About the Authors

Spencer L. Seager

Spencer L. Seager is a professor of chemistry at Weber State University, where he served as chemistry department chairman from 1969 until 1993. He teaches general chemistry at the university and is also active in projects to help improve chemistry and other science education in local elementary schools. He received his B.S. degree in chemistry and Ph.D. degree in physical chemistry from the University of Utah. Other interests include making minor home repairs, reading history of science and technology, listening to classical music, and walking for exercise.

Michael R. Slabaugh

Michael R. Slabaugh is a senior fellow at Weber State University, where he teaches the year-long sequence of general chemistry, organic chemistry, and biochemistry. He received his B.S. degree in chemistry from Purdue University and his Ph.D. degree in organic chemistry from Iowa State University. His interest in plant alkaloids led to a year of postdoctoral study in biochemistry at Texas A&M University. His current professional interests are chemistry education and community involvement in science activities, particularly the State Science and Engineering Fair in Utah. He also enjoys the company of family, hiking in the mountains, and fishing the local streams.

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Brief Contents

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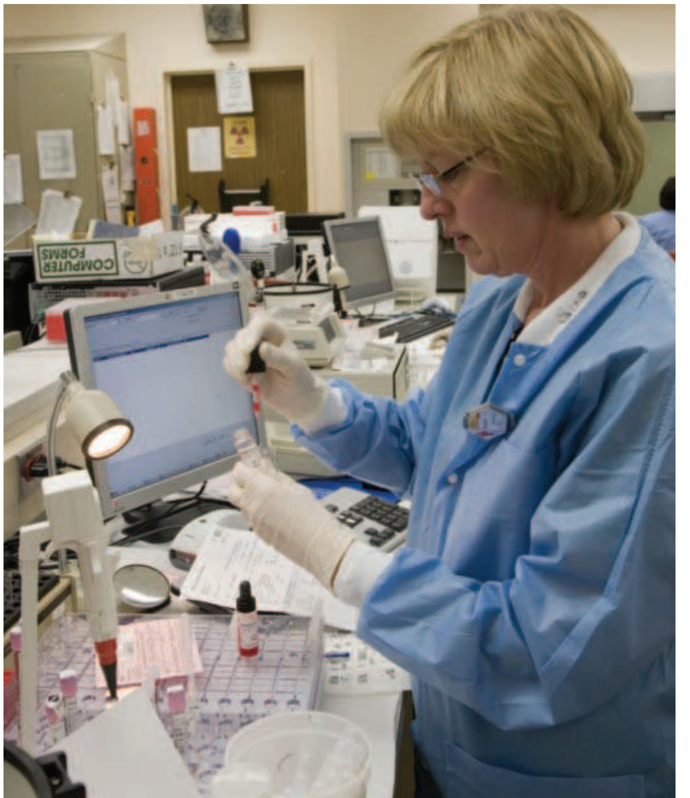
Jochen Sands/Digital Vision/Getty Images



Jim West/PhotoLibrary

3660 Group Inc./Custom Medical Stock Photo

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Preface

The Image of Chemistry

We, as authors, are pleased that the acceptance of the previous six editions of this textbook by students and their teachers has made it possible to publish this seventh edition. In the earlier editions, we expressed our concern about the negative image of chemistry held by many of our students, and their genuine fear of working with chemicals in the laboratory. Unfortunately, this negative image not only persists, but seems to be intensifying. Reports in the media related to chemicals or to chemistry continue to be primarily negative, and in many cases seem to be designed to increase the fear and concern of the general public. With this edition, we continue to hope that those who use this book will gain a more positive understanding and appreciation of the important contributions that chemistry makes in their lives.

Theme and Organization

This edition continues the theme of the positive and useful contributions made by chemistry in our world. Consistent with that theme, we continue to use the chapter opening focus on health care professionals introduced in the second edition. The photos and accompanying brief descriptions of the role of chemistry in each profession continue to emphasize positive contributions of chemistry in our lives.

This text is designed to be used in either a two-semester or three-quarter course of study that provides an introduction to general chemistry, organic chemistry, and biochemistry. Most students who take such courses are majoring in nursing, other health professions, or the life sciences, and consider biochemistry to be the most relevant part of the course of study. However, an understanding of biochemistry depends upon a sound background in organic chemistry, which in turn depends upon a good foundation in general chemistry. We have attempted to present the general and organic chemistry in sufficient depth and breadth to make the biochemistry understandable.

As with previous editions, this textbook is published in a complete hardcover form and a two-volume paperback edition. One volume of the paperback edition contains all the general chemistry and the first two chapters of organic chemistry from the hardcover text. The second volume of the paperback edition contains all the organic chemistry and biochemistry of the hardcover edition. The availability of the textbook in these various forms has been a very popular feature among those who use the text because of the flexibility it affords them.

The decisions about what to include and what to omit from the text were based on our combined 70-plus years of teaching, input from numerous reviewers and adopters, and our philosophy that a textbook functions as a personal tutor to each student. In the role of a personal tutor, a text must be more than just a collection of facts, data, and exercises. It should also help students relate to the material they are studying, carefully guide them through more difficult material, provide them with interesting and relevant examples of chemistry in their lives, and become a reference and a resource that they can use in other courses or their professions.

New to This Edition

In this seventh edition of the text, we have retained features that received a positive reception from our own students, the students of other adopters, other teachers, and reviewers. The retained features are 24 *Study Skills* boxes that include 5 reaction maps; 4 *How Reactions Occur* boxes; 44 *Chemistry Around Us* boxes, including 19 new to this edition. The former feature *Over The Counter* has been changed to *At The Counter* and reflects coverage of both prescription and non-prescription health-related products. Twelve of the 24 *At The Counter* boxes are new to this edition. There are 22 *Chemistry and Your Health* boxes, with 8 new to this edition. A greatly expanded feature of this seventh edition is the *Allied Health Exam Connection* that follows the exercises at the end of each chapter. This feature consists of examples of chemistry questions found on typical entrance examinations used to screen applicants to allied health professional programs. In addition, approximately 20% of the end-of-chapter exercises have been changed.

Allied Health Exam Connection

The following questions are from these sources:

1. *Nursing School Entrance Exam* © 2005, Learning Express, LLC.
2. *McGraw-Hill's Nursing School Entrance Exams* by Thomas A. Evangelist, Tamara B. Orr and Judy Unrein © 2009, The McGraw-Hill Companies, Inc.
3. *NSEE Nursing School Entrance Exams,* 3rd Edition © 2009, Kaplan Publishing.
4. *Cliffs Test Prep: Nursing School Entrance Exams* by Fred N. Grayson © 2004, Wiley Publishing, Inc.
5. *Peterson's Master the Nursing School and Allied Health Entrance Exams,* 18th Edition by Marion F. Gooding © 2008, Peterson's, a Nelnet Company.

5.66 Balance the following redox reaction:

$Mg(s) + H_2O(g) \rightarrow Mg(OH)_2(s) + H_2(g)$

a. $Mg(s) + H_2O(g) \rightarrow Mg(OH)_2(s) + H_2(g)$
b. $Mg(s) + 4H_2O(g) \rightarrow Mg(OH)_2(s) + H_2(g)$
c. $Mg(s) + 2H_2O(g) \rightarrow Mg(OH)_2(s) + H_2(g)$
d. $Mg(s) + H_2O(g) \rightarrow Mg(OH)_2(s) + \frac{1}{2}H_2(g)$

d. −2

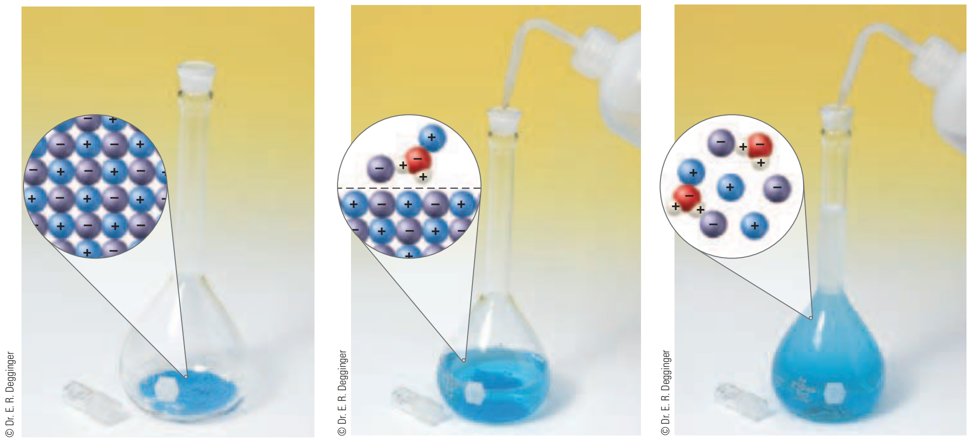
5.72 What is the oxidation number for nitrogen in HNO_3?
a. −2
b. +5
c. −1
d. −5

5.73 The oxidation number of sulfur in the ion SO_4^{-2} is:
a. −2
b. +2
c. +6
d. +10

5.74 Which of the following is the oxidation number of sulfur in the compound sodium thiosulfate, $Na_2S_2O_3$?
a. +1
b. −1
c. +2
d. −2

5.75 Which best describes the following redox reaction:

Also new to this edition are many new photographs and updated art to further enhance student comprehension of key concepts, processes and preparation.



1 *Step 1.* Put 0.125 mol of solute into a 250-mL flask.

2 *Step 2.* Add some water and dissolve the solute.

3 *Step 3.* Fill the flask to the mark with water. Mix thoroughly.

Figure 7.8 Preparation of a 0.500 M solution. Use the data given in the figure and show by a calculation that the resulting solution is 0.500 M.

Revision Summary of Seventh Edition:

Chapter 1:

- Several revised figures
- New photography
- Revised Examples
- New Chemistry Around Us: *Green Chemistry*
- 20% new Exercises
- Numerous new Allied Health Connection Questions

Chapter 2:

- Several revised figures
- New photography
- Revised and new Examples
- New At the Counter: *Calcium Supplements: Which Type is Best?*
- 20% new Exercisese
- Numerous new Allied Health Connection Questions

Chapter 3:

- Several revised figures
- New photography
- 20% new Exercisese
- Numerous new Allied Health Connection Questions

Chapter 4:

- Several revised figures
- New photography
- Revised Examples
- New Chemistry and Your Health: *Fight Hypertension with Potassium*
- 20% new Exercises
- Numerous new Allied Health Connection Questions

Chapter 5:

- Several revised figures
- New photography
- Revised and new Examples
- New Chemistry Around Us: *Ozone: The Good and the Bad*
- 20% new Exercises
- Numerous new Allied Health Connection Questions

Chapter 6:

- Several revised figures
- New photography
- New Chemistry and Your Health: *Huffing: A Potential Introduction of Children to Drug Abuse*
- New At the Counter: *Cutting Drug Costs with Generics*
- 20% new Exercises
- Numerous new Allied Health Connection Questions

Chapter 7:

- Several revised figures
- New photography
- Revised and new Examples
- New Chemistry and Your Health: *The Risk of Dehydration During Vigorous Youth Activities*
- 20% new Exercises
- Numerous new Allied Health Connection Questions

Chapter 8:

- Several revised figures
- New photography
- New Chemistry and Your Health: *Hypothermia: Surviving the Big Chill*
- New Chemistry Around Us: *The True Value of Platinum and Gold*
- 20% new Exercises
- Numerous new Allied Health Connection Questions

Chapter 9:

- Several revised figures
- New photography
- Revised Examples
- New Chemistry Around Us: *Beware the Negative Effects of Acids on Teeth*
- 20% new Exercises
- Numerous new Allied Health Connection Questions

Chapter 10:

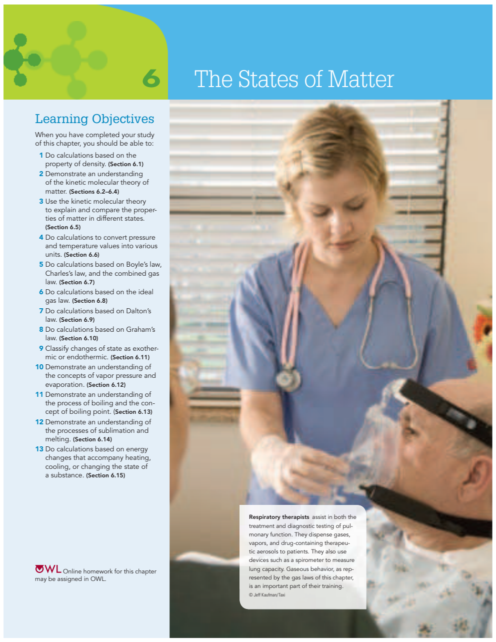
- Several revised figures
- New photography
- New At the Counter: *The Do's and Don'ts of Buying Prescription Drugs Online*
- 20% new Exercises
- Numerous new Allied Health Connection Questions

Chapter 11:

- Several revised figures
- New photography
- 20% new Exercises
- Numerous new Allied Health Connection Questions

Chapter 12:

- Several revised figures
- New photography
- 20% new Exercises
- Numerous new Allied Health Connection Questions



6 The States of Matter

Learning Objectives

When you have completed your study of this chapter, you should be able to:

1 Do calculations based on the property of density. **(Section 6.1)**
2 Demonstrate an understanding of the kinetic molecular theory of matter. **(Sections 6.2–6.4)**
3 Use the kinetic molecular theory to explain and compare the properties of matter in different states. **(Section 6.5)**
4 Do calculations to convert pressure and temperature values into various units. **(Section 6.6)**
5 Do calculations based on Boyle's law, Charles's law, and the combined gas law. **(Section 6.7)**
6 Do calculations based on the ideal gas law. **(Section 6.8)**
7 Do calculations based on Dalton's law. **(Section 6.9)**
8 Do calculations based on Graham's law. **(Section 6.10)**
9 Classify changes of state as exothermic or endothermic. **(Section 6.11)**
10 Demonstrate an understanding of the concepts of vapor pressure and evaporation. **(Section 6.12)**
11 Demonstrate an understanding of the process of boiling and the concept of boiling point. **(Section 6.13)**
12 Demonstrate an understanding of the processes of sublimation and melting. **(Section 6.14)**
13 Do calculations based on energy changes that accompany heating, cooling, or changing the state of a substance. **(Section 6.15)**

OWL Online homework for this chapter may be assigned in OWL.

Respiratory therapists assist in both the treatment and diagnostic testing of pulmonary function. They dispense gases, vapors, and drug-containing therapeutic aerosols to patients. They also use devices such as a spirometer to measure lung capacity. Gaseous behavior, as represented by the gas laws of this chapter, is an important part of their training.

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Features

Each chapter has features especially designed to help students study effectively, as well as organize, understand, and enjoy the material in the course.

Chapter Opening Photos. Each chapter opens with a photo of one of the many health care professionals that provide us with needed services. These professionals represent some of the numerous professions that require an understanding of chemistry.

Chapter Outlines and Learning Objectives. At the beginning of each chapter, a list of learning objectives provides students with a convenient overview of what they should gain by studying the chapter. In order to help students navigate through each chapter and focus on key concepts, these objectives are repeated at the beginning of the section in which the applicable information is discussed. The objectives are referred to again in the concept summary at the end of each chapter along with one or two suggested end-of-chapter exercises. By working the suggested exercises, students get a quick indication of how well they have met the stated learning objectives. Thus, students begin each chapter with a set of objectives and end with an indication of how well they satisfied the objectives.

Key Terms. Identified within the text by the use of bold type, key terms are defined in the margin near the place where they are introduced. Students reviewing a chapter can quickly identify the important concepts on each page with this marginal glossary. A full glossary of key terms and concepts appears at the end of the text.

At the Counter. These boxed features contain useful information about health-related products that are readily available to consumers with or without a prescription. The information in each box provides a connection between the chemical behavior of the product and its effect on the body.

At The Counter 2.1

Calcium Supplements: Which Type Is Best?

In a nutritional context, a supplement provides an amount of a substance that is in addition to the amount normally obtained from the diet.

About 99% of the calcium in the body is used to build bones and teeth. During a lifetime, all bones of the body undergo a natural process of buildup and breakdown. The rate of buildup exceeds the rate of breakdown for the first 25–30 years of life for women and the first 30–35 years of life for men. Beyond these times, the rate of breakdown exceeds the rate of buildup, resulting in a gradual decrease in bone density. Consequently, bones become increasingly weakened, brittle, and susceptible to breaking—a condition called *osteoporosis.* About 50% of women and 13% of men over age 50 suffer a broken bone as a result of osteoporosis.

One of the best ways to reduce the risks associated with osteoporosis is to build as much bone as possible during early life when the rate

If a calcium supplement is needed, which type is best? Most supplements will contain calcium in one of the following three chemical forms: calcium carbonate (often from oyster shells), calcium citrate or calcium phosphate. It really makes little difference which of these three chemical forms the calcium is in, as all three are absorbed quite well by the body. The important factor in a supplement is the amount of calcium contained in each dose. This amount per dose is generally indicated on the label and typically ranges from 333 mg to 630 mg. The maximum benefit from calcium supplements is obtained when the individual dosage is 500 mg or less. So, supplements with individual dosages greater than 500 mg should be divided and taken in portions throughout the day. An additional consideration is that vitamin D is essential for maximum calcium absorption by the body. For this reason, many calcium supplements include vitamin D in their formulation, and clearly indicate this on their labels.

Chemistry Around Us. These boxed features present everyday applications of chemistry that emphasize in a real way the important role of chemistry in our lives. Forty percent of these are new to this edition and emphasize health-related applications of chemistry.

Chemistry and Your Health. These boxed features contain current chemistry-related health issues such as "The Importance of Color in Your Diet," and questions about topics such as safety concerns surrounding genetically modified foods and the relationship between C-reactive protein and heart disease.

Chemistry and Your Health 5.1

The Importance of Color in Your Diet

Scientific evidence accumulated during the 1990s suggested that diets rich in fruits and vegetables had a protective effect against a number of different types of cancer. Studies showed that simply increasing the levels of vitamins and minerals in the diet did not provide the increased protection. This led to research into the nature of other substances found in fruits and vegetables that are important for good health. As a result of this research, a number of chemical compounds found in plants and called *phytonutrients* have been shown to be involved in the maintenance of healthy tissues and organs. The mechanism for their beneficial action in the body is not understood for all phytonutrients, but a significant number are known to work as antioxidants that stop harmful oxidation reactions from occurring.

The colors of fruits and vegetables help identify those containing beneficial compounds. The table below contains a list of some of the more well-known phytonutrients together with sources, colors, and beneficial actions. The amount of evidence supporting the existence of benefits from phytonutrients is not the same for all those listed in the table. In some cases, the experimental evidence is extensive (e.g., the cancer-blocking behavior of isothiocyanates), while in other cases the listed benefits are based on a limited amount of research and more studies are being done (e.g., the contribution to eye health by anthocyanins).

Fruit/Vegetable Color	Fruit/Vegetable Examples	Phytonutrients	Possible Benefits
Red  © iStockphoto.com/ DNY59	Tomatoes, watermelon, pink grapefruit	Lycopene (a carotenoid)	Protect against prostate, cervical, and pancreatic cancer and heart and lung disease

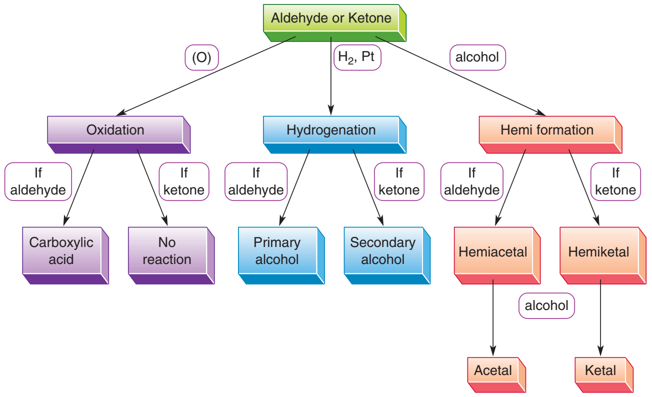
Examples. To reinforce students in their problem-solving skill development, complete step-by-step solutions for numerous examples are included in each chapter.

Learning Checks. Short self-check exercises follow examples and discussions of key or difficult concepts. A complete set of solutions is included in Appendix C. These allow students to measure immediately their understanding and progress.

Study Skills. Most chapters contain a *Study Skills* feature in which a challenging topic, skill, or concept of the chapter is addressed. Study suggestions, analogies, and approaches are provided to help students master these ideas.

Study Skills 14.1 A Reaction Map for Aldehydes and Ketones

This reaction map is designed to help you master organic reactions. Whenever you are trying to complete an organic reaction, use these two basic steps: (1) Identify the functional group that is to react, and (2) identify the reagent that is to react with the functional group. If the reacting functional group is an aldehyde or a ketone, find the reagent in the summary diagram, and use the diagram to predict the correct products.



How Reactions Occur. The mechanisms of representative organic reactions are presented in four boxed inserts to help students dispel the mystery of how these reactions take place.

Concept Summary. Located at the end of each chapter, this feature provides a concise review of the concepts and includes suggested exercises to check achievement of the learning objectives related to the concepts.

Concept Summary

Symbols and Formulas. Symbols based on names have been assigned to every element. Most consist of a single capital letter followed by a lowercase letter. A few consist of a single capital letter. Compounds are represented by formulas made up of elemental symbols. The number of atoms of each element in a molecule is shown by subscripts.

Objective 1, Exercise 2.4

Inside the Atom. Atoms are made up of numerous smaller particles of which the most important to chemical studies are the proton, neutron, and electron. Positively charged protons and neutral neutrons have a relative mass of 1 u each and are located in the nuclei of atoms. Negatively charged electrons with a mass of 1/1836 u are located outside the nuclei of atoms.

Objective 2, Exercises 2.10 and 2.12

Isotopes. Most elements in their natural state are made up of more than one kind of atom. These different kinds of atoms of a specific element are called isotopes and differ from one another only in the number of neutrons in their nuclei. A symbol incorporating atomic number, mass number, and elemental symbol is used to represent specific isotopes.

Objective 3, Exercises 2.16 and 2.22

Relative Masses of Atoms and Molecules. Relative masses tabulated in the periodic table. The units used are atomic mass units, abbreviated u. Relative masses for molecules, called molecular weights, are determined by adding the atomic weights of the atoms making up the molecules.

Objective 4, Exercise 2.32

Isotopes and Atomic Weights. The atomic weights measured for elements are average weights that depend on the percentages and masses of the isotopes in the naturally occurring element. If the isotope percent abundances and isotope masses are known for an element, its atomic weight can be calculated.

Objective 5, Exercise 2.38

Avogadro's Number: The Mole. Avogadro's number of the atoms of an element has a mass in grams equal to the atomic weight of the element. Avogadro's number of molecules has a mass in grams equal to the molecular weight. Avogadro's number of particles is called a mole, abbreviated mol.

