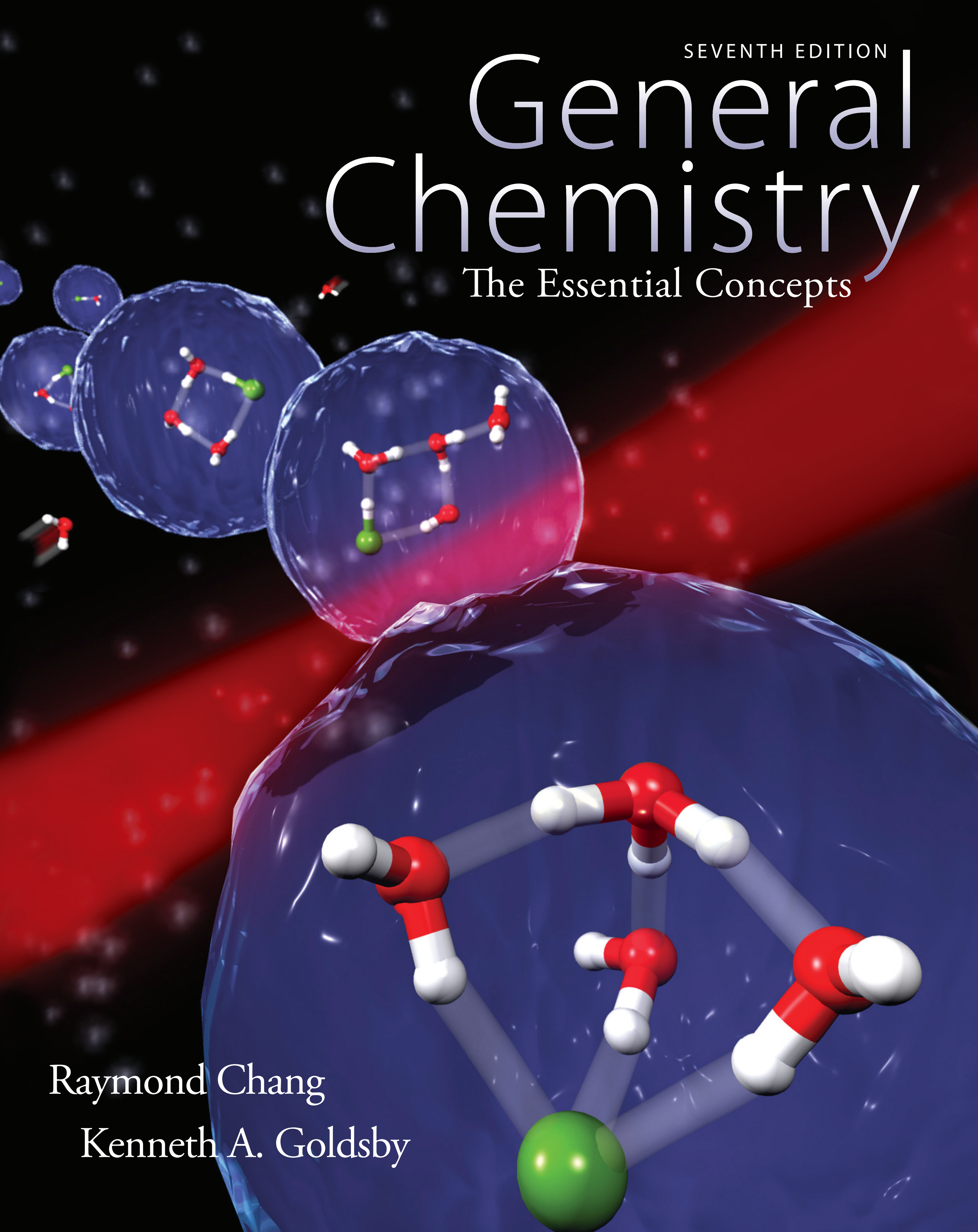
SEVENTH EDITION

General Chemistry

The Essential Concepts



Raymond Chang

Kenneth A. Goldsby

General Chemistry

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The Essential Concepts

SEVENTH EDITION

Raymond Chang

Williams College

Kenneth A. Goldsby

Florida State University

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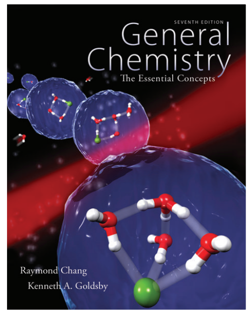
About the Cover

Exactly four water molecules and one hydrogen chloride molecule are necessary to form the smallest droplet of acid. This result was obtained by the groups of Prof. Dr. Martina Havenith (physical chemistry) and Prof. Dr. Dominik Marx (theoretical chemistry) at the Ruhr-Universität Bochum within the research unit FOR 618. They have carried out experiments at ultracold temperatures close to absolute zero temperature using infrared laser spectroscopy to monitor the molecules.



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GENERAL CHEMISTRY: THE ESSENTIAL CONCEPTS, SEVENTH EDITION

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About the Authors

Raymond Chang was born in Hong Kong and grew up in Shanghai and Hong Kong. He received his B.Sc. degree in chemistry from London University, England, and his Ph.D. in chemistry from Yale University. After doing postdoctoral research at Washington University and teaching for a year at Hunter College of the City University of New York, he joined the chemistry department at Williams College, where he has taught since 1968.

Professor Chang has served on the American Chemical Society Examination Committee, the National Chemistry Olympiad Examination Committee, and the Graduate Record Examinations (GRE) Committee. He is an editor of *The Chemical Educator*. Professor Chang has written books on physical chemistry, industrial chemistry, and physical science. He has also coauthored books on the Chinese language, children's picture books, and a novel for young readers.

For relaxation, Professor Chang maintains a forest garden; plays tennis, Ping-Pong, and the harmonica; and practices the piano.

Ken Goldsby was born and raised in Pensacola, Florida. He received his B.A. in chemistry and mathematical science from Rice University. After obtaining his Ph.D. in chemistry from the University of North Carolina at Chapel Hill, Ken carried out postdoctoral research at Ohio State University.

Since joining the Department of Chemistry and Biochemistry at Florida State University in 1986, Ken has received several teaching and advising awards, including the Cottrell Family Professorship for Teaching in Chemistry. In 1998 he was selected as the Florida State University Distinguished Teaching Professor.

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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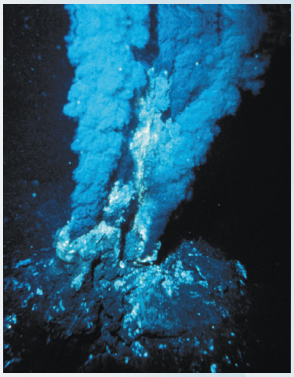

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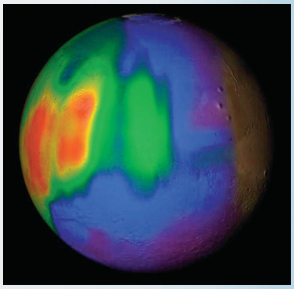

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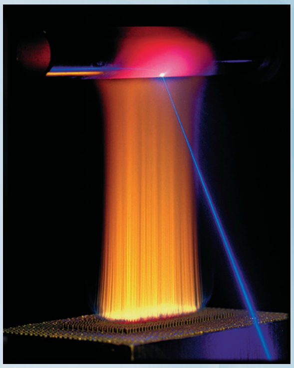

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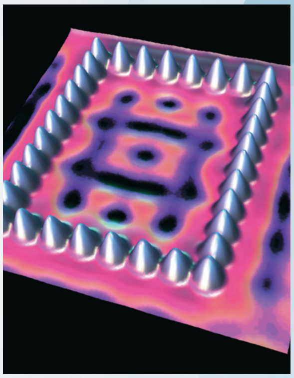

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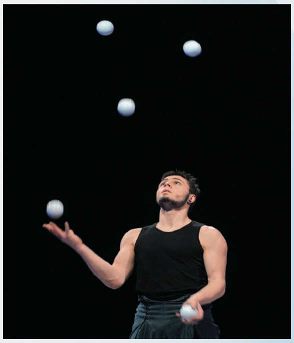

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List of Animations

The animations below are correlated to ***General Chemistry***. Within the chapters are icons letting the student and instructor know that an animation is available for a specific topic and where to find the animation for viewing on our Chang *General Chemistry* companion and Connect websites.

Chang Animations

McGraw-Hill Animations

Simulations

Preface

The seventh edition of *General Chemistry: The Essential Concepts* continues the tradition of presenting the material that is essential to a one-year general chemistry course. It includes all the core topics that are necessary for a solid foundation in general chemistry without sacrificing depth, clarity, or rigor. The positive feedback from users over the years shows that there continues to be a strong need for a concise but thorough text containing all of the core concepts necessary for a solid foundation in general chemistry. *General Chemistry* covers the essential topics in the same depth and at the same level as much longer texts. The reduction in length in this text is achieved in large part by omitting chapters dedicated to descriptive chemistry and boxed essays describing specific applications of chemistry; however, many meaningful and relevant examples of descriptive and applied chemistry are included in the core chapters in the form of end-of-chapter problems.

What's New in This Edition?

