one on calculating hydronium and hydroxide ion concentrations in strong acids and bases and the other on finding the percent dissociation of a weak acid. Several summarizing tables were revised to display reactions, and a key figure was revised. Also, the discussion was rearranged to present the material on acids in a clearer sequence.
- **Chapter 19** incorporates improvements to several key figures, as well as a new figure that summarizes the effects of added acid or base to buffer-component concentration and pH.

- **Chapter 20** has clearer discussions of measuring the change in entropy and of free energy and work, as well as *a new sample problem* on the relationship between  $\Delta G^\circ$  and  $K$ .
- **Chapter 21** includes *a new sample problem* on using  $E^\circ_{\text{half-cell}}$  to find  $E^\circ_{\text{cell}}$ , as well as much clearer discussions of the activity series of the metals, the electrolysis of aqueous salts, and the stoichiometry of electrolysis.
- **Chapter 23** has improvements to many figures as well as *a new sample problem* on finding the coordination number and ion charge in a coordination compound.
- **Chapter 24** provides *a new sample problem* on calculating the specific activity and decay constant of a radionuclide.

## Innovative Topic and Chapter Presentation

While the topic sequence coincides with that used in most mainstream courses, built-in flexibility allows a wide range of differing course structures:

**For courses that follow their own topic sequence**, the general presentation, with its many section and subsection breaks and bulleted lists, allows topics to be rearranged, or even deleted, with minimal loss of continuity.

**For courses that present several chapters, or topics within chapters, in different orders:**

- Redox balancing by the oxidation-number method (formerly covered in Chapter 4) has been removed from the text, and the half-reaction method is covered with electrochemistry in Chapter 21, but it can easily be taught with Chapter 4.
- Gases (Chapter 5) can be covered in sequence to explore the mathematical modeling of physical behavior or, with no loss of continuity, just before liquids and solids (Chapter 12) to show the effects of intermolecular forces on the three states of matter.

**For courses that want an atoms-first approach for some of the material**, Chapters 7 through 13 move smoothly from quantum theory (7) through electron configuration (8), bonding models (9), molecular shape (10), VB and MO bonding theories (11), intermolecular forces in liquids and solids (12), and solutions (13). Immediate applications of these concepts appear in the discussions of periodic patterns in main-group chemistry (Chapter 14) and in the survey of organic chemistry (Chapter 15). Some instructors have also brought forward the coverage of transition elements and coordination compounds (23) as further applications of bonding concepts. (Of course, Chapters 14, 15, and 23 can just as easily remain in their more traditional placement later in the course.)

**For courses that want biological/medical applications**, many chapters highlight these topics, including the role of intermolecular forces in biomolecular structure (12), the

chemistry of polysaccharides, proteins, and nucleic acids (including protein synthesis, DNA replication, and DNA sequencing) (15), as well as introductions to enzyme catalysis (16), biochemical pathways (17), and trace elements in protein function (23).

**For courses that want engineering applications of physical chemistry topics,** Chapters 16 through 21 cover kinetics (16), equilibrium in gases (17), acids and bases (18), and aqueous ionic systems (19) and entropy and free energy (20) as they apply to electrochemical systems (21), all in preparation for coverage of the elements in geochemical cycles, metallurgy, and industry in Chapter 22.



**McGraw-Hill Create™** is another way to implement innovative chapter presentation. With Create, you can easily rearrange chapters, combine material from other content sources, and quickly upload content you have written, such as your course syllabus or teaching notes. Find the content you need in Create by searching through thousands of leading McGraw-Hill textbooks. Create even allows you to personalize your book's appearance by selecting the cover and adding your name, school, and course information. Order a Create book, and you'll receive a complimentary print review copy in 3–5 business days or a complimentary electronic review copy (eComp) via e-mail in minutes. Go to [www.mcgrawhillcreate.com](http://www.mcgrawhillcreate.com) today and register to experience how McGraw-Hill Create empowers you to teach *your* students *your* way. [www.mcgrawhillcreate.com](http://www.mcgrawhillcreate.com)

## LEARNING RESOURCES



**McGraw-Hill Connect® Chemistry** provides online presentation, assignment, and assessment solutions. It connects your students with the tools and resources they'll need to achieve success. With Connect Chemistry, you can deliver assignments, quizzes, and tests online. A robust set of questions, problems, and interactive figures are presented and aligned with the textbook's learning goals. The integration of **ChemDraw by PerkinElmer**, the industry standard in chemical drawing software, allows students to create accurate chemical structures in their online homework assignments. As an instructor, you can edit existing questions and write entirely new problems. Track individual student performance—by question, assignment, or in relation to the class overall—with detailed grade reports. Integrate grade reports easily with Learning Management Systems (LMS), such as WebCT and Blackboard—and much more. **ConnectPlus Chemistry** provides students with all the advantages of Connect Chemistry, plus 24/7 online access to an eBook. This media-rich version of the book is available through the McGraw-Hill Connect platform and allows seamless integration of text, media, and assessments. To learn more, visit [www.mcgrawhillconnect.com](http://www.mcgrawhillconnect.com).

The screenshot displays the McGraw-Hill Connect Chemistry interface. At the top, it shows 'Chapter 10 (2 questions, 20 points)' with options for 'preview', 'policies', and 'student activity'. Below this, a question is presented: '1. Draw a Lewis structure for the species  $\text{PCl}_2^+$ . Place the overall charge on the central atom.' The interface includes a 'draw structure...' button and a 'ChemDraw Communication' window. The ChemDraw window shows a drawing area with a toolbar on the left and a 'Don't forget to save!' warning at the bottom. The drawing area contains a simple Lewis structure of a hydrogen atom ( $\text{H}^{\oplus}$ ).

# LEARNSMART PREP™

Fueled by LearnSmart—the most widely used and intelligent adaptive learning resource—**LearnSmart Prep** is designed to get students ready for a forthcoming course by quickly and effectively addressing prerequisite knowledge gaps that may cause problems down the road.

By distinguishing what students know from what they don't, and honing in on concepts they are most likely to forget, LearnSmart Prep maintains a continuously adapting learning path individualized for each student, and tailors content to focus on what the student needs to master in order to have a successful start in the new class.

This revolutionary technology is available only from McGraw-Hill Education and for hundreds of course areas, including general and organic chemistry, as part of the LearnSmart Advantage series.

