

 CENGAGE

CHEMISTRY

AN ATOMS FIRST APPROACH



ZUMDAHL
ZUMDAHL
DECOSTE

Third Edition

CHEMISTRY

AN ATOMS FIRST APPROACH

THIRD EDITION

Steven S. Zumdahl

University of Illinois

Susan A. Zumdahl

University of Illinois

Donald J. DeCoste

University of Illinois



Australia • Brazil • Mexico • Singapore • United Kingdom • United States

This is an electronic version of the print textbook. Due to electronic rights restrictions, some third party content may be suppressed. Editorial review has deemed that any suppressed content does not materially affect the overall learning experience. The publisher reserves the right to remove content from this title at any time if subsequent rights restrictions require it. For valuable information on pricing, previous editions, changes to current editions, and alternate formats, please visit www.cengage.com/highered to search by ISBN#, author, title, or keyword for materials in your areas of interest.

Important Notice: Media content referenced within the product description or the product text may not be available in the eBook version.

Chemistry: An Atoms First Approach,
Third Edition
Steven S. Zumdahl, Susan A. Zumdahl,
Donald J. DeCoste

Product Director: Thais Alencar

Product Manager: Maureen McLaughlin

Product Assistant: Nellie Mitchell

Marketing Manager: Timothy Cali

Content Manager: Brendan Killion

Learning Designer: Peter McGahey

In-House Subject Matter Expert, Chemistry:
Theresa Dearborn

Project Manager: Sharib Asrar, MPS Ltd.

Manufacturing Planner: Rebecca Cross

Intellectual Property Project Manager:
Betsy Hathaway

Intellectual Property Analyst: Ann Hoffman

Copy Editor: Heather Mann

Designer: Chris Doughman

Art and Production Management,
and Composition: MPS Ltd.

Cover Credit: David McCarthy/Science Source

© 2021, 2016 Cengage Learning, Inc.

Unless otherwise noted, all content is © Cengage.

WCN: 02-300

ALL RIGHTS RESERVED. No part of this work covered by the copyright herein may be reproduced or distributed in any form or by any means, except as permitted by U.S. copyright law, without the prior written permission of the copyright owner.

For product information and technology assistance, contact us at
Cengage Customer & Sales Support, 1-800-354-9706 or
support.cengage.com.

For permission to use material from this text or product,
submit all requests online at
www.cengage.com/permissions.

Library of Congress Control Number: 2019918615

ISBN: 978-0-357-36336-2

Loose-leaf Edition:
ISBN: 978-0-357-36356-0

Cengage
200 Pier 4 Boulevard
Boston, MA 02210
USA

Cengage is a leading provider of customized learning solutions with employees residing in nearly 40 different countries and sales in more than 125 countries around the world. Find your local representative at **www.cengage.com.**

Cengage products are represented in Canada by Nelson Education, Ltd.

To learn more about Cengage platforms and services, register or access your online learning solution, or purchase materials for your course, visit **www.cengage.com.**

Contents

To the Professor ix

To the Student xiii

Review | Measurement and Calculations in Chemistry 1

- R.1** Units of Measurement 2
 - R.2** Uncertainty in Measurement 4
 - R.3** Significant Figures and Calculations 7
 - R.4** Learning to Solve Problems Systematically 11
 - R.5** Dimensional Analysis 11
 - R.6** Temperature 15
 - R.7** Density 19
 - R.8** Classification of Matter 20
 - R.9** Energy 24
 - R.10** The Mole 27
- CHEMICAL CONNECTIONS** Critical Units! 28
For Review 28 | Key Terms 28 | Questions and Exercises 29

Chapter 1 | Chemical Foundations 31

- 1.1** Chemistry: An Atoms-First Approach 32
 - 1.2** The Scientific Method 36
 - 1.3** Fundamental Chemical Laws 38
 - 1.4** Dalton's Atomic Theory 40
 - 1.5** Early Experiments to Characterize the Atom 42
 - 1.6** The Modern View of Atomic Structure: An Introduction 45
- CONNECTING TO ATOMS** Atoms, Ions, and Isotopes: A Pictorial Summary 48
For Review 49 | Key Terms 49 | Questions and Exercises 51

Chapter 2 | Atomic Structure and Periodicity 53

- 2.1** Electromagnetic Radiation 54

- 2.2** The Nature of Matter 57
- CHEMICAL CONNECTIONS** Chemistry That Doesn't Leave You in the Dark 60
- 2.3** The Atomic Spectrum of Hydrogen 62
 - 2.4** The Bohr Model 63
- CHEMICAL CONNECTIONS** 0.035 Femtometer Is a Big Deal 67
- 2.5** The Quantum Mechanical Model of the Atom 68
 - 2.6** Quantum Numbers 71
 - 2.7** Orbital Shapes and Energies 72
 - 2.8** Electron Spin and the Pauli Principle 76
 - 2.9** Polyelectronic Atoms 76
 - 2.10** The History of the Periodic Table 78
- CHEMICAL CONNECTIONS** The Chemistry of Copernicium 81
- 2.11** The Aufbau Principle and the Periodic Table 81
 - 2.12** Periodic Trends in Atomic Properties 87
- CHEMICAL CONNECTIONS** Potassium—Too Much of a Good Thing Can Kill You 94
For Review 94 | Key Terms 94 | Questions and Exercises 97



Manex Catalapiedra/Getty Images



NoSystem Images/Getty Images

Chapter 3 | Bonding: General Concepts 98

- 3.1** Types of Chemical Bonds 99
CHEMICAL CONNECTIONS No Lead Pencils 101
- 3.2** Electronegativity 104
- 3.3** Ions: Electron Configurations and Sizes 106
- 3.4** Partial Ionic Character of Covalent Bonds 110
- 3.5** The Covalent Chemical Bond: A Model 111
- 3.6** The Localized Electron Bonding Model 114
- 3.7** Lewis Structures 115
CONNECTING TO ATOMS Lewis Structures and the Periodic Table 119
- 3.8** Exceptions to the Octet Rule 119
- 3.9** Resonance 123
- 3.10** Naming Simple Compounds 127
For Review 138 | Key Terms 138 | Questions and Exercises 140

Chapter 4 | Molecular Structure and Orbitals 142

- 4.1** Molecular Structure: The VSEPR Model 143
CHEMICAL CONNECTIONS Chemical Structure and Communication: Semiochemicals 150
CONNECTING TO ATOMS Polar Molecules—It's All About Symmetry 155
- 4.2** Bond Polarity and Dipole Moments 156
- 4.3** Hybridization and the Localized Electron Model 159
- 4.4** The Molecular Orbital Model 172
- 4.5** Bonding in Homonuclear Diatomic Molecules 175
- 4.6** Bonding in Heteronuclear Diatomic Molecules 181

- 4.7** Combining the Localized Electron and Molecular Orbital Models 182
CHEMICAL CONNECTIONS What's Hot? 183
For Review 185 | Key Terms 185 | Questions and Exercises 187

Chapter 5 | Stoichiometry 189

- 5.1** Counting by Weighing 190
- 5.2** Atomic Masses 191
- 5.3** Learning to Solve Problems 194
- 5.4** The Mole 195
- 5.5** Molar Mass 200
- 5.6** Percent Composition of Compounds 203
- 5.7** Determining the Formula of a Compound 204
- 5.8** Chemical Equations 211
- 5.9** Balancing Chemical Equations 213
- 5.10** Stoichiometric Calculations: Amounts of Reactants and Products 216
- 5.11** The Concept of Limiting Reactant 221
For Review 230 | Key Terms 230 | Questions and Exercises 233

Chapter 6 | Types of Chemical Reactions and Solution Stoichiometry 234

- 6.1** Water, the Common Solvent 235
- 6.2** The Nature of Aqueous Solutions: Strong and Weak Electrolytes 237
CHEMICAL CONNECTIONS Arrhenius: A Man with Solutions 240
- 6.3** The Composition of Solutions 241
- 6.4** Types of Chemical Reactions 248
- 6.5** Precipitation Reactions 248
- 6.6** Describing Reactions in Solution 254
- 6.7** Stoichiometry of Precipitation Reactions 255
- 6.8** Acid–Base Reactions 258
- 6.9** Oxidation–Reduction Reactions 265
- 6.10** Balancing Oxidation–Reduction Equations 270
For Review 272 | Key Terms 272 | Questions and Exercises 275

Chapter 7 | Chemical Energy 276

- 7.1** The Nature of Chemical Energy 277
- 7.2** Enthalpy 281
- 7.3** Calorimetry 282
CHEMICAL CONNECTIONS Nature Has Hot Plants 284
CHEMICAL CONNECTIONS Firewalking: Magic or Science? 286
- 7.4** Hess's Law 289
- 7.5** Standard Enthalpies of Formation 292

- 7.6** Covalent Bond Energies and Chemical Reactions 299
- 7.7** Present Sources of Energy 303
- 7.8** New Energy Sources 306
- CHEMICAL CONNECTIONS** Veggie Gasoline? 310
For Review 311 | Key Terms 311 | Questions and Exercises 313
- Chapter 8 | Gases** 315
- 8.1** Pressure 316
- 8.2** The Gas Laws of Boyle, Charles, and Avogadro 318
- 8.3** The Ideal Gas Law 324
- 8.4** Gas Stoichiometry 329
CHEMICAL CONNECTIONS High Mountains—Low Octane 330
- 8.5** Dalton's Law of Partial Pressures 334
CHEMICAL CONNECTIONS Separating Gases 336
CHEMICAL CONNECTIONS The Chemistry of Air Bags 340
- 8.6** The Kinetic Molecular Theory of Gases 340
- 8.7** Effusion and Diffusion 348
- 8.8** Real Gases 350
- 8.9** Characteristics of Several Real Gases 352
- 8.10** Chemistry in the Atmosphere 353
For Review 356 | Key Terms 356 | Questions and Exercises 359
- Chapter 9 | Liquids and Solids** 362
- 9.1** Intermolecular Forces 364
CONNECTING TO ATOMS Intermolecular Forces, Physical Properties, and Real Gases 368
- 9.2** The Liquid State 368
- 9.3** Vapor Pressure and Changes of State 371
- 9.4** Phase Diagrams 378
CHEMICAL CONNECTIONS Making Diamonds at Low Pressures: Fooling Mother Nature 380
- 9.5** An Introduction to Structures and Types of Solids 383
CHEMICAL CONNECTIONS Smart Fluids 386
- 9.6** Structure and Bonding in Metals 389
CHEMICAL CONNECTIONS Closest Packing of M & Ms 393
- 9.7** Carbon and Silicon: Network Atomic Solids 394
CHEMICAL CONNECTIONS Graphene—Miracle Substance? 395
- 9.8** Molecular Solids 402
- 9.9** Ionic Solids 403
For Review 406 | Key Terms 406 | Questions and Exercises 409
- Chapter 10 | Properties of Solutions** 411
- 10.1** Solution Composition 412
- 10.2** The Energies of Solution Formation 415
- 10.3** Factors Affecting Solubility 418
- 10.4** The Vapor Pressures of Solutions 422
CHEMICAL CONNECTIONS The Lake Nyos Tragedy 423
- 10.5** Boiling-Point Elevation and Freezing-Point Depression 428
- 10.6** Osmotic Pressure 432
- 10.7** Colligative Properties of Electrolyte Solutions 436
CHEMICAL CONNECTIONS The Drink of Champions—Water 438
- 10.8** Colloids 439
CHEMICAL CONNECTIONS Organisms and Ice Formation 441
For Review 441 | Key Terms 441 | Questions and Exercises 444
- Chapter 11 | Chemical Kinetics** 445
- 11.1** Reaction Rates 447
- 11.2** Rate Laws: An Introduction 450
- 11.3** Determining the Form of the Rate Law 452
- 11.4** The Integrated Rate Law 456
- 11.5** Reaction Mechanisms 467
- 11.6** A Model for Chemical Kinetics 470
- 11.7** Catalysis 476
CHEMICAL CONNECTIONS Enzymes: Nature's Catalysts 480
CONNECTING TO ATOMS Zero-, First-, and Second-Order Reactions and the Collision Model 483
For Review 484 | Key Terms 484 | Questions and Exercises 486



Mark Mawson/Stone/Getty Images



mitzy/Shutterstock.com

Chapter 12 | Chemical Equilibrium 487

- 12.1 The Equilibrium Condition 489
- 12.2 The Equilibrium Constant 491
- 12.3 Equilibrium Expressions Involving Pressures 495
- 12.4 Heterogeneous Equilibria 498
- 12.5 Applications of the Equilibrium Constant 500
- 12.6 Solving Equilibrium Problems 509
- 12.7 Le Châtelier's Principle 514

