

Silberberg ~ Amateis

CHEMISTRY

The Molecular Nature of Matter and Change

Advanced Topics

8e

Mc
Graw
Hill
Education

Periodic Table of the Elements

MAIN-GROUP ELEMENTS

MAIN-GROUP ELEMENTS

1A (1)	1 H 1.008	2A (2)	3A (13)	4A (14)	5A (15)	6A (16)	7A (17)	8A (18)
1	1 H 1.008	2 He 4.003	5 B 10.81	6 C 12.01	7 N 14.01	8 O 16.00	9 F 19.00	2 He 4.003
2	3 Li 6.941	4 Be 9.012	5 B 10.81	6 C 12.01	7 N 14.01	8 O 16.00	9 F 19.00	10 Ne 20.18
3	11 Na 22.99	12 Mg 24.31	13 Al 26.98	14 Si 28.09	15 P 30.97	16 S 32.06	17 Cl 35.45	18 Ar 39.95
4	19 K 39.10	20 Ca 40.08	31 Ga 69.72	32 Ge 72.63	33 As 74.92	34 Se 78.96	35 Br 79.90	36 Kr 83.80
5	37 Rb 85.47	38 Sr 87.62	49 In 114.8	50 Sn 118.7	51 Sb 121.8	52 Te 127.6	53 I 126.9	54 Xe 131.3
6	55 Cs 132.9	56 Ba 137.3	81 Tl 204.4	82 Pb 207.2	83 Bi 209.0	84 Po (209)	85 At (210)	86 Rn (222)
7	87 Fr (223)	88 Ra (226)	113 Nh (284)	114 Fl (289)	115 Mc (288)	116 Lv (293)	117 Ts (294)	118 Og (294)

Legend for element categories:

- Metals (main-group)
- Metals (transition)
- Metals (inner transition)
- Metalloids
- Nonmetals

Legend for element properties:

- Atomic number
- Atomic symbol
- Atomic mass (amu)

TRANSITION ELEMENTS

1B (1)	2B (12)
8B (9)	10 (10)
7B (7)	8 (8)
6B (6)	9 (9)
5B (5)	10 (10)
4B (4)	11 (11)
3B (3)	12 (12)
21 Sc 44.96	22 Ti 47.87
23 V 50.94	24 Cr 52.00
25 Mn 54.94	26 Fe 55.85
27 Co 58.93	28 Ni 58.69
29 Cu 63.55	30 Zn 65.38
39 Y 88.91	40 Zr 91.22
41 Nb 92.91	42 Mo 95.96
43 Tc (98)	44 Ru 101.1
45 Rh 102.9	46 Pd 106.4
47 Cd 112.4	48 Hg 200.6
73 Ta 180.9	74 W 183.8
75 Re 186.2	76 Os 190.2
77 Ir 192.2	78 Pt 195.1
79 Au 197.0	80 Hg 200.6
103 Lr (262)	104 Rf (261)
105 Db (268)	106 Sg (271)
107 Bh (270)	108 Hs (277)
109 Mt (276)	110 Ds (281)
111 Rg (280)	112 Cn (285)

INNER TRANSITION ELEMENTS

6	Lanthanides	58 Ce 140.1	59 Pr 140.9	60 Nd 144.2	61 Pm (145)	62 Sm 150.4	63 Eu 152.0	64 Gd 157.3	65 Tb 158.9	66 Dy 162.5	67 Ho 164.9	68 Er 167.3	69 Tm 168.9	70 Yb 173.1	71 Lu 175.0
7	Actinides	90 Th 232.0	91 Pa (231)	92 U 238.0	93 Np (237)	94 Pu (244)	95 Am (243)	96 Cm (247)	97 Bk (247)	98 Cf (251)	99 Es (252)	100 Fm (257)	101 Md (258)	102 No (259)	103 Lr (262)

Period

The Elements

Name	Symbol	Atomic Number	Atomic Mass*	Name	Symbol	Atomic Number	Atomic Mass*
Actinium	Ac	89	(227)	Mendelevium	Md	101	(256)
Aluminum	Al	13	26.98	Mercury	Hg	80	200.6
Americium	Am	95	(243)	Molybdenum	Mo	42	95.94
Antimony	Sb	51	121.8	Moscovium	Mc	115	(288)
Argon	Ar	18	39.95	Neodymium	Nd	60	144.2
Arsenic	As	33	74.92	Neon	Ne	10	20.18
Astatine	At	85	(210)	Neptunium	Np	93	(244)
Barium	Ba	56	137.3	Nickel	Ni	28	58.70
Berkelium	Bk	97	(247)	Nihonium	Nh	113	(284)
Beryllium	Be	4	9.012	Niobium	Nb	41	92.91
Bismuth	Bi	83	209.0	Nitrogen	N	7	14.01
Bohrium	Bh	107	(267)	Nobelium	No	102	(253)
Boron	B	5	10.81	Oganesson	Og	118	(294)
Bromine	Br	35	79.90	Osmium	Os	76	190.2
Cadmium	Cd	48	112.4	Oxygen	O	8	16.00
Calcium	Ca	20	40.08	Palladium	Pd	46	106.4
Californium	Cf	98	(249)	Phosphorus	P	15	30.97
Carbon	C	6	12.01	Platinum	Pt	78	195.1
Cerium	Ce	58	140.1	Plutonium	Pu	94	(242)
Cesium	Cs	55	132.9	Polonium	Po	84	(209)
Chlorine	Cl	17	35.45	Potassium	K	19	39.10
Chromium	Cr	24	52.00	Praseodymium	Pr	59	140.9
Cobalt	Co	27	58.93	Promethium	Pm	61	(145)
Copernicium	Cn	112	(285)	Protactinium	Pa	91	(231)
Copper	Cu	29	63.55	Radium	Ra	88	(226)
Curium	Cm	96	(247)	Radon	Rn	86	(222)
Darmstadtium	Ds	110	(281)	Rhenium	Re	75	186.2
Dubnium	Db	105	(262)	Rhodium	Rh	45	102.9
Dysprosium	Dy	66	162.5	Roentgenium	Rg	111	(272)
Einsteinium	Es	99	(254)	Rubidium	Rb	37	85.47
Erbium	Er	68	167.3	Ruthenium	Ru	44	101.1
Europium	Eu	63	152.0	Rutherfordium	Rf	104	(263)
Fermium	Fm	100	(253)	Samarium	Sm	62	150.4
Fleborium	Fl	114	(289)	Scandium	Sc	21	44.96
Fluorine	F	9	19.00	Seaborgium	Sg	106	(266)
Francium	Fr	87	(223)	Selenium	Se	34	78.96
Gadolinium	Gd	64	157.3	Silicon	Si	14	28.09
Gallium	Ga	31	69.72	Silver	Ag	47	107.9
Germanium	Ge	32	72.61	Sodium	Na	11	22.99
Gold	Au	79	197.0	Strontium	Sr	38	87.62
Hafnium	Hf	72	178.5	Sulfur	S	16	32.07
Hassium	Hs	108	(277)	Tantalum	Ta	73	180.9
Helium	He	2	4.003	Technetium	Tc	43	(98)
Holmium	Ho	67	164.9	Tellurium	Te	52	127.6
Hydrogen	H	1	1.008	Tennessee	Ts	117	(294)
Indium	In	49	114.8	Terbium	Tb	65	158.9
Iodine	I	53	126.9	Thallium	Tl	81	204.4
Iridium	Ir	77	192.2	Thorium	Th	90	232.0
Iron	Fe	26	55.85	Thulium	Tm	69	168.9
Krypton	Kr	36	83.80	Tin	Sn	50	118.7
Lanthanum	La	57	138.9	Titanium	Ti	22	47.88
Lawrencium	Lr	103	(257)	Tungsten	W	74	183.9
Lead	Pb	82	207.2	Uranium	U	92	238.0
Lithium	Li	3	6.941	Vanadium	V	23	50.94
Livermorium	Lv	116	(293)	Xenon	Xe	54	131.3
Lutetium	Lu	71	175.0	Ytterbium	Yb	70	173.0
Magnesium	Mg	12	24.31	Yttrium	Y	39	88.91
Manganese	Mn	25	54.94	Zinc	Zn	30	65.41
Meitnerium	Mt	109	(268)	Zirconium	Zr	40	91.22

*All atomic masses are given to four significant figures. Values in parentheses represent the mass number of the most stable isotope.

Silberberg ~ Amateis

CHEMISTRY

The Molecular Nature of Matter and Change

Advanced Topics

8e

Mc
Graw
Hill
Education



CHEMISTRY: THE MOLECULAR NATURE OF MATTER AND CHANGE WITH ADVANCED TOPICS, EIGHTH EDITION

Published by McGraw-Hill Education, 2 Penn Plaza, New York, NY 10121. Copyright © 2018 by McGraw-Hill Education. All rights reserved. Printed in the United States of America. Previous editions © 2016, 2012, and 2009. 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 LWI 21 20 19 18 17

ISBN 978-1-259-74109-8

MHID 1-259-74109-5

Chief Product Officer, SVP Products & Markets: *G. Scott Virkler*
Vice President, General Manager, Products & Markets: *Marty Lange*
Vice President, Content Design & Delivery: *Betsy Whalen*
Managing Director: *Thomas Timp*
Director: *David Spurgeon, Ph.D.*
Director, Product Development: *Rose Koos*
Associate Director of Digital Content: *Robin Reed*
Marketing Manager: *Matthew Garcia*
Market Development Manager: *Shannon O'Donnell*
Director of Digital Content: *Shirley Hino, Ph.D.*
Digital Product Developer: *Joan Weber*
Director, Content Design & Delivery: *Linda Avenarius*
Program Manager: *Lora Neyens*
Content Project Managers: *Laura Bies, Tammy Juran & Sandy Schnee*
Buyer: *Sandy Ludovissy*
Design: *David W. Hash*
Content Licensing Specialists: *Ann Marie Jannette & Lorraine Buczek*
Cover Image: © *Don Farrall/Photographer's Choice RF/Getty Images*
Compositor: *Aptara[®], Inc.*
Printer: *LSC Communications*

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

Names: Silberberg, Martin S. (Martin Stuart), 1945- | Amateis, Patricia.
Title: Chemistry : the molecular nature of matter and change : with advanced topics / Silberberg, Amateis.
Description: 8e [8th edition, revised]. | New York, NY : McGraw-Hill Education, [2018] | Includes index.
Identifiers: LCCN 2017009580| ISBN 9781259741098 (alk. paper) | ISBN 1259741095 (alk. paper)
Subjects: LCSH: Chemistry—Textbooks.
Classification: LCC QD33.2 .S55 2018b | DDC 540—dc23 LC record available at <https://lcn.loc.gov/2017009580>

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 xxxii

- 1** *Keys to Studying Chemistry: Definitions, Units, and Problem Solving* 2
- 2** *The Components of Matter* 42
- 3** *Stoichiometry of Formulas and Equations* 94
- 4** *Three Major Classes of Chemical Reactions* 144
- 5** *Gases and the Kinetic-Molecular Theory* 204
- 6** *Thermochemistry: Energy Flow and Chemical Change* 256
- 7** *Quantum Theory and Atomic Structure* 294
- 8** *Electron Configuration and Chemical Periodicity* 330
- 9** *Models of Chemical Bonding* 368
- 10** *The Shapes of Molecules* 404
- 11** *Theories of Covalent Bonding* 442
- 12** *Intermolecular Forces: Liquids, Solids, and Phase Changes* 470
- 13** *The Properties of Mixtures: Solutions and Colloids* 532
- 14** *Periodic Patterns in the Main-Group Elements* 584
- 15** *Organic Compounds and the Atomic Properties of Carbon* 632
- 16** *Kinetics: Rates and Mechanisms of Chemical Reactions* 690
- 17** *Equilibrium: The Extent of Chemical Reactions* 746
- 18** *Acid-Base Equilibria* 792
- 19** *Ionic Equilibria in Aqueous Systems* 842
- 20** *Thermodynamics: Entropy, Free Energy, and Reaction Direction* 894
- 21** *Electrochemistry: Chemical Change and Electrical Work* 938
- 22** *The Elements in Nature and Industry* 996
- 23** *Transition Elements and Their Coordination Compounds* 1036
- 24** *Nuclear Reactions and Their Applications* 1072

Appendix A Common Mathematical Operations in Chemistry A-1

Appendix B Standard Thermodynamic Values for Selected Substances A-5

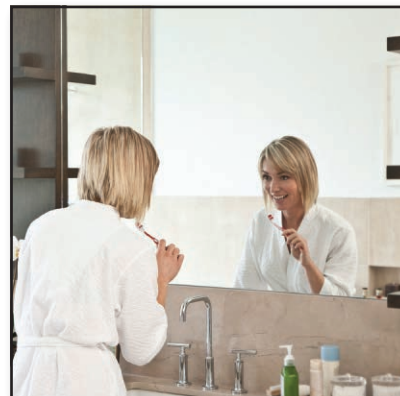
Appendix C Equilibrium Constants for Selected Substances A-8

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

Appendix E Answers to Selected Problems A-15

Glossary G-1

Index I-1



© Fancy Collection/SuperStock RF

CHAPTER

1

Keys to Studying Chemistry: Definitions, Units, and Problem Solving 2

- | | | |
|-------------------------------------------------------------------------------------------------------------------------------------------------------------------------------------------------------------------------------------------------------------------------------------------------------------------------------------------------------------------------------------------------------------------------------------------|---------------------------------------------------------------------------------------------------------------------------------------------------------------------------------------------------------------------------------------------------------------------------------------------------------------------------------------------------------------------------------------------------------------------------------------------------------|--------------------------------------------------------------------------------------------------------------------------------------------------------------------------------------------------------------------------------------------------------------------------------------------------------------------------|
| <p>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</p> <p>1.2 Chemical Arts and the Origins of Modern Chemistry 10
 Prechemical Traditions 10
 The Phlogiston Fiasco and the Impact of Lavoisier 11</p> | <p>1.3 The Scientific Approach: Developing a Model 12</p> <p>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 18
 A Systematic Approach to Solving Chemistry Problems 19
 Temperature Scales 25
 Extensive and Intensive Properties 27</p> | <p>1.5 Uncertainty in Measurement: Significant Figures 28
 Determining Which Digits Are Significant 29
 Significant Figures: Calculations and Rounding Off 30
 Precision, Accuracy, and Instrument Calibration 32</p> <p>CHAPTER REVIEW GUIDE 33
 PROBLEMS 37</p> |
|-------------------------------------------------------------------------------------------------------------------------------------------------------------------------------------------------------------------------------------------------------------------------------------------------------------------------------------------------------------------------------------------------------------------------------------------|---------------------------------------------------------------------------------------------------------------------------------------------------------------------------------------------------------------------------------------------------------------------------------------------------------------------------------------------------------------------------------------------------------------------------------------------------------|--------------------------------------------------------------------------------------------------------------------------------------------------------------------------------------------------------------------------------------------------------------------------------------------------------------------------|