# McGraw Hill Education | LEARNSMART®

McGraw-Hill LearnSmart™ is available as a stand-alone product or as an integrated feature of McGraw-Hill Connect® Chemistry. It is an adaptive learning system designed to help students learn faster, study more efficiently, and retain more knowledge for greater success. LearnSmart assesses a student's knowledge of course content through a series of adaptive questions. It pinpoints concepts the student does not understand and maps out a personalized study plan for success. This innovative study tool also has features that allow instructors to see exactly what students have accomplished and a built-in assessment tool for graded assignments. Visit the following site for a demonstration: [www.mhlearnsmart.com](http://www.mhlearnsmart.com).

General Chemistry - Chemistry: The Molecular Nature of Matter and Change - Silberberg, 6e.

Assignment progress: 48 items left

Standings for all

Which of the following statements correctly describe a polar covalent bond? Select all that apply.

Click ALL answers that you think are right!

- A bond between atoms whose electronegativities differ by more than 0.4 but less than 1.7 is polar.
- In a polar covalent bond the electron density is highest over the more electronegative atom.
- In a polar covalent bond the atom with the lower electronegativity is marked  $\delta^-$  because it has less electron density.
- A polar bond is a covalent bond in which electrons are shared equally.

Do you know the answer? (Be honest)

Home | Open eBook | Reports | Give feedback | Go to Classic View | Go to SmartBook

General Chemistry - Chemistry: The Molecular Nature of Matter and Change - Silberberg, 6e.

Assignment progress: 48 items left

Standings for all

The ideal gas law is expressed as  $PV = nRT$ . In this equation, the symbol  $R$  indicates the \_\_\_\_\_ gas constant, which may have different values depending on the units used to express \_\_\_\_\_ and pressure.

Please fill in the missing words!  
Submit the answers when you are done!

Home | View Intro | Open eBook | Reports | Give feedback | Go to Classic View | Go to SmartBook



# SMARTBOOK™

Powered by the intelligent and adaptive LearnSmart engine, **SmartBook** is the first and only continuously adaptive reading experience available. Distinguishing what students know from what they don't, and honing in on concepts they are most likely to forget, SmartBook personalizes content for each of them. Reading is no longer a passive and linear experience but an engaging and dynamic one, in which students are more likely to master and retain important concepts, coming to class better prepared.

SmartBook includes powerful reports that identify specific topics and learning objectives students need to study. These valuable reports also provide instructors with insight into how students are progressing through textbook content and are useful for identifying class trends, focusing precious class time, providing personalized feedback to students, and tailoring assessment.

**How does SmartBook work?** Each SmartBook contains four components: Preview, Read, Practice, and Recharge. Starting with a preview of each chapter and key learning objectives, students read the material and are guided to topics on which they need the most practice based on their responses to a continuously adapting diagnostic. Reading and practice continue until SmartBook directs students to review—or recharge—important material they are most likely to forget to ensure concept mastery and retention.

The screenshot displays the SmartBook interface for a chemistry chapter titled "Defining the Mole". The interface includes a navigation menu on the left with sections like "SmartBook", "Read", "Practice", and "Recharge". The main content area features a large yellow box with the number "6" and "32.06" inside, representing the molar mass of calcium. Text explains that the molar mass of a monatomic element is the numerical value listed on the periodic table expressed in units of g/mol. It provides an example: "For example, the molar mass of calcium (Ca) is \_\_\_\_\_ with units of g/mol. (Remember to round to 4 significant digits.)". Below this is a form for the user to type their answer, with buttons for "I know it", "Think so", "Missed", and "No idea". The right side of the page contains a section titled "Defining the Mole" with a diagram of a mole of atoms and text explaining the concept of a mole as a unit for counting particles, similar to a dozen for eggs. It also includes a table of molar masses for various elements and compounds.

## LEARNSMART



# LABS™

## THE Virtual Lab Experience

Based on the same world-class, superbly adaptive technology as LearnSmart, **McGraw-Hill LearnSmart Labs** is a must-see, outcomes-based lab simulation. It assesses a student's knowledge and adaptively corrects deficiencies, allowing the student to learn faster and retain more knowledge with greater success.

First, a student's knowledge is adaptively leveled on core learning outcomes: Questioning reveals knowledge deficiencies that are corrected by the delivery of content that is conditional on a student's response. Then, a simulated lab experience requires the student to think and act like a scientist: Recording, interpreting, and analyzing data using simulated equipment found in labs and clinics. The student is allowed to make mistakes—a powerful part of the learning experience! A virtual coach provides subtle hints when needed, asks questions about the student's choices, and allows the student to reflect on and correct those mistakes. Whether your need is to overcome the logistical challenges of a traditional lab, provide better lab prep, improve student performance, or make your online experience one that rivals the real world, LearnSmart Labs accomplishes it all. Learn more at [www.mhlearnsmart.com](http://www.mhlearnsmart.com).

The screenshot shows the LearnSmart Labs interface for an electrochemistry activity. The main area displays a virtual lab setup with a beaker of "DI Water", a beaker of "Zinc (II) sulfate", and a beaker of "Copper (II) sulfate". A voltmeter is connected to the two beakers. In the background, there are five beakers labeled "Copper", "Zinc", "Nickel", "Lead", and "Silver", each containing a metal electrode in its own solution. The interface includes a navigation bar at the bottom with buttons for "Notebook", "Redo Experiment", "Instructions", "Help", "Give Feedback", and "End Experiment". The top of the interface shows the course title "Electrochemistry / Electrochemistry: Activity Series".



**McGraw-Hill Tegrity®** records and distributes your class lecture with just a click of a button. Students can view it anytime and anywhere via computer, iPod, or mobile device. Tegrity indexes as it records your PowerPoint® presentations and anything shown on your computer, so students can use key words to find exactly what they want to study. Tegrity is available as an integrated feature of McGraw-Hill Connect® Chemistry and as a stand-alone product.

## Presentation Tools

Accessed from your textbook's Connect website, the presentation tools include McGraw-Hill–owned photos, artwork, animations, and other types of media that can be used to create customized lectures, visually enhanced tests and quizzes, compelling course websites, or attractive printed support materials for classroom purposes. The visual resources in this collection include the following:

- **Art** Full-color digital files of all illustrations in the book can be readily incorporated into lecture presentations, exams, or custom-made classroom materials. In addition, all files have been inserted into PowerPoint slides for ease of lecture preparation.
- **Photos** The photo collection contains digital files of photographs from the text, which can be reproduced for multiple classroom uses.
- **Tables** Every table that appears in the text has been saved in electronic form for use in classroom presentations and/or materials.
- **Animations** Numerous full-color animations illustrating important processes are also provided. Harness the visual impact of concepts in motion by importing these files into classroom presentations or online course materials.

