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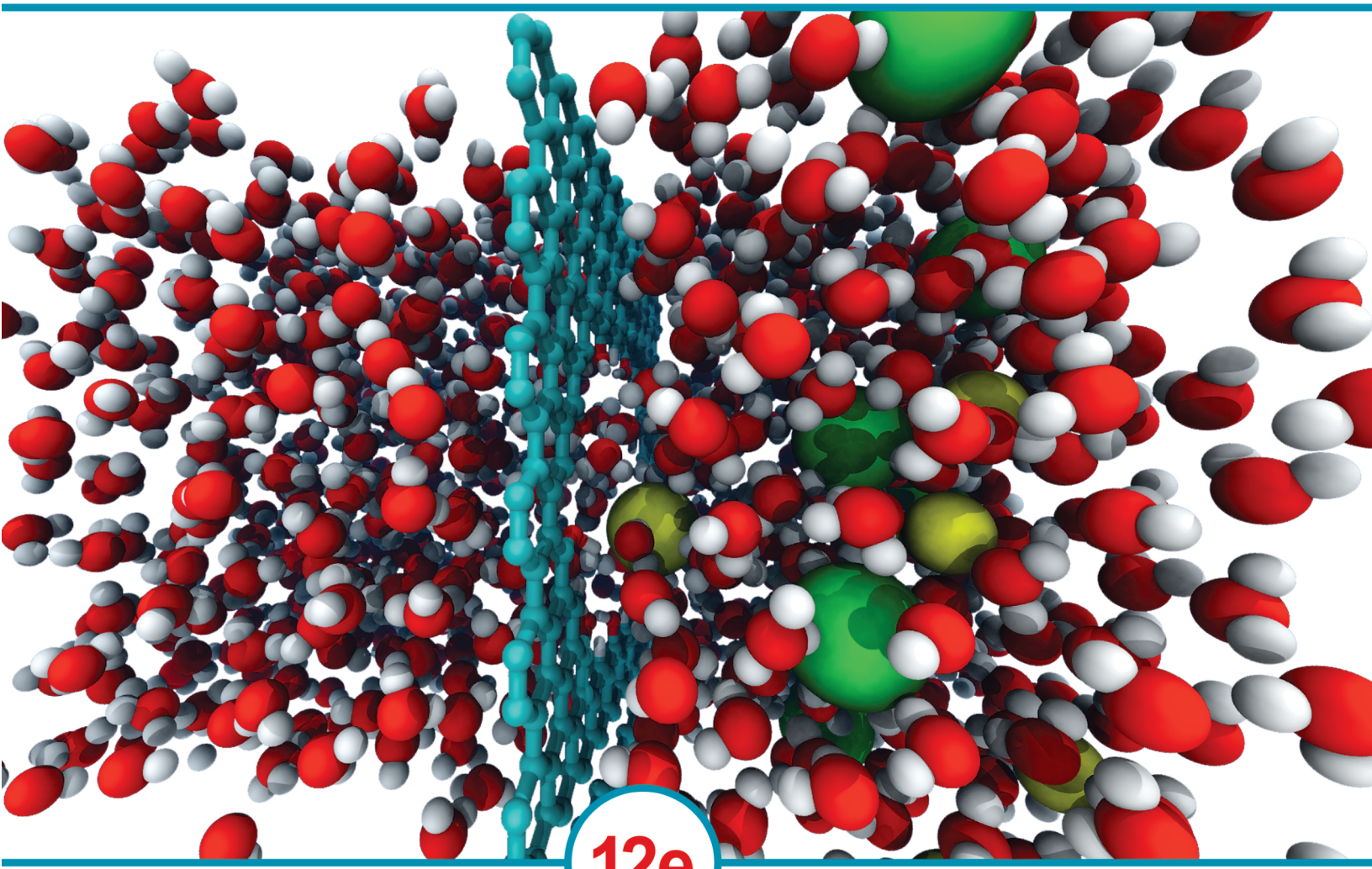

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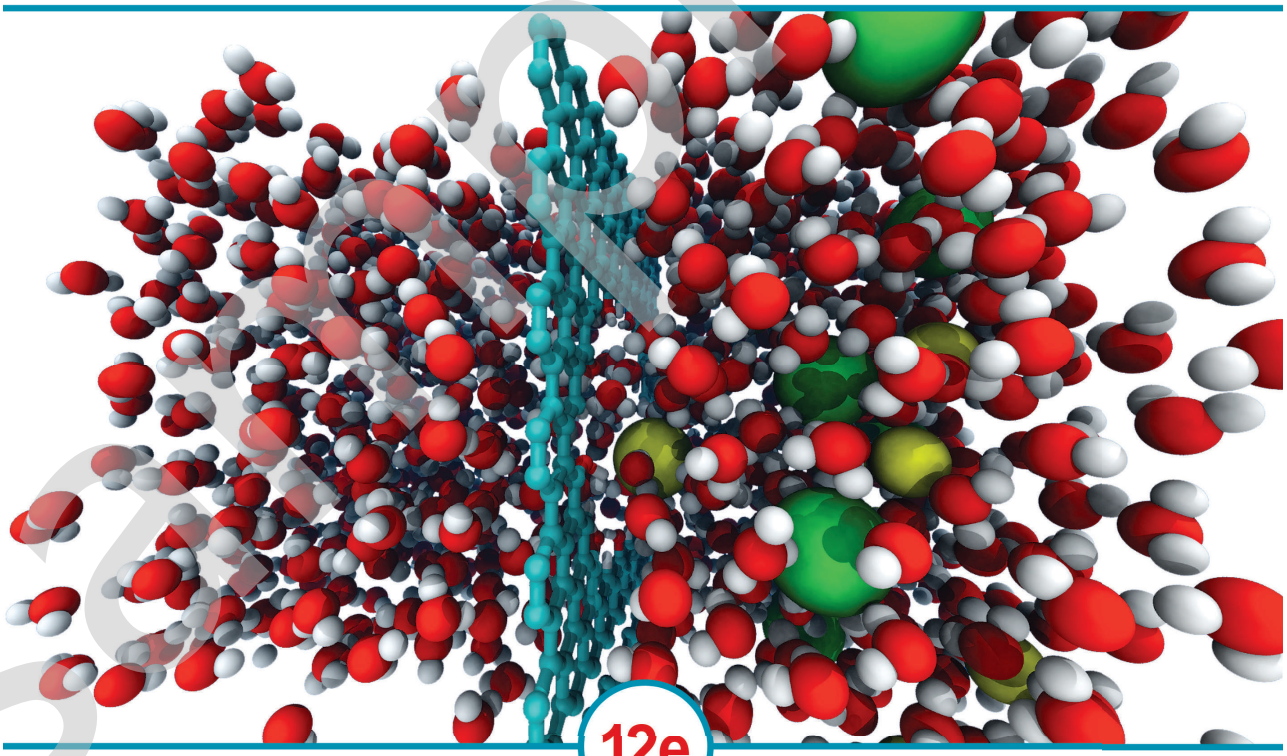

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Raymond Chang

Williams College

Kenneth A. Goldsby

Florida State University

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When he is not working, Ken enjoys hanging out with his family. They especially like spending time together at the coast.

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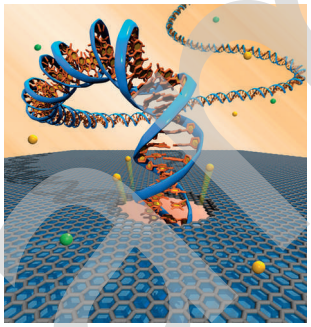

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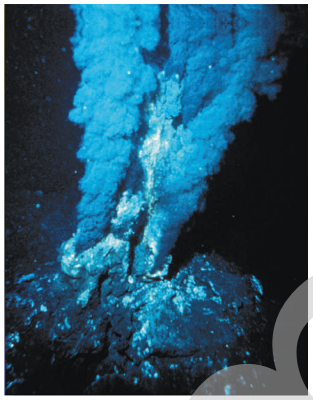

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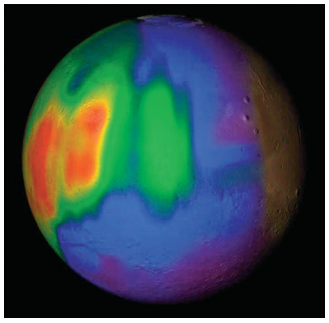

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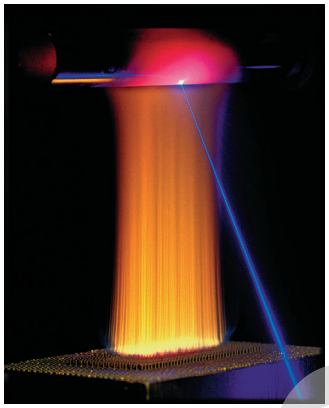

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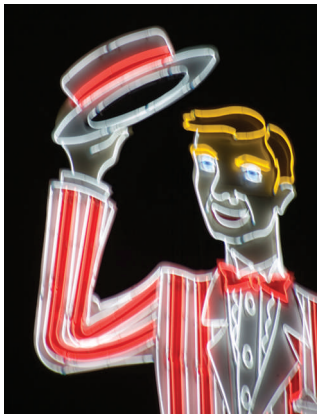

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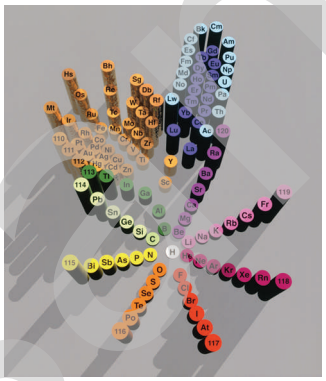

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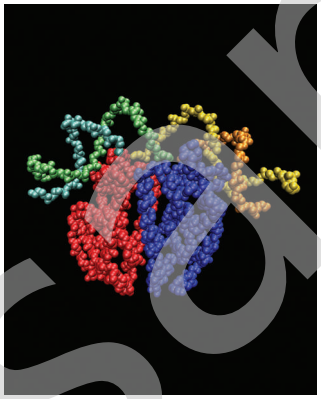

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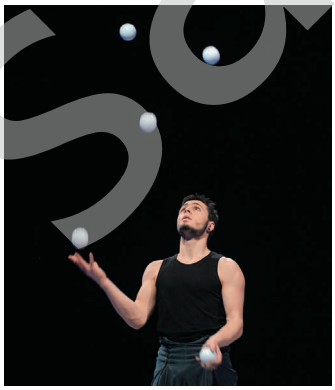

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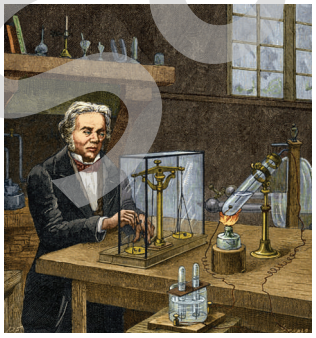

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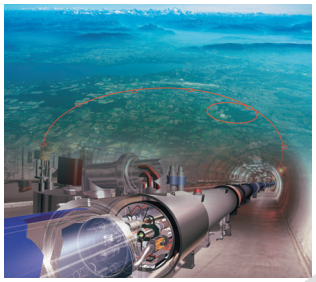

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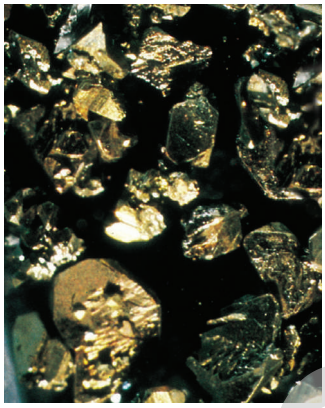

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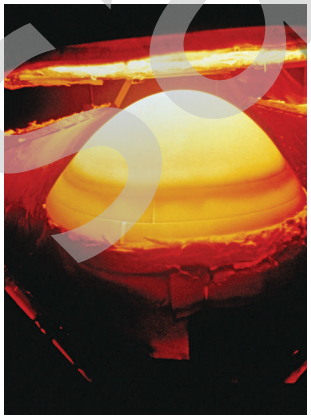

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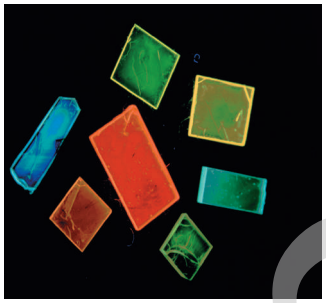

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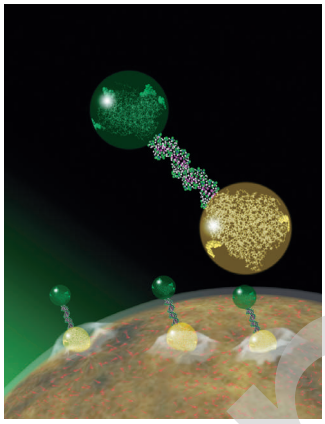

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List of Applications

The opening sentence of this text is, "Chemistry is an active, evolving science that has vital importance to our world, in both the realm of nature and the realm of society." Throughout the text, Chemistry in Action boxes and Chemical Mysteries give specific examples of chemistry as active and evolving in all facets of our lives.

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Preface

The twelfth edition continues the tradition by providing a firm foundation in chemical concepts and principles and to instill in students an appreciation of the vital part chemistry plays in our daily life. It is the responsibility of the textbook authors to assist both instructors and their students in their pursuit of this objective by presenting a broad range of topics in a logical manner. We try to strike a balance between theory and application and to illustrate basic principles with everyday examples whenever possible.

As in previous editions, our goal is to create a text that is clear in explaining abstract concepts, concise so that it does not overburden students with unnecessary extraneous information, yet comprehensive enough so that it prepares students to move on to the next level of learning. The encouraging feedback we have received from instructors and students has convinced us that this approach is effective.

The art program has been extensively revised in this edition. Many of the laboratory apparatuses and scientific instruments were redrawn to enhance the realism of the components. Several of the drawings were updated to reflect advances in the science and applications described in the text; see, for example, the lithium-ion battery depicted in Figure 18.10. Molecular structures were created using ChemDraw, the gold standard in chemical drawing software. Not only do these structures introduce students to the convention used to represent chemical structures in three dimensions that they will see in further coursework, they also provide better continuity with the ChemDraw application they will use in Connect, the online homework and practice system for our text.

In addition to revising the art program, over 100 new photographs are added in this edition. These photos provide a striking look at processes that can be understood by studying the underlying chemistry (see, for example, Figure 19.15, which shows the latest attempt of using lasers to induce nuclear fusion).

Problem Solving

The development of problem-solving skills has always been a major objective of this text. The two major categories of learning are shown next.

Worked examples follow a proven step-by-step strategy and solution.

- **Problem statement** is the reporting of the facts needed to solve the problem based on the question posed.
- **Strategy** is a carefully thought-out plan or method to serve as an important function of learning.
- **Solution** is the process of solving a problem given in a stepwise manner.
- **Check** enables the student to compare and verify with the source information to make sure the answer is reasonable.
- **Practice Exercise** provides the opportunity to solve a similar problem in order to become proficient in this problem type. The Practice Exercises are available in the Connect electronic homework system. The margin note lists additional similar problems to work in the end-of-chapter problem section.

End-of-Chapter Problems are organized in various ways. Each section under a topic heading begins with Review Questions followed by Problems. The Additional Problems section provides more problems not organized by section, followed by the new problem type of Interpreting, Modeling & Estimating.

Many of the examples and end-of-chapter problems present extra tidbits of knowledge and enable the student to solve a chemical problem that a chemist would solve. The examples and problems show students the real world of chemistry and applications to everyday life situations.

Visualization

Graphs and Flow Charts are important in science. In *Chemistry,* flow charts show the thought process of a concept and graphs present data to comprehend the concept. A significant number of Problems and Review of Concepts, including many new to this edition, include graphical data.

Molecular art appears in various formats to serve different needs. Molecular models help to visualize the three-dimensional arrangement of atoms in a molecule. Electrostatic potential maps illustrate the electron density distribution in molecules. Finally, there is the macroscopic to microscopic art helping students understand processes at the molecular level.

Photos are used to help students become familiar with chemicals and understand how chemical reactions appear in reality.

Figures of apparatus enable the student to visualize the practical arrangement in a chemistry laboratory.

Study Aids

Setting the Stage

Each chapter starts with the Chapter Outline and A Look Ahead.

Chapter Outline enables the student to see at a glance the big picture and focus on the main ideas of the chapter.

A Look Ahead provides the student with an overview of concepts that will be presented in the chapter.

Tools to Use for Studying

Useful aids for studying are plentiful in *Chemistry* and should be used constantly to reinforce the comprehension of chemical concepts.

Marginal Notes are used to provide hints and feedback to enhance the knowledge base for the student.

Worked Examples along with the accompanying Practice Exercises are very important tools for learning and mastering chemistry. The problem-solving steps guide the student through the critical thinking necessary for succeeding in chemistry. Using sketches helps student understand the inner workings of a problem. (See Example 6.1 on page 238.) A margin note lists similar problems in the end-of-chapter problems section, enabling the student to apply new skill to other problems of the same type. Answers to the Practice Exercises are listed at the end of the chapter problems.

Review of Concepts enables the student to evaluate if they understand the concept presented in the section.

Key Equations are highlighted within the chapter, drawing the student's eye to material that needs to be understood and retained. The key equations are also presented in the chapter summary materials for easy access in review and study.

Summary of Facts and Concepts provides a quick review of concepts presented and discussed in detail within the chapter.

Key Words are a list of all important terms to help the student understand the language of chemistry.

Testing Your Knowledge

Review of Concepts lets students pause and check to see if they understand the concept presented and discussed in the section occurred. Answers to the Review of Concepts can be found in the Student Solution Manual and online in the accompanying Connect Chemistry companion website.

End-of-Chapter Problems enable the student to practice critical thinking and problem-solving skills. The problems are broken into various types:

- By chapter section. Starting with Review Questions to test basic conceptual understanding, followed by Problems to test the student's skill in solving problems for that particular section of the chapter.
- Additional Problems uses knowledge gained from the various sections and/or previous chapters to solve the problem.
- Interpreting, Modeling & Estimating problems teach students the art of formulating models and estimating ballpark answers based on appropriate assumptions.

Real-Life Relevance

Interesting examples of how chemistry applies to life are used throughout the text. Analogies are used where appropriate to help foster understanding of abstract chemical concepts.

End-of-Chapter Problems pose many relevant questions for the student to solve. Examples include Why do swimming coaches sometimes place a drop of alcohol in a swimmer's ear to draw out water? How does one estimate the pressure in a carbonated soft drink bottle before removing the cap?

Chemistry in Action boxes appear in every chapter on a variety of topics, each with its own story of how chemistry can affect a part of life. The student can learn about the science of scuba diving and nuclear medicine, among many other interesting cases.

Chemical Mystery poses a mystery case to the student. A series of chemical questions provide clues as to how the mystery could possibly be solved. Chemical Mystery will foster a high level of critical thinking using the basic problem-solving steps built up throughout the text.

Digital Resources

McGraw-Hill Education offers various tools and technology products to support *Chemistry,* 12th edition.

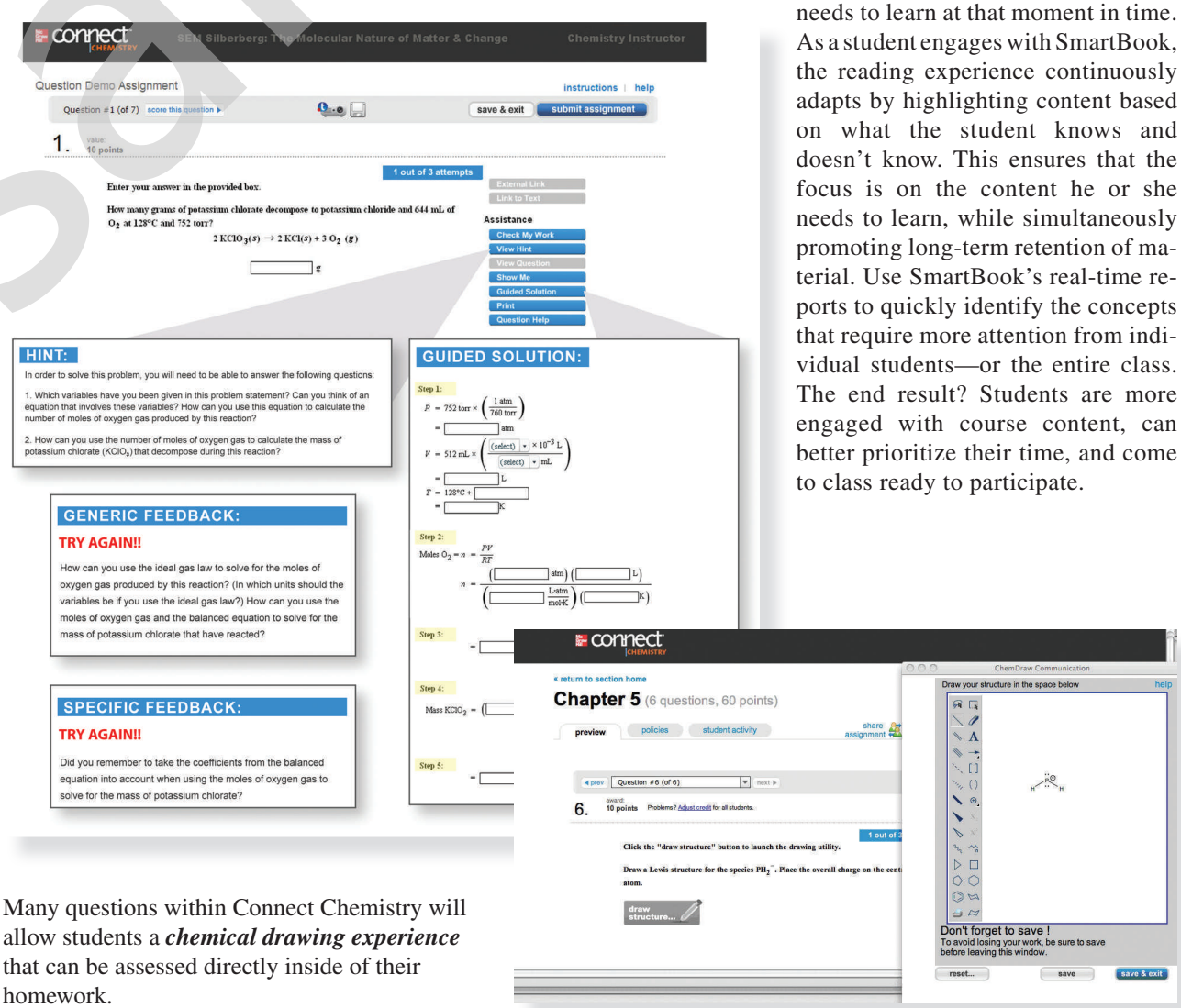
connect plus+ CHEMISTRY

Featuring ***PerkinElmer® ChemDraw***

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

SMARTBOOK®

SmartBook is the first and only adaptive reading experience designed to change the way students read and learn. It creates a personalized reading experience by highlighting the most impactful concepts a student needs to learn at that moment in time. As a student engages with SmartBook, the reading experience continuously adapts by highlighting content based on what the student knows and doesn't know. This ensures that the focus is on the content he or she needs to learn, while simultaneously promoting long-term retention of material. Use SmartBook's real-time reports to quickly identify the concepts that require more attention from individual students—or the entire class. The end result? Students are more engaged with course content, can better prioritize their time, and come to class ready to participate.



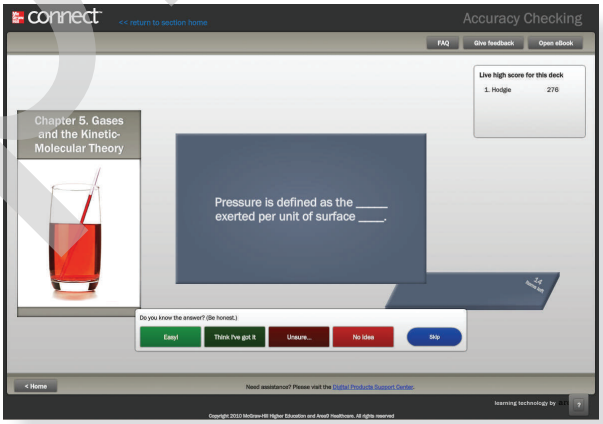
Many questions within Connect Chemistry will allow students a ***chemical drawing experience*** that can be assessed directly inside of their homework.

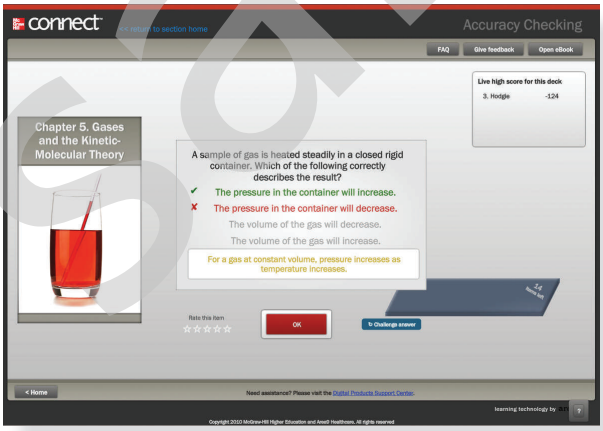
LEARNSMART®

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

Adaptive Probes

A student's knowledge is intelligently probed by asking a series of questions. These questions dynamically change both in the level of difficulty and in content based on the student's weak and strong areas. Each practice session is based on the previous performance, and LearnSmart uses sophisticated models for predicting what the student will forget and how to reinforce that material typically forgotten. This saves students study time and ensures that they have actual mastery of the concepts.



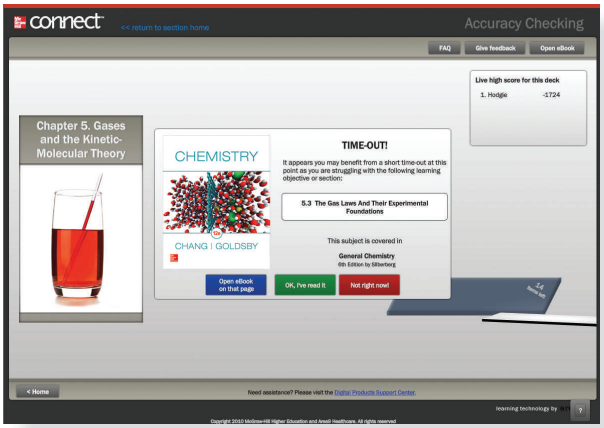


Immediate Feedback

When a student incorrectly answers a probe, the correct answer is provided, along with feedback.

Time Out

When LearnSmart has identified a specific subject area where the student is struggling, he or she is given a "time out" and directed to the textbook section or learning objective for remediation. With ConnectPlus, students are provided with a link to the specific page of the eBook where they can study the material immediately.



Reporting

Dynamically generated reports document student progress and areas for additional reinforcement, offering at-a-glance views of their strengths and weaknesses.

Reports Include:

- Most challenging learning objectives
- Tree of wisdom
- Test results
- Current learning status
- Metacognitive skills
- Missed questions
- Learning plan

LearnSmart Labs for General Chemistry™

LEARNSMART LABS®

THE Virtual Lab Experience.

LearnSmart Labs is a must-see, outcomes-based lab simulation. It assesses a student's knowledge and adaptively corrects deficiencies, allowing the student to learn faster and retain more knowledge with greater success.

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

LEARNSMART PREP®

LearnSmart Prep is an adaptive tool that prepares students for the course they are about to take. It identifies the prerequisite knowledge each student doesn't know or fully understand and provides learning resources to teach essential concepts so he or she enters the classroom prepared to succeed.

ALEKS®

ALEKS (Assessment and LEarning in Knowledge Spaces) is a web-based system for individualized assessment and learning available 24/7 over the Internet. ALEKS uses artificial intelligence to accurately determine a student's knowledge and then guides her to the material that she is most ready to learn. ALEKS offers immediate feedback and access to ALEKSPedia—an interactive text that contains concise entries on chemistry topics. ALEKS is also a full-featured course management system with rich reporting features that allow instructors to monitor individual and class performance, set student goals, assign/grade online quizzes, and more. ALEKS allows instructors to spend more time on concepts while ALEKS teaches students practical problem solving skills. And with ALEKS 360, your student also has access to this text's eBook. Learn more at **www.aleks.com/highered/science**

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McGraw-Hill Create™ is a self-service website that allows you to create customized course materials using McGraw-Hill Education's comprehensive, cross-disciplinary content and digital products. You can even access third party content such as readings, articles, cases, videos, and more. Arrange the content you've selected to match the scope and sequence of your course. Personalize your book with a cover design and choose the best format for your students–eBook, color print, or black-and-white print. And, when you are done, you'll receive a PDF review copy *in just minutes!* **www.mcgrawhillcreate.com**

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Tegrity Campus is a fully automated lecture capture solution used in traditional, hybrid, "flipped classes" and online courses to record lesson, lectures, and skills. Its personalized learning features make study time incredibly efficient and its ability to affordably scale brings this benefit to every student on campus. Patented search technology and real-time LMS integrations make Tegrity the market-leading solution and service. Tegrity is available as an integrated feature of McGraw-Hill Connect® Chemistry and as a standalone.

Presentation Tools

Build instructional materials wherever, whenever, and however you want! Access instructor tools from your text's Connect website to find photo's, artwork, animations, and other media that can be used to create customized lectures, visually enhanced tests and quizzes, compelling course websites, or attractive printed support materials. All assets are copyrighted by McGraw-Hill Higher Education, but can be used by instructors for classroom purposes. The visual resources in this collection include:

- **Art** Full-color digital files of all illustrations in the book.
- **Photos** The photo collection contains digital files of photographs from the text.
- **Tables** Every table that appears in the text is available electronically.
- **Animations** Numerous full-color animations illustrating important processes are also provided.