CHEMICAL CONNECTIONS The Difference Between *Equilibrium Constant* and *Equilibrium Condition* 521

For Review 522 | Key Terms 522 | Questions and Exercises 525

Chapter 13 | Acids and Bases 527

- 13.1 The Nature of Acids and Bases 528
- 13.2 Acid Strength 531
- 13.3 The pH Scale 536
 - CHEMICAL CONNECTIONS** Arnold Beckman, Man of Science 538
- 13.4 Calculating the pH of Strong Acid Solutions 540
- 13.5 Calculating the pH of Weak Acid Solutions 541
- 13.6 Bases 550
 - CHEMICAL CONNECTIONS** Amines 554
- 13.7 Polyprotic Acids 556
- 13.8 Acid–Base Properties of Salts 561
- 13.9 The Effect of Structure on Acid–Base Properties 566
- 13.10 Acid–Base Properties of Oxides 568

13.11 The Lewis Acid–Base Model 569

13.12 Strategy for Solving Acid–Base Problems: A Summary 571

CHEMICAL CONNECTIONS Understanding Various Models of Acids and Bases 572

For Review 573 | Key Terms 573 | Questions and Exercises 577

Chapter 14 | Acid–Base Equilibria 578

- 14.1 Solutions of Acids or Bases Containing a Common Ion 579
- 14.2 Buffered Solutions 582
- 14.3 Buffering Capacity 591
- 14.4 Titrations and pH Curves 594
- 14.5 Acid–Base Indicators 609
 - For Review 615 | Key Terms 615 | Questions and Exercises 617

Chapter 15 | Solubility and Complex Ion Equilibria 619

- 15.1 Solubility Equilibria and the Solubility Product 620
 - CHEMICAL CONNECTIONS** The Chemistry of Teeth 628
- 15.2 Precipitation and Qualitative Analysis 629
- 15.3 Equilibria Involving Complex Ions 635
 - For Review 640 | Key Terms 640 | Questions and Exercises 641a

Chapter 16 | Spontaneity, Entropy, and Free Energy 642

- 16.1 Spontaneous Processes and Entropy 643
 - CHEMICAL CONNECTIONS** Entropy: An Organizing Force? 649
- 16.2 Entropy and the Second Law of Thermodynamics 649
- 16.3 The Effect of Temperature on Spontaneity 650
- 16.4 Free Energy 653
- 16.5 Entropy Changes in Chemical Reactions 656
- 16.6 Free Energy and Chemical Reactions 660
 - CHEMICAL CONNECTIONS** Entropy and Positional Probability 662
- 16.7 The Dependence of Free Energy on Pressure 666
- 16.8 Free Energy and Equilibrium 669
- 16.9 Free Energy and Work 673
 - For Review 676 | Key Terms 676 | Questions and Exercises 678

- Chapter 17 | Electrochemistry** 679
- 17.1** Balancing Oxidation–Reduction Equations 680
- 17.2** Galvanic Cells 686
- 17.3** Standard Reduction Potentials 689
- 17.4** Cell Potential, Electrical Work, and Free Energy 696
- 17.5** Dependence of Cell Potential on Concentration 699
- 17.6** Batteries 705
CHEMICAL CONNECTIONS Fuel Cells—Portable Energy 708
- 17.7** Corrosion 708
- 17.8** Electrolysis 711
CHEMICAL CONNECTIONS The Chemistry of Sunken Treasure 714
- 17.9** Commercial Electrolytic Processes 716
For Review 721 | Key Terms 721 | Questions and Exercises 725
- Chapter 18 | The Nucleus: A Chemist's View** 727
- 18.1** Nuclear Stability and Radioactive Decay 728
- 18.2** The Kinetics of Radioactive Decay 733
CONNECTING TO ATOMS Radioactive Decay Processes 734
- 18.3** Nuclear Transformations 736
- 18.4** Detection and Uses of Radioactivity 739
- 18.5** Thermodynamic Stability of the Nucleus 743
- 18.6** Nuclear Fission and Nuclear Fusion 746
CHEMICAL CONNECTIONS Future Nuclear Power 748
- 18.7** Effects of Radiation 751
For Review 753 | Key Terms 753 | Questions and Exercises 755
- Chapter 19 | The Representative Elements** 757
- 19.1** A Survey of the Representative Elements 758
- 19.2** The Group 1A (1) Elements 763
- 19.3** The Chemistry of Hydrogen 764
- 19.4** The Group 2A (2) Elements 766
- 19.5** The Group 3A (13) Elements 768
- 19.6** The Group 4A (14) Elements 769
CHEMICAL CONNECTIONS Beethoven: Hair Is the Story 771
- 19.7** The Group 5A (15) Elements 772
- 19.8** The Chemistry of Nitrogen 773
CHEMICAL CONNECTIONS Nitrous Oxide: Laughing Gas That Propels Whipped Cream and Cars 779
- 19.9** The Chemistry of Phosphorus 780
- 19.10** The Group 6A (16) Elements 782
- 19.11** The Chemistry of Oxygen 783
- 19.12** The Chemistry of Sulfur 784
- 19.13** The Group 7A (17) Elements 787
- 19.14** The Group 8A (18) Elements 791
For Review 792 | Key Terms 792 | Questions and Exercises 795a
- Chapter 20 | Transition Metals and Coordination Chemistry** 796
- 20.1** The Transition Metals: A Survey 797
- 20.2** The First-Row Transition Metals 802
CHEMICAL CONNECTIONS Titanium Dioxide—Miracle Coating 804
- 20.3** Coordination Compounds 807
- 20.4** Isomerism 811
CHEMICAL CONNECTIONS The Importance of Being *cis* 814
- 20.5** Bonding in Complex Ions: The Localized Electron Model 816
- 20.6** The Crystal Field Model 818
CHEMICAL CONNECTIONS Transition Metal Ions Lend Color to Gems 821
- 20.7** The Biological Importance of Coordination Complexes 824
- 20.8** Metallurgy and Iron and Steel Production 828
For Review 836 | Key Terms 836 | Questions and Exercises 839
- Chapter 21 | Organic and Biological Molecules** 841
- 21.1** Alkanes: Saturated Hydrocarbons 842
CONNECTING TO ATOMS Structural Representations of Organic Molecules 844
- 21.2** Alkenes and Alkynes 851
- 21.3** Aromatic Hydrocarbons 854
- 21.4** Hydrocarbon Derivatives 856
- 21.5** Polymers 863
CHEMICAL CONNECTIONS Wallace Hume Carothers 864
CHEMICAL CONNECTIONS Super-Slippery Slope 865
- 21.6** Natural Polymers 871
CHEMICAL CONNECTIONS Tanning in the Shade 885
For Review 886 | Key Terms 886 | Questions and Exercises 889

Appendix 1 | Mathematical Procedures A1

- A1.1** Exponential Notation A1
- A1.2** Logarithms A4
- A1.3** Graphing Functions A6
- A1.4** Solving Quadratic Equations A7
- A1.5** Uncertainties in Measurements A10

Appendix 2 | The Quantitative Kinetic Molecular Model A13

Appendix 3 | Spectral Analysis A16

Appendix 4 | Selected Thermodynamic Data A19

Appendix 5 | Equilibrium Constants and Reduction Potentials A22

A5.1 Values of K_a for Some Common Monoprotic Acids A22

A5.2 Stepwise Dissociation Constants for Several Common Polyprotic Acids A23

A5.3 Values of K_b for Some Common Weak Bases A23

A5.4 K_{sp} Values at 25°C for Common Ionic Solids A24

A5.5 Standard Reduction Potentials at 25°C (298 K) for Many Common Half-Reactions A25

Appendix 6 | SI Units and Conversion Factors A26

Glossary A27

Answers to Selected Exercises A39

Index A71

To the Professor

Features of *Chemistry: An Atoms First Approach*

Conceptual learning and problem solving have been fundamental to our approach in *Chemistry* through ten successful editions, and *Chemistry: An Atoms First Approach* through two successful editions. Our philosophy is to help students learn to think like chemists so they can apply the process of problem solving to all aspects of their lives. In the third edition of *Chemistry: An Atoms First Approach*, we have continued with this philosophy while making some structural changes. Atoms, molecules, and bonding are still discussed early in the text, and we have included stoichiometry immediately following VSEPR and MO theory. This is to allow students to be able to perform measurement and stoichiometry experiments early in the course. In this edition, we have extended the discussion of energy in the Review chapter so that the students will have sufficient background to better understand electron energy states and the importance of distinguishing between a system and its surroundings.

We found that users of our texts strongly support our approach to general chemistry. We have always written with a student-first approach. That is, in writing every page we ask ourselves: “How can we explain the material in a way that will be most clear to the students?” We always develop concepts in accord with the scientific method. That is, we always consider the observed properties of substances first. We then look for common threads among these properties (formulate laws). Finally we help the students understand why and how the theories of chemistry developed. In describing theories we always make clear that models are works in progress. We expect to find areas where the models fail and, in fact, that this occurrence often leads to the greatest progress in our understanding of how nature operates.

One of the main goals of our treatment of chemistry is to help students learn to be effective problem solvers. We want to go beyond memorized steps to help students think their way through the problems. To do this we take a “think like a chemist” approach. In solving problems we ask students several questions to guide them through the process: Where are we going?, What do we know?, and How do we get there? Our goal is to foster creative, concept-based problem solving, which will serve the students in their lives and careers beyond the general chemistry course.

Over the years, thousands of students and instructors who have used our books have found that these approaches work. The third edition of *Chemistry: An Atoms First Approach* continues this tradition by being a text that effectively explains chemistry to the students and helps them to learn to be creative problem solvers. The most significant difference from our widely used *Chemistry* textbook is that we present atoms, bonding, and molecules at the beginning of the text and have made sure that the rest of the topics flow smoothly from this starting point, while moving stoichiometry to a position earlier in the text.

To strengthen the atoms first approach we have emphasized at every opportunity throughout the text the importance and advantages of thinking about chemistry from an atomic/molecular perspective.

What's New

Considerable effort went into making the third edition of *Chemistry: An Atoms First Approach*.

- › **Core Text**—All chapters were reviewed for clarity, enhanced understanding, and cohesive flow. Many sections were also revised as needed to help enhance the atoms first emphasis throughout the textbook.
- › **Energy**—An increased discussion of energy is included in the Review chapter (Chapter R). This includes discussions of the concepts of *endothermic*, *exothermic*, *system*, *surroundings*, and *chemical energy*.
- › **Chemistry Pioneers**—A new boxed feature “Chemistry Pioneers” contains historical background on many early chemists.
- › **Bond Energy**—We have moved the bond energy discussion from the bonding chapter to the energy chapter. In this way we are able to more coherently discuss different methods for determining the enthalpy change for a reaction. We also provide a direct comparison of using bond energies and standard enthalpies of formations for determining the enthalpy of a given reaction and discuss the results.
- › **Liquids and Solids**—Chapter 9 has been reorganized so that the discussion of vapor pressure and phase diagrams comes immediately after the discussion of liquids. This makes for a smoother transition to discussing phase changes.

- › **ChemWork Problems**—All of end of chapter problems were converted into ChemWork problems. These multi-concept problems (and additional ones) are found interactively online with the same type of assistance a student would get from an instructor.

Hallmarks of *Chemistry: An Atoms First Approach*

- › *Chemistry: An Atoms First Approach* contains numerous discussions, illustrations, and exercises aimed at *overcoming misconceptions*. It has become increasingly clear from our own teaching experience that students often struggle with chemistry because they misunderstand many of the fundamental concepts. In this text, we have gone to great lengths to provide illustrations and explanations aimed at giving students a more accurate picture of the fundamental ideas of chemistry. In particular, we have attempted to represent the microscopic world of chemistry so that students have a picture in their minds of “what the atoms and molecules are doing.” The art program along with the animations emphasize this goal. We have also placed a larger emphasis on the qualitative understanding of concepts before quantitative problems are considered. Because using an algorithm to correctly solve a problem often masks misunderstanding—when students assume they understand the material because they got the right “answer”—it is important to probe their understanding in other ways. In this vein, the text includes a number of *Active Learning Questions* at the end of each chapter that are intended for group discussion. It is our experience that students often learn the most when they teach each other. Students are forced to recognize their own lack of understanding when they try and fail to explain a concept to another student.