CHAPTER

2

The Components of Matter 42

- | | | |
|----------------------------------------------------------------------------------------------------------------------------------------------------------------------------------------------------------------------------------------------------------------------------------------------------------------------------------------------------------------------------------------------------------------------------------------------------------------------------------------------------------------------------------------------------------------------------------------------------------------------------------------------------------------------------------------------|--------------------------------------------------------------------------------------------------------------------------------------------------------------------------------------------------------------------------------------------------------------------------------------------------------------------------------------------------------------------------------------------------------------------------------------------------------------------------------------------------------------------------------------------------------------------------------------|-----------------------------------------------------------------------------------------------------------------------------------------------------------------------------------------------------------------------------------------------------------------------------------------------------------------------------------------------------------------------------------------------------------------------------------------------------------------------------------------------------------------------------------|
| <p>2.1 Elements, Compounds, and Mixtures: An Atomic Overview 44</p> <p>2.2 The Observations That Led to an Atomic View of Matter 46
 Mass Conservation 46
 Definite Composition 47
 Multiple Proportions 49</p> <p>2.3 Dalton's Atomic Theory 50
 Postulates of the Atomic Theory 50
 How the Theory Explains the Mass Laws 50</p> <p>2.4 The Observations That Led to the Nuclear Atom Model 52
 Discovery of the Electron and Its Properties 52
 Discovery of the Atomic Nucleus 54</p> <p>2.5 The Atomic Theory Today 55
 Structure of the Atom 55</p> | <p>Atomic Number, Mass Number, and Atomic Symbol 56
 Isotopes 57
 Atomic Masses of the Elements 57</p> <p>TOOLS OF THE LABORATORY: MASS SPECTROMETRY 60</p> <p>2.6 Elements: A First Look at the Periodic Table 61</p> <p>2.7 Compounds: Introduction to Bonding 64
 The Formation of Ionic Compounds 64
 The Formation of Covalent Substances 66</p> <p>2.8 Compounds: Formulas, Names, and Masses 68
 Binary Ionic Compounds 68
 Compounds That Contain Polyatomic Ions 71</p> | <p>Acid Names from Anion Names 74
 Binary Covalent Compounds 74
 The Simplest Organic Compounds: Straight-Chain Alkanes 76
 Molecular Masses from Chemical Formulas 76
 Representing Molecules with Formulas and Models 78</p> <p>2.9 Mixtures: Classification and Separation 81
 An Overview of the Components of Matter 81</p> <p>TOOLS OF THE LABORATORY: BASIC SEPARATION TECHNIQUES 83</p> <p>CHAPTER REVIEW GUIDE 84
 PROBLEMS 86</p> |
|----------------------------------------------------------------------------------------------------------------------------------------------------------------------------------------------------------------------------------------------------------------------------------------------------------------------------------------------------------------------------------------------------------------------------------------------------------------------------------------------------------------------------------------------------------------------------------------------------------------------------------------------------------------------------------------------|--------------------------------------------------------------------------------------------------------------------------------------------------------------------------------------------------------------------------------------------------------------------------------------------------------------------------------------------------------------------------------------------------------------------------------------------------------------------------------------------------------------------------------------------------------------------------------------|-----------------------------------------------------------------------------------------------------------------------------------------------------------------------------------------------------------------------------------------------------------------------------------------------------------------------------------------------------------------------------------------------------------------------------------------------------------------------------------------------------------------------------------|



Source: NASA

CHAPTER

3

Stoichiometry of Formulas and Equations 94

- 3.1 The Mole 95**
 Defining the Mole 95
 Determining Molar Mass 96
 Converting Between Amount, Mass, and Number of Chemical Entities 97
 The Importance of Mass Percent 102
- 3.2 Determining the Formula of an Unknown Compound 104**
 Empirical Formulas 105
 Molecular Formulas 106
- 3.3 Writing and Balancing Chemical Equations 111**
- 3.4 Calculating Quantities of Reactant and Product 116**
 Stoichiometrically Equivalent Molar Ratios from the Balanced Equation 116
- Chemical Formulas and Molecular Structures; Isomers 110
- Reactions That Occur in a Sequence 120
 Reactions That Involve a Limiting Reactant 122
 Theoretical, Actual, and Percent Reaction Yields 127
- CHAPTER REVIEW GUIDE 130
 PROBLEMS 135

CHAPTER

4

Three Major Classes of Chemical Reactions 144

- 4.1 Solution Concentration and the Role of Water as a Solvent 145**
 The Polar Nature of Water 146
 Ionic Compounds in Water 146
 Covalent Compounds in Water 150
 Expressing Concentration in Terms of Molarity 150
 Amount-Mass-Number Conversions Involving Solutions 151
 Preparing and Diluting Molar Solutions 152
- 4.2 Writing Equations for Aqueous Ionic Reactions 155**
- 4.3 Precipitation Reactions 157**
 The Key Event: Formation of a Solid from Dissolved Ions 157
- Predicting Whether a Precipitate Will Form 157
 Stoichiometry of Precipitation Reactions 162
- 4.4 Acid-Base Reactions 165**
 The Key Event: Formation of H_2O from H^+ and OH^- 167
 Proton Transfer in Acid-Base Reactions 168
 Stoichiometry of Acid-Base Reactions: Acid-Base Titrations 172
- 4.5 Oxidation-Reduction (Redox) Reactions 174**
 The Key Event: Movement of Electrons Between Reactants 174
 Some Essential Redox Terminology 175
- Using Oxidation Numbers to Monitor Electron Charge 176
 Stoichiometry of Redox Reactions: Redox Titrations 179
- 4.6 Elements in Redox Reactions 181**
 Combination Redox Reactions 181
 Decomposition Redox Reactions 182
 Displacement Redox Reactions and Activity Series 184
 Combustion Reactions 186
- 4.7 The Reversibility of Reactions and the Equilibrium State 188**
 CHAPTER REVIEW GUIDE 190
 PROBLEMS 196

CHAPTER

5

Gases and the Kinetic-Molecular Theory 204

- 5.1 An Overview of the Physical States of Matter 205**
- 5.2 Gas Pressure and Its Measurement 207**
 Measuring Gas Pressure: Barometers and Manometers 208
 Units of Pressure 209
- 5.3 The Gas Laws and Their Experimental Foundations 210**
 The Relationship Between Volume and Pressure: Boyle's Law 211
 The Relationship Between Volume and Temperature: Charles's Law 212
 The Relationship Between Volume and Amount: Avogadro's Law 214
 Gas Behavior at Standard Conditions 215
- The Ideal Gas Law 216
 Solving Gas Law Problems 217
- 5.4 Rearrangements of the Ideal Gas Law 222**
 The Density of a Gas 222
 The Molar Mass of a Gas 224
 The Partial Pressure of Each Gas in a Mixture of Gases 225
 The Ideal Gas Law and Reaction Stoichiometry 228
- 5.5 The Kinetic-Molecular Theory: A Model for Gas Behavior 231**
 How the Kinetic-Molecular Theory Explains the Gas Laws 231
 Effusion and Diffusion 236
- The Chaotic World of Gases: Mean Free Path and Collision Frequency 238
- CHEMICAL CONNECTIONS TO ATMOSPHERIC SCIENCE: HOW THE GAS LAWS APPLY TO EARTH'S ATMOSPHERE 239**
- 5.6 Real Gases: Deviations from Ideal Behavior 241**
 Effects of Extreme Conditions on Gas Behavior 241
 The van der Waals Equation: Adjusting the Ideal Gas Law 243
- CHAPTER REVIEW GUIDE 244
 PROBLEMS 247



© Maya Kruchankova/Shutterstock.com

CHAPTER

6

Thermochemistry: Energy Flow and Chemical Change 256

- 6.1 Forms of Energy and Their Interconversion 257**
 Defining the System and Its Surroundings 258
 Energy Change (ΔE): Energy Transfer to or from a System 258
 Heat and Work: Two Forms of Energy Transfer 258
 The Law of Energy Conservation 261
 Units of Energy 261
 State Functions and the Path Independence of the Energy Change 262
 Calculating Pressure-Volume Work (PV Work) 263
- 6.2 Enthalpy: Changes at Constant Pressure 265**
 The Meaning of Enthalpy 265
 Comparing ΔE and ΔH 265
 Exothermic and Endothermic Processes 266
- 6.3 Calorimetry: Measuring the Heat of a Chemical or Physical Change 268**
 Specific Heat Capacity 268
 The Two Major Types of Calorimetry 269
- 6.4 Stoichiometry of Thermochemical Equations 273**
- 6.5 Hess's Law: Finding ΔH of Any Reaction 275**
- 6.6 Standard Enthalpies of Reaction ($\Delta H_{\text{rxn}}^{\circ}$) 277**
 Formation Equations and Their Standard Enthalpy Changes 277
 Determining $\Delta H_{\text{rxn}}^{\circ}$ from ΔH_f° Values for Reactants and Products 279
- CHEMICAL CONNECTIONS TO ENVIRONMENTAL SCIENCE: THE FUTURE OF ENERGY USE 281**
- CHAPTER REVIEW GUIDE 285
 PROBLEMS 288

CHAPTER

7

Quantum Theory and Atomic Structure 294

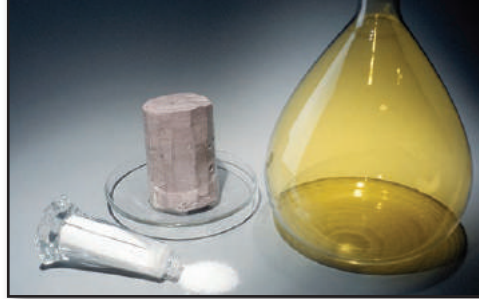
- 7.1 The Nature of Light 295**
 The Wave Nature of Light 296
 The Particle Nature of Light 299
- 7.2 Atomic Spectra 302**
 Line Spectra and the Rydberg Equation 302
 The Bohr Model of the Hydrogen Atom 303
 The Energy Levels of the Hydrogen Atom 305
- 7.3 The Wave-Particle Duality of Matter and Energy 310**
 The Wave Nature of Electrons and the Particle Nature of Photons 310
 Heisenberg's Uncertainty Principle 313
- 7.4 The Quantum-Mechanical Model of the Atom 314**
 The Schrödinger Equation, the Atomic Orbital, and the Probable Location of the Electron 314
- TOOLS OF THE LABORATORY: SPECTROMETRY IN CHEMICAL ANALYSIS 308**
- Quantum Numbers of an Atomic Orbital 319
 Quantum Numbers and Energy Levels 321
 Shapes of Atomic Orbitals 323
 The Special Case of Energy Levels in the Hydrogen Atom 325
- CHAPTER REVIEW GUIDE 326
 PROBLEMS 329

CHAPTER

8

Electron Configuration and Chemical Periodicity 330

- 8.1 Characteristics of Many-Electron Atoms 332**
 The Electron-Spin Quantum Number 332
 The Exclusion Principle 333
 Electrostatic Effects and Energy-Level Splitting 333
- 8.2 The Quantum-Mechanical Model and the Periodic Table 335**
 Building Up Period 1 335
 Building Up Period 2 336
 Building Up Period 3 338
- Building Up Period 4: The First Transition Series 339
 General Principles of Electron Configurations 340
 Intervening Series: Transition and Inner Transition Elements 342
 Similar Electron Configurations Within Groups 342
- 8.3 Trends in Three Atomic Properties 345**
 Trends in Atomic Size 345
- Trends in Ionization Energy 348
 Trends in Electron Affinity 351
- 8.4 Atomic Properties and Chemical Reactivity 353**
 Trends in Metallic Behavior 353
 Properties of Monatomic Ions 355
- CHAPTER REVIEW GUIDE 361
 PROBLEMS 363



© Chip Clark/Fundamental Photographs, NYC

CHAPTER

9

Models of Chemical Bonding 368

9.1 Atomic Properties and Chemical Bonds 369

The Three Ways Elements Combine 369
Lewis Symbols and the Octet Rule 371

9.2 The Ionic Bonding Model 372

Why Ionic Compounds Form:
The Importance of Lattice Energy 373
Periodic Trends in Lattice Energy 375
How the Model Explains the Properties of Ionic Compounds 377

9.3 The Covalent Bonding Model 379

The Formation of a Covalent Bond 379
Bonding Pairs and Lone Pairs 380
Properties of a Covalent Bond:
Order, Energy, and Length 380

How the Model Explains the Properties of Covalent Substances 383

TOOLS OF THE LABORATORY: INFRARED SPECTROSCOPY 384**9.4 Bond Energy and Chemical Change 386**

Changes in Bond Energy: Where Does $\Delta H_{\text{rxn}}^{\circ}$ Come From? 386
Using Bond Energies to Calculate $\Delta H_{\text{rxn}}^{\circ}$ 386
Bond Strengths and the Heat Released from Fuels and Foods 389

9.5 Between the Extremes: Electronegativity and Bond Polarity 390

Electronegativity 390

Bond Polarity and Partial Ionic Character 392
The Gradation in Bonding Across a Period 394

9.6 An Introduction to Metallic Bonding 395

The Electron-Sea Model 395
How the Model Explains the Properties of Metals 396

CHAPTER REVIEW GUIDE 397

PROBLEMS 399

CHAPTER

10

The Shapes of Molecules 404

10.1 Depicting Molecules and Ions with Lewis Structures 405

Applying the Octet Rule to Write Lewis Structures 405
Resonance: Delocalized Electron-Pair Bonding 409
Formal Charge: Selecting the More Important Resonance Structure 411
Lewis Structures for Exceptions to the Octet Rule 413

10.2 Valence-Shell Electron-Pair Repulsion (VSEPR) Theory 417

Electron-Group Arrangements and Molecular Shapes 418
The Molecular Shape with Two Electron Groups (Linear Arrangement) 419

Molecular Shapes with Three Electron Groups (Trigonal Planar Arrangement) 419

Molecular Shapes with Four Electron Groups (Tetrahedral Arrangement) 420

Molecular Shapes with Five Electron Groups (Trigonal Bipyramidal Arrangement) 421

Molecular Shapes with Six Electron Groups (Octahedral Arrangement) 422

Using VSEPR Theory to Determine Molecular Shape 423

Molecular Shapes with More Than One Central Atom 426

10.3 Molecular Shape and Molecular Polarity 428

Bond Polarity, Bond Angle, and Dipole Moment 428
The Effect of Molecular Polarity on Behavior 430

CHEMICAL CONNECTIONS TO SENSORY PHYSIOLOGY: MOLECULAR SHAPE, BIOLOGICAL RECEPTORS, AND THE SENSE OF SMELL 431

CHAPTER REVIEW GUIDE 432

PROBLEMS 437



© Richard Megna/Fundamental Photographs, NYC

CHAPTER

11

Theories of Covalent Bonding 442

11.1 Valence Bond (VB) Theory and Orbital Hybridization 443

The Central Themes of VB Theory 443
Types of Hybrid Orbitals 444

11.2 Modes of Orbital Overlap and the Types of Covalent Bonds 451

Orbital Overlap in Single and Multiple Bonds 451
Orbital Overlap and Rotation Within a Molecule 455

11.3 Molecular Orbital (MO) Theory and Electron Delocalization 455

The Central Themes of MO Theory 455
Homonuclear Diatomic Molecules of Period 2 Elements 458
Two Heteronuclear Diatomic Molecules: HF and NO 462
Two Polyatomic Molecules: Benzene and Ozone 463

CHAPTER REVIEW GUIDE 464

PROBLEMS 466

CHAPTER

12

Intermolecular Forces: Liquids, Solids, and Phase Changes 470

12.1 An Overview of Physical States and Phase Changes 471

12.2 Quantitative Aspects of Phase Changes 474

Heat Involved in Phase Changes 475
The Equilibrium Nature of Phase Changes 478
Phase Diagrams: Effect of Pressure and Temperature on Physical State 482

12.3 Types of Intermolecular Forces 484

How Close Can Molecules Approach Each Other? 484
Ion-Dipole Forces 485
Dipole-Dipole Forces 485
The Hydrogen Bond 486

Polarizability and Induced Dipole Forces 487

Dispersion (London) Forces 488

12.4 Properties of the Liquid State 490

Surface Tension 491
Capillarity 491
Viscosity 492

12.5 The Uniqueness of Water 493

Solvent Properties of Water 493
Thermal Properties of Water 493
Surface Properties of Water 494
The Unusual Density of Solid Water 494

12.6 The Solid State: Structure, Properties, and Bonding 495

Structural Features of Solids 495

TOOLS OF THE LABORATORY: X-RAY DIFFRACTION ANALYSIS AND SCANNING TUNNELING MICROSCOPY 502

Types and Properties of Crystalline Solids 503
Amorphous Solids 506
Bonding in Solids: Molecular Orbital Band Theory 506

12.7 Advanced Materials 509

Electronic Materials 509
Liquid Crystals 511
Ceramic Materials 514
Polymeric Materials 516
Nanotechnology: Designing Materials Atom by Atom 521