- **PowerPoint Lecture Outlines** Ready made presentations for each chapter of the text.
- **PowerPoint Slides** All illustrations, photos, and tables are pre-inserted by chapter into blank PowerPoint slides.

Computerized Test Bank Online

A comprehensive bank of questions is provided within a computerized test bank, enabling professors to prepare and access tests or quizzes. Instructors can create or edit questions, or drag-and-drop questions to prepare tests quickly and easily. Tests can be published to their online course, or printed for paper-based assignments.

Instructor's Solution's Manual

The ***Instructor's Solution Manual,*** written by Raymond Chang and Ken Goldsby, provides the solutions to most end-of-chapter problems. The manual also provides the difficulty level and category type for each problem. This manual is available to instructors online in the text's Connect library tab.

Instructor's Manual

The ***Instructor's Manual*** provides a brief summary of the contents of each chapter, along with the learning goals, references to background concepts in earlier chapters, and teaching tips. This manual can be found online for instructors on the text's Connect library tab.

For the Student

Students can order supplemental study materials by contacting their campus bookstore, calling 1-800-262-4729, or online at http://shop.mheducation.com

Student Solutions Manual
ISBN 1-25-928622-3

The ***Student Solutions Manual*** is written by Raymond Chang and Ken Goldsby. This supplement contains detailed solutions and explanations for even-numbered problems in the main text. The manual also includes a detailed discussion of different types of problems and approaches to solving chemical problems and tutorial solutions for many of the end-of-chapter problems in the text, along with strategies for solving them. Note that solutions to the problems listed under Interpreting, Modeling & Estimating are not provided in the manual.

Student Study Guide ISBN 1-25-928623-1

This valuable ancillary contains material to help the student practice problem-solving skills. For each section of a chapter, the author provides study objectives and a summary of the corresponding text. Following the summary are sample problems with detailed solutions. Each chapter has true–false questions and a self-test, with all answers provided at the end of the chapter.

Animations for MP3/iPod

A number of animations are available for download to your MP3/iPod through the textbook's Connect website.

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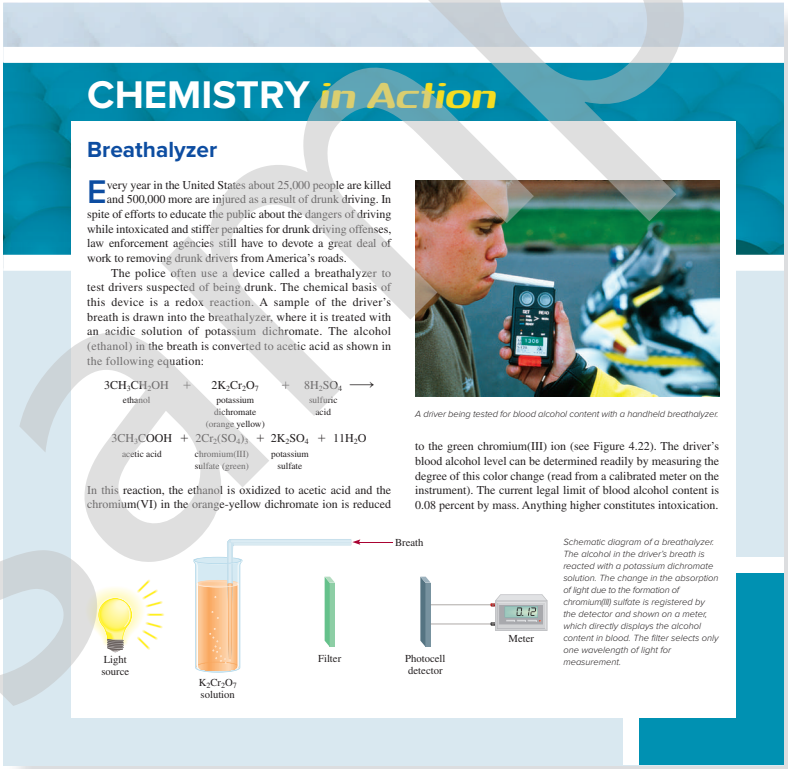
Setting the Stage for Learning

Real-Life Relevance

Interesting examples of how chemistry applies to life are used throughout the text. Analogies are used where appropriate to help foster understanding of abstract chemical concepts.



CHEMISTRY *in Action*

Breathalyzer

Every year in the United States about 25,000 people are killed and 500,000 more are injured as a result of drunk driving. In spite of efforts to educate the public about the dangers of driving while intoxicated and stiffer penalties for drunk driving offenses, law enforcement agencies still have to devote a great deal of work to removing drunk drivers from America's roads.

The police often use a device called a breathalyzer to test drivers suspected of being drunk. The chemical basis of this device is a redox reaction. A sample of the driver's breath is drawn into the breathalyzer, where it is treated with an acidic solution of potassium dichromate. The alcohol (ethanol) in the breath is converted to acetic acid as shown in the following equation:

$$\underset{\text{ethanol}}{3CH_3CH_2OH} + \underset{\substack{\text{potassium dichromate}\\ \text{(orange yellow)}}}{2K_2Cr_2O_7} + \underset{\text{sulfuric acid}}{8H_2SO_4} \longrightarrow \underset{\text{acetic acid}}{3CH_3COOH} + \underset{\substack{\text{chromium(III)}\\ \text{sulfate (green)}}}{2Cr_2(SO_4)_3} + \underset{\substack{\text{potassium}\\ \text{sulfate}}}{2K_2SO_4} + 11H_2O$$

In this reaction, the ethanol is oxidized to acetic acid and the chromium(VI) in the orange-yellow dichromate ion is reduced to the green chromium(III) ion (see Figure 4.22). The driver's blood alcohol level can be determined readily by measuring the degree of this color change (read from a calibrated meter on the instrument). The current legal limit of blood alcohol content is 0.08 percent by mass. Anything higher constitutes intoxication.

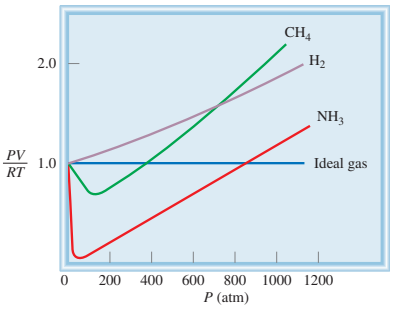
A driver being tested for blood alcohol content with a handheld breathalyzer.

Schematic diagram of a breathalyzer. The alcohol in the driver's breath is reacted with a potassium dichromate solution. The change in the absorption of light due to the formation of chromium(III) sulfate is registered by the detector and shown on a meter, which directly displays the alcohol content in blood. The filter selects only one wavelength of light for measurement.

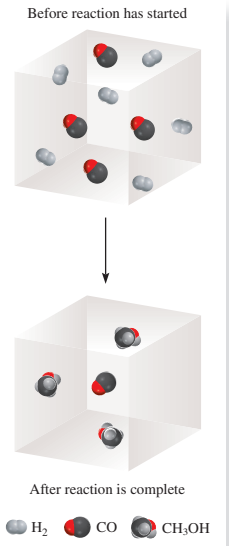
Chemistry in Action boxes appear in every chapter on a variety of topics, each with its own story of how chemistry can affect a part of life. The student can learn about the science of scuba diving and nuclear medicine, among many other interesting cases.

Chemical Mystery poses a mystery case to the student. A series of chemical questions provide clues as to how the mystery could possibly be solved. Chemical Mystery will foster a high level of critical thinking using the basic problem-solving steps built up throughout the text.

Visualization



Graphs and Flow Charts



Molecular art

Study Aids

Key Equations—material to retain

Summary of Facts & Concepts—quick review of important concepts

Key Words—important terms to understand

Key Equations

$\Delta U = q + w$ (6.1)	Mathematical statement of the first law of thermodynamics.
$w = -P\Delta V$ (6.3)	Calculating work done in gas expansion or gas compression.
$H = U + PV$ (6.6)	Definition of enthalpy.
$\Delta H = \Delta U + P\Delta V$ (6.8)	Calculating enthalpy (or energy) change for a constant-pressure process.
$C = ms$ (6.11)	Definition of heat capacity.
$q = ms\Delta t$ (6.12)	Calculating heat change in terms of specific heat.
$q = C\Delta t$ (6.13)	Calculating heat change in terms of heat capacity.
$\Delta H^\circ_{rxn} = \Sigma n\Delta H^\circ_f(\text{products}) - \Sigma m\Delta H^\circ_f(\text{reactants})$ (6.18)	Calculating standard enthalpy of reaction.
$\Delta H_{soln} = U + \Delta H_{hydr}$ (6.20)	Lattice energy and hydration contributions to heat of solution.

Summary of Facts & Concepts

1. Modern chemistry began with Dalton's atomic theory, which states that all matter is composed of tiny, indivisible particles called atoms; that all atoms of the same element are identical; that compounds contain atoms of different elements combined in whole-number ratios; and that atoms are neither created nor destroyed in chemical reactions (the law of conservation of mass).
2. Atoms of constituent elements in a particular compound are always combined in the same proportions by mass (law of definite proportions). When two elements can combine to form more than one type of compound, the masses of one element that combine with a fixed mass of the other element are in a ratio of small whole numbers (law of multiple proportions).
3. An atom consists of a very dense central nucleus containing protons and neutrons, with electrons moving about the nucleus at a relatively large distance from it.
4. Protons are positively charged, neutrons have no charge, and electrons are negatively charged. Protons and neutrons have roughly the same mass, which is about 1840 times greater than the mass of an electron.
5. The atomic number of an element is the number of protons in the nucleus of an atom of the element; it determines the identity of an element. The mass number is the sum of the number of protons and the number of neutrons in the nucleus.
6. Isotopes are atoms of the same element with the same number of protons but different numbers of neutrons.
7. Chemical formulas combine the symbols for the constituent elements with whole-number subscripts to show the type and number of atoms contained in the smallest unit of a compound.
8. The molecular formula conveys the specific number and type of atoms combined in each molecule of a compound. The empirical formula shows the simplest ratios of the atoms combined in a molecule.
9. Chemical compounds are either molecular compounds (in which the smallest units are discrete, individual molecules) or ionic compounds, which are made of cations and anions.
10. The names of many inorganic compounds can be deduced from a set of simple rules. The formulas can be written from the names of the compounds.
11. Organic compounds contain carbon and elements like hydrogen, oxygen, and nitrogen. Hydrocarbon is the simplest type of organic compound.

Key Words

Acid, p. 62
Alkali metals, p. 50
Alkaline earth metals, p. 50
Allotrope, p. 52
Alpha (α) particles, p. 43
Alpha (α) rays, p. 43
Anion, p. 51
Atom, p. 40
Atomic number (Z), p. 46
Base, p. 64
Beta (β) particles, p. 43
Beta (β) rays, p. 43
Binary compound, p. 56
Cation, p. 51
Chemical formula, p. 52
Diatomic molecule, p. 50
Electron, p. 41
Empirical formula, p. 53
Families, p. 48
Gamma (γ) rays, p. 43
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Halogens, p. 50
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Metal, p. 48
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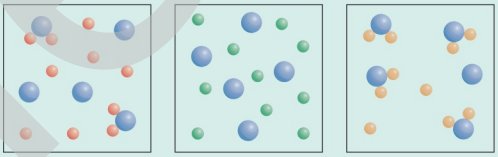
Monatomic ion, p. 51
Neutron, p. 45
Noble gases, p. 50
Nonmetal, p. 48
Nucleus, p. 44
Organic compound, p. 56
Oxoacid, p. 62
Oxoanion, p. 63
Periods, p. 48
Periodic table, p. 48
Polyatomic ion, p. 51
Polyatomic molecule, p. 50
Proton, p. 44
Radiation, p. 41
Radioactivity, p. 43
Structural formula, p. 53
Ternary compound, p. 57

Chang Learning System

Review the section content by using this quick test for acquired knowledge.

Review of Concepts

The diagrams here show three compounds AB_2 (a), AC_2 (b), and AD_2 (c) dissolved in water. Which is the strongest electrolyte and which is the weakest? (For simplicity, water molecules are not shown.)



(a) (b) (c)

Example 4.6

Classify the following redox reactions and indicate changes in the oxidation numbers of the elements:

(a) $2N_2O(g) \longrightarrow 2N_2(g) + O_2(g)$

(b) $6Li(s) + N_2(g) \longrightarrow 2Li_3N(s)$

(c) $Ni(s) + Pb(NO_3)_2(aq) \longrightarrow Pb(s) + Ni(NO_3)_2(aq)$

(d) $2NO_2(g) + H_2O(l) \longrightarrow HNO_2(aq) + HNO_3(aq)$

Strategy Review the definitions of combination reactions, decomposition reactions, displacement reactions, and disproportionation reactions.

Solution

(a) This is a decomposition reaction because one reactant is converted to two different products. The oxidation number of N changes from +1 to 0, while that of O changes from −2 to 0.

(b) This is a combination reaction (two reactants form a single product). The oxidation number of Li changes from 0 to +1 while that of N changes from 0 to −3.

(c) This is a metal displacement reaction. The Ni metal replaces (reduces) the Pb^{2+} ion. The oxidation number of Ni increases from 0 to +2 while that of Pb decreases from +2 to 0.

(d) The oxidation number of N is +4 in NO_2 and it is +3 in HNO_2 and +5 in HNO_3. Because the oxidation number of the *same* element both increases and decreases, this is a disproportionation reaction.

Practice Exercise Identify the following redox reactions by type:

(a) $Fe + H_2SO_4 \longrightarrow FeSO_4 + H_2$

(b) $S + 3F_2 \longrightarrow SF_6$

(c) $2CuCl \longrightarrow Cu + CuCl_2$

(d) $2Ag + PtCl_2 \longrightarrow 2AgCl + Pt$

Learn a problem-solving process of strategizing, solving, and checking your way to a solution.

Use the problem-solving approach on real-world problems. Interpreting, Modeling & Estimating problems provide students the opportunity to solve problems like a chemist.

4.172 Potassium superoxide (KO_2), a useful source of oxygen employed in breathing equipment, reacts with water to form potassium hydroxide, hydrogen peroxide, and oxygen. Furthermore, potassium superoxide also reacts with carbon dioxide to form potassium carbonate and oxygen. (a) Write equations for these two reactions and comment on the effectiveness of potassium superoxide in this application. (b) Focusing only on the reaction between KO_2 and CO_2, estimate the amount of KO_2 needed to sustain a worker in a polluted environment for 30 min. See Problem 1.69 for useful information.

A Note to the Student

General chemistry is commonly perceived to be more difficult than most other subjects. There is some justification for this perception. For one thing, chemistry has a very specialized vocabulary. At first, studying chemistry is like learning a new language. Furthermore, some of the concepts are abstract. Nevertheless, with diligence you can complete this course successfully, and you might even enjoy it. Here are some suggestions to help you form good study habits and master the material in this text.

- Attend classes regularly and take careful notes.
- If possible, always review the topics discussed in class the same day they are covered in class. Use this book to supplement your notes.
- Think critically. Ask yourself if you really understand the meaning of a term or the use of an equation. A good way to test your understanding is to explain a concept to a classmate or some other person.
- Do not hesitate to ask your instructor or your teaching assistant for help.

The twelfth edition tools for *Chemistry* are designed to enable you to do well in your general chemistry course. The following guide explains how to take full advantage of the text, technology, and other tools.

- Before delving into the chapter, read the chapter *outline* and the chapter *introduction* to get a sense of the important topics. Use the outline to organize your note taking in class.
- At the end of each chapter you will find a summary of facts and concepts, the key equations, and a list of key words, all of which will help you review for exams.
- Definitions of the key words can be studied in context on the pages cited in the end-of-chapter list or in the glossary at the back of the book.
- Careful study of the worked-out examples in the body of each chapter will improve your ability to analyze problems and correctly carry out the calculations needed to solve them. Also take the time to work through the practice exercise that follows each example to be sure you understand how to solve the type of problem illustrated in the example. The answers to the practice exercises appear at the end of the chapter, following the questions and problems. For additional practice, you can turn to similar problems referred to in the margin next to the example.
- The questions and problems at the end of the chapter are organized by section.
- The back inside cover shows a list of important figures and tables with page references. This index makes it convenient to quickly look up information when you are solving problems or studying related subjects in different chapters.

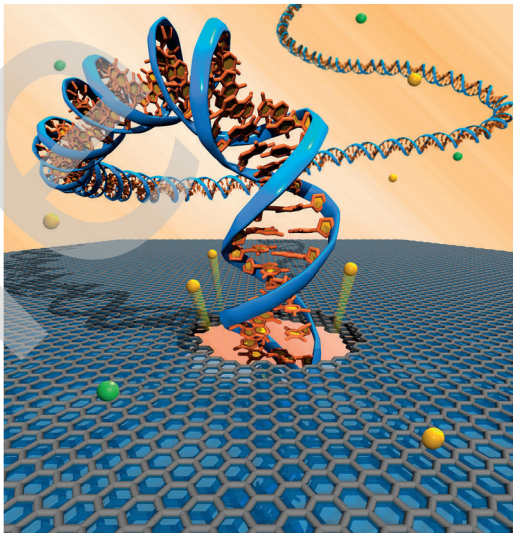
If you follow these suggestions and stay up-to-date with your assignments, you should find that chemistry is challenging, but less difficult and much more interesting than you expected.

—Raymond Chang and Ken Goldsby

CHAPTER

1

Chemistry
The Study of Change



By applying electric fields to push DNA molecules through pores created in graphene, scientists have developed a technique that someday can be used for fast sequencing the four chemical bases according to their unique electrical properties.

CHAPTER OUTLINE

1.1 Chemistry: A Science for the Twenty-First Century
1.2 The Study of Chemistry
1.3 The Scientific Method
1.4 Classifications of Matter
1.5 The Three States of Matter
1.6 Physical and Chemical Properties of Matter
1.7 Measurement
1.8 Handling Numbers
1.9 Dimensional Analysis in Solving Problems
1.10 Real-World Problem Solving: Information, Assumptions, and Simplifications

A LOOK AHEAD

- We begin with a brief introduction to the study of chemistry and describe its role in our modern society. (1.1 and 1.2)
- Next, we become familiar with the scientific method, which is a systematic approach to research in all scientific disciplines. (1.3)
- We define matter and note that a pure substance can either be an element or a compound. We distinguish between a homogeneous mixture and a heterogeneous mixture. We also learn that, in principle, all matter can exist in one of three states: solid, liquid, and gas. (1.4 and 1.5)
- To characterize a substance, we need to know its physical properties, which can be observed without changing its identity and chemical properties, which can be demonstrated only by chemical changes. (1.6)
- Being an experimental science, chemistry involves measurements. We learn the basic SI units and use the SI-derived units for quantities like volume and density. We also become familiar with the three temperature scales: Celsius, Fahrenheit, and Kelvin. (1.7)
- Chemical calculations often involve very large or very small numbers and a convenient way to deal with these numbers is the scientific notation. In calculations or measurements, every quantity must show the proper number of significant figures, which are the meaningful digits. (1.8)
- We learn that dimensional analysis is useful in chemical calculations. By carrying the units through the entire sequence of calculations, all the units will cancel except the desired one. (1.9)
- Solving real-world problems frequently involves making assumptions and simplifications. (1.10)

Chemistry is an active, evolving science that has vital importance to our world, in both the realm of nature and the realm of society. Its roots are ancient, but as we will see, chemistry is every bit a modern science.

We will begin our study of chemistry at the macroscopic level, where we can see and measure the materials of which our world is made. In this chapter, we will discuss the scientific method, which provides the framework for research not only in chemistry but in all other sciences as well. Next we will discover how scientists define and characterize matter. Then we will spend some time learning how to handle numerical results of chemical measurements and solve numerical problems. In Chapter 2, we will begin to explore the microscopic world of atoms and molecules.

化学

The Chinese characters for chemistry mean "The study of change."

1.1 Chemistry: A Science for the Twenty-First Century

Chemistry is *the study of matter and the changes it undergoes.* Chemistry is often called the central science, because a basic knowledge of chemistry is essential for students of biology, physics, geology, ecology, and many other subjects. Indeed, it is central to our way of life; without it, we would be living shorter lives in what we would consider primitive conditions, without automobiles, electricity, computers, CDs, and many other everyday conveniences.

Although chemistry is an ancient science, its modern foundation was laid in the nineteenth century, when intellectual and technological advances enabled scientists to break down substances into ever smaller components and consequently to explain many of their physical and chemical characteristics. The rapid development of increasingly sophisticated technology throughout the twentieth century has given us even greater means to study things that cannot be seen with the naked eye. Using computers and special microscopes, for example, chemists can analyze the structure of atoms and molecules—the fundamental units on which the study of chemistry is based—and design new substances with specific properties, such as drugs and environmentally friendly consumer products.

It is fitting to ask what part the central science will have in the twenty-first century. Almost certainly, chemistry will continue to play a pivotal role in all areas of science and technology. Before plunging into the study of matter and its transformation, let us consider some of the frontiers that chemists are currently exploring (Figure 1.1). Whatever your reasons for taking general chemistry, a good knowledge of the subject will better enable you to appreciate its impact on society and on you as an individual.

1.2 The Study of Chemistry

Compared with other subjects, chemistry is commonly believed to be more difficult, at least at the introductory level. There is some justification for this perception; for one thing, chemistry has a very specialized vocabulary. However, even if this is your first course in chemistry, you already have more familiarity with the subject than you may realize. In everyday conversations we hear words that have a chemical connection, although they may not be used in the scientifically correct sense. Examples are "electronic," "quantum leap," "equilibrium," "catalyst," "chain reaction," and "critical mass." Moreover, if you cook, then you are a practicing chemist! From experience gained in the kitchen, you know that oil and water do not mix and that boiling water left on the stove will evaporate. You apply chemical and physical principles when you use baking soda to leaven bread, choose a pressure cooker to shorten the time it takes to prepare soup, add meat tenderizer to a pot roast, squeeze lemon juice over sliced

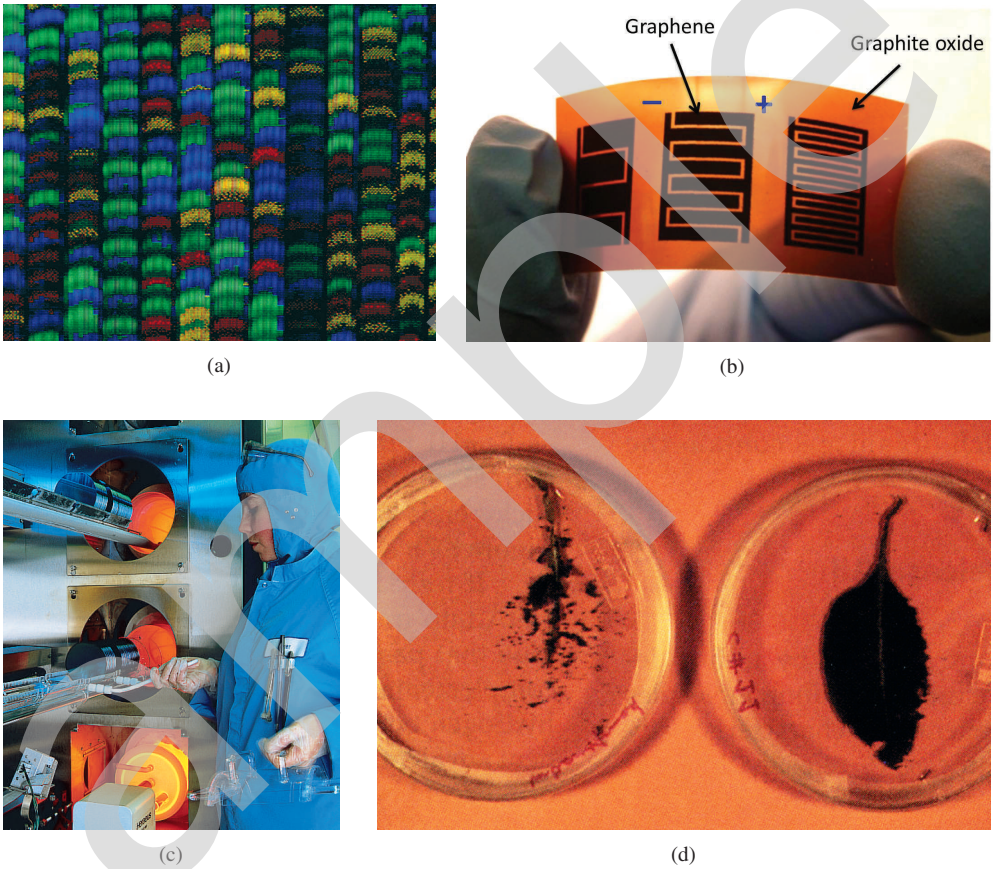


Figure 1.1 *(a) The output from an automated DNA sequencing machine. Each lane displays the sequence (indicated by different colors) obtained with a separate DNA sample. (b) A graphene supercapacitor. These materials provide some of the highest known energy-to-volume ratios and response times. (c) Production of photovoltaic cells, used to convert light into electrical current. (d) The leaf on the left was taken from a tobacco plant that was not genetically engineered but was exposed to tobacco horn worms. The leaf on the right was genetically engineered and is barely attacked by the worms. The same technique can be applied to protect the leaves of other types of plants.*

pears to prevent them from turning brown or over fish to minimize its odor, and add vinegar to the water in which you are going to poach eggs. Every day we observe such changes without thinking about their chemical nature. The purpose of this course is to make you think like a chemist, to look at the *macroscopic world*—the things we can see, touch, and measure directly—and visualize the particles and events of the *microscopic world* that we cannot experience without modern technology and our imaginations.

At first some students find it confusing that their chemistry instructor and textbook seem to be continually shifting back and forth between the macroscopic and microscopic worlds. Just keep in mind that the data for chemical investigations most often come from observations of large-scale phenomena, but the explanations frequently lie in the unseen and partially imagined microscopic world of atoms and molecules. In other words, chemists often *see* one thing (in the macroscopic world) and *think* another (in the microscopic world). Looking at the rusted nails in Figure 1.2, for example, a chemist might think about the basic properties of individual atoms of iron and how these units interact with other atoms and molecules to produce the observed change.

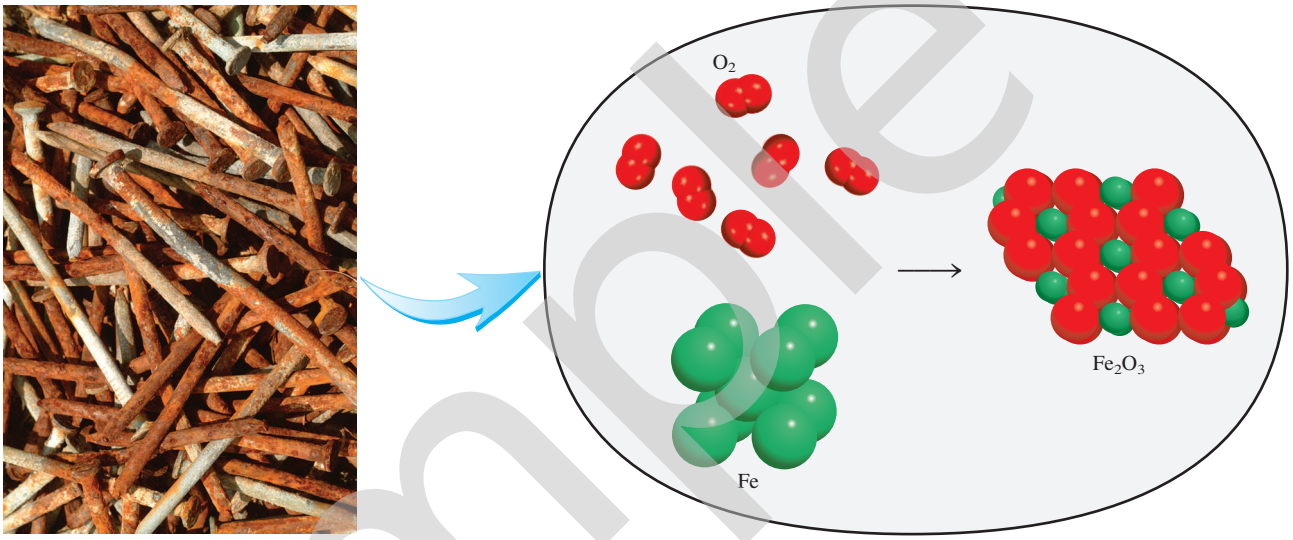


Figure 1.2 *A simplified molecular view of rust (Fe_2O_3) formation from iron (Fe) atoms and oxygen molecules (O_2). In reality, the process requires water and rust also contains water molecules.*

1.3 The Scientific Method

All sciences, including the social sciences, employ variations of what is called the ***scientific method,*** *a systematic approach to research.* For example, a psychologist who wants to know how noise affects people's ability to learn chemistry and a chemist interested in measuring the heat given off when hydrogen gas burns in air would follow roughly the same procedure in carrying out their investigations. The first step is to carefully define the problem. The next step includes performing experiments, making careful observations, and recording information, or *data,* about the system—the part of the universe that is under investigation. (In the examples just discussed, the systems are the group of people the psychologist will study and a mixture of hydrogen and air.)