PowerPoint materials can also be accessed through your textbook's Connect website:

- **PowerPoint Lecture Outlines** Ready-made presentations that combine art and lecture notes are provided for each chapter of the text.
- **PowerPoint Slides** For instructors who prefer to create their lectures from scratch, all illustrations, photos, sample problems, and tables have been inserted into blank PowerPoint slides, arranged by chapter.

## Computerized Test Bank

Prepared by Walter Orchard, Professor Emeritus of Tacoma Community College, over 2300 test questions that accompany *Chemistry: The Molecular Nature of Matter and*

*Change* are available utilizing Brownstone's *Diploma* testing software. *Diploma's* software allows you to quickly create a customized test using McGraw-Hill's supplied questions or by writing your own. *Diploma* allows you to create your tests without an Internet connection—just download the software and question files directly to your computer.

## Instructor's Solutions Manual

This supplement, prepared by MaryKay Orgill of the University of Nevada, Las Vegas, contains complete, worked-out solutions for *all* the end-of-chapter problems in the text. It can be found within the Instructors' Resources, on the Connect website.

## Cooperative Chemistry Laboratory Manual

Prepared by Melanie Cooper of Clemson University, this innovative manual features open-ended problems designed to simulate experience in a research lab. Working in groups, students investigate one problem over a period of several weeks, so they might complete three or four projects during the semester, rather than one preprogrammed experiment per class. The emphasis is on experimental design, analytic problem solving, and communication.

## Student Study Guide

This valuable study guide, prepared by Libby Bent Weberg, is designed to help students recognize learning style; understand how to read, classify, and create a plan for solving a problem; and practice problem-solving skills. For each section of each chapter, the guide provides study objectives and a summary of the corresponding text. Following the summary are sample problems with detailed solutions. Each chapter has true-false questions and a self-test, with all answers provided at the end of the chapter.

## Student Solutions Manual

This supplement, prepared by MaryKay Orgill of the University of Nevada, Las Vegas, contains detailed solutions and explanations for all problems in the main text that have colored numbers.

## Animations for MP3/iPod

A number of animations are available for downloading to an MP3 player or iPod through the textbook's Connect website.

# ACKNOWLEDGMENTS

**B**eginning with the 7th edition, I am delighted to welcome Professor Patricia Amateis of Virginia Tech, who shared equally all the responsibilities for creating the text. Patricia's experience teaching large classes in a research university is the perfect complement to my teaching small classes in a liberal-arts college. She is thorough in her understanding of chemistry, meticulous in her execution of the text, and, as her numerous teaching awards attest, insightful in her appreciation of student learning and devoted to their comprehension of chemistry. I am very fortunate to have her as my co-author.

It would be nearly impossible to put together a more professional, talented, supportive publishing team than my colleagues at McGraw-Hill Education: Managing Director Thomas Timp, Executive Editor David Spurgeon, Develop-

mental Editor Lora Neyens, Project Manager Peggy Selle, Designer David Hash, Marketing Manager Tami Hodge, and Director of Digital Content Shirley Hino. I feel especially grateful that, over the years, we have also become friends.

MaryKay Orgill of the University of Nevada, Las Vegas, provided a thorough accuracy check of all the new sample problems and follow-up problems as part of her superb preparation of both the Student and Instructor's Solutions Manuals.

The seventh edition and its robust learning system are the culmination of numerous chapter reviews, insightful suggestions, brainstorming sessions, symposia, focus groups, interviews, and class tests. Therefore, our deepest gratitude is extended to the following individuals, who participated in preparation of the seventh edition and its associated learning system:

Jeff Allison, *Austin Community College*  
Rebecca Barlag, *Ohio University*  
David Bateman, *Henderson State University*  
James Beil, *Lorain County Community College*  
Stacy Benson, *Oklahoma State University at Stillwater*  
Michael Brelle, *Sierra College*  
Ron Briggs, *Arizona State University*  
Gary Buckley, *Cameron University*  
Tara Carpenter, *University of Maryland, Baltimore County*  
William Case, *University of Richmond*  
Paul Charlesworth, *Michigan Technological University*  
Tabitha Chigwada, *West Virginia University*  
Claire Cohen-Schmidt, *The University of Toledo*  
Gregg Dieckman, *The University of Texas at Dallas*  
Lambert Doezema, *Loyola Marymount University*  
David Esjornson, *Southwestern Oklahoma State University*  
Deborah Exton, *University of Oregon*  
Brandon Fetterly, *University of Wisconsin-Richland*  
Andrew Frazer, *University of Central Florida*  
Kenneth Friedrich, *Portland Community College*

Simon Garrett, *California State University, Northridge*  
John Gelder, *Oklahoma State University at Stillwater*  
Alan Gengenbach, *University of Wisconsin-Eau Claire*  
Eric Goll, *Brookdale Community College*  
Donna Gosnell, *Valdosta State University*  
Brian Gute, *University of Minnesota Duluth*  
Nathan Hammer, *University of Mississippi*  
Mandë Holford, *City University of New York, York College*  
Jason Holland, *University of Central Missouri*  
Burt Hollandsworth, *Harding University*  
Byron Howell, *Tyler Junior College*  
Lenore Hoyt, *University of Louisville*  
Timothy Hughbanks, *Texas A&M University*  
Tara Hurt, *East Mississippi Community College*  
Jason Jadin, *Rochester Community and Technical College*  
Jason Jones, *Francis Marion University*  
Mary Elizabeth Kramer, *University of Delaware*  
Gerald Korenowski, *Rensselaer Polytechnic Institute*  
Mirela Krichten, *The College of New Jersey*  
Stephen Lawrence, *Saginaw Valley State University*

Ronald Marks, *North Greenville University*  
Reza Mohseni, *East Tennessee State University*  
Patricia Muisener, *University of South Florida*  
Edith Osborne, *Angelo State University*  
Greg Oswald, *North Dakota State University*  
Shadrick Paris, *Ohio University*  
Manoj Patil, *Western Iowa Tech Community College*  
Scott Pegan, *University of Denver*  
Joanna Petridou-Fischer, *Spokane Falls Community College*  
Jay Pike, *Providence College*  
Dawn Rickey, *Colorado State University*  
Linda Schultz, *Tarleton State University*  
Jacob Schroeder, *Clemson University*  
John Sibert, *The University of Texas at Dallas*  
Heather Sklenicka, *Rochester Community and Technical College*  
Joshua Telsler, *Roosevelt University*  
Michael Webb, *Central Piedmont Community College*  
Lorelei Wood, *Chandler-Gilbert Community College*  
Christine Yerkes, *University of Illinois at Urbana-Champaign*  
Curtis Zaleski, *Shippensburg University*

The following individuals helped write and review learning goal-oriented content for **LearnSmart for General Chemistry**: Peter de Lijser, California State University-Fullerton; Adam I. Keller, Columbus State Community College; and Erin L. Whitteck.