CHAPTER REVIEW GUIDE 523

PROBLEMS 525



© amnat11/Shutterstock.com

CHAPTER 13 *The Properties of Mixtures: Solutions and Colloids* 532

- 13.1 Types of Solutions: Intermolecular Forces and Solubility** 534
 Intermolecular Forces in Solution 534
 Liquid Solutions and the Role of Molecular Polarity 535
 Gas Solutions and Solid Solutions 537
- 13.2 Intermolecular Forces and Biological Macromolecules** 539
 The Structures of Proteins 539
 Dual Polarity in Soaps, Membranes, and Antibiotics 541
 The Structure of DNA 542
- 13.3 Why Substances Dissolve: Breaking Down the Solution Process** 544
 The Heat of Solution and Its Components 544
- The Heat of Hydration: Dissolving Ionic Solids in Water 545
 The Solution Process and the Change in Entropy 547
- 13.4 Solubility as an Equilibrium Process** 549
 Effect of Temperature on Solubility 549
 Effect of Pressure on Solubility 551
- 13.5 Concentration Terms** 552
 Molarity and Molality 552
 Parts of Solute by Parts of Solution 554
 Interconverting Concentration Terms 556
- 13.6 Colligative Properties of Solutions** 557
 Nonvolatile Nonelectrolyte Solutions 558
- Using Colligative Properties to Find Solute Molar Mass 563
 Volatile Nonelectrolyte Solutions 564
 Strong Electrolyte Solutions 564
 Applications of Colligative Properties 566
- 13.7 The Structure and Properties of Colloids** 568
- CHEMICAL CONNECTIONS TO ENVIRONMENTAL ENGINEERING: SOLUTIONS AND COLLOIDS IN WATER PURIFICATION** 570
- CHAPTER REVIEW GUIDE 572
 PROBLEMS 576

CHAPTER 14 *Periodic Patterns in the Main-Group Elements* 584

- 14.1 Hydrogen, the Simplest Atom** 585
 Where Hydrogen Fits in the Periodic Table 585
 Highlights of Hydrogen Chemistry 586
- 14.2 Trends Across the Periodic Table: The Period 2 Elements** 587
- 14.3 Group 1A(1): The Alkali Metals** 590
 Why the Alkali Metals Are Unusual Physically 590
 Why the Alkali Metals Are So Reactive 592
- 14.4 Group 2A(2)** 592
 How the Alkaline Earth and Alkali Metals Compare Physically 593
 How the Alkaline Earth and Alkali Metals Compare Chemically 593
 Diagonal Relationships: Lithium and Magnesium 595
- 14.5 Group 3A(13): The Boron Family** 595
 How the Transition Elements Influence This Group's Properties 595
 Features That First Appear in This Group's Chemical Properties 595
- Highlights of Boron Chemistry 597
 Diagonal Relationships: Beryllium and Aluminum 598
- 14.6 Group 4A(14): The Carbon Family** 598
 How Type of Bonding Affects Physical Properties 598
 How Bonding Changes in This Group's Compounds 601
 Highlights of Carbon Chemistry 601
 Highlights of Silicon Chemistry 603
 Diagonal Relationships: Boron and Silicon 604
- 14.7 Group 5A(15): The Nitrogen Family** 604
 The Wide Range of Physical Behavior 606
 Patterns in Chemical Behavior 606
 Highlights of Nitrogen Chemistry 607
 Highlights of Phosphorus Chemistry 610
- 14.8 Group 6A(16): The Oxygen Family** 612
 How the Oxygen and Nitrogen Families Compare Physically 612
 How the Oxygen and Nitrogen Families Compare Chemically 614
- Highlights of Oxygen Chemistry:
 Range of Oxide Properties 615
 Highlights of Sulfur Chemistry 615
- 14.9 Group 7A(17): The Halogens** 617
 Physical Behavior of the Halogens 617
 Why the Halogens Are So Reactive 617
 Highlights of Halogen Chemistry 619
- 14.10 Group 8A(18): The Noble Gases** 622
 How the Noble Gases and Alkali Metals Contrast Physically 622
 How Noble Gases Can Form Compounds 624
- CHAPTER REVIEW GUIDE 624
 PROBLEMS 625



© Miroslav Hlavko/Shutterstock.com

CHAPTER

15

Organic Compounds and the Atomic Properties of Carbon 632

15.1 The Special Nature of Carbon and the Characteristics of Organic Molecules 633

The Structural Complexity of Organic Molecules 634

The Chemical Diversity of Organic Molecules 634

15.2 The Structures and Classes of Hydrocarbons 636

Carbon Skeletons and Hydrogen Skins 636

Alkanes: Hydrocarbons with Only Single Bonds 639

Dispersion Forces and the Physical Properties of Alkanes 641

Constitutional Isomerism 641

Chiral Molecules and Optical Isomerism 642

Alkenes: Hydrocarbons with Double Bonds 644

Restricted Rotation and Geometric (*cis-trans*) Isomerism 645

Alkynes: Hydrocarbons with Triple Bonds 646

Aromatic Hydrocarbons: Cyclic Molecules with Delocalized π Electrons 647

Variations on a Theme: Catenated Inorganic Hydrides 648

TOOLS OF THE LABORATORY: NUCLEAR MAGNETIC RESONANCE (NMR) SPECTROSCOPY 649**15.3 Some Important Classes of Organic Reactions 651**

Types of Organic Reactions 651

The Redox Process in Organic Reactions 653

15.4 Properties and Reactivities of Common Functional Groups 654

Functional Groups with Only Single Bonds 654

Functional Groups with Double Bonds 659

Functional Groups with Both Single and Double Bonds 662

Functional Groups with Triple Bonds 666

15.5 The Monomer-Polymer Theme I: Synthetic Macromolecules 668

Addition Polymers 668

Condensation Polymers 669

15.6 The Monomer-Polymer Theme II: Biological Macromolecules 670

Sugars and Polysaccharides 670

Amino Acids and Proteins 672

Nucleotides and Nucleic Acids 674

CHEMICAL CONNECTIONS TO GENETICS AND FORENSICS: DNA SEQUENCING AND FINGERPRINTING 679

CHAPTER REVIEW GUIDE 681

PROBLEMS 683

CHAPTER

16

Kinetics: Rates and Mechanisms of Chemical Reactions 690

16.1 Focusing on Reaction Rate 691**16.2 Expressing the Reaction Rate 694**

Average, Instantaneous, and Initial Reaction Rates 694

Expressing Rate in Terms of Reactant and Product Concentrations 696

16.3 The Rate Law and Its Components 698

Some Laboratory Methods for Determining the Initial Rate 699

Determining Reaction Orders 699

Determining the Rate Constant 704

16.4 Integrated Rate Laws: Concentration Changes over Time 708

Integrated Rate Laws for First-, Second-, and Zero-Order Reactions 708

Determining Reaction Orders from an Integrated Rate Law 710

Reaction Half-Life 712

16.5 Theories of Chemical Kinetics 716

Collision Theory: Basis of the Rate Law 716

Transition State Theory: What the Activation Energy Is Used For 719

16.6 Reaction Mechanisms: The Steps from Reactant to Product 722

Elementary Reactions and Molecularity 722

The Rate-Determining Step of a Reaction Mechanism 724

Correlating the Mechanism with the Rate Law 725

16.7 Catalysis: Speeding Up a Reaction 729

The Basis of Catalytic Action 730

Homogeneous Catalysis 730

Heterogeneous Catalysis 731

Kinetics and Function of Biological Catalysts 732

CHEMICAL CONNECTIONS TO ATMOSPHERIC SCIENCE: DEPLETION OF EARTH'S OZONE LAYER 735

CHAPTER REVIEW GUIDE 736

PROBLEMS 740



© hxdbzxy/Shutterstock.com

CHAPTER

17

Equilibrium: The Extent of Chemical Reactions 746

- 17.1 The Equilibrium State and the Equilibrium Constant 747**
- 17.2 The Reaction Quotient and the Equilibrium Constant 750**
The Changing Value of the Reaction Quotient 750
Writing the Reaction Quotient in Its Various Forms 751
- 17.3 Expressing Equilibria with Pressure Terms: Relation Between K_c and K_p 756**
- 17.4 Comparing Q and K to Determine Reaction Direction 757**
- 17.5 How to Solve Equilibrium Problems 760**
Using Quantities to Find the Equilibrium Constant 760
Using the Equilibrium Constant to Find Quantities 763
Problems Involving Mixtures of Reactants and Products 768
- 17.6 Reaction Conditions and Equilibrium: Le Châtelier's Principle 770**
The Effect of a Change in Concentration 770
The Effect of a Change in Pressure (Volume) 773
- The Effect of a Change in Temperature 775
The Lack of Effect of a Catalyst 777
Applying Le Châtelier's Principle to the Synthesis of Ammonia 779
- CHEMICAL CONNECTIONS TO CELLULAR METABOLISM: DESIGN AND CONTROL OF A METABOLIC PATHWAY 781**
- CHAPTER REVIEW GUIDE 782
PROBLEMS 785

CHAPTER

18

Acid-Base Equilibria 792

- 18.1 Acids and Bases in Water 794**
Release of H^+ or OH^- and the Arrhenius Acid-Base Definition 794
Variation in Acid Strength: The Acid-Dissociation Constant (K_a) 795
Classifying the Relative Strengths of Acids and Bases 797
- 18.2 Autoionization of Water and the pH Scale 798**
The Equilibrium Nature of Autoionization: The Ion-Product Constant for Water (K_w) 799
Expressing the Hydronium Ion Concentration: The pH Scale 800
- 18.3 Proton Transfer and the Brønsted-Lowry Acid-Base Definition 803**
Conjugate Acid-Base Pairs 804
Relative Acid-Base Strength and the Net Direction of Reaction 805
- 18.4 Solving Problems Involving Weak-Acid Equilibria 808**
Finding K_a Given Concentrations 809
Finding Concentrations Given K_a 810
The Effect of Concentration on the Extent of Acid Dissociation 811
The Behavior of Polyprotic Acids 813
- 18.5 Molecular Properties and Acid Strength 816**
Acid Strength of Nonmetal Hydrides 816
Acid Strength of Oxoacids 816
Acidity of Hydrated Metal Ions 817
- 18.6 Weak Bases and Their Relation to Weak Acids 818**
Molecules as Weak Bases: Ammonia and the Amines 818
Anions of Weak Acids as Weak Bases 820
The Relation Between K_a and K_b of a Conjugate Acid-Base Pair 821
- 18.7 Acid-Base Properties of Salt Solutions 823**
Salts That Yield Neutral Solutions 823
Salts That Yield Acidic Solutions 823
Salts That Yield Basic Solutions 824
Salts of Weakly Acidic Cations and Weakly Basic Anions 824
Salts of Amphiprotic Anions 825
- 18.8 Generalizing the Brønsted-Lowry Concept: The Leveling Effect 827**
- 18.9 Electron-Pair Donation and the Lewis Acid-Base Definition 827**
Molecules as Lewis Acids 828
Metal Cations as Lewis Acids 829
An Overview of Acid-Base Definitions 830
- CHAPTER REVIEW GUIDE 831
PROBLEMS 834



© Joe Scherschel/Getty Images

CHAPTER

19

Ionic Equilibria in Aqueous Systems 842

19.1 Equilibria of Acid-Base Buffers 843

What a Buffer Is and How It Works: The Common-Ion Effect 843
 The Henderson-Hasselbalch Equation 848
 Buffer Capacity and Buffer Range 849
 Preparing a Buffer 851

19.2 Acid-Base Titration Curves 853

Strong Acid–Strong Base Titration Curves 853
 Weak Acid–Strong Base Titration Curves 855
 Weak Base–Strong Acid Titration Curves 859
 Monitoring pH with Acid-Base Indicators 860

Titration Curves for Polyprotic Acids 862
 Amino Acids as Biological Polyprotic Acids 863

19.3 Equilibria of Slightly Soluble Ionic Compounds 864

The Ion-Product Expression (Q_{sp}) and the Solubility-Product Constant (K_{sp}) 864
 Calculations Involving the Solubility-Product Constant 865
 Effect of a Common Ion on Solubility 868
 Effect of pH on Solubility 869
 Applying Ionic Equilibria to the Formation of a Limestone Cave 870
 Predicting the Formation of a Precipitate: Q_{sp} vs. K_{sp} 871

Separating Ions by Selective Precipitation and Simultaneous Equilibria 874

CHEMICAL CONNECTIONS TO ENVIRONMENTAL SCIENCE: THE ACID-RAIN PROBLEM 875**19.4 Equilibria Involving Complex Ions 877**

Formation of Complex Ions 877
 Complex Ions and the Solubility of Precipitates 879
 Complex Ions of Amphoteric Hydroxides 881

CHAPTER REVIEW GUIDE 883

PROBLEMS 887

CHAPTER

20

Thermodynamics: Entropy, Free Energy, and Reaction Direction 894

20.1 The Second Law of Thermodynamics: Predicting Spontaneous Change 895

The First Law of Thermodynamics Does Not Predict Spontaneous Change 896
 The Sign of ΔH Does Not Predict Spontaneous Change 896
 Freedom of Particle Motion and Dispersal of Kinetic Energy 897
 Entropy and the Number of Microstates 898
 Quantitative Meaning of an Entropy Change—Measuring Thermodynamic Variables 900
 Entropy and the Second Law of Thermodynamics 905

Standard Molar Entropies and the Third Law 905
 Predicting Relative S° of a System 906

20.2 Calculating the Change in Entropy of a Reaction 910

Entropy Changes in the System: Standard Entropy of Reaction (ΔS_{rxn}°) 910
 Entropy Changes in the Surroundings: The Other Part of the Total 912
 The Entropy Change and the Equilibrium State 914
 Spontaneous Exothermic and Endothermic Changes 915

20.3 Entropy, Free Energy, and Work 916

Free Energy Change and Reaction Spontaneity 916

Calculating Standard Free Energy Changes 917

The Free Energy Change and the Work a System Can Do 919
 The Effect of Temperature on Reaction Spontaneity 920
 Coupling of Reactions to Drive a Nonspontaneous Change 924

CHEMICAL CONNECTIONS TO BIOLOGICAL ENERGETICS: THE UNIVERSAL ROLE OF ATP 925**20.4 Free Energy, Equilibrium, and Reaction Direction 926**

CHAPTER REVIEW GUIDE 932

PROBLEMS 936



© Griffin Technology

CHAPTER

21

Electrochemistry: Chemical Change and Electrical Work 938

21.1 Redox Reactions and Electrochemical Cells 939

A Quick Review of Oxidation-Reduction Concepts 939

Half-Reaction Method for Balancing Redox Reactions 940

An Overview of Electrochemical Cells 944

21.2 Voltaic Cells: Using Spontaneous Reactions to Generate Electrical Energy 945

Construction and Operation of a Voltaic Cell 946

Notation for a Voltaic Cell 948

Why Does a Voltaic Cell Work? 949

21.3 Cell Potential: Output of a Voltaic Cell 950Standard Cell Potential (E_{cell}°) 950

Relative Strengths of Oxidizing and Reducing Agents 953

Using $E_{\text{half-cell}}^{\circ}$ Values to Write Spontaneous Redox Reactions 954
Explaining the Activity Series of the Metals 958**21.4 Free Energy and Electrical Work 959**

Standard Cell Potential and the

Equilibrium Constant 959

The Effect of Concentration on Cell Potential 961

Following Changes in Potential During Cell Operation 963

Concentration Cells 964

21.5 Electrochemical Processes in Batteries 968

Primary (Nonrechargeable) Batteries 968

Secondary (Rechargeable) Batteries 969

Fuel Cells 970

21.6 Corrosion: An Environmental Voltaic Cell 972

The Corrosion of Iron 972

Protecting Against the Corrosion of Iron 973

21.7 Electrolytic Cells: Using Electrical Energy to Drive Nonspontaneous Reactions 974

Construction and Operation of an Electrolytic Cell 974

Predicting the Products of Electrolysis 976

Stoichiometry of Electrolysis: The Relation Between Amounts of Charge and Products 980

CHEMICAL CONNECTIONS TO BIOLOGICAL ENERGETICS: CELLULAR ELECTROCHEMISTRY AND THE PRODUCTION OF ATP 982

CHAPTER REVIEW GUIDE 984

PROBLEMS 987

CHAPTER

22

The Elements in Nature and Industry 996

22.1 How the Elements Occur in Nature 997

Earth's Structure and the Abundance of the Elements 997

Sources of the Elements 1000

22.2 The Cycling of Elements Through the Environment 1002

The Carbon Cycle 1002

The Nitrogen Cycle 1004

The Phosphorus Cycle 1005

22.3 Metallurgy: Extracting a Metal from Its Ore 1008

Pretreating the Ore 1009

Converting Mineral to Element 1010

Refining and Alloying the Element 1012

22.4 Tapping the Crust: Isolation and Uses of Selected Elements 1014

Producing the Alkali Metals: Sodium and Potassium 1014

The Indispensable Three: Iron, Copper, and Aluminum 1015

Mining the Sea for Magnesium 1021

The Sources and Uses of

Hydrogen 1022

22.5 Chemical Manufacturing: Two Case Studies 1025

Sulfuric Acid, the Most Important Chemical 1025

The Chlor-Alkali Process 1028

CHAPTER REVIEW GUIDE 1029

PROBLEMS 1030



© PjrStudio/Alamy

CHAPTER 23 Transition Elements and Their Coordination Compounds 1036

23.1 Properties of the Transition Elements 1037

Electron Configurations of the Transition Metals and Their Ions 1038
 Atomic and Physical Properties of the Transition Elements 1040
 Chemical Properties of the Transition Elements 1042