The data obtained in a research study may be both ***qualitative,*** *consisting of general observations about the system,* and ***quantitative,*** *comprising numbers obtained by various measurements of the system.* Chemists generally use standardized symbols and equations in recording their measurements and observations. This form of representation not only simplifies the process of keeping records, but also provides a common basis for communication with other chemists.

When the experiments have been completed and the data have been recorded, the next step in the scientific method is interpretation, meaning that the scientist attempts to explain the observed phenomenon. Based on the data that were gathered, the researcher formulates a ***hypothesis,*** *a tentative explanation for a set of observations.* Further experiments are devised to test the validity of the hypothesis in as many ways as possible, and the process begins anew. Figure 1.3 summarizes the main steps of the research process.

After a large amount of data has been collected, it is often desirable to summarize the information in a concise way, as a law. In science, a ***law*** is *a concise verbal or mathematical statement of a relationship between phenomena that is always the same under the same conditions.* For example, Sir Isaac Newton's second law of motion, which you may remember from high school science, says that force equals mass times acceleration ($F = ma$). What this law means is that an

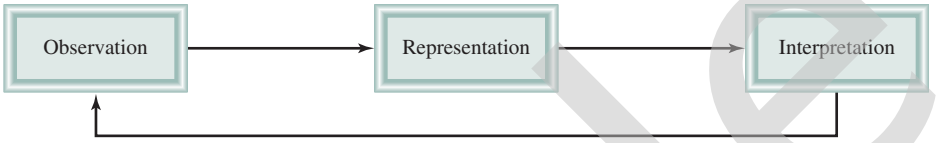


Figure 1.3 *The three levels of studying chemistry and their relationships. Observation deals with events in the macroscopic world; atoms and molecules constitute the microscopic world. Representation is a scientific shorthand for describing an experiment in symbols and chemical equations. Chemists use their knowledge of atoms and molecules to explain an observed phenomenon.*

increase in the mass or in the acceleration of an object will always increase its force proportionally, and a decrease in mass or acceleration will always decrease the force.

Hypotheses that survive many experimental tests of their validity may evolve into theories. A ***theory*** is *a unifying principle that explains a body of facts and/or those laws that are based on them.* Theories, too, are constantly being tested. If a theory is disproved by experiment, then it must be discarded or modified so that it becomes consistent with experimental observations. Proving or disproving a theory can take years, even centuries, in part because the necessary technology may not be available. Atomic theory, which we will study in Chapter 2, is a case in point. It took more than 2000 years to work out this fundamental principle of chemistry proposed by Democritus, an ancient Greek philosopher. A more contemporary example is the search for the Higgs boson discussed on page 6.

Scientific progress is seldom, if ever, made in a rigid, step-by-step fashion. Sometimes a law precedes a theory; sometimes it is the other way around. Two scientists may start working on a project with exactly the same objective, but will end up taking drastically different approaches. Scientists are, after all, human beings, and their modes of thinking and working are very much influenced by their background, training, and personalities.

The development of science has been irregular and sometimes even illogical. Great discoveries are usually the result of the cumulative contributions and experience of many workers, even though the credit for formulating a theory or a law is usually given to only one individual. There is, of course, an element of luck involved in scientific discoveries, but it has been said that "chance favors the prepared mind." It takes an alert and well-trained person to recognize the significance of an accidental discovery and to take full advantage of it. More often than not, the public learns only of spectacular scientific breakthroughs. For every success story, however, there are hundreds of cases in which scientists have spent years working on projects that ultimately led to a dead end, and in which positive achievements came only after many wrong turns and at such a slow pace that they went unheralded. Yet even the dead ends contribute something to the continually growing body of knowledge about the physical universe. It is the love of the search that keeps many scientists in the laboratory.

Review of Concepts

Which of the following statements is true?

(a) A hypothesis always leads to the formulation of a law.
(b) The scientific method is a rigid sequence of steps in solving problems.
(c) A law summarizes a series of experimental observations; a theory provides an explanation for the observations.

CHEMISTRY *in Action*

The Search for the Higgs Boson

In this chapter, we identify mass as a fundamental property of matter, but have you ever wondered: Why does matter even have mass? It might seem obvious that "everything" has mass, but is that a requirement of nature? We will see later in our studies that light is composed of particles that do not have mass when at rest, and physics tells us under different circumstances the universe might not contain *anything* with mass. Yet we know that *our* universe is made up of an uncountable number of particles with mass, and these building blocks are necessary to form the elements that make up the people to ask such questions. The search for the answer to this question illustrates nicely the process we call the scientific method.

Current theoretical models tell us that everything in the universe is based on two types of elementary particles: bosons and fermions. We can distinguish the roles of these particles by considering the building blocks of matter to be constructed from fermions, while bosons are particles responsible for the force that holds the fermions together. In 1964, three different research teams independently proposed mechanisms in which a field of energy permeates the universe, and the interaction of matter with this field is due to a specific boson associated with the field. The greater the number of these bosons, the greater the interaction will be with the field. This interaction is the property we call mass, and the field and the associated boson came to be named for Peter Higgs, one of the original physicists to propose this mechanism.

This theory ignited a frantic search for the "Higgs boson" that became one of the most heralded quests in modern science. The Large Hadron Collider at CERN in Geneva, Switzerland (described on p. 875) was constructed to carry out experiments designed to find evidence for the Higgs boson. In these experiments, protons are accelerated to nearly the speed of light in opposite directions in a circular 17-mile tunnel, and then allowed to collide, generating even more fundamental particles at very high energies. The data are examined for evidence of an excess of particles at an energy consistent with theoretical predictions for the Higgs boson. The ongoing process of theory suggesting experiments that give results used to evaluate and ultimately refine the theory, and so on, is the essence of the scientific method.

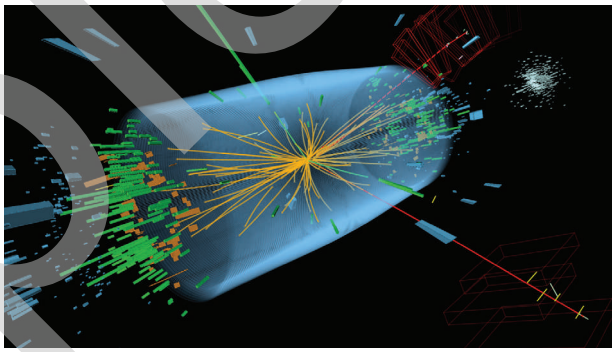


Illustration of the data obtained from decay of the Higgs boson into other particles following an 8-TeV collision at the Large Hadron Collider at CERN.

On July 4, 2012, scientists at CERN announced the discovery of the Higgs boson. It takes about 1 trillion proton-proton collisions to produce one Higgs boson event, so it requires a tremendous amount of data obtained from two independent sets of experiments to confirm the findings. In science, the quest for answers is never completely done. Our understanding can always be improved or refined, and sometimes entire tenets of accepted science are replaced by another theory that does a better job explaining the observations. For example, scientists are not sure if the Higgs boson is the only particle that confers mass to matter, or if it is only one of several such bosons predicted by other theories.

But over the long run, the scientific method has proven to be our best way of understanding the physical world. It took 50 years for experimental science to validate the existence of the Higgs boson. This discovery was greeted with great fanfare and recognized the following year with a 2013 Nobel Prize in Physics for Peter Higgs and François Englert, another one of the six original scientists who first proposed the existence of a universal field that gives particles their mass. It is impossible to imagine where science will take our understanding of the universe in the next 50 years, but we can be fairly certain that many of the theories and experiments driving this scientific discovery will be very different than the ones we use today.

1.4 Classifications of Matter

We defined chemistry in Section 1.1 as the study of matter and the changes it undergoes. ***Matter*** is *anything that occupies space and has mass.* Matter includes things we can see and touch (such as water, earth, and trees), as well as things we cannot (such as air). Thus, everything in the universe has a "chemical" connection.

Chemists distinguish among several subcategories of matter based on composition and properties. The classifications of matter include substances, mixtures, elements, and compounds, as well as atoms and molecules, which we will consider in Chapter 2.

Substances and Mixtures

A ***substance*** is *a form of matter that has a definite (constant) composition and distinct properties.* Examples are water, ammonia, table sugar (sucrose), gold, and oxygen. Substances differ from one another in composition and can be identified by their appearance, smell, taste, and other properties.

A ***mixture*** is *a combination of two or more substances in which the substances retain their distinct identities.* Some familiar examples are air, soft drinks, milk, and cement. Mixtures do not have constant composition. Therefore, samples of air collected in different cities would probably differ in composition because of differences in altitude, pollution, and so on.

Mixtures are either homogeneous or heterogeneous. When a spoonful of sugar dissolves in water we obtain a ***homogeneous mixture*** in which *the composition of the mixture is the same throughout.* If sand is mixed with iron filings, however, the sand grains and the iron filings remain separate (Figure 1.4). This type of mixture is called a ***heterogeneous mixture*** because *the composition is not uniform.*

Any mixture, whether homogeneous or heterogeneous, can be created and then separated by physical means into pure components without changing the identities of the components. Thus, sugar can be recovered from a water solution by heating the solution and evaporating it to dryness. Condensing the vapor will give us back the water component. To separate the iron-sand mixture, we can use a magnet to remove the iron filings from the sand, because sand is not attracted to the magnet [see Figure 1.4(b)]. After separation, the components of the mixture will have the same composition and properties as they did to start with.

Elements and Compounds

Substances can be either elements or compounds. An ***element*** is *a substance that cannot be separated into simpler substances by chemical means.* To date, 118 elements have been positively identified. Most of them occur naturally on Earth. The others have been created by scientists via nuclear processes, which are the subject of Chapter 19 of this text.

(a) (b)

Figure 1.4 *(a) The mixture contains iron filings and sand. (b) A magnet separates the iron filings from the mixture. The same technique is used on a larger scale to separate iron and steel from nonmagnetic objects such as aluminum, glass, and plastics.*

Table 1.1 Some Common Elements and Their Symbols

Name	Symbol	Name	Symbol	Name	Symbol
Aluminum	Al	Fluorine	F	Oxygen	O
Arsenic	As	Gold	Au	Phosphorus	P
Barium	Ba	Hydrogen	H	Platinum	Pt
Bismuth	Bi	Iodine	I	Potassium	K
Bromine	Br	Iron	Fe	Silicon	Si
Calcium	Ca	Lead	Pb	Silver	Ag
Carbon	C	Magnesium	Mg	Sodium	Na
Chlorine	Cl	Manganese	Mn	Sulfur	S
Chromium	Cr	Mercury	Hg	Tin	Sn
Cobalt	Co	Nickel	Ni	Tungsten	W
Copper	Cu	Nitrogen	N	Zinc	Zn

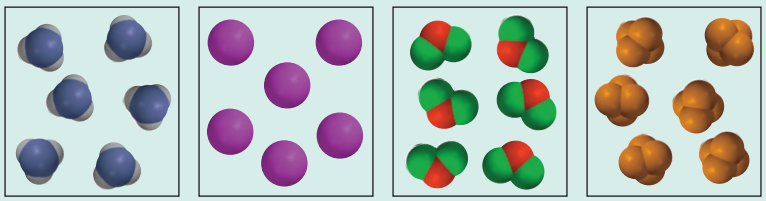
For convenience, chemists use symbols of one or two letters to represent the elements. The first letter of a symbol is *always* capitalized, but any following letters are not. For example, Co is the symbol for the element cobalt, whereas CO is the formula for the carbon monoxide molecule. Table 1.1 shows the names and symbols of some of the more common elements; a complete list of the elements and their symbols appears inside the front cover of this book. The symbols of some elements are derived from their Latin names—for example, Au from *aurum* (gold), Fe from *ferrum* (iron), and Na from *natrium* (sodium)—whereas most of them come from their English names. Appendix 1 gives the origin of the names and lists the discoverers of most of the elements.

Atoms of most elements can interact with one another to form compounds. Hydrogen gas, for example, burns in oxygen gas to form water, which has properties that are distinctly different from those of the starting materials. Water is made up of two parts hydrogen and one part oxygen. This composition does not change, regardless of whether the water comes from a faucet in the United States, a lake in Outer Mongolia, or the ice caps on Mars. Thus, water is a ***compound,*** *a substance composed of atoms of two or more elements chemically united in fixed proportions.* Unlike mixtures, compounds can be separated only by chemical means into their pure components.

The relationships among elements, compounds, and other categories of matter are summarized in Figure 1.5.

Review of Concepts

Which of the following diagrams represent elements and which represent compounds? Each color sphere (or truncated sphere) represents an atom. Different colored atoms indicate different elements.



(a) (b) (c) (d)

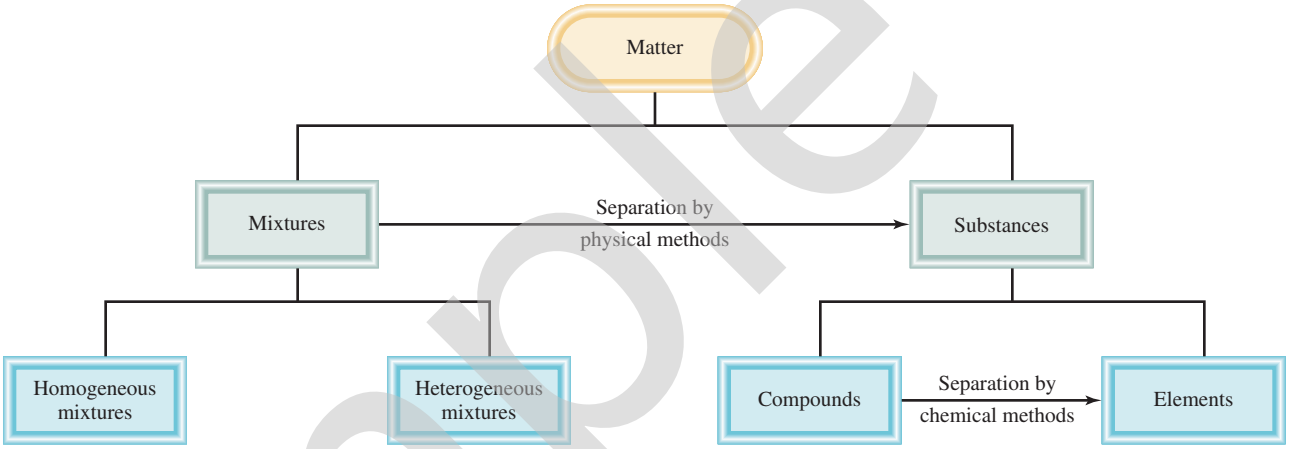


Figure 1.5 *Classification of matter.*

1.5 The Three States of Matter

All substances, at least in principle, can exist in three states: solid, liquid, and gas. As Figure 1.6 shows, gases differ from liquids and solids in the distances between the molecules. In a solid, molecules are held close together in an orderly fashion with little freedom of motion. Molecules in a liquid are close together but are not held so rigidly in position and can move past one another. In a gas, the molecules are separated by distances that are large compared with the size of the molecules.

The three states of matter can be interconverted without changing the composition of the substance. Upon heating, a solid (for example, ice) will melt to form a liquid (water). (The temperature at which this transition occurs is called the *melting point.*) Further heating will convert the liquid into a gas. (This conversion takes place at the *boiling point* of the liquid.) On the other hand, cooling a gas will cause it to condense into a liquid. When the liquid is cooled further, it will freeze into the solid form.

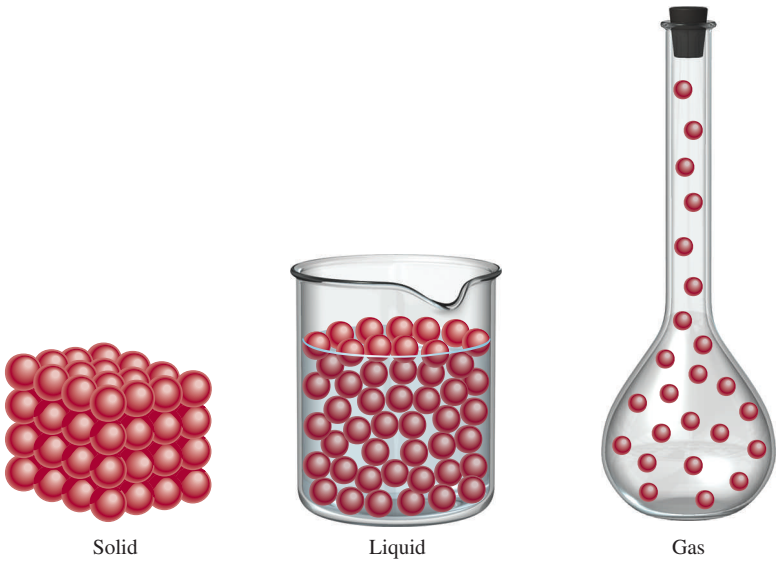


Figure 1.6 *Microscopic views of a solid, a liquid, and a gas.*

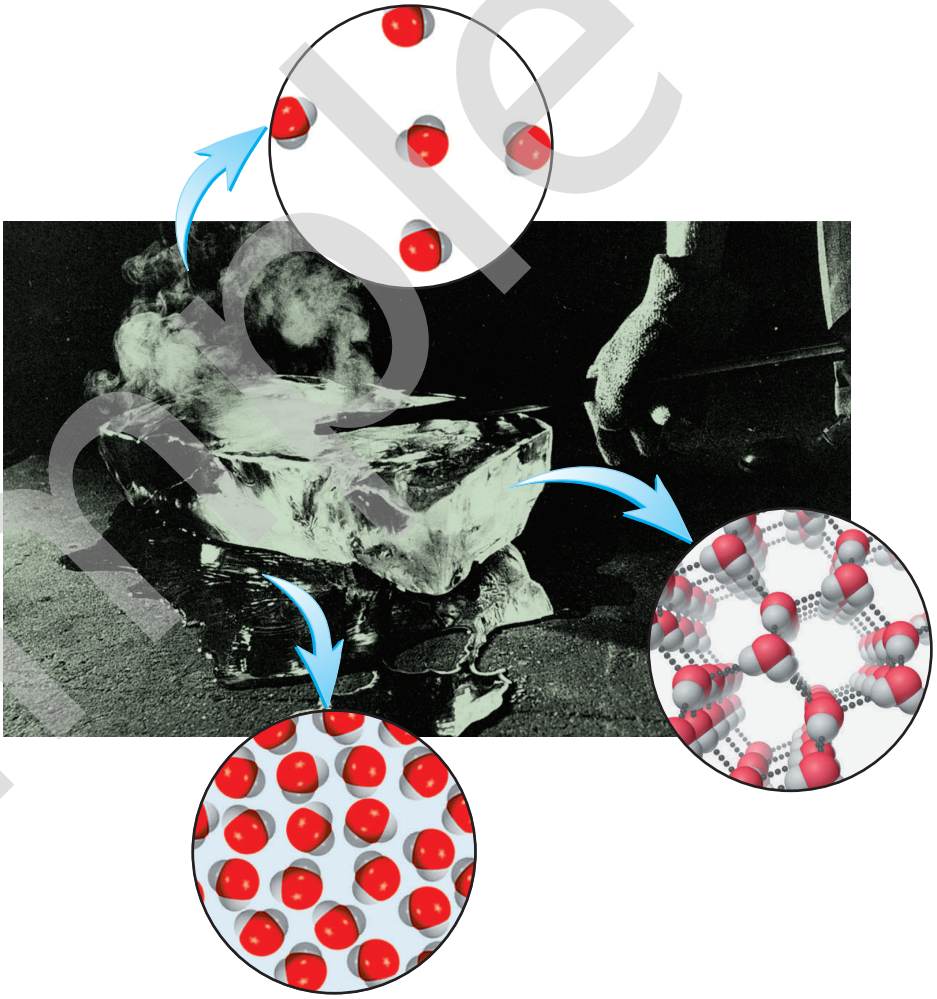


Figure 1.7 *The three states of matter. A hot poker changes ice into water and steam.*

Figure 1.7 shows the three states of water. Note that the properties of water are unique among common substances in that the molecules in the liquid state are more closely packed than those in the solid state.

Review of Concepts

An ice cube is placed in a closed container. On heating, the ice cube first melts and the water then boils to form steam. Which of the following statements is true?

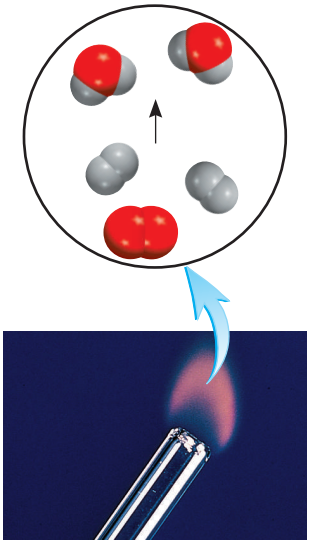
(a) The physical appearance of the water is different at every stage of change.
(b) The mass of water is greatest for the ice cube and least for the steam.

1.6 Physical and Chemical Properties of Matter

Substances are identified by their properties as well as by their composition. Color, melting point, and boiling point are physical properties. A ***physical property*** *can be measured and observed without changing the composition or identity of a substance.* For example, we can measure the melting point of ice by heating a block of ice and recording the temperature at which the ice is converted to water. Water differs from ice only in appearance, not in composition, so this is a physical change; we can freeze

the water to recover the original ice. Therefore, the melting point of a substance is a physical property. Similarly, when we say that helium gas is lighter than air, we are referring to a physical property.

On the other hand, the statement "Hydrogen gas burns in oxygen gas to form water" describes a ***chemical property*** of hydrogen, because *to observe this property we must carry out a chemical change,* in this case burning. After the change, the original chemical substance, the hydrogen gas, will have vanished, and all that will be left is a different chemical substance—water. We *cannot* recover the hydrogen from the water by means of a physical change, such as boiling or freezing.



Hydrogen burning in air to form water.

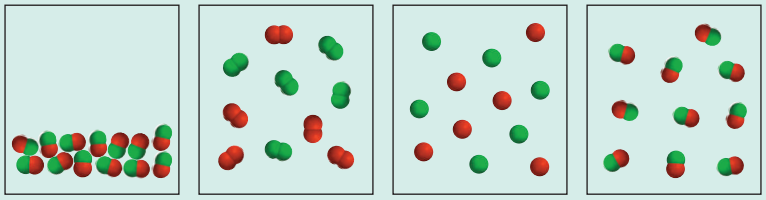
Every time we hard-boil an egg, we bring about a chemical change. When subjected to a temperature of about 100°C, the yolk and the egg white undergo changes that alter not only their physical appearance but their chemical makeup as well. When eaten, the egg is changed again, by substances in our bodies called *enzymes.* This digestive action is another example of a chemical change. What happens during digestion depends on the chemical properties of both the enzymes and the food.

All measurable properties of matter fall into one of two additional categories: extensive properties and intensive properties. The measured value of an ***extensive property*** *depends on how much matter is being considered.* ***Mass,*** which is *the quantity of matter in a given sample of a substance,* is an extensive property. More matter means more mass. Values of the same extensive property can be added together. For example, two copper pennies will have a combined mass that is the sum of the masses of each penny, and the length of two tennis courts is the sum of the lengths of each tennis court. ***Volume,*** defined as *length cubed,* is another extensive property. The value of an extensive quantity depends on the amount of matter.

The measured value of an ***intensive property*** *does not depend on how much matter is being considered.* ***Density,*** defined as *the mass of an object divided by its volume,* is an intensive property. So is temperature. Suppose that we have two beakers of water at the same temperature. If we combine them to make a single quantity of water in a larger beaker, the temperature of the larger quantity of water will be the same as it was in two separate beakers. Unlike mass, length, and volume, temperature and other intensive properties are not additive.

Review of Concepts

The diagram in (a) shows a compound made up of atoms of two elements (represented by the green and red spheres) in the liquid state. Which of the diagrams in (b)–(d) represents a physical change and which diagrams represent a chemical change?



(a) (b) (c) (d)

1.7 Measurement

The measurements chemists make are often used in calculations to obtain other related quantities. Different instruments enable us to measure a substance's properties: The meterstick measures length or scale; the buret, the pipet, the graduated cylinder, and

the volumetric flask measure volume (Figure 1.8); the balance measures mass; the thermometer measures temperature. These instruments provide measurements of ***macroscopic properties,*** which *can be determined directly.* ***Microscopic properties,*** *on the atomic or molecular scale, must be determined by an indirect method,* as we will see in Chapter 2.

A measured quantity is usually written as a number with an appropriate unit. To say that the distance between New York and San Francisco by car along a certain route is 5166 is meaningless. We must specify that the distance is 5166 kilometers. The same is true in chemistry; units are essential to stating measurements correctly.

SI Units

For many years, scientists recorded measurements in *metric units,* which are related decimally, that is, by powers of 10. In 1960, however, the General Conference of Weights and Measures, the international authority on units, proposed a revised metric system called the ***International System of Units*** (abbreviated ***SI,*** from the French *S*ystème *I*nternationale d'Unites). Table 1.2 shows the seven SI base units. All other units of measurement can be derived from these base units. Like metric units, SI units are modified in decimal fashion by a series of prefixes, as shown in Table 1.3. We will use both metric and SI units in this book.

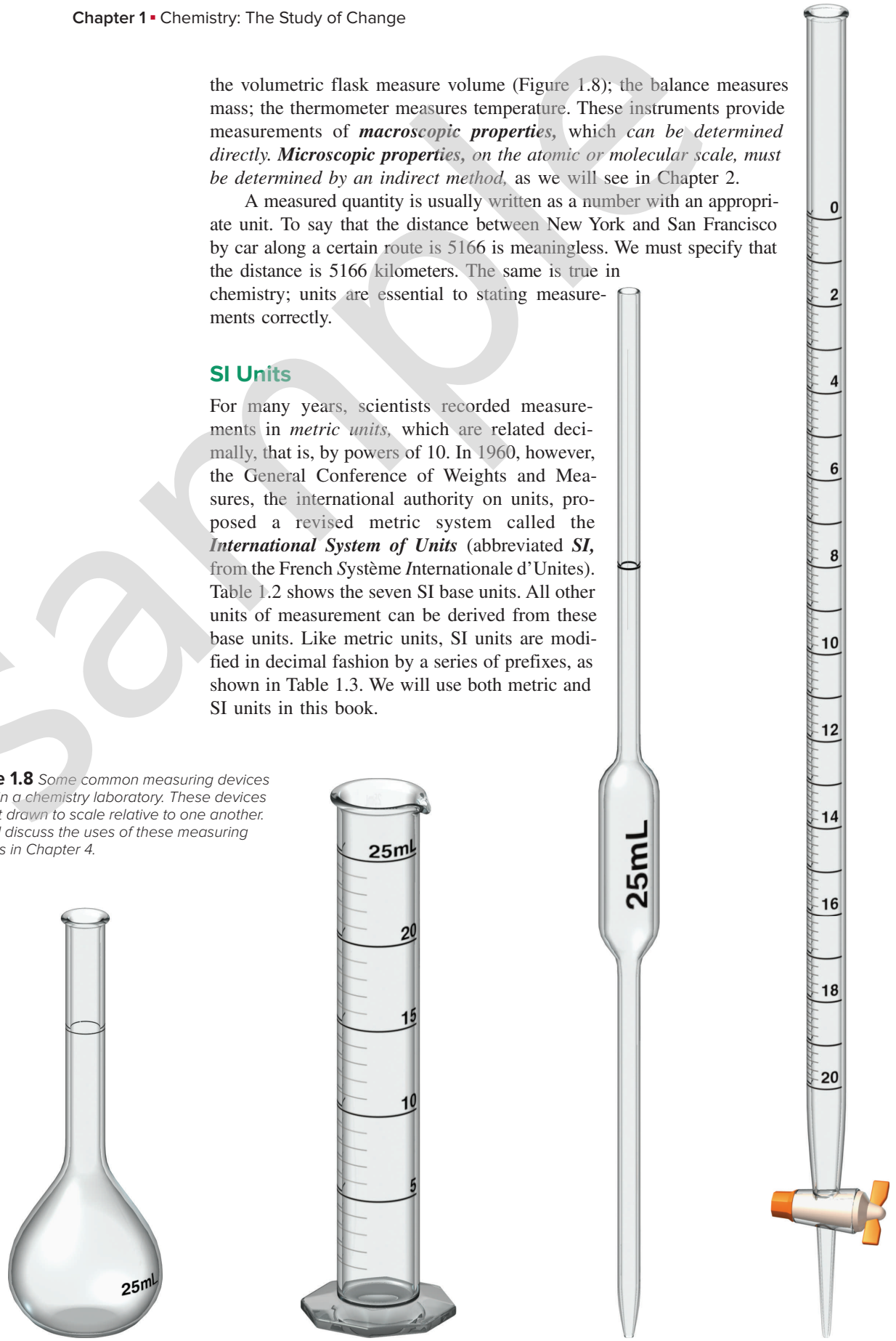


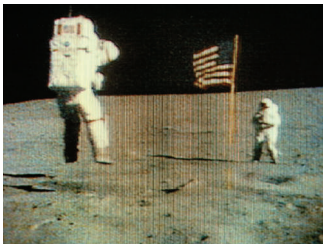
Figure 1.8 *Some common measuring devices found in a chemistry laboratory. These devices are not drawn to scale relative to one another. We will discuss the uses of these measuring devices in Chapter 4.*

Table 1.2 SI Base Units

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

Note that a metric prefix simply represents a number:

$1 \text{ mm} = 1 \times 10^{-3} \text{ m}$



An astronaut jumping on the surface of the moon.