Objective 6, Exercises 2.44 a & b and 2.46 a & b

The Mole and Chemical Formulas. The mole concept when applied to molecular formulas gives numerous relationships that yield useful factors for factor-unit calculations.

Key Terms and Concepts. These are listed at the end of the chapter for easy review, with a reference to the chapter section in which they are presented.

Key Equations. This feature provides a useful summary of general equations and reactions from the chapter. This feature is particularly helpful to students in the organic chemistry chapters.

Exercises. Nearly 1,700 end-of-chapter exercises are arranged by section. Approximately half of the exercises are answered in the back of the text. Complete solutions to these answered exercises are included in the Student Study Guide. Solutions and answers to the remaining exercises are provided in the Instructor's Manual. We have included a significant number of clinical and other familiar applications of chemistry in the exercises.

Allied Health Exam Connection. These examples of chemistry questions from typical entrance exams used to screen applicants to allied health professional programs help students focus their attention on the type of chemical concepts considered important in such programs.

Chemistry for Thought. Included at the end of each chapter are special questions designed to encourage students to expand their reasoning skills. Some of these exercises are based on photographs found in the chapter, while others emphasize clinical or other useful applications of chemistry.

Possible Course Outlines

This text may be used effectively in either a two-semester or three-quarter course of study:

First semester: Chapters 1–13 (general chemistry and three chapters of organic chemistry)

Second semester: Chapters 14–25 (organic chemistry and biochemistry)

First semester: Chapters 1–10 (general chemistry)

Second semester: Chapters 11–21 (organic chemistry and some biochemistry)

First quarter: Chapters 1–10 (general chemistry)

Second quarter: Chapters 11–18 (organic chemistry)

Third quarter: Chapters 19–25 (biochemistry)

Supporting Materials

Supporting instructor materials are available to qualified adopters. Please consult your local Cengage Learning Brooks/Cole representative for details. Go to **www.cengage.com/chemistry/seager** and click your textbook's Faculty Companion Site to:

See samples of materials
Request a desk copy
Locate your local representative
Download digital resources for instructors and students

Print Resources

Safety-Scale Laboratory Experiments for *Chemistry for Today*: General, Organic, and Biochemistry, 7th Edition. ISBN 0-538-73454-X
Prepared by Spencer L. Seager and Michael R. Slabaugh, this well-tested collection of experiments has been developed during more than 35 years of laboratory instruction with students at Weber State University. This manual provides a blend of training in laboratory skills and experiences that illustrate concepts from the authors' textbook. The experiments are designed to use small quantities of chemicals, and emphasize safety and proper disposal of used materials.

Instructor's Guide for Safety-Scale Laboratory Experiments. ISBN 0-538-73525-2 Prepared by the authors of the laboratory manual, this useful resource gives complete directions for preparing the reagents and other materials used in each experiment. It also contains useful comments concerning the experiments, answers to questions included in the experiments, and suggestions for the proper disposal of used materials. The Instructor's Guide is available online, accessible from **www.cengage.com/chemistry/seager.**

Study Guide and Solutions Manual. Prepared by Jennifer P. Harris of Portland Community College, each chapter contains a chapter outline, learning objectives, detailed solutions to the even-numbered exercises answered in the text, and self-test questions. ISBN 0-538-73458-2. Download a sample chapter from the Student Companion Website, which is accessible from **www.cengage.com/chemistry/seager.**

Media Resources

OWL (Online Web Learning) for General, Organic, and Biochemistry

Instant Access to OWL (four semesters) ISBN 0-495-11105-8
Instant Access to OWL (one semester) ISBN 0-495-11121-X
Instant Access to OWL eBook (four semesters) ISBN 0-538-73587-2
Instant Access to OWL eBook (one semester) ISBN 0-538-79351-1

Authored by Roberta Day, Beatrice Botch, and David Gross of the University of Massachusetts, Amherst; William Vining of the State University of New York at Oneonta; and Susan Young of Hartwick College. Featuring an updated and more intuitive instructor interface, OWL offers more assignable, gradable content (including end-of-chapter questions specific to each textbook), more reliability, and more flexibility than any other system. Developed by chemistry instructors for teaching chemistry, OWL makes homework management a breeze and has already helped hundreds of thousands of students master chemistry through tutorials, interactive simulations, and algorithmically generated homework questions that provide instant, answer-specific feedback. In addition, when you become an OWL user, you can expect service that goes far beyond the ordinary. OWL is continually enhanced with online learning tools to address the various learning styles of today's students such as:

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Thinkwell Video Lessons that teach key concepts through video, audio, and whiteboard examples.

Jmol molecular visualization program for rotating molecules and measuring bond distances and angles.

For *Chemistry for Today,* OWL includes parameterized end-of-chapter questions from the text. To view an OWL demo, and for more information, visit **www.cengage.com/owl** or contact your Cengage Learning Brooks/Cole representative.

OWL for General Chemistry, Organic Chemistry, and Biochemistry. See the above description in the instructor support materials section.

PowerLecture with JoinIn™ and ExamView® Instructor's CD and DVD-ROM Package. ISBN 0-538-73461-2

PowerLecture is a one-stop digital library and presentation tool that includes:

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Image libraries in PowerPoint and JPEG formats that contain **digital files for all text art, most photographs, and all numbered tables** in the text. These files can be used to create your own transparencies or PowerPoint lectures.

Digital files for the complete Instructor's Solutions Manual and Test Bank.

Sample chapters from the Study Guide and Student Solutions Manual.

ExamView Computerized Testing by James K. Hardy of the University of Akron enables you to create customized tests of up to 250 items in print or online using more than 1000 questions carefully matched to the corresponding text sections. Tests can be taken electronically or printed for class distribution.

JoinIn™ STUDENT RESPONSE SYSTEM

JoinIn™ clicker questions authored by Jennifer Harris specifically for this text, for use with the classroom response system of your choice. Assess student progress with instant quizzes and polls, and display student answers seamlessly within the Microsoft PowerPoint slides of your own lecture questions. Please consult your Brooks/Cole representative for more details.

- **Instructor's Manual and Testbank.** ISBN 0-538-73459-0
- Prepared by James K. Hardy of the University of Akron, each chapter contains a summary of the chapter in outline form, learning objectives, lecture hints and suggestions, solutions to *Chemistry for Thought* questions, answers and solutions to odd-numbered exercises not answered in the text, and more than 1,300 exam questions. Digital files for the Instructor's Manual and Testbank are on the PowerLecture Instructor's CD.

Student Companion Website. Accessible from **www.cengage.com/chemistry/seager,** this site provides online study tools including an online glossary and flashcards, interactive versions of Active Figures from the text, and samples of the *Study Guide* and *Student Solutions Manual.*

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Faculty Companion Website. Go to **www.cengage.com/chemistry/moore** and click this book's Faculty Companion Site to access the *Instructor's Manual,* sample chapters from the Student Study Guide, and Blackboard and WebCT versions of ExamView.

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Spencer L. Seager

Michael R. Slabaugh

Matter, Measurements, and Calculations

1

The health care system is one of the largest employing industries in the United States. **Nurses** are an essential component of that system. Here a nurse assists in a delicate surgical procedure to place a stent into a coronary artery. This is one of many situations in which careful attention to measurement (a topic of this chapter) is crucial.

Jochen Sands/Digital Vision/Getty Images

Learning Objectives

When you have completed your study of this chapter, you should be able to:

1. Explain what matter is. **(Section 1.1)**
2. Explain the difference between the terms *physical* and *chemical* as applied to the properties of matter and changes in matter. **(Section 1.2)**
3. Describe matter in terms of the accepted scientific model. **(Section 1.3)**
4. On the basis of observation or information given to you, classify matter into the correct category of each of the following pairs: heterogeneous or homogeneous, solution or pure substance, and element or compound. **(Section 1.4)**
5. Recognize the use of measurement units in everyday activities. **(Section 1.5)**
6. Recognize units of the metric system, and convert measurements done using the metric system into related units. **(Section 1.6)**
7. Express numbers using scientific notation, and do calculations with numbers expressed in scientific notation. **(Section 1.7)**
8. Express the results of measurements and calculations using the correct number of significant figures. **(Section 1.8)**
9. Use the factor-unit method to solve numerical problems. **(Section 1.9)**
10. Do calculations involving percentages. **(Section 1.10)**
11. Do calculations involving densities. **(Section 1.11)**

OWL Online homework for this chapter may be assigned in OWL.

Chemistry is often described as the scientific study of matter. In a way, almost any study is a study of matter, because matter is the substance of everything. Chemists, however, are especially interested in matter; they study it and attempt to understand it from nearly every possible point of view.

The chemical nature of all matter makes an understanding of chemistry useful and necessary for individuals who are studying in a wide variety of areas, including the health sciences, the natural sciences, home economics, education, environmental science, and law enforcement.

Matter comes in many shapes, sizes, and colors that are interesting to look at and describe. Early chemists did little more than describe what they observed, and their chemistry was a descriptive science that was severely limited in scope. It became a much more useful science when chemists began to make quantitative measurements, do calculations, and incorporate the results into their descriptions. Some fundamental ideas about matter are presented in this chapter, along with some ideas about quantitative measurement, the scientific measurement system, and calculations.

1.1 What Is Matter?

Learning Objective

1. Explain what matter is.

Definitions are useful in all areas of knowledge; they provide a common vocabulary for both presentations to students and discussions between professionals. You will be expected to learn a number of definitions as you study chemistry, and the first one is a definition of *matter*. Earlier, we said that matter is the substance of everything. That isn't very scientific, even though we think we know what it means. If you stop reading for a moment and look around, you will see a number of objects that might include people, potted plants, walls, furniture, books, windows, and a TV set or radio. The objects you see have at least two things in common: Each one has mass, and each one occupies space. These two common characteristics provide the basis for the scientific definition of matter. **Matter** is anything that has mass and occupies space. You probably understand what is meant by an object occupying space, especially if you have tried to occupy the same space as some other object. The resulting physical bruises leave a lasting mental impression.

matter Anything that has mass and occupies space.

You might not understand the meaning of the term *mass* quite as well, but it can also be illustrated painfully. Imagine walking into a very dimly lit room and being able to just barely see two large objects of equal size on the floor. You know that one is a bowling ball and the other is an inflated plastic ball, but you can't visually identify which is which. However, a hard kick delivered to either object easily allows you to identify each one. The bowling ball resists being moved much more strongly than does the inflated ball. Resistance to movement depends on the amount of matter in an object, and **mass** is an actual measurement of the amount of matter present.

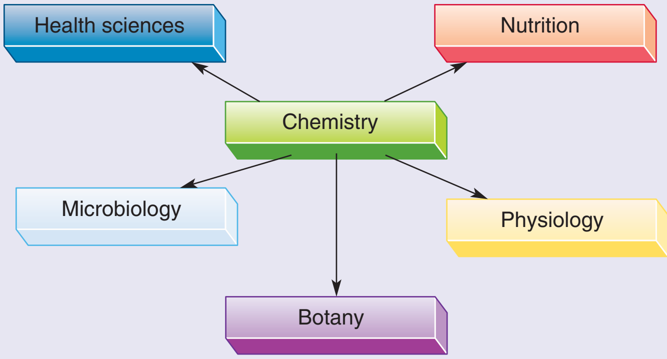
mass A measurement of the amount of matter in an object.

The term *weight* is probably more familiar to you than *mass,* but the two are related. All objects are attracted to each other by gravity, and the greater their mass, the stronger the attraction between them. The **weight** of an object on Earth is a measurement of the gravitational force pulling the object toward Earth. An object with twice the mass of a second object is attracted with twice the force, and therefore has twice the weight of the second object. The mass of an object is constant no matter where it is located (even if it is in a weightless condition in outer space). However, the weight of an object depends on

weight A measurement of the gravitational force acting on an object.

Chemistry Around Us 1.1
A Central Science

Chemistry is often referred to as the "central science" because it serves as a necessary foundation for many other scientific disciplines. Regardless of which scientific field you are interested in, every single substance you will discuss or work with is made up of chemicals. Also, many processes important to those fields will be based on an understanding of chemistry.



Chemistry is the foundation for many other scientific disciplines.

We also consider chemistry a central science because of its crucial role in responding to the needs of society. We use chemistry to discover new processes, develop new sources of energy, produce new products and materials, provide more food, and ensure better health.

As you read this text, you will encounter chapter opening photos dealing with applications of chemistry in the health care professions. Within the chapters, other Chemistry Around Us boxes focus on specific substances that play essential roles in meeting the needs of society.

Chemicals are present in everything we can touch, smell, or see. Chemistry is all around us.

the strength of the gravitational attraction to which it is subjected. For example, a rock that weighs 16 pounds on Earth would weigh about 2.7 pounds on the moon because the gravitational attraction is only about one-sixth that of Earth. However, the rock contains the same amount of matter and thus has the same mass whether it is located on Earth or on the moon.

Despite the difference in meaning between mass and weight, the determination of mass is commonly called "weighing." We will follow that practice in this book, but we will use the correct term *mass* when referring to an amount of matter.

1.2 Properties and Changes

Learning Objective

2. Explain the difference between the terms *physical* and *chemical* as applied to the properties of matter and changes in matter.

When you looked at your surroundings earlier, you didn't have much trouble identifying the various things you saw. For example, unless the decorator of your room had unusual tastes, you could easily tell the difference between a TV set and a potted plant by observing such characteristics as shape, color, and size. Our ability to identify objects or materials and discriminate between them depends on such characteristics. Scientists prefer to use the term *property* instead of characteristic, and they classify properties into two categories, physical and chemical.

Chemistry Around Us 1.2

Cosmetics: Complex Mixtures and Complex Regulations

The federal Food, Drug, and Cosmetic (FD&C) Act defines a cosmetic as anything applied directly to the human body for cleansing, beautifying, promoting attractiveness, or altering the appearance without affecting the body's structure or functions. According to this definition, mixtures as diverse as a modern roll-on deodorant and henna, a colored plant extract used in ancient times as well as today to dye hair, are classified as cosmetics. However, it is interesting to note that according to the FD&C Act, soap is not legally considered to be a cosmetic.

The sale of cosmetics in the United States is regulated by the federal Food and Drug Administration (FDA), but the regulatory requirements applied to the sale of cosmetics are not nearly as stringent as those applied to other FDA-regulated products. With the exception of color additives and a few prohibited substances, cosmetics manufacturers may use any ingredient or raw material in their products and market the products without obtaining FDA approval. The regulation that provides consumers with the greatest amount of information about the chemical composition of cosmetics comes not from the FDA, but from the Fair Packaging and Labeling Act. This act requires that every cosmetic product must be labeled with a list of all ingredients in order of decreasing quantity. For example, many skin-care products contain more water than any other ingredient, so water is listed first.

Any cosmetic product that is also designed to treat or prevent disease, or otherwise affect the structure or functions of the human body, is regulated as both a drug and a cosmetic, and must meet the labeling requirements for both. Some well-known examples of this type of product are dandruff shampoos, fluoride toothpastes, and antiperspirants/deodorants. A good way to tell if you are buying a cosmetic that is also regulated as a drug is to see if the first item on the ingredient label is listed as an "active ingredient." Regulations require that the active ingredient be identified and listed first, followed by the cosmetic ingredients in order of decreasing amounts.

© Maren Slabaugh

Many different types of products are classified as cosmetics. Each one must have a list of ingredients on the label.

physical properties Properties of matter that can be observed or measured without trying to change the composition of the matter being studied.

chemical properties Properties matter demonstrates when attempts are made to change it into new substances.

Physical properties are those that can be observed or measured without changing or trying to change the composition of the matter in question—no original substances are destroyed, and no new substances appear. For example, you can observe the color or measure the size of a sheet of paper without attempting to change the paper into anything else. Color and size are physical properties of the paper. **Chemical properties** are the properties matter demonstrates when attempts are made to change it into other kinds of matter. For example, a sheet of paper can be burned; in the process, the paper is changed into new substances. On the other hand, attempts to burn a piece of glass under similar conditions fail. The ability of paper to burn is a chemical property, as is the inability of glass to burn.

physical changes Changes matter undergoes without changing composition.

chemical changes Changes matter undergoes that involve changes in composition.

You can easily change the size of a sheet of paper by cutting off a piece. The paper sheet is not converted into any new substance by this change, but it is simply made smaller. **Physical changes** can be carried out without changing the composition of a substance. However, there is no way you can burn a sheet of paper without changing it into new substances. Thus, the change that occurs when paper burns is called a **chemical change.** Active Figure 1.1 shows an example of a chemical change, the burning of magnesium metal. The bright light produced by this chemical change led to the use of magnesium in the flash powder used in early photography. Magnesium is still used in fireworks to produce a brilliant white light.

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1 A strip of magnesium metal.

2 After being ignited with a flame, the magnesium burns with a blinding white light.

3 The white ash of magnesium oxide from the burning of several magnesium strips.

Active Figure 1.1 A chemical change occurs when magnesium metal burns. **Go to www.cengage.com/chemistry/seager or OWL to explore an interactive version of this figure.**

Example 1.1

Classify each of the following changes as physical or chemical: (a) a match is burned; (b) iron is melted; (c) limestone is crushed; (d) limestone is heated, producing lime and carbon dioxide; (e) an antacid seltzer tablet is dissolved in water; and (f) a rubber band is stretched.

Solution

Changes b, c, and f are physical changes because no composition changes occurred and no new substances were formed.

The others are chemical changes because new substances were formed. A match is burned—combustion gases are given off, and matchstick wood is converted to ashes. Limestone is heated—lime and carbon dioxide are the new substances. A seltzer tablet is dissolved in water—the fizzing that results is evidence that at least one new material (a gas) is produced.

Learning Check 1.1 Classify each of the following changes as physical or chemical, and, in the cases of chemical change, describe one observation or test that indicates new substances have been formed: (a) milk sours, (b) a wet handkerchief dries, (c) fruit ripens, (d) a stick of dynamite explodes, (e) air is compressed into a steel container, and (f) water boils.

Among the most common physical changes are changes in state, such as the melting of solids to form liquids, the sublimation of solids to form gases, or the evaporation of liquids to form gases. These changes take place when heat is added to or removed from matter, as represented in Figure 1.2. We will discuss changes in state in more detail in Chapter 6.

1.3 A Model of Matter

Learning Objective

3. Describe matter in terms of the accepted scientific model.

Model building is a common activity of scientists, but the results in many cases would not look appropriate on a fireplace mantle. **Scientific models** are explanations for observed behavior. Some, such as the well-known representation of the solar system, can easily be depicted in a physical way. Others are so abstract that they can be represented only by mathematical equations.

scientific models Explanations for observed behavior in nature.

A

B

Figure 1.2 Examples of physical change: Solid iodine becomes gaseous iodine when heated A; liquid benzene becomes solid benzene when cooled B.

Our present understanding of the nature of matter is a model that has been developed and refined over many years. Based on careful observations and measurements of the properties of matter, the model is still being modified as more is learned. In this book, we will concern ourselves with only some very basic concepts of this model, but even these basic ideas will provide a powerful tool for understanding the behavior of matter.

The study of the behavior of gases such as air, oxygen, and carbon dioxide by some of the earliest scientists led to a number of important ideas about matter. The volume of a gas kept at a constant temperature was found to change with pressure. An increase in pressure caused the gas volume to decrease, whereas a decrease in pressure permitted the gas volume to increase. It was also discovered that the volume of a gas maintained at constant pressure increased as the gas temperature was increased. Gases were also found to have mass and to mix rapidly with one another when brought together.

Figure 1.3 A hang glider soars far above the ground. How does this feat confirm that air is matter?

A simple model for matter was developed that explained these gaseous properties, as well as many properties of solids and liquids. Some details of the model are discussed in Chapter 6, but one conclusion is important to us now. All matter is made up of particles that are too small to see (see ▶ Figure 1.3). The early framers of this model called the small particles *molecules.* It is now known that molecules are the constituent particles of many, but not all, substances. In this chapter, we will limit our discussion to substances made up of molecules. Substances that are not made of molecules are discussed in Sections 4.3 and 4.11.

The results of some simple experiments will help us formally define the term *molecule.* Suppose you have a container filled with oxygen gas and you perform a number of experiments with it. You find that a glowing splinter of wood bursts into flames when placed in the gas. A piece of moist iron rusts much faster in the oxygen than it does in air. A mouse or other animal can safely breathe the gas.

Now suppose you divide another sample of oxygen the same size as the first into two smaller samples. The results of similar experiments done with these samples would be the same as before. Continued subdivision of an oxygen sample into smaller and smaller samples does not change the ability of the oxygen in the samples to behave just like the oxygen in the original sample. We conclude that the physical division of a sample of oxygen gas into smaller and smaller samples does not change the oxygen into anything else—it is still oxygen. Is there a limit to such divisions? What is the smallest sample of oxygen that will behave like the larger sample? We hope you have concluded that the smallest sample must be a single molecule. Although its very small size would make a one-molecule sample difficult to handle, it would nevertheless behave just as a larger sample would—it could be stored in a container, it would make wood burn rapidly, it would rust iron, and it could be breathed safely by a mouse.

We are now ready to formally define the term *molecule.* A **molecule** is the smallest particle of a pure substance that has the properties of that substance and is capable of a stable independent existence. Alternatively, a molecule is defined as the limit of physical subdivision for a pure substance.

molecule The smallest particle of a pure substance that has the properties of that substance and is capable of a stable independent existence. Alternatively, a molecule is the limit of physical subdivision for a pure substance.

In less formal terms, these definitions indicate that a sample of pure substance—such as oxygen, carbon monoxide, or carbon dioxide—can be physically separated into smaller and smaller samples only until there is a single molecule. Any further separation cannot be done physically, but if it were done (chemically), the resulting sample would no longer have the same properties as the larger samples.

The idea that it might be possible to chemically separate a molecule into smaller particles grew out of continued study and experimentation by early scientists. In modern terminology, the smaller particles that make up molecules are called atoms. John Dalton (1766−1844) is generally credited with developing the first atomic theory containing ideas that are still used today. The main points of his theory, which he proposed in 1808, can be summarized in the following five statements:

1. All matter is made up of tiny particles called atoms.
2. Substances called elements are made up of atoms that are all identical.
3. Substances called compounds are combinations of atoms of two or more elements.
4. Every molecule of a specific compound always contains the same number of atoms of each kind of element found in the compound.
5. In chemical reactions, atoms are rearranged, separated, or combined, but are never created nor destroyed.

Early scientists used graphic symbols such as circles and squares to represent the few different atoms that were known at the time. Instead of different shapes, we will use representations such as those in Figure 1.4 for oxygen, carbon monoxide, and carbon dioxide molecules.

The three pure substances just mentioned illustrate three types of molecules found in matter. Oxygen molecules consist of two oxygen atoms, and are called **diatomic molecules** to indicate that fact. Molecules such as oxygen that contain only one kind of atom are also called **homoatomic molecules** to indicate that the atoms are all of the same kind. Carbon

diatomic molecules Molecules that contain two atoms.

homoatomic molecules Molecules that contain only one kind of atom.