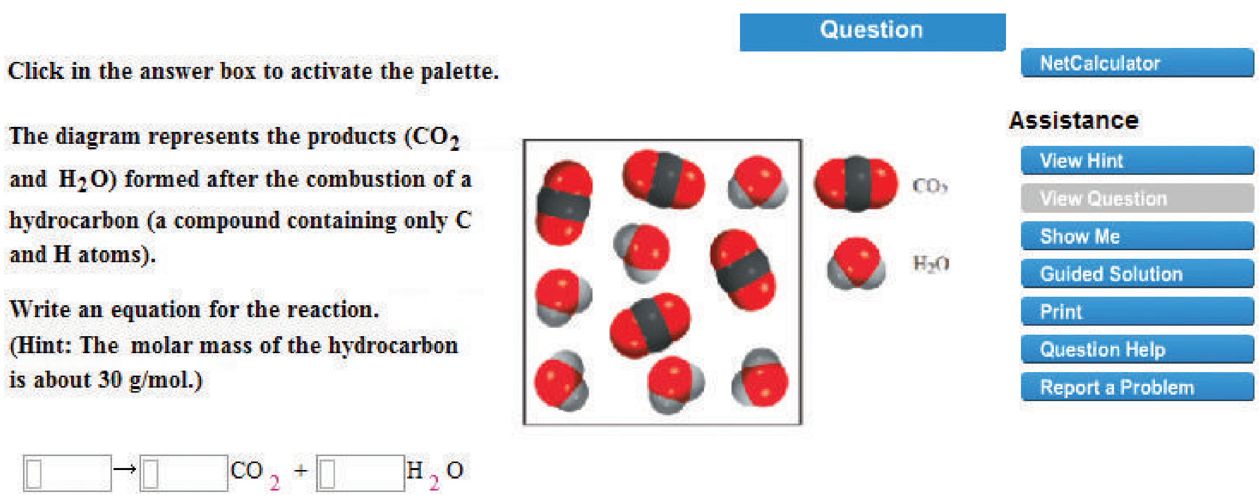
Kenneth Goldsby, Florida State University, has joined Raymond Chang as an author on the seventh edition of *General Chemistry*. Ken's background in inorganic chemistry has added insight into content and problems, and his extensive work with undergraduate students, both in the classroom and in the laboratory, reinforces Raymond's long tradition of understanding and respecting the student's view of the material as well as that of the instructor's.

Many new **End-of-Chapter Problems** have been added to this edition of *General Chemistry*, with an emphasis on interpreting graphs and solving problems based on visual information. 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.

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. In particular, numerous problems are based on descriptive and applied chemistry that would be found in boxed essays and the later chapters of a longer text; see Problems 1.72, 4.115, 6.108, 8.108, 11.73, 13.105, and 19.127. These examples and problems show students the real world of chemistry and applications to everyday life situations.

New is the creation and versatility of our **Connect Chemistry** system. McGraw-Hill has initiated a rigorous process to ensure high-quality electronic homework. Through careful observation of real students and active instructors, we have advanced online homework to an online learning and engagement environment. The goal of Connect is to usher in a new era of meaningful online learning that balances the conceptual and quantitative problem solving aspect of this most vital discipline.

McGraw-Hill is offering students and instructors an enhanced digital homework experience using **Connect**



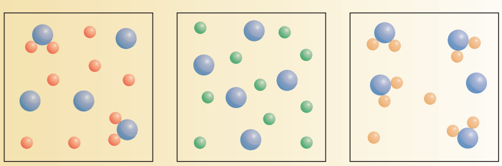
Chemistry. Each problem within Connect Chemistry carries the text problem-solving methodology and is tailored with specific hints, as well as answer-specific feedback for common incorrect answers. Each question has been accuracy checked by two or more chemistry professors. Several rounds of editorial and chemical accuracy checking, in addition to numerous instructor and student tests of all problems, ensure the accuracy of all content.

In addition to the specific hints and feedback provided for all questions, many questions allow students a chemical drawing experience that can be assessed directly within their online homework. Connect Chemistry utilizes ChemDraw, which is widely considered the "gold standard" of scientific drawing programs and the cornerstone application for scientists who draw and annotate molecules, reactions, and pathways. This collaboration of Connect and ChemDraw features an easy-to-use, intuitive, and comprehensive course management and homework system with professional-grade drawing capabilities.

New **Review of Concepts** have been added to many chapters. Review of Concepts are quick conceptual exercises spread throughout the chapters to enable the student to gauge his or her understanding of the concept just presented. The answers to the Review of Concepts are available in the Problem-Solving Workbook and on the companion website in Connect.

Review of Concepts

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



(a) (b) (c)

The entire text has been revised to improve clarity and readability, hallmark characteristics of *General Chemistry*. New and substantial revisions to **chapters and sections** include

- Chapter 3—summary of solving stoichiometry problems based on the mole method.
- Chapter 4—new Example 4.4 on writing molecular, ionic, and net-ionic equations addressing common misconceptions for diprotic and triprotic acids.
- Chapter 9—Example 9.11 provides insight into drawing Lewis structures for compounds containing elements in the third period and beyond, and addresses the controversies in drawing these structures.
- Chapter 19—discussion of the increasingly important lithium-ion battery has been updated, including a new figure highlighting the role of graphene in these systems.
- Chapter 21—up-to-date information on the nuclear power plant accident in Fukushima, Japan, and its implications for the nuclear power industry.

Visualization

Graphs and Flow Charts are important in science. In *General Chemistry*, flow charts illustrate a conceptual thought process or an approach to solving a problem. A significant number of Problems and Review of Concepts, many new to this edition, include graphical data; for example, see the Review of Concepts on page 215 and Problems 4.118, 5.120, 13.113, 17.73, and 21.77.

Study Aids

Setting the Stage

Each chapter starts with the chapter outline and a list of the essential concepts in the chapter.

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

Essential Concepts summarizes the main topics to be presented in the chapter.

Tools to Use for Studying

Study aids are abundant in *General Chemistry*, enabling students to reinforce the comprehension of chemical concepts and learn problem-solving skills.

Worked Examples, along with the accompanying Practice Exercise, is a very important tool for learning and mastering chemistry. The problem-solving steps guide the student through the critical thinking necessary for succeeding in chemistry. Using sketches helps the student understand the inner workings of a problem. 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. Answers to the Review of Concepts can be found in the Problem-Solving Workbook and online in the accompanying Connect Chemistry companion website.

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.

Testing Your Knowledge

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.

Real-Life Relevance

Interesting examples of how chemistry applies to life, both around the home and "on the job," are used throughout the text. Analogies based on common experiences such as banking (Chapter 6) and driving (Chapter 14) are used to help foster understanding of abstract chemical concepts. End-of-chapter problems ask students to apply the concepts presented in the text to answer questions drawn from common experiences, including: Why do swimming coaches sometimes place a drop of alcohol in a swimmer's ear? How does one determine if it is "safe" to open a carbonated soft drink bottle before removing the cap?

Enhanced Support for Faculty and Students

To the Instructor:

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McGraw-Hill Connect® is a Web-based, interactive assignment and assessment platform that incorporates cognitive science principles to customize the learning process. The chemical drawing tool found within Connect Chemistry is PerkinElmer ChemDraw, which is widely considered the "gold standard" of scientific drawing programs and the cornerstone application for scientists who draw and annotate molecules, reactions, and pathways. This combination of Connect and ChemDraw features an easy-to-use, intuitive and comprehensive course management and homework system with professional-grade drawing capabilities.

End-of-chapter problems from this textbook are available in Connect Chemistry for instructors to build assignments that are automatically graded and tracked through reports that export easily to Excel. Instructors can edit existing problems and write entirely new problems; track individual student performance—by problem, assignment, concepts, or in relation to the class overall—with automatic grading; provide instant feedback to students; and store detailed grade reports securely online. Grade reports can be easily integrated with learning management systems such as WebCT and Blackboard. Single sign-on integration is available with Blackboard course management systems. Within Connect, instructors can also create and share materials with colleagues. Ask your McGraw-Hill representative for more information, and then check it out at www.mcgrawhillconnect.com/chemistry.

With ConnectPlus, if you or your students are ready for an alternative version of the traditional textbook, McGraw-Hill has your solution. E-books from McGraw-Hill are smart, interactive, searchable, and portable. Included is a powerful suite of built-in tools that allow detailed searching, highlighting, note taking, and instructor-to-student note sharing. In addition, the media-rich E-book for ***Chemistry*** integrates relevant animations and videos into the textbook content for a true multimedia learning experience.

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McGraw-Hill LearnSmart™ This adaptive diagnostic learning system, based on artificial intelligence, constantly assesses the student's knowledge of the course material. As the student works within the system, LearnSmart develops a personal learning path adapted to what the student has actively learned and retained. This innovative study tool also has features to allow the instructor to see exactly what the student has accomplished, with a built-in assessment tool for graded assignments. LearnSmart for general chemistry can be accessed by going to www.mcgrawhillconnect.com/chemistry.

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- **Art** Full-color art digital files of all illustrations in the book.
- **Photos** 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 in PowerPoint for each chapter of the text.
- **PowerPoint Slides** All illustrations, photos, and tables are preinserted by chapter into blank PowerPoint slides.