Table 1.3 Prefixes Used with SI Units

Prefix	Symbol	Meaning	Example
tera-	T	1,000,000,000,000, or 10^{12}	1 terameter (Tm) $= 1 \times 10^{12}$ m
giga-	G	1,000,000,000, or 10^{9}	1 gigameter (Gm) $= 1 \times 10^{9}$ m
mega-	M	1,000,000, or 10^{6}	1 megameter (Mm) $= 1 \times 10^{6}$ m
kilo-	k	1,000, or 10^{3}	1 kilometer (km) $= 1 \times 10^{3}$ m
deci-	d	1/10, or 10^{-1}	1 decimeter (dm) = 0.1 m
centi-	c	1/100, or 10^{-2}	1 centimeter (cm) = 0.01 m
milli-	m	1/1,000, or 10^{-3}	1 millimeter (mm) = 0.001 m
micro-	μ	1/1,000,000, or 10^{-6}	1 micrometer (μm) $= 1 \times 10^{-6}$ m
nano-	n	1/1,000,000,000, or 10^{-9}	1 nanometer (nm) $= 1 \times 10^{-9}$ m
pico-	p	1/1,000,000,000,000, or 10^{-12}	1 picometer (pm) $= 1 \times 10^{-12}$ m
femto-	f	1/1,000,000,000,000,000, or 10^{-15}	1 femtometer (fm) $= 1 \times 10^{-15}$ m
atto-	a	1/1,000,000,000,000,000,000 or 10^{-18}	1 attometer (am) $= 1 \times 10^{-18}$ m

Measurements that we will utilize frequently in our study of chemistry include time, mass, volume, density, and temperature.

Mass and Weight

The terms "mass" and "weight" are often used interchangeably, although, strictly speaking, they are different quantities. Whereas mass is a measure of the amount of matter in an object, ***weight,*** technically speaking, is *the force that gravity exerts on an object.* An apple that falls from a tree is pulled downward by Earth's gravity. The mass of the apple is constant and does not depend on its location, but its weight does. For example, on the surface of the moon the apple would weigh only one-sixth what it does on Earth, because the moon's gravity is only one-sixth that of Earth. The moon's smaller gravity enabled astronauts to jump about rather freely on its surface despite their bulky suits and equipment. Chemists are interested primarily in mass, which can be determined readily with a balance; the process of measuring mass, oddly, is called *weighing.*

The SI unit of mass is the *kilogram* (kg). Unlike the units of length and time, which are based on natural processes that can be repeated by scientists anywhere, the kilogram is defined in terms of a particular object (Figure 1.9). In chemistry, however, the smaller *gram* (g) is more convenient:

$$1 \text{ kg} = 1000 \text{ g} = 1 \times 10^{3} \text{ g}$$

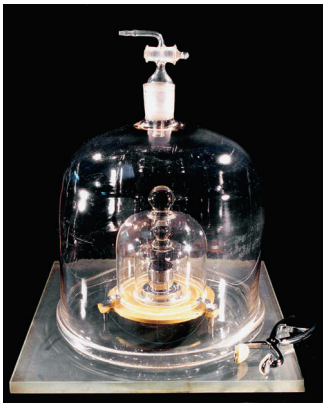


Figure 1.9 *The prototype kilogram is made of a platinum-iridium alloy. It is kept in a vault at the International Bureau of Weights and Measures in Sèvres, France. In 2007 it was discovered that the alloy has mysteriously lost about 50 μg!*

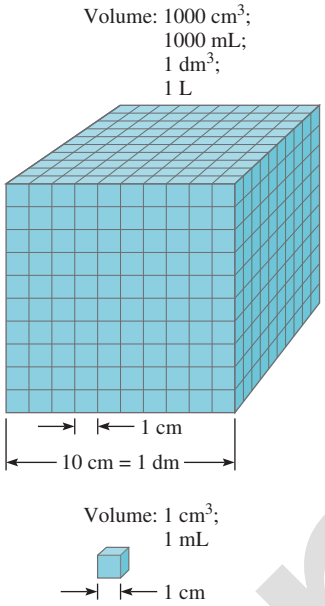


Figure 1.10 *Comparison of two volumes, 1 mL and 1000 mL.*

Volume

The SI unit of length is the *meter* (m), and the SI-derived unit for volume is the *cubic meter* (m^3). Generally, however, chemists work with much smaller volumes, such as the cubic centimeter (cm^3) and the cubic decimeter (dm^3):

$$1\ \text{cm}^3 = (1 \times 10^{-2}\ \text{m})^3 = 1 \times 10^{-6}\ \text{m}^3$$
$$1\ \text{dm}^3 = (1 \times 10^{-1}\ \text{m})^3 = 1 \times 10^{-3}\ \text{m}^3$$

Another common unit of volume is the liter (L). A ***liter*** is *the volume occupied by one cubic decimeter.* One liter of volume is equal to 1000 milliliters (mL) or 1000 cm^3:

$$\begin{aligned} 1\ \text{L} &= 1000\ \text{mL} \\ &= 1000\ \text{cm}^3 \\ &= 1\ \text{dm}^3 \end{aligned}$$

and one milliliter is equal to one cubic centimeter:

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

Figure 1.10 compares the relative sizes of two volumes. Even though the liter is not an SI unit, volumes are usually expressed in liters and milliliters.

Density

The equation for density is

$$\text{density} = \frac{\text{mass}}{\text{volume}}$$

or

$$d = \frac{m}{V} \tag{1.1}$$

where *d, m,* and *V* denote density, mass, and volume, respectively. Because density is an intensive property and does not depend on the quantity of mass present, for a given substance the ratio of mass to volume always remains the same; in other words, *V* increases as *m* does. Density usually decreases with temperature.

The SI-derived unit for density is the kilogram per cubic meter (kg/m^3). This unit is awkwardly large for most chemical applications. Therefore, grams per cubic centimeter (g/cm^3) and its equivalent, grams per milliliter (g/mL), are more commonly used for solid and liquid densities. Because gas densities are often very low, we express them in units of grams per liter (g/L):

$$1\ \text{g/cm}^3 = 1\ \text{g/mL} = 1000\ \text{kg/m}^3$$
$$1\ \text{g/L} = 0.001\ \text{g/mL}$$

Table 1.4 lists the densities of several substances.

Table 1.4

Densities of Some Substances at 25°C

Substance	Density (g/cm^3)
Air*	0.001
Ethanol	0.79
Water	1.00
Graphite	2.2
Table salt	2.2
Aluminum	2.70
Diamond	3.5
Iron	7.9
Lead	11.3
Mercury	13.6
Gold	19.3
Osmium†	22.6

*Measured at 1 atmosphere.
†Osmium (Os) is the densest element known.

Examples 1.1 and 1.2 show density calculations.

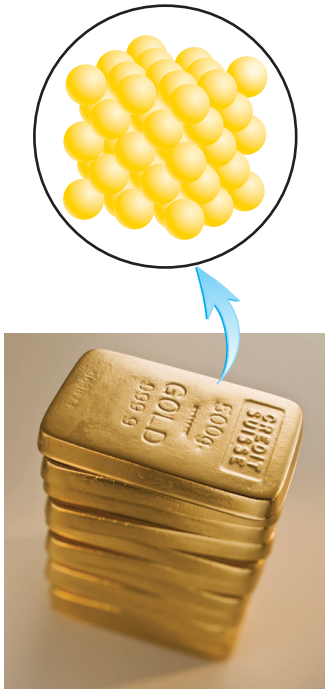
Example 1.1

Gold is a precious metal that is chemically unreactive. It is used mainly in jewelry, dentistry, and electronic devices. A piece of gold ingot with a mass of 301 g has a volume of 15.6 cm^3. Calculate the density of gold.

Solution We are given the mass and volume and asked to calculate the density. Therefore, from Equation (1.1), we write

$$\begin{aligned} d &= \frac{m}{V} \\ &= \frac{301\ \text{g}}{15.6\ \text{cm}^3} \\ &= 19.3\ \text{g/cm}^3 \end{aligned}$$

Practice Exercise A piece of platinum metal with a density of 21.5 g/cm^3 has a volume of 4.49 cm^3. What is its mass?



Gold bars and the solid-state arrangement of the gold atoms.

Similar problems: 1.21, 1.22.

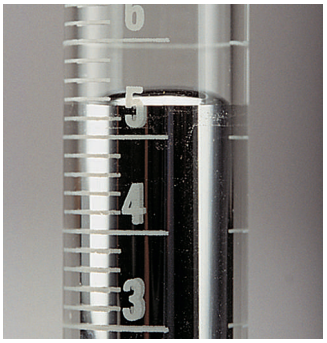
Example 1.2

The density of mercury, the only metal that is a liquid at room temperature, is 13.6 g/mL. Calculate the mass of 5.50 mL of the liquid.

Solution We are given the density and volume of a liquid and asked to calculate the mass of the liquid. We rearrange Equation (1.1) to give

$$\begin{aligned} m &= d \times V \\ &= 13.6\ \frac{\text{g}}{\cancel{\text{mL}}} \times 5.50\ \cancel{\text{mL}} \\ &= 74.8\ \text{g} \end{aligned}$$

Practice Exercise The density of sulfuric acid in a certain car battery is 1.41 g/mL. Calculate the mass of 242 mL of the liquid.



Mercury.

Similar problems: 1.21, 1.22.

Temperature Scales

Three temperature scales are currently in use. Their units are °F (degrees Fahrenheit), °C (degrees Celsius), and K (kelvin). The Fahrenheit scale, which is the most commonly used scale in the United States outside the laboratory, defines the normal freezing and boiling points of water to be exactly 32°F and 212°F, respectively. The Celsius scale divides the range between the freezing point (0°C) and boiling point (100°C) of water into 100 degrees. As Table 1.2 shows, the ***kelvin*** is *the SI base unit of temperature:* It is the *absolute* temperature scale. By absolute we mean that the zero on the Kelvin scale, denoted by 0 K, is the lowest temperature that can be attained theoretically. On the other hand, 0°F and 0°C are based on the behavior of an arbitrarily chosen substance, water. Figure 1.11 compares the three temperature scales.

Note that the Kelvin scale does not have the degree sign. Also, temperatures expressed in kelvins can never be negative.

The size of a degree on the Fahrenheit scale is only 100/180, or 5/9, of a degree on the Celsius scale. To convert degrees Fahrenheit to degrees Celsius, we write

$$?°\text{C} = (°\text{F} - 32°\text{F}) \times \frac{5°\text{C}}{9°\text{F}} \tag{1.2}$$

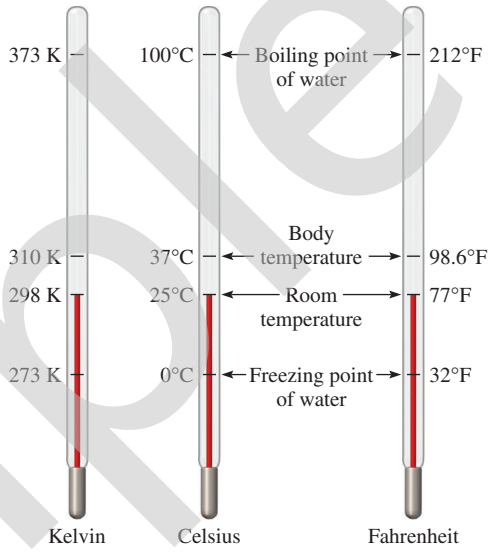


Figure 1.11 *Comparison of the three temperature scales: Celsius, and Fahrenheit, and the absolute (Kelvin) scales. Note that there are 100 divisions, or 100 degrees, between the freezing point and the boiling point of water on the Celsius scale, and there are 180 divisions, or 180 degrees, between the same two temperature limits on the Fahrenheit scale. The Celsius scale was formerly called the centigrade scale.*

The following equation is used to convert degrees Celsius to degrees Fahrenheit:

$$?°F = \frac{9°F}{5°C} \times (°C) + 32°F \tag{1.3}$$

Both the Celsius and the Kelvin scales have units of equal magnitude; that is, one degree Celsius is equivalent to one kelvin. Experimental studies have shown that absolute zero on the Kelvin scale is equivalent to −273.15°C on the Celsius scale. Thus, we can use the following equation to convert degrees Celsius to kelvin:

$$? K = (°C + 273.15°C) \frac{1 K}{1°C} \tag{1.4}$$

We will frequently find it necessary to convert between degrees Celsius and degrees Fahrenheit and between degrees Celsius and kelvin. Example 1.3 illustrates these conversions.

The Chemistry in Action essay on page 17 shows why we must be careful with units in scientific work.

Magnet suspended above superconductor cooled below its transition temperature by liquid nitrogen.

Example 1.3

(a) Below the transition temperature of −141°C, a certain substance becomes a superconductor; that is, it can conduct electricity with no resistance. What is the temperature in degrees Fahrenheit? (b) Helium has the lowest boiling point of all the elements at −452°F. Convert this temperature to degrees Celsius. (c) Mercury, the only metal that exists as a liquid at room temperature, melts at −38.9°C. Convert its melting point to kelvins.

Solution These three parts require that we carry out temperature conversions, so we need Equations (1.2), (1.3), and (1.4). Keep in mind that the lowest temperature on the Kelvin scale is zero (0 K); therefore, it can never be negative.

(a) This conversion is carried out by writing

$$\frac{9°F}{5°C} \times (-141°C) + 32°F = -222°F$$

(Continued)

CHEMISTRY *in Action*

The Importance of Units

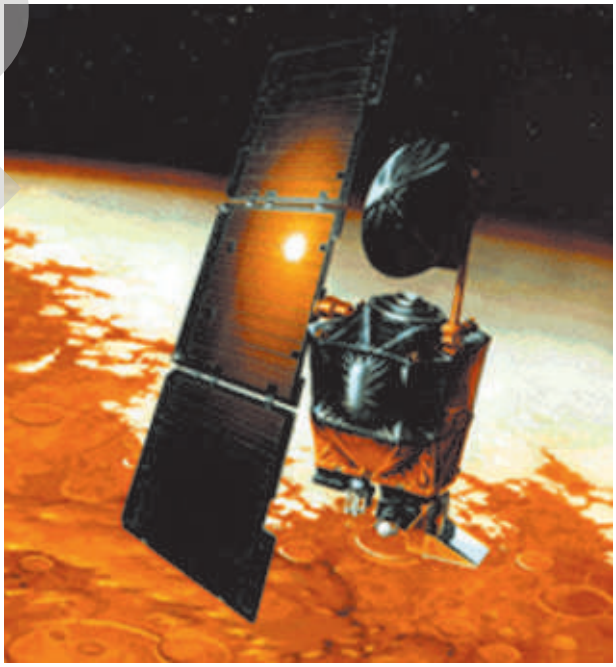
In December 1998, NASA launched the 125-million dollar Mars Climate Orbiter, intended as the red planet's first weather satellite. After a 416-million mi journey, the spacecraft was supposed to go into Mars' orbit on September 23, 1999. Instead, it entered Mars' atmosphere about 100 km (62 mi) lower than planned and was destroyed by heat. The mission controllers said the loss of the spacecraft was due to the failure to convert English measurement units into metric units in the navigation software.

Engineers at Lockheed Martin Corporation who built the spacecraft specified its thrust in pounds, which is an English unit. Scientists at NASA's Jet Propulsion Laboratory, on the other hand, had assumed that thrust data they received were expressed in metric units, as newtons. Normally, pound is the unit for mass. Expressed as a unit for force, however, 1 lb is the force due to gravitational attraction on an object of that mass. To carry out the conversion between pound and newton, we start with 1 lb = 0.4536 kg and from Newton's second law of motion,

$$\begin{aligned} \text{force} &= \text{mass} \times \text{acceleration} \\ &= 0.4536\ \text{kg} \times 9.81\ \text{m/s}^2 \\ &= 4.45\ \text{kg m/s}^2 \\ &= 4.45\ \text{N} \end{aligned}$$

because 1 newton (N) = 1 kg m/s^2. Therefore, instead of converting 1 lb of force to 4.45 N, the scientists treated it as 1 N.

The considerably smaller engine thrust expressed in newtons resulted in a lower orbit and the ultimate destruction of the spacecraft. Commenting on the failure of the Mars mission, one scientist said: "This is going to be the cautionary tale that will be embedded into introduction to the metric system in elementary school, high school, and college science courses till the end of time."



Artist's conception of the Martian Climate Orbiter.

(b) Here we have

$$(-452°\text{F} - 32°\text{F}) \times \frac{5°\text{C}}{9°\text{F}} = -269°\text{C}$$

(c) The melting point of mercury in kelvins is given by

$$(-38.9°\text{C} + 273.15°\text{C}) \times \frac{1\ \text{K}}{1°\text{C}} = 234.3\ \text{K}$$

Similar problems: 1.24, 1.25, 1.26.

Practice Exercise Convert (a) 327.5°C (the melting point of lead) to degrees Fahrenheit; (b) 172.9°F (the boiling point of ethanol) to degrees Celsius; and (c) 77 K, the boiling point of liquid nitrogen, to degrees Celsius.

Review of Concepts

The density of copper is 8.94 g/cm^3 at 20°C and 8.91 g/cm^3 at 60°C. This density decrease is the result of which of the following?

(a) The metal expands.
(b) The metal contracts.
(c) The mass of the metal increases.
(d) The mass of the metal decreases.

1.8 Handling Numbers

Having surveyed some of the units used in chemistry, we now turn to techniques for handling numbers associated with measurements: scientific notation and significant figures.

Scientific Notation

Chemists often deal with numbers that are either extremely large or extremely small. For example, in 1 g of the element hydrogen there are roughly

$$602{,}200{,}000{,}000{,}000{,}000{,}000{,}000$$

hydrogen atoms. Each hydrogen atom has a mass of only

$$0.000000000000000000000000166 \text{ g}$$

These numbers are cumbersome to handle, and it is easy to make mistakes when using them in arithmetic computations. Consider the following multiplication:

$$0.0000000056 \times 0.00000000048 = 0.000000000000000002688$$

It would be easy for us to miss one zero or add one more zero after the decimal point. Consequently, when working with very large and very small numbers, we use a system called *scientific notation.* Regardless of their magnitude, all numbers can be expressed in the form

$$N \times 10^n$$

where N is a number between 1 and 10 and n, the exponent, is a positive or negative integer (whole number). Any number expressed in this way is said to be written in scientific notation.

Suppose that we are given a certain number and asked to express it in scientific notation. Basically, this assignment calls for us to find n. We count the number of places that the decimal point must be moved to give the number N (which is between 1 and 10). If the decimal point has to be moved to the left, then n is a positive integer; if it has to be moved to the right, n is a negative integer. The following examples illustrate the use of scientific notation:

(1) Express 568.762 in scientific notation:

$$568.762 = 5.68762 \times 10^2$$

Note that the decimal point is moved to the left by two places and $n = 2$.

(2) Express 0.00000772 in scientific notation:

$$0.00000772 = 7.72 \times 10^{-6}$$

Here the decimal point is moved to the right by six places and $n = -6$.

Keep in mind the following two points. First, $n = 0$ is used for numbers that are not expressed in scientific notation. For example, 74.6×10^0 ($n = 0$) is equivalent to 74.6. Second, the usual practice is to omit the superscript when $n = 1$. Thus, the scientific notation for 74.6 is 7.46×10 and not 7.46×10^1.

Any number raised to the power zero is equal to one.

Next, we consider how scientific notation is handled in arithmetic operations.

Addition and Subtraction

To add or subtract using scientific notation, we first write each quantity—say, N_1 and N_2—with the same exponent n. Then we combine N_1 and N_2; the exponents remain the same. Consider the following examples:

$$\begin{aligned}
(7.4 \times 10^3) + (2.1 \times 10^3) &= 9.5 \times 10^3 \\
(4.31 \times 10^4) + (3.9 \times 10^3) &= (4.31 \times 10^4) + (0.39 \times 10^4) \\
&= 4.70 \times 10^4 \\
(2.22 \times 10^{-2}) - (4.10 \times 10^{-3}) &= (2.22 \times 10^{-2}) - (0.41 \times 10^{-2}) \\
&= 1.81 \times 10^{-2}
\end{aligned}$$

Multiplication and Division

To multiply numbers expressed in scientific notation, we multiply N_1 and N_2 in the usual way, but *add* the exponents together. To divide using scientific notation, we divide N_1 and N_2 as usual and subtract the exponents. The following examples show how these operations are performed:

$$\begin{aligned}
(8.0 \times 10^4) \times (5.0 \times 10^2) &= (8.0 \times 5.0)(10^{4+2}) \\
&= 40 \times 10^6 \\
&= 4.0 \times 10^7 \\
(4.0 \times 10^{-5}) \times (7.0 \times 10^3) &= (4.0 \times 7.0)(10^{-5+3}) \\
&= 28 \times 10^{-2} \\
&= 2.8 \times 10^{-1} \\
\frac{6.9 \times 10^7}{3.0 \times 10^{-5}} &= \frac{6.9}{3.0} \times 10^{7-(-5)} \\
&= 2.3 \times 10^{12} \\
\frac{8.5 \times 10^4}{5.0 \times 10^9} &= \frac{8.5}{5.0} \times 10^{4-9} \\
&= 1.7 \times 10^{-5}
\end{aligned}$$

Significant Figures

Except when all the numbers involved are integers (for example, in counting the number of students in a class), it is often impossible to obtain the exact value of the quantity under investigation. For this reason, it is important to indicate the margin of error in a measurement by clearly indicating the number of ***significant figures,*** which are *the meaningful digits in a measured or calculated quantity.* When significant figures are used, the last digit is understood to be uncertain. For example, we might measure the volume of a given amount of liquid using a graduated cylinder with a scale that gives an uncertainty of 1 mL in the measurement. If the volume is found to be 6 mL, then the actual volume is in the range of 5 mL to 7 mL. We represent the volume of the liquid as (6 ± 1) mL. In this case, there is only one significant figure (the digit 6) that is uncertain by either plus or minus 1 mL. For greater accuracy, we might use a graduated cylinder that has finer divisions, so that the volume we measure is now uncertain by only 0.1 mL. If the volume of the liquid is now found to be 6.0 mL, we may express the quantity as (6.0 ± 0.1) mL, and the actual value is somewhere between 5.9 mL and 6.1 mL.

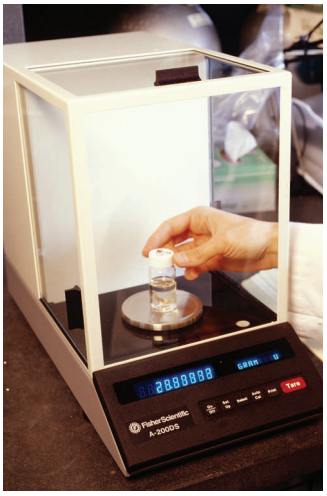


Figure 1.12 *A Fisher Scientific A-200DS Digital Recorder Precision Balance.*

We can further improve the measuring device and obtain more significant figures, but in every case, the last digit is always uncertain; the amount of this uncertainty depends on the particular measuring device we use.

Figure 1.12 shows a modern balance. Balances such as this one are available in many general chemistry laboratories; they readily measure the mass of objects to four decimal places. Therefore, the measured mass typically will have four significant figures (for example, 0.8642 g) or more (for example, 3.9745 g). Keeping track of the number of significant figures in a measurement such as mass ensures that calculations involving the data will reflect the precision of the measurement.

Guidelines for Using Significant Figures

We must always be careful in scientific work to write the proper number of significant figures. In general, it is fairly easy to determine how many significant figures a number has by following these rules:

1. Any digit that is not zero is significant. Thus, 845 cm has three significant figures, 1.234 kg has four significant figures, and so on.
2. Zeros between nonzero digits are significant. Thus, 606 m contains three significant figures, 40,501 kg contains five significant figures, and so on.
3. Zeros to the left of the first nonzero digit are not significant. Their purpose is to indicate the placement of the decimal point. For example, 0.08 L contains one significant figure, 0.0000349 g contains three significant figures, and so on.
4. If a number is greater than 1, then all the zeros written to the right of the decimal point count as significant figures. Thus, 2.0 mg has two significant figures, 40.062 mL has five significant figures, and 3.040 dm has four significant figures. If a number is less than 1, then only the zeros that are at the end of the number and the zeros that are between nonzero digits are significant. This means that 0.090 kg has two significant figures, 0.3005 L has four significant figures, 0.00420 min has three significant figures, and so on.
5. For numbers that do not contain decimal points, the trailing zeros (that is, zeros after the last nonzero digit) may or may not be significant. Thus, 400 cm may have one significant figure (the digit 4), two significant figures (40), or three significant figures (400). We cannot know which is correct without more information. By using scientific notation, however, we avoid this ambiguity. In this particular case, we can express the number 400 as 4×10^2 for one significant figure, 4.0×10^2 for two significant figures, or 4.00×10^2 for three significant figures.

Example 1.4 shows the determination of significant figures.

Example 1.4

Determine the number of significant figures in the following measurements: (a) 394 cm, (b) 5.03 g, (c) 0.714 m, (d) 0.052 kg, (e) 2.720×10^{22} atoms, (f) 3000 mL.

Solution (a) Three, because each digit is a nonzero digit. (b) Three, because zeros between nonzero digits are significant. (c) Three, because zeros to the left of the first nonzero digit do not count as significant figures. (d) Two. Same reason as in (c). (e) Four. Because the number is greater than one, all the zeros written to the right of the decimal point count as significant figures. (f) This is an ambiguous case. The number of significant figures may be four (3.000×10^3), three (3.00×10^3), two

(Continued)

(3.0×10^3), or one (3×10^3). This example illustrates why scientific notation must be used to show the proper number of significant figures.

Similar problems: 1.33, 1.34.

Practice Exercise Determine the number of significant figures in each of the following measurements: (a) 35 mL, (b) 2008 g, (c) 0.0580 m^3, (d) 7.2×10^4 molecules, (e) 830 kg.

A second set of rules specifies how to handle significant figures in calculations.

1. In addition and subtraction, the answer cannot have more digits to the right of the decimal point than either of the original numbers. Consider these examples:

$$\begin{array}{rl} 89.332 & \\ +\ 1.1 & \longleftarrow \text{one digit after the decimal point} \\ \hline 90.432 & \longleftarrow \text{round off to } 90.4 \end{array}$$

$$\begin{array}{rl} 2.097 & \\ -\ 0.12 & \longleftarrow \text{two digits after the decimal point} \\ \hline 1.977 & \longleftarrow \text{round off to } 1.98 \end{array}$$

The rounding-off procedure is as follows. To round off a number at a certain point we simply drop the digits that follow if the first of them is less than 5. Thus, 8.724 rounds off to 8.72 if we want only two digits after the decimal point. If the first digit following the point of rounding off is equal to or greater than 5, we add 1 to the preceding digit. Thus, 8.727 rounds off to 8.73, and 0.425 rounds off to 0.43.

2. In multiplication and division, the number of significant figures in the final product or quotient is determined by the original number that has the *smallest* number of significant figures. The following examples illustrate this rule:

$$2.8 \times 4.5039 = 12.61092 \longleftarrow \text{round off to } 13$$

$$\frac{6.85}{112.04} = 0.0611388789 \longleftarrow \text{round off to } 0.0611$$

3. Keep in mind that *exact numbers* obtained from definitions or by counting numbers of objects can be considered to have an infinite number of significant figures. For example, the inch is defined to be exactly 2.54 centimeters; that is,

$$1 \text{ in} = 2.54 \text{ cm}$$

Thus, the "2.54" in the equation should not be interpreted as a measured number with three significant figures. In calculations involving conversion between "in" and "cm," we treat both "1" and "2.54" as having an infinite number of significant figures. Similarly, if an object has a mass of 5.0 g, then the mass of nine such objects is

$$5.0 \text{ g} \times 9 = 45 \text{ g}$$

The answer has two significant figures because 5.0 g has two significant figures. The number 9 is exact and does not determine the number of significant figures.

Example 1.5 shows how significant figures are handled in arithmetic operations.

Example 1.5

Carry out the following arithmetic operations to the correct number of significant figures: (a) 12,343.2 g + 0.1893 g, (b) 55.67 L − 2.386 L, (c) 7.52 m × 6.9232, (d) 0.0239 kg ÷ 46.5 mL, (e) 5.21×10^3 cm + 2.92×10^2 cm.

(Continued)

Solution In addition and subtraction, the number of decimal places in the answer is determined by the number having the lowest number of decimal places. In multiplication and division, the significant number of the answer is determined by the number having the smallest number of significant figures.

(a)
$$\begin{array}{r} 12{,}343.2\ \text{g} \\ +\quad 0.1893\ \text{g} \\ \hline 12{,}343.3893\ \text{g} \end{array} \longleftarrow \text{round off to } 12{,}343.4\ \text{g}$$

(b)
$$\begin{array}{r} 55.67\ \text{L} \\ -\ 2.386\ \text{L} \\ \hline 53.284\ \text{L} \end{array} \longleftarrow \text{round off to } 53.28\ \text{L}$$

(c) $7.52\ \text{m} \times 6.9232 = 52.06246\ \text{m} \longleftarrow$ round off to $52.1\ \text{m}$

(d) $\dfrac{0.0239\ \text{kg}}{46.5\ \text{mL}} = 0.0005139784946\ \text{kg/mL} \longleftarrow$ round off to $0.000514\ \text{kg/mL}$ or $5.14 \times 10^{-4}\ \text{kg/mL}$

(e) First we change 2.92×10^2 cm to 0.292×10^3 cm and then carry out the addition $(5.21\ \text{cm} + 0.292\ \text{cm}) \times 10^3$. Following the procedure in (a), we find the answer is 5.50×10^3 cm.

Similar problems: 1.35, 1.36.