Figure 1.4 Symbolic representations of molecules.

heteroatomic molecules Molecules that contain two or more kinds of atoms.

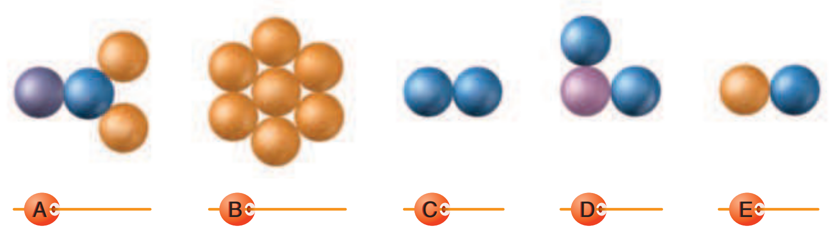
triatomic molecules Molecules that contain three atoms.

polyatomic molecules Molecules that contain more than three atoms.

monoxide molecules also contain two atoms and therefore are diatomic molecules. However, in this case the atoms are not identical, a fact indicated by the term **heteroatomic molecule.** Carbon dioxide molecules consist of three atoms that are not all identical, so carbon dioxide molecules are described by the terms **triatomic** and heteroatomic. The words *diatomic* and *triatomic* are commonly used to indicate two- or three-atom molecules, but the word **polyatomic** is usually used to describe molecules that contain more than three atoms.

Example 1.2

Use the terms *diatomic, triatomic, polyatomic, homoatomic,* or *heteroatomic* to classify the following molecules correctly:



Solution

A. Polyatomic and heteroatomic (more than three atoms, and the atoms are not all identical)
B. Polyatomic and homoatomic (more than three atoms, and the atoms are identical)
C. Diatomic and homoatomic (two atoms, and the atoms are identical)
D. Triatomic and heteroatomic (three atoms, and the atoms are not identical)
E. Diatomic and heteroatomic (two atoms, and the atoms are not identical)

Learning Check 1.2 Use the terms *diatomic, triatomic, polyatomic, homoatomic,* or *heteroatomic* to classify the following molecules correctly:

a. Water molecules have been found to contain two hydrogen atoms and one oxygen atom.
b. Molecules of ozone contain three oxygen atoms.
c. Natural gas is made up primarily of methane molecules which contain one atom of carbon and four atoms of hydrogen.

atom The limit of chemical subdivision for matter.

The subdivision of molecules into smaller particles is a chemical change. How far can such subdivisions of molecules go? You are probably a step ahead of us and have guessed that the answer is atoms. In fact, this provides us with a definition of atoms. An **atom** is the limit of chemical subdivision. In less formal terms, atoms are the smallest particles of matter that can be produced as a result of chemical changes. However, all chemical changes do not necessarily break molecules into atoms. In some cases, chemical changes might just divide a large molecule into two or more smaller molecules. Also, as we will see later, some chemical changes form larger molecules from smaller ones. The important point is that only chemical changes will produce a division of molecules, and the smallest particles of matter that can possibly be produced by such a division are called atoms.

1.4 Classifying Matter

Learning Objective

4. On the basis of observation or information given to you, classify matter into the correct category of each of the following pairs: heterogeneous or homogeneous, solution or pure substance, and element or compound.

Unknown substances are often analyzed to determine their compositions. An analyst, upon receiving a sample to analyze, will always ask an important question: Is the sample a pure substance or a mixture? Any sample of matter must be one or the other. Pure water and sugar are both pure substances, but you can create a mixture by stirring together some sugar and pure water.

What is the difference between a pure substance and a mixture? Two differences are that a **pure substance** has a constant composition and a fixed set of physical and chemical properties. Pure water, for example, always contains the same proportions of hydrogen and oxygen and freezes at a specific temperature. A **mixture** of sugar and water, however, can vary in composition, and the properties will be different for the different compositions. For example, a glass of sugar water could contain a few crystals of sugar or several spoonfuls. Properties such as the sweetness and freezing point would vary depending on the amount of sugar present in the mixture.

pure substance Matter that has a constant composition and fixed properties.

mixture A physical blend of matter that can theoretically be physically separated into two or more components.

Another difference is that a pure substance cannot be physically separated into simpler substances, whereas a mixture can theoretically be separated into its components. For example, if we heat a sugar-and-water mixture, the water evaporates, and the sugar remains. We say mixtures can theoretically be separated because some separations are very difficult to achieve. ◗ Figure 1.5 summarizes these ideas.

Pure substances, and mixtures such as sugar water, are examples of **homogeneous matter**—matter that has a uniform appearance and the same properties throughout. Homogeneous mixtures such as sugar water are called **solutions** (see ◗ Figure 1.6). Mixtures in which the properties and appearance are not uniform throughout the sample are examples of **heterogeneous matter.** The mixture of rock salt and sand that is spread on snowy roads during the winter is an example.

homogeneous matter Matter that has the same properties throughout the sample.

solutions Homogeneous mixtures of two or more pure substances.

heterogeneous matter Matter with properties that are not the same throughout the sample.

Commonly, the word *solution* is used to describe homogeneous liquid mixtures such as sugar water, but solutions of gases and solids also exist. The air around us is a gaseous solution, containing primarily nitrogen and oxygen. The alloys of some metals are solid solutions. For example, small amounts of copper are often added to the gold used in making jewelry. The resulting solid solution is harder than gold and has greater resistance to wear.

Most matter is found in nature in the form of heterogeneous mixtures. The properties of such mixtures depend on the location from which samples are taken. In some cases, the heterogeneity is obvious. In a slice of tomato, for example, the parts representing the skin, juice, seeds, and pulp can be easily seen and identified because they look different. Thus, at least one property (e.g., color or texture) is different for the different parts. However, a sample of clean sand from a seashore must be inspected very closely before slight differences in appearance can be seen for different grains. At this point, you might be thinking that even the solutions described earlier as homogeneous mixtures would appear to be heterogeneous if they were looked at closely enough. We could differentiate between sugar and water molecules if sugar solutions were observed under sufficient magnification. We will generally limit ourselves to differences normally visible when we classify matter as heterogeneous on the basis of appearance.

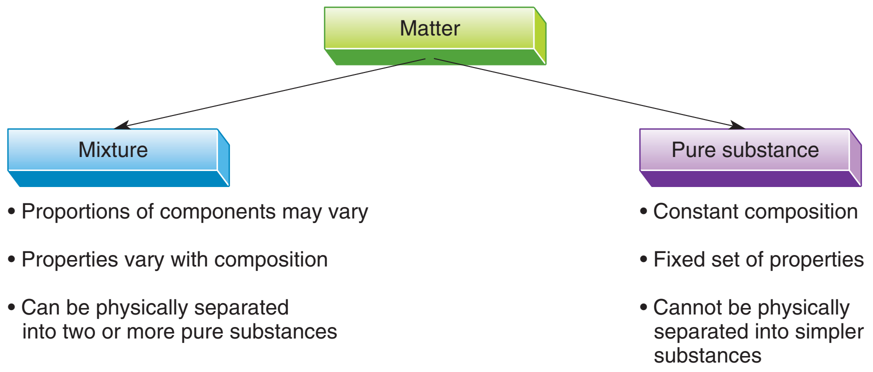


Figure 1.5 Mixtures and pure substances.

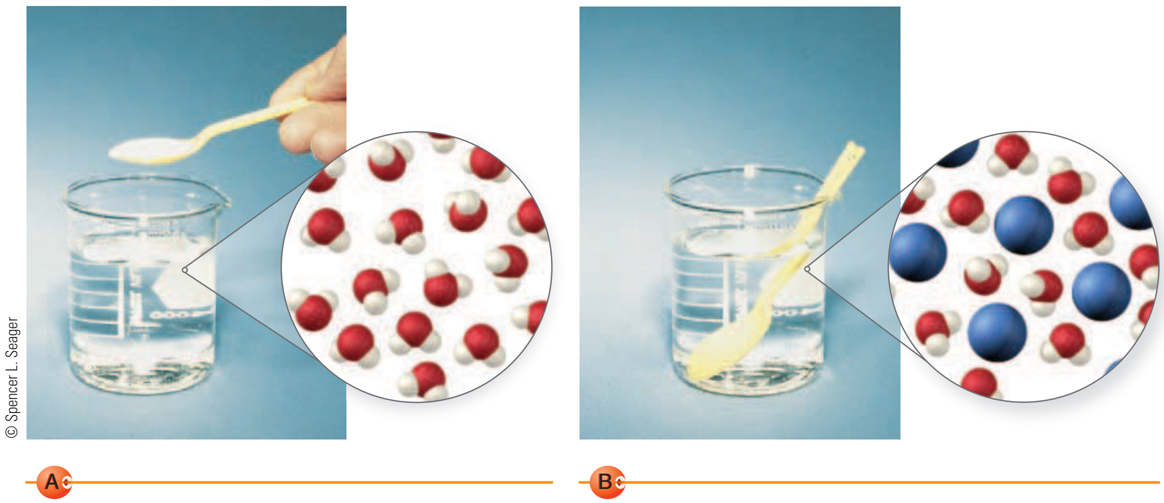


Figure 1.6 Sugar and water (A) form a solution when mixed (B).

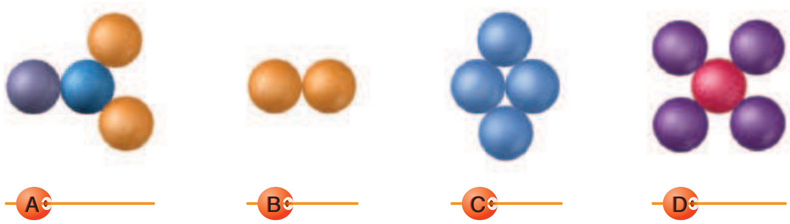
Earlier, we looked at three examples of pure substances—oxygen, carbon monoxide, and carbon dioxide—and found that the molecules of these substances are of different types. Oxygen molecules are diatomic and homoatomic, carbon monoxide molecules are diatomic and heteroatomic, and carbon dioxide molecules are triatomic and heteroatomic. Many pure substances have been found to consist of either homoatomic or heteroatomic molecules—a characteristic that permits them to be classified into one of two categories. Pure substances made up of homoatomic molecules are called **elements,** and those made up of heteroatomic molecules are called **compounds.** Thus, oxygen is an element, whereas carbon monoxide and carbon dioxide are compounds.

element A pure substance consisting of only one kind of atom in the form of homoatomic molecules or individual atoms.

compound A pure substance consisting of two or more kinds of atoms in the form of heteroatomic molecules or individual atoms.

It is useful to note a fact here that is discussed in more detail later in Section 4.11. The smallest particles of some elements and compounds are individual atoms rather than molecules. However, in elements of this type, the individual atoms are all of the same kind, whereas in compounds, two or more kinds of atoms are involved. Thus, the classification of a pure substance as an element or a compound is based on the fact that only one kind of atom is found in elements and two or more kinds are found in compounds. In both cases, the atoms may be present individually or in the form of homoatomic molecules (elements) or heteroatomic molecules (compounds). Some common household materials are pure substances (elements or compounds), such as aluminum foil, baking soda, and table salt.

Learning Check 1.3 Classify the molecules represented below as those of an element or a compound:



The characteristics of the molecules of elements and compounds lead us to some conclusions about their chemical behavior. Elements cannot be chemically subdivided into simpler pure substances, but compounds can. Because elements contain only one kind of atom and the atom is the limit of chemical subdivision, there is no chemical way to

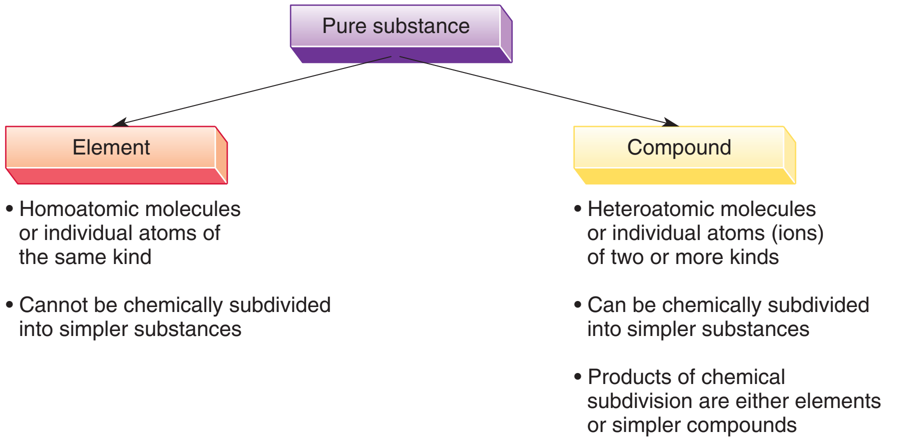


Figure 1.7 Elements and compounds.

break an element into any simpler pure substance—the simplest pure substance is an element. On the other hand, because the molecules of compounds contain more than one kind of atom, breaking such molecules into simpler pure substances is possible. For example, a molecule of table sugar can be chemically changed into two simpler molecules (which are also sugars) or into atoms or molecules of the elements carbon, hydrogen, and oxygen. Thus, compounds can be chemically subdivided into simpler compounds or elements. ◗ Figure 1.7 summarizes these ideas, and ◗ Figure 1.8 illustrates a classification scheme for matter based on the ideas we have discussed.

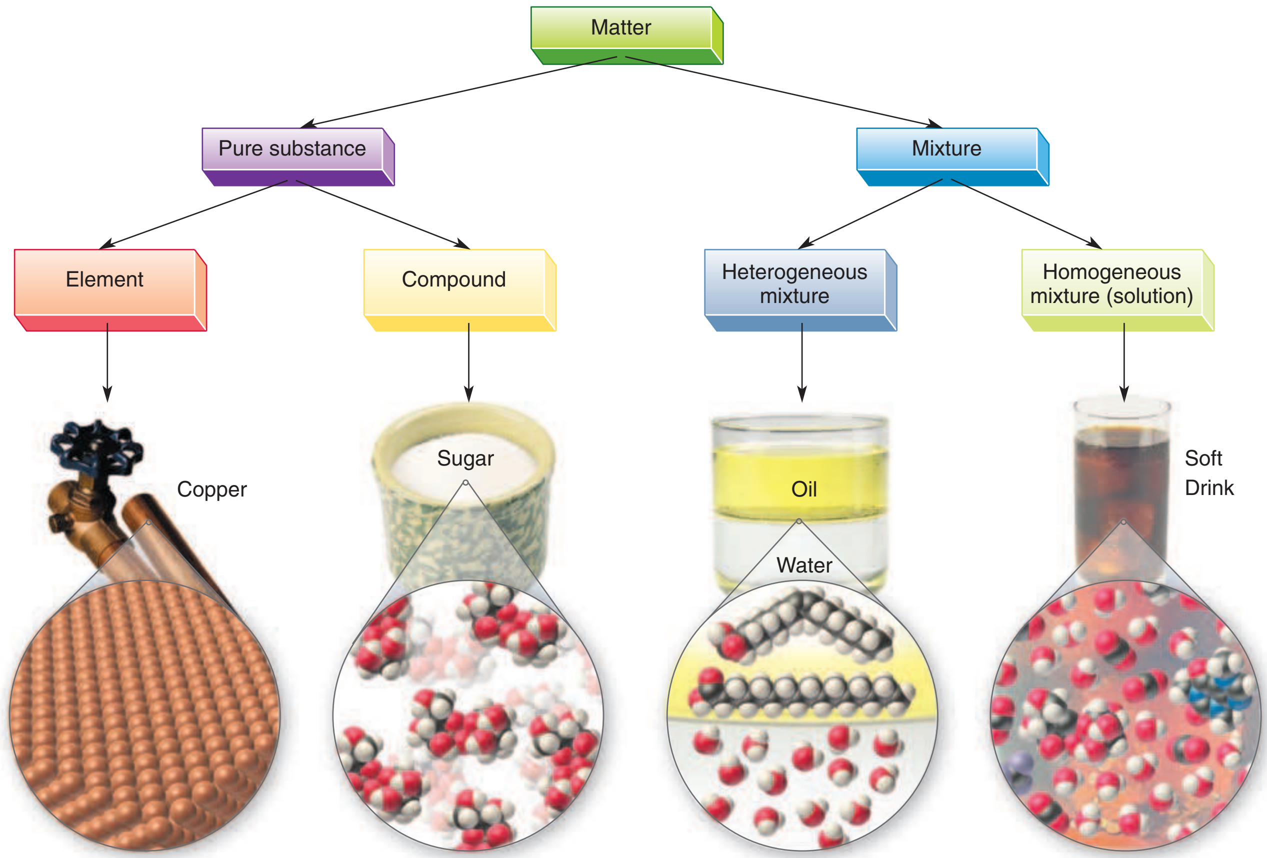


Figure 1.8 A classification scheme for matter.

Example 1.3

When sulfur, an element, is heated in air, it combines with oxygen to form sulfur dioxide. Classify sulfur dioxide as an element or a compound.

Solution

Because sulfur and oxygen are both elements and they combine to form sulfur dioxide, the molecules of sulfur dioxide must contain atoms of both sulfur and oxygen. Thus, sulfur dioxide is a compound because its molecules are heteroatomic.

Learning Check 1.4 Suppose an element and a compound combine to form only one product. Classify the product as an element or a compound.

1.5 Measurement Units

Learning Objective

5. Recognize the use of measurement units in everyday activities.

Matter can be classified and some physical or chemical properties can be observed without making any measurements. However, the use of quantitative measurements and calculations greatly expands our ability to understand the chemical nature of the world around us. A measurement consists of two parts, a number and an identifying unit. A number expressed without a unit is generally useless, especially in scientific work. We constantly make and express measurements in our daily lives. We measure the gallons of gasoline put into our cars, the time it takes to drive a certain distance, and the temperature on a hot or cold day. In some of our daily measurements, the units might be implied or understood. For example, if someone said the temperature outside was 39, you would probably assume this was 39 degrees Fahrenheit if you lived in the United States, but in most other parts of the world, it would be 39 degrees Celsius. Such confusion is avoided by expressing both the number and the unit of a measurement.

All measurements are based on units agreed on by those making and using the measurements. When a measurement is made in terms of an agreed-on unit, the result is expressed as some multiple of that unit. For example, when you purchase 10 pounds of potatoes, you are buying a quantity of potatoes equal to 10 times the standard quantity called 1 pound. Similarly, 3 feet of string is a length of string 3 times as long as the standard length that has been agreed on and called 1 foot.

The earliest units used for measurements were based on the dimensions of the human body. For example, the foot was the length of some important person's foot, and the biblical cubit was the length along the forearm from the elbow to the tip of the middle finger. One problem with such units is obvious; the size of the units changed when the person on whom they were based changed because of death, change in political power, and so on.

As science became more quantitative, scientists found that the lack of standard units became more and more of a problem. A standard system of units was developed in France about the time of the French Revolution and was soon adopted by scientists throughout the world. This system, called the *metric system,* has since been adopted and is used by almost all nations of the world. The United States adopted the system but has not yet put it into widespread use.

In an attempt to further standardize scientific measurements, an international agreement in 1960 established certain basic metric units, and units derived from them, as preferred units to be used in scientific measurements. Measurement units in this system are known as SI units after the French Système International d'Unités. SI units have not yet been totally put into widespread use. Many scientists continue to express certain quantities, such as volume, in non-SI units. The metric system in this book is generally based on accepted SI units but also includes a few of the commonly used non-SI units.

1.6 The Metric System

Learning Objective

6. Recognize units of the metric system, and convert measurements done using the metric system into related units.

The metric system has a number of advantages compared with other measurement systems. One of the most useful of these advantages is that the metric system is a decimal system in which larger and smaller units of a quantity are related by factors of 10. See Table 1.1 for a comparison between the metric and English units of length—a meter is slightly longer than a yard. Notice in Table 1.1 that the units of length in the metric system are related by multiplying a specific number of times by 10—remember, $100 = 10 \times 10$ and $1000 = 10 \times 10 \times 10$. The relationships between the units of the English system show no such pattern.

The relationships between units of the metric system that are larger or smaller than a **basic** (defined) **unit** are indicated by prefixes attached to the name of the basic unit. Thus, 1 kilometer (km) is a unit of length that is 1000 times longer than the basic unit of 1 meter (m), and a millimeter (mm) is only $\frac{1}{1000}$ the length of 1 m. Some commonly used prefixes are given in Table 1.2.

basic unit of measurement A specific unit from which other units for the same quantity are obtained by multiplication or division.

Table 1.1 Metric and English Units of Length

	Base Unit	Larger Unit	Smaller Unit
Metric	1 meter	1 kilometer = 1000 meters	10 decimeters = 1 meter 100 centimeters = 1 meter 1000 millimeters = 1 meter
English	1 yard	1 mile = 1760 yards	3 feet = 1 yard 36 inches = 1 yard

Table 1.2 Common Prefixes of the Metric System

Prefix[a]	Abbreviation	Relationship to Basic Unit	Exponential Relationship to Basic Unit[b]
mega-	M	1,000,000 × basic unit	10^{6} × basic unit
kilo-	k	1000 × basic unit	10^{3} × basic unit
deci-	d	1/10 × basic unit	10^{-1} × basic unit
centi-	c	1/100 × basic unit	10^{-2} × basic unit
milli-	m	1/1000 × basic unit	10^{-3} × basic unit
micro-	μ	1/1,000,000 × basic unit	10^{-6} × basic unit
nano-	n	1/1,000,000,000 × basic unit	10^{-9} × basic unit
pico-	p	1/1,000,000,000,000 × basic unit	10^{-12} × basic unit

[a]The prefixes in boldface (heavy) type are the most common ones. [b]The use of exponents to express large and small numbers is discussed in Section 1.7.

derived unit of measurement A unit obtained by multiplication or division of one or more basic units.

Area and volume are examples of **derived units** of measurement; they are obtained or derived from the basic unit of length:

$$\text{area} = (\text{length})(\text{length}) = (\text{length})^2$$

$$\text{volume} = (\text{length})(\text{length})(\text{length}) = (\text{length})^3$$

The unit used to express an area depends on the unit of length used.

Example 1.4

Calculate the area of a rectangle that has sides of 1.5 and 2.0 m. Express the answer in units of square meters and square centimeters.

Solution

$$\text{area} = (\text{length})(\text{length})$$

In terms of meters, area = (1.5 m)(2.0 m) = 3.0 m^2. Note that m^2 represents meter squared, or square meters. In terms of centimeters, area = (150 cm)(200 cm) = 30,000 cm^2. The lengths expressed in centimeters were obtained by remembering that 1 m = 100 cm.

Learning Check 1.5 The area of a circle is given by the formula $A = \pi r^2$, where r is the radius and $\pi = 3.14$. Calculate the area of a circle that has a radius of 3.5 cm.

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Figure 1.9 A liter is slightly larger than a quart.