An instructor can access the Presentation Center through the *General Chemistry*, Seventh Edition, companion website or through the Library Tab within Instructor version of Connect.

Computerized Test Bank Online

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

Instructor's Solution Manual is written by Raymond Chang and Ken Goldsby. The solutions to all of the end-of-chapter problems are given in the manual. This manual is 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, reference to background concepts in earlier chapters, and teaching tips. This manual is online in the text's Connect Library tab.

To the Student:

Students can order supplemental study materials by contacting their campus bookstore, calling 1-800-262-4729, or online at www.shopmcgraw-hill.com.

Designed to help students maximize their learning experience in chemistry, we offer the following options to students:

LearnSmart

McGraw-Hill LearnSmart™ is an adaptive diagnostic learning system, based on artificial intelligence, constantly assessing the students knowledge of the course material. As the students work within the system, LearnSmart develops a personal learning path adapted to what they have actively learned and retained.

Animations are available on the Web through the *General Chemistry*, Seventh Edition, companion website or through Connect. The animations are also formatted for digital devices.

Problem-Solving Workbook with Selected Solutions is a valuable resource containing material to help the student practice problem-solving skills. It also contains the detailed solutions and explanations for the even-numbered problems for each chapter.

Acknowledgements

We would like to thank the following reviewers and symposium participants, whose comments were of great help to us in preparing this revision:

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We have benefited much from discussions with our colleagues at Williams College and Florida State, and from correspondence with many instructors here and abroad.

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—Raymond Chang and Ken Goldsby

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 seventh edition tools for *General 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.
- Connect Chemistry houses an extraordinary amount of resources. Go to www.mhhe.com/chang and click on the appropriate cover to explore animations, download content to your Media Player, and do your homework electronically and more.
- 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 homework problems. For additional practice, you can turn to similar homework 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.

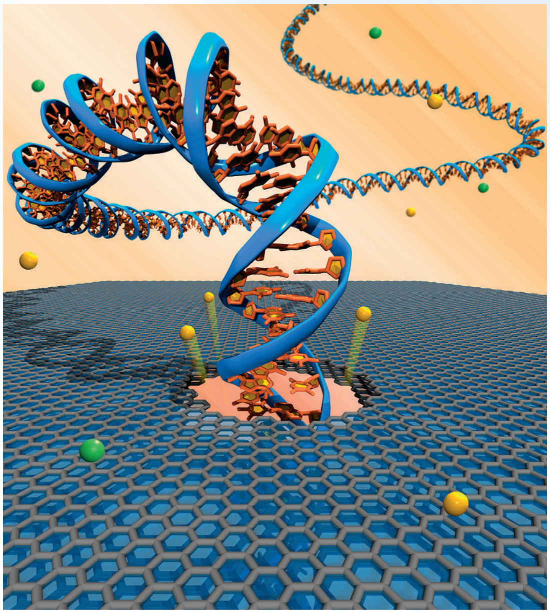
If you follow these suggestions and stay up-to-date with your assignments, you should find that chemistry is challenging, but less difficult and much more interesting than you expected.

—Raymond Chang and Ken Goldsby

CHAPTER 1

Introduction

Chapter Outline



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.

Essential Concepts

The Study of Chemistry Chemistry is the study of the properties of matter and the changes it undergoes. Elements and compounds are substances that take part in chemical transformation.

Physical and Chemical Properties 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.

Measurements and Units Chemistry is a quantitative science and requires measurements. The measured quantities (for example, mass, volume, density, and temperature) usually have units associated with them. The units used in chemistry are based on the international system (SI) of units.

Handling Numbers Scientific notation is used to express large and small numbers, and each number in a measurement must indicate the meaningful digits, called significant figures.

Doing Chemical Calculations A simple and effective way to perform chemical calculations is dimensional analysis. In this procedure, an equation is set up in such a way that all the units cancel except the ones for the final answer.

1.1 The Study of Chemistry

Whether or not this is your first course in chemistry, you undoubtedly have some preconceived ideas about the nature of this science and about what chemists do. Most likely, you think chemistry is practiced in a laboratory by someone in a white coat who studies things in test tubes. This description is fine, up to a point. Chemistry is largely an experimental science, and a great deal of knowledge comes from laboratory research. In addition, however, today's chemists may use a computer to study the microscopic structure and chemical properties of substances or employ sophisticated electronic equipment to analyze pollutants from auto emissions or toxic substances in the soil. Many frontiers in biology and medicine are currently being explored at the level of atoms and molecules—the structural units on which the study of chemistry is based. Chemists participate in the development of new drugs and in agricultural research. What's more, they are seeking solutions to the problem of environmental pollution along with replacements for energy sources. And most industries, whatever their products, have a basis in chemistry. For example, chemists developed the polymers (very large molecules) that manufacturers use to make a wide variety of goods, including clothing, cooking utensils, artificial organs, and toys. Indeed, because of its diverse applications, chemistry is often called the "central science."

How to Study Chemistry

Compared with other subjects, chemistry is commonly perceived 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. 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 perhaps even pleasurably. Listed here are some suggestions to help you form good study habits and master the material:

- Attend classes regularly and take careful notes.
- If possible, always review the topics you learned in class the *same* day the topics 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 for you to explain a concept to a classmate or some other person.
- Do not hesitate to ask your instructor or your teaching assistant for help.

You will find that chemistry is much more than numbers, formulas, and abstract theories. It is a logical discipline brimming with interesting ideas and applications.

1.2 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 follow roughly the same procedure in carrying out their investigations. The first step is carefully defining 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 these examples, 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 communications with other chemists. Figure 1.1 summarizes the main steps of the research process.

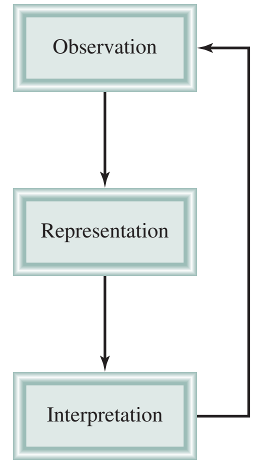


Figure 1.1 *The three levels of studying chemistry and their relationships. Observation deals with events in the macroscopic world; atoms and molecules constitute the microscopic world. Representation is a scientific shorthand for describing an experiment in symbols and chemical equations. Chemists use their knowledge of atoms and molecules to explain an observed phenomenon.*

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,*** or *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.

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 increase in the mass or in the acceleration of an object always increases the object's force proportionally, and a decrease in mass or acceleration always decreases 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.

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 may take drastically different approaches. They may be led in vastly different directions. Scientists are, after all, human beings, and their modes of thinking and working are very much influenced by their backgrounds, 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 spent years working on projects that ultimately led to a dead end. Many 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 formation 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.

1.3 Classifications of Matter

化学

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

Matter is *anything that occupies space and has mass,* and ***chemistry*** is *the study of matter and the changes it undergoes.* All matter, at least in principle, can exist in three states: solid, liquid, and gas. Solids are rigid objects with definite shapes. Liquids are less rigid than solids and are fluid—they are able to flow and assume the shape of their containers. Like liquids, gases are fluid, but unlike liquids, they can expand indefinitely.

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 1.2 shows the three states of water. Note that the properties of water are unique

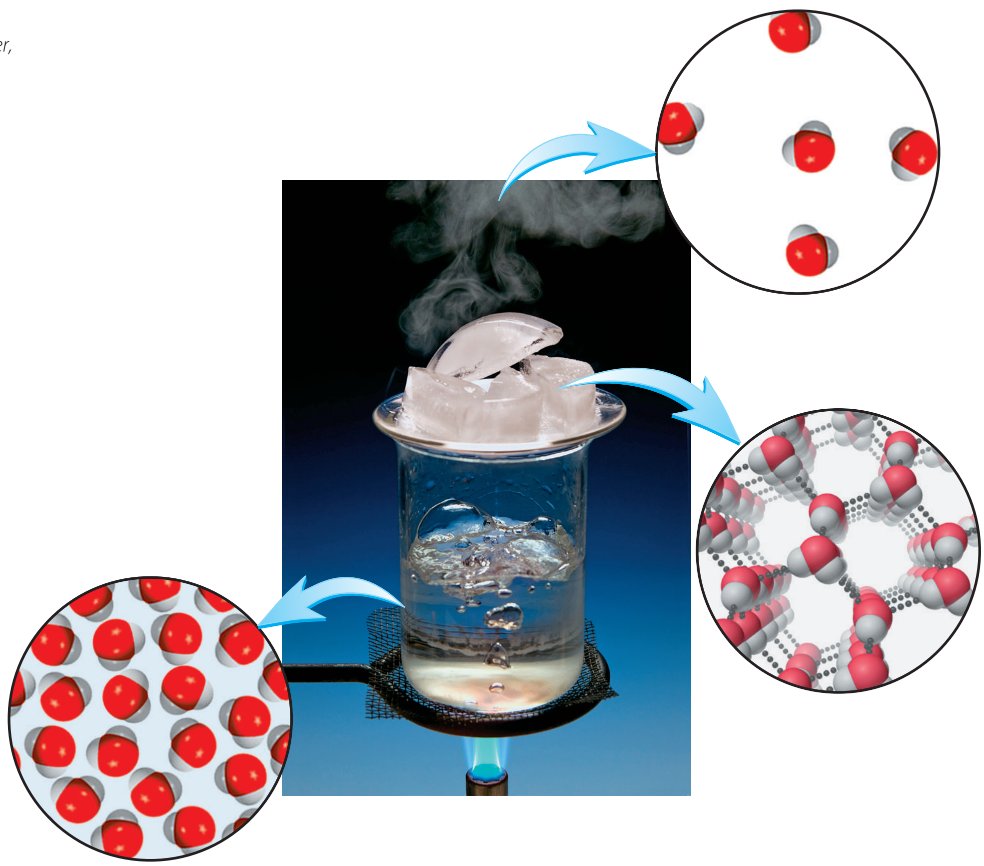


Figure 1.2 *The three states of matter for water: solid ice, liquid water, and gaseous steam.*

(a) (b)

Figure 1.3 *(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.*

among common substances in that the molecules in the liquid state are more closely packed than those in the solid state.