Practice Exercise Carry out the following arithmetic operations and round off the answers to the appropriate number of significant figures: (a) 26.5862 L + 0.17 L, (b) 9.1 g − 4.682 g, (c) 7.1×10^4 dm $\times$ 2.2654×10^2 dm, (d) 6.54 g ÷ 86.5542 mL, (e) $(7.55 \times 10^4\ \text{m}) - (8.62 \times 10^3\ \text{m})$.

The preceding rounding-off procedure applies to one-step calculations. In *chain calculations,* that is, calculations involving more than one step, we can get a different answer depending on how we round off. Consider the following two-step calculations:

$$\begin{array}{ll} \text{First step:} & A \times B = C \\ \text{Second step:} & C \times D = E \end{array}$$

Let's suppose that A = 3.66, B = 8.45, and D = 2.11. Depending on whether we round off C to three or four significant figures, we obtain a different number for E:

Method 1	Method 2
$3.66 \times 8.45 = 30.9$	$3.66 \times 8.45 = 30.93$
$30.9 \times 2.11 = 65.2$	$30.93 \times 2.11 = 65.3$

However, if we had carried out the calculation as $3.66 \times 8.45 \times 2.11$ on a calculator without rounding off the intermediate answer, we would have obtained 65.3 as the answer for E. Although retaining an additional digit past the number of significant figures for intermediate steps helps to eliminate errors from rounding, this procedure is not necessary for most calculations because the difference between the answers is usually quite small. Therefore, for most examples and end-of-chapter problems where intermediate answers are reported, all answers, intermediate and final, will be rounded.

Accuracy and Precision

In discussing measurements and significant figures, it is useful to distinguish between *accuracy* and *precision.* ***Accuracy*** tells us *how close a measurement is to the true value of the quantity that was measured.* ***Precision*** *refers to how closely two or more measurements of the same quantity agree with one another* (Figure 1.13).

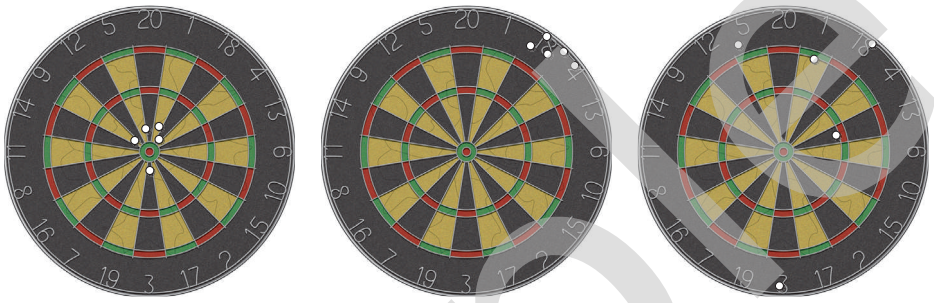


Figure 1.13 *The distribution of holes formed by darts on a dart board shows the difference between precise and accurate. (a) Good accuracy and good precision. (b) Poor accuracy and good precision. (c) Poor accuracy and poor precision.*

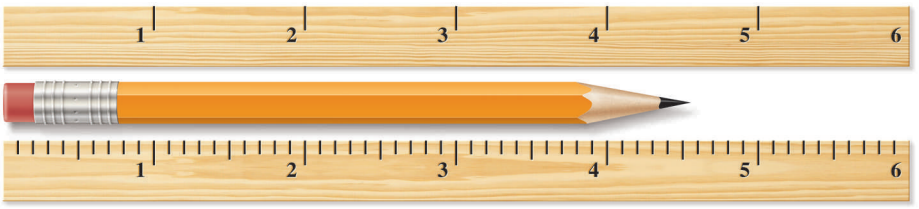
The difference between accuracy and precision is a subtle but important one. Suppose, for example, that three students are asked to determine the mass of a piece of copper wire. The results of two successive weighings by each student are

	Student A	Student B	Student C
	1.964 g	1.972 g	2.000 g
	1.978 g	1.968 g	2.002 g
Average value	1.971 g	1.970 g	2.001 g

The true mass of the wire is 2.000 g. Therefore, Student B's results are more *precise* than those of Student A (1.972 g and 1.968 g deviate less from 1.970 g than 1.964 g and 1.978 g from 1.971 g), but neither set of results is very *accurate.* Student C's results are not only the most *precise,* but also the most *accurate,* because the average value is closest to the true value. Highly accurate measurements are usually precise too. On the other hand, highly precise measurements do not necessarily guarantee accurate results. For example, an improperly calibrated meterstick or a faulty balance may give precise readings that are in error.

Review of Concepts

Give the length of the pencil with proper significant figures according to which ruler you use for the measurement.



1.9 Dimensional Analysis in Solving Problems

Careful measurements and the proper use of significant figures, along with correct calculations, will yield accurate numerical results. But to be meaningful, the answers also must be expressed in the desired units. The procedure we use to convert between units in solving chemistry problems is called *dimensional analysis* (also called the *factor-label method*). A simple technique requiring little memorization, dimensional analysis is based on the relationship between different units that express the same

physical quantity. For example, by definition 1 in = 2.54 cm (exactly). This equivalence enables us to write a conversion factor as follows:

$$\frac{1 \text{ in}}{2.54 \text{ cm}}$$

Because both the numerator and the denominator express the same length, this fraction is equal to 1. Similarly, we can write the conversion factor as

$$\frac{2.54 \text{ cm}}{1 \text{ in}}$$

which is also equal to 1. Conversion factors are useful for changing units. Thus, if we wish to convert a length expressed in inches to centimeters, we multiply the length by the appropriate conversion factor.

$$12.00 \cancel{\text{in}} \times \frac{2.54 \text{ cm}}{1 \cancel{\text{in}}} = 30.48 \text{ cm}$$

We choose the conversion factor that cancels the unit inches and produces the desired unit, centimeters. Note that the result is expressed in four significant figures because 2.54 is an exact number.

Next let us consider the conversion of 57.8 meters to centimeters. This problem can be expressed as

$$? \text{ cm} = 57.8 \text{ m}$$

By definition,

$$1 \text{ cm} = 1 \times 10^{-2} \text{ m}$$

Because we are converting "m" to "cm," we choose the conversion factor that has meters in the denominator,

$$\frac{1 \text{ cm}}{1 \times 10^{-2} \text{ m}}$$

and write the conversion as

$$\begin{aligned} ? \text{ cm} &= 57.8 \cancel{\text{m}} \times \frac{1 \text{ cm}}{1 \times 10^{-2} \cancel{\text{m}}} \\ &= 5780 \text{ cm} \\ &= 5.78 \times 10^{3} \text{ cm} \end{aligned}$$

Note that scientific notation is used to indicate that the answer has three significant figures. Again, the conversion factor 1 cm/1×10^{-2} m contains exact numbers; therefore, it does not affect the number of significant figures.

In general, to apply dimensional analysis we use the relationship

$$\text{given quantity} \times \text{conversion factor} = \text{desired quantity}$$

and the units cancel as follows:

Remember that the unit we want appears in the numerator and the unit we want to cancel appears in the denominator.

$$\cancel{\text{given unit}} \times \frac{\text{desired unit}}{\cancel{\text{given unit}}} = \text{desired unit}$$

In dimensional analysis, the units are carried through the entire sequence of calculations. Therefore, if the equation is set up correctly, then all the units will cancel except the desired one. If this is not the case, then an error must have been made somewhere, and it can usually be spotted by reviewing the solution.

A Note on Problem Solving

At this point you have been introduced to scientific notation, significant figures, and dimensional analysis, which will help you in solving numerical problems. Chemistry is an experimental science and many of the problems are quantitative in nature. The key to success in problem solving is practice. Just as a marathon runner cannot prepare for a race by simply reading books on running and a pianist cannot give a successful concert by only memorizing the musical score, you cannot be sure of your understanding of chemistry without solving problems. The following steps will help to improve your skill at solving numerical problems.

1. Read the question carefully. Understand the information that is given and what you are asked to solve. Frequently it is helpful to make a sketch that will help you to visualize the situation.
2. Find the appropriate equation that relates the given information and the unknown quantity. Sometimes solving a problem will involve more than one step, and you may be expected to look up quantities in tables that are not provided in the problem. Dimensional analysis is often needed to carry out conversions.
3. Check your answer for the correct sign, units, and significant figures.
4. A very important part of problem solving is being able to judge whether the answer is reasonable. It is relatively easy to spot a wrong sign or incorrect units. But if a number (say, 9) is incorrectly placed in the denominator instead of in the numerator, the answer would be too small even if the sign and units of the calculated quantity were correct.
5. One quick way to check the answer is to round off the numbers in the calculation in such a way so as to simplify the arithmetic. The answer you get will not be exact, but it will be close to the correct one.

Example 1.6

A person's average daily intake of glucose (a form of sugar) is 0.0833 pound (lb). What is this mass in milligrams (mg)? (1 lb = 453.6 g.)

Strategy The problem can be stated as

$$? \text{ mg} = 0.0833 \text{ lb}$$

The relationship between pounds and grams is given in the problem. This relationship will enable conversion from pounds to grams. A metric conversion is then needed to convert grams to milligrams ($1 \text{ mg} = 1 \times 10^{-3}$ g). Arrange the appropriate conversion factors so that pounds and grams cancel and the unit milligrams is obtained in your answer.

Solution The sequence of conversions is

$$\text{pounds} \longrightarrow \text{grams} \longrightarrow \text{milligrams}$$

Using the following conversion factors

$$\frac{453.6 \text{ g}}{1 \text{ lb}} \quad \text{and} \quad \frac{1 \text{ mg}}{1 \times 10^{-3} \text{ g}}$$

(Continued)

Glucose tablets can provide diabetics with a quick method for raising their blood sugar levels.

Conversion factors for some of the English system units commonly used in the United States for nonscientific measurements (for example, pounds and inches) are provided inside the back cover of this book.

we obtain the answer in one step:

$$? \text{ mg} = 0.0833 \text{ lb} \times \frac{453.6 \text{ g}}{1 \text{ lb}} \times \frac{1 \text{ mg}}{1 \times 10^{-3} \text{ g}} = 3.78 \times 10^4 \text{ mg}$$

Check As an estimate, we note that 1 lb is roughly 500 g and that 1 g = 1000 mg. Therefore, 1 lb is roughly 5×10^5 mg. Rounding off 0.0833 lb to 0.1 lb, we get 5×10^4 mg, which is close to the preceding quantity.

Similar problem: 1.45.

Practice Exercise A roll of aluminum foil has a mass of 1.07 kg. What is its mass in pounds?

A cryogenic storage tank for liquid helium.

As Examples 1.7 and 1.8 illustrate, conversion factors can be squared or cubed in dimensional analysis.

Example 1.7

A liquid helium storage tank has a volume of 275 L. What is the volume in m^3?

Strategy The problem can be stated as

$$? \text{ m}^3 = 275 \text{ L}$$

How many conversion factors are needed for this problem? Recall that 1 L = 1000 cm^3 and 1 cm = 1×10^{-2} m.

Solution We need two conversion factors here: one to convert liters to cm^3 and one to convert centimeters to meters:

$$\frac{1000 \text{ cm}^3}{1 \text{ L}} \quad \text{and} \quad \frac{1 \times 10^{-2} \text{ m}}{1 \text{ cm}}$$

Because the second conversion deals with length (cm and m) and we want volume here, it must therefore be cubed to give

$$\frac{1 \times 10^{-2} \text{ m}}{1 \text{ cm}} \times \frac{1 \times 10^{-2} \text{ m}}{1 \text{ cm}} \times \frac{1 \times 10^{-2} \text{ m}}{1 \text{ cm}} = \left(\frac{1 \times 10^{-2} \text{ m}}{1 \text{ cm}}\right)^3$$

This means that 1 cm^3 = 1×10^{-6} m^3. Now we can write

$$? \text{ m}^3 = 275 \text{ L} \times \frac{1000 \text{ cm}^3}{1 \text{ L}} \times \left(\frac{1 \times 10^{-2} \text{ m}}{1 \text{ cm}}\right)^3 = 0.275 \text{ m}^3$$

Remember that when a unit is raised to a power, any conversion factor you use must also be raised to that power.

Check From the preceding conversion factors you can show that 1 L = 1×10^{-3} m^3. Therefore, a 275-L storage tank would be equal to 275×10^{-3} m^3 or 0.275 m^3, which is the answer.

Similar problem: 1.50(d).

Practice Exercise The volume of a room is 1.08×10^8 dm^3. What is the volume in m^3?

Example 1.8

Liquid nitrogen is obtained from liquefied air and is used to prepare frozen goods and in low-temperature research. The density of the liquid at its boiling point (−196°C or 77 K) is 0.808 g/cm^3. Convert the density to units of kg/m^3.

(Continued)

Strategy The problem can be stated as

$$? \text{ kg/m}^3 = 0.808 \text{ g/cm}^3$$

Two separate conversions are required for this problem: g $\longrightarrow$ kg and $\text{cm}^3 \longrightarrow \text{m}^3$. Recall that 1 kg = 1000 g and 1 cm = 1×10^{-2} m.

Solution In Example 1.7 we saw that $1 \text{ cm}^3 = 1 \times 10^{-6} \text{ m}^3$. The conversion factors are

$$\frac{1 \text{ kg}}{1000 \text{ g}} \quad \text{and} \quad \frac{1 \text{ cm}^3}{1 \times 10^{-6} \text{ m}^3}$$

Finally,

$$? \text{ kg/m}^3 = \frac{0.808 \text{ g}}{1 \text{ cm}^3} \times \frac{1 \text{ kg}}{1000 \text{ g}} \times \frac{1 \text{ cm}^3}{1 \times 10^{-6} \text{ m}^3} = 808 \text{ kg/m}^3$$

Check Because $1 \text{ m}^3 = 1 \times 10^6 \text{ cm}^3$, we would expect much more mass in 1 m^3 than in 1 cm^3. Therefore, the answer is reasonable.

Practice Exercise The density of the lightest metal, lithium (Li), is $5.34 \times 10^2 \text{ kg/m}^3$. Convert the density to g/cm^3.

Liquid nitrogen is used for frozen foods and low-temperature research.

Similar problem: 1.51.

Review of Concepts

The Food and Drug Administration recommends no more than 65 g of daily intake of fat. What is this mass in pounds? (1 lb = 453.6 g.)

1.10 Real-World Problem Solving: Information, Assumptions, and Simplifications

In chemistry, as in other scientific disciplines, it is not always possible to solve a numerical problem exactly. There are many reasons why this is the case. For example, our understanding of a situation is not complete or data are not fully available. In these cases, we must learn to make an intelligent guess. This approach is sometimes called "ball-park estimates," which are simple, quick calculations that can be done on the "back of an envelope." As you can imagine, in many cases the answers are only order-of-magnitude estimates.[†]

In most of the example problems that you have seen so far, as well as the questions given at the end of this and subsequent chapters, the necessary information is provided; however, in order to solve important real-world problems such as those related to medicine, energy, and agriculture, you must be able to determine what information is needed and where to find it. Much of the information you might need can be found in the various tables located throughout the text, and a list of tables and important figures is given on the inside back cover. In many cases, however, you will need to go to outside sources to find the information you need. Although the Internet is a fast way to find information, you must take care that the source is reliable and well referenced. One excellent source is the National Institute of Standards and Technology (NIST).

In order to know what information you need, you will first have to formulate a plan for solving the problem. In addition to the limitations of the theories used in science, typically assumptions are made in setting up and solving the problems based on those theories. These assumptions come at a price, however, as the accuracy of the answer is reduced with increasing simplifications of the problem, as illustrated in Example 1.9.

[†]An order of magnitude is a factor of 10.

Example 1.9

A modern pencil "lead" is actually composed primarily of graphite, a form of carbon. Estimate the mass of the graphite core in a standard No. 2 pencil before it is sharpened.

Strategy Assume that the pencil lead can be approximated as a cylinder. Measurement of a typical unsharpened pencil gives a length of about 18 cm (subtracting the length of the eraser head) and a diameter of roughly 2 mm for the lead. The volume of a cylinder V is given by $V = \pi r^2 l$, where r is the radius and l is the length. Assuming that the lead is pure graphite, you can calculate the mass of the lead from the volume using the density of graphite given in Table 1.4.

Solution Converting the diameter of the lead to units of cm gives

$$2 \text{ mm} \times \frac{1 \text{ cm}}{10 \text{ mm}} = 0.2 \text{ cm}$$

which, along with the length of the lead, gives

$$V = \pi\left(\frac{0.2 \text{ cm}}{2}\right)^2 \times 18 \text{ cm} = 0.57 \text{ cm}^3$$

Rearranging Equation (1.1) gives

$$m = d \times V = 2.2 \frac{\text{g}}{\text{cm}^3} \times 0.57 \text{ cm}^3 = 1 \text{ g}$$

Check Rounding off the values used to calculate the volume of the lead gives $3 \times (0.1 \text{ cm})^2 \times 20 \text{ cm} = 0.6 \text{ cm}^3$. Multiplying that volume by roughly 2 g/cm^3 gives around 1 g, which agrees with the value just calculated.

Similar problems: 1.105, 1.106, 1.114.

Practice Exercise Estimate the mass of air in a ping pong ball.

Considering Example 1.9, even if the dimensions of the pencil lead were measured with greater precision, the accuracy of the final answer would be limited by the assumptions made in modeling this problem. The pencil lead is actually a mixture of graphite and clay, where the relative amounts of the two materials determine the softness of the lead, so the density of the material is likely to be different than 2.2 g/cm^3. You could probably find a better value for the density of the mixture used to make No. 2 pencils, but it is not worth the effort in this case.

Key Equations

$d = \frac{m}{V}$ (1.1)	Equation for density
$?°\text{C} = (°\text{F} - 32°\text{F}) \times \frac{5°\text{C}}{9°\text{F}}$ (1.2)	Converting °F to °C
$?°\text{F} = \frac{9°\text{F}}{5°\text{C}} \times (°\text{C}) + 32°\text{F}$ (1.3)	Converting °C to °F
$?\text{ K} = (°\text{C} + 273.15°\text{C}) \frac{1 \text{ K}}{1°\text{C}}$ (1.4)	Converting °C to K

Summary of Facts & Concepts

1. The study of chemistry involves three basic steps: observation, representation, and interpretation. Observation refers to measurements in the macroscopic world; representation involves the use of shorthand notation symbols and equations for communication; interpretations are based on atoms and molecules, which belong to the microscopic world.
2. The scientific method is a systematic approach to research that begins with the gathering of information through observation and measurements. In the process, hypotheses, laws, and theories are devised and tested.
3. Chemists study matter and the changes it undergoes. The substances that make up matter have unique physical properties that can be observed without changing their identity and unique chemical properties that, when they are demonstrated, do change the identity of the substances. Mixtures, whether homogeneous or heterogeneous, can be separated into pure components by physical means.
4. The simplest substances in chemistry are elements. Compounds are formed by the chemical combination of atoms of different elements in fixed proportions.
5. All substances, in principle, can exist in three states: solid, liquid, and gas. The interconversion between these states can be effected by changing the temperature.
6. SI units are used to express physical quantities in all sciences, including chemistry.
7. Numbers expressed in scientific notation have the form $N \times 10^n$, where N is between 1 and 10, and n is a positive or negative integer. Scientific notation helps us handle very large and very small quantities.

Key Words

Accuracy, p. 22
Chemical property, p. 11
Chemistry, p. 2
Compound, p. 8
Density, p. 11
Element, p. 7
Extensive property, p. 11
Heterogeneous mixture, p. 7
Homogeneous mixture, p. 7
Hypothesis, p. 4
Intensive property, p. 11
International System of Units (SI), p. 12
Kelvin, p. 15
Law, p. 4
Liter, p. 14
Macroscopic property, p. 12
Mass, p. 11
Matter, p. 6
Microscopic property, p. 12
Mixture, p. 7
Physical property, p. 10
Precision, p. 22
Qualitative, p. 4
Quantitative, p. 4
Scientific method, p. 4
Significant figures, p. 19
Substance, p. 7
Theory, p. 5
Volume, p. 11
Weight, p. 13

Questions & Problems

• *Problems available in Connect Plus*
Red numbered problems solved in Student Solutions Manual

The Scientific Method

Review Questions

1.1 Explain what is meant by the scientific method.

1.2 What is the difference between qualitative data and quantitative data?

Problems

1.3 Classify the following as qualitative or quantitative statements, giving your reasons. (a) The sun is approximately 93 million mi from Earth. (b) Leonardo da Vinci was a better painter than Michelangelo. (c) Ice is less dense than water. (d) Butter tastes better than margarine. (e) A stitch in time saves nine.

• **1.4** Classify each of the following statements as a hypothesis, a law, or a theory. (a) Beethoven's contribution to music would have been much greater if he had married. (b) An autumn leaf gravitates toward the ground because there is an attractive force between the leaf and Earth. (c) All matter is composed of very small particles called atoms.

Classification and Properties of Matter

Review Questions

1.5 Give an example for each of the following terms: (a) matter, (b) substance, (c) mixture.

• 1.6 Give an example of a homogeneous mixture and an example of a heterogeneous mixture.

1.7 Using examples, explain the difference between a physical property and a chemical property.

1.8 How does an intensive property differ from an extensive property? Which of the following properties are intensive and which are extensive? (a) length, (b) volume, (c) temperature, (d) mass.

1.9 Give an example of an element and a compound. How do elements and compounds differ?

1.10 What is the number of known elements?

Problems

1.11 Do the following statements describe chemical or physical properties? (a) Oxygen gas supports combustion. (b) Fertilizers help to increase agricultural production. (c) Water boils below 100°C on top of a mountain. (d) Lead is denser than aluminum. (e) Uranium is a radioactive element.

1.12 Does each of the following describe a physical change or a chemical change? (a) The helium gas inside a balloon tends to leak out after a few hours. (b) A flashlight beam slowly gets dimmer and finally goes out. (c) Frozen orange juice is reconstituted by adding water to it. (d) The growth of plants depends on the sun's energy in a process called photosynthesis. (e) A spoonful of table salt dissolves in a bowl of soup.

1.13 Give the names of the elements represented by the chemical symbols Li, F, P, Cu, As, Zn, Cl, Pt, Mg, U, Al, Si, Ne. (See Table 1.1 and the inside front cover.)

1.14 Give the chemical symbols for the following elements: (a) cesium, (b) germanium, (c) gallium, (d) strontium, (e) uranium, (f) selenium, (g) neon, (h) cadmium. (See Table 1.1 and the inside front cover.)

1.15 Classify each of the following substances as an element or a compound: (a) hydrogen, (b) water, (c) gold, (d) sugar.

1.16 Classify each of the following as an element, a compound, a homogeneous mixture, or a heterogeneous mixture: (a) water from a well, (b) argon gas, (c) sucrose, (d) a bottle of red wine, (e) chicken noodle soup, (f) blood flowing in a capillary, (g) ozone.

Measurement

Review Questions

1.17 Name the SI base units that are important in chemistry. Give the SI units for expressing the following: (a) length, (b) volume, (c) mass, (d) time, (e) energy, (f) temperature.

1.18 Write the numbers represented by the following prefixes: (a) mega-, (b) kilo-, (c) deci-, (d) centi-, (e) milli-, (f) micro-, (g) nano-, (h) pico-.

1.19 What units do chemists normally use for density of liquids and solids? For gas density? Explain the differences.

1.20 Describe the three temperature scales used in the laboratory and in everyday life: the Fahrenheit scale, the Celsius scale, and the Kelvin scale.

Problems

1.21 Bromine is a reddish-brown liquid. Calculate its density (in g/mL) if 586 g of the substance occupies 188 mL.

1.22 The density of methanol, a colorless organic liquid used as solvent, is 0.7918 g/mL. Calculate the mass of 89.9 mL of the liquid.

1.23 Convert the following temperatures to degrees Celsius or Fahrenheit: (a) 95°F, the temperature on a hot summer day; (b) 12°F, the temperature on a cold winter day; (c) a 102°F fever; (d) a furnace operating at 1852°F; (e) −273.15°C (theoretically the lowest attainable temperature).

1.24 (a) Normally the human body can endure a temperature of 105°F for only short periods of time without permanent damage to the brain and other vital organs. What is this temperature in degrees Celsius? (b) Ethylene glycol is a liquid organic compound that is used as an antifreeze in car radiators. It freezes at −11.5°C. Calculate its freezing temperature in degrees Fahrenheit. (c) The temperature on the surface of the sun is about 6300°C. What is this temperature in degrees Fahrenheit? (d) The ignition temperature of paper is 451°F. What is the temperature in degrees Celsius?

1.25 Convert the following temperatures to kelvin: (a) 113°C, the melting point of sulfur, (b) 37°C, the normal body temperature, (c) 357°C, the boiling point of mercury.

1.26 Convert the following temperatures to degrees Celsius: (a) 77 K, the boiling point of liquid nitrogen, (b) 4.2 K, the boiling point of liquid helium, (c) 601 K, the melting point of lead.

Handling Numbers

Review Questions

1.27 What is the advantage of using scientific notation over decimal notation?

1.28 Define significant figure. Discuss the importance of using the proper number of significant figures in measurements and calculations.

Problems

1.29 Express the following numbers in scientific notation: (a) 0.000000027, (b) 356, (c) 47,764, (d) 0.096.

1.30 Express the following numbers as decimals: (a) 1.52×10^{-2}, (b) 7.78×10^{-8}.

1.31 Express the answers to the following calculations in scientific notation:

(a) $145.75 + (2.3 \times 10^{-1})$

(b) $79{,}500 \div (2.5 \times 10^{2})$

(c) $(7.0 \times 10^{-3}) - (8.0 \times 10^{-4})$

(d) $(1.0 \times 10^{4}) \times (9.9 \times 10^{6})$

1.32 Express the answers to the following calculations in scientific notation:

(a) $0.0095 + (8.5 \times 10^{-3})$

(b) $653 \div (5.75 \times 10^{-8})$

(c) $850{,}000 - (9.0 \times 10^{5})$

(d) $(3.6 \times 10^{-4}) \times (3.6 \times 10^{6})$

1.33 What is the number of significant figures in each of the following measurements?

(a) 4867 mi

(b) 56 mL

(c) 60,104 tons

(d) 2900 g

(e) 40.2 g/cm^3

(f) 0.0000003 cm

(g) 0.7 min

(h) 4.6×10^{19} atoms

1.34 How many significant figures are there in each of the following? (a) 0.006 L, (b) 0.0605 dm, (c) 60.5 mg, (d) 605.5 cm^2, (e) 960×10^{-3} g, (f) 6 kg, (g) 60 m.

1.35 Carry out the following operations as if they were calculations of experimental results, and express each answer in the correct units with the correct number of significant figures:

(a) 5.6792 m + 0.6 m + 4.33 m

(b) 3.70 g − 2.9133 g

(c) 4.51 cm × 3.6666 cm

(d) $(3 \times 10^{4}$ g + 6.827 g)/(0.043 cm^3 − 0.021 cm^3)

1.36 Carry out the following operations as if they were calculations of experimental results, and express each answer in the correct units with the correct number of significant figures:

(a) 7.310 km ÷ 5.70 km

(b) $(3.26 \times 10^{-3}$ mg) − $(7.88 \times 10^{-5}$ mg)

(c) $(4.02 \times 10^{6}$ dm) + $(7.74 \times 10^{7}$ dm)

(d) (7.8 m − 0.34 m)/(1.15 s + 0.82 s)

1.37 Three students (A, B, and C) are asked to determine the volume of a sample of ethanol. Each student measures the volume three times with a graduated cylinder. The results in milliliters are: A (87.1, 88.2, 87.6); B (86.9, 87.1, 87.2); C (87.6, 87.8, 87.9). The true volume is 87.0 mL. Comment on the precision and the accuracy of each student's results.

1.38 Three apprentice tailors (X, Y, and Z) are assigned the task of measuring the seam of a pair of trousers. Each one makes three measurements. The results in inches are X (31.5, 31.6, 31.4); Y (32.8, 32.3, 32.7); Z (31.9, 32.2, 32.1). The true length is 32.0 in. Comment on the precision and the accuracy of each tailor's measurements.

Dimensional Analysis

Problems

1.39 Carry out the following conversions: (a) 22.6 m to decimeters, (b) 25.4 mg to kilograms, (c) 556 mL to liters, (d) 10.6 kg/m^3 to g/cm^3.

1.40 Carry out the following conversions: (a) 242 lb to milligrams, (b) 68.3 cm^3 to cubic meters, (c) 7.2 m^3 to liters, (d) 28.3 μg to pounds.

1.41 The average speed of helium at 25°C is 1255 m/s. Convert this speed to miles per hour (mph).

1.42 How many seconds are there in a solar year (365.24 days)?

1.43 How many minutes does it take light from the sun to reach Earth? (The distance from the sun to Earth is 93 million mi; the speed of light = 3.00×10^{8} m/s.)

1.44 A jogger runs a mile in 8.92 min. Calculate the speed in (a) in/s, (b) m/min, (c) km/h. (1 mi = 1609 m; 1 in = 2.54 cm.)

1.45 A 6.0-ft person weighs 168 lb. Express this person's height in meters and weight in kilograms. (1 lb = 453.6 g; 1 m = 3.28 ft.)

1.46 The speed limit on parts of the German autobahn was once set at 286 kilometers per hour (km/h). Calculate the speed limit in miles per hour (mph).

1.47 For a fighter jet to take off from the deck of an aircraft carrier, it must reach a speed of 62 m/s. Calculate the speed in miles per hour (mph).

1.48 The "normal" lead content in human blood is about 0.40 part per million (that is, 0.40 g of lead per million grams of blood). A value of 0.80 part per million (ppm) is considered to be dangerous. How many grams of lead are contained in 6.0×10^{3} g of blood (the amount in an average adult) if the lead content is 0.62 ppm?