The unit used to express volume also depends on the unit of length used in the calculation. Thus, a volume could have such units as cubic meters (m^3), cubic decimeters (dm^3), or cubic centimeters (cm^3). The abbreviation cc is also used to represent cubic centimeters, especially in medical work. The liter (L), a non-SI unit of volume, has been used as a basic unit of volume by chemists for many years (see Figure 1.9). For all practical purposes, 1 L and 1 dm^3 are equal volumes. This also means that 1 milliliter (mL) is equal to 1 cm^3. Most laboratory glassware is calibrated in liters or milliliters.

Example 1.5

A circular Petri dish with vertical sides has a radius of 7.50 cm. You want to fill the dish with a liquid medium to a depth of 2.50 cm. What volume of medium in milliliters and liters will be required?

Solution

The volume of medium required will equal the area of the circular dish (in square centimeters, cm^2) multiplied by the liquid depth (in centimeters, cm). Note that the unit of this product will be cubic centimeters (cm^3). According to Learning Check 1.5, the area of a circle is equal to πr^2, where $\pi = 3.14$. Thus, the liquid volume will be

$$V = (3.14)(7.50\text{ cm})^2(2.50\text{ cm}) = 442\text{ cm}^3$$

Because 1 cm^3 = 1 mL, the volume equals 442 mL. Also, because 1 L = 1000 mL, the volume can be converted to liters:

$$(442\ \cancel{\text{mL}})\left(\frac{1\text{ L}}{1000\ \cancel{\text{mL}}}\right) = 0.442\text{ L}$$

Notice that the milliliter units canceled in the calculation. This conversion to liters is an example of the factor-unit method of problem solving, which is discussed in Section 1.9.

Learning Check 1.6 A rectangular aquarium has sides with lengths of 30.0 cm and 20.0 cm, and a height of 15.0 cm. Calculate the volume of the aquarium, and express the answer in milliliters and liters.

The basic unit of mass in the metric system is 1 kilogram (kg), which is equal to about 2.2 pounds in the English system. A kilogram is too large to be conveniently used in some applications, so it is subdivided into smaller units. Two of these smaller units that are often used in chemistry are the gram (g) and milligram (mg) (see Figure 1.10). The prefixes *kilo-* (k) and *milli-* (m) indicate the following relationships between these units:

$$1\ \text{kg} = 1000\ \text{g}$$

$$1\ \text{g} = 1000\ \text{mg}$$

$$1\ \text{kg} = 1{,}000{,}000\ \text{mg}$$

Example 1.6

All measurements in international track and field events are made using the metric system. Javelins thrown by female competitors must have a mass of no less than 600 g. Express this mass in kilograms and milligrams.

Solution

Because 1 kg = 1000 g, 600 g can be converted to kilograms as follows:

$$600\ \cancel{\text{g}} \times \frac{1\ \text{kg}}{1000\ \cancel{\text{g}}} = 0.600\ \text{kg}$$

Also, because 1 g = 1000 mg,

$$600\ \cancel{\text{g}} \times \frac{1000\ \text{mg}}{1\ \cancel{\text{g}}} = 600{,}000\ \text{mg}$$

Once again, the units of the original quantity (600 g) were canceled, and the desired units were generated by this application of the factor-unit method (see Section 1.9).

Learning Check 1.7 The javelin thrown by male competitors in track and field meets must have a minimum mass of 0.800 kg. A javelin is weighed and found to have a mass of 0.819 kg. Express the mass of the weighed javelin in grams.

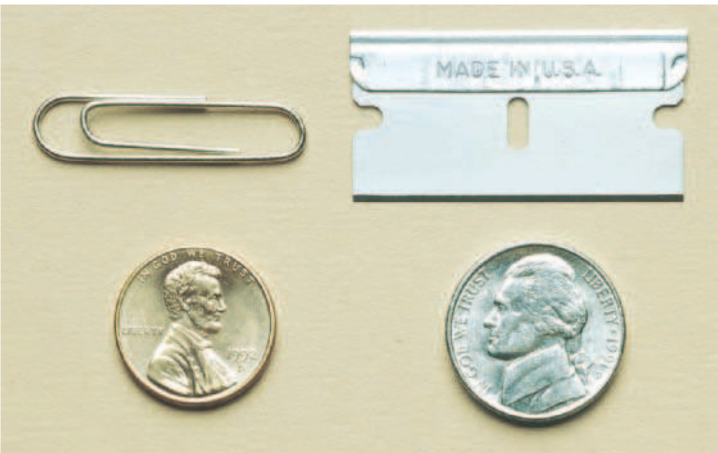

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Figure 1.10 Metric masses of some common items as found in a 0.4-g paper clip, 3.0-g razor blade, 3.1-g penny, and 4.7-g nickel.

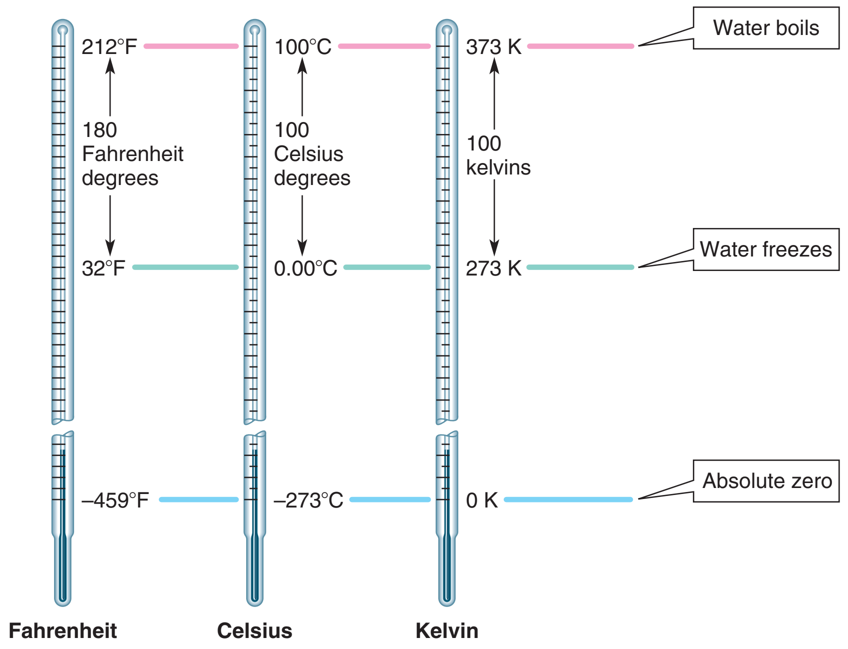


Figure 1.11 Fahrenheit, Celsius, and Kelvin temperature scales. The lowest temperature possible is absolute zero, 0 K.

Temperature is difficult to define but easy for most of us to measure—we just read a thermometer. However, thermometers can have temperature scales that represent different units. For example, a temperature of 293 would probably be considered quite high until it was pointed out that it is just room temperature as measured using the Kelvin temperature scale. Temperatures on this scale are given in kelvins, K. (Notice that the abbreviation is K, not °K.)

The Celsius scale (formerly known as the centigrade scale) is the temperature scale used in most scientific work. On this scale, water freezes at 0°C and boils at 100°C under normal atmospheric pressure. A Celsius degree (division) is the same size as a kelvin of the Kelvin scale, but the two scales have different zero points. ◗ Figure 1.11 compares the two scientific temperature scales and the familiar Fahrenheit scale. There are 100 Celsius degrees (divisions) between the freezing point (0°C) and the boiling point (100°C) of water. On the Fahrenheit scale, these same two temperatures are 180 degrees (divisions) apart (the freezing point is 32°F and the boiling point is 212°F). Readings on these two scales are related by the following equations:

$$^{\circ}\mathrm{C} = \frac{5}{9}(^{\circ}\mathrm{F} - 32) \tag{1.1}$$

$$^{\circ}\mathrm{F} = \frac{9}{5}(^{\circ}\mathrm{C}) + 32 \tag{1.2}$$

As mentioned, the difference between the Kelvin and Celsius scales is simply the zero point; consequently, readings on the two scales are related as follows:

$$^{\circ}\mathrm{C} = \mathrm{K} - 273 \tag{1.3}$$

$$\mathrm{K} = {^{\circ}\mathrm{C}} + 273 \tag{1.4}$$

Notice that Equation 1.2 can be obtained by solving Equation 1.1 for Fahrenheit degrees, and Equation 1.4 can be obtained by solving Equation 1.3 for kelvins. Thus, you need to remember only Equations 1.1 and 1.3, rather than all four.

Example 1.7

A temperature reading of 77°F is measured with a Fahrenheit thermometer. What reading would this temperature give if a Celsius thermometer were used?

Solution

The change is from a Fahrenheit reading to a Celsius reading, so Equation 1.1 is used:

$$°C = \frac{5}{9}(°F - 32) = \frac{5}{9}(77° - 32) = \frac{5}{9}(45°) = 25°$$

Thus, the reading on a Celsius thermometer would be 25°C.

Learning Check 1.8 What Kelvin thermometer reading would correspond to the 77°F reading described in Example 1.7?

The last units discussed at this point are derived units of energy. Other units will be introduced later in the book as they are needed. The metric system unit of energy is a joule (J), pronounced "jewl." A joule is quite small, as shown by the fact that a 50-watt light bulb uses 50 J of energy every second. A typical household in the United States uses several billion joules of electrical energy in a month.

The calorie (cal), a slightly larger unit of energy, is sometimes used by chemists. One calorie is the amount of heat energy required to increase the temperature of 1 g of water by 1°C. The calorie and joule are related as follows:

$$1 \text{ cal} = 4.184 \text{ J}$$

The nutritional calorie of the weight watcher is actually 1000 scientific calories, or 1 kcal. It is represented by writing *calorie* with a capital C (Calorie, abbreviated Cal). Table 1.3 contains a list of the commonly used metric units, their relationship to basic units, and their relationship to English units.

Table 1.3 Commonly Used Metric Units

Quantity	Metric Unit	Relationship to Metric Basic Unit	Relationship to English Unit
Length	meter (m)	Basic unit	1 m = 1.094 yd
	centimeter (cm)	100 cm = 1 m	1 cm = 0.394 in.
	millimeter (mm)	1000 mm = 1 m	1 mm = 0.0394 in.
	kilometer (km)	1 km = 1000 m	1 km = 0.621 mi
Volume	cubic decimeter (dm^3)	Basic unit	1 dm^3 = 1.057 qt
	cubic centimeter (cm^3 or cc)	1000 cm^3 = 1 dm^3	1 cm^3 = 0.0338 fl oz
	liter (L)	1 L = 1 dm^3	1 L = 1.057 qt
	milliliter (mL)[a]	1000 mL = 1 dm^3	1 mL = 0.0338 fl oz
Mass	gram (g)	1000 g = 1 kg	1 g = 0.035 oz
	milligram (mg)	1,000,000 mg = 1 kg	1 mg = 0.015 grain
	kilogram (kg)	Basic unit	1 kg = 2.20 lb
Temperature	degree Celsius (°C)	1°C = 1 K	1°C = 1.80°F
	kelvin (K)	Basic unit	1 K = 1.80°F
Energy	calorie (cal)	1 cal = 4.184 J	1 cal = 0.00397 BTU[b]
	kilocalorie (kcal)	1 kcal = 4184 J	1 kcal = 3.97 BTU
	joule (J)	Basic unit	1 J = 0.000949 BTU
Time	second (s)	Basic unit	Same unit used

Note: [a]1 mL = 1 cm^3. [b]A BTU (British thermal unit) is the amount of heat required to increase the temperature of 1 pound of water 1°F.

Chemistry Around Us 1.3

Green Chemistry

At first glance, the title "Green Chemistry" might generate thoughts of the very successful world wide synthetic chemical dye industry that is primarily responsible for the colorful society in which we live. However, today the words "Green Chemistry" have much more significance than the chemistry of color.

Begun in the 1990s, "Green Chemistry" represents growing movements with the goals of reducing the ecological damage done by the world wide chemical industry as it provides us with the many products on which we depend. In its simplest form, the chemical industry functions by mixing ingredients together in order to manufacture desired products. However, the dark side to this process is that there is almost always one or more by-products generated in addition to the desired product.

As an extreme example, it has been estimated that the production of each pound of useful drugs by the pharmaceutical industry also results in the generation of 25 to 100 pounds of waste materials—some of which pose serious health hazards. Some chemical industries are less wasteful than this example, but the United States Environmental Protection Agency (EPA) maintains that at least 40 million tons of hazardous waste is generated each year in the United States alone.

The Green Chemistry movement has a goal of reducing this ecological damage by redesigning how chemical products are made. As a result of such efforts, plywood that was formerly manufactured using glue made from formaldehyde is now made with a glue derived from less harmful materials. Also, energy-efficient cushions made from vegetable-oil-derived polyurethane foam are now available for purchase.

A fundamental idea of Green Chemistry is to focus on ways to reduce the formation of waste products rather than trying to figure out how to deal with waste after it has been generated. Some basic concepts central to this idea are:

- Design chemical processes to maximize the amount of raw material that actually ends up in the product.
- Use safe, environmentally benign substances in the manufacturing process whenever possible.
- Design and use energy-efficient manufacturing processes.
- Create as little waste as possible, and use environmentally appropriate ways to dispose of it.

A number of awards have been established to recognize innovations created by individuals and/or businesses that follow these concepts. The Presidential Green Chemistry Challenge Awards were begun in the United States in 1995. Other countries that have established similar awards include Australia, Canada, Italy, Japan, and the United Kingdom. Even the Nobel Prize Committee recognized the importance of Green Chemistry in 2005 by awarding the Nobel Prize in Chemistry to three researchers who developed a Green Chemistry procedure that was useful in synthesizing organic chemicals with minimal waste production.

So, the next time you hear or read the words "Green Chemistry," don't think only of dyes or colors. Instead, remember this important movement to devise safer, more ecologically friendly and more energy-efficient chemical processes to use in manufacturing the many chemical products upon which we all depend.

Environmental Protection Agency

"Green chemistry" products

1.7 Large and Small Numbers

Learning Objective

7. Express numbers using scientific notation, and do calculations with numbers expressed in scientific notation.

scientific notation A way of representing numbers consisting of a product between a nonexponential number and 10 raised to a whole-number exponent that may be positive or negative.

Numbers are used in all measurements and calculations. Many numbers are readily understood and represented because of common experience with them. A price of 10 dollars, a height of 7 feet, a weight of 165 pounds, and a time of 40 seconds are examples of such numbers. But how do we handle numbers like the diameter of a hydrogen atom (about one hundred-millionth of a centimeter) or the distance light travels in 1 year—a light-year (about 6 trillion miles)? These numbers are so small and large, respectively, that they defy understanding in terms of relationships to familiar distances. Even if we can't totally relate to them, it is important in scientific work to be able to conveniently represent and work with such numbers.

Scientific notation provides a method for conveniently representing any number including those that are very large or very small. In **scientific notation,** numbers are represented as the product of a nonexponential term and an exponential term in the general form $M \times 10^n$. The nonexponential term M is a number between 1 and 10 (but not equal to 10) written with a decimal to the right of the first nonzero digit in the number. This position of the decimal is called the **standard position.** The exponential term is a 10 raised to a whole number exponent n that may be positive or negative. The value of n is the number of places the decimal must be moved from the standard position in M to be at the original position in the number when the number is written normally without using scientific notation. If n is positive, the original decimal position is to the right of the standard position. If n is negative, the original decimal position is to the left of the standard position.

standard position for a decimal In scientific notation, the position to the right of the first nonzero digit in the nonexponential number.

Example 1.8

The following numbers are written using scientific notation. Write them without using scientific notation.

a. 3.72×10^5 **b.** 8.513×10^{-7}

Solution

a. The exponent 5 indicates that the original position of the decimal is located 5 places to the right of the standard position. Zeros are added to accommodate this change:

$$3.72 \times 10^5 = 372{,}000. = 372{,}000$$

standard position original position

b. The exponent -7 indicates that the original position of the decimal is 7 places to the left of the standard position. Again, zeros are added as needed:

$$8.513 \times 10^{-7} = .0000008513$$

original position standard position

Example 1.9

Write the following numbers using scientific notation:

a. 8725.6 **b.** 0.000729

Solution

a. The standard decimal position is between the 8 and 7: 8.7256. However, the original position of the decimal is 3 places to the right of the standard position. Therefore, the exponent must be $+3$:

$$8725.6 = 8.7256 \times 10^3$$

b. The standard decimal position is between the 7 and 2: 7.29. However, the original position is 4 places to the left of standard. Therefore, the exponent must be -4:

$$0.000729 = 7.29 \times 10^{-4}$$

The zeros to the left of the 7 are dropped because they are not significant figures (see Section 1.8); they only locate the decimal in the nonscientific notation and are not needed in the scientific notation.

Learning Check 1.9 Some of the following numbers are written using scientific notation, and some are not. In each case, rewrite the number using the notation in which it is not written.

a. 5.88×10^2 **c.** 3.915×10^{-4} **e.** 36.77
b. 0.000439 **d.** 9870 **f.** 0.102

Example 1.10

Determine which of the following numbers are written correctly using scientific notation. For those that are not, rewrite them correctly.

a. 001.5×10^{-3} **b.** 28.0×10^{2} **c.** 0.35×10^{4}

Solution

a. Incorrect; the zeros to the left are not needed. The correct answer is 1.5×10^{-3}.
b. Incorrect; the decimal is not in the standard position. Move the decimal 1 position to the left and increase the exponent by 1 to give the correct answer of 2.80×10^{3}.
c. Incorrect; the decimal is not in the standard position. Move the decimal 1 position to the right and decrease the exponent by 1 to give the correct answer of 3.5×10^{3}.

Learning Check 1.10 Determine which of the following numbers are written correctly using scientific notation. For those that are not, rewrite them correctly.

a. 62.5×10^{4} **b.** 0.0098 **c.** 0.0041×10^{-3} **d.** 7.85×10^{2}

The multiplication and division of numbers written in scientific notation can be done quite simply by using some characteristics of exponentials. Consider the following multiplication:

$$(a \times 10^{y})(b \times 10^{z})$$

The multiplication is done in two steps. First, the nonexponential terms a and b are multiplied in the usual way. The exponential terms 10^{y} and 10^{z} are multiplied by adding the exponents y and z and using the resulting sum as a new exponent of 10. Thus, we can write

$$(a \times 10^{y})(b \times 10^{z}) = (a \times b)(10^{y+z})$$

Division is done similarly. The nonexponential terms are divided in the usual way, and the exponential terms are divided by subtracting the exponent of the bottom term from that of the top term. The final answer is then written as a product of the resulting nonexponential and exponential terms:

$$\frac{a \times 10^{y}}{b \times 10^{z}} = \left(\frac{a}{b}\right)(10^{y-z})$$

Multiplication and division calculations involving scientific notation are easily done using a hand calculator. Table 1.4 gives the steps, the typical calculator procedures (buttons to press), and typical calculator readout or display for the division of 7.2×10^{-3} by 1.2×10^{4}.

Example 1.11

Do the following operations:

a. $(3.5 \times 10^{4})(2.0 \times 10^{2})$

b. $\dfrac{3.8 \times 10^{5}}{1.9 \times 10^{2}}$

c. $(4.6 \times 10^{-7})(5.0 \times 10^{3})$

d. $\dfrac{1.2 \times 10^{3}}{3.0 \times 10^{-2}}$

Table 1.4 Using a Calculator for Scientific Notation Calculations

Step	Procedure	Calculator Display
1. Enter 7.2	Press buttons 7, ., 2	7.2
2. Enter 10^{-3}	Press button that activates exponential mode (EE, Exp, etc.)	7.2^{00}
	Press 3	7.2^{03}
	Press change-sign button (±, etc.)	7.2^{-03}
3. Divide	Press divide button (÷)	7.2^{-03}
4. Enter 1.2	Press buttons 1, ., 2	1.2
5. Enter 10^4	Press button that activates exponential mode (EE, Exp, etc.)	1.2^{00}
	Press 4	1.2^{04}
6. Obtain answer	Press equals button (=)	$6.^{-07}$

Solution

a. $$(3.5 \times 10^4)(2.0 \times 10^2) = (3.5 \times 2.0)(10^4 \times 10^2) = (7.0)(10^{4+2}) = 7.0 \times 10^6$$

b. $$\frac{3.8 \times 10^5}{1.9 \times 10^2} = \frac{3.8}{1.9} \times \frac{10^5}{10^2} = (2.0)(10^{5-2}) = 2.0 \times 10^3$$

c. $$(4.6 \times 10^{-7})(5.0 \times 10^3) = (4.6 \times 5.0)(10^{-7} \times 10^3) = (23)(10^{-7+3}) = 23 \times 10^{-4}$$

To get the decimal into the standard position, move it 1 place to the left. This changes the exponent from -4 to -3, so the final result is 2.3×10^{-3}. (This number in decimal form, 0.0023, can be written correctly as either 23×10^{-4} or 2.3×10^{-3}, but scientific notation requires that the decimal point be to the right of the first nonzero number.)

d. $$\frac{1.2 \times 10^3}{3.0 \times 10^{-2}} = \frac{1.2}{3.0} \times \frac{10^3}{10^{-2}} = (0.40)(10^{3-(-2)}) = 0.40 \times 10^5$$

Adjust the decimal to the standard position and get 4.0×10^4. If these examples were done using a calculator, the displayed answers would normally be given in correct scientific notation.

Learning Check 1.11 Perform the following operations, and express the result in correct scientific notation:

a. $(2.4 \times 10^3)(1.5 \times 10^4)$

b. $(3.5 \times 10^2)(2.0 \times 10^{-3})$

c. $\dfrac{6.3 \times 10^5}{2.1 \times 10^3}$

d. $\dfrac{4.4 \times 10^{-2}}{8.8 \times 10^{-3}}$

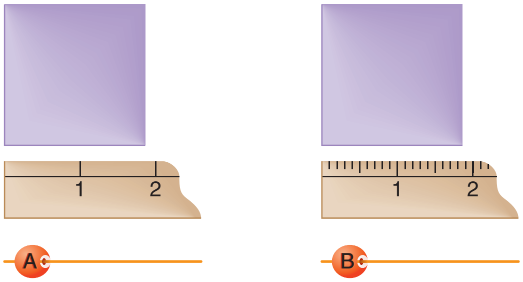
The diameter of a hydrogen atom, mentioned earlier as one hundred-millionth of a centimeter, is written in scientific notation as 1.0×10^{-8} cm. Similarly, 1 light-year of 6 trillion miles is 6.0×10^{12} mi.