23.2 The Inner Transition Elements 1044

The Lanthanides 1044
 The Actinides 1045

23.3 Coordination Compounds 1046

Complex Ions: Coordination Numbers, Geometries, and Ligands 1046
 Formulas and Names of Coordination Compounds 1048
 Isomerism in Coordination Compounds 1051

23.4 Theoretical Basis for the Bonding and Properties of Complex Ions 1055

Applying Valence Bond Theory to Complex Ions 1055
 Crystal Field Theory 1056

CHEMICAL CONNECTIONS TO NUTRITIONAL SCIENCE: TRANSITION METALS AS ESSENTIAL DIETARY TRACE ELEMENTS 1063

CHAPTER REVIEW GUIDE 1065
 PROBLEMS 1067

CHAPTER 24 Nuclear Reactions and Their Applications 1072

24.1 Radioactive Decay and Nuclear Stability 1073

Comparing Chemical and Nuclear Change 1074
 The Components of the Nucleus: Terms and Notation 1074
 The Discovery of Radioactivity and the Types of Emissions 1075
 Modes of Radioactive Decay; Balancing Nuclear Equations 1075
 Nuclear Stability and the Mode of Decay 1079

24.2 The Kinetics of Radioactive Decay 1083

Detection and Measurement of Radioactivity 1083
 The Rate of Radioactive Decay 1084
 Radioisotopic Dating 1087

24.3 Nuclear Transmutation: Induced Changes in Nuclei 1090

Early Transmutation Experiments; Nuclear Shorthand Notation 1090
 Particle Accelerators and the Transuranium Elements 1091

24.4 Ionization: Effects of Nuclear Radiation on Matter 1093

Effects of Ionizing Radiation on Living Tissue 1093
 Background Sources of Ionizing Radiation 1095
 Assessing the Risk from Ionizing Radiation 1096

24.5 Applications of Radioisotopes 1098

Radioactive Tracers 1098
 Additional Applications of Ionizing Radiation 1100

24.6 The Interconversion of Mass and Energy 1101

The Mass Difference Between a Nucleus and Its Nucleons 1101
 Nuclear Binding Energy and Binding Energy per Nucleon 1102

24.7 Applications of Fission and Fusion 1104

The Process of Nuclear Fission 1105
 The Promise of Nuclear Fusion 1109

CHEMICAL CONNECTIONS TO COSMOLOGY: ORIGIN OF THE ELEMENTS IN THE STARS 1110

CHAPTER REVIEW GUIDE 1112
 PROBLEMS 1114

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
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 22
- 1.6 Converting Units Raised to a Power 23
- 1.7 Calculating Density from Mass and Volume 24
- 1.8 Converting Units of Temperature 27
- 1.9 Determining the Number of Significant Figures 29
- 1.10 Significant Figures and Rounding 32

Chapter 2

- 2.1 Distinguishing Elements, Compounds, and Mixtures at the Atomic Scale 45
- 2.2 Calculating the Mass of an Element in a Compound 48
- 2.3 Visualizing the Mass Laws 51
- 2.4 Determining the Numbers of Subatomic Particles in the Isotopes of an Element 57
- 2.5 Calculating the Atomic Mass of an Element 58
- 2.6 Identifying an Element from Its Z Value 62
- 2.7 Predicting the Ion an Element Forms 66
- 2.8 Naming Binary Ionic Compounds 69
- 2.9 Determining Formulas of Binary Ionic Compounds 70
- 2.10 Determining Names and Formulas of Ionic Compounds of Metals That Form More Than One Ion 71
- 2.11 Determining Names and Formulas of Ionic Compounds Containing Polyatomic Ions (Including Hydrates) 73
- 2.12 Recognizing Incorrect Names and Formulas of Ionic Compounds 73
- 2.13 Determining Names and Formulas of Anions and Acids 74
- 2.14 Determining Names and Formulas of Binary Covalent Compounds 75
- 2.15 Recognizing Incorrect Names and Formulas of Binary Covalent Compounds 75
- 2.16 Calculating the Molecular Mass of a Compound 77
- 2.17 Using Molecular Depictions to Determine Formula, Name, and Mass 77

Chapter 3

- 3.1 Converting Between Mass and Amount of an Element 98
- 3.2 Converting Between Number of Entities and Amount of an Element 99
- 3.3 Converting Between Number of Entities and Mass of an Element 99
- 3.4 Converting Between Number of Entities and Mass of a Compound I 100
- 3.5 Converting Between Number of Entities and Mass of a Compound II 101
- 3.6 Calculating the Mass Percent of Each Element in a Compound from the Formula 102
- 3.7 Calculating the Mass of an Element in a Compound 104
- 3.8 Determining an Empirical Formula from Amounts of Elements 105
- 3.9 Determining an Empirical Formula from Masses of Elements 106
- 3.10 Determining a Molecular Formula from Elemental Analysis and Molar Mass 107
- 3.11 Determining a Molecular Formula from Combustion Analysis 108
- 3.12 Balancing a Chemical Equation 114

- 3.13 Writing a Balanced Equation from a Molecular Scene 115
- 3.14 Calculating Quantities of Reactants and Products: Amount (mol) to Amount (mol) 118
- 3.15 Calculating Quantities of Reactants and Products: Amount (mol) to Mass (g) 119
- 3.16 Calculating Quantities of Reactants and Products: Mass to Mass 120
- 3.17 Writing an Overall Equation for a Reaction Sequence 121
- 3.18 Using Molecular Depictions in a Limiting-Reactant Problem 123
- 3.19 Calculating Quantities in a Limiting-Reactant Problem: Amount to Amount 125
- 3.20 Calculating Quantities in a Limiting-Reactant Problem: Mass to Mass 125
- 3.21 Calculating Percent Yield 128

Chapter 4

- 4.1 Using Molecular Scenes to Depict an Ionic Compound in Aqueous Solution 148
- 4.2 Determining Amount (mol) of Ions in Solution 149
- 4.3 Calculating the Molarity of a Solution 150
- 4.4 Calculating Mass of Solute in a Given Volume of Solution 151
- 4.5 Determining Amount (mol) of Ions in a Solution 151
- 4.6 Preparing a Dilute Solution from a Concentrated Solution 153
- 4.7 Visualizing Changes in Concentration 154
- 4.8 Predicting Whether a Precipitation Reaction Occurs; Writing Ionic Equations 159
- 4.9 Using Molecular Depictions in Precipitation Reactions 160
- 4.10 Calculating Amounts of Reactants and Products in a Precipitation Reaction 162
- 4.11 Solving a Limiting-Reactant Problem for a Precipitation Reaction 163
- 4.12 Determining the Number of H^+ (or OH^-) Ions in Solution 166
- 4.13 Writing Ionic Equations for Acid-Base Reactions 167
- 4.14 Writing Proton-Transfer Equations for Acid-Base Reactions 171
- 4.15 Calculating the Amounts of Reactants and Products in an Acid-Base Reaction 172
- 4.16 Finding the Concentration of an Acid from a Titration 173
- 4.17 Determining the Oxidation Number of Each Element in a Compound (or Ion) 177
- 4.18 Identifying Redox Reactions and Oxidizing and Reducing Agents 178
- 4.19 Finding the Amount of Reducing Agent by Titration 180
- 4.20 Identifying the Type of Redox Reaction 187

Chapter 5

- 5.1 Converting Units of Pressure 210
- 5.2 Applying the Volume-Pressure Relationship 217
- 5.3 Applying the Volume-Temperature and Pressure-Temperature Relationships 218
- 5.4 Applying the Volume-Amount and Pressure-Amount Relationships 218
- 5.5 Applying the Volume-Pressure-Temperature Relationship 219
- 5.6 Solving for an Unknown Gas Variable at Fixed Conditions 220
- 5.7 Using Gas Laws to Determine a Balanced Equation 221
- 5.8 Calculating Gas Density 223
- 5.9 Finding the Molar Mass of a Volatile Liquid 225
- 5.10 Applying Dalton's Law of Partial Pressures 226
- 5.11 Calculating the Amount of Gas Collected over Water 228

- 5.12 Using Gas Variables to Find Amounts of Reactants or Products I 229
- 5.13 Using Gas Variables to Find Amounts of Reactants or Products II 230
- 5.14 Applying Graham's Law of Effusion 236

Chapter 6

- 6.1 Determining the Change in Internal Energy of a System 262
- 6.2 Calculating Pressure-Volume Work Done by or on a System 264
- 6.3 Drawing Enthalpy Diagrams and Determining the Sign of ΔH 267
- 6.4 Relating Quantity of Heat and Temperature Change 269
- 6.5 Determining the Specific Heat Capacity of a Solid 270
- 6.6 Determining the Enthalpy Change of an Aqueous Reaction 270
- 6.7 Calculating the Heat of a Combustion Reaction 272
- 6.8 Using the Enthalpy Change of a Reaction (ΔH) to Find the Amount of a Substance 274
- 6.9 Using Hess's Law to Calculate an Unknown ΔH 276
- 6.10 Writing Formation Equations 278
- 6.11 Calculating $\Delta H_{\text{rxn}}^{\circ}$ from ΔH_f° Values 280

Chapter 7

- 7.1 Interconverting Wavelength and Frequency 297
- 7.2 Interconverting Energy, Wavelength, and Frequency 301
- 7.3 Determining ΔE and λ of an Electron Transition 307
- 7.4 Calculating the de Broglie Wavelength of an Electron 311
- 7.5 Applying the Uncertainty Principle 313
- 7.6 Determining ΔE and λ of an Electron Transition Using the Particle-in-a-Box Model 316
- 7.7 Determining Quantum Numbers for an Energy Level 320
- 7.8 Determining Sublevel Names and Orbital Quantum Numbers 321
- 7.9 Identifying Incorrect Quantum Numbers 322

Chapter 8

- 8.1 Correlating Quantum Numbers and Orbital Diagrams 337
- 8.2 Determining Electron Configurations 344
- 8.3 Ranking Elements by Atomic Size 347
- 8.4 Ranking Elements by First Ionization Energy 350
- 8.5 Identifying an Element from Its Ionization Energies 351
- 8.6 Writing Electron Configurations of Main-Group Ions 356
- 8.7 Writing Electron Configurations and Predicting Magnetic Behavior of Transition Metal Ions 358
- 8.8 Ranking Ions by Size 360

Chapter 9

- 9.1 Depicting Ion Formation 372
- 9.2 Predicting Relative Lattice Energy from Ionic Properties 376
- 9.3 Comparing Bond Length and Bond Strength 382
- 9.4 Using Bond Energies to Calculate $\Delta H_{\text{rxn}}^{\circ}$ 389
- 9.5 Determining Bond Polarity from EN Values 393

Chapter 10

- 10.1 Writing Lewis Structures for Species with Single Bonds and One Central Atom 407
- 10.2 Writing Lewis Structures for Molecules with Single Bonds and More Than One Central Atom 408
- 10.3 Writing Lewis Structures for Molecules with Multiple Bonds 409
- 10.4 Writing Resonance Structures and Assigning Formal Charges 412
- 10.5 Writing Lewis Structures for Octet-Rule Exceptions 416
- 10.6 Examining Shapes with Two, Three, or Four Electron Groups 425

- 10.7 Examining Shapes with Five or Six Electron Groups 426
- 10.8 Predicting Molecular Shapes with More Than One Central Atom 427
- 10.9 Predicting the Polarity of Molecules 429

Chapter 11

- 11.1 Postulating Hybrid Orbitals in a Molecule 449
- 11.2 Describing the Types of Orbitals and Bonds in Molecules 454
- 11.3 Predicting Stability of Species Using MO Diagrams 457
- 11.4 Using MO Theory to Explain Bond Properties 461

Chapter 12

- 12.1 Finding the Heat of a Phase Change Depicted by Molecular Scenes 477
- 12.2 Applying the Clausius-Clapeyron Equation 480
- 12.3 Using a Phase Diagram to Predict Phase Changes 483
- 12.4 Drawing Hydrogen Bonds Between Molecules of a Substance 487
- 12.5 Identifying the Types of Intermolecular Forces 489
- 12.6 Determining the Number of Particles per Unit Cell and the Coordination Number 497
- 12.7 Determining Atomic Radius 500
- 12.8 Determining Atomic Radius from the Unit Cell 501

Chapter 13

- 13.1 Predicting Relative Solubilities 537
- 13.2 Calculating an Aqueous Ionic Heat of Solution 546
- 13.3 Using Henry's Law to Calculate Gas Solubility 552
- 13.4 Calculating Molality 553
- 13.5 Expressing Concentrations in Parts by Mass, Parts by Volume, and Mole Fraction 555
- 13.6 Interconverting Concentration Terms 556
- 13.7 Using Raoult's Law to Find ΔP 559
- 13.8 Determining Boiling and Freezing Points of a Solution 561
- 13.9 Determining Molar Mass from Colligative Properties 563
- 13.10 Depicting Strong Electrolyte Solutions 565

Chapter 15

- 15.1 Drawing Hydrocarbons 637
- 15.2 Naming Hydrocarbons and Understanding Chirality and Geometric Isomerism 646
- 15.3 Recognizing the Type of Organic Reaction 652
- 15.4 Predicting the Reactions of Alcohols, Alkyl Halides, and Amines 658
- 15.5 Predicting the Steps in a Reaction Sequence 661
- 15.6 Predicting Reactions of the Carboxylic Acid Family 665
- 15.7 Recognizing Functional Groups 667

Chapter 16

- 16.1 Expressing Rate in Terms of Changes in Concentration with Time 697
- 16.2 Determining Reaction Orders from Rate Laws 701
- 16.3 Determining Reaction Orders and Rate Constants from Rate Data 705
- 16.4 Determining Reaction Orders from Molecular Scenes 706
- 16.5 Determining the Reactant Concentration After a Given Time 709
- 16.6 Using Molecular Scenes to Find Quantities at Various Times 713
- 16.7 Determining the Half-Life of a First-Order Reaction 714
- 16.8 Determining the Energy of Activation 718
- 16.9 Drawing Reaction Energy Diagrams and Transition States 721
- 16.10 Determining Molarities and Rate Laws for Elementary Steps 723
- 16.11 Identifying Intermediates and Correlating Rate Laws and Reaction Mechanisms 726

Chapter 17

- 17.1 Writing the Reaction Quotient from the Balanced Equation 752
- 17.2 Finding K for Reactions Multiplied by a Common Factor or Reversed and for an Overall Reaction 754
- 17.3 Converting Between K_c and K_p 757
- 17.4 Using Molecular Scenes to Determine Reaction Direction 758
- 17.5 Using Concentrations to Determine Reaction Direction 759
- 17.6 Calculating K_c from Concentration Data 762
- 17.7 Determining Equilibrium Concentrations from K_c 763
- 17.8 Determining Equilibrium Concentrations from Initial Concentrations and K_c 763
- 17.9 Making a Simplifying Assumption to Calculate Equilibrium Concentrations 766
- 17.10 Predicting Reaction Direction and Calculating Equilibrium Concentrations 768
- 17.11 Predicting the Effect of a Change in Concentration on the Equilibrium Position 772
- 17.12 Predicting the Effect of a Change in Volume (Pressure) on the Equilibrium Position 774
- 17.13 Predicting the Effect of a Change in Temperature on the Equilibrium Position 776
- 17.14 Determining Equilibrium Parameters from Molecular Scenes 778