With a strong problem-solving orientation, this text talks to students about how to approach and solve chemical problems. We emphasize a thoughtful, logical approach rather than simply memorizing procedures. This approach is thoroughly described in Section 5.3 (Learning to Solve Problems), which promotes the importance of thoughtful, creative problem solving. This section emphasizes to students that thinking through a problem produces more long-term, meaningful learning that can be applied to “real life” than memorizing steps that apply only to a particular type of problem. To help students adopt this way of thinking we have organized the problem-solving process in terms of:

- › Where are we going?
- › What do we know?
- › How do we get there?
- › Reality check, which prompts students to check whether their answer makes sense

As we proceed in the text, we gradually shift more responsibility to the students to think through the examples so that they do not become overly dependent on our help.

One of the characteristics of this text is an innovative method for dealing with acid–base equilibria, the material the typical student finds most difficult and frustrating. The key to this approach involves first deciding what species are present in solution, then thinking about the chemical properties of these species. This method provides a general framework for approaching all types of solution equilibria.

- › The text contains almost 250 *Examples*, with more given in the text discussions, to illustrate general problem-solving strategies. When a specific strategy is presented, it is summarized in a Problem-Solving Strategy box and the *Example* that follows it reinforces the use of the strategy to solve the problem. In general, we emphasize the use of conceptual understanding to solve problems rather than an algorithm-based approach. This approach is strongly reinforced by the inclusion of 204 *Interactive Examples*, which encourage students to thoughtfully consider the examples step-by-step.
- › We have presented a thorough *treatment of reactions* that occur in solution, including acid–base reactions. This material appears in Chapter 6, “Types of Chemical Reactions and Solution Stoichiometry,” directly after the chapter on chemical stoichiometry, to emphasize the connection between solution reactions and chemical reactions in general. Chapter 6 also includes oxidation–reduction reactions and balancing by oxidation state, because a large number of interesting and important chemical reactions involve redox processes.
- › *Descriptive chemistry* and chemical principles are thoroughly integrated in this text. Chemical models may appear sterile and confusing without the observations that stimulated their invention. On the other hand, facts without organizing principles may seem overwhelming. A combination of observation and models can make chemistry both interesting and understandable. In the chapter on the chemistry of the elements we have used tables and charts to show how properties and models correlate. Descriptive chemistry is presented in a variety of ways—as applications of principles in separate sections, in *Examples* and exercises, in photographs, and in *Chemical Connections*.
- › Throughout the book a strong *emphasis on models* prevails. Coverage includes how they are constructed, how they are tested, and what we learn when they inevitably fail. Models are developed naturally, with pertinent observation always presented first to show why a particular model was invented.
- › *Chemical Connections* boxes present applications of chemistry in various fields and in our daily lives. Margin notes in the *Instructor’s Annotated Edition* also highlight

many more *Chemical Connections* available on the student Web site.

- › We offer end-of-chapter exercises for every type of student and for every kind of homework assignment: questions that promote group learning, exercises that reinforce student understanding, and problems that present the ultimate challenge with increased rigor and by integrating multiple concepts. To further encourage this approach we have included a selection of ChemWork Problems in the text. These multiconcept problems (and additional ones) are found interactively online with the same type of assistance a student would get from an instructor. We have also included biochemistry problems to make the connection for students in the course who are not chemistry majors.
- › Judging from the favorable comments of instructors and students who have used our books, the text seems to work very well in a variety of courses. We are especially pleased that *readability* is cited as a key strength when students are asked to assess our textbooks.

Supporting Materials

Please visit <https://www.cengage.com/c/chemistry-an-atoms-first-approach-3e-zumdahl/> for information about student and instructor resources for this book and about custom versions.

Acknowledgments

The successful completion of the third edition of *Chemistry: An Atoms First Approach* is due to the efforts of many talented and dedicated people. We especially want to thank Maureen McLaughlin, Senior Product Manager, for her support and oversight of the project. We also want to thank Peter McGahey, Learning Designer, and Theresa Dearborn, In-House Subject Matter Expert for their knowledge of the field and helpful input. Special thanks as well go to Vanessa Desiato, Associate Content Manager, for her ideas and attention to detail. Special thanks also go to Brendan Killion, who came aboard as Senior Content Manager and was able to see this edition to its successful completion with his expertise and efficiency.

We are especially grateful to Tom Hummel, University of Illinois at Urbana–Champaign, who managed the end-of-chapter problems and the solutions manuals. We greatly appreciate the outstanding work of Gretchen Adams, University of Illinois at Urbana–Champaign, who created the *Interactive Examples* and who converted many of the end-of-chapter problems to ChemWork problems. We are deeply indebted to these two colleagues who contribute in so many ways to our various projects. We owe a large debt of appreciation to

Professor Tom McGrath of Baylor University for providing a great deal of useful insight as we prepared the third edition. We followed many of Professor McGrath's suggestions as we revised the text.

Thanks to others who supplied valuable assistance for this text: Nellie Mitchell, Product Assistant; Lizz Anderson, Designer; Timothy Cali, Marketing Manager; Nicole Hurst, Senior Market Development Manager; Beth McCracken, Digital Delivery Lead; and Ann Hoffman, Intellectual Property Analyst.

We are especially thankful to all of the reviewers who participated in different aspects of the development process from reviewing chapters to providing feedback on the development of new features. We sincerely appreciate all of these suggestions.

Reviewers

Third Edition Reviewers

Robin Bond, *The Evergreen State College*

Sujata Guha, *Tennessee State University*

Ryan Hayes, *Andrews University*

Jeremy Kua, *University of San Diego*

Craig P. McClure, *Drexel University*

Douglas R. Mulford, *Emory University*

Llanie Nobile, *Marquette University*

Van Quach, *Seminole State College*

Ashely Corrigan Steffey, *University of San Diego*

Shuai Sun, *University of Kansas*

Second Edition Reviewers

David Boatright, *University of West Georgia*

Nancy Breen, *Roger Williams University*

Elzbieta Cook, *Louisiana State University*

Tracy McGill, *Emory University*

Douglas Mulford, *Emory University*

Edith Osborne, *Angelo State University*

Yasmin Patell, *Kansas State University*

Jerry Poteat, *Georgia Perimeter College*

Michael Sakuta, *Georgia Perimeter College*

Jeff Statler, *University of Utah*

Shane Street, *University of Alabama*

Dennis Taylor, *Clemson University*

Joshua Wallach, *Old Dominion University*

First Edition Reviewers

David Boatright, *University of West Georgia*

Gene Carlisle, *West Texas A&M University*

Tsun-Mei Chang, *University of Wisconsin–Parkside*

Charles Cox, *Georgia Institute of Technology*

Michelle Driessen, *University of Minnesota*

Amina El-Ashmawy, *Collin College*

Deborah Exton, *University of Oregon*

David Geiger, *SUNY-College at Geneseo*

Jon Hardesty, *Collin County Community College*

Jennie Mayer, *Bellevue College*

Katherine Parks, *Motlow State Community College*

Casey Raymond, *SUNY Oswego*

David Smith, *New Mexico State University*

Lydia Tien, *Monroe Community College*

Wayne Wesolowski, *University of Arizona*

David Zax, *Cornell University*

To the Student

As you jump into the study of chemistry we hope that you will find our text helpful and interesting. Our job is to present the concepts and ideas of chemistry in a way you can understand. We hope to encourage you in your studies and to help you learn to solve problems in ways you can apply in all areas of your professional and personal lives.

Our main goal is to help you learn to become a truly creative problem solver. Our world badly needs people who can “think outside the box.” Our focus is to help you learn to think like a chemist. Why would you want to do that? Chemists are great problem solvers. They use logic, trial and error, and intuition—along with lots of patience—to work through complex problems. Chemists make mistakes, as we all do in our lives. The important thing that a chemist does is to learn from the mistakes and to try again. This “can do” attitude is useful in all careers.

In this book we develop the concepts in a natural way: The observations come first and then we develop models to explain the observed behavior. Models help us to understand and explain our world. They are central to scientific thinking. Models are very useful, but they also have limitations, which we will point out. By understanding the basic concepts in chemistry we lay the foundation for solving problems.

Our main goal is to help you learn a thoughtful method of problem solving. True learning is more than memorizing facts. Truly educated people use their factual knowledge as a starting point—a basis for creative problem solving. Our strategy for solving problems is explained in Section 5.3. To solve a problem we ask ourselves questions, which help us think through the problem. We let the problem guide us to the solution. This process can be applied to all types of problems in all areas of life.

As you study the text, use the *Examples* and the problem-solving strategies to help you. The strategies are boxed to highlight them for you, and the *Examples* show how these strategies are applied.

After you have read and studied each chapter of the text you’ll need to practice your problem-solving skills. To do this we have provided plenty of review questions and end-of-chapter exercises. Your instructor may assign these on paper or online; in either case, you’ll want to work with your fellow students. One of the most effective ways to learn chemistry is through the exchange of ideas that comes from helping one another. The online homework assignments will give you instant feedback, and, in print, we have provided answers to some of the exercises in the back of the text. In all cases, your main goal is not just to get the correct answer, but to understand the process for getting the answer. Memorizing solutions for specific problems is not a very good way to prepare for an exam (or to solve problems in the real world!).

To become a great problem solver you’ll need these skills:

1. Look within the problem for the solution. (Let the problem guide you.)
2. Use the concepts you have learned along with a systematic, logical approach to find the solution.
3. Solve the problem by asking questions and learn to trust yourself to think it out.

You will make mistakes, but the important thing is to learn from these errors. The only way to gain confidence is to practice, practice, practice and to use your mistakes to find your weaknesses. Be patient with yourself and work hard to understand rather than simply memorize.

We hope you’ll have an interesting and successful year learning to think like a chemist!

Steve and Susan Zumdahl and Don DeCoste

Connecting To Atoms This new boxed feature contains atoms first summaries that cover key atoms first concepts at point-of-use.

Chemical Connections Interesting applications of modern chemistry show students the relevance of chemistry to the real world.

Connecting To Atoms is a new boxed feature that contains summaries that cover key atoms first concepts at point-of-use. They have visuals and text to help students understand how an atomic/molecular approach clarifies chemical concepts. Examples of topics include Atoms, Ions, and Isotopes: A Pictorial Summary, Lewis Structures and the Periodic Table, Polar Molecules—It's All About Symmetry, and Entropy and Positional Probability.

CONNECTING TO ATOMS 11.1

Zero-, First-, and Second-Order Reactions and the Collision Model

In Section 11.4 we derived integrated rate laws for zero-, first-, and second-order reactions, and in Section 11.5 we discussed mechanisms for such reactions. Let's consider what each of these types of reactions would "look" like at the molecular level.

Let's consider the generic reaction $2AB \rightarrow AB_2 + A$, which we can represent as $\bullet\bullet + \bullet\bullet \rightarrow \bullet\bullet\bullet + \bullet$. How could various possible mechanisms lead to different rate laws for this reaction?

Possible mechanism that yields a second order rate law

As we discussed in Section 11.6, the molecules must collide in order to react. The most obvious rate law, then, would seem to be second order since we need two molecules to collide. A simple mechanism for a second order reaction could be:

- $\bullet\bullet + \bullet\bullet \rightarrow \bullet\bullet\bullet + \bullet$ (slow)
- $\bullet\bullet\bullet \rightarrow \bullet\bullet\bullet + \bullet$ (fast)

The first step is the rate-determining step, making the rate = $k[AB]^2$, or second order.

Possible mechanism that yields a zero-order rate law

Zero-order reactions are often the result of catalysis. For example, suppose the reaction occurs on the surface of a metal catalyst, such as platinum (Pt). Once the surface is completely covered with AB molecules, increasing the concentration of AB would have no effect on the rate

since only the AB molecules on the surface can react.

collisions between AB molecules only transfer energy from one AB to another. If the first and third steps in the above mechanism occur at a much faster rate than the second step, the decomposition of the excited state is the rate-determining step. Thus the second step dictates the rate and the rate law = $k[AB]$, which is first order.

In all three mechanisms, then, the molecules must collide for a reaction to occur.

Collision model

In Section 11.6, you learned about the collision model, which tells us that in order for a reaction to occur, molecules must collide. We can see that even with this simple model, we can develop different mechanisms that give rise to different rate laws. As we have discussed, however, while a collision is necessary, it is not the only factor required to result in a reaction. The molecules must be oriented correctly (which is one way in which a catalyst works to speed up a reaction), and activation energy must be overcome (which is why reactions proceed at a higher rate with increased temperature). To better understand how reactions proceed and what factors are important in the rate of a reaction, we need to take a molecular-level approach to consider how the molecules are interacting.