Substances and Mixtures

A ***substance*** is *matter that has a definite or constant composition and distinct properties.* Examples are water, silver, ethanol, table salt (sodium chloride), and carbon dioxide. Substances differ from one another in composition and can be identified by their appearance, smell, taste, and other properties. At present, over 66 million substances are known, and the list is growing rapidly.

A ***mixture*** is *a combination of two or more substances in which the substances retain their distinct identities.* Some 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, *the composition of the mixture,* after sufficient stirring, *is the same throughout the solution.* This solution is a ***homogeneous mixture.*** If sand is mixed with iron filings, however, the sand grains and the iron filings remain visible and separate (Figure 1.3). This type of mixture, in which *the composition is not uniform,* is called a ***heterogeneous mixture.*** Adding oil to water creates another heterogeneous mixture because the liquid does not have a constant composition.

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 water 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 1.3(b)]. After separation, the components of the mixture will have the same composition and properties as they did to start with.

Elements and Compounds

A substance can be either an element or a compound. An ***element*** is *a substance that cannot be separated into simpler substances by chemical means.* At present, 118 elements have been positively identified. (See the list inside the front cover of this book.)

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
Bromine	Br	Iodine	I	Potassium	K
Calcium	Ca	Iron	Fe	Silicon	Si
Carbon	C	Lead	Pb	Silver	Ag
Chlorine	Cl	Magnesium	Mg	Sodium	Na
Chromium	Cr	Mercury	Hg	Sulfur	S
Cobalt	Co	Nickel	Ni	Tin	Sn
Copper	Cu	Nitrogen	N	Zinc	Zn

Chemists use alphabetical symbols to represent the names of the elements. The first letter of the symbol for an element is *always* capitalized, but the second letter is *never* capitalized. For example, Co is the symbol for the element cobalt, whereas CO is the formula for carbon monoxide, which is made up of the elements carbon and oxygen. Table 1.1 shows some of the more common elements. The symbols for some elements are derived from their Latin names—for example, Au from *aurum* (gold), Fe from *ferrum* (iron), and Na from *natrium* (sodium)—although most of them are abbreviated forms of their English names.

Figure 1.4 shows the most abundant elements in Earth's crust and in the human body. As you can see, only five elements (oxygen, silicon, aluminum, iron, and calcium) comprise over 90 percent of Earth's crust. Of these five elements, only oxygen is among the most abundant elements in living systems.

Most elements can interact with one or more other elements to form compounds. We define a ***compound*** as *a substance composed of two or more elements chemically united in fixed proportions.* Hydrogen gas, for example, burns in oxygen gas to form water, a compound whose properties are distinctly different from those of the starting materials. Water is made up of two parts of hydrogen and one part of oxygen. This composition does not change, regardless of whether the water comes from a faucet in the United States, the Yangtze River in China, or the ice caps on Mars. 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.

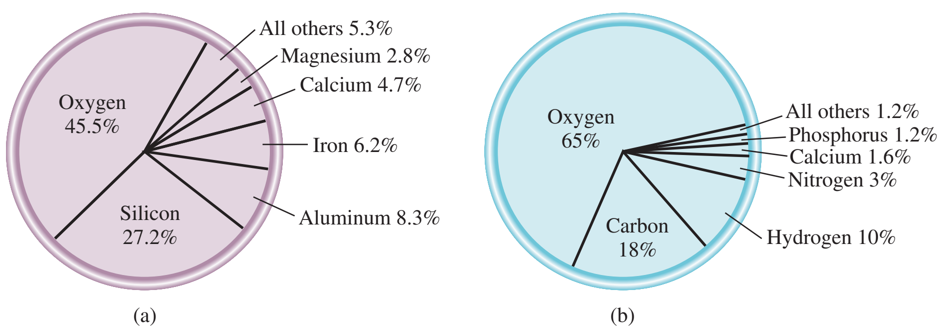


Figure 1.4 *(a) Natural abundance of the elements in percent by mass. For example, oxygen's abundance is 45.5 percent. This means that in a 100-g sample of Earth's crust there are, on the average, 45.5 g of the element oxygen. (b) Abundance of elements in the human body in percent by mass.*

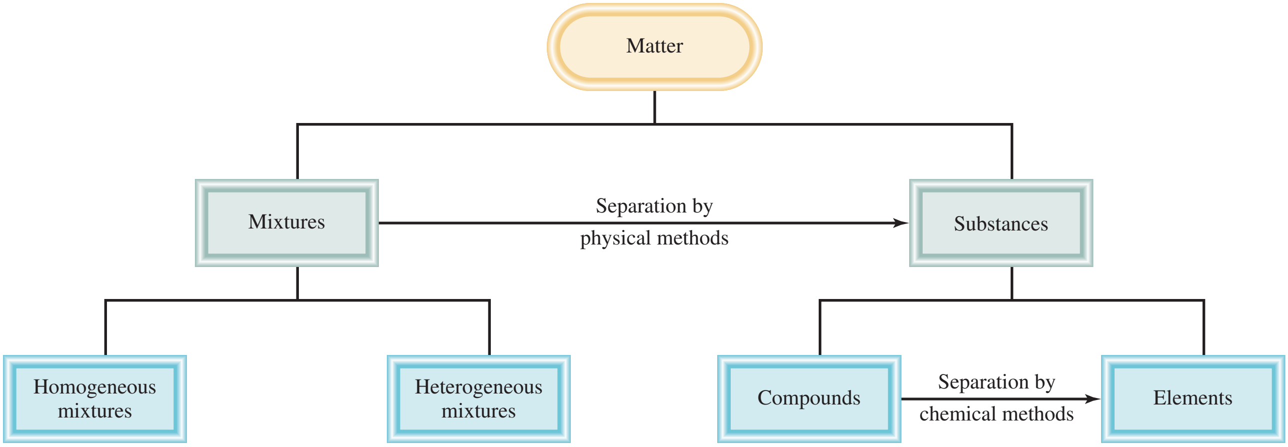
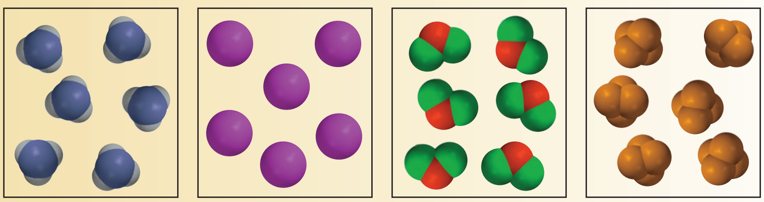


Figure 1.5 *Classification of matter.*

Review of Concepts

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



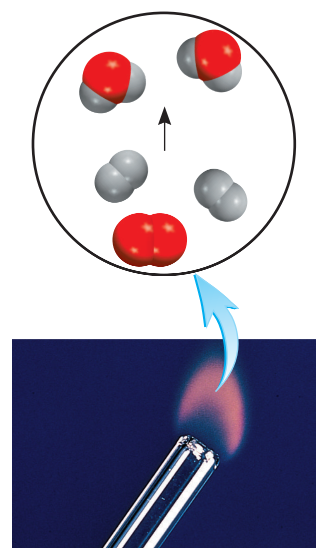
(a) (b) (c) (d)

1.4 Physical and Chemical Properties of Matter

Substances are identified by their properties as well as by their composition. Color, melting point, boiling point, and density 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 and 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 substances, hydrogen and oxygen gas, will have vanished and a chemically different substance—water—will have taken their place. We *cannot* recover hydrogen and oxygen from water by a physical change such as boiling or freezing.

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 reactions that alter not only their physical appearance but their chemical makeup as well. When eaten, the egg is changed again, by substances in the body called *enzymes.* This digestive action



Hydrogen burning in air to form water.

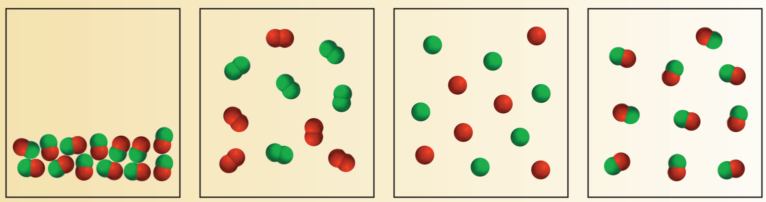
is another example of a chemical change. What happens during such a process depends on the chemical properties of the specific enzymes and of the food involved.