Chapter 18

- 18.1 Classifying Acid and Base Strength from the Chemical Formula 798
- 18.2 Calculating $[\text{H}_3\text{O}^+]$ or $[\text{OH}^-]$ in Aqueous Solution 800
- 18.3 Calculating $[\text{H}_3\text{O}^+]$, pH, $[\text{OH}^-]$, and pOH for Strong Acids and Bases 802
- 18.4 Identifying Conjugate Acid-Base Pairs 805
- 18.5 Predicting the Net Direction of an Acid-Base Reaction 807
- 18.6 Using Molecular Scenes to Predict the Net Direction of an Acid-Base Reaction 807
- 18.7 Finding K_a of a Weak Acid from the Solution pH 809
- 18.8 Determining Concentration and pH from K_a and Initial $[\text{HA}]$ 810
- 18.9 Finding the Percent Dissociation of a Weak Acid 812
- 18.10 Calculating Equilibrium Concentrations for a Polyprotic Acid 814
- 18.11 Determining pH from K_b and Initial $[\text{B}]$ 819
- 18.12 Determining the pH of a Solution of A^- 822
- 18.13 Predicting Relative Acidity of Salt Solutions from Reactions of the Ions with Water 824
- 18.14 Predicting the Relative Acidity of a Salt Solution from K_a and K_b of the Ions 826
- 18.15 Identifying Lewis Acids and Bases 830

Chapter 19

- 19.1 Calculating the Effect of Added H_3O^+ or OH^- on Buffer pH 846
- 19.2 Using Molecular Scenes to Examine Buffers 850
- 19.3 Preparing a Buffer 851
- 19.4 Finding the pH During a Weak Acid–Strong Base Titration 857
- 19.5 Writing Ion-Product Expressions 865
- 19.6 Determining K_{sp} from Solubility 866
- 19.7 Determining Solubility from K_{sp} 867
- 19.8 Calculating the Effect of a Common Ion on Solubility 869
- 19.9 Predicting the Effect on Solubility of Adding Strong Acid 870
- 19.10 Predicting Whether a Precipitate Will Form 871
- 19.11 Using Molecular Scenes to Predict Whether a Precipitate Will Form 872

- 19.12 Separating Ions by Selective Precipitation 874
- 19.13 Calculating the Concentration of a Complex Ion 878
- 19.14 Calculating the Effect of Complex-Ion Formation on Solubility 880

Chapter 20

- 20.1 Calculating the Change in Entropy During an Isothermal Volume Change of an Ideal Gas 902
- 20.2 Calculating the Change in Entropy During a Phase Change 903
- 20.3 Calculating the Entropy Change Resulting from a Change in Temperature 904
- 20.4 Predicting Relative Entropy Values 909
- 20.5 Calculating the Standard Entropy of Reaction, $\Delta S_{\text{rxn}}^\circ$ 911
- 20.6 Determining Reaction Spontaneity 913
- 20.7 Calculating $\Delta G_{\text{rxn}}^\circ$ from Enthalpy and Entropy Values 917
- 20.8 Calculating $\Delta G_{\text{rxn}}^\circ$ from ΔG_f° Values 919
- 20.9 Using Molecular Scenes to Determine the Signs of ΔH , ΔS , and ΔG 921
- 20.10 Determining the Effect of Temperature on ΔG 922
- 20.11 Finding the Temperature at Which a Reaction Becomes Spontaneous 923
- 20.12 Exploring the Relationship Between ΔG° and K 927
- 20.13 Using Molecular Scenes to Find ΔG for a Reaction at Nonstandard Conditions 928
- 20.14 Calculating ΔG at Nonstandard Conditions 930

Chapter 21

- 21.1 Balancing a Redox Reaction in Basic Solution 942
- 21.2 Describing a Voltaic Cell with a Diagram and Notation 948
- 21.3 Using $E_{\text{half-cell}}^\circ$ Values to Find E_{cell}° 951
- 21.4 Calculating an Unknown $E_{\text{half-cell}}^\circ$ from E_{cell}° 953
- 21.5 Writing Spontaneous Redox Reactions and Ranking Oxidizing and Reducing Agents by Strength 956
- 21.6 Calculating K and ΔG° from E_{cell}° 961
- 21.7 Using the Nernst Equation to Calculate E_{cell} 962
- 21.8 Calculating the Potential of a Concentration Cell 966
- 21.9 Predicting the Electrolysis Products of a Molten Salt Mixture 977
- 21.10 Predicting the Electrolysis Products of Aqueous Salt Solutions 979
- 21.11 Applying the Relationship Among Current, Time, and Amount of Substance 981

Chapter 23

- 23.1 Writing Electron Configurations of Transition Metal Atoms and Ions 1040
- 23.2 Finding the Number of Unpaired Electrons 1045
- 23.3 Finding the Coordination Number and Charge of the Central Metal Ion in a Coordination Compound 1049
- 23.4 Writing Names and Formulas of Coordination Compounds 1050
- 23.5 Determining the Type of Stereoisomerism 1054
- 23.6 Ranking Crystal Field Splitting Energies (Δ) for Complex Ions of a Metal 1059
- 23.7 Identifying High-Spin and Low-Spin Complex Ions 1061

Chapter 24

- 24.1 Writing Equations for Nuclear Reactions 1078
- 24.2 Predicting Nuclear Stability 1080
- 24.3 Predicting the Mode of Nuclear Decay 1082
- 24.4 Calculating the Specific Activity and the Decay Constant of a Radioactive Nuclide 1085
- 24.5 Finding the Number of Radioactive Nuclei 1086
- 24.6 Applying Radiocarbon Dating 1089
- 24.7 Calculating the Binding Energy per Nucleon 1103

ABOUT THE AUTHORS



Courtesy of Martin S. Silberberg

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 traveling, gardening, and singing.



Courtesy of Patricia G. Amateis

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 31 years, teaching General Chemistry and Analytical Chemistry. For the past 16 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

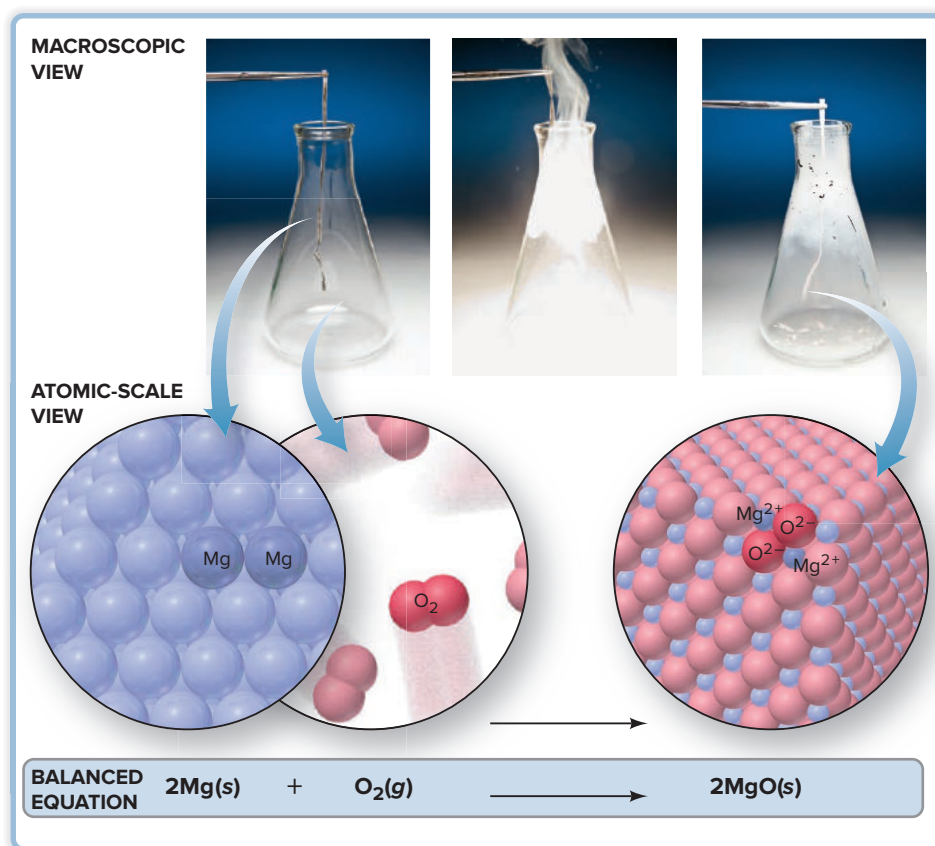
Chemistry is so crucial to an understanding of medicine and biology, environmental science, and many areas of engineering and industrial processing that it has become a requirement for an increasing number of academic majors. Furthermore, chemical principles lie at the core of some of the key societal issues we face in the 21st 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 eighth 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.



(three photos): © McGraw-Hill Education/Charles Winters/Timeframe Photography, Inc.

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 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:

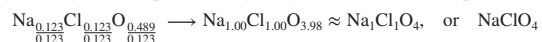
$$\text{Amount (mol) of Na} = 2.82 \text{ g Na} \times \frac{1 \text{ mol Na}}{22.99 \text{ g Na}} = 0.123 \text{ mol Na}$$

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

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

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 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

of an unknown compound is found to contain 1.23 g of H, 12.64 g of O. What is the empirical formula and the name of the compound?

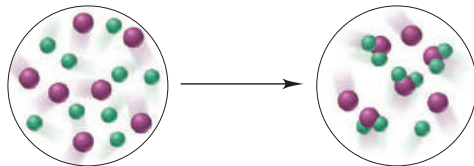
When metal M reacts with sulfur to form a compound with the formula M_2S_3 , what are the names of M and M_2S_3 ? [*Hint*: amount (mol) of S, and use the formula to find the amount (mol) of M.]

PROBLEMS 3.42(b), 3.43(b), 3.46, and 3.47

SAMPLE PROBLEM 2.3

Visualizing the Mass Laws

Problem The scenes below represent an atomic-scale view of a chemical reaction:



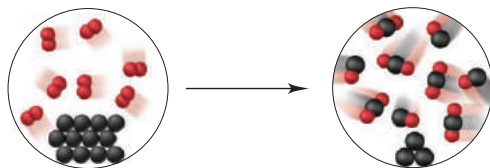
Which of the mass laws—mass conservation, definite composition, and/or multiple proportions—is (are) illustrated?

Plan From the depictions, we note the numbers, colors, and combinations of atoms (spheres) to see which mass laws pertain. If the numbers of each atom are the same before and after the reaction, the total mass did not change (mass conservation). If a compound forms that always has the same atom ratio, the elements are present in fixed parts by mass (definite composition). If the same elements form different compounds and the ratio of the atoms of one element that combine with one atom of the other element is a small whole number, the ratio of their masses is a small whole number as well (multiple proportions).

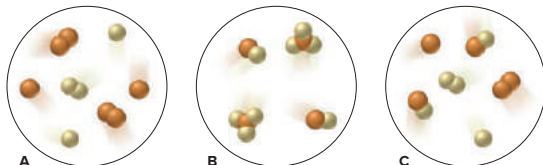
Solution There are seven purple and nine green atoms in each circle, so mass is conserved. The compound formed has one purple and two green atoms, so it has definite composition. Only one compound forms, so the law of multiple proportions does not pertain.

FOLLOW-UP PROBLEMS

2.3A The following scenes represent a chemical change. Which of the mass laws is (are) illustrated?



2.3B Which sample(s) best display(s) the fact that compounds of bromine (*orange*) and fluorine (*yellow*) exhibit the law of multiple proportions? Explain.



SOME SIMILAR PROBLEMS 2.14 and 2.15

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

Treating water involves several steps (Figure B13.1):

- 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; $Al_2(SO_4)_3$] or iron(III) chloride ($FeCl_3$), which supply Al^{3+} or Fe^{3+} ions 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) have pore sizes of 0.1–10 μm are thin tubes bundled together inside a vessel. The water is forced into these tubes, and the colloid is retained. Filtration is resistant to disinfectants.

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

- Chlorine, as aqueous bleach (ClO^-) or Cl_2 , is most common, but carcinogenic chlorinated organic compounds can form.
- UV light emitted by high-intensity fluorescent tubes disinfects by disrupting microorganisms' DNA.
- Ozone (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 Ca^{2+} and Mg^{2+} , is called **hard water**. Combined with fatty-acid anions in soap, these cations form solid deposits on clothes, washing machines, and sinks:

$$Ca^{2+}(aq) + 2C_{17}H_{35}COONa(aq) \xrightarrow{\text{soap}} (C_{17}H_{35}COO)_2Ca(s) + 2Na^{+}(aq)$$

insoluble deposit

When a large amount of HCO_3^- is present, the cations form scale, a carbonate deposit in boilers and hot-water pipes that interferes with the transfer of heat:

$$Ca^{2+}(aq) + 2HCO_3^-(aq) \xrightarrow{\text{heat}} CaCO_3(s) + CO_2(g) + H_2O(l)$$

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:
 - Nitrification.** Certain bacteria oxidize ammonia (electron donor) with O_2 (electron acceptor) to form nitrate ion:

$$NH_4^+ + 2O_2 \rightarrow NO_3^- + 2H^+ + H_2O$$
 - Denitrification.** Other bacteria oxidize an added compound like methanol (CH_3OH) using the NO_3^- :

$$5CH_3OH + 6NO_3^- \rightarrow 3N_2 + 5CO_2 + 7H_2O + 6OH^-$$
 Thus, the process converts NH_3 in wastewater to N_2 , which is released to the atmosphere.

Membrane Processes and Reverse Osmosis

Membranes with 0.0001–0.01 μm 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 Pb^{2+} , Cd^{2+} , and 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).

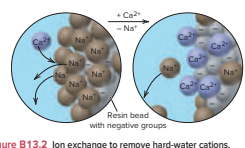


Figure B13.2 Ion exchange to remove hard-water cations.

anionic groups, such as $-SO_3^-$ or $-COO^-$, and Na^+ ions for charge balance (Figure B13.2). The hard-water cations displace the Na^+ ions and bind to the anionic groups. When all resin sites are occupied, the resin is regenerated with concentrated Na^+ solution that exchanges Na^+ ions for bound Ca^{2+} and Mg^{2+} .

Reverse osmosis is used for desalination, which can convert seawater (40,000 ppm of ions) to drinking water (400 ppm) (Figure B13.3).

Problems

B13.1 Briefly answer each of the following:

- Why is cake alum $[Al_2(SO_4)_3]$ added during water purification?
- Why is water that contains large amounts of Ca^{2+} and Mg^{2+} difficult to use for cleaning?
- What is the meaning of "reverse" in reverse osmosis?
- Why might a water treatment plant use ozone as a disinfectant instead of chlorine?
- How does passing a saturated NaCl solution through a "spen" ion-exchange resin regenerate the resin?

B13.2 Wastewater discharged into a stream by a sugar refinery contains 3.55 g of sucrose ($C_{12}H_{22}O_{11}$) per liter. A government-sponsored study is testing the feasibility of removing the sugar by reverse osmosis. What pressure must be applied to the wastewater solution at 20. °C to produce pure water?

TOOLS OF THE LABORATORY

Nuclear Magnetic Resonance (NMR) Spectroscopy

Figure B13.1

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}C , ^{19}F , 1H , and 31P , 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 1H -NMR spectroscopy, which measures changes in the nuclei of the most common isotope of hydrogen. Oriented randomly, the magnetic fields of all the 1H 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 (ΔE) between the two energy states (spin states) lies in the radio-frequency (rf) region of the electromagnetic spectrum (Figure B15.1).

When an 1H (blue arrow) in the lower energy (parallel) spin state absorbs a photon in the radio-frequency region with an energy equal to ΔE , 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 1H -NMR spectrometer. The ΔE between the two states depends on the actual magnetic field acting on each 1H nucleus, which is affected by the tiny magnetic fields of the electrons of atoms adjacent to that nucleus. Thus, the ΔE required for resonance of each 1H nucleus depends on its specific molecular environment—the C atoms, electronegative atoms, multiple bonds, and aromatic rings around it. 1H nuclei in different molecular environments produce different peaks in the 1H -NMR spectrum.

An 1H -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 1H nuclei in a given environment is where a peak appears. Chemical shifts are shown relative to that of an added standard, tetramethylsilane [(CH_3) $_4$ Si, or TMS]. TMS has 12 1H 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 1H -NMR spectrum of acetone. The six 1H nuclei of acetone have identical environments: all six are bonded to two C atoms that are each bonded to the C atom involved in the 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 1H nuclei have two different environments. The taller peak is due to the six 1H nuclei in the two CH_3 groups, and the shorter peak is due to the two 1H nuclei in the CH_2 group. The area under each peak (given as a number of chart-paper grid spaces) is proportional to the number of 1H 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 CH_3 groups to two in the 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)

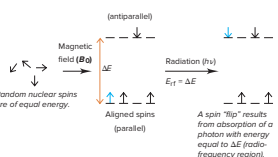


Figure B15.1 The basis of 1H spin resonance.