Note that the AB molecules must collide in order to react. However, the

CONNECTING TO ATOMS 1.1

Atoms, Ions, and Isotopes: A Pictorial Summary

Atom Number
(Number of protons)

Protons
Relative charge = +1
Relative mass = 1

Neutrons
Relative charge = 0
Relative mass = 1

Electrons
Relative charge = -1
Relative mass = 1/1840

Mean Number
Total number of protons and neutrons

Isotopes
Same number protons (same atomic number)
Different number neutrons (different mass numbers)

Line of Isotopes

Loss of electron
Sodium ion (Na^+)

Gain of electron
Chloride ion (Cl^-)

Cation
Positive charge
Number of protons > Number of electrons

Anion
Negative charge
Number of electrons > Number of protons

CHEMICAL CONNECTIONS 3.1

No Lead Pencils

If you've ever wondered why the part of a pencil that makes the mark is called the "lead," pencils have no lead in them now—and they never have. Apparently the association between writing and the element lead arose during the Roman Empire, when lead rods were used as writing utensils because they leave a gray mark on paper. Many centuries later, in 1564, a deposit of a black substance found to be very useful for writing was discovered in Borrowdale, England. This substance, originally called "black lead," was shown in 1803 by Swedish chemist Carl Scheele to be a form of carbon and was subsequently named graphite (after the Greek graphos, meaning "to write").

Originally, chunks of graphite from Borrowdale, called marking stones, were used as writing instruments. Later, sticks of graphite were used. Because graphite is brittle, the sticks needed reinforcement. At first they were wrapped in string, which was unwound as the core wore down. Eventually, graphite rods were tied between two wooden slats or inserted into hollowed-out wooden sticks to form the first crude pencils.

Although Borrowdale graphite was pure enough to use directly, most graphite must be mixed with other materials to be useful for writing instruments. In 1795, the French chemist Nicolas-Jacques Conte invented a process in which graphite is mixed with clay and water to produce pencil "lead," a recipe that is still used today. In modern pencil manufacture, graphite and clay are mixed and crushed into a fine powder to which water is added. After the gray sludge is blended for several days, it is dried, ground up again, and mixed with more water to give a gray paste. The paste is extruded through a metal tube to form thin rods, which are then cut into pencil length pieces called "leads." These leads are heated in an oven to 1000°C, until they are smooth and hard.

The ratio of clay to graphite is adjusted to vary the hardness of the lead—the more clay in the mix, the harder the lead and the lighter the line it makes. Pencils are made from a slat of wood with several grooves cut in to hold the leads. A similar grooved slat is then placed on top and glued to form a "sandwich" from which individual pencils are cut, sanded smooth, and painted.

Pencils are made from a slat of wood with several grooves cut in to hold the leads. A similar grooved slat is then placed on top and glued to form a "sandwich" from which individual pencils are cut, sanded smooth, and painted.

Pencils are very useful—especially for doing chemistry problems—because we can erase our mistakes. Most pencils used in the United States have erasers (first attached to pencils in 1858), although most European pencils do not. Lead and to end, the number of pencils made in the United States each year would circle the earth about 15 times. Pencils illustrate how useful a simple substance like graphite can be.

CHEMICAL CONNECTIONS 11.1

Enzymes: Nature's Catalysts

The most impressive examples of homogeneous catalysts occur in nature, where the complex reactions necessary for plants and animal life are made possible by enzymes. Enzymes are large molecules specifically tailored to facilitate a given type of reaction. Usually enzymes are proteins, an important class of biomolecules constructed from α -amino acids that have the general structure

where R represents any one of 20 different substituents. These amino acid molecules can be "hooked together" to form a polymer (a word meaning "many parts") called a protein. The general structure of a protein can be represented as follows:

containing one less amino acid. Without the enzymes found in human cells, this reaction would be much too slow to be useful. One of these enzymes is carboxypeptidase-A, a zinc-containing protein (Fig. 11.18). Carboxypeptidase-A captures the protein to be acted on (called the substrate) in a special groove and positions carboxypeptidase-A so that the end is in the active site, where the catalysis occurs (Fig. 11.19). Note that the Zn^{2+} ion bonds to the oxygen of the C=O (carbonyl) group. This polarizes the electron density in the carbonyl group, allowing the neighboring C-N bond to be broken much more easily. When the reaction is completed, the remaining portion of the substrate protein and the newly formed amino acid are released by the enzyme.

The process just described for carboxypeptidase-A is characteristic of the behavior of other enzymes. Enzyme catalysis can be represented by the series of reactions shown below:

$$E + S \rightleftharpoons E \cdot S \rightarrow E + P$$

where E represents the enzyme, S represents the substrate, $E \cdot S$ represents the enzyme-substrate complex, and P represents the products. The enzyme and substrate form a complex, where the reaction occurs. The enzyme then releases the product and is ready to repeat the process. The most amazing thing about enzymes is their efficiency. Because an enzyme plays its catalytic role over and over and very rapidly, only a tiny amount of enzyme is required. This makes the isolation of enzymes for study quite difficult.

Enzymes: Nature's Catalysts

FIGURE 11.17 The removal of the end amino acid from a protein by reaction with a molecule of water. The products are an amino acid and a new, smaller protein.

FIGURE 11.18 (a) The structure of the enzyme carboxypeptidase-A, which contains 301 amino acids. The zinc ion is shown above as a black sphere in the center. (b) Carboxypeptidase-A with a substrate (pink) in place.

FIGURE 11.19 Proton-substrate interaction. The substrate is shown in black and red, with the red representing the terminal amino acid. Blue indicates side chains from the enzyme that help bind the substrate.

Chemical Connections describe current applications of chemistry. These special-interest boxes cover such topics as the invention of Post-it Notes, farming the wind, and the use of iron metal to clean up contaminated groundwater. Additional *Chemical Connections* are available on the student Web site.

Conceptual Understanding Conceptual learning and problem solving are fundamental to the approach of **Chemistry**. The text gives students the tools to become critical thinkers: to ask questions, to apply rules and models, and to evaluate the outcome.

*“The **first** principles of the universe are **atoms** and empty space; everything else is merely thought to exist.”*
—Democritus

The authors’ **emphasis on modeling** (or chemical theories) throughout the text addresses the problem of rote memorization by helping students better understand and appreciate the process of scientific thinking. By stressing the limitations and uses of scientific models, the authors show students how chemists think and work.

Molecular Structure: The VSEPR Model

The structures of molecules play a very important role in determining their chemical properties. As we will see later, this is particularly important for biomolecules; a slight change in the structure of a large biomolecule can

CRITICAL THINKING We now have evidence that electron energies in the atoms are quantized. Some of this evidence is discussed in this chapter. What if energies in atoms were not quantized? What are some differences we would notice?

The text includes a number of open-ended **Critical Thinking** questions that emphasize the importance of conceptual learning. These questions are particularly useful for generating group discussion.

LET'S REVIEW A Summary of the Hydrogen Atom

- » In the quantum (wave) mechanical model, the electron is viewed as a standing wave. This representation leads to a series of wave functions (orbitals) that describe the possible energies and spatial distributions available to the electron.
- » In agreement with the Heisenberg uncertainty principle, the model cannot specify the detailed electron motions. Instead, the square of the wave function represents the probability distribution of the electron in that orbital. This allows us to picture orbitals in terms of probability distributions, or electron density maps.
- » The size of an orbital is arbitrarily defined as the surface that contains 90% of the total electron probability.
- » The hydrogen atom has many types of orbitals. In the ground state, the single electron resides in the 1s orbital. The electron can be excited to higher-energy orbitals if energy is put into the atom.

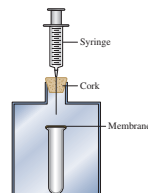
Let's Review boxes help students organize their thinking about the crucial chemical concepts that they encounter.

The text includes a number of **Active Learning Questions** at the end of each chapter that are intended for group discussion, as students often learn the most when they teach each other.

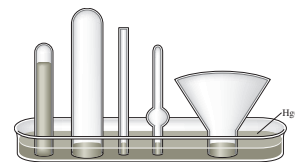
Active Learning Questions

These questions are designed to be used by groups of students in class.

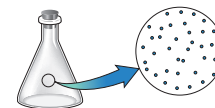
1. Consider the following apparatus: a test tube covered with a nonpermeable elastic membrane inside a container that is closed with a cork. A syringe goes through the cork.



- a. As you push down on the syringe, how does the membrane covering the test tube change?
 - b. You stop pushing the syringe but continue to hold it down. In a few seconds, what happens to the membrane?
2. Fig. 8.2 shows a picture of a barometer. Which of the following statements is the best explanation of how this barometer works?
- a. Air pressure outside the tube causes the mercury to move in the tube until the air pressure inside and outside the tube is equal.
 - b. Air pressure inside the tube causes the mercury to move in the tube until the air pressure inside and outside the tube is equal.



4. As you increase the temperature of a gas in a sealed, rigid container, what happens to the density of the gas? Would the results be the same if you did the same experiment in a container with a piston at constant pressure? (See Fig. 8.17.)
5. A diagram in a chemistry book shows a magnified view of a flask of air as follows:



- What do you suppose is between the dots (the dots represent air molecules)?
- a. air
 - b. dust
 - c. pollutants
 - d. oxygen
 - e. nothing
6. If you put a drinking straw in water, place your finger over the

Problem Solving This text talks to the student about how to approach and solve chemical problems, as one of the main goals of general chemistry is to help students become creative problem solvers. The authors emphasize a thoughtful, logical approach rather than simply memorizing procedures.

Learning to Solve Problems

One of the great rewards of studying chemistry is to become a good problem-solver. Being able to solve complex problems is a talent that will serve you well in all walks of life. It is our purpose in this text to help you learn to solve problems in a flexible, creative way based on understanding the fundamental ideas of chemistry. We call this approach **conceptual problem solving**.

The ultimate goal is to be able to solve new problems (that is, problems you have not seen before) on your own. In this text we will provide problems and offer solutions by explaining how to think about the problems. While the answers to these problems are important, it is perhaps even more important to understand the process—the thinking necessary to get the answer. Although at first we will be solving the problem for you, do not take a passive role. While studying the solution, it is crucial that you interactively think through the problem with us. Do not skip the discussion and jump to the answer. Usually, the solution will involve asking a series of questions. Make sure that you understand each step in the process. This active approach should apply to problems outside of chemistry as well. For example, imagine riding with someone in a car to an unfamiliar destination. If your goal is simply to have the other person get you to that destination, you will probably not pay much attention to how to get there (passive), and if you have to find this same place in the future on your own, you will probably not be able to do it. If, however, your goal is to learn how to get there, you would pay attention to distances, signs, and turns (active). This is how you should read the solutions in the text (and the text in general).

While actively studying our solutions to problems is helpful, at some point you will need to know how to think through these problems on your own. If we help you too much as you solve a problem, you won't really learn effectively. If we always "drive," you won't interact as meaningfully with the material. Eventually you need to learn to drive yourself. We will provide more help at the beginning of the text and less as we proceed to later chapters.

There are two fundamentally different ways you might use to approach a problem. One way emphasizes memorization. We might call this the "pigeonholing method." In this approach, the first step is to label the problem—to decide in which pigeonhole it fits. The pigeonholing method requires that we provide you with a set of steps that you memorize and store in the appropriate slot for each different problem you encounter. The difficulty with this method is that it requires a new pigeonhole each time a problem is changed by even a small amount.

Consider the driving analogy again. Suppose you have memorized how to drive from your house to the grocery store. Do you know how to drive back from the grocery store to your house? Not necessarily. If you have only memorized the directions and do not understand fundamental principles such as "I traveled north to get to the store,

In Chapter 5, "Stoichiometry," the authors dedicate a section, **Learning to Solve Problems**, that emphasizes the importance of problem solving. This section helps students understand that thinking their way through a problem produces more long-term, meaningful learning than simply memorizing steps, which are soon forgotten.

Chapters 1–8 introduce a series of questions into the in-chapter **Examples** to engage students in the process of problem solving, such as **Where are we going?** and **How do we get there?** This more active approach helps students think their way through the solution to the problem.

EXAMPLE 5.1 The Average Mass of an Element

When a sample of natural copper is vaporized and injected into a mass spectrometer, the results shown in Fig. 5.3 are obtained. Use these data to compute the average mass of natural copper. (The mass values for ^{63}Cu and ^{65}Cu are 62.93 u and 64.93 u, respectively.)