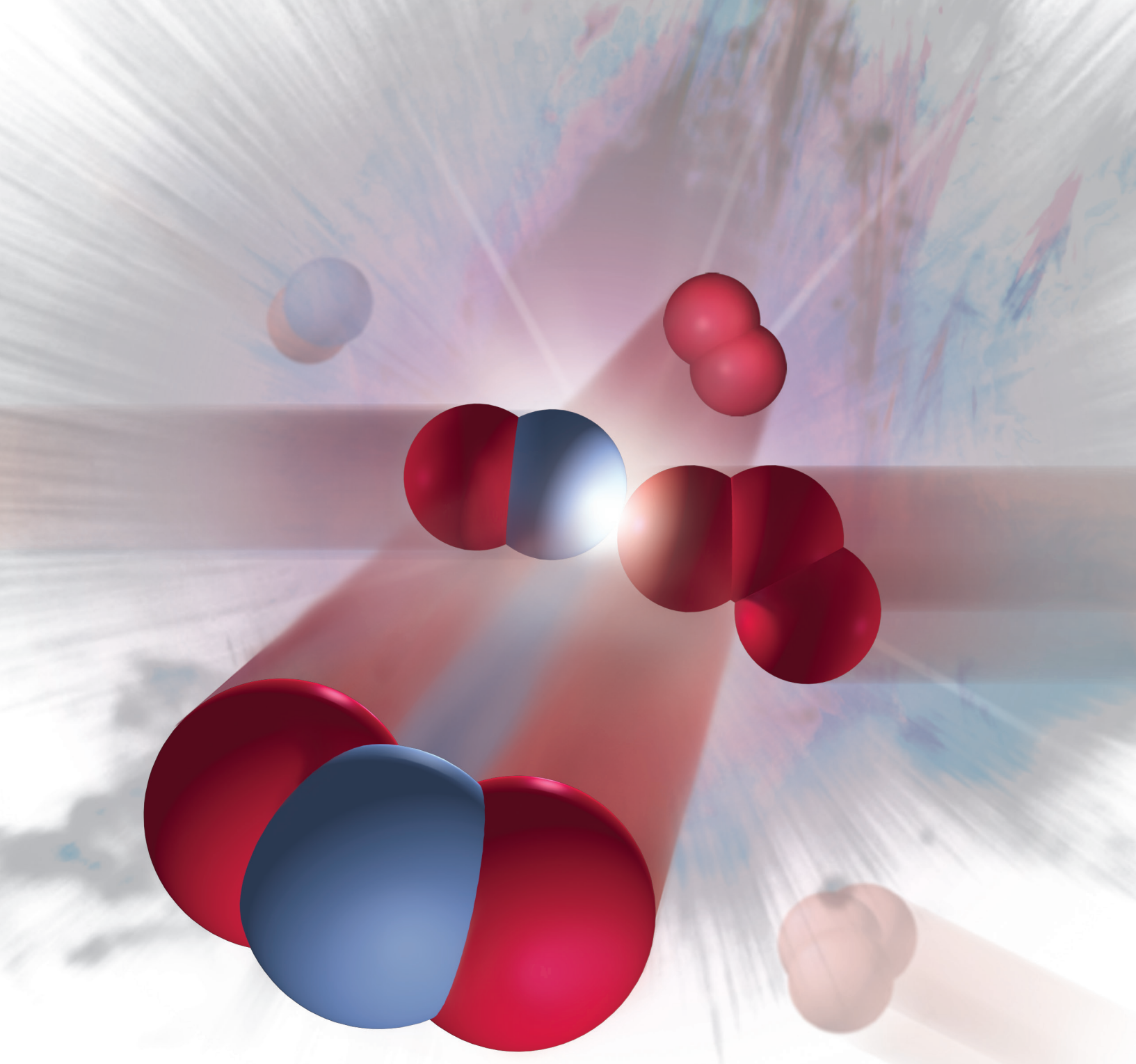
Several expert freelancers contributed as well. Jane Hoover did her usual excellent job in copyediting the text, and Janelle Pregler and Judy Bantebein followed with meticulous proof-

reading. And many thanks to Jerry Marshall, who patiently researched new stock and studio photos, and to Charles Winters, who shot the new lab photos.

In this edition, perhaps more than in any other after the first, my wife Ruth's dedication and expertise shaped the quality of the text. She checked the style, revised the layout, and monitored changes in the pages, among countless other details—all while she kept smiling and supporting me.



*This page intentionally left blank*



# CHEMISTRY

# 1

## Keys to the Study of Chemistry

### 1.1 Some Fundamental Definitions

States of Matter  
Properties of Matter and Its Changes  
Central Theme in Chemistry  
Importance of Energy

### 1.2 Chemical Arts and the Origins of Modern Chemistry

Prechemical Traditions  
Impact of Lavoisier

### 1.3 The Scientific Approach: Developing a Model

### 1.4 Measurement and Chemical Problem Solving

Features of SI Units  
SI Units in Chemistry  
Units and Conversion Factors  
Systematic Problem-Solving Approach  
Temperature Scales  
Extensive and Intensive Properties

### 1.5 Uncertainty in Measurement: Significant Figures

Determining Significant Digits  
Calculations and Rounding Off  
Precision, Accuracy, and  
Instrument Calibration



- › exponential (scientific) notation (Appendix A)