1.8 Significant Figures

Learning Objective

8. Express the results of measurements and calculations using the correct number of significant figures.

Every measurement contains an uncertainty that is characteristic of the device used to make the measurement. These uncertainties are represented by the numbers used to record the measurement. Consider the following small square:



In (A), the length of one side of the square is measured with a ruler divided into centimeters. It is easy to see that the length is greater than 1 cm, but not quite 2 cm. The length is recorded by writing the number that is known with certainty to be correct (the 1) and writing an estimate for the uncertain number. The result is 1.9 cm, where the .9 is the estimate. In (B), the ruler is divided into tenths of centimeters. It is easy to see that the length is at least 1.8 cm, but not quite 1.9 cm. Once again the certain numbers (1.8) are written, and an estimate is made for the uncertain part. The result is 1.86 cm. When measurements are recorded this way, the numbers representing the certain measurement plus one number representing the estimate are called **significant figures.** Thus, the first measurement of 1.9 cm contains two significant figures, and the second measurement of 1.86 cm contains three significant figures.

significant figures The numbers in a measurement that represent the certainty of the measurement, plus one number representing an estimate.

The maximum number of significant figures possible in a measurement is determined by the design of the measuring device and cannot be changed by expressing the measurement in different units. The 1.8-cm length determined earlier can also be represented in terms of meters and millimeters as follows:

$$1.8 \text{ cm} = 0.018 \text{ m} = 18 \text{ mm}$$

In this form, it appears that the length expressed as 0.018 m contains four significant figures, but this is impossible; a measurement made with a device doesn't become more certain simply by changing the unit used to express the number. Thus, the zeros are not significant figures; their only function is to locate the correct position for the decimal. Zeros located to the left of nonzero numbers, such as the two zeros in 0.018 cm, are never considered to be significant. Thus, 12.5 mg, 0.0125 g, and 0.0000125 kg all represent the same measured mass, and all contain three significant figures.

Zeros located between nonzero numbers or trailing zeros located at the end of numbers will be considered significant. Thus, 2050 μL, 2.050 mL, and 0.002050 L all represent the same volume measurement, and all contain four significant figures.

The rule that specifies counting trailing zeros as significant is generally followed by scientists, but some quantities are expressed with trailing zeros that are not significant. For example, suppose you read in a newspaper that the population of a city is 1,250,000 people. Should the four trailing zeros be considered significant? If they are, it means that the population is known with certainty to the nearest 10 people and that the measurement has an uncertainty of only plus or minus 1 person. A more reasonable conclusion is that the census is correct to the nearest 1000 people. This could be represented as 1250 thousand or, using scientific notation, 1.250×10^6. In either of these representations, only four significant figures are used. In scientific notation, the correct number of significant figures is used in the nonexponential term, and the location of the decimal is determined by the exponent. In this book, large numbers will always be represented by scientific notation instead of using nonsignificant trailing zeros. However, you are likely to encounter nonsignificant trailing zeros in other reading materials. The rules for determining the significance of zeros are summarized as follows:

1. Zeros not preceded by nonzero numbers are not significant figures. These zeros are sometimes called *leading zeros.*
2. Zeros located between nonzero numbers are significant figures. These zeros are sometimes called *buried* or *confined zeros.*
3. Zeros located at the end of a number are significant figures. These zeros are sometimes called *trailing zeros.*

Example 1.12

Determine the number of significant figures in each of the following measurements, and use scientific notation to express each measurement using the correct number of significant figures:

a. 24.6°C **b.** 0.036 g **c.** 15.0 mL **d.** 0.0020 m

Solution

a. All the numbers are significant: three significant figures, 2.46×10^1 °C.
b. The leading zeros are not significant: two significant figures, 3.6×10^{-2} g.
c. The trailing zero is significant: three significant figures, 1.50×10^1 mL.
d. The leading zeros are not significant, but the trailing zero is: two significant figures, 2.0×10^{-3} m.

Learning Check 1.12 Determine the number of significant figures in each of the following measurements:

a. 250 mg **c.** 0.0108 kg **e.** 0.001 mm
b. 18.05 mL **d.** 37°C **f.** 101.0 K

Learning Check 1.13 Use scientific notation to express each of the following measurements using the correct number of significant figures:

a. 101 m **c.** 0.00230 kg **e.** 21.65 mL
b. 1200 g **d.** 1296°C **f.** 0.015 km

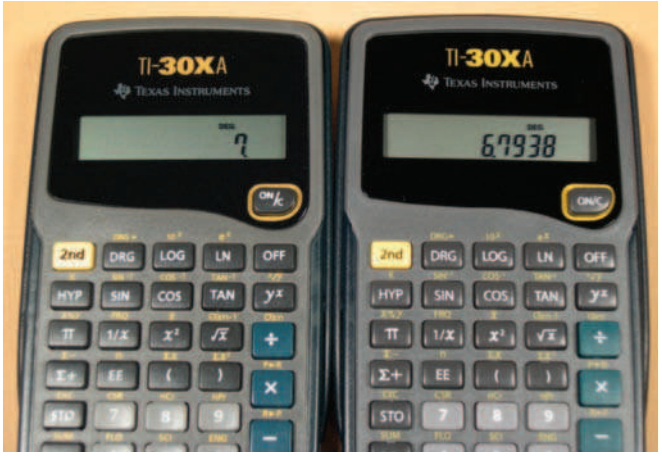
Most measurements that are made do not stand as final answers. Instead, they are usually used to make calculations involving multiplication, division, addition, or subtraction. The answer obtained from such a calculation cannot have more certainty than the least certain measurement used in the calculation. It should be written to reflect an uncertainty equal to that of the most uncertain measurement. This is accomplished by the following rules:

1. The answer obtained by multiplication or division must contain the same number of significant figures as the quantity with the fewest significant figures used in the calculation.
2. The answer obtained by addition or subtraction must contain the same number of places to the right of the decimal as the quantity in the calculation with the fewest number of places to the right of the decimal.

To follow these rules, it is often necessary to reduce the number of significant figures by rounding answers. The following are rules for rounding:

1. If the first of the nonsignificant figures to be dropped from an answer is 5 or greater, all the nonsignificant figures are dropped, and the last significant figure is increased by 1.

Figure 1.12 Calculators usually display the sum of 4.362 and 2.638 as 7 (too few figures), and the product of 0.67 and 10.14 as 6.7938 (too many figures).



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2. If the first of the nonsignificant figures to be dropped from an answer is less than 5, all nonsignificant figures are dropped, and the last significant figure is left unchanged.

Remember, if you use a calculator, it will often express answers with too few or too many figures (see Figure 1.12). It will be up to you to determine the proper number of significant figures to use and to round the calculator answer correctly.

Example 1.13

Do the following calculations, and round the answers to the correct number of significant figures:

a. $(4.95)(12.10)$ **b.** $\dfrac{3.0954}{0.0085}$ **c.** $\dfrac{(9.15)(0.9100)}{3.117}$ **d.** $\dfrac{320}{4.00}$

Solution

All calculations are done with a hand calculator, and the calculator answer is written first. Appropriate rounding is done to get the final answer.

a. Calculator answer: 59.895

The number 4.95 has three significant figures, and 12.10 has four. Thus, the answer must have three significant figures:

$$\underbrace{59.8}_{\text{significant figures}}\underbrace{95}_{\text{nonsignificant figures}}$$

The first of the nonsignificant figures to be dropped is 9, so after both are dropped, the last significant figure is increased by 1. The final answer containing three significant figures is 59.9.

b. Calculator answer: 364.16471

The number 3.0954 has five significant figures, and 0.0085 has two. Thus, the answer must have only two:

$$\underbrace{36}_{\text{significant figures}}\underbrace{4.16471}_{\text{nonsignificant figures}}$$

The first of the nonsignificant figures to be dropped is 4, so the last significant figure remains unchanged after the nonsignificant figures are dropped. The correct answer then is 360. If this is written as 360, it will contain three significant figures. However, the answer can be written with the proper number of significant figures by using scientific notation. The final correct answer containing two significant figures is 3.6×10^2.

c. Calculator answer: 2.6713186
The number 9.15 has three significant figures, 0.9100 has four, and 3.117 has four. Thus, the answer must have only three:

$$\underbrace{2.67}_{\substack{\text{significant}\\\text{figures}}}\underbrace{13186}_{\substack{\text{nonsignificant}\\\text{figures}}}$$

Appropriate rounding gives 2.67 as the correct answer.

d. Calculator answer: 80
Both the numbers 320 and 4.00 contain three significant figures, so the answer should also have three. However, modern calculators usually do not display trailing zeros that follow the decimal. So, the appropriate number of trailing zeros must be added. The correct answer should be expressed as 80.0, which contains three significant figures.

Learning Check 1.14 Do the following calculations, and round the answers to the correct number of significant figures:

a. $(0.0019)(21.39)$ **b.** $\dfrac{8.321}{4.1}$ **c.** $\dfrac{(0.0911)(3.22)}{1.379}$

Example 1.14

Do the following additions and subtractions, and write the answers with the correct number of significant figures:

a. $1.9 + 18.65$ **b.** $15.00 - 8.0$ **c.** $1500 + 10.9 + 0.005$ **d.** $5.1196 - 5.02$

Solution

In each case, the numbers are arranged vertically with the decimals in a vertical line. The answer is then rounded so it contains the same number of places to the right of the decimal as the smallest number of places in the quantities added or subtracted.

a.
$$\begin{array}{r} 1.9 \\ 18.65 \\ \hline 20.55 \end{array}$$
The answer must be expressed with one place to the right of the decimal to match the one place in 1.9.

Correct answer: 20.6 (Why was the final 5 increased to 6?)

b.
$$\begin{array}{r} 15.00 \\ -8.0 \\ \hline 7.0 \end{array}$$
The answer must be expressed using one place to the right of the decimal to match the 8.0. A typical calculator answer would probably be 7, requiring that a zero be added to the right of the decimal to provide the correct number of significant figures.

c.
$$\begin{array}{r} 1500 \\ 10.9 \\ 0.005 \\ \hline 1510.905 \end{array}$$
The answer must be expressed with no places to the right of the decimal to match the 1500.

Correct answer: 1511 (Why was the final 0 of the answer increased to 1?)

d. $\begin{array}{r} 5.1196 \\ -5.02 \\ \hline 0.0996 \end{array}$ The answer must be expressed with two places to the right of the decimal to match the 5.02.

Correct answer: 0.10 (What rounding rule was followed?)

Notice that the answer to (a) has more significant figures than the least significant number used in the calculation. The answer to (d), on the other hand, has fewer significant figures than either number used in the calculation. This happened because the rule for dealing with addition and subtraction focuses on the number of figures located to the right of the decimal and is not concerned with the figures to the left of the decimal. Thus, the number of significant figures in the answer is sometimes increased as a result of addition and decreased as a result of subtraction. You should be aware of this and not be confused when it happens.

Learning Check 1.15 Do the following additions and subtractions, and round the answers to the correct number of significant figures:

a. $8.01 + 3.2$ **c.** $4.33 - 3.12$
b. $3000 + 20.3 + 0.009$ **d.** $6.023 - 2.42$

exact numbers Numbers that have no uncertainty; numbers from defined relationships, counting numbers, and numbers that are part of simple fractions.

Some numbers used in calculations are **exact numbers** that have no uncertainty associated with them and are considered to contain an unlimited number of significant trailing zeros. Such numbers are not used when the appropriate number of significant figures is determined for calculated answers. In other words, exact numbers do not limit the number of significant figures in calculated answers. One kind of exact number is a number used as part of a defined relationship between quantities. For example, 1 m contains exactly 100 cm:

$$1 \text{ m} = 100 \text{ cm}$$

Thus, the numbers 1 and 100 are exact. A second kind of exact number is a counting number obtained by counting individual objects. A dozen eggs contains exactly 12 eggs, not 11.8 and not 12.3. The 12 is an exact counting number. A third kind of exact number is one that is part of a simple fraction such as $\frac{1}{4}$, $\frac{2}{3}$, or the $\frac{5}{9}$ used in Equation 1.1 to convert Fahrenheit temperature readings into Celsius readings.

1.9 Using Units in Calculations

Learning Objective

9. Use the factor-unit method to solve numerical problems.

Some beginning chemistry students are concerned about not being able to solve numerical chemistry problems. They may say, “I can work the problems, I just can’t set them up.” What they are really saying is “I can do the arithmetic once the numbers are properly arranged, but I can’t do the arranging.”

This section presents a method for arranging numbers that will work for most of the numerical problems you will encounter in this course. This method has a number of names, including the factor-unit method, the factor-label method, and dimensional analysis. We will call it the factor-unit method. It is a systematic approach to solving numerical problems and consists of the following steps:

Step 1. Write down the known or given quantity. Include both the numerical value and units of the quantity.
Step 2. Leave some working space and set the known quantity equal to the units of the unknown quantity.

Step 3. Multiply the known quantity by one or more factors, such that the units of the factor cancel the units of the known quantity and generate the units of the unknown quantity. These **factors** are fractions derived from numerical relationships between quantities. These relationships can be definitions or experimentally measured quantities. For example, the defined relationship 1 m = 100 cm provides the following two factors:

$$\frac{1\text{ m}}{100\text{ cm}} \quad \text{and} \quad \frac{100\text{ cm}}{1\text{ m}}$$

factors used in the factor-unit method Fractions obtained from numerical relationships between quantities.

Step 4. After you get the desired units, do the necessary arithmetic to produce the final answer.

Example 1.15

Use the factor-unit method and numerical relationships from Table 1.3 to calculate the number of yards in 100 m.

Solution

The known quantity is 100 m, and the unit of the unknown quantity is yard (yd).

Step 1. 100 m

Step 2. 100 m $\qquad = \text{yd}$

Step 3. $100\cancel{\text{m}} \times \dfrac{1.094\text{ yd}}{1\cancel{\text{m}}} = \text{yd}$

The factor $\dfrac{1.094\text{ yd}}{1\text{ m}}$ came from the numerical relationship 1 m = 1.094 yd found in Table 1.3.

Step 4. $\dfrac{(100)(1.094)\text{ yd}}{1} = 109.4\text{ yd}$

This answer should be rounded to 109 yd, an answer that contains three significant figures, just as 100 m does. The 1 m in the factor is an exact number used as part of a defined relationship, so it doesn't influence the number of significant figures in the answer.

Example 1.16

A laboratory technician uses a micropipet to measure a 50-μL (50-microliter) sample of blood serum for analysis. Express the sample volume in liters (L).

Solution

The known quantity is 50 μL, and the unit of the unknown quantity is liters.

Step 1. 50 μL

Step 2. 50 μL $\qquad = \text{L}$

Step 3. $50\,\cancel{\mu\text{L}} \times \dfrac{1 \times 10^{-6}\text{ L}}{1\,\cancel{\mu\text{L}}} = \text{L}$

The factor $\dfrac{1 \times 10^{-6}\text{ L}}{1\,\mu\text{L}}$ came from the numerical relationship $1\,\mu\text{L} = 1 \times 10^{-6}\text{ L}$ described in Table 1.2.

Step 4. $\dfrac{(50)(1 \times 10^{-6}\text{ L})}{1} = 5.0 \times 10^{-5}\text{ L}$

The answer is expressed using the same number of significant figures as the 50 μL because 1 and 1×10^{-6} are exact numbers by definition.

Learning Check 1.16 Creatinine is a substance found in the blood. An analysis of a blood serum sample detected 1.1 mg of creatinine. Express this amount in grams by using the factor-unit method. Remember, the prefix milli means $\frac{1}{1000}$, so 1 g = 1000 mg.

Example 1.17

One of the fastest-moving nerve impulses in the body travels at a speed of 400 feet per second (ft/s). What is the speed in miles per hour?

Solution

The known quantity is 400 ft/s, and the unit of the unknown quantity is miles per hour (mi/h).

Step 1. $\dfrac{400\ \text{ft}}{\text{s}}$

Step 2. $\dfrac{400\ \text{ft}}{\text{s}} \qquad = \dfrac{\text{mi}}{\text{h}}$

Step 3. $\left(\dfrac{400\ \cancel{\text{ft}}}{\cancel{\text{s}}}\right)\left(\dfrac{1\ \text{mi}}{5280\ \cancel{\text{ft}}}\right)\left(\dfrac{60\ \cancel{\text{s}}}{1\ \cancel{\text{min}}}\right)\left(\dfrac{60\ \cancel{\text{min}}}{1\ \text{h}}\right) = \dfrac{\text{mi}}{\text{h}}$

The factors came from the following well-known numerical relationships: 1 mi = 5280 ft, 1 min = 60 s, 1 h = 60 min. All numbers in these factors are exact numbers based on definitions.

Step 4. $\dfrac{(400)(1)(60)(60)\,\text{mi}}{(5280)(1)(1)\,\text{h}} = 272.7\ \dfrac{\text{mi}}{\text{h}}$

Rounding to three significant figures, the same as in 400 ft/s, gives 273 mi/h.

Learning Check 1.17 A world-class sprinter can run 100 m in 10.0 s. This corresponds to a speed of 10.0 m/s. Convert this speed to miles per hour. Use information from Table 1.3.

1.10 Calculating Percentages

Learning Objective

10. Do calculations involving percentages.

The word *percent* literally means per one hundred. It is the number of specific items in a group of 100 such items. Since items are seldom found in groups of exactly 100, we usually have to calculate the number of specific items that would be in the group if it did contain exactly 100 items. This number is the percentage, and the calculation follows:

$$\text{percent} = \frac{\text{number of specific items}}{\text{total items in the group}} \times 100 \tag{1.5}$$

$$\% = \frac{\text{part}}{\text{total}} \times 100 \tag{1.6}$$

In Equation 1.6, the word *part* is used to represent the number of specific items included in the total.

Example 1.18

A college has 4517 female and 3227 male students enrolled. What percentage of the student body is female?

Solution

The total student body consists of 7744 people, of which 4517 are female.

$$\% \text{ female} = \frac{\text{number of females}}{\text{total number of students}} \times 100$$

$$\% \text{ female} = \frac{4517}{7744} \times 100 = 58.33$$

Rearrangement of Equation 1.6 gives another useful percent relationship. Because $\% = \text{part/total} \times 100$,

$$\text{part} = \frac{(\%)(\text{total})}{100} \tag{1.7}$$

According to Equation 1.7, the number of specific items corresponding to a percentage can be calculated by multiplying the percentage and total, then dividing the product by 100.

Example 1.19

The human body is approximately 70% water by mass. What is the mass of water in a 170-pound (lb) person?

Study Skills 1.1 Help with Calculations

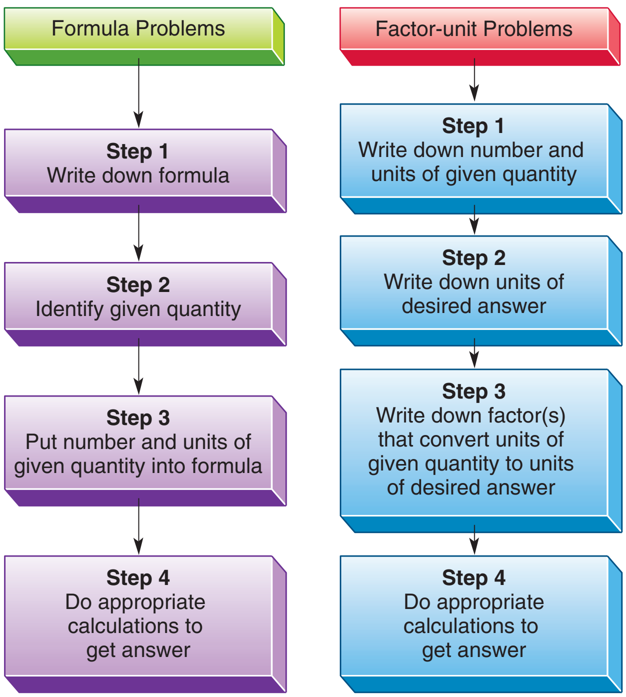
Many students feel uneasy about working chemistry problems that involve the use of mathematics. The uneasiness is often increased if the problem to be solved is a story problem. One tip that will help you solve such problems in this textbook is to remember that almost all of these problems are one of two types: those for which a specific formula applies and those where the factor-unit method is used. When you do homework or take quizzes or examinations and encounter a math-type problem, your first task should be to decide which type of problem it is, formula or factor-unit.

In this chapter, formula-type problems were those that dealt with percentage calculations (see Examples 1.18 and 1.19) and the conversion from one temperature scale to another (see Example 1.7). If you decide a problem is a formula type, Step 1 in solving it is to write down the formula that applies. This important step makes it easier to do the next step because a formula is like a road map that tells you how to proceed from one point to another. Step 2 is to identify the given quantity and put its number and units into the formula. This will leave only the answer missing in most formula-type problems. You can then obtain the answer by doing the appropriate calculations.

In this textbook, the factor-unit method discussed in Section 1.9 is used for most problems that require mathematical calculations. This method simplifies problem solving and should be mastered so it can be used where it applies. The beauty of this method is that it mimics your natural, everyday way of solving problems. This real-life method usually involves identifying where you are, where you want to go, and how to get there. The factor-unit method follows the same pattern: Step 1, identify the given number and its units; Step 2, write down the unit of the desired answer; Step 3, put in factors that will convert the units of the given quantity into the units of the desired answer.

Thus, we see that working story problems is not as difficult a task as it might first appear. A key is to see through all the words and find what is given (number and units). Then look for what is wanted by focusing on key words or phrases like "how much," "what is," and "calculate." Finally, use one of the two methods, formula or factor-unit, to solve the problem.

The steps to follow in solving both types of problems are summarized in the following flow charts.



Solution

In this problem, what we would classify as the part is the mass of water in a person who weighs a total of 170 lb. Substitution into Equation 1.7 gives

$$\text{mass of water} = \frac{(\%)(\text{total})}{100} = \frac{(70)(170\text{ lb})}{100} = 119\text{ lb}$$

We should round our answer to only two significant figures to match the two in the 70%. Do this by using scientific notation; the answer is 1.2×10^2 lb.

Learning Check 1.18

a. A student is saving money to buy a computer that will cost a total of $1200. The student has saved $988. What percentage of the purchase price has been saved?
b. In a chemistry class of 83 students, 90.4% voted not to have the final exam. How many students wanted to take the exam?

1.11 Density

Learning Objective

11. Do calculations involving densities.

A discussion of the density of matter will conclude the topics of this chapter. Density is a physical property of matter, so it can be measured without changing the composition of the sample of matter under investigation. **Density** is the number obtained by dividing the mass of a sample of matter by the volume of the same sample.

density The number given when the mass of a sample of a substance is divided by the volume of the same sample.

$$\text{density} = \frac{\text{mass}}{\text{volume}} \tag{1.8}$$

$$d = \frac{m}{V} \tag{1.9}$$

We see from these equations that once a numerical value has been obtained for the density, two factors are available that relate mass and volume, and these may be used to solve problems.

Example 1.20

The density of iron metal has been determined to be 7.2 g/cm^3.

a. Use the density value to calculate the mass of an iron sample that has a volume of 35.0 cm^3.
b. Use the density value to calculate the volume occupied by 138 g of iron.