SOLUTION

Where are we going?
To calculate the average mass of natural copper

What do we know?

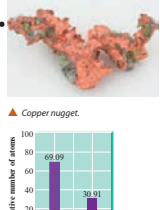
- > ^{63}Cu mass = 62.93 u
- > ^{65}Cu mass = 64.93 u

How do we get there?

As shown by the graph, of every 100 atoms of natural copper, 69.09 are ^{63}Cu and 30.91 are ^{65}Cu . Thus the mass of 100 atoms of natural copper is

$$(69.09 \text{ atoms}) \left(62.93 \frac{\text{u}}{\text{atom}} \right) + (30.91 \text{ atoms}) \left(64.93 \frac{\text{u}}{\text{atom}} \right) = 6355 \text{ u}$$

The average mass of a copper atom is

$$\frac{6355 \text{ u}}{100 \text{ atoms}} = 63.55 \text{ u/atom}$$


INTERACTIVE EXAMPLE 5.5 Calculating the Number of Moles and Mass

Cobalt (Co) is a metal that is added to steel to improve its resistance to corrosion. Calculate both the number of moles in a sample of cobalt containing 5.00×10^{23} atoms and the mass of the sample.

SOLUTION

Where are we going?
To calculate the number of moles and the mass of a sample of Co

What do we know?

- > Sample contains 5.00×10^{23} atoms of Co

How do we get there?

Note that the sample of 5.00×10^{23} atoms of cobalt is less than 1 mole (6.022×10^{23} atoms) of cobalt. What fraction of a mole it represents can be determined as follows:


$$5.00 \times 10^{23} \text{ atoms Co} \times \frac{1 \text{ mol Co}}{6.022 \times 10^{23} \text{ atoms Co}} = 8.30 \times 10^{-4} \text{ mol Co}$$

Since the mass of 1 mole of cobalt atoms is 58.93 g, the mass of 5.00×10^{23} atoms can be determined as follows:

$$8.30 \times 10^{-4} \text{ mol Co} \times \frac{58.93 \text{ g Co}}{1 \text{ mol Co}} = 4.89 \times 10^{-2} \text{ g Co}$$

Reality Check In this case the sample contains 5×10^{23} atoms, which is approximately 1/1000 of a mole. Thus the sample should have a mass of about (1/1000) (58.93) = 0.06. Our answer of ~0.05 makes sense.

See Exercise 5.48



Interactive Examples engage students in the problem-solving process by requiring them to think through the example step-by-step rather than simply scanning the written example in the text as many students do.

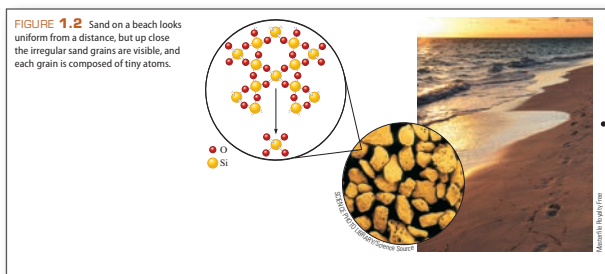
PROBLEM-SOLVING STRATEGY

Steps to Apply the VSEPR Model

1. Draw the Lewis structure for the molecule.
2. Count the electron pairs and arrange them in the way that minimizes repulsion (that is, put the pairs as far apart as possible).
3. Determine the positions of the atoms from the way the electron pairs are shared.
4. Determine the name of the molecular structure from the positions of the atoms.

Problem-Solving Strategy boxes focus students' attention on the very important process of problem solving.

Dynamic Art Program Most of the glassware, orbitals, graphs, flowcharts, and molecules have been redrawn to better serve visual learners and enhance the text book.



The art program emphasizes molecular-level interactions that help students visualize the “micro/macro” connection.

Realistic drawings of glassware and instrumentation found in the lab help students make real connections.

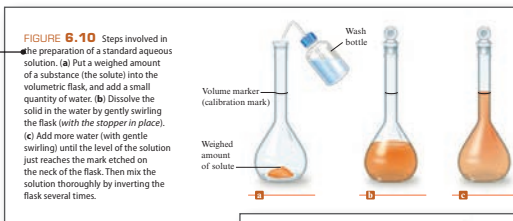


FIGURE 6.10 Steps involved in the preparation of a standard aqueous solution. (a) Put a weighed amount of a substance (the solute) into the volumetric flask, and add a small quantity of water. (b) Dissolve the solid in the water by gently swirling the flask (with the stopper in place). (c) Add more water (with gentle swirling) until the level of the solution just reaches the mark etched on the neck of the flask. Then mix the solution thoroughly by inverting the flask several times.

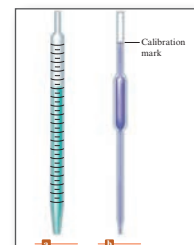


FIGURE 6.11 (a) A measuring pipet is graduated and can be used to measure various volumes of liquid accurately. (b) A volumetric (transfer) pipet is designed to measure one volume accurately. When filled to the mark, it delivers the volume indicated on the pipet.

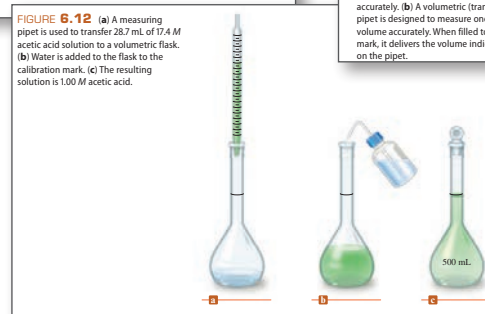


FIGURE 6.12 (a) A measuring pipet is used to transfer 28.7 mL of 17.4 M acetic acid solution to a volumetric flask. (b) Water is added to the flask to the calibration mark. (c) The resulting solution is 1.00 M acetic acid.

TABLE 9.3 Intermolecular Forces				
Type of Interaction	Nature of Attraction	Model	Range of Energies (kJ/mol)	Examples
Ion-dipole	Attraction between charge on anion and end of dipole with opposite partial charge		35–400	
Dipole-dipole	Attraction between opposite partially charged ends of polar molecules		5–30	
Hydrogen bond	Attraction between partially positively charged H atom attached to a highly electronegative atom and a lone pair on another atom	 where \bullet is N, O, or F and \circ is an atom with a lone pair of electrons	10–40	
London dispersion	Attraction between an instantaneous dipole and a dipole induced in a neighboring atom or molecule		<1–40	

Tables containing visuals help students understand and compare intramolecular and intermolecular forces between ions, atoms, and molecules.

For Review

Key Terms

Section 14.1
common ion
common ion effect

Section 14.2
buffered solution
Henderson-Hasselbalch
equation

Section 14.3
buffering capacity

Section 14.4
pH curve (titration curve)
millimole (mmol)
equivalence point
(stoichiometric point)

Section 14.5
acid-base indicator
phenolphthalein

Buffered Solutions

- Contains a weak acid (HA) and its salt (NaA) or a weak base (B) and its salt (BHCl)
- Resists a change in its pH when H^+ or OH^- is added
- For a buffered solution containing HA and A^-
 - The Henderson-Hasselbalch equation is useful:

$$pH = pK_a + \log \left(\frac{[A^-]}{[HA]} \right)$$

- The capacity of the buffered solution depends on the amounts of HA and A^- present
- The most efficient buffering occurs when the $\frac{[A^-]}{[HA]}$ ratio is close to 1
- Buffering works because the amounts of HA (which reacts with added OH^-) and A^- (which reacts with added H^+) are large enough that the $\frac{[A^-]}{[HA]}$ ratio does not change significantly when strong acids or bases are added

Acid-Base Titrations

- The progress of a titration is represented by plotting the pH of the solution versus the volume of added titrant; the resulting graph is called a pH curve or titration curve
- Strong acid-strong base titrations show a sharp change in pH near the equivalence point
- The shape of the pH curve for a strong base-strong acid titration before the equivalence point is quite different from the shape of the pH curve for a strong base-weak acid titration
 - The strong base-weak acid pH curve shows the effects of buffering before the equivalence point
 - For a strong base-weak acid titration, the pH is greater than 7 at the equivalence point because of the basic properties of A^-
- Indicators are sometimes used to mark the equivalence point of an acid-base titration
 - The end point is where the indicator changes color
 - The goal is to have the end point and the equivalence point be as close as possible

Review Questions

Answers to the Review Questions can be found on the Student Web site (accessible from www.cengagebrain.com).

1. What is meant by the presence of a common ion? How does the presence of a common ion affect an equilibrium such as



What is an acid-base solution called that contains a common ion?

2. Define a buffer solution. What makes up a buffer solution? How do buffers absorb added H^+ or OH^- with little pH change?

Is it necessary that the concentrations of the weak acid and the weak base in a buffered solution be equal? Explain. What is the pH of a buffer when the weak acid and conjugate base concentrations are equal?

A buffer generally contains a weak acid and its weak conjugate base, or a weak base and its weak conjugate acid, in water. You can solve for the pH by setting up the equilibrium problem using the K_a reaction of the weak acid or the K_b reaction of the conjugate base. Both reactions give the same answer for the pH of the solution. Explain.

Each chapter has a **For Review** section to reinforce key concepts and includes review questions for students to practice independently.

Active Learning Questions are designed to promote discussion among groups of students in class.

140 CHAPTER 3 Bonding: General Concepts

the bonding in the substance. Why do we draw resonance structures?

7. Define formal charge and explain how to calculate it. What is the purpose of the formal charge? Organic compounds are composed mostly of carbon and hydrogen but also may have oxygen, nitrogen, and/or halogens in the formula. Formal charge arguments work very well for organic compounds when drawing the best Lewis structure. How do C, H, N, O, and Cl satisfy the

octet rule in organic compounds so as to have a formula charge of zero?

8. The compounds $AlCl_3$, $CrCl_3$, and ICl_3 have similar formulas, yet each follows a different set of rules to name it. Name these compounds, and then compare and contrast the nomenclature rules used in each case.
9. How would you name $HBrO_4$, KIO_3 , $NaBrO_2$, and HIO ? Refer to Table 3.5 and the acid nomenclature discussion in the text.

Active Learning Questions

These questions are designed to be used by groups of students in class.

- Explain the electronegativity trends across a row and down a column of the periodic table. Compare these trends with those of ionization energies and atomic radii. How are they related?
- Using only the periodic table, predict the most stable ion for Na, Mg, Al, S, Cl, K, Ca, and Ga. Arrange these from largest to smallest radius, and explain why the radius varies as it does. Compare your predictions with Fig. 3.5.
- Which has the greater bond lengths: NO_2^- or NO_3^- ? Explain.
- The following ions are best described with resonance structures. Draw the resonance structures, and using formal charge arguments, predict the best Lewis structure for each ion.
 - NCO^-
 - CNO^-
- What is meant by a chemical bond? Why do atoms form bonds with each other? Why do some elements exist as molecules in nature instead of as free atoms?
- Why are some bonds ionic and some covalent?
- How does a bond between Na and Cl differ from a bond between C and O? What about a bond between N and N?
- Does a Lewis structure tell which electrons come from which atoms? Explain.
- Evaluate each of the following as an acceptable name for water:
 - dihydrogen oxide
 - hydroxide hydride
 - hydrogen hydroxide
 - oxygen dihydride
- Why do we call $Ba(NO_3)_2$ barium nitrate, but we call $Fe(NO_3)_2$ iron(II) nitrate?
- Why is calcium dichloride not the correct systematic name for $CaCl_2$?
- The common name for NH_3 is ammonia. What would be the systematic name for NH_3 ? Support your answer.

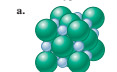
A blue question or exercise number indicates that the answer to that question or exercise appears at the back of this book and a solution appears in the Student Solutions Manual.

14. The following electrostatic potential diagrams represent H_2 , HCl , or $NaCl$. Label each and explain your choices.



15. Describe the type of bonding that exists in the $Cl_2(g)$ molecule. How does this type of bonding differ from that found in the $HCl(g)$ molecule? How is it similar?
16. Some of the important properties of ionic compounds are as follows:
 - low electrical conductivity as solids and high conductivity in solution or when molten
 - relatively high melting and boiling points
 - brittleness
 - solubility in polar solvents
 How does the concept of ionic bonding discussed in this chapter account for these properties?