1.49 Carry out the following conversions: (a) 1.42 light-years to miles (a light-year is an astronomical measure of distance—the distance traveled by light in a year, or 365 days; the speed of light is 3.00×10^{8} m/s). (b) 32.4 yd to centimeters. (c) 3.0×10^{10} cm/s to ft/s.

1.50 Carry out the following conversions: (a) 70 kg, the average weight of a male adult, to pounds. (b) 14 billion years (roughly the age of the universe) to seconds. (Assume there are 365 days in a year.) (c) 7 ft 6 in, the height of the basketball player Yao Ming, to meters. (d) 88.6 m^3 to liters.

1.51 Aluminum is a lightweight metal (density = 2.70 g/cm^3) used in aircraft construction, high-voltage transmission lines, beverage cans, and foils. What is its density in kg/m^3?

1.52 Ammonia gas is used as a refrigerant in large-scale cooling systems. The density of ammonia gas under certain conditions is 0.625 g/L. Calculate its density in g/cm^3.

Additional Problems

1.53 Give one qualitative and one quantitative statement about each of the following: (a) water, (b) carbon, (c) iron, (d) hydrogen gas, (e) sucrose (cane sugar), (f) table salt (sodium chloride), (g) mercury, (h) gold, (i) air.

1.54 Which of the following statements describe physical properties and which describe chemical properties? (a) Iron has a tendency to rust. (b) Rainwater in industrialized regions tends to be acidic. (c) Hemoglobin molecules have a red color. (d) When a glass of water is left out in the sun, the water gradually disappears. (e) Carbon dioxide in air is converted to more complex molecules by plants during photosynthesis.

1.55 In 2008, about 95.0 billion lb of sulfuric acid were produced in the United States. Convert this quantity to tons.

1.56 In determining the density of a rectangular metal bar, a student made the following measurements: length, 8.53 cm; width, 2.4 cm; height, 1.0 cm; mass, 52.7064 g. Calculate the density of the metal to the correct number of significant figures.

1.57 Calculate the mass of each of the following: (a) a sphere of gold with a radius of 10.0 cm [the volume of a sphere with a radius r is $V = (4/3)\pi r^3$; the density of gold = 19.3 g/cm^3], (b) a cube of platinum of edge length 0.040 mm (the density of platinum = 21.4 g/cm^3), (c) 50.0 mL of ethanol (the density of ethanol = 0.798 g/mL).

1.58 A cylindrical glass bottle 21.5 cm in length is filled with cooking oil of density 0.953 g/mL. If the mass of the oil needed to fill the bottle is 1360 g, calculate the inner diameter of the bottle.

1.59 The following procedure was used to determine the volume of a flask. The flask was weighed dry and then filled with water. If the masses of the empty flask and filled flask were 56.12 g and 87.39 g, respectively, and the density of water is 0.9976 g/cm^3, calculate the volume of the flask in cm^3.

1.60 The speed of sound in air at room temperature is about 343 m/s. Calculate this speed in miles per hour. (1 mi = 1609 m.)

1.61 A piece of silver (Ag) metal weighing 194.3 g is placed in a graduated cylinder containing 242.0 mL of water. The volume of water now reads 260.5 mL. From these data calculate the density of silver.

1.62 The experiment described in Problem 1.61 is a crude but convenient way to determine the density of some solids. Describe a similar experiment that would enable you to measure the density of ice. Specifically, what would be the requirements for the liquid used in your experiment?

1.63 A lead sphere of diameter 48.6 cm has a mass of 6.852×10^5 g. Calculate the density of lead.

1.64 Lithium is the least dense metal known (density: 0.53 g/cm^3). What is the volume occupied by 1.20×10^3 g of lithium?

1.65 The medicinal thermometer commonly used in homes can be read ±0.1°F, whereas those in the doctor's office may be accurate to ±0.1°C. In degrees Celsius, express the percent error expected from each of these thermometers in measuring a person's body temperature of 38.9°C.

1.66 Vanillin (used to flavor vanilla ice cream and other foods) is the substance whose aroma the human nose detects in the smallest amount. The threshold limit is 2.0×10^{-11} g per liter of air. If the current price of 50 g of vanillin is $112, determine the cost to supply enough vanillin so that the aroma could be detected in a large aircraft hangar with a volume of 5.0×10^7 ft^3.

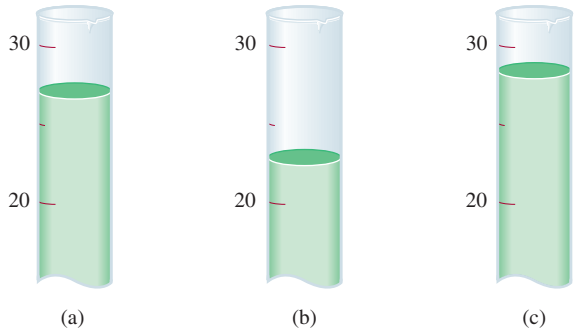
1.67 At what temperature does the numerical reading on a Celsius thermometer equal that on a Fahrenheit thermometer?

1.68 Suppose that a new temperature scale has been devised on which the melting point of ethanol (−117.3°C) and the boiling point of ethanol (78.3°C) are taken as 0°S and 100°S, respectively, where S is the symbol for the new temperature scale. Derive an equation relating a reading on this scale to a reading on the Celsius scale. What would this thermometer read at 25°C?

1.69 A resting adult requires about 240 mL of pure oxygen/min and breathes about 12 times every minute. If inhaled air contains 20 percent oxygen by volume and exhaled air 16 percent, what is the volume of air per breath? (Assume that the volume of inhaled air is equal to that of exhaled air.)

1.70 (a) Referring to Problem 1.69, calculate the total volume (in liters) of air an adult breathes in a day. (b) In a city with heavy traffic, the air contains 2.1×10^{-6} L of carbon monoxide (a poisonous gas) per liter. Calculate the average daily intake of carbon monoxide in liters by a person.

1.71 Three different 25.0-g samples of solid pellets are added to 20.0 mL of water in three different measuring cylinders. The results are shown here. Given the densities of the three metals used, identify the cylinder that contains each sample of solid pellets: A (2.9 g/cm^3), B (8.3 g/cm^3), and C (3.3 g/cm^3).



1.72 The circumference of an NBA-approved basketball is 29.6 in. Given that the radius of Earth is about 6400 km, how many basketballs would it take to circle around the equator with the basketballs touching one another? Round off your answer to an integer with three significant figures.

1.73 A student is given a crucible and asked to prove whether it is made of pure platinum. She first weighs the crucible in air and then weighs it suspended in water (density = 0.9986 g/mL). The readings are 860.2 g and 820.2 g, respectively. Based on these measurements and given that the density of platinum is 21.45 g/cm^3, what should her conclusion be? (*Hint:* An object suspended in a fluid is buoyed up by the mass of the fluid displaced by the object. Neglect the buoyance of air.)

1.74 The surface area and average depth of the Pacific Ocean are 1.8×10^8 km^2 and 3.9×10^3 m, respectively. Calculate the volume of water in the ocean in liters.

1.75 The unit "troy ounce" is often used for precious metals such as gold (Au) and platinum (Pt). (1 troy ounce = 31.103 g.) (a) A gold coin weighs 2.41 troy ounces. Calculate its mass in grams. (b) Is a troy ounce heavier or lighter than an ounce? (1 lb = 16 oz; 1 lb = 453.6 g.)

1.76 Osmium (Os) is the densest element known (density = 22.57 g/cm^3). Calculate the mass in pounds and in kilograms of an Os sphere 15 cm in diameter (about the size of a grapefruit). See Problem 1.57 for volume of a sphere.

1.77 Percent error is often expressed as the absolute value of the difference between the true value and the experimental value, divided by the true value:

$$\text{percent error} = \frac{|\text{true value} - \text{experimental value}|}{|\text{true value}|} \times 100\%$$

The vertical lines indicate absolute value. Calculate the percent error for the following measurements: (a) The density of alcohol (ethanol) is found to be 0.802 g/mL. (True value: 0.798 g/mL.) (b) The mass of gold in an earring is analyzed to be 0.837 g. (True value: 0.864 g.)

1.78 The natural abundances of elements in the human body, expressed as percent by mass, are: oxygen (O), 65 percent; carbon (C), 18 percent; hydrogen (H), 10 percent; nitrogen (N), 3 percent; calcium (Ca), 1.6 percent; phosphorus (P), 1.2 percent; all other elements, 1.2 percent. Calculate the mass in grams of each element in the body of a 62-kg person.

1.79 The men's world record for running a mile outdoors (as of 1999) is 3 min 43.13 s. At this rate, how long would it take to run a 1500-m race? (1 mi = 1609 m.)

1.80 Venus, the second closest planet to the sun, has a surface temperature of 7.3×10^2 K. Convert this temperature to °C and °F.

1.81 Chalcopyrite, the principal ore of copper (Cu), contains 34.63 percent Cu by mass. How many grams of Cu can be obtained from 5.11×10^3 kg of the ore?

1.82 It has been estimated that 8.0×10^4 tons of gold (Au) have been mined. Assume gold costs $948 per ounce. What is the total worth of this quantity of gold?

1.83 A 1.0-mL volume of seawater contains about 4.0×10^{-12} g of gold. The total volume of ocean water is 1.5×10^{21} L. Calculate the total amount of gold (in grams) that is present in seawater, and the worth of the gold in dollars (see Problem 1.82). With so much gold out there, why hasn't someone become rich by mining gold from the ocean?

1.84 Measurements show that 1.0 g of iron (Fe) contains 1.1×10^{22} Fe atoms. How many Fe atoms are in 4.9 g of Fe, which is the total amount of iron in the body of an average adult?

1.85 The thin outer layer of Earth, called the crust, contains only 0.50 percent of Earth's total mass and yet is the source of almost all the elements (the atmosphere provides elements such as oxygen, nitrogen, and a few other gases). Silicon (Si) is the second most abundant element in Earth's crust (27.2 percent by mass). Calculate the mass of silicon in kilograms in Earth's crust. (The mass of Earth is 5.9×10^{21} tons. 1 ton = 2000 lb; 1 lb = 453.6 g.)

1.86 The radius of a copper (Cu) atom is roughly 1.3×10^{-10} m. How many times can you divide evenly a piece of 10-cm copper wire until it is reduced to two separate copper atoms? (Assume there are appropriate tools for this procedure and that copper atoms are lined up in a straight line, in contact with each other. Round off your answer to an integer.)

1.87 One gallon of gasoline in an automobile's engine produces on the average 9.5 kg of carbon dioxide, which is a greenhouse gas, that is, it promotes the warming of Earth's atmosphere. Calculate the annual production of carbon dioxide in kilograms if there are 250 million cars in the United States and each car covers a distance of 5000 mi at a consumption rate of 20 miles per gallon.

1.88 A sheet of aluminum (Al) foil has a total area of 1.000 ft^2 and a mass of 3.636 g. What is the thickness of the foil in millimeters? (Density of Al = 2.699 g/cm^3.)

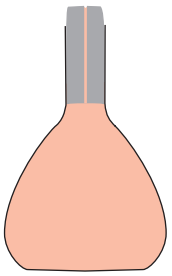
1.89 Comment on whether each of the following is a homogeneous mixture or a heterogeneous mixture: (a) air in a closed bottle and (b) air over New York City.

1.90 Chlorine is used to disinfect swimming pools. The accepted concentration for this purpose is 1 ppm chlorine, or 1 g of chlorine per million grams of water. Calculate the volume of a chlorine solution (in milliliters) a homeowner should add to her

swimming pool if the solution contains 6.0 percent chlorine by mass and there are 2.0×10^4 gallons of water in the pool. (1 gallon = 3.79 L; density of liquids = 1.0 g/mL.)

1.91 An aluminum cylinder is 10.0 cm in length and has a radius of 0.25 cm. If the mass of a single Al atom is 4.48×10^{-23}g, calculate the number of Al atoms present in the cylinder. The density of aluminum is 2.70 g/cm^3.

1.92 A pycnometer is a device for measuring the density of liquids. It is a glass flask with a close-fitting ground glass stopper having a capillary hole through it. (a) The volume of the pycnometer is determined by using distilled water at 20°C with a known density of 0.99820 g/mL. First, the water is filled to the rim. With the stopper in place, the fine hole allows the excess liquid to escape. The pycnometer is then carefully dried with filter paper. Given that the masses of the empty pycnometer and the same one filled with water are 32.0764 g and 43.1195 g, respectively, calculate the volume of the pycnometer. (b) If the mass of the pycnometer filled with ethanol at 20°C is 40.8051 g, calculate the density of ethanol. (c) Pycnometers can also be used to measure the density of solids. First, small zinc granules weighing 22.8476 g are placed in the pycnometer, which is then filled with water. If the combined mass of the pycnometer plus the zinc granules and water is 62.7728 g, what is the density of zinc?



1.93 In 1849 a gold prospector in California collected a bag of gold nuggets plus sand. Given that the density of gold and sand are 19.3 g/cm^3 and 2.95 g/cm^3, respectively, and that the density of the mixture is 4.17 g/cm^3, calculate the percent by mass of gold in the mixture.

1.94 The average time it takes for a molecule to diffuse a distance of x cm is given by

$$t = \frac{x^2}{2D}$$

where t is the time in seconds and D is the diffusion coefficient. Given that the diffusion coefficient of glucose is 5.7×10^{-7} cm^2/s, calculate the time it would take for a glucose molecule to diffuse 10 μm, which is roughly the size of a cell.

1.95 A human brain weighs about 1 kg and contains about 10^{11} cells. Assuming that each cell is completely filled with water (density = 1 g/mL), calculate the length of one side of such a cell if it were a cube. If the cells are spread out in a thin layer that is a single cell thick, what is the surface area in square meters?

1.96 (a) Carbon monoxide (CO) is a poisonous gas because it binds very strongly to the oxygen carrier hemoglobin in blood. A concentration of 8.00×10^2 ppm by volume of carbon monoxide is considered lethal to humans. Calculate the volume in liters occupied by carbon monoxide in a room that measures 17.6 m long, 8.80 m wide, and 2.64 m high at this concentration. (b) Prolonged exposure to mercury (Hg) vapor can cause neurological disorders and respiratory problems. For safe air quality control, the concentration of mercury vapor must be under 0.050 mg/m^3. Convert this number to g/L. (c) The general test for type II diabetes is that the blood sugar (glucose) level should be below 120 mg per deciliter (mg/dL). Convert this number to micrograms per milliliter (μg/mL).

1.97 A bank teller is asked to assemble "one-dollar" sets of coins for his clients. Each set is made of three quarters, one nickel, and two dimes. The masses of the coins are: quarter: 5.645 g; nickel: 4.967 g; dime: 2.316 g. What is the maximum number of sets that can be assembled from 33.871 kg of quarters, 10.432 kg of nickels, and 7.990 kg of dimes? What is the total mass (in g) of the assembled sets of coins?

1.98 A graduated cylinder is filled to the 40.00-mL mark with a mineral oil. The masses of the cylinder before and after the addition of the mineral oil are 124.966 g and 159.446 g, respectively. In a separate experiment, a metal ball bearing of mass 18.713 g is placed in the cylinder and the cylinder is again filled to the 40.00-mL mark with the mineral oil. The combined mass of the ball bearing and mineral oil is 50.952 g. Calculate the density and radius of the ball bearing. [The volume of a sphere of radius r is $(4/3)\pi r^3$.]

1.99 A chemist in the nineteenth century prepared an unknown substance. In general, do you think it would be more difficult to prove that it is an element or a compound? Explain.

1.100 Bronze is an alloy made of copper (Cu) and tin (Sn) used in applications that require low metal-on-metal friction. Calculate the mass of a bronze cylinder of radius 6.44 cm and length 44.37 cm. The composition of the bronze is 79.42 percent Cu and 20.58 percent Sn and the densities of Cu and Sn are 8.94 g/cm^3 and 7.31 g/cm^3, respectively. What assumption should you make in this calculation?

1.101 You are given a liquid. Briefly describe steps you would take to show whether it is a pure substance or a homogeneous mixture.

1.102 A chemist mixes two liquids A and B to form a homogeneous mixture. The densities of the liquids are 2.0514 g/mL for A and 2.6678 g/mL for B. When she drops a small object into the mixture, she finds that the object becomes suspended in the liquid; that is, it neither sinks nor floats. If the mixture is made of 41.37 percent A and 58.63 percent B by volume, what is the density of the metal? Can this procedure be used in general to determine the densities of solids? What assumptions must be made in applying this method?

1.103 Tums is a popular remedy for acid indigestion. A typical Tums tablet contains calcium carbonate plus some inert substances. When ingested, it reacts with the gastric juice (hydrochloric acid) in the stomach to give off carbon dioxide gas. When a 1.328-g tablet reacted with 40.00 mL of hydrochloric acid (density: 1.140 g/mL), carbon dioxide gas was given off and the resulting solution weighed 46.699 g. Calculate the number of liters of carbon dioxide gas released if its density is 1.81 g/L.

1.104 A 250-mL glass bottle was filled with 242 mL of water at 20°C and tightly capped. It was then left outdoors overnight, where the average temperature was −5°C. Predict what would happen. The density of water at 20°C is 0.998 g/cm^3 and that of ice at −5°C is 0.916 g/cm^3.

Interpreting, Modeling & Estimating

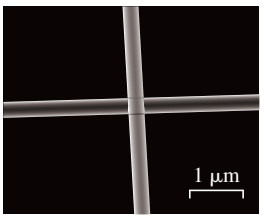
1.105 What is the mass of one mole of ants? (*Useful information:* A mole is the unit used for atomic and subatomic particles. It is approximately 6×10^{23}. A 1-cm-long ant weighs about 3 mg.)

1.106 How much time (in years) does an 80-year-old person spend sleeping during his or her life span?

1.107 Estimate the daily amount of water (in gallons) used indoors by a family of four in the United States.

1.108 Public bowling alleys generally stock bowling balls from 8 to 16 lb, where the mass is given in whole numbers. Given that regulation bowling balls have a diameter of 8.6 in, which (if any) of these bowling balls would you expect to float in water?

1.109 Fusing "nanofibers" with diameters of 100–300 nm gives junctures with very small volumes that would potentially allow the study of reactions involving only a few molecules. Estimate the volume in liters of the junction formed between two such fibers with internal diameters of 200 nm. The scale reads 1 μm.



1.110 Estimate the annual consumption of gasoline by passenger cars in the United States.

1.111 Estimate the total amount of ocean water in liters.

1.112 Estimate the volume of blood in an adult in liters.

1.113 How far (in feet) does light travel in one nanosecond?

1.114 Estimate the distance (in miles) covered by an NBA player in a professional basketball game.

1.115 In water conservation, chemists spread a thin film of a certain inert material over the surface of water to cut down on the rate of evaporation of water in reservoirs. This technique was pioneered by Benjamin Franklin three centuries ago. Franklin found that 0.10 mL of oil could spread over the surface of water about 40 m^2 in area. Assuming that the oil forms a *monolayer,* that is, a layer that is only one molecule thick, estimate the length of each oil molecule in nanometers. (1 nm = 1×10^{-9} m.)

Answers to Practice Exercises

1.1 96.5 g. **1.2** 341 g. **1.3** (a) 621.5°F, (b) 78.3°C, (c) −196°C. **1.4** (a) Two, (b) four, (c) three, (d) two, (e) three or two. **1.5** (a) 26.76 L, (b) 4.4 g, (c) 1.6×10^7 dm^2, (d) 0.0756 g/mL, (e) 6.69×10^4 m. **1.6** 2.36 lb. **1.7** 1.08×10^5 m^3. **1.8** 0.534 g/cm^3. **1.9** Roughly 0.03 g.

CHEMICAL *MYSTERY*

The Disappearance of the Dinosaurs

Dinosaurs dominated life on Earth for millions of years and then disappeared very suddenly. To solve the mystery, paleontologists studied fossils and skeletons found in rocks in various layers of Earth's crust. Their findings enabled them to map out which species existed on Earth during specific geologic periods. They also revealed no dinosaur skeletons in rocks formed immediately after the Cretaceous period, which dates back some

65 million years. It is therefore assumed that the dinosaurs became extinct about 65 million years ago.

Among the many hypotheses put forward to account for their disappearance were disruptions of the food chain and a dramatic change in climate caused by violent volcanic eruptions. However, there was no convincing evidence for any one hypothesis until 1977. It was then that a group of paleontologists working in Italy obtained some very puzzling data at a site near Gubbio. The chemical analysis of a layer of clay deposited above sediments formed during the Cretaceous period (and therefore a layer that records events occurring *after* the Cretaceous period) showed a surprisingly high content of the element iridium (Ir). Iridium is very rare in Earth's crust but is comparatively abundant in asteroids.

This investigation led to the hypothesis that the extinction of dinosaurs occurred as follows. To account for the quantity of iridium found, scientists suggested that a large asteroid several miles in diameter hit Earth about the time the dinosaurs disappeared. The impact of the asteroid on Earth's surface must have been so tremendous that it literally vaporized a large quantity of surrounding rocks, soils, and other objects. The resulting dust and debris floated through the air and blocked the sunlight for months or perhaps years. Without ample sunlight most plants could not grow, and the fossil record confirms that many types of plants did indeed die out at this time. Consequently, of course, many plant-eating animals perished, and then, in turn, meat-eating animals began to starve. Dwindling food sources would obviously affect large animals needing great amounts of food more quickly and more severely than small animals. Therefore, the huge dinosaurs, the largest of which might have weighed as much as 30 tons, vanished due to lack of food.

Chemical Clues

1. How does the study of dinosaur extinction illustrate the scientific method?
2. Suggest two ways that would enable you to test the asteroid collision hypothesis.
3. In your opinion, is it justifiable to refer to the asteroid explanation as the theory of dinosaur extinction?
4. Available evidence suggests that about 20 percent of the asteroid's mass turned to dust and spread uniformly over Earth after settling out of the upper atmosphere. This dust amounted to about 0.02 g/cm^2 of Earth's surface. The asteroid very likely had a density of about 2 g/cm^3. Calculate the mass (in kilograms and tons) of the asteroid and its radius in meters, assuming that it was a sphere. (The area of Earth is 5.1×10^{14} m^2; 1 lb = 453.6 g.) (Source: *Consider a Spherical Cow—A Course in Environmental Problem Solving* by J. Harte, University Science Books, Mill Valley, CA 1988. Used with permission.)

CHAPTER

2

Atoms, Molecules, and Ions

Illustration depicting Marie and Pierre Curie at work in their laboratory. The Curies studied and identified many radioactive elements.

CHAPTER OUTLINE

2.1 The Atomic Theory
2.2 The Structure of the Atom
2.3 Atomic Number, Mass Number, and Isotopes
2.4 The Periodic Table
2.5 Molecules and Ions
2.6 Chemical Formulas
2.7 Naming Compounds
2.8 Introduction to Organic Compounds

A LOOK AHEAD

- We begin with a historical perspective of the search for the fundamental units of matter. The modern version of atomic theory was laid by John Dalton in the nineteenth century, who postulated that elements are composed of extremely small particles, called atoms. All atoms of a given element are identical, but they are different from atoms of all other elements. (2.1)
- We note that, through experimentation, scientists have learned that an atom is composed of three elementary particles: proton, electron, and neutron. The proton has a positive charge, the electron has a negative charge, and the neutron has no charge. Protons and neutrons are located in a small region at the center of the atom, called the nucleus, while electrons are spread out about the nucleus at some distance from it. (2.2)
- We will learn the following ways to identify atoms. Atomic number is the number of protons in a nucleus; atoms of different elements have different atomic numbers. Isotopes are atoms of the same element having a different number of neutrons. Mass number is the sum of the number of protons and neutrons in an atom. Because an atom is electrically neutral, the number of protons is equal to the number of electrons in it. (2.3)
- Next we will see how elements can be grouped together according to their chemical and physical properties in a chart called the periodic table. The periodic table enables us to classify elements (as metals, metalloids, and nonmetals) and correlate their properties in a systematic way. (2.4)
- We will see that atoms of most elements interact to form compounds, which are classified as molecules or ionic compounds made of positive (cations) and negative (anions) ions. (2.5)
- We learn to use chemical formulas (molecular and empirical) to represent molecules and ionic compounds and models to represent molecules. (2.6)
- We learn a set of rules that help us name the inorganic compounds. (2.7)
- Finally, we will briefly explore the organic world to which we will return in a later chapter. (2.8)

Since ancient times humans have pondered the nature of matter. Our modern ideas of the structure of matter began to take shape in the early nineteenth century with Dalton's atomic theory. We now know that all matter is made of atoms, molecules, and ions. All of chemistry is concerned in one way or another with these species.

2.1 The Atomic Theory

In the fifth century B.C. the Greek philosopher Democritus expressed the belief that all matter consists of very small, indivisible particles, which he named *atomos* (meaning uncuttable or indivisible). Although Democritus' idea was not accepted by many of his contemporaries (notably Plato and Aristotle), somehow it endured. Experimental evidence from early scientific investigations provided support for the notion of "atomism" and gradually gave rise to the modern definitions of elements and compounds. In 1808 an English scientist and school teacher, John Dalton,[†] formulated a precise definition of the indivisible building blocks of matter that we call atoms.

Dalton's work marked the beginning of the modern era of chemistry. The hypotheses about the nature of matter on which Dalton's atomic theory is based can be summarized as follows:

1. Elements are composed of extremely small particles called atoms.
2. All atoms of a given element are identical, having the same size, mass, and chemical properties. The atoms of one element are different from the atoms of all other elements.
3. Compounds are composed of atoms of more than one element. In any compound, the ratio of the numbers of atoms of any two of the elements present is either an integer or a simple fraction.
4. A chemical reaction involves only the separation, combination, or rearrangement of atoms; it does not result in their creation or destruction.

Figure 2.1 is a schematic representation of the last three hypotheses.

Dalton's concept of an atom was far more detailed and specific than Democritus'. The second hypothesis states that atoms of one element are different from atoms of all other elements. Dalton made no attempt to describe the structure or composition of atoms—he had no idea what an atom is really like. But he did realize that the

[†]John Dalton (1766–1844). English chemist, mathematician, and philosopher. In addition to the atomic theory, he also formulated several gas laws and gave the first detailed description of color blindness, from which he suffered. Dalton was described as an indifferent experimenter, and singularly wanting in the language and power of illustration. His only recreation was lawn bowling on Thursday afternoons. Perhaps it was the sight of those wooden balls that provided him with the idea of the atomic theory.

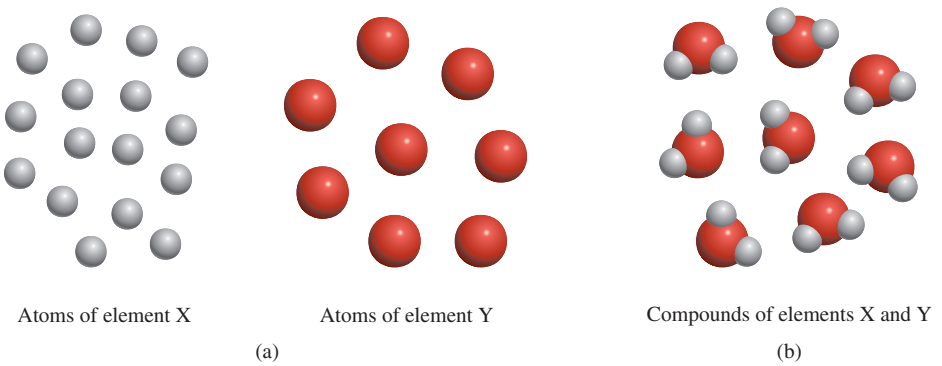


Figure 2.1 *(a) According to Dalton's atomic theory, atoms of the same element are identical, but atoms of one element are different from atoms of other elements. (b) Compound formed from atoms of elements X and Y. In this case, the ratio of the atoms of element X to the atoms of element Y is 2:1. Note that a chemical reaction results only in the rearrangement of atoms, not in their destruction or creation.*

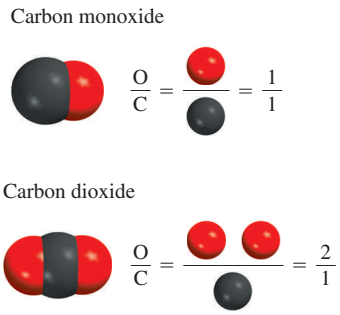


Figure 2.2 *An illustration of the law of multiple proportions.*

different properties shown by elements such as hydrogen and oxygen can be explained by assuming that hydrogen atoms are not the same as oxygen atoms.

The third hypothesis suggests that, to form a certain compound, we need not only atoms of the right kinds of elements, but specific numbers of these atoms as well. This idea is an extension of a law published in 1799 by Joseph Proust,[†] a French chemist. Proust's ***law of definite proportions*** states that *different samples of the same compound always contain its constituent elements in the same proportion by mass.* Thus, if we were to analyze samples of carbon dioxide gas obtained from different sources, we would find in each sample the same ratio by mass of carbon to oxygen. It stands to reason, then, that if the ratio of the masses of different elements in a given compound is fixed, the ratio of the atoms of these elements in the compound also must be constant.

Dalton's third hypothesis supports another important law, the ***law of multiple proportions***. According to the law, *if two elements can combine to form more than one compound, the masses of one element that combine with a fixed mass of the other element are in ratios of small whole numbers.* Dalton's theory explains the law of multiple proportions quite simply: Different compounds made up of the same elements differ in the number of atoms of each kind that combine. For example, carbon forms two stable compounds with oxygen, namely, carbon monoxide and carbon dioxide. Modern measurement techniques indicate that one atom of carbon combines with one atom of oxygen in carbon monoxide and with two atoms of oxygen in carbon dioxide. Thus, the ratio of oxygen in carbon monoxide to oxygen in carbon dioxide is 1:2. This result is consistent with the law of multiple proportions (Figure 2.2).