Solution

The value of the density tells us that one cubic centimeter (cm^3) of iron has a mass of 7.2 g. This may be written as 7.2 g = 1.0 cm^3, using two significant figures for the volume. This relationship gives two factors that can be used to solve our problems:

$$\frac{7.2\text{ g}}{1.0\text{ cm}^3} \quad \text{and} \quad \frac{1.0\text{ cm}^3}{7.2\text{ g}}$$

a. The sample volume is 35.0 cm^3, and we wish to use a factor to convert this to grams. The first factor given above will work.

$$35.0\ \cancel{\text{cm}^3} \times \frac{7.2\text{ g}}{1.0\ \cancel{\text{cm}^3}} = 252\text{ g (calculator answer)}$$
$$= 2.5 \times 10^2\text{ g (properly rounded answer)}$$

Chemistry and Your Health 1.1

Health Information on the Web

The World Wide Web contains at least 20,000 sites that are related to health issues or health problems or that sell health products. Many of these sites are very good resources for individuals concerned about such topics. Unfortunately however, there are also sites run by scam artists who are simply interested in trying to make money at the expense of gullible or uneducated web surfers. The information provided by these sites is not only useless but might also be dangerous. There have been reports of sites run by individuals claiming to have a "miracle cure" for serious diseases. The sites encourage individuals to stop taking their prescription medication and instead buy the new miracle product. Another characteristic of a site to be avoided is one that claims to have a physician who will diagnose or treat you without requiring you to have a proper examination and consultation. As a general rule, you should use common sense. Another good idea is to find websites that are already linked with organizations you are familiar with or recognize as being legitimate. The following sites are very helpful and contain accurate and complete information.

Healthfinder: *www.healthfinder.gov*
USDA's Food & Nutrition Research Briefs: *www.nal.usda.gov/fnic/usda/fnrb*
National Cancer Institute: *www.cancer.gov*
Centers for Disease Control and Prevention: *www.cdc.gov*
U.S. Department of Health & Human Services: *www.healthfinder.gov*
U.S. National Library of Medicine: *www.nlm.nih.gov/medlineplus/*
New England Journal of Medicine: *www.nejm.org*
Journal of the American Medical Association: *www.ama-assn.org/public/journals/jama/*
American Cancer Society: *www.cancer.org*
Online mental health ("Psych Central"): *www.coil.com/~grohol/*
Mediconsult: *www.mediconsult.com*
American Academy of Family Physicians: *www.familydoctor.org*
Medfusion: an information partnership of medical societies: *www.medfusion.new/ihealth*
Medical Library Association: *www.mlanet.org/resources/userguide.html*
NOAH: New York Online Access to Health: *www.noah-health.org/*

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An enormous amount of health information is available through home computers and the internet.

b. The sample mass is 138 g, and we wish to convert this to cubic centimeters (cm^3). The second factor given above will work.

$$138\ \cancel{g} \times \frac{1.0\ cm^3}{7.2\ \cancel{g}} = 19.17\ cm^3\ \text{(calculator answer)}$$

$$= 19\ cm^3\ \text{(properly rounded answer)}$$

Learning Check 1.19 Aluminum metal has a density of 2.7 g/cm^3.

a. Calculate the mass of an aluminum sample with a volume of 60.0 cm^3.
b. Calculate the volume of an aluminum sample that has a mass of 98.5 g.

For some substances, density is rather easily determined experimentally by direct measurement. The mass of a sample is obtained by weighing the sample. The sample volume can be calculated if the sample is a regular solid such as a cube. If the sample is a liquid or an irregular solid, the volume can be measured by using volumetric apparatus such as those shown in Figure 1.13. Densities of solids are often given in units of g/cm^3, and those of liquids in units of g/mL because of the different ways the volumes are determined. However, according to Table 1.3, 1 cm^3 = 1 mL, so the numerical value is the same regardless of which of the two volume units is used.

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Figure 1.13 Glassware for measuring volumes of liquids. Clockwise from top center: buret, graduated cylinder, syringe, pipet, and volumetric flask.

Example 1.21

a. A hypodermic syringe was used to deliver 5.0 cc (cm^3) of alcohol into an empty container that had a mass of 25.12 g when empty. The container with the alcohol sample weighed 29.08 g. Calculate the density of the alcohol.
b. A cube of copper metal measures 2.00 cm on each edge and weighs 71.36 g. What is the density of the copper sample?
c. According to its owner, a chain necklace is made of pure gold. In order to check this, the chain was weighed and found to have a mass of 19.21 g. The chain was then put into a graduated cylinder that contained 20.8 mL of water. After the chain was put into the cylinder, the water level rose to 21.9 mL. The density of pure gold was looked up in a handbook and found to be 19.2 g/cm^3. Is the chain made of pure gold?

Solution

a. According to Table 1.3, 1 cc (or cm^3) = 1 mL, so the volume is 5.0 mL. The sample mass is equal to the difference between the mass of the container with the sample inside and the mass of the empty container:

$$m = 29.08 \text{ g} - 25.12 \text{ g} = 3.96 \text{ g}$$

The density of the sample is equal to the sample mass divided by the sample volume:

$$d = \frac{m}{V} = \frac{3.96 \text{ g}}{5.0 \text{ mL}} = 0.79 \text{ g/mL (rounded value)}$$

b. The volume of a cube is equal to the product of the three sides:

$$V = (2.00 \text{ cm})(2.00 \text{ cm})(2.00 \text{ cm}) = 8.00 \text{ cm}^3$$

The density is equal to the mass divided by the volume:

$$d = \frac{m}{V} = \frac{71.36 \text{ g}}{8.00 \text{ cm}^3} = 8.92 \text{ g/cm}^3 \text{ (rounded value)}$$

c. The volume of the chain is equal to the difference in the water level in the cylinder with and without the chain present:

$$V = 21.9 \text{ mL} - 20.8 \text{ mL} = 1.1 \text{ mL} = 1.1 \text{ cm}^3$$

The density is equal to the mass divided by the volume:

$$d = \frac{m}{V} = \frac{19.21 \text{ g}}{1.1 \text{ cm}^3} = 17 \text{ g/cm}^3 \text{ (rounded value)}$$

The experimentally determined density is less than the density of pure gold, so the chain is not made of pure gold.

Learning Check 1.20

a. A pipet was used to put a 10.00-mL sample of a liquid into an empty container. The empty container weighed 51.22 g, and the container with the liquid sample weighed 64.93 g. Calculate the density of the liquid in g/mL.
b. A box of small irregular pieces of metal was found in a storage room. It was known that the metal was either nickel or chromium. A 35.66-g sample of the metal was weighed and put into a graduated cylinder that contained 21.2 mL of water (Figure 1.14, left). The water level after the metal was added was 25.2 mL (Figure 1.14, right). Was the metal nickel (density = 8.9 g/cm^3) or chromium (density = 7.2 g/cm^3)?

At the Counter 1.1

Nonprescription Medicines

The important role of pharmacy in the modern practice of medicine is well known. In a simplified but familiar scenario, a sick patient consults a physician, who diagnoses the ailment and writes a prescription for medicine. The patient takes the prescription to a pharmacist, who prepares and packages the medication, which the patient takes according to the directions on the container. However, it is estimated that nearly 40% of the common, everyday health problems of people in the United States are treated without the aid of either a physician or a pharmacist. This is possible because of the availability of an estimated 100 thousand medicinal products that may be purchased by consumers without a prescription. These nonprescription medications are often called over-the-counter (OTC) drugs or medicines. They are available not only in pharmacies, but also in such places as supermarkets and convenience stores.

The range of health problems that can be treated with OTC medicines has grown significantly in recent years as a result of action by the FDA that transferred more than 600 effective prescription-only medicines to nonprescription (OTC) status. However, the widespread availability and use of these products raised significant questions about how best to provide consumers with the adequate directions for proper use that are required by law. In an attempt to improve upon the nonstandardized labeling practices used for OTC medicines, the FDA in 1997 proposed new OTC medicine labeling rules:

1. All wording must be in "plain English."
2. A large, easy-to-read type must be used.
3. Labels must follow a consistent design style.
4. Information must be given in a standardized order.
5. Standardized headings and subheadings must be used.

As is often true, progress of this type does not come without a price tag. The FDA estimated it would cost about $14 million to comply with the new labeling rules. However, the Nonprescription Drug Manufacturers Association estimated implementation costs at $155 million for compliance within the required two years and more than $400 million if packaging changes were required to accommodate the new labeling format. The actual costs have not yet been determined, but consumers will certainly pay a significant part of these costs through increased prices in OTC products.

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The active ingredients in all these products were originally available only by prescription.

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Figure 1.14 Measuring the volume of irregular metal pieces.

Concept Summary

What Is Matter? Matter, the substance of everything, is defined as anything that has mass and occupies space. Mass is a measurement of the amount of matter present in an object. Weight is a measure of the gravitational force pulling on an object.

Objective 1, Exercise 1.2

Properties and Changes. Chemical properties cannot be determined without attempting to change one kind of matter into another. Physical properties can be determined without attempting such composition changes. Any change in matter that is accompanied by a composition change is a chemical change. Physical changes take place without the occurrence of any composition changes.

Objective 2, Exercises 1.8 a & b and 1.10 b & c

A Model of Matter. Scientific models are explanations for observed behavior. The results of many observations led scientists to a model for matter in which all matter is composed of tiny particles. In many substances, these particles are called molecules, and they represent the smallest piece of such substances that is capable of a stable existence. Molecules, in turn, are made up of atoms, which represent the limit of chemical subdivision for matter. The terms *diatomic, triatomic, polyatomic, homoatomic,* and *heteroatomic* are commonly used to describe the atomic composition of molecules.

Objective 3, Exercise 1.12

Classifying Matter. It often simplifies things to classify the items being studied. Some useful categories into which matter can be classified are heterogeneous, homogeneous, solution, pure substance, element, and compound. All matter is either heterogeneous or homogeneous. Heterogeneous matter is a mixture in which the properties and appearance are not uniform. Homogeneous matter is either a mixture of two or more pure substances (and the mixture is called a solution), or it is a pure substance. If it is a pure substance, it is either an element (containing atoms of only one kind) or a compound (containing two or more kinds of atoms).

Objective 4, Exercises 1.18, 1.22, and 1.24

Measurement Units. All measurements are based on standard units that have been agreed on and adopted. The earliest measurements were based on human body dimensions, but the changeable nature of such basic units made the adoption of a worldwide standard desirable.

Objective 5, Exercise 1.28

The Metric System. The metric system of measurement is used by most scientists worldwide and all major nations except the United States. It is a decimal system in which larger and smaller units of a quantity are related by factors of 10. Prefixes are used to designate relationships between the basic unit and larger or smaller units of a quantity.

Objective 6, Exercises 1.30 and 1.40

Large and Small Numbers. Because of difficulties in working with very large or very small numbers in calculations, a system of scientific notation has been devised to represent such numbers. In scientific notation, numbers are represented as products of a nonexponential number and 10 raised to some power. The nonexponential number is always written with the decimal in the standard position (to the right of the first nonzero digit in the number). Numbers written in scientific notation can be manipulated in calculations by following a few rules.

Objective 7, Exercises 1.48 and 1.60

Significant Figures. In measured quantities, the significant figures are the numbers representing the part of the measurement that is certain, plus one number representing an estimate. The maximum number of significant figures possible in a measurement is determined by the design of the measuring device. The results of calculations made using numbers from measurements can be expressed with the proper number of significant figures by following simple rules.

Objective 8, Exercises 1.64 and 1.66

Using Units in Calculations. The factor-unit method for doing calculations is based on a specific set of steps. One crucial step involves the use of factors that are obtained from fixed numerical relationships between quantities. The units of the factor must always cancel the units of the known quantity and generate the units of the unknown or desired quantity.

Objective 9, Exercise 1.82

Calculating Percentages. The word *percent* means per one hundred, and a percentage is literally the number of specific items contained in a group of 100 items. Few items always occur in groups of exactly 100, so a calculation can be done that determines how many specific items would be in a group if the group actually did contain exactly 100 items.

Objective 10, Exercise 1.92

Density. The density of a substance is the number obtained by dividing the mass of a sample by the volume of the same sample. Measured values of density provide two factors that can be used with the factor-unit method to calculate the mass of a substance if the volume is known, or the volume if the mass is known.

Objective 11, Exercise 1.98

Key Terms and Concepts

Atom (1.3)
Basic unit of measurement (1.6)
Chemical changes (1.2)
Chemical properties (1.2)
Compound (1.4)
Density (1.11)
Derived unit of measurement (1.6)
Diatomic molecules (1.3)
Element (1.4)
Exact numbers (1.8)
Factors used in the factor-unit method (1.9)
Heteroatomic molecules (1.3)
Heterogeneous matter (1.4)
Homoatomic molecules (1.3)
Homogeneous matter (1.4)
Mass (1.1)
Matter (1.1)
Mixture (1.4)
Molecule (1.3)
Physical changes (1.2)
Physical properties (1.2)
Polyatomic molecules (1.3)
Pure substance (1.4)
Scientific models (1.3)
Scientific notation (1.7)
Significant figures (1.8)
Solutions (1.4)
Standard position for a decimal (1.7)
Triatomic molecules (1.3)
Weight (1.1)

Key Equations

1. Conversion of temperature readings from one scale to another (Section 1.6)

$$°C = \frac{5}{9}(°F - 32) \qquad \text{Equation 1.1}$$

$$°F = \frac{9}{5}(°C) + 32 \qquad \text{Equation 1.2}$$

$$°C = K - 273 \qquad \text{Equation 1.3}$$

$$K = °C + 273 \qquad \text{Equation 1.4}$$

2. Calculation of percentage (Section 1.10)

$$\text{percent} = \frac{\text{number of specific items}}{\text{total items in the group}} \times 100 \qquad \text{Equation 1.5}$$

$$\% = \frac{\text{part}}{\text{total}} \times 100 \qquad \text{Equation 1.6}$$

3. Calculation of number of items representing a specific percentage of a total (Section 1.10)

$$\text{part} = \frac{(\%)(\text{total})}{100} \qquad \text{Equation 1.7}$$

4. Calculation of density from mass and volume data (Section 1.11)

$$d = \frac{m}{V} \qquad \text{Equation 1.9}$$

Exercises

OWL Interactive versions of these problems are assignable in OWL.

Even-numbered exercises are answered in Appendix B.

Blue-numbered exercises are more challenging.

What Is Matter? (Section 1.1)

1.1 A heavy steel ball is suspended by a thin wire. The ball is hit from the side with a hammer but hardly moves. Describe what you think would happen if this identical experiment were carried out on the moon.

1.2 Explain how the following are related to each other: matter, mass, and weight.

1.3 Tell how you would try to prove to a doubter that air is matter.

1.4 Which of the following do you think is likely to change the most when done on Earth and then on the moon? Carefully explain your reasoning.

- **a.** The distance you can throw a bowling ball through the air.
- **b.** The distance you can roll a bowling ball on a flat, smooth surface.

1.5 The attractive force of gravity for objects near Earth's surface increases as you move toward Earth's center. Suppose you are transported from a deep mine to the top of a tall mountain.

- **a.** How would your mass be changed by the move?
- **b.** How would your weight be changed by the move?

1.6 Earth's rotation causes it to bulge at the equator. How would the weights of people of equal mass differ when one was determined at the equator and one at the North Pole? (See Exercise 1.5.)

Properties and Changes (Section 1.2)

1.7 Classify each of the following as a physical or chemical change, and give at least one observation, fact, or reason to support your answer.

- **a.** A plum ripens
- **b.** Water boils
- **c.** A glass window breaks
- **d.** Food is digested

1.8 Classify each of the following as a physical or chemical change, and give at least one observation, fact, or reason to support your answer.

- **a.** A stick is broken into two pieces
- **b.** A candle burns
- **c.** Rock salt is crushed by a hammer
- **d.** Tree leaves change color in autumn

1.9 Classify each of the following properties as physical or chemical. Explain your reasoning in each case.

- **a.** Iron melts at 1535°C
- **b.** Alcohol is very flammable
- **c.** The metal used in artificial hip-joint implants is not corroded by body fluids
- **d.** A 1-in. cube of aluminum weighs less than a 1-in. cube of lead
- **e.** An antacid tablet neutralizes stomach acid

1.10 Classify each of the following properties as physical or chemical. Explain your reasoning in each case.

a. Mercury metal is a liquid at room temperature

b. Sodium metal reacts vigorously with water

c. Water freezes at 0°C

d. Gold does not rust

e. Chlorophyll molecules are green in color

A Model of Matter (Section 1.3)

1.11 A sample of liquid alcohol is frozen to a solid, then allowed to melt back to a liquid. Have the alcohol molecules been changed by the process? Explain your answer.

1.12 Succinic acid, a white solid that melts at 182°C, is heated gently, and a gas is given off. After the gas evolution stops, a white solid remains that melts at a temperature different from 182°C.

a. Have the succinic acid molecules been changed by the process? Explain your answer.

b. Is the white solid that remains after heating still succinic acid? Explain your answer.

c. In terms of the number of atoms contained, how do you think the size of succinic acid molecules compares with the size of the molecules of the white solid produced by this process? Explain your answer.

d. Classify molecules of succinic acid by using the term *homoatomic* or *heteroatomic*. Explain your reasoning.

1.13 A sample of solid elemental phosphorus that is deep red in color is burned. While the phosphorus is burning, a white smoke is produced that is actually a finely divided solid that is collected.

a. Have the molecules of phosphorus been changed by the process of burning? Explain your answer.

b. Is the collected white solid a different substance from the phosphorus? Explain your answer.

c. In terms of the number of atoms contained, how do you think the size of the molecules of the white solid compares with the size of the molecules of phosphorus? Explain your answer.

d. Classify molecules of the collected white solid using the term *homoatomic* or *heteroatomic*. Explain your reasoning.

1.14 Oxygen gas and solid carbon are both made up of homoatomic molecules. The two react to form a single substance, carbon dioxide. Use the term *homoatomic or heteroatomic* to classify molecules of carbon dioxide. Explain your reasoning.

1.15 Under appropriate conditions, hydrogen peroxide can be changed to water and oxygen gas. Use the term *homoatomic* or *heteroatomic* to classify molecules of hydrogen peroxide. Explain your reasoning.

1.16 Water can be decomposed to hydrogen gas and oxygen gas by passing electricity through it. Use the term *homoatomic* or *heteroatomic* to classify molecules of water. Explain your reasoning.

1.17 Methane gas, a component of natural gas, is burned in pure oxygen. The only products of the process are water and carbon dioxide. Use the term *homoatomic* or *heteroatomic* to classify molecules of methane. Explain your reasoning.

Classifying Matter (Section 1.4)

1.18 Classify each pure substance represented below by a capital letter as an element or a compound. Indicate when such a classification cannot be made, and explain why

a. Substance A is composed of heteroatomic molecules

b. Substance D is composed of homoatomic molecules

c. Substance E is changed into substances G and J when it is heated

1.19 Classify each pure substance represented below by a capital letter as an element or a compound. Indicate when such a classification cannot be made, and explain why.

a. Two elements when mixed combine to form only substance L

b. An element and a compound when mixed form substances M and Q

c. Substance X is not changed by heating it

1.20 Consider the following experiments, and answer the questions pertaining to classification:

a. A pure substance R is heated, cooled, put under pressure, and exposed to light but does not change into anything else. What can be said about classifying R as an element or a compound? Explain your reasoning.

b. Upon heating, solid pure substance T gives off a gas and leaves another solid behind. What can be said about classifying T as an element or compound? Explain your reasoning.

c. What can be said about classifying the solid left in part b as an element or compound? Explain your reasoning.

1.21 Early scientists incorrectly classified calcium oxide (lime) as an element for a number of years. Discuss one or more reasons why you think they might have done this.

1.22 Classify each of the following as homogeneous or heterogeneous:

a. a pure gold chain

b. liquid eyedrops

c. chunky peanut butter

d. a slice of watermelon

e. cooking oil

f. Italian salad dressing

g. window glass

1.23 Classify each of the following as homogeneous or heterogeneous:

a. muddy flood water

b. gelatin dessert

c. normal urine

d. smog-filled air

e. an apple

f. mouthwash

g. petroleum jelly

1.24 Classify as pure substance or solution each of the materials of Exercise 1.22 that you classified as homogeneous.

1.25 Classify as pure substance or solution each of the materials of Exercise 1.23 that you classified as homogeneous.

Measurement Units (Section 1.5)

1.26 Briefly discuss why a system of measurement units is an important part of our modern society.

1.27 In the distant past, 1 in. was defined as the length resulting from laying a specific number of grain kernels such as corn in a row. Discuss the disadvantages of such a system.

1.28 An old British unit used to express weight is a stone. It is equal to 14 1b. What sort of weighings might be expressed in stones? Suggest some standard that might have been used to establish the unit.

The Metric System (Section 1.6)

1.29 Which of the following quantities are expressed in metric units?

- **a.** The amount of aspirin in a tablet: 5 grains
- **b.** The distance between two cities: 55 km
- **c.** The internal displacement of an auto engine: 5 L
- **d.** The time for a race: 4 min, 5.2 s
- **e.** The area of a field: 3.6 acres
- **f.** The temperature on a hot day: 104°F

1.30 Which of the following are expressed in metric units?

- **a.** Normal body temperature: 37°C
- **b.** The amount of soft drink in a bottle: 2 L
- **c.** The height of a ceiling in a room: 8.0 ft
- **d.** The amount of aspirin in a tablet: 81 mg
- **e.** The volume of a cooking pot: 4 qt
- **f.** The time for a short race to be won: 10.2 s

1.31 Referring to Table 1.3, suggest an appropriate metric system unit for each nonmetric unit in Exercise 1.29.

1.32 Referring to Table 1.3, suggest an appropriate metric system unit for each nonmetric unit in Exercise 1.30.

1.33 Referring only to Table 1.2, answer the following questions:

- **a.** A computer has 12 megabytes of memory storage. How many bytes of storage is this?
- **b.** A 10-km race is 6.2 mi long. How many meters long is it?
- **c.** A chemical balance can detect a mass as small as 0.1 mg. What is this detection limit in grams?
- **d.** A micrometer is a device used to measure small lengths. If it lives up to its name, what is the smallest metric length that could be measured using a micrometer?

1.34 Referring only to Table 1.2, answer the following questions:

- **a.** Devices are available that allow liquid volumes as small as one microliter (mL) to be measured. How many microliters would be contained in 1.00 liter?
- **b.** Electrical power is often measured in kilowatts. How many watts would equal 75 kilowatts?
- **c.** Ultrasound is sound of such high frequency that it cannot be heard. The frequency is measured in hertz (vibrations per second). How many hertz correspond to 15 megahertz?
- **d.** A chlorine atom has a diameter of 200 picometers. How many meters is this diameter?

1.35 One inch is approximately equal to 2.54 cm. Express this length in millimeters and meters.

1.36 Cookbooks are going metric. In such books, 1 cup is equal to 240 mL. Express 1 cup in terms of liters and cubic centimeters.

1.37 The so-called metric mile race is 1500 m long. What is its length in miles?

1.38 The shotput used by female track and field athletes has a mass of 4.0 kg. What would be the weight of such a shotput in pounds?