17. Label the type of bonding for each of the following.

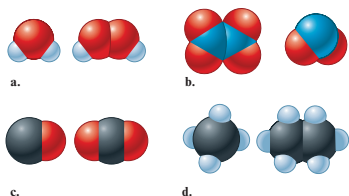


18. Distinguish between the following terms.
 - molecule versus ion
 - covalent bonding versus ionic bonding

Comprehensive End-of-Chapter Practice and Review

Questions

23. Reference Section 5.2 to find the atomic masses of ^{12}C and ^{13}C , the relative abundance of ^{12}C and ^{13}C in natural carbon, and the average mass (in u) of a carbon atom. If you had a sample of natural carbon containing exactly 10,000 atoms, determine the number of ^{12}C and ^{13}C atoms present. What would be the average mass (in u) and the total mass (in u) of the carbon atoms in this 10,000-atom sample? If you had a sample of natural carbon containing 6.0221×10^{23} atoms, determine the number of ^{12}C and ^{13}C atoms present. What would be the average mass (in u) and the total mass (in u) of this 6.0221×10^{23} atom sample? Given that $1 \text{ g} = 6.0221 \times 10^{23} \text{ u}$, what is the total mass of 1 mole of natural carbon in units of grams?
24. Avogadro's number, molar mass, and the chemical formula of a compound are three useful conversion factors. What unit conversions can be accomplished using these conversion factors?
25. If you had a mole of U.S. dollar bills and equally distributed the money to all of the people of the world, how rich would every person be? Assume a world population of 7 billion.
26. Describe 1 mole of CO_2 in as many ways as you can.
27. Which of the following compounds have the same empirical formulas?



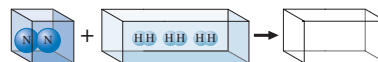
28. What is the difference between the molar mass and the empirical formula mass of a compound? When are these masses the same, and when are they different? When different, how is the molar mass related to the empirical formula mass?
29. How is the mass percent of elements in a compound different for a 1.0-g sample versus a 100.-g sample versus a 1-mole sample of the compound?
30. A balanced chemical equation contains a large amount of information. What information is given in a balanced equation?
31. The reaction of an element X with element Y is represented in the following diagram. Which of the equations best describes this reaction?

Questions are homework problems directed at concepts within the chapter and in general don't require calculation.

52 CHAPTER 1 Chemical Foundations

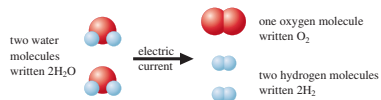
17. The vitamin niacin (nicotinic acid, $\text{C}_6\text{H}_5\text{NO}_2$) can be isolated from a variety of natural sources such as liver, yeast, milk, and whole grain. It also can be synthesized from commercially available materials. From a nutritional point of view, which source of nicotinic acid is best for use in a multivitamin tablet? Why?
18. Section 1.4 describes the postulates of Dalton's atomic theory. With some modifications, these postulates hold up very well regarding how we view elements, compounds, and chemical reactions today. Answer the following questions concerning Dalton's atomic theory and the modifications made today.
- The atom can be broken down into smaller parts. What are the smaller parts?
 - How are atoms of hydrogen identical to each other and how can they be different from each other?
 - How are atoms of hydrogen different from atoms of helium? How can H atoms be similar to He atoms?
 - How is water different from hydrogen peroxide (H_2O_2) even though both compounds are composed of only hydrogen and oxygen?
 - What happens in a chemical reaction and why is mass conserved in a chemical reaction?
19. The contributions of J. J. Thomson and Ernest Rutherford led the way to today's understanding of the structure of the atom. What were their contributions?
20. What is the modern view of the structure of the atom?
21. The number of protons in an atom determines the identity of the atom. What do the number and arrangement of the electrons in an atom determine? What does the number of neutrons in an atom determine?
22. If the volume of a proton is similar to the volume of an electron, how will the densities of these two particles compare to each other?
23. For lighter, stable isotopes, the ratio of the mass number to the atomic number is close to a certain value. What is the value? What happens to the value of the mass number to atomic number ratio as stable isotopes become heavier?
24. What refinements had to be made in Dalton's atomic theory to account for Gay-Lussac's results on the combining volumes of gases?

26. Observations of the reaction between nitrogen gas and hydrogen gas show us that 1 volume of nitrogen reacts with 3 volumes of hydrogen to make 2 volumes of gaseous product, as shown below:



Determine the formula of the product and justify your answer.

27. A sample of chloroform is found to contain 12.0 g of carbon, 106.4 g of chlorine, and 1.01 g of hydrogen. If a second sample of chloroform is found to contain 30.0 g of carbon, what is the total mass of chloroform in the second sample?
28. A sample of H_2SO_4 contains 2.02 g of hydrogen, 32.07 g of sulfur, and 64.00 g of oxygen. How many grams of sulfur and grams of oxygen are present in a second sample of H_2SO_4 containing 7.27 g of hydrogen?
29. In Section 1.1 of the text, the concept of a chemical reaction was introduced with the example of the decomposition of water, represented as follows:



Use ideas from Dalton's atomic theory to explain how the above representation illustrates the law of conservation of mass.

30. In a combustion reaction, 46.0 g of ethanol reacts with 96.0 g of oxygen to produce water and carbon dioxide. If 54.0 g of water is produced, what mass of carbon dioxide is produced?
31. Early tables of atomic weights (masses) were generated by measuring the mass of a substance that reacts with 1.00 g of oxygen. Given the following data and taking the atomic mass of hydrogen as 1.00, generate a table of relative atomic masses for oxygen, sodium, and magnesium.

Element	Mass That Combines with 1.00 g Oxygen	Assumed Formula
Hydrogen	0.126 g	HO
Sodium	2.875 g	NaO
Magnesium	1.500 g	MgO

32. Indium oxide contains 4.784 g of indium for every 1.000 g of oxygen. In 1869, when Mendeleev first presented his version of the periodic table, he proposed the formula In_2O_3 for indium oxide. Before that time it was thought that the formula was InO . What values for the atomic mass of indium are obtained using these two formulas? Assume that oxygen has an atomic mass of 16.00.

Exercises

In this section, similar exercises are paired.

Development of the Atomic Theory

25. When mixtures of gaseous H_2 and gaseous Cl_2 react, a product forms that has the same properties regardless of the relative amounts of H_2 and Cl_2 used.
- How is this result interpreted in terms of the law of definite proportion?
 - When a volume of H_2 reacts with an equal volume of Cl_2 at the same temperature and pressure, what volume of product having the formula HCl is formed?

There are numerous **Exercises** to reinforce students' understanding of each section. These problems are paired and organized by topic so that instructors can review them in class and assign them for homework.

Wealth of End-of-Chapter Problems The text offers an unparalleled variety of end-of-chapter content with problems that increase in rigor and integrate multiple concepts.

New **ChemWork** end-of-chapter multiconcept problems, with many additional problems, are found interactively online with the same type of assistance a student would get from an instructor. These problems are available to assign online for more practice.

ChemWork Problems

These multiconcept problems (and additional ones) are found interactively online with the same type of assistance a student would get from an instructor.

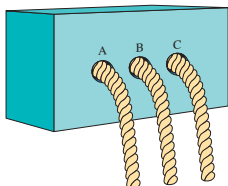
49. Complete the following table, including the mass number and the atomic number with the symbol for the isotope.

Number of Protons	Number of Neutrons	Symbol
9	10	
13	14	
53	74	
34	45	
16	16	

57. In a reaction, 34.0 g of chromium(III) oxide reacts with 12.1 g of aluminum to produce chromium and aluminum oxide. If 23.3 g of chromium is produced, what mass of aluminum oxide is produced?
58. Using the information in Table 1.1, answer the following questions. In an ion with an unknown charge, the total mass of all the electrons was determined to be 2.55×10^{-26} g, while the total mass of its protons was 5.34×10^{-23} g. What is the identity and charge of this ion? What is the symbol and mass number of a neutral atom whose total mass of its electrons is 3.92×10^{-26} g, while its neutrons have a mass of 9.35×10^{-23} g?
59. A single molecule has a mass of 7.31×10^{-23} g. Provide an example of a real molecule that can have this mass. Assume the elements that make up the molecule are made of light isotopes where the number of protons equals the number of neutrons in the nucleus of each element.

Challenge Problems

60. Confronted with the box shown in the diagram, you wish to discover something about its internal workings. You have no tools and cannot open the box. You pull on rope B, and it moves rather freely. When you pull on rope A, rope C appears to be pulled slightly into the box. When you pull on rope C, rope A almost disappears into the box.



- a. Based on these observations, construct a model for the interior mechanism of the box.
- b. What further experiments could you do to refine your model?

Challenge Problems take students one step further and challenge them more rigorously than the Additional Exercises.

Marathon Problem

This problem is designed to incorporate several concepts and techniques into one situation.

66. You have gone back in time and are working with Dalton on a table of relative masses. Following are his data.
- 0.602 g gas A reacts with 0.295 g gas B
 0.172 g gas B reacts with 0.401 g gas C
 0.320 g gas A reacts with 0.374 g gas C
- a. Assuming simplest formulas (AB, BC, and AC), construct a table of relative masses for Dalton.
- b. Knowing some history of chemistry, you tell Dalton that if he determines the volumes of the gases reacted at constant temperature and pressure, he need not assume simplest formulas. You collect the following data:
- 6 volumes gas A + 1 volume gas B → 4 volumes product
 1 volume gas B + 4 volumes gas C → 4 volumes product
 3 volumes gas A + 2 volumes gas C → 6 volumes product

Marathon Problems also combine concepts from multiple chapters; they are the most challenging problems in the end-of-chapter material.

About the Authors



Steven S. Zumdahl earned a B.S. in Chemistry from Wheaton College (IL) and a Ph.D. from the University of Illinois, Urbana–Champaign. He has been a faculty member at the University of Colorado–Boulder, Parkland College (IL), and the University of Illinois at Urbana–Champaign (UIUC), where he is Professor Emeritus. He has received numerous awards, including the National Catalyst Award for Excellence in Chemical Education, the University of Illinois Teaching Award, the UIUC Liberal Arts and Sciences Award for Excellence in Teaching, UIUC Liberal Arts and Sciences Advising Award, and the School of Chemical Sciences Teaching Award (five times). He is the author of several chemistry textbooks. In his leisure time he enjoys traveling and collecting classic cars.



Susan A. Zumdahl earned a B.S. and M.A. in Chemistry at California State University–Fullerton. She has taught science and mathematics at all levels, including middle school, high school, community college, and university. At the University of Illinois at Urbana–Champaign, she developed a program for increasing the retention of minorities and women in science and engineering. This program focused on using active learning and peer teaching to encourage students to excel in the sciences. She has coordinated and led workshops and programs for science teachers from elementary through college levels. These programs encourage and support active learning and creative techniques for teaching science. For several years she was director of an Institute for Chemical Education (ICE) field center in Southern California, and she has authored several chemistry textbooks. Susan spearheaded the development of a sophisticated Web-based electronic homework system for teaching chemistry. She enjoys traveling, classic cars, and gardening in her spare time—when she is not playing with her grandchildren.



Donald J. DeCoste is the Associate Director of General Chemistry at the University of Illinois, Urbana–Champaign and has been teaching chemistry at the high school and college levels for 30 years. He earned his B.S. in Chemistry and Ph.D. from the University of Illinois, Urbana–Champaign. At Illinois, he teaches courses in general chemistry and the teaching of chemistry. He has also developed chemistry courses for non-science majors, preservice secondary teachers, and preservice elementary/middle school teachers. He has received the LAS Award for Excellence in Undergraduate Teaching, the Provost’s Excellence in Undergraduate Teaching Award, and the School of Chemistry Sciences Teaching Award (five times). Don has led workshops for secondary teachers and graduate student teaching assistants, discussing the methods and benefits of getting students more actively involved in class. When not involved in teaching and advising, Don enjoys spending time with his wife and three children.



REVIEW

Glassware used to measure volumes of liquids. (Alexander Raths/Shutterstock.com)

Measurement and Calculations in Chemistry

R.1 Units of Measurement

R.2 Uncertainty in Measurement

Precision and Accuracy

R.3 Significant Figures and Calculations

R.4 Learning to Solve Problems Systematically

R.5 Dimensional Analysis

R.6 Temperature

R.7 Density

R.8 Classification of Matter

R.9 Energy

Chemical Energy

R.10 The Mole

Making observations is fundamental to all science. These observations can be qualitative or quantitative. A quantitative observation is called a **measurement**, which always has two parts: a **number** and a scale (called a **unit**). Both parts must be present for a measurement to be meaningful. A qualitative observation does not involve a number. Examples of qualitative observations are “the substance is blue” and “the sun is very hot.”