Dalton's fourth hypothesis is another way of stating the ***law of conservation of mass,***[‡] which is that *matter can be neither created nor destroyed.* Because matter is made of atoms that are unchanged in a chemical reaction, it follows that mass must be conserved as well. Dalton's brilliant insight into the nature of matter was the main stimulus for the rapid progress of chemistry during the nineteenth century.

Review of Concepts

The atoms of elements A (blue) and B (orange) form two compounds shown here. Do these compounds obey the law of multiple proportions?

2.2 The Structure of the Atom

On the basis of Dalton's atomic theory, we can define an ***atom*** as *the basic unit of an element that can enter into chemical combination.* Dalton imagined an atom that was both extremely small and indivisible. However, a series of investigations that began in the 1850s and extended into the twentieth century clearly demonstrated that atoms actually possess internal structure; that is, they are made up of even smaller particles, which are called *subatomic particles.* This research led to the discovery of three such particles—electrons, protons, and neutrons.

[†]Joseph Louis Proust (1754–1826). French chemist. Proust was the first person to isolate sugar from grapes.

[‡]According to Albert Einstein, mass and energy are alternate aspects of a single entity called *mass-energy.* Chemical reactions usually involve a gain or loss of heat and other forms of energy. Thus, when energy is lost in a reaction, for example, mass is also lost. Except for nuclear reactions (see Chapter 19), however, changes of mass in chemical reactions are too small to detect. Therefore, for all practical purposes mass is conserved.

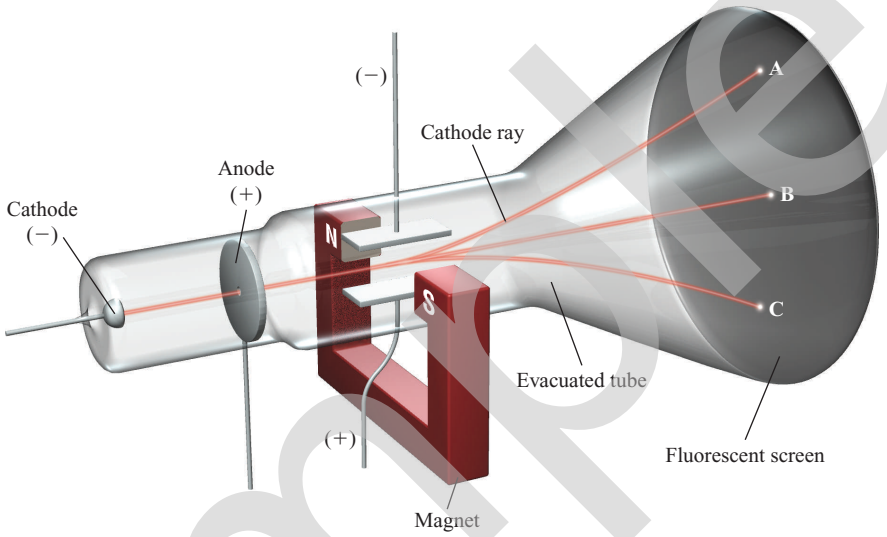


Figure 2.3 *A cathode ray tube with an electric field perpendicular to the direction of the cathode rays and an external magnetic field. The symbols N and S denote the north and south poles of the magnet. The cathode rays will strike the end of the tube at A in the presence of a magnetic field, at C in the presence of an electric field, and at B when there are no external fields present or when the effects of the electric field and magnetic field cancel each other.*

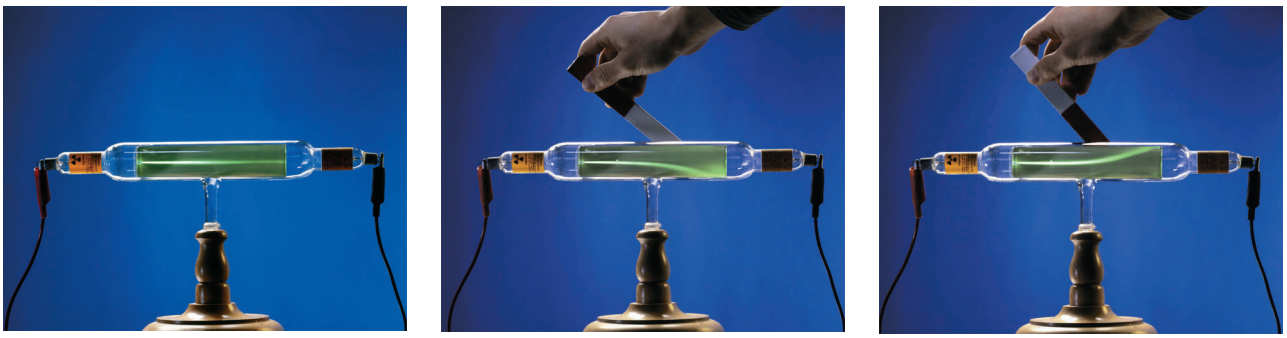
The Electron

In the 1890s, many scientists became caught up in the study of ***radiation,*** *the emission and transmission of energy through space in the form of waves.* Information gained from this research contributed greatly to our understanding of atomic structure. One device used to investigate this phenomenon was a cathode ray tube, the forerunner of the television tube (Figure 2.3). It is a glass tube from which most of the air has been evacuated. When the two metal plates are connected to a high-voltage source, the negatively charged plate, called the *cathode,* emits an invisible ray. The cathode ray is drawn to the positively charged plate, called the *anode,* where it passes through a hole and continues traveling to the other end of the tube. When the ray strikes the specially coated surface, it produces a strong fluorescence, or bright light.

Animation
Cathode Ray Tube

In some experiments, two electrically charged plates and a magnet were added to the *outside* of the cathode ray tube (see Figure 2.3). When the magnetic field is on and the electric field is off, the cathode ray strikes point A. When only the electric field is on, the ray strikes point C. When both the magnetic and the electric fields are off or when they are both on but balanced so that they cancel each other's influence, the ray strikes point B. According to electromagnetic theory, a moving charged body behaves like a magnet and can interact with electric and magnetic fields through which it passes. Because the cathode ray is attracted by the plate bearing positive charges and repelled by the plate bearing negative charges, it must consist of negatively charged particles. We know these *negatively charged particles* as ***electrons.*** Figure 2.4 shows the effect of a bar magnet on the cathode ray.

Electrons are normally associated with atoms. However, they can also be studied individually.



(a) (b) (c)

Figure 2.4 *(a) A cathode ray produced in a discharge tube traveling from the cathode (left) to the anode (right). The ray itself is invisible, but the fluorescence of a zinc sulfide coating on the glass causes it to appear green. (b) The cathode ray is bent downward when a bar magnet is brought toward it. (c) When the polarity of the magnet is reversed, the ray bends in the opposite direction.*

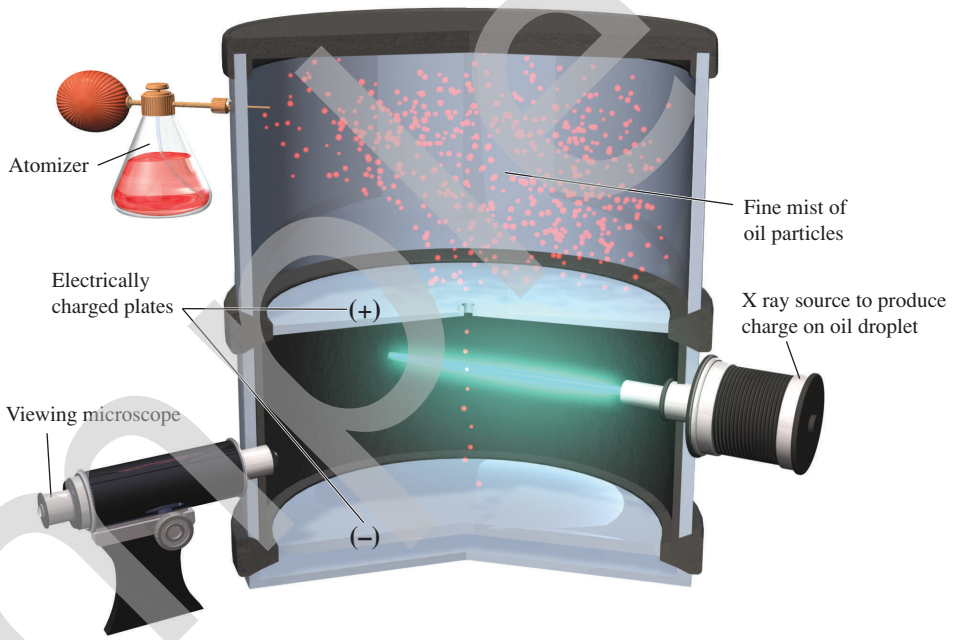


Figure 2.5 *Schematic diagram of Millikan's oil drop experiment.*

Animation
Millikan Oil Drop

An English physicist, J. J. Thomson,[†] used a cathode ray tube and his knowledge of electromagnetic theory to determine the ratio of electric charge to the mass of an individual electron. The number he came up with was -1.76×10^8 C/g, where C stands for *coulomb,* which is the unit of electric charge. Thereafter, in a series of experiments carried out between 1908 and 1917, R. A. Millikan[‡] succeeded in measuring the charge of the electron with great precision. His work proved that the charge on each electron was exactly the same. In his experiment, Millikan examined the motion of single tiny drops of oil that picked up static charge from ions in the air. He suspended the charged drops in air by applying an electric field and followed their motions through a microscope (Figure 2.5). Using his knowledge of electrostatics, Millikan found the charge of an electron to be -1.6022×10^{-19} C. From these data he calculated the mass of an electron:

$$\begin{aligned}\text{mass of an electron} &= \frac{\text{charge}}{\text{charge/mass}} \\ &= \frac{-1.6022 \times 10^{-19}\,\text{C}}{-1.76 \times 10^{8}\,\text{C/g}} \\ &= 9.10 \times 10^{-28}\,\text{g}\end{aligned}$$

This is an exceedingly small mass.

Radioactivity

In 1895 the German physicist Wilhelm Röntgen[§] noticed that cathode rays caused glass and metals to emit very unusual rays. This highly energetic radiation penetrated matter, darkened covered photographic plates, and caused a variety of substances to

[†]Joseph John Thomson (1856–1940). British physicist who received the Nobel Prize in Physics in 1906 for discovering the electron.

[‡]Robert Andrews Millikan (1868–1953). American physicist who was awarded the Nobel Prize in Physics in 1923 for determining the charge of the electron.

[§]Wilhelm Konrad Röntgen (1845–1923). German physicist who received the Nobel Prize in Physics in 1901 for the discovery of X rays.

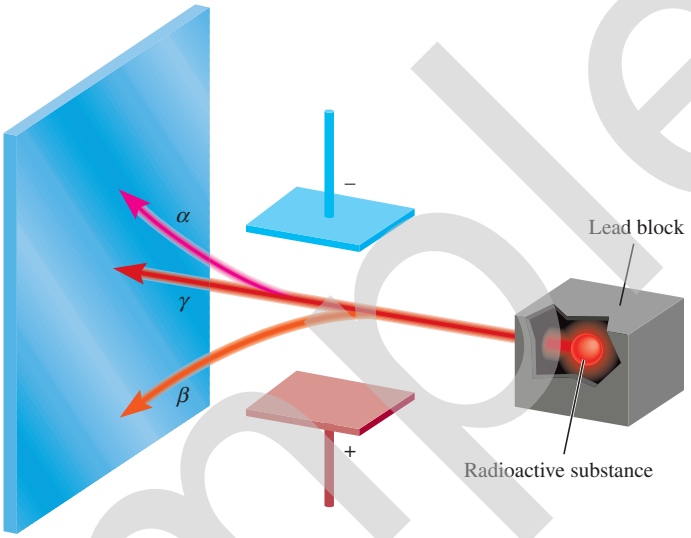


Figure 2.6 *Three types of rays emitted by radioactive elements. β rays consist of negatively charged particles (electrons) and are therefore attracted by the positively charged plate. The opposite holds true for α rays—they are positively charged and are drawn to the negatively charged plate. Because γ rays have no charges, their path is unaffected by an external electric field.*

fluoresce. Because these rays could not be deflected by a magnet, they could not contain charged particles as cathode rays do. Röntgen called them X rays because their nature was not known.

Not long after Röntgen's discovery, Antoine Becquerel,[†] a professor of physics in Paris, began to study the fluorescent properties of substances. Purely by accident, he found that exposing thickly wrapped photographic plates to a certain uranium compound caused them to darken, even without the stimulation of cathode rays. Like X rays, the rays from the uranium compound were highly energetic and could not be deflected by a magnet, but they differed from X rays because they arose spontaneously. One of Becquerel's students, Marie Curie,[‡] suggested the name ***radioactivity*** to describe this *spontaneous emission of particles and/or radiation.* Since then, any element that spontaneously emits radiation is said to be *radioactive.*

Three types of rays are produced by the *decay,* or breakdown, of radioactive substances such as uranium. Two of the three are deflected by oppositely charged metal plates (Figure 2.6). ***Alpha (α) rays*** consist of *positively charged particles,* called ***α particles,*** and therefore are deflected by the positively charged plate. ***Beta (β) rays,*** or ***β particles,*** are electrons and are deflected by the negatively charged plate. The third type of radioactive radiation consists of high-energy rays called ***gamma (γ) rays.*** Like X rays, γ rays have no charge and are not affected by an external field.

Animation
Alpha, Beta, and Gamma Rays

The Proton and the Nucleus

By the early 1900s, two features of atoms had become clear: They contain electrons, and they are electrically neutral. To maintain electric neutrality, an atom must contain an equal number of positive and negative charges. Therefore, Thomson proposed that an atom could be thought of as a uniform, positive sphere of matter in which electrons are embedded like raisins in a cake (Figure 2.7). This so-called "plum-pudding" model was the accepted theory for a number of years.

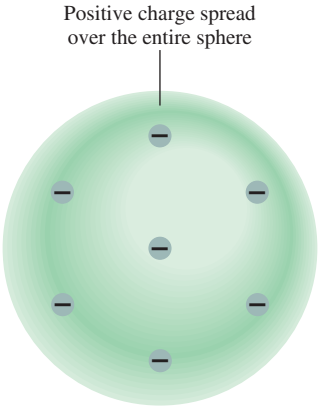


Figure 2.7 *Thomson's model of the atom, sometimes described as the "plum-pudding" model, after a traditional English dessert containing raisins. The electrons are embedded in a uniform, positively charged sphere.*

[†]Antoine Henri Becquerel (1852–1908). French physicist who was awarded the Nobel Prize in Physics in 1903 for discovering radioactivity in uranium.

[‡]Marie (Marya Sklodowska) Curie (1867–1934). Polish-born chemist and physicist. In 1903 she and her French husband, Pierre Curie, were awarded the Nobel Prize in Physics for their work on radioactivity. In 1911, she again received the Nobel prize, this time in chemistry, for her work on the radioactive elements radium and polonium. She is one of only three people to have received two Nobel prizes in science. Despite her great contribution to science, her nomination to the French Academy of Sciences in 1911 was rejected by one vote because she was a woman! Her daughter Irene, and son-in-law Frederic Joliot-Curie, shared the Nobel Prize in Chemistry in 1935.

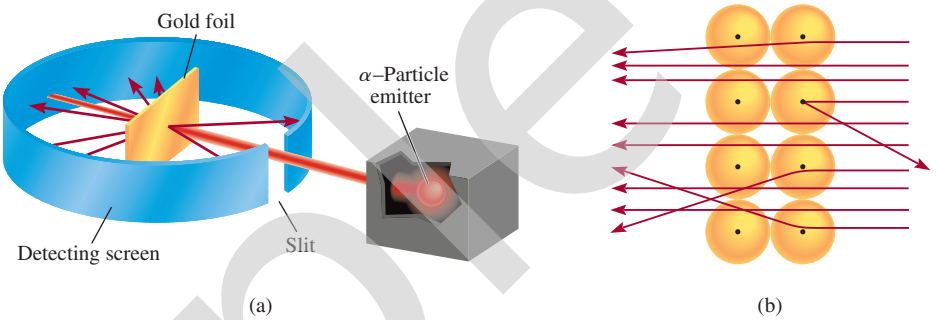


Figure 2.8 *(a) Rutherford's experimental design for measuring the scattering of α particles by a piece of gold foil. Most of the α particles passed through the gold foil with little or no deflection. A few were deflected at wide angles. Occasionally an α particle was turned back. (b) Magnified view of α particles passing through and being deflected by nuclei.*

Animation
α-Particle Scattering

Animation
Rutherford's Experiment

In 1910 the New Zealand physicist Ernest Rutherford,[†] who had studied with Thomson at Cambridge University, decided to use α particles to probe the structure of atoms. Together with his associate Hans Geiger[‡] and an undergraduate named Ernest Marsden,[§] Rutherford carried out a series of experiments using very thin foils of gold and other metals as targets for α particles from a radioactive source (Figure 2.8). They observed that the majority of particles penetrated the foil either undeflected or with only a slight deflection. But every now and then an α particle was scattered (or deflected) at a large angle. In some instances, an α particle actually bounced back in the direction from which it had come! This was a most surprising finding, for in Thomson's model the positive charge of the atom was so diffuse that the positive α particles should have passed through the foil with very little deflection. To quote Rutherford's initial reaction when told of this discovery: "It was as incredible as if you had fired a 15-inch shell at a piece of tissue paper and it came back and hit you."

Rutherford was later able to explain the results of the α-scattering experiment in terms of a new model for the atom. According to Rutherford, most of the atom must be empty space. This explains why the majority of α particles passed through the gold foil with little or no deflection. The atom's positive charges, Rutherford proposed, are all concentrated in the ***nucleus,*** which is *a dense central core within the atom.* Whenever an α particle came close to a nucleus in the scattering experiment, it experienced a large repulsive force and therefore a large deflection. Moreover, an α particle traveling directly toward a nucleus would be completely repelled and its direction would be reversed.

The positively charged particles in the nucleus are called ***protons.*** In separate experiments, it was found that each proton carries the same *quantity* of charge as an electron and has a mass of 1.67262×10^{-24} g—about 1840 times the mass of the oppositely charged electron.

A common non-SI unit for atomic length is the angstrom (Å; 1 Å = 100 pm).

At this stage of investigation, scientists perceived the atom as follows: The mass of a nucleus constitutes most of the mass of the entire atom, but the nucleus occupies only about $1/10^{13}$ of the volume of the atom. We express atomic (and molecular) dimensions in terms of the SI unit called the *picometer (pm),* where

$$1 \text{ pm} = 1 \times 10^{-12} \text{ m}$$

[†]Ernest Rutherford (1871–1937). New Zealand physicist. Rutherford did most of his work in England (Manchester and Cambridge Universities). He received the Nobel Prize in Chemistry in 1908 for his investigations into the structure of the atomic nucleus. His often-quoted comment to his students was that "all science is either physics or stamp-collecting."

[‡]Johannes Hans Wilhelm Geiger (1882–1945). German physicist. Geiger's work focused on the structure of the atomic nucleus and on radioactivity. He invented a device for measuring radiation that is now commonly called the Geiger counter.

[§]Ernest Marsden (1889–1970). English physicist. It is gratifying to know that at times an undergraduate can assist in winning a Nobel prize. Marsden went on to contribute significantly to the development of science in New Zealand.

A typical atomic radius is about 100 pm, whereas the radius of an atomic nucleus is only about 5×10^{-3} pm. You can appreciate the relative sizes of an atom and its nucleus by imagining that if an atom were the size of a sports stadium, the volume of its nucleus would be comparable to that of a small marble. Although the protons are confined to the nucleus of the atom, the electrons are conceived of as being spread out about the nucleus at some distance from it.

The concept of atomic radius is useful experimentally, but we should not infer that atoms have well-defined boundaries or surfaces. We will learn later that the outer regions of atoms are relatively "fuzzy."

If the size of an atom were expanded to that of this sports stadium, the size of the nucleus would be that of a marble.

The Neutron

Rutherford's model of atomic structure left one major problem unsolved. It was known that hydrogen, the simplest atom, contains only one proton and that the helium atom contains two protons. Therefore, the ratio of the mass of a helium atom to that of a hydrogen atom should be 2:1. (Because electrons are much lighter than protons, their contribution to atomic mass can be ignored.) In reality, however, the ratio is 4:1. Rutherford and others postulated that there must be another type of subatomic particle in the atomic nucleus; the proof was provided by another English physicist, James Chadwick,[†] in 1932. When Chadwick bombarded a thin sheet of beryllium with α particles, a very high-energy radiation similar to γ rays was emitted by the metal. Later experiments showed that the rays actually consisted of a third type of subatomic particles, which Chadwick named ***neutrons,*** because they proved to be *electrically neutral particles having a mass slightly greater than that of protons.* The mystery of the mass ratio could now be explained. In the helium nucleus there are two protons and two neutrons, but in the hydrogen nucleus there is only one proton and no neutrons; therefore, the ratio is 4:1.

Figure 2.9 shows the location of the elementary particles (protons, neutrons, and electrons) in an atom. There are other subatomic particles, but the electron, the

[†]James Chadwick (1891–1972). British physicist. In 1935 he received the Nobel Prize in Physics for proving the existence of neutrons.

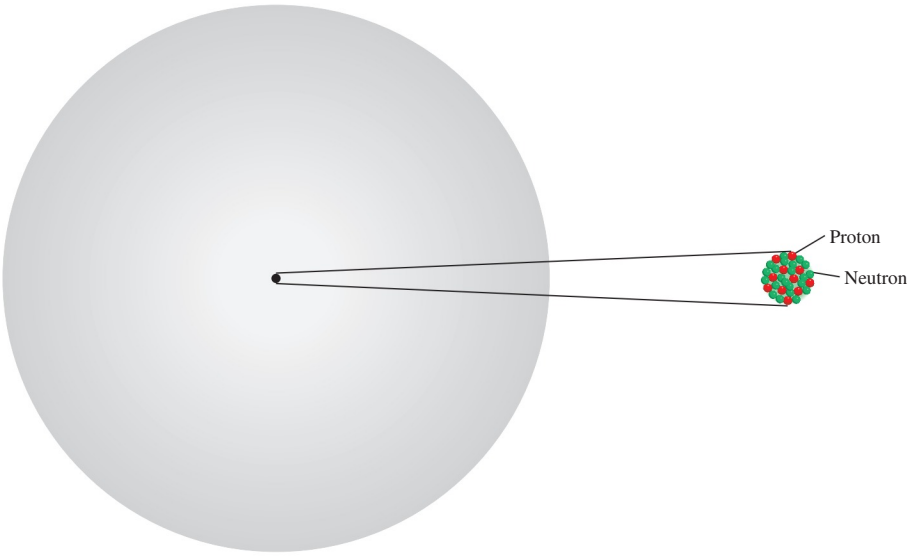


Figure 2.9 *The protons and neutrons of an atom are packed in an extremely small nucleus. Electrons are shown as "clouds" around the nucleus.*

Table 2.1 Mass and Charge of Subatomic Particles

Particle	Mass (g)	Charge: Coulomb	Charge: Charge Unit
Electron*	9.10938×10^{-28}	-1.6022×10^{-19}	-1
Proton	1.67262×10^{-24}	$+1.6022 \times 10^{-19}$	$+1$
Neutron	1.67493×10^{-24}	0	0

*More refined measurements have given us a more accurate value of an electron's mass than Millikan's.

proton, and the neutron are the three fundamental components of the atom that are important in chemistry. Table 2.1 shows the masses and charges of these three elementary particles.

2.3 Atomic Number, Mass Number, and Isotopes

All atoms can be identified by the number of protons and neutrons they contain. The ***atomic number (Z)*** is *the number of protons in the nucleus of each atom of an element.* In a neutral atom the number of protons is equal to the number of electrons, so the atomic number also indicates the number of electrons present in the atom. The chemical identity of an atom can be determined solely from its atomic number. For example, the atomic number of fluorine is 9. This means that each fluorine atom has 9 protons and 9 electrons. Or, viewed another way, every atom in the universe that contains 9 protons is correctly named "fluorine."

Protons and neutrons are collectively called *nucleons*.

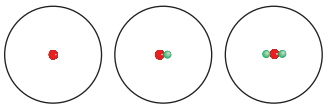
The ***mass number (A)*** is *the total number of neutrons and protons present in the nucleus of an atom of an element.* Except for the most common form of hydrogen, which has one proton and no neutrons, all atomic nuclei contain both protons and neutrons. In general, the mass number is given by

$$\begin{aligned}\text{mass number} &= \text{number of protons} + \text{number of neutrons}\\ &= \text{atomic number} + \text{number of neutrons}\end{aligned} \tag{2.1}$$

The number of neutrons in an atom is equal to the difference between the mass number and the atomic number, or $(A - Z)$. For example, if the mass number of a particular boron atom is 12 and the atomic number is 5 (indicating 5 protons in the nucleus), then the number of neutrons is $12 - 5 = 7$. Note that all three quantities (atomic number, number of neutrons, and mass number) must be positive integers, or whole numbers.

Atoms of a given element do not all have the same mass. Most elements have two or more ***isotopes,*** *atoms that have the same atomic number but different mass numbers.* For example, there are three isotopes of hydrogen. One, simply known as hydrogen, has one proton and no neutrons. The *deuterium* isotope contains one proton and one neutron, and *tritium* has one proton and two neutrons. The accepted way to denote the atomic number and mass number of an atom of an element (X) is as follows:

mass number → $^{A}_{Z}X$ ← atomic number



$^{1}_{1}H$ $^{2}_{1}H$ $^{3}_{1}H$

Thus, for the isotopes of hydrogen, we write

$^{1}_{1}H$	$^{2}_{1}H$	$^{3}_{1}H$
hydrogen	deuterium	tritium

As another example, consider two common isotopes of uranium with mass numbers of 235 and 238, respectively:

$$^{235}_{92}\text{U} \qquad ^{238}_{92}\text{U}$$

The first isotope is used in nuclear reactors and atomic bombs, whereas the second isotope lacks the properties necessary for these applications. With the exception of hydrogen, which has different names for each of its isotopes, isotopes of elements are identified by their mass numbers. Thus, the preceding two isotopes are called uranium-235 (pronounced "uranium two thirty-five") and uranium-238 (pronounced "uranium two thirty-eight").

The chemical properties of an element are determined primarily by the protons and electrons in its atoms; neutrons do not take part in chemical changes under normal conditions. Therefore, isotopes of the same element have similar chemistries, forming the same types of compounds and displaying similar reactivities.

Example 2.1 shows how to calculate the number of protons, neutrons, and electrons using atomic numbers and mass numbers.

Example 2.1

Give the number of protons, neutrons, and electrons in each of the following species: (a) $^{20}_{11}\text{Na}$, (b) $^{22}_{11}\text{Na}$, (c) ^{17}O, and (d) carbon-14.

Strategy Recall that the superscript denotes the mass number (A) and the subscript denotes the atomic number (Z). Mass number is always greater than atomic number. (The only exception is $^{1}_{1}\text{H}$, where the mass number is equal to the atomic number.) In a case where no subscript is shown, as in parts (c) and (d), the atomic number can be deduced from the element symbol or name. To determine the number of electrons, remember that because atoms are electrically neutral, the number of electrons is equal to the number of protons.

Solution

(a) The atomic number is 11, so there are 11 protons. The mass number is 20, so the number of neutrons is $20 - 11 = 9$. The number of electrons is the same as the number of protons; that is, 11.

(b) The atomic number is the same as that in (a), or 11. The mass number is 22, so the number of neutrons is $22 - 11 = 11$. The number of electrons is 11. Note that the species in (a) and (b) are chemically similar isotopes of sodium.

(c) The atomic number of O (oxygen) is 8, so there are 8 protons. The mass number is 17, so there are $17 - 8 = 9$ neutrons. There are 8 electrons.

(d) Carbon-14 can also be represented as ^{14}C. The atomic number of carbon is 6, so there are $14 - 6 = 8$ neutrons. The number of electrons is 6.

Similar problems: 2.15, 2.16.

Practice Exercise How many protons, neutrons, and electrons are in the following isotope of copper: ^{63}Cu?

Review of Concepts

(a) What is the atomic number of an element if one of its isotopes has 117 neutrons and a mass number of 195?

(b) Which of the following two symbols provides more information? ^{17}O or $_{8}\text{O}$.

2.4 The Periodic Table

More than half of the elements known today were discovered between 1800 and 1900. During this period, chemists noted that many elements show strong similarities to one another. Recognition of periodic regularities in physical and chemical behavior and the need to organize the large volume of available information about the structure and properties of elemental substances led to the development of the ***periodic table,*** *a chart in which elements having similar chemical and physical properties are grouped together.* Figure 2.10 shows the modern periodic table in which the elements are arranged by atomic number (shown above the element symbol) in *horizontal rows* called ***periods*** and in *vertical columns* known as ***groups*** or ***families,*** according to similarities in their chemical properties. Note that elements 113–118 have recently been synthesized, although they have not yet been named.

The elements can be divided into three categories—metals, nonmetals, and metalloids. A ***metal*** is *a good conductor of heat and electricity* while a ***nonmetal*** is usually *a poor conductor of heat and electricity.* A ***metalloid*** *has properties that are intermediate between those of metals and nonmetals.* Figure 2.10 shows that the majority of known elements are metals; only 17 elements are nonmetals, and 8 elements are metalloids. From left to right across any period, the physical and chemical properties of the elements change gradually from metallic to nonmetallic.