All measurable properties of matter fall into two categories: extensive properties and intensive properties. The measured value of an ***extensive property*** *depends on how much matter is being considered.* Mass, length, and volume are extensive properties. More matter means more mass. Values of the same extensive property can be added together. For example, two copper pennies have a combined mass that is the sum of the masses of each penny, and the total volume occupied by the water in two beakers is the sum of the volumes of the water in each of the beakers.

The measured value of an ***intensive property*** *does not depend on the amount of matter being considered.* Temperature is an intensive property. 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 amount of water will be the same as it was in two separate beakers. Unlike mass and volume, temperature and other intensive properties such as melting point, boiling point, and density 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)

1.5 Measurement

The study of chemistry depends heavily on measurement. For instance, chemists use measurements to compare the properties of different substances and to assess changes resulting from an experiment. A number of common devices enable us to make simple measurements of a substance's properties: The meterstick measures length; the buret, the pipet, the graduated cylinder, and the volumetric flask measure volume (Figure 1.6); 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. In science, 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

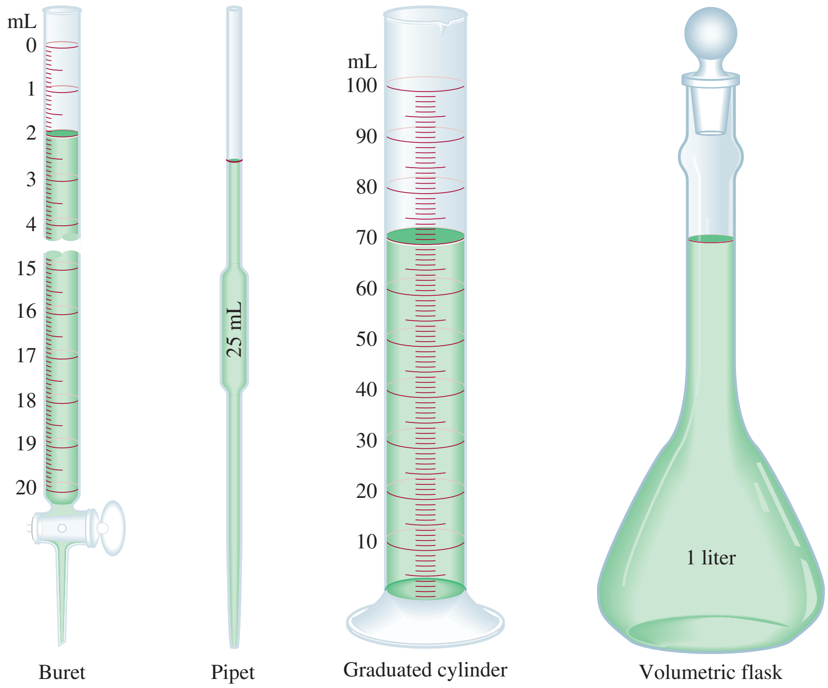


Figure 1.6 *Some common measuring devices found in a chemistry laboratory. These devices are not drawn to scale relative to one another. We will discuss the use of these measuring devices in Chapter 4.*

system called the ***International System of Units*** (abbreviated *SI,* from the French *S*ystem *I*nternational d'Unites). Table 1.2 shows the seven SI base units. All other SI 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 use both metric and SI units in this book.

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

Mass and Weight

Mass is *a measure of the quantity of matter in an object.* The terms "mass" and "weight" are often used interchangeably, although, strictly speaking, they refer to different quantities. In scientific terms, ***weight*** 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

Table 1.2 SI Base Units

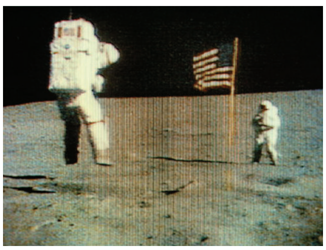
Base Quantity	Name of Unit	Symbol
Length	meter	m
Mass	kilogram	kg
Time	second	s
Electrical current	ampere	A
Temperature	kelvin	K
Amount of substance	mole	mol
Luminous intensity	candela	cd

Note that a metric prefix simply represents a number:

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

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



An astronaut jumping on the surface of the moon.

what it does on Earth, because of the smaller mass of the moon. This is why astronauts were able to jump about rather freely on the moon's surface despite their bulky suits and equipment. The mass of an object can be determined readily with a balance, and this process, oddly, is called weighing.

The SI base unit of mass is the *kilogram* (kg), but in chemistry the smaller gram (g) is more convenient:

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

Volume

Volume is *length (m) cubed,* so its SI-derived unit 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, non-SI unit of volume is the liter (L). A ***liter*** is *the volume occupied by one cubic decimeter.* Chemists generally use L and mL for liquid volume. One liter is equal to 1000 milliliters (mL) or 1000 cubic centimeters:

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

and one milliliter is equal to one cubic centimeter:

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

Figure 1.7 compares the relative sizes of two volumes.

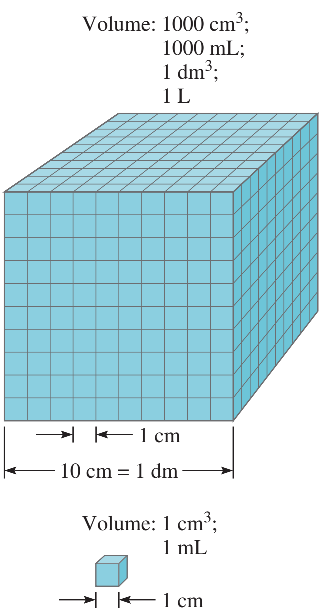


Figure 1.7 *Comparison of two volumes, 1 mL and 1000 mL.*

Density

Density is *the mass of an object divided by its volume:*

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

or

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

where d, m, and V denote density, mass, and volume, respectively. Note that density is an intensive property that does not depend on the quantity of mass present. The reason is that V increases as m does, so the ratio of the two quantities always remains the same for a given material.

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. Table 1.4 lists the densities of several substances.

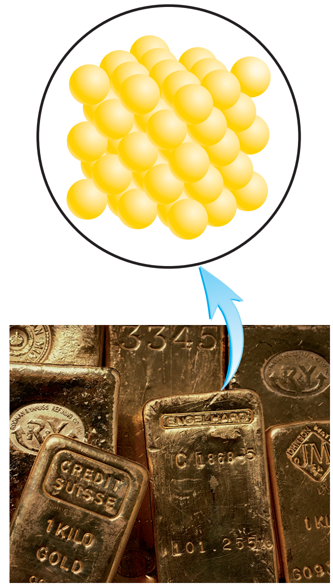
Example 1.1

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 257 g has a volume of 13.3 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

$$\begin{aligned} d &= \frac{m}{V} \\ &= \frac{257\ \text{g}}{13.3\ \text{cm}^3} \\ &= \boxed{19.3\ \text{g/cm}^3} \end{aligned}$$

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: 1.17, 1.18.

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

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 1.8 compares the three temperature scales.

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

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

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

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

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.

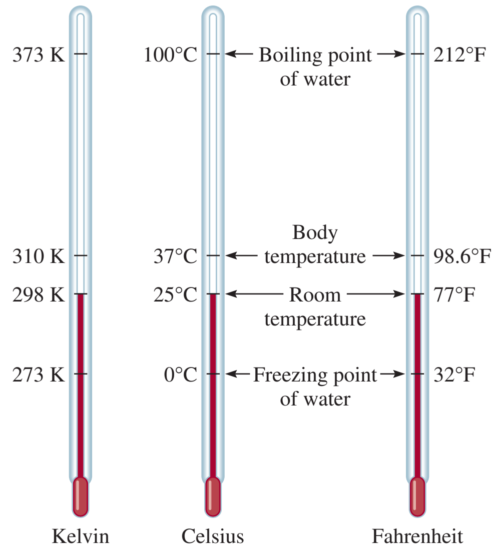
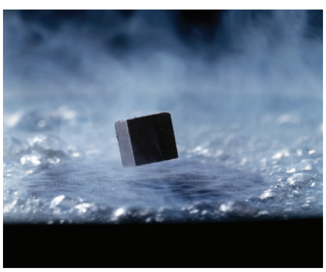


Figure 1.8 *Comparison of the three temperature scales: Celsius, Fahrenheit, and the absolute (Kelvin) scales. Note that there are 100 divisions, or 100 degrees, between the freezing point and the boiling point of water on the Celsius scale, and there are 180 divisions, or 180 degrees, between the same two temperature limits on the Fahrenheit scale. The Celsius scale was formerly called the centigrade scale. Note that the Kelvin scale does not have the degree sign. Also, temperature expressed in kelvins can never be negative.*

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:

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



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

Example 1.2

(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

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

(b) Here we have

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

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

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

Similar problems: 1.19, 1.20.

(Continued)

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. The decrease in density is the result of which of the following?

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

1.6 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

$$0.00000000000000000000000166 \text{ g}$$

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, then 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 $n = -6$.

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

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.