1.39 Referring to Table 1.3, answer the following questions:

- **a.** Which is larger, a liter or a quart?
- **b.** How many milliliters are in a 12.0-fl-oz soft drink?
- **c.** Which is larger, a BTU or a kilocalorie?

1.40 Referring to Table 1.3, answer the following questions:

- **a.** Approximately how many inches longer is a meter stick than a yardstick?
- **b.** A temperature increases by 65°C. How many kelvins would this increase be?
- **c.** You have a 5-lb bag of sugar. Approximately how many kilograms of sugar do you have?

1.41 Do the following, using appropriate values from Table 1.3:

- **a.** Calculate the area in square meters of a circular skating rink that has a 12.5-m radius. For a circle, the area (A) is related to the radius (r) by $A = \pi r^2$, where $\pi = 3.14$.
- **b.** Calculate the floor area and volume of a rectangular room that is 5.0 m long, 2.8 m wide, and 2.1 m high. Express your answers in square meters and cubic meters (meters cubed).
- **c.** A model sailboat has a triangular sail that is 25 cm high (h) and has a base (b) of 15 cm. Calculate the area (A) of the sail in square centimeters. $A = \dfrac{(b)(h)}{2}$ for a triangle.

1.42 Using appropriate values from Table 1.3, answer the following questions:

- **a.** One kilogram of water has a volume of 1.0 dm^3. What is the mass of 1.0 cm^3 of water?
- **b.** One quart is 32 fl oz. How many fluid ounces are contained in a 2.0-L bottle of soft drink?
- **c.** Approximately how many milligrams of aspirin are contained in a 5-grain tablet?

1.43 The weather report says the temperature is 23°F. What is this temperature on the Celsius scale? On the Kelvin scale?

1.44 Recall from Chemistry Around Us 1.3 that a normal body temperature might be as low as 36.1°C in the morning and as high as 37.2°C at bedtime. What are these temperatures on the Fahrenheit scale?

1.45 One pound of body fat releases approximately 4500 kcal of energy when it is metabolized. How many joules of energy is this? How many BTUs?

Large and Small Numbers (Section 1.7)

1.46 Which of the following numbers are written using scientific notation correctly? For those that are not, explain what is wrong.

a. 02.7×10^{-3}

b. 4.1×10^{2}

c. 71.9×10^{-6}

d. 10^{3}

e. $.0405 \times 10^{-2}$

f. 0.119

1.47 Which of the following numbers are written using scientific notation correctly? For those that are not, explain what is wrong.

a. 3.6×10^{-5}

b. 3.9^{-2}

c. 295×10^{3}

d. 0.05×10^{-3}

e. 10^{-4}

f. 13.1×10^{6}

1.48 Write each of the following numbers using scientific notation:

a. 14 thousand

b. 365

c. 0.00204

d. 461.8

e. 0.00100

d. 9.11 hundred

1.49 Write each of the following numbers using scientific notation:

a. 1.02 thousand

b. 0.07102

c. 3050

d. 1.51 million

e. three thousand

f. 31.05

1.50 The speed of light is about 186 thousand mi/s, or 1100 million km/h. Write both numbers using scientific notation.

1.51 A sheet of paper is 0.0106 cm, or 0.0042 in., thick. Write both numbers using scientific notation.

1.52 A single copper atom has a mass of 1.05×10^{-22} g. Write this number in a decimal form without using scientific notation.

1.53 In 2.0 g of hydrogen gas, there are approximately 6.02×10^{23} hydrogen molecules. Write this number without using scientific notation.

1.54 Do the following multiplications, and express each answer using scientific notation:

a. $(8.2 \times 10^{-3})(1.1 \times 10^{-2})$

b. $(2.7 \times 10^{2})(5.1 \times 10^{4})$

c. $(3.3 \times 10^{-4})(2.3 \times 10^{2})$

d. $(9.2 \times 10^{-4})(2.1 \times 10^{4})$

e. $(4.3 \times 10^{6})(6.1 \times 10^{5})$

1.55 Do the following multiplications, and express each answer using scientific notation:

a. $(5.0 \times 10^{-5})(7.1 \times 10^{-2})$

b. $(6.3 \times 10^{-9})(3.7 \times 10^{7})$

c. $(3.2 \times 10^{-4})(1.0 \times 10^{4})$

d. $(2.7 \times 10^{2})(3.8 \times 10^{4})$

e. $(7.1 \times 10^{4})(6.9 \times 10^{7})$

1.56 Express each of the following numbers using scientific notation, then carry out the multiplication. Express each answer using scientific notation.

a. (144)(0.0876)

b. (751)(106)

c. (0.0422)(0.00119)

d. (128,000)(0.0000316)

1.57 Express each of the following numbers using scientific notation, then carry out the multiplication. Express each answer using scientific notation.

a. (835)(0.00245)

b. (300)(245)

c. (68.3)(421)

d. (32.9)(0.115)

1.58 Do the following divisions, and express each answer using scientific notation:

a. $\dfrac{3.1 \times 10^{-3}}{1.2 \times 10^{2}}$

b. $\dfrac{7.9 \times 10^{4}}{3.6 \times 10^{2}}$

c. $\dfrac{4.7 \times 10^{-1}}{7.4 \times 10^{2}}$

d. $\dfrac{0.00229}{3.16}$

e. $\dfrac{119}{3.8 \times 10^{3}}$

1.59 Do the following divisions, and express each answer using scientific notation:

a. $\dfrac{154}{2.82}$

b. $\dfrac{7.6 \times 10^{2}}{5.5 \times 10^{1}}$

c. $\dfrac{9.1 \times 10^{-5}}{3.4 \times 10^{-2}}$

d. $\dfrac{7.6 \times 10^{3}}{3.8 \times 10^{-4}}$

e. $\dfrac{3.8 \times 10^{-3}}{4.7 \times 10^{4}}$

1.60 Do the following calculations, and express each answer using scientific notation:

a. $\dfrac{(5.3)(0.22)}{(6.1)(1.1)}$

b. $\dfrac{(3.8 \times 10^{-4})(1.7 \times 10^{-2})}{6.3 \times 10^{3}}$

c. $\dfrac{4.8 \times 10^6}{(7.4 \times 10^3)(2.5 \times 10^{-4})}$

d. $\dfrac{5.6}{(0.022)(109)}$

e. $\dfrac{(4.6 \times 10^{-3})(2.3 \times 10^2)}{(7.4 \times 10^{-4})(9.4 \times 10^{-5})}$

1.61 Do the following calculations, and express each answer using scientific notation:

a. $\dfrac{(7.4 \times 10^{-3})(1.3 \times 10^4)}{5.5 \times 10^{-2}}$

b. $\dfrac{6.4 \times 10^5}{(8.8 \times 10^3)(1.9 \times 10^{-4}}$

c. $\dfrac{(6.4 \times 10^{-2})(1.1 \times 10^{-8})}{(2.7 \times 10^{-4})(3.4 \times 10^{-4})}$

d. $\dfrac{(963)(1.03)}{(0.555)(412)}$

e. $\dfrac{1.15}{(0.12)(0.73)}$

Significant Figures (Section 1.8)

1.62 Indicate to what decimal position readings should be estimated and recorded (nearest 0.1, .01, etc.) for measurements made with the following devices:

a. A ruler with a smallest scale marking of 0.1 cm

b. A measuring telescope with a smallest scale marking of 0.1 mm

c. A protractor with a smallest scale marking of 1°

d. A tire pressure gauge with a smallest scale marking of 1 lb/in$^{2.}$

1.63 Indicate to what decimal position readings should be estimated and recorded (nearest 0.1, .01, etc.) for measurements made with the following devices:

a. A buret with a smallest scale marking of 0.1 mL

b. A graduated cylinder with a smallest scale marking of 1 mL

c. A thermometer with a smallest scale marking of 0.1°C

d. A barometer with a smallest scale marking of 1 torr

1.64 Write the following measured quantities as you would record them, using the correct number of significant figures based on the device used to make the measurement:

a. Exactly 6 mL of water measured with a graduated cylinder that has a smallest scale marking of 0.1 mL.

b. A temperature that appears to be exactly 37 degrees using a thermometer with a smallest scale marking of 1°C.

c. A time of exactly nine seconds measured with a stopwatch that has a smallest scale marking of 0.1 second.

d. Fifteen and one-half degrees measured with a protractor that has 1-degree scale markings.

1.65 Write the following measured quantities as you would record them, using the correct number of significant figures based on the device used to make the measurements.

a. A length of two and one-half centimeters measured with a measuring telescope with a smallest scale marking of 0.1 mm.

b. An initial reading of exactly 0 for a buret with a smallest scale marking of 0.1 mL.

c. A length of four and one-half centimeters measured with a ruler that has a smallest scale marking of 0.1 cm.

d. An atmospheric pressure of exactly 690 torr measured with a barometer that has a smallest scale marking of 1 torr.

1.66 In each of the following, identify the measured numbers and exact numbers. Do the indicated calculation, and write your answer using the correct number of significant figures.

a. A bag of potatoes is found to weigh 5.06 lb. The bag contains 16 potatoes. Calculate the weight of an average potato.

b. The foul-shooting percentages for the five starting players of a women's basketball team are 71.2%, 66.9%, 74.1%, 80.9% and 63.6%. What is the average shooting percentage of the five players?

1.67 In each of the following, identify the measured numbers and exact numbers. Do the indicated calculations, and write your answer using the correct number of significant figures.

a. An individual has a job of counting the number of people who enter a store between 1 P.M. and 2 P.M. each day for 5 days. The counts were 19, 24, 17, 31, and 40. What was the average number of people entering the store per day for the 5-day period?

b. The starting five members of a women's basketball team have the following heights: 6′9″, 5′8″, 5′6″, 5′1″, and 4′11″. What is the average height of the starting five?

1.68 Determine the number of significant figures in each of the following:

a. 0.0400

b. 309

c. 4.006

d. 4.4×10^{-3}

e. 1.002

f. 255.02

1.69 Determine the number of significant figures in each of the following:

a. 132.0

b. 2.00×10^3

c. 0.0004

d. 4796

e. 0.00200

f. 1769.0

170. Do the following calculations and use the correct number of significant figures in your answers. Assume all numbers are the results of measurements.

a. (3.71)(1.4)

b. (0.0851)(1.2262)

c. $\dfrac{(0.1432)(2.81)}{(0.7762)}$

d. $(3.3 \times 10^4)(3.09 \times 10^{-3})$

e. $\dfrac{(760)(2.00)}{6.02 \times 10^{20}}$

171. Do the following calculations and use the correct number of significant figures in your answers. Assume all numbers are the results of measurements.

a. (4.09)(3.0)

b. $\frac{(3.192 \times 10^6)(0.0041)}{105}$

c. $\frac{(19.3)(100)}{1000}$

d. $(1.02 \times 10^{-21})(1.1 \times 10^9)^2$

e. $\frac{(251)(3.1 \times 10^{-1})}{(24)(3.0)}$

1.72 Do the following calculations and use the correct number of significant figures in your answers. Assume all numbers are the results of measurements.

a. 0.208 + 4.9 + 1.11

b. 228 + 0.999 + 1.02

c. 8.543 − 7.954

d. $(3.2 \times 10^{-2}) + (5.5 \times 10^{-1})$
(HINT: Write in decimal form first, then add.)

e. 336.86 − 309.11

f. 21.66 − 0.02387

1.73 Do the following calculations and use the correct number of significant figures in your answers. Assume all numbers are the results of measurements.

a. 2.1 + 5.07 + 0.119

b. 0.051 + 8.11 + 0.02

c. 4.337 − 3.211

d. $(2.93 \times 10^{-1}) + (6.2 \times 10^{-2})$
(HINT: Write in decimal form first, then add.)

e. 471.19 − 365.09

f. 17.76 − 0.0479

1.74 Do the following calculations and use the correct number of significant figures in your answers. Assume all numbers are the results of measurements. In calculations involving both addition/subtraction and multiplication/division, it is usually better to do additions/subtractions first.

a. $\frac{(0.0267 + 0.0019)(4.626)}{28.7794}$

b. $\frac{212.6 - 21.88}{86.37}$

c. $\frac{27.99 - 18.07}{4.63 - 0.88}$

d. $\frac{18.87}{2.46} - \frac{18.07}{0.88}$
(HINT: Do divisions first, then subtract.)

e. $\frac{(8.46 - 2.09)(0.51 + 0.22)}{(3.74 + 0.07)(0.16 + 0.2)}$

f. $\frac{12.06 - 11.84}{0.271}$

1.75 Do the following calculations and use the correct number of significant figures in your answers. Assume all numbers are the results of measurements. In calculations involving both addition/subtraction and multiplication/division, it is usually better to do additions/subtractions first.

a. $\frac{132.15 - 32.16}{87.55}$

b. $\frac{(0.0844 + 0.1021)(7.174)}{19.1101}$

c. $\frac{(2.78 - 0.68)(0.42 + 0.4)}{(1.058 + 0.06)(0.22 + 0.2)}$

d. $\frac{27.65 - 21.71}{4.97 - 0.36}$

e. $\frac{12.47}{6.97} - \frac{203.4}{201.8}$
(HINT: Do divisions first, then subtract.)

f. $\frac{19.37 - 18.49}{0.822}$

1.76 The following measurements were obtained for the length and width of a series of rectangles. Each measurement was made using a ruler with a smallest scale marking of 0.1 cm.

Black rectangle:	$l = 12.00$ cm, $w = 10.40$ cm
Red rectangle:	$l = 20.20$ cm, $w = 2.42$ cm
Green rectangle:	$l = 3.18$ cm, $w = 2.55$ cm
Orange rectangle:	$l = 13.22$ cm, $w = 0.68$ cm

a. Calculate the area (length × width) and perimeter (sum of all four sides) for each rectangle and express your results in square centimeters and centimeters, respectively, and give the correct number of significant figures in the result.

b. Change all measured values to meters and then calculate the area and perimeter of each rectangle. Express your answers in square meters and meters, respectively, and give the correct number of significant figures.

c. Does changing the units used change the number of significant figures in the answers?

Using Units in Calculations (Section 1.9)

1.77 Determine a single factor derived from Table 1.3 that could be used as a multiplier to make each of the following conversions:

a. 4 yd to meters

b. 125,000 BTU to kilocalories

c. 400 mm to inches

d. 200 cm to inches

1.78 Determine a single factor from Table 1.3 that could be used as a multiplier to make each of the following conversions:

a. 20 mg to grains

b. 350 mL to fl oz

c. 4 qt to liters

d. 5 yd to meters

1.79 Obtain a factor from Table 1.3 and calculate the number of liters in 1.00 gal (4 qt) by using the factor-unit method of calculation.

1.80 A marathon race is about 26 miles. Obtain a factor from Table 1.3 and use the factor-unit method to calculate the distance of a marathon in kilometers.

1.81 A metric cookbook calls for a baking temperature of 200°C. Your oven settings are in degrees Fahrenheit. What Fahrenheit setting should you use?

1.82 A metric cookbook calls for 250 mL of milk. Your measuring cup is in English units. About how many cups of milk should you use? (NOTE: You will need two factors, one from Table 1.3 and one from the fact that 1 cup = 8 fl oz.)

1.83 An Olympic competitor threw the javelin 96.33 m. What is this distance in feet?

1.84 You have a 40-lb baggage limit for a transatlantic flight. When your baggage is put on the scale, you think you are within the limits because it reads 18.0. But then you realize that weight is in kilograms. Do a calculation to determine whether your baggage is overweight.

1.85 You need 3.00 lb of meat that sells for $3.41/lb (i.e., 1 lb = $3.41). Use this price to determine a factor to calculate the cost of the meat you need using the factor-unit method.

1.86 During a glucose tolerance test, the serum glucose concentration of a patient was found to be 131 mg/dL. Convert the concentration to grams per liter.

Calculating Percentages (Section 1.10)

1.87 Retirement age is 65 years in many companies. What percentage of the way from birth to retirement is a 45-year-old person?

1.88 A salesperson made a sale of $467.80 and received a commission of $32.75. What percent commission was paid?

1.89 After drying, 140 lb of grapes yields 28 lb of raisins. What percentage of the grapes' mass was lost during the drying process?

1.90 The recommended daily intake of thiamin is 1.4 mg for a male adult. Suppose such a person takes in only 1.0 mg/day. What percentage of the recommended intake is he receiving?

1.91 The recommended daily caloric intake for a 20-year-old woman is 2000. How many Calories should her breakfast contain if she wants it to be 45% of her recommended daily total?

1.92 Immunoglobulin antibodies occur in five forms. A sample of serum is analyzed with the following results. Calculate the percentage of total immunoglobulin represented by each type.

Type:	IgG	IgA	IgM	IgD	IgE
Amount (mg):	987.1	213.3	99.7	14.4	0.1

Density (Section 1.11)

1.93 Calculate the density of the following materials for which the mass and volume of samples have been measured. Express the density of liquids in g/mL, the density of solids in g/cm^3, and the density of gases in g/L.

a. 250 mL of liquid mercury metal (Hg) has a mass of 3400 g.

b. 500 mL of concentrated liquid sulfuric acid (H_2SO_4) has a mass of 925 g.

c. 5.00 L of oxygen gas has a mass of 7.15 g.

d. A 200-cm^3 block of magnesium metal (Mg) has a mass of 350 g.

1.94 Calculate the density of the following materials for which the mass and volume of samples have been measured. Express the density of liquids in g/mL, the density of solids in g/cm^3, and the density of gases in g/L.

a. A 50.0-mL sample of liquid acetone has a mass of 39.6 g.

b. A 1.00-cup (236-mL) sample of homogenized milk has a mass of 243 g.

c. 20.0 L of dry carbon dioxide gas (CO_2) has a mass of 39.54 g.

d. A 25.0-cm^3 block of nickel metal (Ni) has a mass of 222.5 g.

1.95 Calculate the volume and density of a rectangular block of platinum metal (Pt) with edges of 7.50 cm, 10.9 cm, and 3.00 cm. The block weighs 5273 g.

1.96 Calculate the volume and density of a cube of lead metal (Pb) that has a mass of 718.3 g and has edges that measure 3.98 cm.

1.97 The volume of an irregularly shaped solid can be determined by immersing the solid in a liquid and measuring the volume of liquid displaced. Find the volume and density of the following:

a. An irregular piece of the mineral quartz is found to weigh 12.4 g. It is then placed into a graduated cylinder that contains some water. The quartz does not float. The water in the cylinder was at a level of 25.2 mL before the quartz was added and at 29.9 mL afterward.

b. The volume of a sample of lead shot is determined using a graduated cylinder, as in part (a). The cylinder readings are 16.3 mL before the shot is added and 21.7 mL after. The sample of shot weighs 61.0 g.

c. A sample of coarse rock salt is found to have a mass of 11.7 g. The volume of the sample is determined by the graduated-cylinder method described in (a), but kerosene is substituted for water because the salt will not dissolve in kerosene. The cylinder readings are 20.7 mL before adding the salt and 26.1 mL after.

1.98 The density of ether is 0.736 g/mL. What is the volume in mL of 280 g of ether?

1.99 Calculate the mass in grams of 125 mL of chloroform (d = 1.49 g/mL).

Additional Exercises

1.100 Do the following metric system conversions by changing only the power of 10. For example, convert 2.5 L to mL: 2.5 L = 2.5×10^3 mL

a. Convert 4.5 km to mm

b. Convert 6.0×10^6 mg to g

c. Convert 9.86×10^{15} m to km

d. Convert 1.91×10^{-4} kg to mg

e. Convert 5.0 ng to mg

1.101 A single water molecule has a mass of 2.99×10^{-23} g. Each molecule contains two hydrogen atoms that together make up 11.2% of the mass of the water molecule. What is the mass in grams of a single hydrogen atom?

1.102 It has been found that fat makes up 14% of a 170-lb person's body weight. One pound of body fat provides 4500 kcal of energy when it is metabolized. The person requires 2000 kcal of energy per day to survive. Assume all the energy needed for survival comes from the metabolism of body fat, and calculate the number of days the person could survive without eating before depleting the entire body fat reserve.

1.103 Cooking oil has a density of 0.812 g/mL. What is the mass in grams of 1.00 quart of cooking oil? Use Table 1.3 for any necessary factors.

1.104 A 175-lb patient is to undergo surgery and will be given an anesthetic intravenously. The safe dosage of anesthetic is 12 mg/kg of body weight. Determine the maximum dose of anesthetic in mg that should be used.

1.105 At 4.0°C, pure water has a density of 1.00 g/mL. At 60.0°C, the density is 0.98 g/mL. Calculate the volume in mL of 1.00 g of water at each temperature, and then calculate the percentage increase in volume that occurs as water is heated from 4.0°C to 60.0°C.

Allied Health Exam Connection

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1.106 Which of the following properties is considered a physical property?

a. flammability
b. boiling point
c. reactivity
d. osmolarity

1.107 Which of the following depicts a chemical process?

a. Helium is combined with neon
b. Iron forms rust
c. Water causes soil erosion
d. Ice melts

1.108 Which of the following is a mixture?

a. sodium chloride
b. rice and beans
c. magnesium sulfate
d. water

1.109 If it is 90°F, approximately what temperature is it on the Celsius scale?

a. 18°C
b. 32°C
c. 58°C
d. 104°C

1.110 If the temperature is 25°C, what is the temperature in °F?

a. 25°F
b. 298°F
c. 0°F
d. 77°F

1.111 The correct formula for converting Fahrenheit to Celsius is given by: $°C = \frac{5}{9}(°F - 32)$. Convert 72°F into temperature in Celsius.

a. 72°C
b. 40°C
c. 25°C
d. 22.2°C

1.112 In degrees Kelvin, the freezing pointing of water is:

a. −273°
b. 0°
c. 100°
d. 273°

1.113 The number of degrees on the Fahrenheit thermometer between the freezing point and the boiling point of water is:

a. 100 degrees
b. 180 degrees
c. 212 degrees
d. 273 degrees

1.114 A calorie is a form of:

a. light
b. heat
c. darkness
d. sound

1.115 How many millimeters are there in one centimeter?

a. 10,000
b. 1,000
c. 100
d. 10

1.116 Convert 4.50×10^2 nm into ____ pm.

a. 4.50×10^2 pm
b. 4.50×10^{-2} pm
c. 4.50×10^{11} pm
d. 4.50×10^5 pm

1.117 One millimeter contains how many μm?

a. 10
b. 100
c. 1,000
d. 10,000

1.118 Convert 4.50×10^2 nm into ___ m.

a. 4.50×10^2 m

b. 4.50×10^{11} m

c. 4.50×10^{-7} m

d. 4.50×10^8 m

1.119 The quantity 6,180 meters can be rewritten as:

a. 6.180×10^3 meter.

b. 6,180 kilometer.

c. $6{,}180 \times 10^3$ meter.

d. 180×10^3 meter.