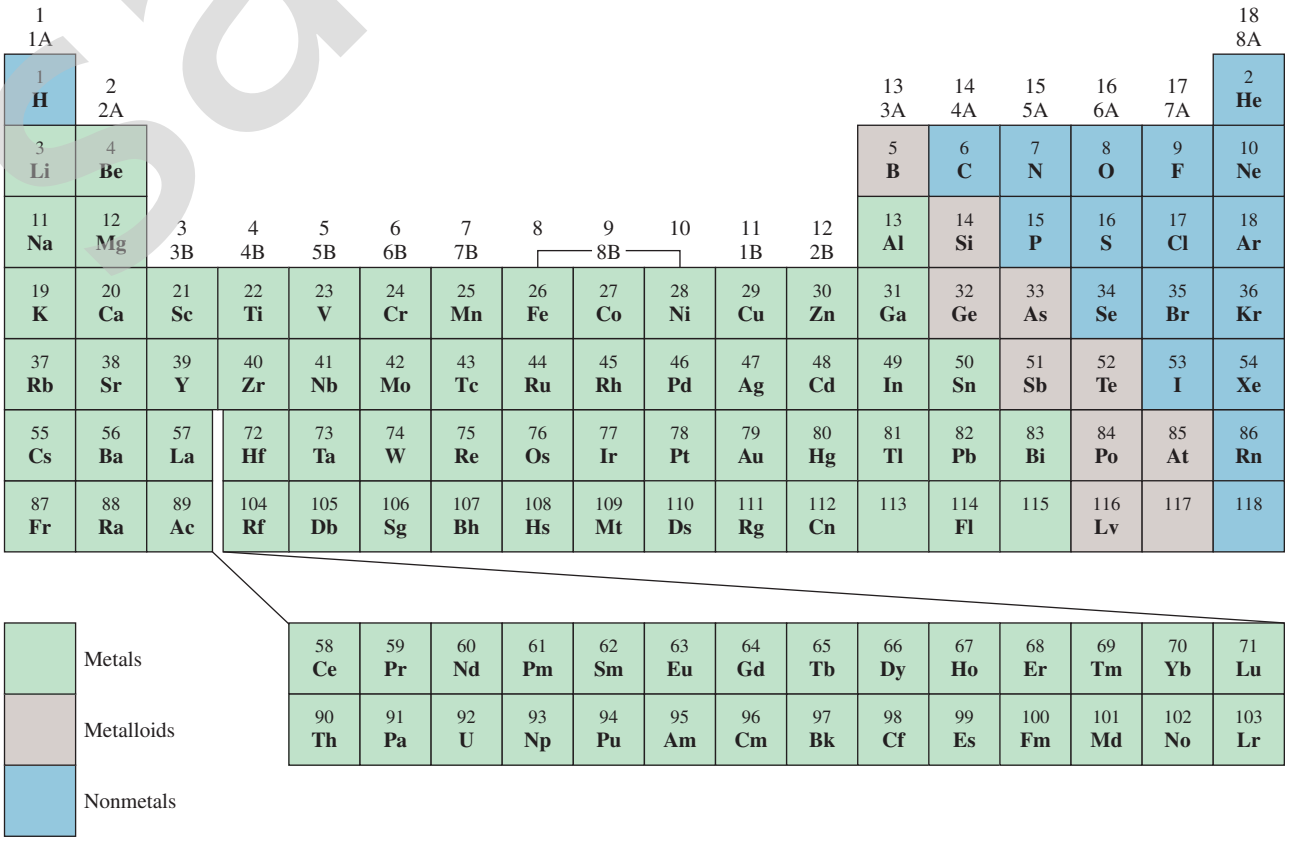


Figure 2.10 *The modern periodic table. The elements are arranged according to the atomic numbers above their symbols. With the exception of hydrogen (H), nonmetals appear at the far right of the table. The two rows of metals beneath the main body of the table are conventionally set apart to keep the table from being too wide. Actually, cerium (Ce) should follow lanthanum (La), and thorium (Th) should come right after actinium (Ac). The 1–18 group designation has been recommended by the International Union of Pure and Applied Chemistry (IUPAC) but is not yet in wide use. In this text, we use the standard U.S. notation for group numbers (1A–8A and 1B–8B). No names have yet been assigned to elements 113, 115, 117, and 118.*

CHEMISTRY *in Action*

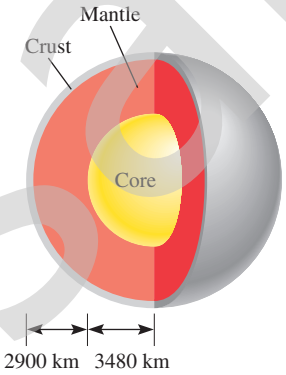
Distribution of Elements on Earth and in Living Systems

The majority of elements are naturally occurring. How are these elements distributed on Earth, and which are essential to living systems?

Earth's crust extends from the surface to a depth of about 40 km (about 25 mi). Because of technical difficulties, scientists have not been able to study the inner portions of Earth as easily as the crust. Nevertheless, it is believed that there is a solid core consisting mostly of iron at the center of Earth. Surrounding the core is a layer called the *mantle,* which consists of hot fluid containing iron, carbon, silicon, and sulfur.

Of the 83 elements that are found in nature, 12 make up 99.7 percent of Earth's crust by mass. They are, in decreasing order of natural abundance, oxygen (O), silicon (Si), aluminum (Al), iron (Fe), calcium (Ca), magnesium (Mg), sodium (Na), potassium (K), titanium (Ti), hydrogen (H), phosphorus (P), and manganese (Mn). In discussing the natural abundance of the elements, we should keep in mind that (1) the elements are not evenly distributed throughout Earth's crust, and (2) most elements occur in combined forms. These facts provide the basis for most methods of obtaining pure elements from their compounds, as we will see in later chapters.

The accompanying table lists the essential elements in the human body. Of special interest are the *trace elements,* such as iron (Fe), copper (Cu), zinc (Zn), iodine (I), and cobalt (Co), which together make up about 0.1 percent of the body's mass. These elements are necessary for biological functions such as growth, transport of oxygen for metabolism, and defense against disease. There is a delicate balance in the amounts of these elements in our bodies. Too much or too little over an extended period of time can lead to serious illness, retardation, or even death.

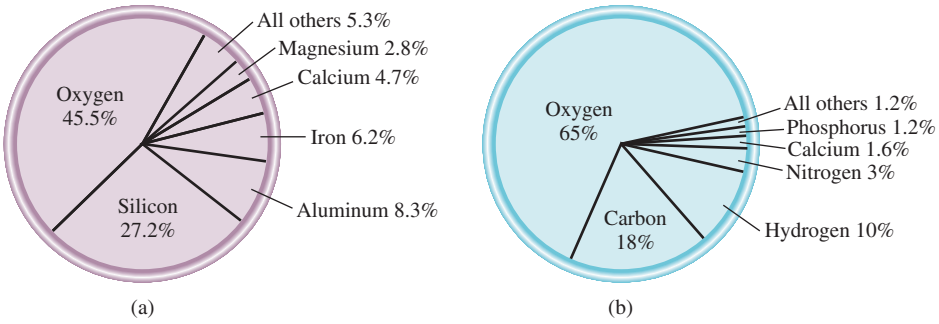


Structure of Earth's interior.

Essential Elements in the Human Body

Element	Percent by Mass*	Element	Percent by Mass*
Oxygen	65	Sodium	0.1
Carbon	18	Magnesium	0.05
Hydrogen	10	Iron	<0.05
Nitrogen	3	Cobalt	<0.05
Calcium	1.6	Copper	<0.05
Phosphorus	1.2	Zinc	<0.05
Potassium	0.2	Iodine	<0.05
Sulfur	0.2	Selenium	<0.01
Chlorine	0.2	Fluorine	<0.01

*Percent by mass *gives the mass of the element in grams present in a 100-g sample.*



(a) Natural abundance of the elements in percent by mass. For example, oxygen's abundance is 45.5 percent. This means that in a 100-g sample of Earth's crust there are, on the average, 45.5 g of the element oxygen.
(b) Abundance of elements in the human body in percent by mass.

Elements are often referred to collectively by their periodic table group number (Group 1A, Group 2A, and so on). However, for convenience, some element groups have been given special names. *The Group 1A elements (Li, Na, K, Rb, Cs, and Fr)* are called ***alkali metals,*** and *the Group 2A elements (Be, Mg, Ca, Sr, Ba, and Ra)* are called ***alkaline earth metals.*** *Elements in Group 7A (F, Cl, Br, I, and At)* are known as ***halogens,*** and *elements in Group 8A (He, Ne, Ar, Kr, Xe, and Rn)* are called ***noble gases,*** or *rare gases.*

The periodic table is a handy tool that correlates the properties of the elements in a systematic way and helps us to make predictions about chemical behavior. We will take a closer look at this keystone of chemistry in Chapter 8.

The Chemistry in Action essay on p. 49 describes the distribution of the elements on Earth and in the human body.

Review of Concepts

In viewing the periodic table, do chemical properties change more markedly across a period or down a group?

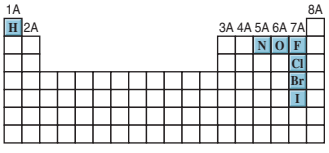
2.5 Molecules and Ions

Of all the elements, only the six noble gases in Group 8A of the periodic table (He, Ne, Ar, Kr, Xe, and Rn) exist in nature as single atoms. For this reason, they are called *monatomic* (meaning a single atom) gases. Most matter is composed of molecules or ions formed by atoms.

Molecules

We will discuss the nature of chemical bonds in Chapters 9 and 10.

A ***molecule*** is an *aggregate of at least two atoms in a definite arrangement held together by chemical forces* (also called *chemical bonds*). A molecule may contain atoms of the same element or atoms of two or more elements joined in a fixed ratio, in accordance with the law of definite proportions stated in Section 2.1. Thus, a molecule is not necessarily a compound, which, by definition, is made up of two or more elements (see Section 1.4). Hydrogen gas, for example, is a pure element, but it consists of molecules made up of two H atoms each. Water, on the other hand, is a molecular compound that contains hydrogen and oxygen in a ratio of two H atoms and one O atom. Like atoms, molecules are electrically neutral.



Elements that exist as diatomic molecules.

The hydrogen molecule, symbolized as H_2, is called a ***diatomic molecule*** because it *contains only two atoms.* Other elements that normally exist as diatomic molecules are nitrogen (N_2) and oxygen (O_2), as well as the Group 7A elements—fluorine (F_2), chlorine (Cl_2), bromine (Br_2), and iodine (I_2). Of course, a diatomic molecule can contain atoms of different elements. Examples are hydrogen chloride (HCl) and carbon monoxide (CO).

The vast majority of molecules contain more than two atoms. They can be atoms of the same element, as in ozone (O_3), which is made up of three atoms of oxygen, or they can be combinations of two or more different elements. *Molecules containing more than two atoms* are called ***polyatomic molecules.*** Like ozone, water (H_2O) and ammonia (NH_3) are polyatomic molecules.

Ions

An ***ion*** is *an atom or a group of atoms that has a net positive or negative charge.* The number of positively charged protons in the nucleus of an atom remains the same during ordinary chemical changes (called chemical reactions), but negatively charged

electrons may be lost or gained. The loss of one or more electrons from a neutral atom results in a ***cation,*** *an ion with a net positive charge.* For example, a sodium atom (Na) can readily lose an electron to become a sodium cation, which is represented by Na^+:

In Chapter 8 we will see why atoms of different elements gain (or lose) a specific number of electrons.

Na Atom	Na^+ Ion
11 protons	11 protons
11 electrons	10 electrons

On the other hand, an ***anion*** is *an ion whose net charge is negative* due to an increase in the number of electrons. A chlorine atom (Cl), for instance, can gain an electron to become the chloride ion Cl^-:

Cl Atom	Cl^- Ion
17 protons	17 protons
17 electrons	18 electrons

Sodium chloride (NaCl), ordinary table salt, is called an ***ionic compound*** because it is *formed from cations and anions.*

An atom can lose or gain more than one electron. Examples of ions formed by the loss or gain of more than one electron are Mg^{2+}, Fe^{3+}, S^{2-}, and N^{3-}. These ions, as well as Na^+ and Cl^-, are called ***monatomic ions*** because they *contain only one atom.* Figure 2.11 shows the charges of a number of monatomic ions. With very few exceptions, metals tend to form cations and nonmetals form anions.

In addition, two or more atoms can combine to form an ion that has a net positive or net negative charge. ***Polyatomic ions*** such as OH^- (hydroxide ion), CN^- (cyanide ion), and NH_4^+ (ammonium ion) are *ions containing more than one atom.*

Review of Concepts

(a) What does S_8 signify? How does it differ from 8S?
(b) Determine the number of protons and electrons for the following ions: (a) P^{3-} and (b) Ti^{4+}.

1 1A	2 2A	3 3B	4 4B	5 5B	6 6B	7 7B	8	9 8B	10	11 1B	12 2B	13 3A	14 4A	15 5A	16 6A	17 7A	18 8A
Li^+													C^{4-}	N^{3-}	O^{2-}	F^-	
Na^+	Mg^{2+}											Al^{3+}		P^{3-}	S^{2-}	Cl^-	
K^+	Ca^{2+}				Cr^{2+} Cr^{3+}	Mn^{2+} Mn^{3+}	Fe^{2+} Fe^{3+}	Co^{2+} Co^{3+}	Ni^{2+} Ni^{3+}	Cu^+ Cu^{2+}	Zn^{2+}				Se^{2-}	Br^-	
Rb^+	Sr^{2+}									Ag^+	Cd^{2+}		Sn^{2+} Sn^{4+}		Te^{2-}	I^-	
Cs^+	Ba^{2+}									Au^+ Au^{3+}	Hg_2^{2+} Hg^{2+}		Pb^{2+} Pb^{4+}				

Figure 2.11 *Common monatomic ions arranged according to their positions in the periodic table. Note that the Hg_2^{2+} ion contains two atoms.*

2.6 Chemical Formulas

Chemists use ***chemical formulas*** to *express the composition of molecules and ionic compounds in terms of chemical symbols.* By composition we mean not only the elements present but also the ratios in which the atoms are combined. Here we are concerned with two types of formulas: molecular formulas and empirical formulas.

Molecular Formulas

A ***molecular formula*** *shows the exact number of atoms of each element in the smallest unit of a substance.* In our discussion of molecules, each example was given with its molecular formula in parentheses. Thus, H_2 is the molecular formula for hydrogen, O_2 is oxygen, O_3 is ozone, and H_2O is water. The subscript numeral indicates the number of atoms of an element present. There is no subscript for O in H_2O because there is only one atom of oxygen in a molecule of water, and so the number "one" is omitted from the formula. Note that oxygen (O_2) and ozone (O_3) are allotropes of oxygen. An ***allotrope*** is *one of two or more distinct forms of an element.* Two allotropic forms of the element carbon—diamond and graphite—are dramatically different not only in properties but also in their relative cost.

Molecular Models

Molecules are too small for us to observe directly. An effective means of visualizing them is by the use of molecular models. Two standard types of molecular models are currently in use: *ball-and-stick* models and *space-filling* models (Figure 2.12). In ball-and-stick model kits, the atoms are wooden or plastic balls with holes in them. Sticks or springs are used to represent chemical bonds. The angles they form between atoms approximate the bond angles in actual molecules. With the exception of the H atom, the balls are all the same size and each type of atom is represented by a specific color. In space-filling models, atoms are represented by truncated balls held together by snap

See back endpaper for color codes for atoms.

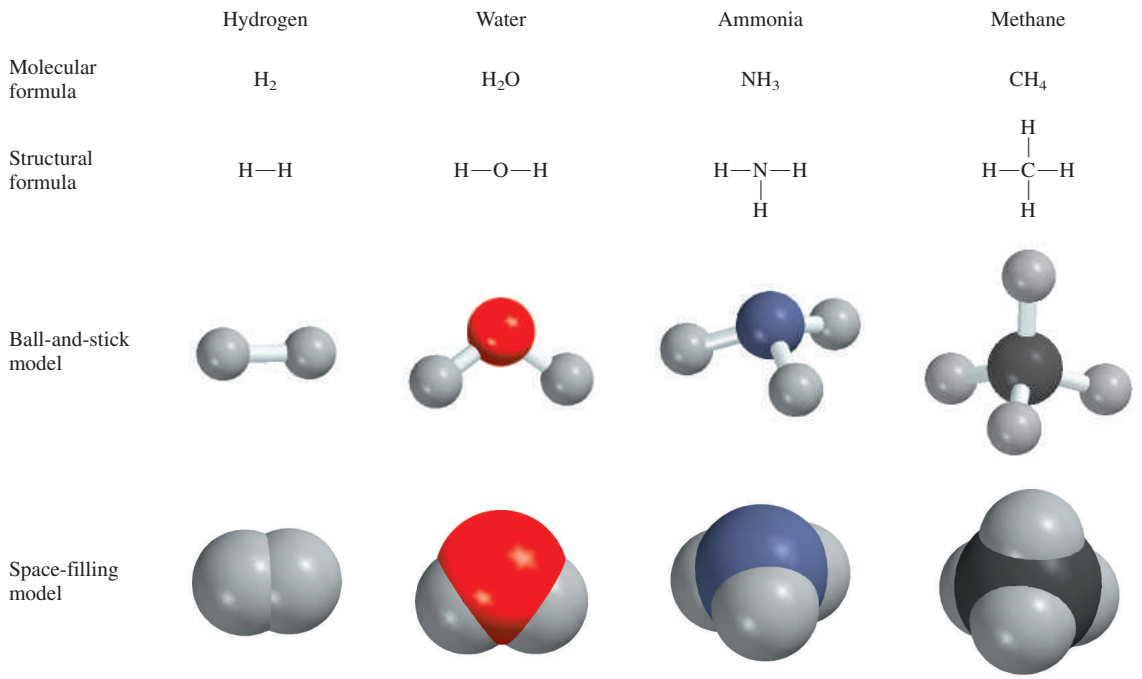


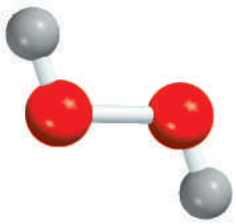
Figure 2.12 *Molecular and structural formulas and molecular models of four common molecules.*

fasteners, so that the bonds are not visible. The balls are proportional in size to atoms. The first step toward building a molecular model is writing the ***structural formula,*** which *shows how atoms are bonded to one another in a molecule.* For example, it is known that each of the two H atoms is bonded to an O atom in the water molecule. Therefore, the structural formula of water is H—O—H. A line connecting the two atomic symbols represents a chemical bond.

Ball-and-stick models show the three-dimensional arrangement of atoms clearly, and they are fairly easy to construct. However, the balls are not proportional to the size of atoms. Furthermore, the sticks greatly exaggerate the space between atoms in a molecule. Space-filling models are more accurate because they show the variation in atomic size. Their drawbacks are that they are time-consuming to put together and they do not show the three-dimensional positions of atoms very well. Molecular modeling software can also be used to create ball-and-stick and space-filling models. We will use both models extensively in this text.

Empirical Formulas

The molecular formula of hydrogen peroxide, a substance used as an antiseptic and as a bleaching agent for textiles and hair, is H_2O_2. This formula indicates that each hydrogen peroxide molecule consists of two hydrogen atoms and two oxygen atoms. The ratio of hydrogen to oxygen atoms in this molecule is 2:2 or 1:1. The empirical formula of hydrogen peroxide is HO. Thus, the ***empirical formula*** *tells us which elements are present and the simplest whole-number ratio of their atoms,* but not necessarily the actual number of atoms in a given molecule. As another example, consider the compound hydrazine (N_2H_4), which is used as a rocket fuel. The empirical formula of hydrazine is NH_2. Although the ratio of nitrogen to hydrogen is 1:2 in both the molecular formula (N_2H_4) and the empirical formula (NH_2), only the molecular formula tells us the actual number of N atoms (two) and H atoms (four) present in a hydrazine molecule.



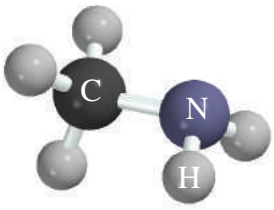
H_2O_2

Empirical formulas are the *simplest* chemical formulas; they are written by reducing the subscripts in the molecular formulas to the smallest possible whole numbers. Molecular formulas are the *true* formulas of molecules. If we know the molecular formula, we also know the empirical formula, but the reverse is not true. Why, then, do chemists bother with empirical formulas? As we will see in Chapter 3, when chemists analyze an unknown compound, the first step is usually the determination of the compound's empirical formula. With additional information, it is possible to deduce the molecular formula.

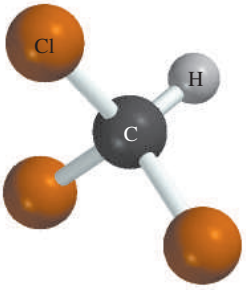
The word "empirical" means "derived from experiment." As we will see in Chapter 3, empirical formulas are determined experimentally.

For many molecules, the molecular formula and the empirical formula are one and the same. Some examples are water (H_2O), ammonia (NH_3), carbon dioxide (CO_2), and methane (CH_4).

Examples 2.2 and 2.3 deal with writing molecular formulas from molecular models and writing empirical formulas from molecular formulas.



Methylamine



Chloroform

Example 2.2

Write the molecular formula of methylamine, a colorless gas used in the production of pharmaceuticals and pesticides, from its ball-and-stick model, shown in the margin.

Solution Refer to the labels (also see back end papers). There are five H atoms, one C atom, and one N atom. Therefore, the molecular formula is CH_5N. However, the standard way of writing the molecular formula for methylamine is CH_3NH_2 because it shows how the atoms are joined in the molecule.

Practice Exercise Write the molecular formula of chloroform, which is used as a solvent and a cleaning agent. The ball-and-stick model of chloroform is shown in the margin.

Similar problems: 2.47, 2.48.

Example 2.3

Write the empirical formulas for the following molecules: (a) diborane (B_2H_6), used in rocket propellants; (b) dimethyl fumarate ($C_8H_{12}O_4$), a substance used to treat psoriasis, a skin disease; and (c) vanillin ($C_8H_8O_3$), a flavoring agent used in foods and beverages.

Strategy Recall that to write the empirical formula, the subscripts in the molecular formula must be converted to the smallest possible whole numbers.

Solution

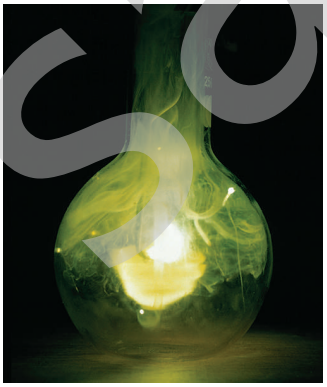
(a) There are two boron atoms and six hydrogen atoms in diborane. Dividing the subscripts by 2, we obtain the empirical formula BH_3.

(b) In dimethyl fumarate there are 8 carbon atoms, 12 hydrogen atoms, and 4 oxygen atoms. Dividing the subscripts by 4, we obtain the empirical formula C_2H_3O. Note that if we had divided the subscripts by 2, we would have obtained the formula $C_4H_6O_2$. Although the ratio of carbon to hydrogen to oxygen atoms in $C_4H_6O_2$ is the same as that in C_2H_3O (2:3:1), $C_4H_6O_2$ is not the simplest formula because its subscripts are not in the smallest whole-number ratio.

(c) Because the subscripts in $C_8H_8O_3$ are already the smallest possible whole numbers, the empirical formula for vanillin is the same as its molecular formula.

Similar problems: 2.45, 2.46.

Practice Exercise Write the empirical formula for caffeine ($C_8H_{10}N_4O_2$), a stimulant found in tea and coffee.

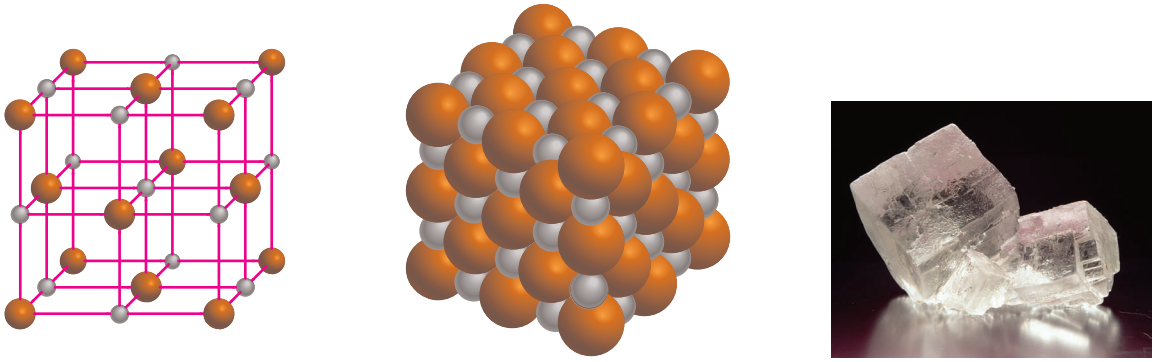


Sodium metal reacting with chlorine gas to form sodium chloride.

Formula of Ionic Compounds

The formulas of ionic compounds are usually the same as their empirical formulas because ionic compounds do not consist of discrete molecular units. For example, a solid sample of sodium chloride (NaCl) consists of equal numbers of Na^+ and Cl^- ions arranged in a three-dimensional network (Figure 2.13). In such a compound there is a 1:1 ratio of cations to anions so that the compound is electrically neutral. As you can see in Figure 2.13, no Na^+ ion in NaCl is associated with just one particular Cl^- ion. In fact, each Na^+ ion is equally held by six surrounding Cl^- ions and vice versa. Thus, NaCl is the empirical formula for sodium chloride. In other ionic compounds, the actual structure may be different, but the arrangement of cations and anions is such that the compounds are all electrically neutral. Note that the charges on the cation and anion are not shown in the formula for an ionic compound.

For ionic compounds to be electrically neutral, the sum of the charges on the cation and anion in each formula unit must be zero. If the charges on the cation and anion are numerically different, we apply the following rule to make the formula



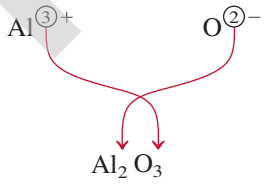
(a) (b) (c)

Figure 2.13 *(a) Structure of solid NaCl. (b) In reality, the cations are in contact with the anions. In both (a) and (b), the smaller spheres represent Na^+ ions and the larger spheres, Cl^- ions. (c) Crystals of NaCl.*

electrically neutral: *The subscript of the cation is numerically equal to the charge on the anion, and the subscript of the anion is numerically equal to the charge on the cation.* If the charges are numerically equal, then no subscripts are necessary. This rule follows from the fact that because the formulas of ionic compounds are usually empirical formulas, the subscripts must always be reduced to the smallest ratios. Let us consider some examples.

Refer to Figure 2.11 for charges of cations and anions.

- **Potassium Bromide.** The potassium cation K^+ and the bromine anion Br^- combine to form the ionic compound potassium bromide. The sum of the charges is $+1 + (-1) = 0$, so no subscripts are necessary. The formula is KBr.
- **Zinc Iodide.** The zinc cation Zn^{2+} and the iodine anion I^- combine to form zinc iodide. The sum of the charges of one Zn^{2+} ion and one I^- ion is $+2 + (-1) = +1$. To make the charges add up to zero we multiply the -1 charge of the anion by 2 and add the subscript "2" to the symbol for iodine. Therefore the formula for zinc iodide is ZnI_2.
- **Aluminum Oxide.** The cation is Al^{3+} and the oxygen anion is O^{2-}. The following diagram helps us determine the subscripts for the compound formed by the cation and the anion:



The sum of the charges is $2(+3) + 3(-2) = 0$. Thus, the formula for aluminum oxide is Al_2O_3.

Note that in each of the three examples, the subscripts are in the smallest ratios.

Example 2.4

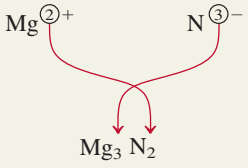
Magnesium nitride is used to prepare Borazon, a very hard compound employed in cutting tools and machine parts. Write the formula of magnesium nitride, containing the Mg^{2+} and N^{3-} ions.

Strategy Our guide for writing formulas for ionic compounds is electrical neutrality; that is, the total charge on the cation(s) must be equal to the total charge on the anion(s). Because the charges on the Mg^{2+} and N^{3-} ions are not equal, we know the formula cannot be MgN. Instead, we write the formula as Mg_xN_y, where x and y are subscripts to be determined.

Solution To satisfy electrical neutrality, the following relationship must hold:

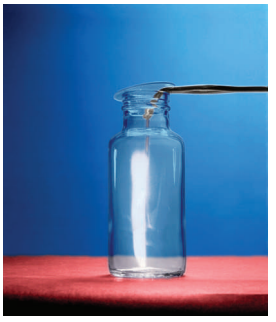
$$(+2)x + (-3)y = 0$$

Solving, we obtain $x/y = 3/2$. Setting $x = 3$ and $y = 2$, we write



Check The subscripts are reduced to the smallest whole-number ratio of the atoms because the chemical formula of an ionic compound is usually its empirical formula.

Practice Exercise Write the formulas of the following ionic compounds: (a) chromium sulfate (containing the Cr^{3+} and SO_4^{2-} ions) and (b) titanium oxide (containing the Ti^{4+} and O^{2-} ions).



When magnesium burns in air, it forms both magnesium oxide and magnesium nitride.

Similar problems: 2.43, 2.44.