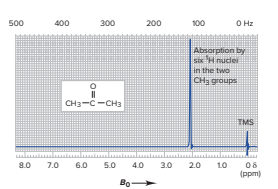


Figure B15.2 The 1H -NMR spectrum of acetone.

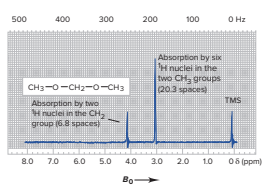


Figure B15.3 The 1H -NMR spectrum of dimethoxymethane.

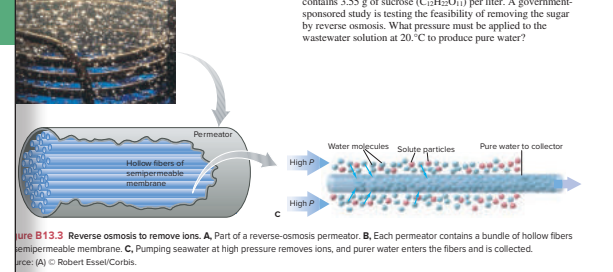


Figure B13.3 Reverse osmosis to remove ions. **A**, Part of a reverse-osmosis permeator. **B**, Each permeator contains a bundle of hollow fibers of semipermeable membrane. **C**, Pumping seawater at high pressure removes ions, and pure water enters the fibers and is collected. (A) © Robert Essel/Corbis.

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

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.

CHAPTER REVIEW GUIDE

Learning Objectives

Relevant section (S) 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 ΔH_{sol} (§13.3)
- The dependence of ΔH_{sol} 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)
- Ideal solutions and the importance of Raoult's law (§13.6)
- How the phase diagram of a solution differs from that of the pure solvent (§13.6)
- Why the vapor over a solution of a volatile nonelectrolyte is richer in the more volatile component (§13.6)
- Why strong electrolyte solutions are not ideal and the meanings of the van't Hoff factor and ionic atmosphere (§13.6)
- How particle size distinguishes suspensions, colloids, and solutions (§13.7)
- How colloidal behavior is demonstrated by the Tyndall effect and Brownian motion (§13.7)

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 and freezing points 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 depiction to determine colligative properties (SP 13.10)

Key Terms

Page numbers appear in parentheses.

alloy (538)	freezing point depression (ΔT_f) (561)	mass percent [% (w/w)] (554)	soap (541)
amino acid (539)	hard water (570)	miscible (534)	solubility (S) (534)
boiling point elevation (ΔT_b) (559)	heat of hydration (ΔH_{hyd}) (545)	molality (<i>m</i>) (553)	solute (534)
charge density (545)	heat of solution (ΔH_{sol}) (544)	mole fraction (<i>X</i>) (554)	solution (545)
colligative property (557)	Henry's law (551)	mononucleotide (543)	solvent (534)
colloid (568)	hydration (545)	nonelectrolyte (557)	supersaturated solution (549)
desalination (571)	hydration shell (534)	nucleic acid (542)	suspension (568)
dipole-induced dipole force (535)	ideal solution (558)	osmosis (562)	Tyndall effect (569)
double helix (543)	ion exchange (570)	osmotic pressure (Π) (562)	unsaturated solution (549)
electrolyte (557)	ionic atmosphere (565)	Raoult's law (558)	vapor pressure lowering (ΔP) (558)
entropy (S) (547)	ion-induced dipole force (534)	reverse osmosis (571)	volume percent [% (v/v)] (554)
fractional distillation (564)	like-dissolves-like rule (534)	saturated solution (549)	wastewater (571)
	lipid bilayer (542)	semipermeable membrane (562)	water softening (570)

Key Equations and Relationships

Page numbers appear in parentheses.

- Dividing the general heat of solution into component enthalpies (544):

$$\Delta H_{\text{sol}} = \Delta H_{\text{solvent}} + \Delta H_{\text{solute}} + \Delta H_{\text{mix}}$$
- Dividing the heat of solution of an ionic compound in water into component enthalpies (545):

$$\Delta H_{\text{sol}} = \Delta H_{\text{mice}} + \Delta H_{\text{hyd}}$$

- Relating gas solubility to its partial pressure (Henry's law) (551):

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

- Defining concentration in terms of molarity (552):

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

- Defining concentration in terms of molality (553):

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

- Defining concentration in terms of mass percent (554):

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

- Defining concentration in terms of volume percent (554):

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

- Defining concentration in terms of mole fraction (554):

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

- Expressing the relationship between the vapor pressure of solvent above a solution and its mole fraction in the solution (Raoult's law) (558):

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

- Calculating the vapor pressure lowering due to solute (558):

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

- Calculating the boiling point elevation of a solution (560):

$$\Delta T_b = K_b m$$

- Calculating the freezing point depression of a solution (561):

$$\Delta T_f = K_f m$$

- Calculating the osmotic pressure of a solution (562):

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

BRIEF SOLUTIONS TO FOLLOW-UP PROBLEMS

- (a) 1-Butanol has one —OH group/molecule, while 1,4-butanediol has two —OH groups/molecule. 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 between the polar CHCl_3 molecules and water. The forces between nonpolar CCl_4 molecules and water are weaker dipole-induced dipole forces, which do not effectively replace H bonds between water molecules.

- (a) Chloroform dissolves more chloromethane due to similar dipole-dipole forces between the polar molecules of these two substances. CH_2Cl_2 molecules do not exhibit H bonding and so do not effectively replace H bonds between methanol molecules. (b) Hexane dissolves more pentanol due to dispersion forces between the hydrocarbon chains in each molecule.

- From Equation 13.2, we have

$$\begin{aligned} \Delta H_{\text{soln}} \text{ of KNO}_3 &= \Delta H_{\text{mice}} \text{ of KNO}_3 \\ &+ (\Delta H_{\text{hyd}} \text{ of K}^+ + \Delta H_{\text{hyd}} \text{ of NO}_3^-) \\ &= 34.89 \text{ kJ/mol} + 685 \text{ kJ/mol} + (\Delta H_{\text{hyd}} \text{ of K}^+ + \Delta H_{\text{hyd}} \text{ of NO}_3^-) \\ \Delta H_{\text{hyd}} \text{ of K}^+ + \Delta H_{\text{hyd}} \text{ of NO}_3^- &= 34.89 \text{ kJ/mol} - 685 \text{ kJ/mol} \\ &= -650. \text{ kJ/mol} \end{aligned}$$

- From Equation 13.2, we have

$$\begin{aligned} \Delta H_{\text{soln}} \text{ of NaCN} &= \Delta H_{\text{mice}} \text{ of NaCN} \\ &+ (\Delta H_{\text{hyd}} \text{ of Na}^+ + \Delta H_{\text{hyd}} \text{ of CN}^-) \\ \Delta H_{\text{soln}} \text{ of NaCN} &= 766 \text{ kJ/mol} + (-410. \text{ kJ/mol} + \Delta H_{\text{hyd}} \text{ of CN}^-) \\ \Delta H_{\text{hyd}} \text{ of CN}^- &= 1.21 \text{ kJ/mol} - 766 \text{ kJ/mol} + 410. \text{ kJ/mol} \\ &= -355 \text{ kJ/mol} \end{aligned}$$

- The partial pressure of N_2 in air is the volume percent divided by 100 times the total pressure (Dalton's law, Section 5.4):
 $P_{\text{N}_2} = 0.78 \times 1 \text{ atm} = 0.78 \text{ atm}$.

$$\begin{aligned} S_{\text{gas}} &= k_H \times P_{\text{gas}} \\ S_{\text{N}_2} &= (7 \times 10^{-4} \text{ mol/L}\cdot\text{atm})(0.78 \text{ atm}) \\ &= 5 \times 10^{-4} \text{ mol/L} \end{aligned}$$

- In a mixture of gases, the volume percent of a gas divided by 100 times the total pressure equals the gas's partial pressure (Dalton's law, Section 5.4):

$$P_{\text{gas}} = 0.40 \times 1.2 \text{ atm} = 0.48 \text{ atm}$$

$$k_H = \frac{S_{\text{gas}}}{P_{\text{gas}}} = \frac{1.2 \times 10^{-2} \text{ mol/L}}{0.48 \text{ atm}} = 2.5 \times 10^{-2} \text{ mol/L}\cdot\text{atm}$$

- Convert mass (g) of ethanol to kg, multiply by the molality to obtain amount (mol) of glucose, and then multiply amount (mol) of glucose by the molar mass to obtain mass of glucose.

$$\begin{aligned} \text{Amount (mol) of glucose} &= 563 \text{ g ethanol} \times \frac{1 \text{ kg}}{10^3 \text{ g}} \times \frac{2.40 \times 10^{-2} \text{ mol glucose}}{1 \text{ kg ethanol}} \\ &= 1.35 \times 10^{-2} \text{ mol glucose} \end{aligned}$$

$$\begin{aligned} \text{Mass (g) glucose} &= 1.35 \times 10^{-2} \text{ mol C}_6\text{H}_{12}\text{O}_6 \times \frac{180.16 \text{ g C}_6\text{H}_{12}\text{O}_6}{1 \text{ mol C}_6\text{H}_{12}\text{O}_6} \\ &= 2.43 \text{ g glucose} \end{aligned}$$

- Convert mass (g) of 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 I_2 by kg of $(\text{CH}_3\text{CH}_2)_2\text{O}$.

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

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

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

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.

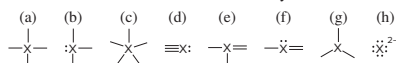
PROBLEMS

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 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 structures 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) SiF_4 ; (b) SeCl_2 ; (c) COF_2 (C is the central atom).
- 10.6** Draw a Lewis structure for (a) PH_3 ; (b) C_2F_4 ; (c) SbH_3 .
- 10.7** Draw a Lewis structure for (a) PF_3 ; (b) H_2CO_3 (both H atoms are attached to O atoms); (c) CS_2 .
- 10.8** Draw a Lewis structure for (a) CH_4S ; (b) S_2Cl_2 ; (c) CHCl_3 .
- 10.9** Draw Lewis structures of all the important resonance forms of (a) NO_2^- ; (b) NO_2F (N is central).
- 10.10** Draw Lewis structures of all the important resonance forms of (a) HNO_3 (HONO_2); (b) HAsO_4^{2-} (HOAsO_3^{2-}).
- 10.11** Draw Lewis structures of all the important resonance forms of (a) N_3^- ; (b) NO_2^- .
- 10.12** Draw Lewis structures of all the important resonance forms of (a) HCO_2^- (H is attached to C); (b) HBrO_4 (HOBrO_3).
- 10.13** Draw the Lewis structure with lowest formal charges, and determine the charge of each atom in (a) IF_5 ; (b) AlH_4^- .
- 10.14** Draw the Lewis structure with lowest formal charges, and determine the charge of each atom in (a) OCS ; (b) NO .
- 10.15** Draw the Lewis structure with lowest formal charges, and determine the charge of each atom in (a) CN^- ; (b) ClO^- .
- 10.16** Draw the Lewis structure with lowest formal charges, and

- draw a Lewis structure for a resonance form of each ion with the lowest possible formal charges, show the charges, and give oxidation numbers of the atoms: (a) AsO_4^{3-} ; (b) ClO_2^- .

- 10.19** These species do not obey the octet rule. Draw a Lewis structure for each, and state the type of octet-rule exception: (a) BH_3 (b) AsF_7^- (c) SeCl_4

- 10.20** These species do not obey the octet rule. Draw a Lewis structure for each, and state the type of octet-rule exception: (a) PF_6^- (b) ClO_3 (c) H_3PO_3 (one P—H bond)

- 10.21** These species do not obey the octet rule. Draw a Lewis structure for each, and state the type of octet-rule exception: (a) BrF_3 (b) ICl_2^- (c) BeF_2

- 10.22** These species do not obey the octet rule. Draw a Lewis structure for each, and state the type of octet-rule exception: (a) O_3^- (b) XeF_2 (c) SbF_4^-

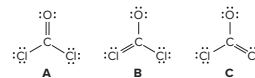
Problems in Context

- 10.23** Molten beryllium chloride reacts with chloride ion from molten NaCl to form the 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 (BrO_7^-) 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 BrO_7^- in which all atoms have lowest formal charges.

- 10.25** Cryolite (Na_3AlF_6) is an indispensable component in the electrochemical production of aluminum. Draw a Lewis structure for the 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 (AX_mE_n) and the ideal bond angle.
- 10.30** Name all the molecular shapes that have a tetrahedral electron-group arrangement.

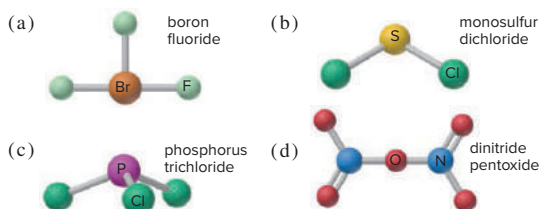
Comprehensive Problems

- 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 3.1×10^{-11} m; the radius of its nucleus is 2.5×10^{-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 6.64648×10^{-24} g, and each of its two electrons has a mass of 9.10939×10^{-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:	Mg^{2+}	K^+	Rb^+	Ba^{2+}	Cl^-	O^{2-}	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 digital 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, 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 the eighth edition focused on continued optimization of the text. To aid us in this process, we were able to use data from literally thousands of student responses to questions in LearnSmart, the adaptive learning system that assesses student knowledge of course content. The data, such as average time spent answering each question and the percentage of students who correctly answered the question on the first attempt, revealed the learning objectives that students found particularly difficult. We utilized several approaches to present these difficult concepts in a clearer, more straightforward way in the eighth edition of *Chemistry: The Molecular Nature of Matter and Change*.

Making the concepts clearer through digital learning resources. Students will be able to access over 2,000 digital learning resources throughout this text's SmartBook. These learning resources present summaries of concepts and worked examples, including over 400 videos of chemistry instructors solving problems or modeling concepts that students can view over and over again. Thus, students can have an "office hour" moment at any time.

NEW! Student Hot Spot

We are very pleased to incorporate real student data points and input, derived from thousands of our LearnSmart users, to help guide our revision. LearnSmart Heat Maps provided a quick visual snapshot of usage of portions of the text and the relative difficulty students experienced in mastering the content. With these data, we were able to both hone our text content when needed and, for particularly challenging concepts, point students to the learning resources that can elucidate and reinforce those concepts. You'll see these marginal features throughout the text. Students should log into Connect and view the resources through our SmartBook.

50 CHAPTER 2 Atoms, Molecules, and Ions

2.4 The Periodic Table

More than half of the elements known today were discovered between 1800 and 1900. During this period, chemists noted that the physical and chemical properties of certain groups of elements were similar to one another. These similarities, together with the need to organize the large volume of available information about the structure and properties of elemental substances, led to the development of the **periodic table**, a chart in which elements having similar chemical and physical properties are grouped together. Figure 2.10 shows the modern periodic table in which the elements are arranged by atomic number (shown above the element symbol) in horizontal rows called **periods** and in vertical columns called **groups** or **families**. Elements in the same group tend to have similar physical and chemical properties.

The elements can be categorized as metals, nonmetals, or metalloids. A **metal** is a good conductor of heat and electricity, whereas a **nonmetal** is usually a poor conductor of heat and electricity. A **metalloid** has properties that are intermediate between those of metals and nonmetals.

Figure 2.10 shows that the majority of known elements are metals; only 17 elements are nonmetals, and fewer than 10 elements are metalloids. Although most sources, including this text, designate the elements B, Si, Ge, As, Sb, and Te as metalloids, sources vary for the elements Po and At. In this text, we classify both Po and At as metalloids. From left to right across any period, the physical and chemical properties of the elements change gradually from metallic to nonmetallic.

Elements are often referred to collectively by their periodic table group number (Group 1A, Group 2A, and so on). For convenience, however, some element groups have been given special names. The Group 1A elements, with the exception of H (i.e., Li, Na, K, Rb, Cs, and Fr), are called **alkali metals**, and the Group 2A elements (Be, Mg, Ca, Sr, Ba, and Ra) are called **alkaline earth metals**. Elements in Group 6A (O, S, Se, Te, and Po) are sometimes referred to as the **chalcogens**. Elements in Group 7A (F, Cl, Br, I, and At) are known as **halogens**, and elements in Group 8A (He, Ne, Ar, Kr, Xe, and Rn) are called **noble gases**, or rare gases. The elements in Group 1B and Groups 3B–8B collectively are called the **transition elements** or **transition metals**.