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

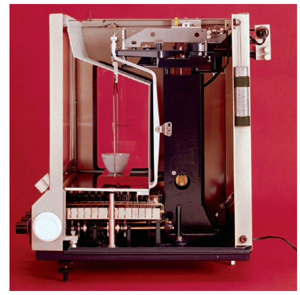


Figure 1.9 *A single-pan balance.*

Except when all the numbers involved are integers (for example, in counting the number of students in a class), obtaining the exact value of the quantity under investigation is often impossible. 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. 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 1.9 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 1.3

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 (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: 1.27, 1.28.

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:

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

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

   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 (such as 1 ft = 12 in, where 12 is an exact number) or by counting numbers of objects can be considered to have an infinite number of significant figures.

Example 1.4

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

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

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

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

(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 $\boxed{5.50 \times 10^3\ \text{cm.}}$

Similar problems: 1.29, 1.30.

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:

First step: $A \times B = C$

Second step: $C \times D = E$

Let's suppose that A = 3.66, B = 8.45, and D = 2.11. Depending on whether we round off C to three (Method 1) or four (Method 2) significant figures, we obtain a different number for E:

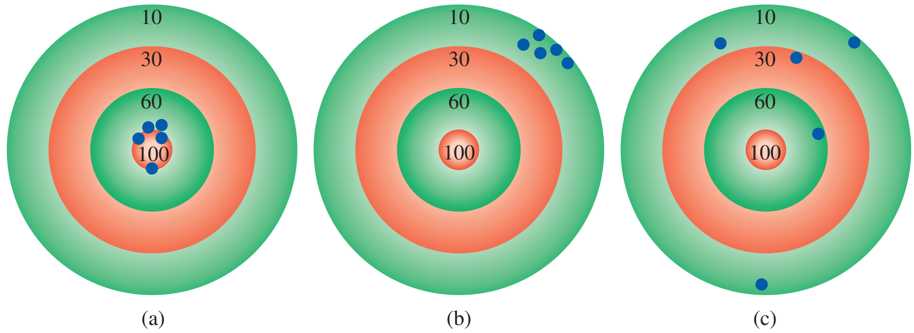
Method 1	Method 2
$3.66 \times 8.45 = 30.9$	$3.66 \times 8.45 = 30.93$
$30.9 \times 2.11 = 65.2$	$30.93 \times 2.11 = 65.3$

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*

Figure 1.10 *The distribution of darts on a dartboard 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 blue dots show the positions of the darts.*



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 1.10).

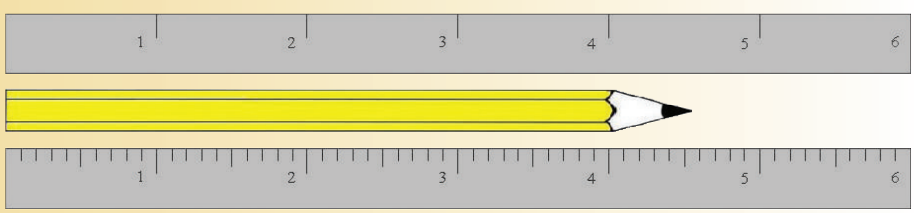
The difference between accuracy and precision is a subtle but important one. Suppose, for example, that three students are asked to determine the mass of a piece of copper wire. The results of two successive weighings by each student are

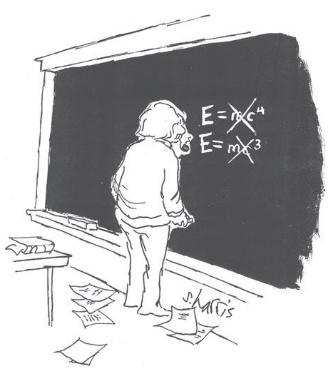
	Student A	**Student B**	**Student C**
	1.964 g	1.972 g	2.000 g
	1.978 g	1.968 g	2.002 g
Average value	1.971 g	1.970 g	2.001 g

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.





Dimensional analysis might also have led Einstein to his famous mass-energy equation $E = mc^2$.

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1.7 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 multiple 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 $\times$ 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:

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

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 violinist 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 8) 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 way to quickly check the answer is to make a "ballpark" estimate. The idea here is to round off the numbers in the calculation in such a way that we simplify the arithmetic. This approach is sometimes called the "back-of-the-envelope calculation" because it can be done easily without using a calculator. The answer you get will not be exact, but it will be close to the correct one.

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.

Example 1.5

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 conversion is

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

(Continued)

Using the following conversion factors:

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

we obtain the answer in one step:

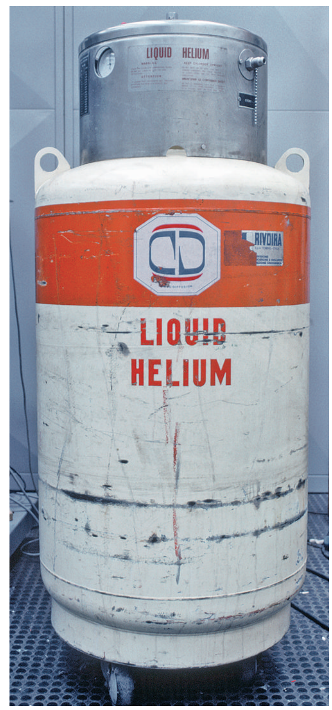
$$? \text{ mg} = 0.0833 \cancel{\text{lb}} \times \frac{453.6 \cancel{\text{g}}}{1 \cancel{\text{lb}}} \times \frac{1 \text{ mg}}{1 \times 10^{-3} \cancel{\text{g}}} = \boxed{3.78 \times 10^4 \text{ mg}}$$

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: 1.37(a).

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

As Examples 1.6 and 1.7 illustrate, conversion factors can be squared or cubed in dimensional analysis.



A cryogenic storage tank for liquid helium.

Example 1.6

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

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

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: 1.38(g).

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

Example 1.7

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)


Liquid nitrogen.

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 $cm^3 \longrightarrow m^3$. Recall that 1 kg = 1000 g and 1 cm = 1×10^{-2} m.

Solution In Example 1.6 we saw that 1 cm^3 = 1×10^{-6} 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,

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

Check Because 1 m^3 = 1×10^6 cm^3, we would expect much more mass in 1 m^3 than in 1 cm^3. Therefore, the answer is reasonable.

Similar problem: 1.39.

Practice Exercise The density of the lightest metal, lithium (Li), is 5.34×10^2 kg/m^3. Convert the density to g/cm^3.

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.)

Key Equations

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

Summary of Facts and Concepts

1. 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.
2. Chemists study matter and the substances of which it is composed. All substances, in principle, can exist in three states: solid, liquid, and gas. The interconversion between these states can be effected by a change in temperature.
3. The simplest substances in chemistry are elements. Compounds are formed by the combination of atoms of different elements. Substances have both unique physical properties that can be observed without changing the identity of the substances and unique chemical properties that, when they are demonstrated, do change the identity of the substances.
4. SI units are used to express physical quantities in all sciences, including chemistry. 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. Most measured quantities are inexact to some extent. The number of significant figures indicates the exactness of the measurement.

5. In the dimensional analysis method of solving problems the units are multiplied together, divided into each other, or canceled like algebraic quantities. Obtaining the correct units for the final answer ensures that the calculation has been carried out properly.

Key Words

Accuracy, p. 17
Chemical property, p. 7
Chemistry, p. 4
Compound, p. 6
Density, p. 10
Element, p. 5
Extensive property, p. 8
Heterogeneous mixture, p. 5
Homogeneous mixture, p. 5
Hypothesis, p. 3
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Physical property, p. 7
Precision, p. 18
Qualitative, p. 3
Quantitative, p. 3
Scientific method, p. 2
Significant figures, p. 15
Substance, p. 5
Theory, p. 3
Volume, p. 10
Weight, p. 9

Questions and Problems

Basic Definitions

Review Questions

1.1 Define these terms: (a) matter, (b) mass, (c) weight, (d) substance, (e) mixture.

1.2 Which of these statements is scientifically correct?
"The mass of the student is 56 kg."
"The weight of the student is 56 kg."

1.3 Give an example of a homogeneous mixture and an example of a heterogeneous mixture.

1.4 What is the difference between a physical property and a chemical property?

1.5 Give an example of an intensive property and an example of an extensive property.

1.6 Define these terms: (a) element, (b) compound.

Problems

1.7 Do these 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.

1.8 Does each of these 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.

1.9 Which of these properties are intensive and which are extensive? (a) length, (b) volume, (c) temperature, (d) mass.

1.10 Which of these properties are intensive and which are extensive? (a) area, (b) color, (c) density.

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

1.12 Classify each of these as an element or a compound: (a) sodium chloride (table salt), (b) helium, (c) alcohol, (d) platinum.

Units

Review Questions

1.13 Give the SI units for expressing these: (a) length, (b) area, (c) volume, (d) mass, (e) time, (f) force, (g) energy, (h) temperature.

1.14 Write the numbers for these prefixes: (a) mega-, (b) kilo-, (c) deci-, (d) centi-, (e) milli-, (f) micro-, (g) nano-, (h) pico-.

1.15 Define density. What units do chemists normally use for density? Is density an intensive or extensive property?

1.16 Write the equations for converting degrees Celsius to degrees Fahrenheit and degrees Fahrenheit to degrees Celsius.

Problems

1.17 A lead sphere has a mass of 1.20×10^4 g, and its volume is 1.05×10^3 cm^3. Calculate the density of lead.

1.18 Mercury is the only metal that is a liquid at room temperature. Its density is 13.6 g/mL. How many grams of mercury will occupy a volume of 95.8 mL?