In this chapter we will discuss measurements in detail and explain the various properties of the numbers and units associated with measurements. This material should be familiar to you from previous science courses, but we include it here to provide a review of these topics that are critical to the operations of chemistry.

R.1 Units of Measurement



▲ Soda is commonly sold in 2-liter bottles—an example of the use of SI units in everyday life.

A unit such as volume that is based on a fundamental unit is called a “derived unit.”

In our study of chemistry we will use measurements of mass, length, time, temperature, electric current, and the amount of a substance, among others. Scientists recognized long ago that standard systems of units had to be adopted if measurements were to be useful. If every scientist had a different set of units, complete chaos would result. Unfortunately, different standards were adopted in different parts of the world. The two major systems are the *English system* used in the United States and the *metric system* used by most of the rest of the industrialized world. This duality causes a good deal of trouble; for example, parts as simple as bolts are not interchangeable between machines built according to the two systems. As a result, the United States has begun to adopt the metric system.

Most scientists in all countries have for many years used the metric system. In 1960, an international agreement set up a system of units called the *International System (le Système International in French)*, or the **SI system**. This system is based on the metric system and units derived from the metric system. The fundamental SI units are listed in Table R.1. We will discuss how to manipulate these units later in this chapter.

Because the fundamental units are not always convenient (expressing the mass of a pin in kilograms is awkward), prefixes are used to change the size of the unit. These are listed in Table R.2. Some common objects and their measurements in SI units are listed in Table R.3.

One physical quantity that is very important in chemistry is *volume*, which is not a fundamental SI unit but is derived from length. A cube that measures 1 meter (m) on each edge is represented in Fig. R.1. This cube has a volume of $(1\text{ m})^3 = 1\text{ m}^3$. Because there are 10 decimeters (dm) in a meter, the volume of this cube is $(1\text{ m})^3 = (10\text{ dm})^3 = 1000\text{ dm}^3$. A cubic decimeter, that is $(1\text{ dm})^3$, is commonly called a *liter (L)*, which is a unit of volume slightly larger than a quart. As shown in Fig. R.1, 1000 liters are contained in a cube with a volume of 1 cubic meter. Similarly, since 1 decimeter equals

TABLE R.1 | The Fundamental SI Units

Physical Quantity	Name of Unit	Abbreviation
Mass	kilogram	kg
Length	meter	m
Time	second	s
Temperature	kelvin	K
Electric current	ampere	A
Amount of substance	mole	mol
Luminous intensity	candela	cd

TABLE R.2 | The Prefixes Used in the SI System (Those most commonly encountered are shown in blue.)

Prefix	Symbol	Meaning	Exponential Notation*
exa	E	1,000,000,000,000,000,000	10^{18}
peta	P	1,000,000,000,000,000	10^{15}
tera	T	1,000,000,000,000	10^{12}
giga	G	1,000,000,000	10^9
mega	M	1,000,000	10^6
kilo	k	1000	10^3
hecto	h	100	10^2
deka	da	10	10^1
—	—	1	10^0
deci	d	0.1	10^{-1}
centi	c	0.01	10^{-2}
milli	m	0.001	10^{-3}
micro	μ	0.000001	10^{-6}
nano	n	0.000000001	10^{-9}
pico	p	0.000000000001	10^{-12}
femto	f	0.000000000000001	10^{-15}
atto	a	0.000000000000000001	10^{-18}

*See Appendix 1.1 if you need a review of exponential notation.

TABLE R.3 | Some Examples of Commonly Used Units

Unit	Example
Length	A dime is 1 mm thick. A quarter is 2.5 cm in diameter. The average height of an adult man is 1.8 m.
Mass	A nickel has a mass of about 5 g. A 120-lb person has a mass of about 55 kg.
Volume	A 12-oz can of soda has a volume of about 360 mL.

10 centimeters (cm), the liter can be divided into 1000 cubes each with a volume of 1 cubic centimeter:

$$1 \text{ liter} = (1 \text{ dm})^3 = (10 \text{ cm})^3 = 1000 \text{ cm}^3$$

Also, since $1 \text{ cm}^3 = 1 \text{ milliliter (mL)}$,

$$1 \text{ liter} = 1000 \text{ cm}^3 = 1000 \text{ mL}$$

Thus 1 liter contains 1000 cubic centimeters, or 1000 milliliters.

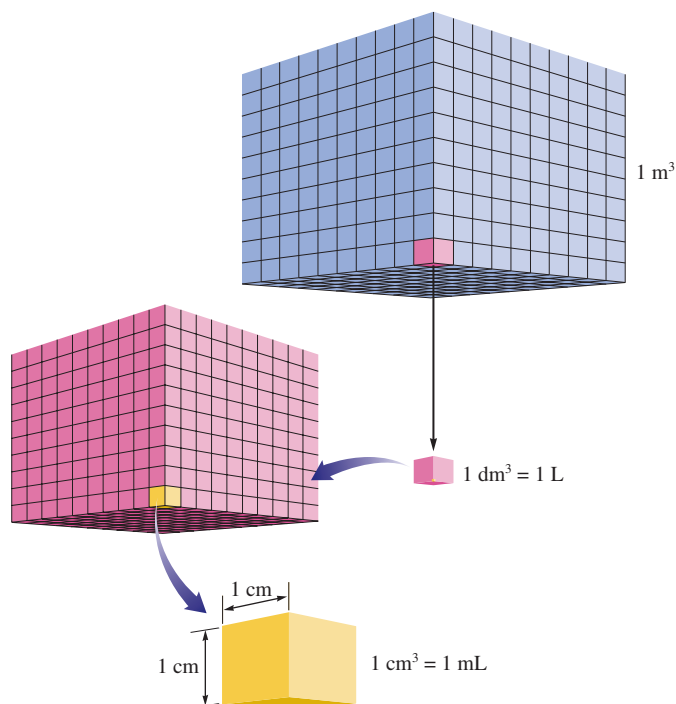
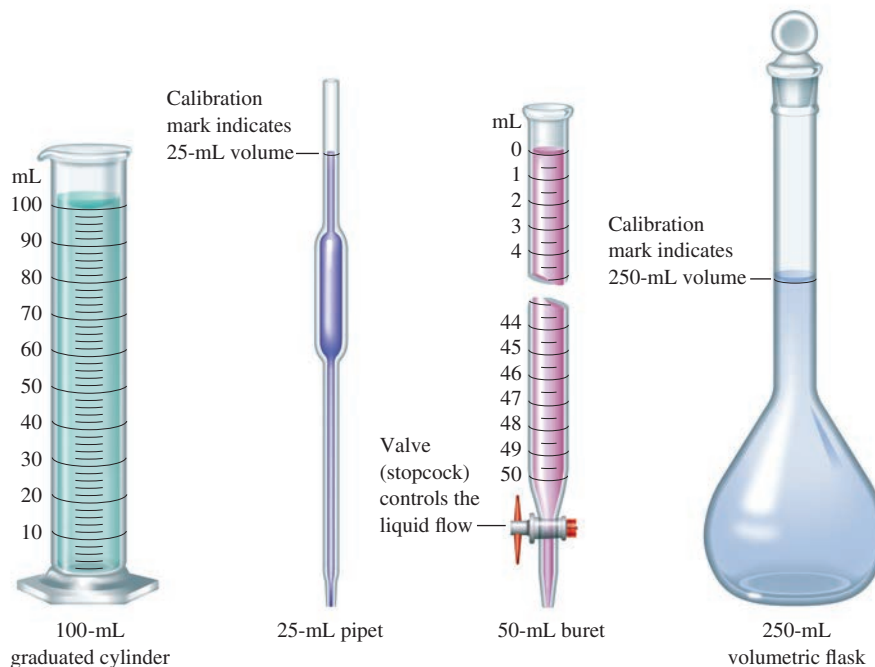


FIGURE R.1 The largest cube has sides 1 m in length and a volume of 1 m^3 . The middle-sized cube has sides 1 dm in length and a volume of 1 dm^3 , or 1 L. The smallest cube has sides 1 cm in length and a volume of 1 cm^3 , or 1 mL.

FIGURE R.2 Common types of laboratory equipment used to measure liquid volume.



Chemical laboratory work frequently requires measurement of the volumes of liquids. Several devices for the accurate determination of liquid volume are shown in Fig. R.2.

An important point concerning measurements is the relationship between mass and weight. Although these terms are sometimes used interchangeably, they are *not* the same. *Mass is a measure of the resistance of an object to a change in its state of motion.* Mass is measured by the force necessary to give an object a certain acceleration. On the earth we use the force that gravity exerts on an object to measure its mass. We call this force the object's **weight**. Since weight is the response of mass to gravity, it varies with the strength of the gravitational field. Therefore, your body mass is the same on the earth or on the moon, but your weight would be much less on the moon than on the earth because of the moon's smaller gravitational field.

Because weighing something on a chemical balance involves comparing the mass of that object to a standard mass, the terms *weight* and *mass* are sometimes used interchangeably, although this is incorrect.

R.2 Uncertainty in Measurement

The number associated with a measurement is obtained using some measuring device. For example, consider the measurement of the volume of a liquid using a buret (shown in Fig. R.3 with the scale greatly magnified). Notice that the meniscus of the liquid occurs at about 19.85 milliliters. This means that about 19.85 mL of liquid has been delivered from the buret (if the initial position of the liquid meniscus was 0.00 mL). Note that we must estimate the last number of the volume reading by interpolating between the 0.1-mL marks. Since the last number is estimated, its value may be different if another person makes the same measurement. If five different people read the same volume, the results might be as follows:

Person	Results of Measurement
1	19.85 mL
2	19.84 mL
3	19.86 mL
4	19.87 mL
5	19.86 mL

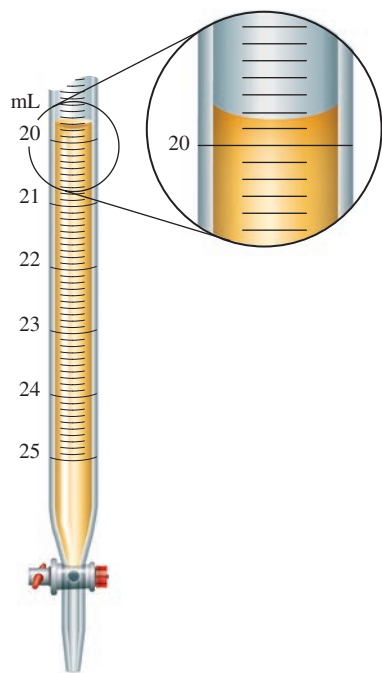


FIGURE R.3 Measurement of volume using a buret. The volume is read at the bottom of the liquid curve (called the meniscus).

A measurement always has some degree of uncertainty.

Uncertainty in measurement is discussed in more detail in Appendix 1.5.

These results show that the first three numbers (19.8) remain the same regardless of who makes the measurement; these are called *certain* digits. However, the digit to the right of the 1 must be estimated and therefore varies; it is called an *uncertain* digit. We customarily report a measurement by recording all the certain digits plus the *first* uncertain digit. In our example it would not make any sense to try to record the volume of thousandths of a milliliter, because the value for hundredths of a milliliter must be estimated when using the buret.

It is very important to realize that a *measurement always has some degree of uncertainty*. The uncertainty of a measurement depends on the precision of the measuring device. For example, using a bathroom scale, you might estimate the mass of a grapefruit to be approximately 1.5 pounds. Weighing the same grapefruit on a highly precise balance might produce a result of 1.476 pounds. In the first case, the uncertainty occurs in the tenths of a pound place; in the second case, the uncertainty occurs in the thousandths of a pound place. Suppose we weigh two similar grapefruits on the two devices and obtain the following results:

	Bathroom Scale	Balance
Grapefruit 1	1.5 lb	1.476 lb
Grapefruit 2	1.5 lb	1.518 lb

Do the two grapefruits have the same mass? The answer depends on which set of results you consider. Thus a conclusion based on a series of measurements depends on the certainty of those measurements. For this reason, it is important to indicate the uncertainty in any measurement. This is done by always recording the certain digits and the first uncertain digit (the estimated number). These numbers are called the **significant figures** of a measurement.

The convention of significant figures automatically indicates something about the uncertainty in a measurement. The uncertainty in the last number (the estimated number) is usually assumed to be ± 1 unless otherwise indicated. For example, the measurement 1.86 kilograms can be taken to mean 1.86 ± 0.01 kilograms.