The periodic table is a handy tool that correlates the properties of the elements in a systematic way and helps us to predict chemical behavior. At the turn of the twentieth century, the periodic table was deemed "the most predictive tool in all of science." We will take a more detailed look at this keystone of chemistry in Chapter 7.

SUGGESTED RESOURCES

1. Slide

2. Slide

3. Video

$$\begin{aligned} \text{Mass (kg) of uranium} &= \text{mass (kg) of pitchblende} \times \frac{\text{mass (kg) of uranium in pitchblende}}{\text{mass (kg) of pitchblende}} \\ &= 102 \text{ kg pitchblende} \times \frac{71.4 \text{ kg uranium}}{84.2 \text{ kg pitchblende}} = 86.5 \text{ kg uranium} \end{aligned}$$

Converting the mass of uranium from kg to g:

$$\text{Mass (g) of uranium} = 86.5 \text{ kg uranium} \times \frac{1000 \text{ g}}{1 \text{ kg}} = 8.65 \times 10^4 \text{ g uranium}$$

Finding the mass (in kg) of oxygen in 102 kg of pitchblende:

$$\begin{aligned} \text{Mass (kg) of oxygen} &= \text{mass (kg) of pitchblende} - \text{mass (kg) of uranium} \\ &= 102 \text{ kg} - 86.5 \text{ kg} = 15.5 \text{ kg oxygen} \end{aligned}$$

Converting the mass of oxygen from kg to g:

$$\text{Mass (g) of oxygen} = 15.5 \text{ kg oxygen} \times \frac{1000 \text{ g}}{1 \text{ kg}} = 1.55 \times 10^4 \text{ g oxygen}$$

Check The analysis showed that most of the mass of pitchblende is due to uranium, so the large mass of uranium makes sense. Rounding off to check the math gives

$$\sim 100 \text{ kg pitchblende} \times \frac{70}{85} = 82 \text{ kg uranium}$$

FOLLOW-UP PROBLEMS

2.2A The mineral "fool's gold" does not contain any gold; instead it is a compound composed only of the elements iron and sulfur. A 110.0-g sample of fool's gold contains 51.2 g of iron. What mass of sulfur is in a sample of fool's gold that contains 86.2 g of iron?

2.2B Silver bromide is the light-sensitive compound coated onto black-and-white film. A 26.8-g sample contains 15.4 g of silver, with bromine as the only other element. How many grams of each element are on a roll of film that contains 3.57 g of silver bromide?

SOME SIMILAR PROBLEMS 2.22–2.25

Student Hot Spot

Student data indicate that you may struggle with using mass fraction to calculate the mass of an element in a compound. Access the Smartbook to view additional Learning Resources on this topic.

Applying ideas with enhanced problems throughout the chapters. The much admired four-part problem-solving format (plan, solution, check, follow-up) is retained in the eighth edition, in both data-based and molecular-scene *Sample Problems*. Two *Follow-up Problems* are included with each sample problem, as well as a list of *Similar Problems* within the end-of-chapter problem set. *Brief Solutions* for all of the follow-up problems appear at the end of each chapter (rather than providing just a numerical answer in a distant end-of-book appendix, as is typical). The eighth edition has over 250 sample problems and over 500 follow-up problems. In almost every chapter, several sample and follow-up problems (and their brief solutions) were revised in this edition with two goals in mind. We sought to provide students with a variety of problems that would clearly elucidate concepts and demonstrate problem solving techniques, while giving students the opportunity to be challenged and gain competence. We also included more intermediate steps in the solutions to both sample and follow-up problems so that students could more easily follow the solutions.

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. Several existing figures have been revised and several new ones added to create an even better teaching tool. We continue to streamline figure legends by placing their content into clarifying annotations with the figures themselves.

Mastering the content with abundant end-of-chapter problem sets. New problems were added to several chapter problem sets, providing students and teachers with abundant choices in a wide range of difficulty and real-life scenarios. The problem sets are more extensive than in most other texts.

Content Changes to Individual Chapters

In addition to the general optimization of concept explanations and problem solutions throughout the text, specific improvements were made to most chapters:

- **Chapter 1** has a revised table of decimal prefixes and SI units to make conversion among SI units clearer, a revised discussion on intensive and extensive properties, and a revised sample problem on density.
- **Chapter 2** includes revised sample problems on mass percent and naming of compounds.
- **Chapter 3** has several new end-of-chapter problems: one new problem on the determination of a molecular formula, two new problems on writing a balanced reaction and determining the limiting reactant from molecular scenes, and two new stoichiometric problems involving limiting reactants.
- **Chapter 4** includes a new figure illustrating the activity series of the halogens. Sample problems on stoichiometry in precipitation and acid-base reactions were revised to include reactions that do not have 1:1 mole ratios.
- **Chapter 5** has two revised sample problems that provide students with additional opportunities for pressure unit conversions and stoichiometry calculations for gas reactions.

- **Chapter 6** has a clearer and more detailed discussion on pressure-volume work and a revised sample problem on the calorimetric determination of heat of combustion. Also included are new end-of-chapter problems on the calculation of enthalpy change for an aqueous reaction and determination of heat of combustion with bomb calorimetry.
- **Chapter 7** contains a new table summarizing the relationships between the quantum numbers and orbitals for the first four main energy levels.
- **Chapter 8** contains a new figure on electron spin; orbital diagrams have been added to the solutions of several sample problems.
- **Chapter 9** has improvements to several figures, a more detailed discussion of relationship between difference in electronegativity and ionic character, and some new follow-up problems.
- **Chapter 10** includes more detailed examples of depicting molecules with double bonds and ions with Lewis structures. Sample and follow-up problems have been revised to provide more opportunities to calculate formal charges and use those to evaluate resonance structures.
- **Chapter 11** has new art to illustrate formation of sigma and pi bonds and a new figure to show the placement of lone pairs in hybrid orbitals.
- **Chapter 12** includes additional information about viscosity and intermolecular forces.
- **Chapter 13** includes a more challenging sample problem on Henry's law, as well as revisions to several follow-up problems. There are new problems on the calculation of molar mass from freezing point depression.
- **Chapter 15** incorporates new art to make nomenclature clearer and a revised figure to show the key stages in protein synthesis.
- **Chapter 16** has a revised sample problem using the first-order integrated rate law, a revised figure on reaction mechanisms, and a new molecular scene problem on first-order reactions.
- **Chapter 17** contains a revised table on concentration ratios in an equilibrium system and two new sample problems, one on finding the equilibrium constant for an overall reaction, and the other on converting between K_p and K_c .
- **Chapter 18** has a new table on magnitude of K_a and percent dissociation and two revised sample problems.
- **Chapter 19** has a revised sample problem on buffer pH that reflects a more realistic lab procedure, a new molecular scene problem involving buffer solutions, a clearer presentation of pH calculations during acid-base titrations, and revised figures of pH titration curves. The section on acid-base indicators has been expanded, including the addition of a new figure about choosing an indicator for each type of acid-base titration. The discussion of aqueous solutions of metal sulfides was simplified.
- **Chapter 20** incorporates a new table that summarizes Q , K , ΔG , and reaction spontaneity.
- **Chapter 21** has several revised follow-up problems.
- **Chapter 23** has a new figure illustrating chelate complex ions and several revised figures. A new equation for calculating the charge of the metal ion in a complex ion has been added.

- **Chapter 24** has a new table summarizing changes in mass and atomic numbers during radioactive decay; a table on stability of even vs. odd numbers of nucleons has been revised. The discussion about mode of decay and neutron/proton ratio has been expanded.

Addition of Advanced Topics

In this special version of the 8th edition, advanced topics have been added to three chapters for use in classes in which a deeper and more rigorous level of discussion is appropriate. Problems on these advanced topics have been added to the end-of-chapter problem sets and to the online homework question bank.

- **Chapter 7** includes an expanded discussion on the development of the Schrödinger equation and the particle-in-a-box model. A new sample problem gives students an opportunity to apply the particle-in-a-box model to electron transitions.
- **Chapter 16** incorporates the calculus involved in the derivation of the integrated rate laws for zero-, first-, and second-order reactions. Also now included are discussions of pseudo-first-order reactions, steady-state approximation, and the Michaelis-Menten equation for enzyme kinetics.
- **Chapter 20** has a significantly expanded section on entropy. The calculations of entropy changes during isothermal gas expansion or contraction, phase changes, and changes in temperature have been added to enhance the current content; three new sample problems demonstrating these entropy change calculations are included.

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 use 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 emphasize 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 stress 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 today and register to experience how McGraw-Hill Create empowers you to teach *your* students *your* way. www.mcgrawhillcreate.com



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.



© Hero Images/Getty Images

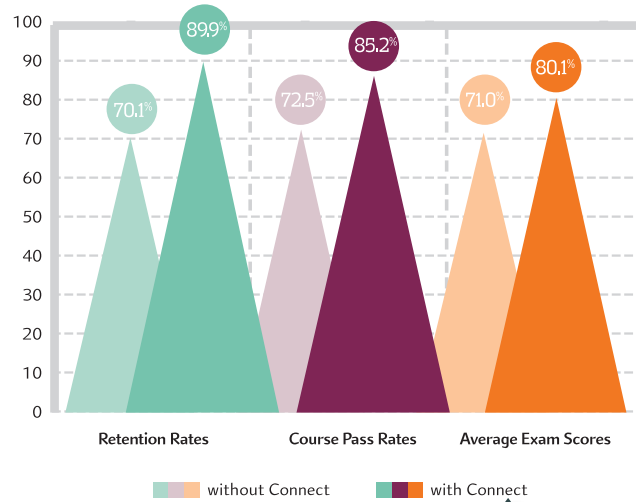
McGraw-Hill Connect® Learn Without Limits

Connect is a teaching and learning platform that is proven to deliver better results for students and instructors.

Connect empowers students by continually adapting to deliver precisely what they need, when they need it, and how they need it, so your class time is more engaging and effective.

73% of instructors who use **Connect** require it; instructor satisfaction **increases by 28%** when **Connect** is required.

Connect's Impact on Retention Rates, Pass Rates, and Average Exam Scores



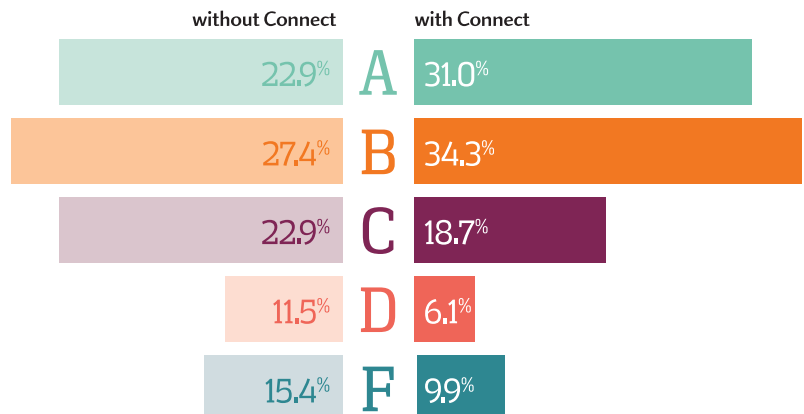
Using **Connect** improves retention rates by **19.8%**, passing rates by **12.7%**, and exam scores by **9.1%**.

Analytics

Connect Insight®

Connect Insight is Connect's new one-of-a-kind visual analytics dashboard that provides at-a-glance information regarding student performance, which is immediately actionable. By presenting assignment, assessment, and topical performance results together with a time metric that is easily visible for aggregate or individual results, Connect Insight gives the user the ability to take a just-in-time approach to teaching and learning, which was never before available. Connect Insight presents data that helps instructors improve class performance in a way that is efficient and effective.

Impact on Final Course Grade Distribution



Adaptive



THE ADAPTIVE READING EXPERIENCE DESIGNED TO TRANSFORM THE WAY STUDENTS READ

More students earn **A's** and **B's** when they use McGraw-Hill Education **Adaptive** products.

SmartBook®

Proven to help students improve grades and study more efficiently, SmartBook contains the same content within the print book, but actively tailors that content to the needs of the individual. SmartBook's adaptive technology provides precise, personalized instruction on what the student should do next, guiding the student to master and remember key concepts, targeting gaps in knowledge and offering customized feedback, and driving the student toward comprehension and retention of the subject matter. Available on tablets, SmartBook puts learning at the student's fingertips—anywhere, anytime.

Over **8 billion** questions have been answered, making McGraw-Hill Education products more intelligent, reliable, and precise.

www.mheducation.com

STUDENTS WANT
McGraw Hill Education SMARTBOOK®



of students reported **SmartBook** to be a more effective way of reading material.



of students want to use the Practice Quiz feature available within **SmartBook** to help them study.



of students reported having reliable access to off-campus wifi.



of students say they would purchase **SmartBook** over print alone.



of students reported that **SmartBook** would impact their study skills in a positive way.

Mc
Graw
Hill
Education

*Findings based on 2015 focus group results administered by McGraw-Hill Education

ADDITIONAL INSTRUCTOR AND STUDENT RESOURCES FOR YOUR COURSE!



MCGRAW-HILL CONNECT CHEMISTRY

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. Also available within Connect, our adaptive SmartBook has been supplemented with additional learning resources tied to each learning objective to provide point-in-time help to students who need it. To learn more, visit www.mheducation.com.

Instructors have access to the following instructor resources through Connect.

- **Art** Full-color digital files of all illustrations, photos, and tables 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.
- **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 Lecture Outlines** Ready-made presentations that combine art and lecture notes are provided for each chapter of the text.
- **Computerized Test Bank** Over 2300 test questions that accompany *Chemistry: The Molecular Nature of Matter and Change* are available utilizing the industry-leading test generation software TestGen. These same questions are also available and assignable through Connect for online tests.
- **Instructor's Solutions Manual** This supplement, prepared by Mara Vorachek-Warren of St. Charles Community College, contains complete, worked-out solutions for *all* the end-of-chapter problems in the text.

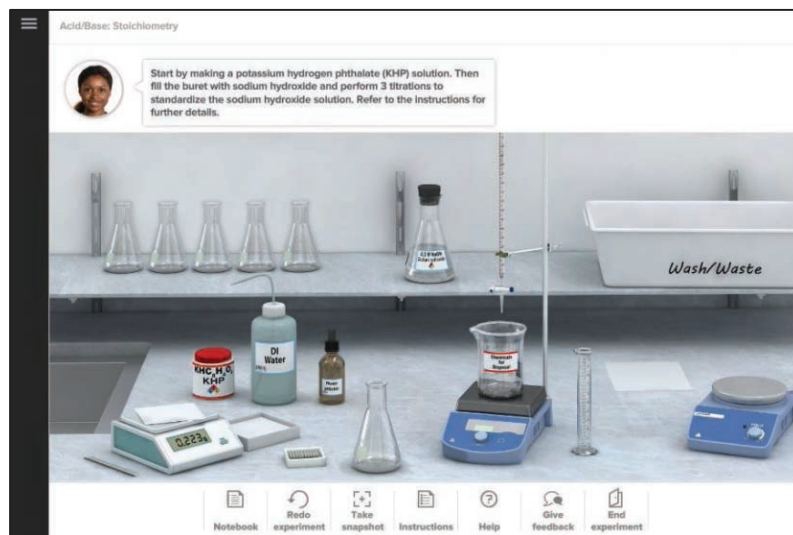
LEARNSMART



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 gaps in prerequisite knowledge 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.

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 students' online experience one that rivals the real world, LearnSmart Labs accomplishes it all.



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 SOLUTIONS MANUAL

This supplement, prepared by Mara Vorachek-Warren of St. Charles Community College, contains detailed solutions and explanations for all problems in the main text that have colored numbers.

ACKNOWLEDGMENTS

It would be nearly impossible to put together a more professional, talented, and supportive publishing team than our colleagues at McGraw-Hill Education: Managing Director Thomas Timp, Director of Chemistry David Spurgeon, Ph.D., Associate Director of Digital Content Robin Reed, Program Manager Lora Neyens, Content Project Manager Laura Bies, Designer David Hash, Marketing Manager Matthew Garcia, and Director of Digital Content Shirley Hino. It is a pleasure to work with them; their leadership, knowledge, and encouragement have helped to make this latest edition a reality.