**M**aybe you're taking this course because chemistry is fundamental to understanding other natural sciences. Maybe it's required for your medical or engineering major. Or maybe you just want to learn more about the impact of chemistry on society or even on a personal, every-day level. For example, were you awakened today by your alarm clock? In chemical terms, the sound of the alarm was created when molecules aligned in the liquid-crystal display of your clock and electrons flowed to create a noise. You might have thrown off a thermal insulator of manufactured polymer (blanket) and jumped in the shower to emulsify fatty substances on your skin and hair with purified water and formulated detergents. Perhaps you next adorned yourself in an array of pleasant-smelling pigmented gels, dyed polymeric fibers, synthetic footwear, and metal-alloy jewelry. After a breakfast of nutrient-enriched, spoilage-retarded carbohydrates (cereal) in a white emulsion of fats, proteins, and monosaccharides (milk) and a cup of hot aqueous extract containing a stimulating alkaloid (coffee), you probably abraded your teeth with a colloidal dispersion of artificially flavored, dental-hardening agents (toothpaste), grabbed your portable electronic device containing ultrathin, microetched semiconductor layers powered by a series of voltaic cells (laptop), collected some objects made from processed cellulose and plastic, electronically printed with light- and oxygen-resistant inks (books), hopped in your hydrocarbon-fueled, metal-vinyl-ceramic vehicle, electrically ignited a synchronized series of controlled gaseous explosions (started your car), and took off for class!

But the true impact of chemistry extends much farther than the commercial products of daily life. The truth is that the most profound biological and environmental questions ultimately have chemical answers: How does an organism reproduce, grow, and age? What are the underlying explanations for health and disease? How can we sustain a planetary ecosystem in which plant, animal, and human populations thrive? Is there life on other worlds?

So, no matter what your reason for studying chemistry, you're going to learn some amazing ideas. And, this course comes with a bonus for developing two mental skills. The first, common to all science courses, is the ability to solve problems systematically. The second is specific to chemistry, for as you comprehend its ideas, you begin to view a hidden reality, one filled with incredibly minute particles moving at fantastic speeds, colliding billions of times a second, and interacting in ways that allow your brain to translate fluxes of electric charge into thoughts and that determine how all the matter inside and outside of you behaves. This chapter holds the keys to unlock and enter this new world.

**IN THIS CHAPTER . . .** *We discuss some central ideas about matter and energy, the process of science, units of measurement, and how scientists handle data.*

- › We begin with fundamental concepts about matter and energy and their changes.
- › A brief discussion of chemistry's origins, including some major missteps, leads to an overview of how scientists build models to study nature.
- › We examine modern units for mass, length, volume, density, and temperature and apply systematic chemical problem solving to unit conversions.
- › We see that data collection always includes some uncertainty and examine the distinction between accuracy and precision.

## 1.1 SOME FUNDAMENTAL DEFINITIONS

A good place to begin our exploration of chemistry is to define it and a few central concepts. **Chemistry** is the study of matter and its properties, the changes that matter undergoes, and the energy associated with those changes. **Matter** is the “stuff” of the universe: air, glass, planets, students—*anything that has mass and volume*. (In Section 1.4, we discuss the meanings of mass and volume in terms of how they are measured.) Chemists want to know the **composition** of matter, *the types and amounts of simpler substances that make it up*. A **substance** is a type of matter that has a defined, fixed composition.

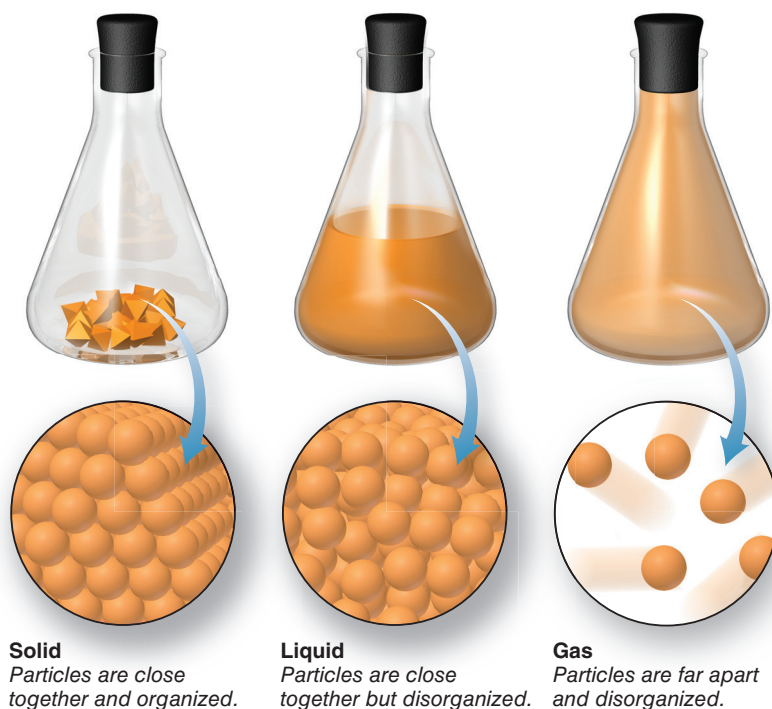
### The States of Matter

Matter occurs commonly in *three physical forms* called **states**: solid, liquid, and gas. On the macroscopic scale, each state of matter is defined by the way the sample fills a container (Figure 1.1, *flasks at top*):

- A **solid** has a fixed shape that does not conform to the container shape. Solids are *not* defined by rigidity or hardness: solid iron is rigid and hard, but solid lead is flexible, and solid wax is soft.
- A **liquid** has a varying shape that conforms to the container shape, but only to the extent of the liquid’s volume; that is, a liquid has *an upper surface*.
- A **gas** also has a varying shape that conforms to the container shape, but it fills the entire container and, thus, does *not* have a surface.

On the atomic scale, each state is defined by the relative positions of its particles (Figure 1.1, *circles at bottom*):

- In a *solid*, the particles lie next to each other in a regular, three-dimensional *array*.
- In a *liquid*, the particles also lie close together but move randomly around each other.
- In a *gas*, the particles have large distances between them and move randomly throughout the container.



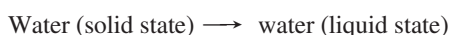
**Figure 1.1** The physical states of matter.

## The Properties of Matter and Its Changes

We learn about matter by observing its **properties**, *the characteristics that give each substance its unique identity*. To identify a person, we might observe height, weight, hair and eye color, fingerprints, and, now, even DNA pattern, until we arrive at a unique identification. To identify a substance, we observe two types of properties, physical and chemical, which are closely related to two types of change that matter undergoes.