1.19 (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?

1.20 (a) Convert the following temperatures to kelvin: (i) 113°C, the melting point of sulfur, (ii) 37°C, the normal body temperature, (iii) 357°C, the boiling point of mercury. (b) Convert the following temperatures to degrees Celsius: (i) 77 K, the boiling point of liquid nitrogen, (ii) 4.2 K, the boiling point of liquid helium, (iii) 601 K, the melting point of lead.

Scientific Notation

Problems

1.21 Express these numbers in scientific notation: (a) 0.000000027, (b) 356, (c) 0.096.

1.22 Express these numbers in scientific notation: (a) 0.749, (b) 802.6, (c) 0.000000621.

1.23 Convert these numbers to nonscientific notation: (a) 1.52×10^4, (b) 7.78×10^{-8}.

1.24 Convert these numbers to nonscientific notation: (a) 3.256×10^{-5}, (b) 6.03×10^6.

1.25 Express the answers to these operations 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) $(1.0 \times 10^4) \times (9.9 \times 10^6)$

1.26 Express the answers to these operations 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) $(3.6 \times 10^{-4}) \times (3.6 \times 10^6)$

Significant Figures

Problems

1.27 What is the number of significant figures in each of these measured quantities? (a) 4867 miles, (b) 56 mL, (c) 60,104 tons, (d) 2900 g.

1.28 What is the number of significant figures in each of these measured quantities? (a) 40.2 g/cm^3, (b) 0.0000003 cm, (c) 70 min, (d) 4.6×10^{19} atoms.

1.29 Carry out these operations as if they were calculations of experimental results, and express each answer in the correct units and 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) $(3 \times 10^4\ \text{g} + 6.827\ \text{g})/(0.043\ \text{cm}^3 - 0.021\ \text{cm}^3)$

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

(a) 7.310 km ÷ 5.70 km

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

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

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

Dimensional Analysis

Problems

1.31 Carry out these conversions: (a) 22.6 m to decimeters, (b) 25.4 mg to kilograms.

1.32 Carry out these conversions: (a) 242 lb to milligrams, (b) 68.3 cm^3 to cubic meters.

1.33 The price of gold on a certain day in 2009 was $932 per troy ounce. How much did 1.00 g of gold cost that day? (1 troy ounce = 31.03 g.)

1.34 Three students (A, B, and C) are asked to determine the volume of a sample of methanol. Each student measures the volume three times with a graduated cylinder. The results in milliliters are A (47.2, 48.2, 47.6); B (46.9, 47.1, 47.2); C (47.8, 47.8, 47.9). The true volume of methanol is 47.0 mL. Which student is the most accurate? Which student is the most precise?

1.35 Three students (X, Y, and Z) are assigned the task of determining the mass of a sample of iron. Each student makes three determinations with a balance. The results in grams are X (61.5, 61.6, 61.4); Y (62.8, 62.2, 62.7); Z (61.9, 62.2, 62.1). The actual mass of the iron is 62.0 g. Which student is the least precise? Which student is the most accurate?

1.36 A slow jogger runs a mile in 13 min. Calculate the speed in (a) in/s, (b) m/min, (c) km/h. (1 mi = 1609 m; 1 in = 2.54 cm.)

1.37 Carry out these conversions: (a) 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.) (b) The current speed limit in some states in the United States is 70 miles per hour. What is the speed limit in kilometers per hour? (c) The speed of light is 3.0×10^{10} cm/s. How many miles does light travel in 1 hour? (d) Lead is a toxic substance. 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?

1.38 Carry out these 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), (b) 32.4 yd to centimeters, (c) 3.0×10^{10} cm/s to ft/s, (d) 47.4°F to degrees Celsius, (e) −273.15°C (the lowest temperature) to degrees Fahrenheit, (f) 71.2 cm^3 to m^3, (g) 7.2 m^3 to liters.

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

1.40 The density of ammonia gas under certain conditions is 0.625 g/L. Calculate its density in g/cm^3.

Additional Problems

1.41 Which of these describe physical 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.

1.42 In 2010 about 132 billion pounds of sulfuric acid were produced in the United States. Convert this quantity to tons.

1.43 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?

1.44 In the determination of 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.

1.45 Calculate the mass of each of these: (a) a sphere of gold of radius 10.0 cm [the volume of a sphere of radius r is $V = (\frac{4}{3})\pi r^3$; 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).

1.46 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.

1.47 This procedure was carried out 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 the 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 cubic centimeters.

1.48 A silver (Ag) object 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.

1.49 The experiment described in Problem 1.48 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?

1.50 How far (in feet) does light travel in 1.00 ns? The speed of light is 3.00×10^8 m/s.

1.51 The medicinal thermometer commonly used in homes can be read to ±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.

1.52 A thermometer gives a reading of 24.2°C ± 0.1°C. Calculate the temperature in degrees Fahrenheit. What is the uncertainty?

1.53 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 detectable in a large aircraft hangar of volume 5.0×10^7 ft^3.

1.54 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.)

1.55 The total volume of seawater is 1.5×10^{21} L. Assume that seawater contains 3.1 percent sodium chloride by mass and that its density is 1.03 g/mL. Calculate the total mass of sodium chloride in kilograms and in tons. (1 ton = 2000 lb; 1 lb = 453.6 g.)

1.56 Magnesium (Mg) is a valuable metal used in alloys, in batteries, and in chemical synthesis. It is obtained mostly from seawater, which contains about 1.3 g of Mg for every kilogram of seawater. Calculate the volume of seawater (in liters) needed to extract 8.0×10^4 tons of Mg, which is roughly the annual production in the United States. (Density of seawater = 1.03 g/mL.)

1.57 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/cm^3). The readings are 860.2 g and 820.2 g, respectively. Given that the density of platinum is 21.45 g/cm^3, what should her conclusion be based on these measurements? (*Hint:* An object suspended in a fluid is buoyed up by the mass of the fluid displaced by the object. Neglect the buoyancy of air.)

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

1.59 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.

1.60 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:

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

where the vertical lines indicate absolute value. Calculate the percent error for these 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.)

1.61 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 three significant figures.

1.62 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 its worth in dollars, assuming that the price of gold is \$930 an ounce. With so much gold out there, why hasn't someone become rich by mining gold from the ocean?

1.63 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.)

1.64 The diameter 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.

1.65 One gallon of gasoline burned 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 mi per gallon.

1.66 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.)

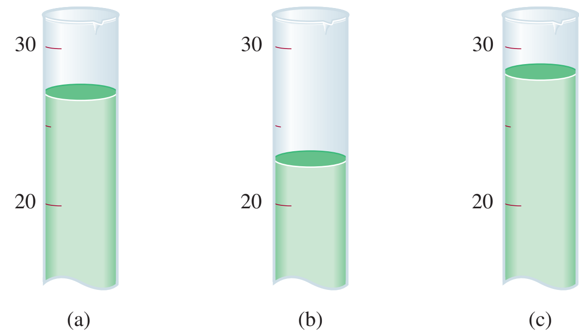
1.67 Chlorine is used to disinfect swimming pools. The accepted concentration for this purpose is 1 ppm chlorine or 1 g of chlorine per million g of water. Calculate the volume of a chlorine solution (in milliliters) a homeowner should add to her swimming pool if the solution contains 6.0 percent chlorine by mass and there are 2×10^4 gallons of water in the pool. (1 gallon = 3.79 L; density of liquids = 1.0 g/mL.)

1.68 Fluoridation is the process of adding fluorine compounds to drinking water to help fight tooth decay. A concentration of 1 ppm of fluorine is sufficient for the purpose. (1 ppm means 1 g of fluorine per 1 million g of water.) The compound normally chosen for fluoridation is sodium fluoride, which is also added to some toothpastes. Calculate the quantity of sodium fluoride in kilograms needed per year for a city of 50,000 people if the daily consumption of water per person is 150 gallons. What percent of the sodium fluoride is "wasted" if each person uses only 6.0 L of water a day for drinking and cooking? (Sodium fluoride is 45.0 percent fluorine by mass. 1 gallon = 3.79 L; 1 year = 365 days; density of water = 1.0 g/mL.)

1.69 In water conservation, chemists spread a thin film of a certain inert material over the surface of water to cut down 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 of 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.)

1.70 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.

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



Special Problems

1.72 Dinosaurs dominated life on Earth for millions of years and then disappeared very suddenly. In the experimentation and data-collecting stage, 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. 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 gradually perished, and then, in turn, meat-eating animals began to starve. Limitation of food sources obviously affects large animals needing great amounts of food more quickly and more severely than small animals. Therefore, the huge dinosaurs vanished because of lack of food.

(a) How does the study of dinosaur extinction illustrate the scientific method?

(b) Suggest two ways to test the hypothesis.

(c) In your opinion, is it justifiable to refer to the asteroid explanation as the theory of dinosaur extinction?

(d) Available evidence suggests that about 20 percent of the asteroid's mass turned to dust and spread uniformly over Earth after eventually 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.)