Mara Vorachek-Warren of St. Charles Community College provided a thorough accuracy check of all the new

sample problems, follow-up problems, and end-of-chapter problems as part of her superb preparation of both the Student and Instructor's Solutions Manuals.

The following individuals helped write and review learning goal-oriented content for **LearnSmart for General Chemistry**: Margaret Ruth Leslie, Kent State University and Adam I. Keller, Columbus State Community College.

Several expert freelancers contributed as well. Jane Hoover did her usual excellent job in copyediting the text, and Lauren Timmer and Louis Poncz followed with meticulous proofreading. And many thanks to Jerry Marshall, who patiently researched new stock and studio photos.

CHEMISTRY

Advanced Topics



1

Keys to Studying Chemistry: Definitions, Units, and Problem Solving

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



Source: © Fancy Collection/SuperStock

- › exponential (scientific) notation (Appendix A)

Maybe 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 your everyday life. For example, does the following morning routine (described in chemical terms) sound familiar? You are awakened by the buzzing of your alarm clock, a sound created when molecules align in the liquid-crystal display of your clock and electrons flow to create a noise. You throw off a thermal insulator of manufactured polymer (blanket) and jump in the shower to emulsify fatty substances on your skin and hair with purified water and formulated detergents. Next you adorn 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 abrade your teeth with a colloidal dispersion of artificially flavored, dental-hardening agents (toothpaste), grab your portable electronic device containing ultrathin, microetched semiconductor layers powered by a series of voltaic cells (laptop), collect some objects made from processed cellulose and plastic, electronically printed with light- and oxygen-resistant inks (books), hop in your hydrocarbon-fueled, metal-vinyl-ceramic vehicle, electrically ignite a synchronized series of controlled gaseous explosions (start your car), and take 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 things. 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 by defining it and a few central concepts. **Chemistry** is *the scientific 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 pattern, or *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.

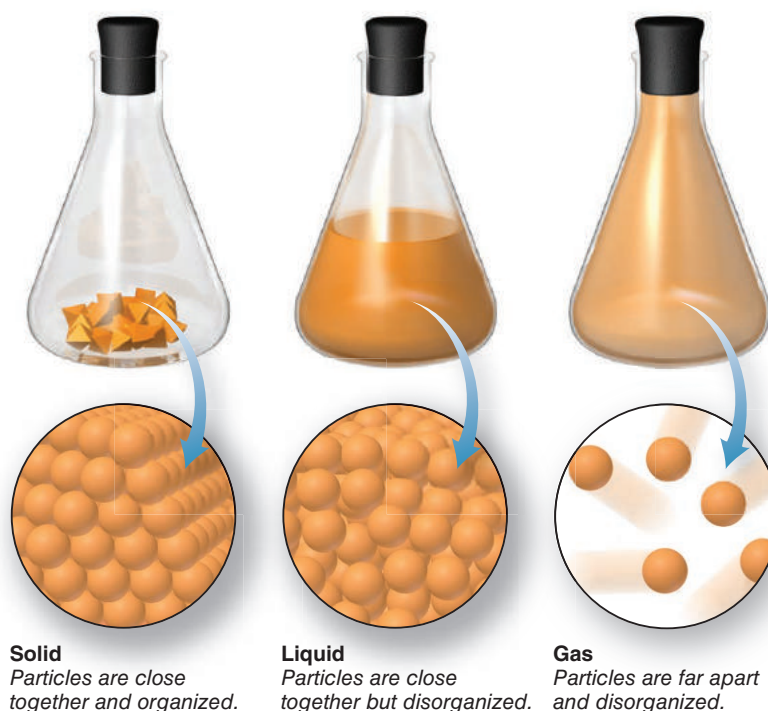


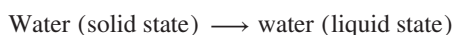
Figure 1.1 The physical states of matter.

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 color, 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; however, *the particles are the same* in both states of water.

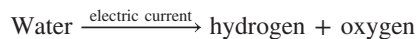
Physical change (same substance before and after):



All changes of state of matter are physical changes.

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 *one or more substances are converted into one or more substances with different composition and properties*. 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.

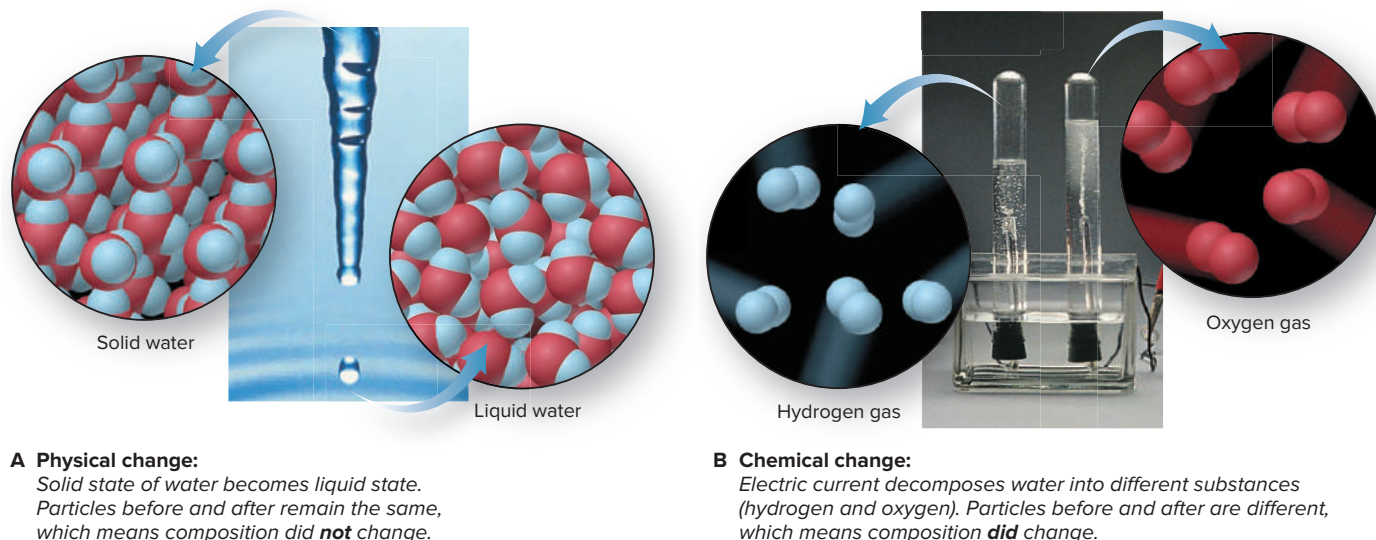


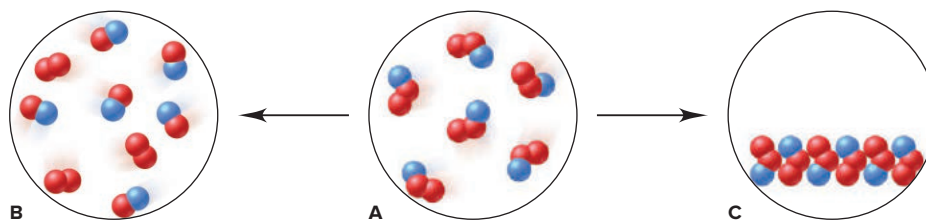
Figure 1.2 The distinction between physical and chemical change.

Source: (A) © Paul Morrell/Stone/Getty Images; (B) © McGraw-Hill Education/Stephen Frisch, photographer

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.

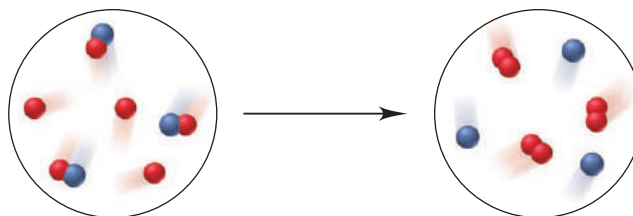
Plan Given depictions of two changes, we have to determine whether each represents a physical or a chemical change. The number and colors of the little spheres that make up each particle tell its “composition.” Samples with particles of the *same* composition but in a different arrangement depict a *physical*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

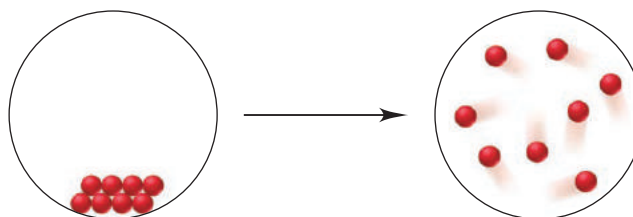
FOLLOW-UP PROBLEMS

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?



SOME SIMILAR PROBLEMS 1.1 and 1.84

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*).

Table 1.1

Some Characteristic Properties of Copper

Physical Properties

Easily hammered into sheets (malleable) and drawn into wires (ductile)



Can be melted and mixed with zinc to form brass



Density = 8.95 g/cm³
Melting point = 1083°C
Boiling point = 2570°C

Chemical Properties

Slowly forms a blue-green carbonate in moist air



Reacts with nitric or sulfuric acid

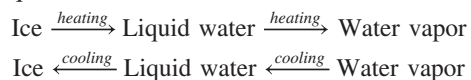


Slowly forms deep-blue solution in aqueous ammonia



Source: (copper) © McGraw-Hill Education/Mark Dierker, photographer; (candlestick) © Ruth Melnick; (copper carbonate, copper reacting with acid, copper and ammonia) © McGraw-Hill Education/Stephen Frisch, photographer

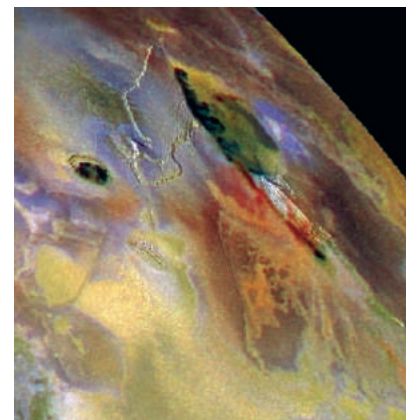
Similarly, as the temperature drops, water vapor condenses to liquid water, and with further cooling, the liquid freezes to ice:



In a steel plant, solid iron melts to liquid (molten) iron and then cools to the solid again. And, far beyond the confines of a laboratory or steel plant, lakes of molten sulfur (a solid on Earth at room temperature) 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 substances that are common on Earth occur in unusual states on other worlds.

Source: JPL-NASA

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 (a) Frost forming is a **physical change**: the drop in temperature changes water vapor (gaseous water) in humid air to ice crystals (solid water).

(b) 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.

(c) The match burning is a **chemical change**: the combustible substances in the match head are converted into other substances.

(d) Perspiration evaporating is a **physical change**: the water in sweat changes its state, from liquid to gas, but not its composition.

(e) 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:

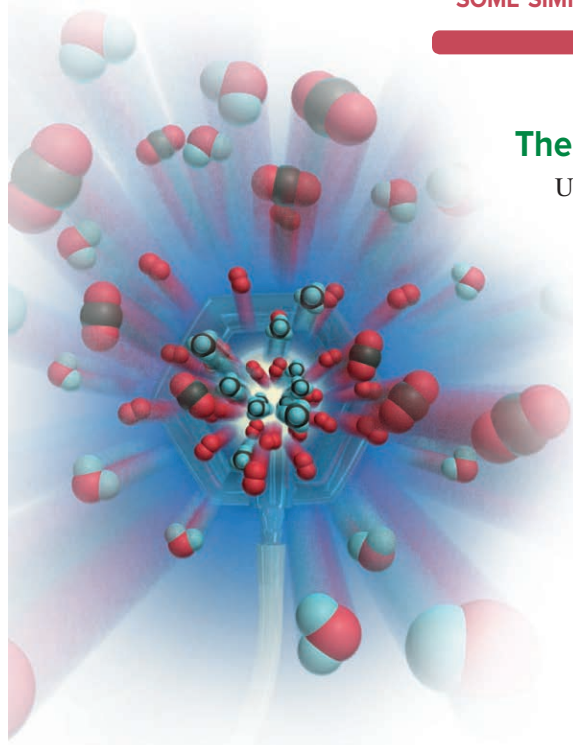
*We study **observable** changes in matter to understand their **unobservable** causes.*

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.



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

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 are more stable and are favored over situations of higher energy, which are 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

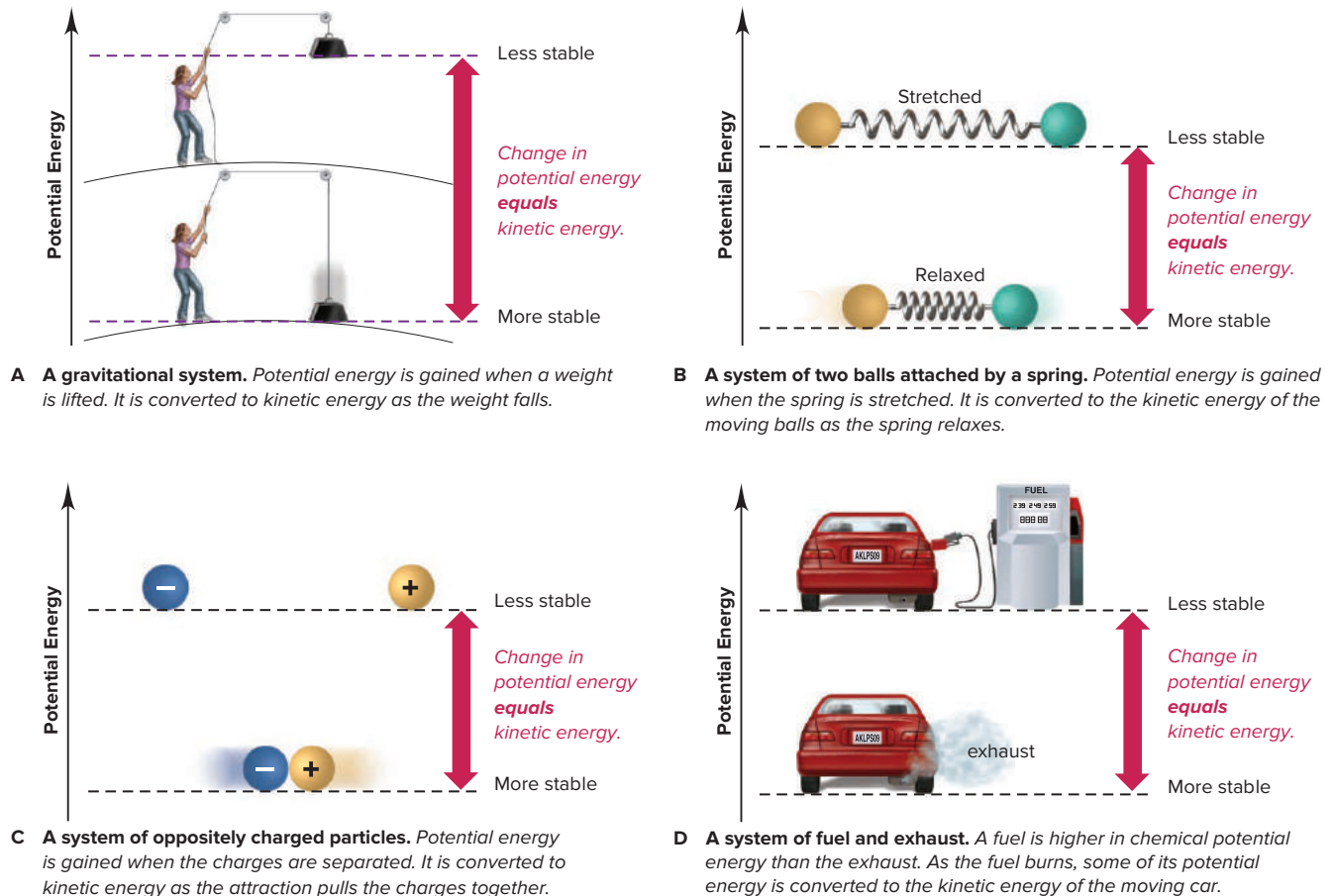


Figure 1.3 Potential energy is converted to kinetic energy. The dashed horizontal lines indicate the potential energy of each system before and after the change.