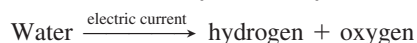
**Physical Change: No Change in Composition** **Physical properties** are characteristics a substance shows *by itself, without changing into or interacting with another substance*. These properties include melting point, electrical conductivity, and density. A **physical change** occurs when a substance *alters its physical properties, not its composition*. For example, when ice melts, several physical properties change, such as hardness, density, and ability to flow. But the composition of the sample does *not* change: it is still water. The photograph in Figure 1.2A shows what this change looks like in everyday life. The “blow-up” circles depict a magnified view of the particles making up the sample. In the icicle, the particles lie in the repeating pattern characteristic of a solid, whereas they are jumbled in the liquid droplet, but *the particles are the same* in both states of water.

*Physical change (same substance before and after):*

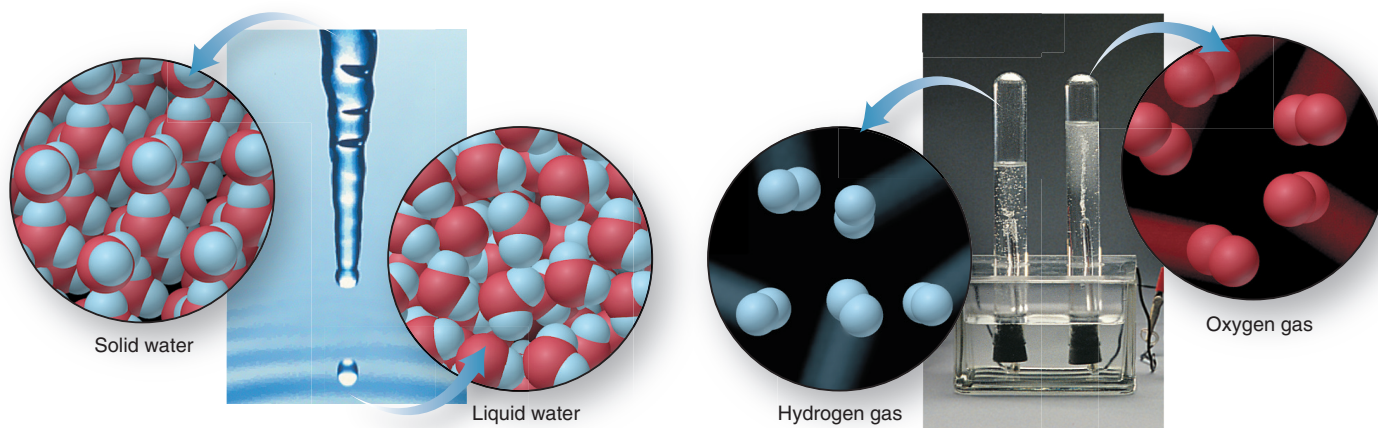


**Chemical Change: A Change in Composition** **Chemical properties** are characteristics a substance shows *as it changes into or interacts with another substance (or substances)*. Chemical properties include flammability, corrosiveness, and reactivity with acids. A **chemical change**, also called a **chemical reaction**, occurs when *a substance (or substances) is converted into a different substance (or substances)*. Figure 1.2B shows the chemical change (reaction) that occurs when you pass an electric current through water: the water decomposes (breaks down) into two other substances, hydrogen and oxygen, that bubble into the tubes. The composition *has* changed: the final sample is no longer water.

*Chemical change (different substances before and after):*



Let’s work through a sample problem that uses atomic-scale scenes to distinguish between physical and chemical change.



### A Physical change:

*Solid state of water becomes liquid state. Particles before and after remain the same, which means composition did **not** change.*

### B Chemical change:

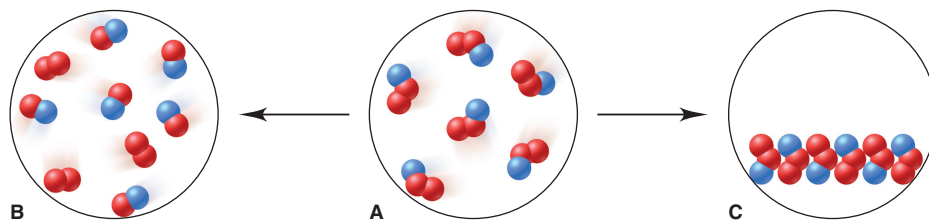
*Electric current decomposes water into different substances (hydrogen and oxygen). Particles before and after are different, which means composition **did** change.*

**Figure 1.2** The distinction between physical and chemical change.

## SAMPLE PROBLEM 1.1

## Visualizing Change on the Atomic Scale

**Problem** The scenes below represent an atomic-scale view of a sample of matter, A, undergoing two different changes, left to B and right to C:



Decide whether each depiction shows a physical or a chemical change.

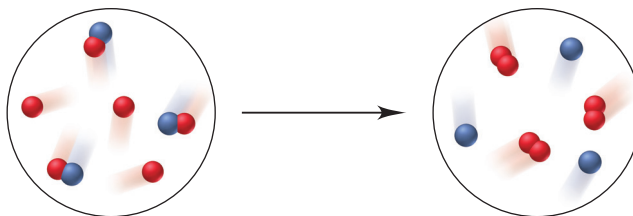
**Plan** Given depictions of the changes, we have to determine whether each represents a physical or a chemical change. The number and colors of the little spheres that make up each particle tell its “composition.” Samples with particles of the *same* composition but in a different arrangement depict a *physical* change, whereas samples with particles of a *different* composition depict a *chemical* change.

**Solution** In A, each particle consists of one blue and two red spheres. The particles in A change into two types in B, one made of red and blue spheres and the other made of two red spheres; therefore, they have undergone a **chemical change** to form different particles. The particles in C are the same as those in A, but they are closer together and arranged in a different state of matter; therefore, they have undergone a **physical change**.

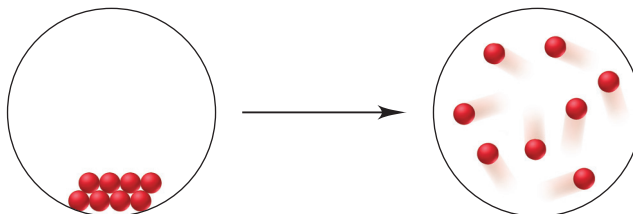
## FOLLOW-UP PROBLEMS

*Brief Solutions for all Follow-up Problems appear at the end of the chapter.*

**1.1A** Is the following change chemical or physical?



**1.1B** Is the following change chemical or physical?



## A SIMILAR PROBLEM 1.1

Like water, hydrogen, oxygen, or any other real substance, copper is also identified by *its own set* of physical and chemical properties (Table 1.1).