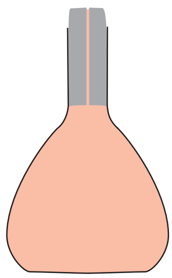
1.73 You are given a liquid. Briefly describe steps you would take to show whether it is a pure substance or a homogeneous mixture.

1.74 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 this collection of coins?

1.75 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 $(4/3)\pi r^3$.]

1.76 Bronze is an alloy made of copper (Cu) and tin (Sn). 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?

1.77 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 20ºC 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 20ºC 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?



1.78 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.

1.79 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.

1.80 (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).

1.81 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?

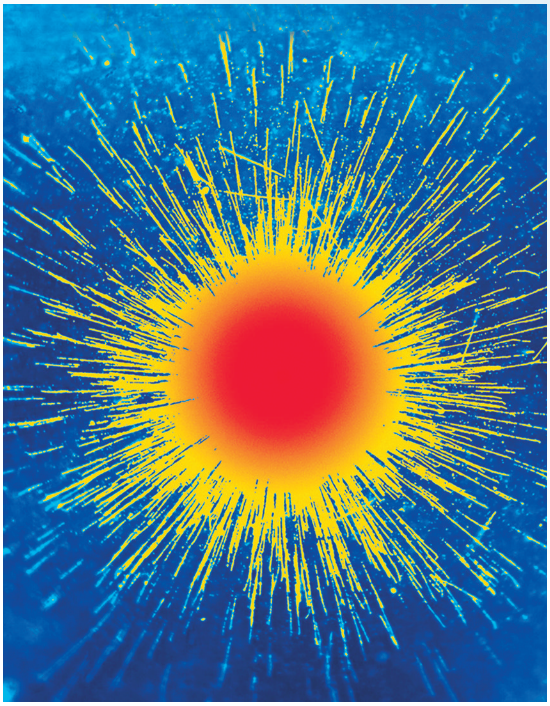
Answers to Practice Exercises

1.1 96.5 g. **1.2** (a) 621.5°F, (b) 78.3°C, (c) −196°C. **1.3** (a) Two, (b) four, (c) three, (d) two, (e) three or two. **1.4** (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. **1.5** 2.36 lb. **1.6** 1.08×10^5 m^3. **1.7** 0.534 g/cm^3.

CHAPTER 2

Atoms, Molecules, and Ions

Chapter Outline



Colored images of the radioactive emission of radium (Ra). Study of radioactivity helped to advance scientists' knowledge about atomic structure.

Essential Concepts

Development of the Atomic Theory The search for the fundamental units of matter began in ancient times. The modern version of atomic theory was laid out by John Dalton, who postulated that elements are made of extremely small particles, called atoms, and that all atoms of a given element are identical, but they are different from atoms of all other elements.

The Structure of the Atom 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, and electrons are spread out about the nucleus at some distance from it.

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 different numbers of neutrons. Mass number is the sum of the number of protons and neutrons in an atom.

The Periodic Table 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.

From Atoms to Molecules and Ions Atoms of most elements interact to form compounds, which are classified as molecules or ionic compounds made of positive (cations) and negative (anions) ions. Chemical formulas tell us the type and number of atoms present in a molecule or compound.

Naming Compounds The names of many inorganic compounds can be deduced from a set of simple rules.

Organic Compounds The simplest type of organic compounds is the hydrocarbons.

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. It was in 1808 that an English scientist and schoolteacher, 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

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 hypotheses 2 and 3.

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 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 the 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

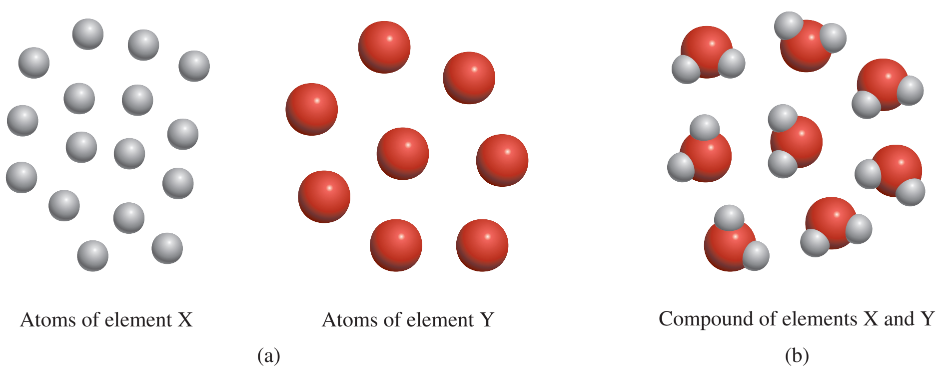


Figure 2.1 *(a) According to Dalton's atomic theory, atoms of the same element are identical, but atoms of one element are different from atoms of other elements. (b) Compounds formed from atoms of elements X and Y. In this case, the ratio of the atoms of element X to the atoms of element Y is 2:1.*

compound is fixed, the ratio of the atoms of these elements in the compound also must be constant.

Dalton's third hypothesis also supports another important law, the ***law of multiple proportions.*** According to this 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: The compounds 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 that one atom of carbon combines with two oxygen atoms 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 because the mass of an element in a compound is proportional to the number of atoms of the element present (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.

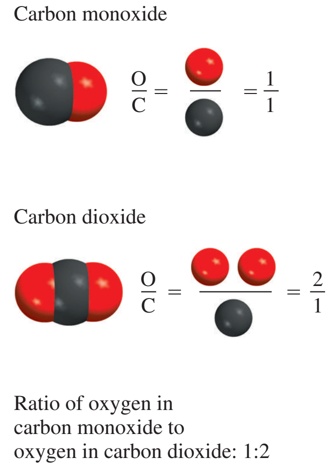


Figure 2.2 *An illustration of the law of multiple proportions.*

Review of Concepts

The atoms of elements A (blue) and B (yellow) 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.

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

Animation
Cathode Ray Tube

[†]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 21), however, changes of mass in chemical reactions are too small to detect. Therefore, for all practical purposes mass is conserved.

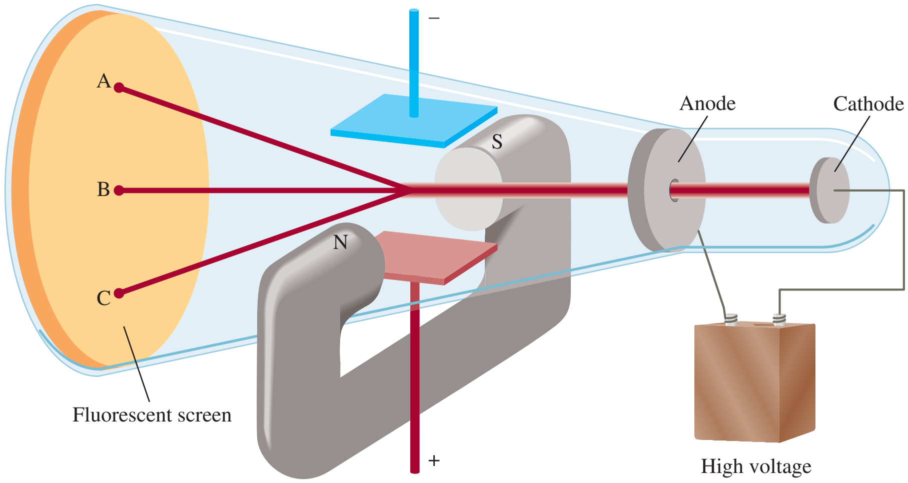


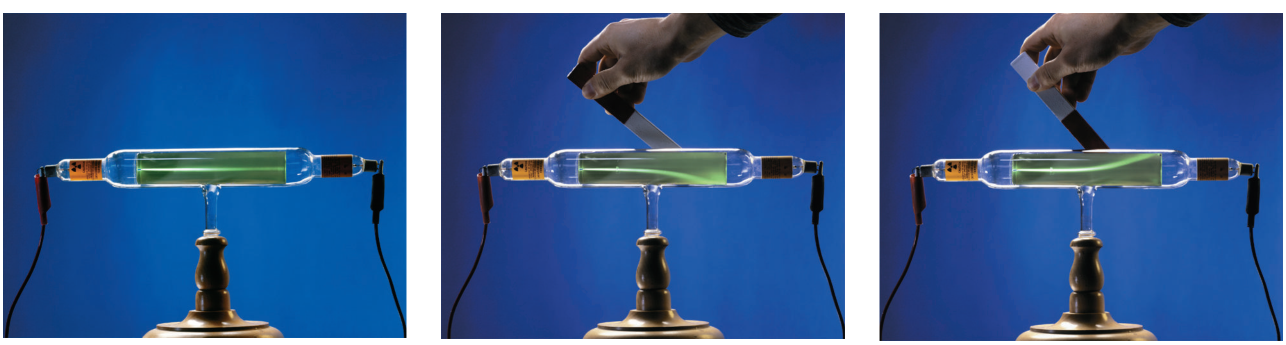
Figure 2.3 *A cathode ray tube with an electric field perpendicular to the direction of the cathode rays and an external magnetic field. The symbols N and S denote the north and south poles of the magnet. The cathode rays will strike the end of the tube at A in the presence of a magnetic field, at C 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.*

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.

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 be studied individually.

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 is -1.76×10^8 C/g, where C stands



(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.*