INTERACTIVE EXAMPLE R.1

Uncertainty in Measurement

In analyzing a sample of polluted water, a chemist measured out a 25.00-mL water sample with a pipet (see Fig. R.2). At another point in the analysis, the chemist used a graduated cylinder (see Fig. R.2) to measure 25 mL of a solution. What is the difference between the measurements 25.00 mL and 25 mL?

SOLUTION

Even though the two volume measurements appear to be equal, they really convey different information. The quantity 25 mL means that the volume is between 24 mL and 26 mL, whereas the quantity 25.00 mL means that the volume is between 24.99 mL and 25.01 mL. The pipet measures volume with much greater precision than does the graduated cylinder.

See Question R.21

When making a measurement, it is important to record the results to the appropriate number of significant figures. For example, if a certain buret can be read to ± 0.01 mL, you should record a reading of twenty-five milliliters as 25.00 mL, not 25 mL. This way, at some later time when you are using your results to do calculations, the uncertainty in the measurement will be known to you.

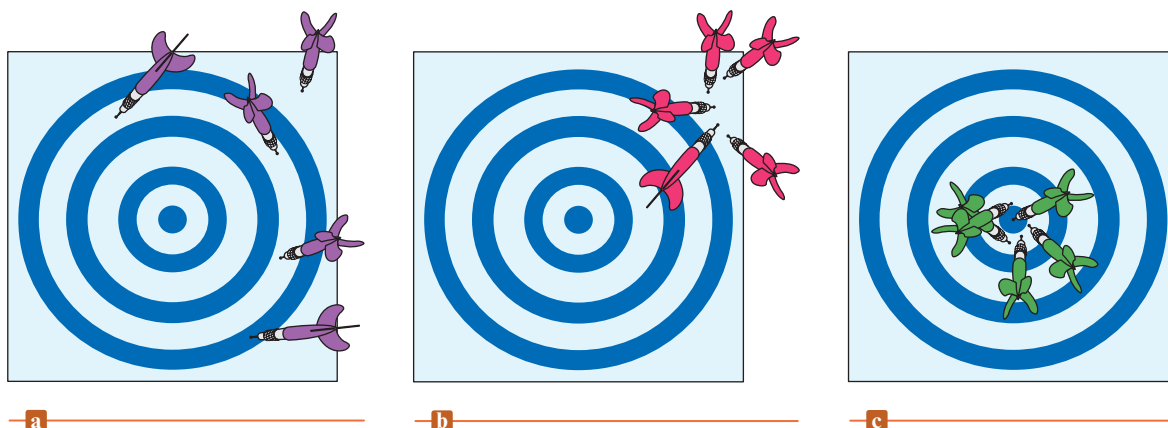


FIGURE R.4 The results of several dart throws show the difference between precise and accurate. **(a)** Neither accurate nor precise (large random errors). **(b)** Precise but not accurate (small random errors, large systematic error). **(c)** Bull's-eye! Both precise and accurate (small random errors, no systematic error).

Precision and Accuracy

Two terms often used to describe the reliability of measurements are *precision* and *accuracy*. Although these words are frequently used interchangeably in everyday life, they have different meanings in the scientific context. **Accuracy** refers to the agreement of a particular value with the true value. **Precision** refers to the degree of agreement among several measurements of the same quantity. Precision reflects the *reproducibility* of a given type of measurement. The difference between these terms is illustrated by the results of three different dart throws shown in Fig. R.4.

Two different types of errors are illustrated in Fig. R.4. A **random error** (also called an *indeterminate error*) means that a measurement has an equal probability of being high or low. This type of error occurs in estimating the value of the last digit of a measurement. The second type of error is called **systematic error** (or *determinate error*). This type of error occurs in the same direction each time; it is either always high or always low. Fig. R.4(a) indicates large random errors (poor technique). Fig. R.4(b) indicates small random errors but a large systematic error, and Fig. R.4(c) indicates small random errors and no systematic error.

In quantitative work, precision is often used as an indication of accuracy; we assume that the *average* of a series of precise measurements (which should “average out” the random errors because of their equal probability of being high or low) is accurate, or close to the “true” value. However, this assumption is valid only if systematic errors are absent. Suppose we weigh a piece of brass five times on a very precise balance and obtain the following results:

Weighing	Result
1	2.486 g
2	2.487 g
3	2.485 g
4	2.484 g
5	2.488 g

Normally, we would assume that the true mass of the piece of brass is very close to 2.486 grams, which is the average of the five results:

$$\frac{2.486 \text{ g} + 2.487 \text{ g} + 2.485 \text{ g} + 2.484 \text{ g} + 2.488 \text{ g}}{5} = 2.486 \text{ g}$$

However, if the balance has a defect causing it to give a result that is consistently 1.000 gram too high (a systematic error of +1.000 gram), then the measured value of

2.486 grams would be seriously in error. The point here is that high precision among several measurements is an indication of accuracy *only* if systematic errors are absent.

EXAMPLE R.2

Precision and Accuracy

To check the accuracy of a graduated cylinder, a student filled the cylinder to the 25-mL mark using water delivered from a buret (see Fig. R.2) and then read the volume delivered. Following are the results of five trials:

Trial	Volume Shown by Graduated Cylinder	Volume Shown by the Buret
1	25 mL	26.54 mL
2	25 mL	26.51 mL
3	25 mL	26.60 mL
4	25 mL	26.49 mL
5	25 mL	26.57 mL
Average	25 mL	26.54 mL

Is the graduated cylinder accurate?

SOLUTION

Precision is an indication of accuracy only if there are no systematic errors.

The results of the trials show very good precision (for a graduated cylinder). The student has good technique. However, note that the average value measured using the buret is significantly different from 25 mL. Thus this graduated cylinder is not very accurate. It produces a systematic error (in this case, the indicated result is low for each measurement).

See Question R.2

R.3 Significant Figures and Calculations

Calculating the final result for an experiment usually involves adding, subtracting, multiplying, or dividing the results of various types of measurements. Since it is very important that the uncertainty in the final result is known correctly, we have developed rules for counting the significant figures in each number and for determining the correct number of significant figures in the final result.

Rules for Counting Significant Figures

- 1. Nonzero integers.** Nonzero integers always count as significant figures.
- 2. Zeros.** There are three classes of zeros:
 - a. Leading zeros** are zeros that *precede* all the nonzero digits. These do not count as significant figures. In the number 0.0025, the three zeros simply indicate the position of the decimal point. This number has only two significant figures.
 - b. Captive zeros** are zeros *between* nonzero digits. These always count as significant figures. The number 1.008 has four significant figures.
 - c. Trailing zeros** are zeros at the *right end* of the number. They are significant only if the number contains a decimal point. The number 100 has only one significant figure, whereas the number 1.00×10^2 has three significant figures. The number one hundred written as 100. also has three significant figures.

(Box continues on the following page)

Leading zeros are never significant figures.

Captive zeros are always significant figures.

Trailing zeros are sometimes significant figures.

Exact numbers never limit the number of significant figures in a calculation.

Exponential notation is reviewed in Appendix 1.1.

3. *Exact numbers.* Many times calculations involve numbers that were not obtained using measuring devices but were determined by counting: 10 experiments, 3 apples, 8 molecules. Such numbers are called *exact numbers*. They can be assumed to have an infinite number of significant figures. Other examples of exact numbers are the 2 in $2\pi r$ (the circumference of a circle) and the 4 and the 3 in $\frac{4}{3}\pi r^3$ (the volume of a sphere). Exact numbers also can arise from definitions. For example, one inch is defined as *exactly* 2.54 centimeters. Thus, in the statement $1 \text{ in} = 2.54 \text{ cm}$, neither the 2.54 nor the 1 limits the number of significant figures when used in a calculation.

Note that the number 1.00×10^2 above is written in **exponential notation**. This type of notation has at least two advantages: the number of significant figures can be easily indicated, and fewer zeros are needed to write a very large or very small number. For example, the number 0.000060 is much more conveniently represented as 6.0×10^{-5} . (The number has two significant figures.)

INTERACTIVE EXAMPLE R.3

Significant Figures

Give the number of significant figures for each of the following results.

- A student's extraction procedure on tea yields 0.0105 g of caffeine.
- A chemist records a mass of 0.050080 g in an analysis.
- In an experiment a span of time is determined to be 8.050×10^{-3} s.

SOLUTION

- The number contains three significant figures. The zeros to the left of the 1 are leading zeros and are not significant, but the remaining zero (a captive zero) is significant.
- The number contains five significant figures. The leading zeros (to the left of the 5) are not significant. The captive zeros between the 5 and the 8 are significant, and the trailing zero to the right of the 8 is significant because the number contains a decimal point.
- This number has four significant figures. Both zeros are significant.

See Exercises R.15 through R.18

To this point we have learned to count the significant figures in a given number. Next, we must consider how uncertainty accumulates as calculations are carried out. The detailed analysis of the accumulation of uncertainties depends on the type of calculation involved and can be complex. However, in this textbook we will employ the following simple rules that have been developed for determining the appropriate number of significant figures in the result of a calculation.

Rules for Significant Figures in Mathematical Operations

- For multiplication or division, the number of significant figures in the result is the same as the number in the least precise measurement used in the calculation. For example, consider the calculation

$$\begin{array}{ccc}
 4.56 \times 1.4 = 6.38 & \xrightarrow{\text{Corrected}} & 6.4 \\
 \uparrow & & \uparrow \\
 \text{Limiting term has} & & \text{Two significant} \\
 \text{two significant} & & \text{figures} \\
 \text{figures} & &
 \end{array}$$

The product should have only two significant figures, since 1.4 has two significant figures.

(Box continues on the following page)

2. For addition or subtraction, the result has the same number of decimal places as the least precise measurement used in the calculation. For example, consider the sum

$$\begin{array}{r}
 12.11 \\
 18.0 \\
 \hline
 31.123
 \end{array}
 \begin{array}{l}
 \leftarrow \text{Limiting term has one decimal place} \\
 \text{Corrected} \rightarrow 31.1 \\
 \uparrow \\
 \text{One decimal place}
 \end{array}$$

The correct result is 31.1, since 18.0 has only one decimal place.

For multiplication and division: significant figures are counted.

For addition and subtraction: decimal places are counted.

Rule 2 is consistent with the operation of electronic calculators.

Note that for multiplication and division, significant figures are counted. For addition and subtraction, the decimal places are counted.

In most calculations you will need to round numbers to obtain the correct number of significant figures. The following rules should be applied when rounding.

Rules for Rounding

1. In a series of calculations, carry the extra digits through to the final result, *then* round.
2. If the digit to be removed
 - a. is less than 5, the preceding digit stays the same. For example, 1.33 rounds to 1.3.
 - b. is equal to or greater than 5, the preceding digit is increased by 1. For example, 1.36 rounds to 1.4.

Although rounding is generally straightforward, one point requires special emphasis. As an illustration, suppose that the number 4.348 needs to be rounded to two significant figures. In doing this, we look *only* at the *first number* to the right of the 3:

$$\begin{array}{c}
 4.348 \\
 \uparrow \\
 \text{Look at this number to} \\
 \text{round to two significant figures.}
 \end{array}$$

Do not round sequentially. The number 6.8347 rounded to three significant figures is 6.83, not 6.84.

The number is rounded to 4.3 because 4 is less than 5. It is incorrect to round sequentially. For example, do *not* round the 4 to 5 to give 4.35 and then round the 3 to 4 to give 4.4.

When rounding, *use only the first number to the right of the last significant figure.*

It is important to note that Rule 1 above usually will not be followed in the Examples in this text because we want to show the correct number of significant figures in *each step* of a problem. This same practice is followed for the detailed solutions given in the *Solutions Guide*. However, when you are doing problems, you should carry extra digits throughout a series of calculations and round to the correct number of significant figures only at the end. This is the practice you should follow. The fact that your rounding procedures are different from those used in this text must be taken into account when you check your answer with the one given at the end of the book or in the *Solutions Guide*. Your answer (based on rounding only at the end of a calculation) may differ in the last place from that given here as the “correct” answer because we have rounded after each step. To help you understand the difference between these rounding procedures, we will consider them further in Example R.4.

INTERACTIVE EXAMPLE R.4

Significant Figures in Mathematical Operations

Carry out the following mathematical operations, and give each result with the correct number of significant figures.

- a. $1.05 \times 10^{-3} \div 6.135$
- b. $21 - 13.8$