**Temperature and Changes in Matter** Depending on the temperature and pressure of the surroundings, many substances can exist in each of the three physical states and undergo changes in state as well. For example, as the temperature increases, solid water melts to liquid water, which boils to gaseous water (also called *water vapor*). Similarly, as the temperature drops, water vapor condenses to liquid water, and with further cooling, the liquid freezes to ice. In a steel plant, solid iron melts to liquid

Table 1.1

## Some Characteristic Properties of Copper

## Physical Properties

## Chemical Properties

Easily shaped into sheets (malleable) and wires (ductile)



Can be melted and mixed with zinc to form brass



Density = 8.95 g/cm<sup>3</sup>  
Melting point = 1083°C  
Boiling point = 2570°C

Slowly forms a blue-green carbonate in moist air



Reacts with nitric or sulfuric acid



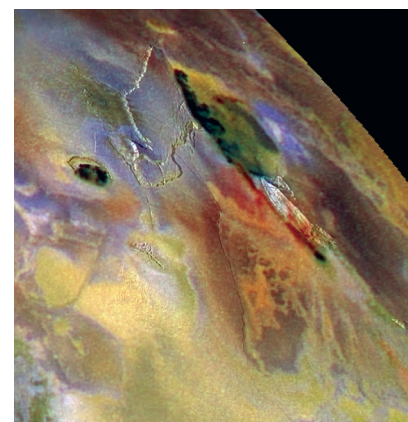
Slowly forms deep-blue solution in aqueous ammonia



(molten) iron and then cools to the solid again. And, far beyond the confines of a laboratory or steel plant, lakes of molten sulfur lie on Jupiter's moon Io (*see photo*), which is capped by poles of frozen hydrogen sulfide, a gas on Earth.

The main point is that *a physical change caused by heating can generally be reversed by cooling*. This is *not* generally true for a chemical change. For example, heating iron in moist air causes a chemical reaction that yields the brown, crumbly substance known as rust. Cooling does not reverse this change; rather, another chemical change (or series of them) is required.

The following sample problem provides practice in distinguishing some familiar examples of physical and chemical change.



Many common substances occur in unusual states on other worlds.

## SAMPLE PROBLEM 1.2

## Distinguishing Between Physical and Chemical Change

**Problem** Decide whether each of the following processes is primarily a physical or a chemical change, and explain briefly:

- Frost forms as the temperature drops on a humid winter night.
- A cornstalk grows from a seed that is watered and fertilized.
- A match ignites to form ash and a mixture of gases.
- Perspiration evaporates when you relax after jogging.
- A silver fork tarnishes slowly in air.

**Plan** The basic question we ask to decide whether a change is chemical or physical is, "Does the substance change composition or just change form?"

- Solution**
- Frost forming is a **physical change**: the drop in temperature changes water vapor (gaseous water) in humid air to ice crystals (solid water).
  - A seed growing involves **chemical change**: the seed uses water, substances from air, fertilizer, and soil, and energy from sunlight to make complex changes in composition.
  - The match burning is a **chemical change**: the combustible substances in the match head are converted into other substances.
  - Perspiration evaporating is a **physical change**: the water in sweat changes its state, from liquid to gas, but not its composition.
  - Tarnishing is a **chemical change**: silver changes to silver sulfide by reacting with sulfur-containing substances in the air.



**FOLLOW-UP PROBLEMS**

**1.2A** Decide whether each of the following processes is primarily a physical or a chemical change, and explain briefly:

- (a) Purple iodine vapor appears when solid iodine is warmed.
- (b) Gasoline fumes are ignited by a spark in an automobile engine's cylinder.
- (c) A scab forms over an open cut.

**1.2B** Decide whether each of the following processes is primarily a physical or a chemical change, and explain briefly:

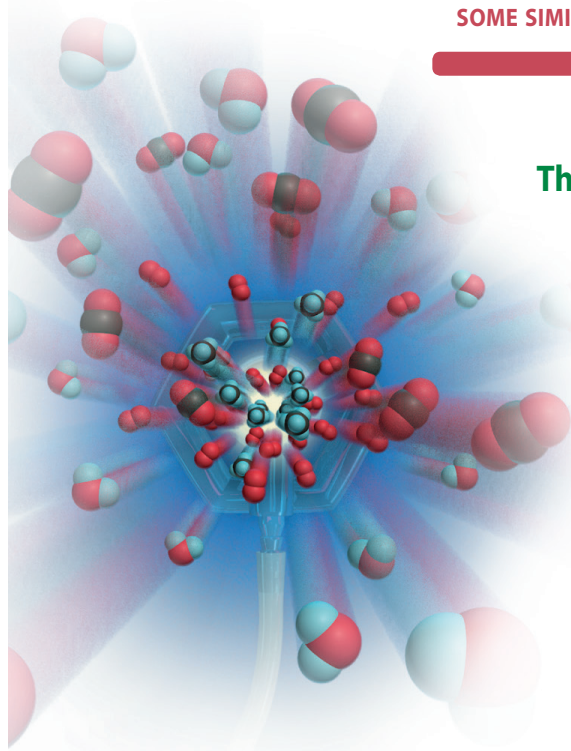
- (a) Clouds form in the sky.
- (b) Old milk turns sour.
- (c) Butter is melted to use on popcorn.

**SOME SIMILAR PROBLEMS** 1.6 and 1.7

## The Central Theme in Chemistry

Understanding the properties of a substance and the changes it undergoes leads to the central theme in chemistry: *macroscopic-scale* properties and behavior, those we can see, are the results of *atomic-scale* properties and behavior that we cannot see. The distinction between chemical and physical change is defined by composition, which we study macroscopically. But composition ultimately depends on the makeup of substances at the atomic scale. Similarly, macroscopic properties of substances in any of the three states arise from atomic-scale behavior of their particles. Picturing a chemical event on the molecular scale, even one as common as the flame of a laboratory burner (*see margin*), helps clarify what is taking place. What is happening when water boils or copper melts? What events occur in the invisible world of minute particles that cause a seed to grow, a neon light to glow, or a nail to rust? Throughout the text, we return to this central idea:

*We study **observable** changes in matter to understand their **unobservable** causes.*



Methane and oxygen form carbon dioxide and water in the flame of a lab burner. (Carbon is black, oxygen red, and hydrogen blue.)

## The Importance of Energy in the Study of Matter

Physical and chemical changes are accompanied by energy changes. **Energy** is often defined as *the ability to do work*. Essentially, all work involves moving something. Work is done when your arm lifts a book, when a car's engine moves the wheels, or when a falling rock moves the ground as it lands. The object doing the work (arm, engine, rock) transfers some of the energy it possesses to the object on which the work is done (book, wheels, ground).

*The total energy an object possesses is the sum of its potential energy and its kinetic energy.*

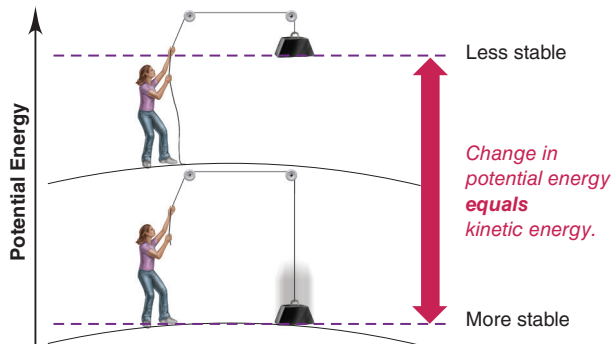
- **Potential energy** is the energy due to the **position** of the object relative to other objects.
- **Kinetic energy** is the energy due to the **motion** of the object.

Let's examine four systems that illustrate the relationship between these two forms of energy: a weight raised above the ground, two balls attached by a spring, two electrically charged particles, and a fuel and its waste products. Two concepts central to all these cases are

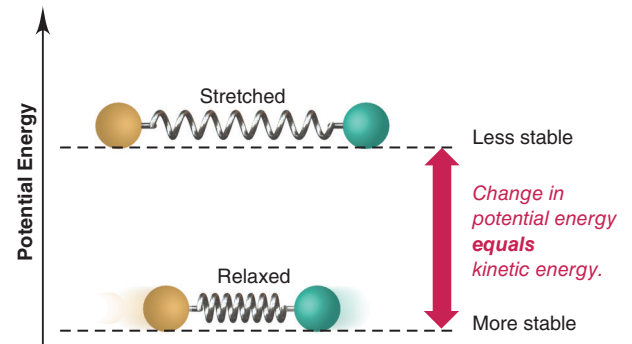
1. *When energy is converted from one form to the other, it is conserved, not destroyed.*
2. *Situations of lower energy (more stable) are favored over situations of higher energy (less stable).*

The four cases are

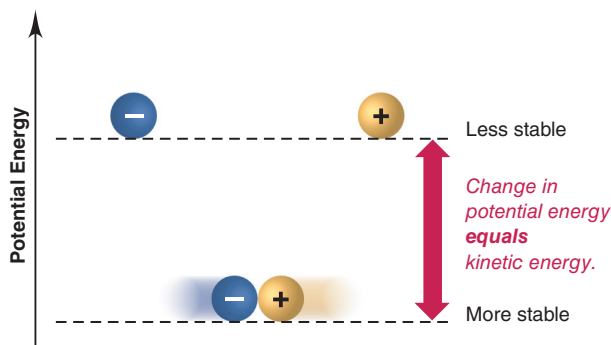
- *A weight raised above the ground* (Figure 1.3A). The energy you exert to lift a weight against gravity increases the weight's potential energy (energy due to its position). When you drop the weight, that additional potential energy is converted to kinetic energy (energy due to motion). The situation with the weight elevated and higher in potential energy is *less stable*, so the weight will fall when released, resulting in a situation that is lower in potential energy and *more stable*.
- *Two balls attached by a spring* (Figure 1.3B). When you pull the balls apart, the energy you exert to stretch the relaxed spring increases the system's potential energy. This change in potential energy is converted to kinetic energy when you release the balls. The system of balls and spring is less stable (has more potential energy) when the spring is stretched than when it is relaxed.
- *Two electrically charged particles* (Figure 1.3C). Due to interactions known as *electrostatic forces*, *opposite charges attract each other*, and *like charges repel each other*. When energy is exerted to move a positive particle away from a negative one, the potential energy of the system increases, and that increase is converted to kinetic energy when the particles are pulled together by the electrostatic attraction. Similarly, when energy is used to move two positive (or two negative) particles together, their potential energy increases and changes to kinetic energy when they are pushed apart by the electrostatic repulsion. Charged particles move naturally to a more stable situation (lower energy).
- *A fuel and its waste products* (Figure 1.3D). Matter is composed of positively and negatively charged particles. *The chemical potential energy of a substance results*



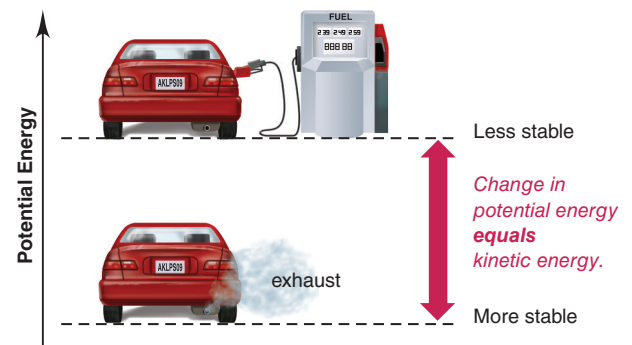
**A** A gravitational system. Potential energy is gained when a weight is lifted. It is converted to kinetic energy as the weight falls.



**B** A system of two balls attached by a spring. Potential energy is gained when the spring is stretched. It is converted to the kinetic energy of the moving balls as the spring relaxes.



**C** A system of oppositely charged particles. Potential energy is gained when the charges are separated. It is converted to kinetic energy as the attraction pulls the charges together.



**D** A system of fuel and exhaust. A fuel is higher in chemical potential energy than the exhaust. As the fuel burns, some of its potential energy is converted to the kinetic energy of the moving car.

**Figure 1.3** Potential energy is converted to kinetic energy. The dashed horizontal lines indicate the potential energy of each system before and after the change.