1.120 The number 1,000,000 is what power of 10?

a. 10^{-6}

b. 10^6

c. 1^6

d. 0.000001

1.121 What exponent or power of ten would you use to express how many meters are in a kilometer?

a. 10^5

b. 10^3

c. 10^4

d. 10^2

1.122 Express 0.05620 in exponential notation.

a. 0.057×10^{-3}

b. 57×10^{-3}

c. 563×10^{-4}

d. 5.62×10^{-2}

1.123 Write the correct answer (correct number of significant figures) for the following calculation: $(27 + 93) \times 5.1558$

a. 618.697

b. 618.7

c. 619

d. 618.6970

1.124 The oxidation of 1 gram of CHO (carbohydrate) produces 4 calories. How much CHO must be oxidized in the body to produce 36 calories?

a. 4 grams

b. 7 grams

c. 9 grams

d. 12 grams

1.125 The percentage of oxygen by weight in $Al_2(SO_4)_3$ (atomic weights: Al = 27, S = 32, O = 16) is approximately:

a. 19

b. 21

c. 56

d. 92

1.126 A 10-percent solution of glucose will contain:

a. 1 gram of glucose per 1,000 milliliters of solution

b. 1 gram of glucose per 100 milliliters of solution

c. 1 gram of glucose per 10 microliters of solution

d. 10 grams of glucose per 100 milliliters of solution

1.127 The density of gold (Au) is 19.3g/cm^3 and that of iron (Fe) is 7.9g/cm^3. A comparison of the volumes (V) of 50 gram samples of each metal would show that:

a. $V_{Au} = V_{Fe}$

b. $V_{Au} < V_{Fe}$

c. $V_{Au} > V_{Fe}$

d. There is no predictable relationship between volumes.

Chemistry for Thought

1.128 The following pairs of substances represent heterogeneous mixtures. For each pair, describe the steps you would follow to separate the components and collect them.

a. wood sawdust and sand

b. sugar and sand

c. iron filings and sand

d. sand soaked with oil

1.129 Explain why a bathroom mirror becomes foggy when someone takes a hot shower. Classify any changes that occur as physical or chemical.

1.130 A 20-year-old student was weighed and found to have a mass of 44.5 kg. She converted this to pounds and got an answer of 20.2 lb. Describe the mistake she probably made in doing the calculation.

1.131 Liquid mercury metal freezes to a solid at a temperature of −38.9°C. Suppose you want to measure a temperature that is at least as low as −45°C. Can you use a mercury thermometer? If not, propose a way to make the measurement.

1.132 Answer the question contained in Figure 1.3. How does hang gliding confirm that air is an example of matter?

1.133 Show how the factor-unit method can be used to prepare an oatmeal breakfast for 27 guests at a family reunion. The directions on the oatmeal box say that 1 cup of dry oatmeal makes 3 servings.

1.134 A chemist is brought a small solid figurine. The owner wants to know if it is made of silver but doesn't want it damaged during the analysis. The chemist decides to determine the density, knowing that silver has a density of 10.5 g/mL. The figurine is put into a graduated cylinder that contains 32.6 mL of water. The reading while the figurine is in the water is 60.1 mL. The mass of the figurine is 240.8 g. Is the figurine made of silver? Explain your reasoning.

1.135 Refer to Chemistry Around Us 1.2, then check the labels on your toiletries and see if you can identify one product that is regulated as both a drug and a cosmetic.

1.136 Refer to Figure 1.6, then use the model of matter described in Section 1.3 to propose an explanation for the following observation. When two teaspoons of sugar are dissolved in a small glass of water, the volume of the resulting solution is not significantly larger than the original volume of the water.

2 Atoms and Molecules

Learning Objectives

When you have completed your study of this chapter, you should be able to:

1. Use symbols for chemical elements to write formulas for chemical compounds. **(Section 2.1)**
2. Identify the characteristics of protons, neutrons, and electrons. **(Section 2.2)**
3. Use the concepts of atomic number and mass number to determine the number of subatomic particles in isotopes and to write correct symbols for isotopes. **(Section 2.3)**
4. Use atomic weights of the elements to calculate molecular weights of compounds. **(Section 2.4)**
5. Use isotope percent abundances and masses to calculate atomic weights of elements. **(Section 2.5)**
6. Use the mole concept to obtain relationships between number of moles, number of grams, and number of atoms for elements, and use those relationships to obtain factors for use in factor-unit calculations. **(Section 2.6)**
7. Use the mole concept and molecular formulas to obtain relationships between number of moles, number of grams, and number of atoms or molecules for compounds, and use those relationships to obtain factors for use in factor-unit calculations. **(Section 2.7)**

OWL Online homework for this chapter may be assigned in OWL.

As a result of advances in medical technologies, many specialties have been created in the health care industry. Here a **nuclear medicine technologist** works with a patient who is undergoing a body scan. In order to understand the scanning process, the technologist must also have an understanding of chemical symbols, formulas, and isotopes, which are topics introduced in this chapter.

We introduced some fundamental ideas about matter, atoms, molecules, measurements, and calculations in Chapter 1. In this chapter, these ideas are applied, the mole is defined, and the quantitative nature of chemistry becomes more apparent. A system of symbols is introduced that simplifies the way atoms and molecules are represented.

2.1 Symbols and Formulas

Learning Objective

1. Use symbols for chemical elements to write formulas for chemical compounds.

In Chapter 1, we defined elements as homogeneous pure substances made up of identical atoms. At least 115 different elements are known to exist. This leads to the conclusion that a minimum of 115 different kinds of atoms exist. Eighty-eight of the elements are naturally occurring and therefore are found in Earth's crust, oceans, or atmosphere. The others are synthetic elements produced in the laboratory.

Each element can be characterized and identified by its unique set of physical and chemical properties, but it would be very cumbersome to list all these properties each time a specific element was discussed. For this reason, each element has been assigned a unique name and symbol. Many elements have been named by their discoverer, and as a result the names are varied. Some are based on elemental properties, others come from the names of famous scientists or places, while others are derived from the names of astronomical bodies or mythological characters.

An **elemental symbol** is based on the element's name and consists of a single capital letter or a capital letter followed by a lowercase letter. The symbols for 11 elements are based on the element's name in Latin or German. An elemental symbol is sometimes used to represent an element in a general way or to represent a single atom of an element. Table 2.1 lists the elements whose names and symbols have been agreed on.

elemental symbol A symbol assigned to an element based on the name of the element, consisting of one capital letter or a capital letter followed by a lowercase letter.

Compounds are pure substances made up of two or more different kinds of atoms. The atoms found in compounds are the same ones found in elements. Thus, the symbols used to represent elements can be combined and used to represent compounds. A molecular compound is depicted by a **compound formula,** in which each atom is represented by an appropriate elemental symbol. When more than one atom of an element is present in a molecule, a subscript is used to indicate the number.

compound formula A representation of the molecule of a compound, consisting of the symbols of the atoms found in the molecule. Atoms present in numbers greater than 1 have the number indicated by a subscript.

Notice the similarity between this practice and the molecular representations in Figure 1.4. The carbon dioxide molecules in Figure 1.4 are represented by the formula CO_2, where C represents an atom of carbon and O an atom of oxygen. The subscript 2 on the oxygen indicates that the molecule contains two atoms of oxygen. Notice that the single carbon atom in the molecule is not indicated by a subscript 1. Subscript 1 is never used in molecular formulas; it is understood. Formulas for molecular compounds are sometimes used to represent a compound in a general way, or to represent a single molecule of a compound. See Table 2.2 and Figure 2.1 for other examples of compound formulas.

Example 2.1

Write formulas for the following compounds:

a. Nitrogen dioxide: one nitrogen (N) atom and two oxygen (O) atoms
b. Sulfuric acid: two hydrogen (H) atoms, one sulfur (S) atom, and four oxygen (O) atoms

Table 2.1 The Chemical Elements and Their Symbols

Ac	actinium	Dy	dysprosium	Mn	manganese	Rn	radon
Ag	silver *(argentum)*[a]	Er	erbium	Mo	molybdenum	Ru	ruthenium
Al	aluminum	Es	einsteinium	Mt	meitnerium	S	sulfur
Am	americium	Eu	europium	N	nitrogen	Sb	antimony *(stibium)*[a]
Ar	argon	F	flourine	Na	sodium *(natrium)*[a]	Sc	scandium
As	arsenic	Fe	iron *(ferrum)*[a]	Nb	niobium	Se	selenium
At	astatine	Fm	fermium	Nd	neodymium	Sg	seaborgium
Au	gold *(aurum)*[a]	Fr	francium	Ne	neon	Si	silicon
B	boron	Ga	gallium	Ni	nickel	Sm	samarium
Ba	barium	Gd	gadolinium	No	nobelium	Sn	tin *(stannum)*[a]
Be	beryllium	Ge	germanium	Np	neptunium	Sr	strontium
Bh	bohrium	H	hydrogen	O	oxygen	Ta	tantalum
Bi	bismuth	He	helium	Os	osmium	Tb	terbium
Bk	berkelium	Hf	hafnium	P	phosphorus	Tc	technetium
Br	bromine	Hg	mercury *(hydrargyrum)*[a]	Pa	protactinium	Te	tellurium
C	carbon	Ho	holmium	Pb	lead *(plumbum)*[a]	Th	thorium
Ca	calcium	Hs	hassium	Pd	palladium	Ti	titanium
Cd	cadmium	I	iodine	Pm	promethium	Tl	thallium
Ce	cerium	In	indium	Po	polonium	Tm	thulium
Cf	californium	Ir	iridium	Pr	praseodymium	U	uranium
Cl	chlorine	K	potassium *(kalium)*[a]	Pt	platinum	V	vanadium
Cm	curium	Kr	krypton	Pu	plutonium	W	tungsten *(wolfram)*[a]
Co	cobalt	La	lanthanum	Ra	radium	Xe	xenon
Cr	chromium	Li	lithium	Rb	rubidium	Y	yttrium
Cs	cesium	Lr	lawrencium	Re	rhenium	Yb	ytterbium
Cu	copper *(cuprum)*[a]	Lu	lutetium	Rf	rutherfordium	Zn	zinc
Db	dubnium	Md	mendelevium	Rg	roentgenium	Zr	zirconium
Ds	darmstadtium	Mg	magnesium	Rh	rhodium		

[a]Elements with symbols not derived from their English names.

Table 2.2 Examples of Compound Formulas

Compound Name	Molecular Representation	Molecular Formula
Methane		CH_4
Water		H_2O
Carbon monoxide		CO
Hydrogen peroxide		H_2O_2

Solution

a. The symbols for the atoms are obtained from Table 2.1. The single N atom will not have a subscript because ones are understood and never written. The two O atoms will be represented by writing a subscript 2. The molecular formula is NO_2.

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Figure 2.1 Iron pyrite is a mineral that contains iron and sulfur atoms in a 1:2 ratio, respectively. The golden crystals, called "fool's gold" by experienced miners, have caused (temporary) excitement for many novice prospectors. What is the formula for iron pyrite?

b. Using similar reasoning, the H atom will have a subscript 2, the S atom will have no subscript, and the O atom will have a subscript 4. Therefore, the molecular formula is H_2SO_4.

Learning Check 2.1 Write molecular formulas for the following compounds:

a. Phosphoric acid: three hydrogen (H) atoms, one phosphorus (P) atom, and four oxygen (O) atoms
b. Sulfur trioxide: one sulfur (S) atom and three oxygen (O) atoms
c. Glucose: six carbon (C) atoms, twelve hydrogen (H) atoms, and six oxygen (O) atoms

2.2 Inside the Atom

Learning Objective

2. Identify the characteristics of protons, neutrons, and electrons.

An atom has been defined as the limit of chemical subdivision for matter. On the basis of the characteristics of atoms that have been discussed, you probably have a general (and correct) idea that atoms can be considered to be the units from which matter is made. However, the question of how atoms interact to form matter has not yet been addressed. This interesting topic is discussed in Chapters 3 and 4, but a bit more must be learned about atoms first.

Extensive experimental evidence collected since the middle of the 19th century indicates that atoms are made up of many smaller particles. More than 100 of these subatomic particles have been discovered, and the search for more continues. As yet there is no single theory that can explain all observations involving subatomic particles, but three fundamental particles are included in all current theories: the proton, the neutron, and the electron. Most chemical behavior of matter can be explained in terms of a few of the well-known characteristics of these particles. These important characteristics are mass, electrical charge, and location in atoms. They are summarized in Table 2.3. The atomic mass unit (u) shown in the table is discussed in Section 2.4.

In Chapter 1, atoms were described as being very tiny particles. The masses of single atoms of the elements are now known to fall within the range of 1.67×10^{-24} g for the least massive to 5.00×10^{-22} g for the most massive. Thus, it is not surprising to find that

Chemistry Around Us 2.1

Diamonds: From Gems to iPods

Most people associate the word *diamond* with images of expensive, sparkling rings or other jewelry. Because of their beauty, diamonds have been prized throughout history. However, in the modern world, the extreme hardness of this crystalline form of carbon makes it an indispensable part of industrial tools designed to cut, grind, or drill holes in other hard materials such as steel or stone. Until 1955, all diamonds used industrially or as gems came from natural sources. In that year, scientists first produced synthetic diamonds by subjecting ordinary carbon to very high pressures and temperatures. In 1970, these techniques had improved to the point that the first tiny gem-quality diamonds were produced.

For years, scientists have known about other very useful properties of diamond in addition to beauty and hardness. For example, electronic applications of diamond have the potential to help computers run at higher speeds without overheating, produce lasers of extreme power, downsize cell phones to fit in a wristwatch, and make iPods that store 10,000 movies instead of 10,000 songs. However, research into such applications has been limited by the price and rarity of natural diamonds.

Those limitations are rapidly disappearing as the ability to produce diamonds synthetically has increased significantly. Techniques using high pressures and temperatures have progressed to the point that synthetic diamonds of gem quality can now be produced and sold at a price of about one-fourth the price of natural stones. In addition, a process called chemical vapor deposition (CVD) has been developed. In this process, carbon is vaporized by heating it to a high temperature, and the vapor condenses on a tiny chip of diamond, causing the chip to grow. Small amounts of other elements such as boron can be included in diamond produced by this method, making the diamond into a semiconductor that is useful in electronic devices such as iPods.

Diamond-based semiconductors might greatly increase the performance of future electronic devices.

Table 2.3 Characteristics of the Fundamental Subatomic Particles

Particle	Common Symbols	Charge (±)	Mass (g)	Mass (u)	Location
		Characteristics			
Electron	e^-	1−	9.07×10^{-28}	1/1836	Outside nucleus
Proton	p, p^+, H^+	1+	1.67×10^{-24}	1	Inside nucleus
Neutron	n	0	1.67×10^{-24}	1	Inside nucleus

the particles that make up atoms have very small masses as well. Even though atomic masses are very small, the mass information given in Table 2.3 indicates that most of the mass of an atom comes from the protons and neutrons it contains. The protons and neutrons are tightly bound together to form the central portion of an atom called the **nucleus.** Because protons each have a +1 electrical charge, and neutrons have no charge, the nucleus of an atom has a positive electrical charge that is equal to the number of protons it contains.

nucleus The central core of atoms that contains protons, neutrons, and most of the mass of atoms.

Electrons are negatively charged particles located outside the nucleus of an atom. Protons and electrons carry equal but opposite electrical charges, so a neutral atom that has no electrical charge must have the same number of protons in its nucleus as it has electrons outside the nucleus. The electrons of an atom are thought to move very rapidly throughout a relatively large volume of space surrounding the small but very heavy nucleus (see Figure 2.2).

Even though subatomic particles exist, the atom itself is the particle of primary interest in chemistry because subatomic particles do not lead an independent existence for any appreciable length of time. The only way they gain long-term stability is by combining with other particles to form an atom.

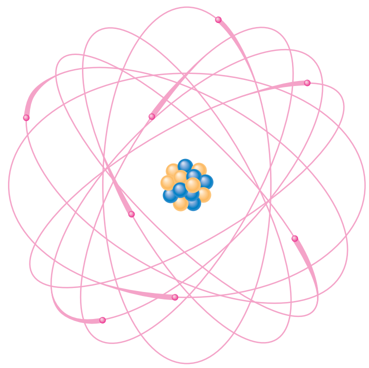

Figure 2.2 Electrons move rapidly around a massive nucleus. This figure is not drawn to scale. For a nucleus of the size shown, the closest electrons would be at least 80 m away.

2.3 Isotopes

Learning Objective

3. Use the concepts of atomic number and mass number to determine the number of subatomic particles in isotopes and to write correct symbols for isotopes.

Atoms of elements have no electrical charge and so must contain identical numbers of positive protons and negative electrons. However, because neutrons have no electrical charge, their numbers in an atom do not have to be the same as the numbers of protons or electrons. The number of protons in the nucleus of an atom is given by the **atomic number** for the atom. Atomic numbers are represented by the symbol Z. All atoms of a specific element must have the same atomic number. The atomic numbers for each element are the numbers above the elemental symbols of the periodic table inside the front cover of this book. Remember, this is also the number of electrons in the atoms of each element.

atomic number of an atom A number equal to the number of protons in the nucleus of an atom. Symbolically it is represented by Z.

The possibility of having different numbers of neutrons combined with one given number of protons to form nuclei leads to some interesting results. For example, three different kinds of hydrogen atoms are known to exist. Each kind of atom contains one proton and one electron, so all have an atomic number of 1. However, the nuclei of the different kinds of hydrogen atoms contain different numbers of neutrons. The most common kind has no neutrons in the nucleus, the next most common has one, and the least common kind has two. The sum of the number of protons and the number of neutrons in a nucleus is called the **mass number** and is represented by the symbol A. Thus, the three kinds of hydrogen atoms all have atomic numbers of 1 and mass numbers of 1, 2, and 3, respectively. Atoms that have the same atomic number but different mass numbers are called **isotopes.** Most elements are made up of mixtures of two or more isotopes. When it is important to distinguish between them, the following notation is used for each isotope: ${}^{A}_{Z}E$, where E is the symbol for the element.

mass number of an atom A number equal to the sum of the number of protons and neutrons in the nucleus of an atom. Symbolically it is represented by A.

isotopes Atoms that have the same atomic number but different mass numbers. That is, they are atoms of the same element that contain different numbers of neutrons in their nuclei.

The three isotopes of hydrogen are represented as follows using this notation: ${}^{1}_{1}H$, ${}^{2}_{1}H$, and ${}^{3}_{1}H$. When these symbols are not convenient to use, as in written or spoken references to the isotopes, the elemental name followed by the mass number is used. Thus, the three hydrogen isotopes are hydrogen-1, hydrogen-2, and hydrogen-3. These three isotopes have specific names: protium ($A = 1$), deuterium ($A = 2$), and tritium ($A = 3$).

Example 2.2

Use the periodic table inside the front cover to answer the following questions about isotopes:

a. What are the mass number, atomic number, and isotope symbol $\left({}^{A}_{Z}E\right)$ for an atom that contains 7 protons and 8 neutrons?
b. How many neutrons are contained in an atom of nickel-60?
c. How many protons and how many neutrons are contained in an atom with a mass number of 26 and the symbol Mg?

Solution

a. The mass number, A, equals the sum of the number of protons and the number of neutrons: $A = 7 + 8 = 15$. The atomic number, Z, equals the number of protons: $Z = 7$. According to the periodic table, the element with an atomic number of 7 is nitrogen, with the symbol N. The isotope symbol is ${}^{15}_{7}N$.

b. According to the periodic table, nickel has the symbol Ni, and an atomic number, Z, of 28. The mass number, 60, is equal to the sum of the number of protons and the number of neutrons. The number of protons is equal to the atomic number, 28. Therefore, the number of neutrons is $60 - 28 = 32$. The atom contains 32 neutrons.

c. According to the periodic table, the element with the symbol Mg is magnesium, which has an atomic number of 12. Therefore, the atom contains 12 protons. Since A, the number of protons plus neutrons is equal to 26, the number of neutrons is $26 - 12$, or 14. The atom contains 14 neutrons.

Learning Check 2.2 Use the periodic table inside the front cover to answer the following questions about isotopes:

a. What are the atomic number, mass number, and isotope symbol for an atom that contains 4 protons and 5 neutrons?

b. How many neutrons are contained in an atom of chlorine-37?

c. How many protons and how many neutrons are contained in an atom with a mass number of 28 and the symbol Si?

2.4 Relative Masses of Atoms and Molecules

Learning Objective

4. Use atomic weights of the elements to calculate molecular weights of compounds.

Because of their extremely small size, it is very inconvenient to use the actual masses of atoms when the atoms are being characterized or when quantitative calculations are done. In fact, the earliest chemists had no way of determining the actual masses of atoms. For this reason, a system was devised that utilized relative or comparative masses for the atoms. These relative masses are the numbers that are given beneath the symbol and name for each element in the periodic table inside the front cover.

Relative masses provide a simple way of comparing the masses of atoms. For example, the mass of neon atoms, Ne, from the periodic table is given as 20.18. Similarly, the mass of calcium atoms, Ca, is given as 40.08. These numbers simply indicate that calcium atoms have a mass that is about twice the mass of neon atoms. The exact relationship between the two masses is calculated as follows, using the correct number of significant figures:

$$\frac{\text{Ca atom mass}}{\text{Ne atom mass}} = \frac{40.08}{20.18} = 1.986$$

In a similar way, we arrive at the conclusion that helium atoms are about four times as massive as hydrogen atoms:

$$\frac{\text{He atom mass}}{\text{H atom mass}} = \frac{4.003}{1.008} = 3.971$$

In each case, we have been able to determine the relationship between the masses of the atoms without using the actual masses.

Modern instruments called mass spectrometers allow the actual masses of individual atoms to be measured. These measured masses show the same relationships to each other as do the relative masses:

$$\frac{\text{Ca atom mass}}{\text{Ne atom mass}} = \frac{6.655 \times 10^{-23}\text{ g}}{3.351 \times 10^{-23}\text{ g}} = 1.986$$

$$\frac{\text{He atom mass}}{\text{H atom mass}} = \frac{6.647 \times 10^{-24}\text{ g}}{1.674 \times 10^{-24}\text{ g}} = 3.971$$

At The Counter 2.1

Calcium Supplements: Which Type Is Best?

In a nutritional context, a supplement provides an amount of a substance that is in addition to the amount normally obtained from the diet.

About 99% of the calcium in the body is used to build bones and teeth. During a lifetime, all bones of the body undergo a natural process of buildup and breakdown. The rate of buildup exceeds the rate of breakdown for the first 25–30 years of life for women and the first 30–35 years of life for men. Beyond these times, the rate of breakdown exceeds the rate of buildup, resulting in a gradual decrease in bone density. Consequently, bones become increasingly weakened, brittle, and susceptible to breaking—a condition called *osteoporosis.* About 50% of women and 13% of men over age 50 suffer a broken bone as a result of osteoporosis.

One of the best ways to reduce the risks associated with osteoporosis is to build as much bone as possible during early life when the rate of buildup exceeds the rate of breakdown. With this goal in mind, the following daily calcium intakes have been suggested:

Age	Amount of Calcium
birth to 6 months	400 mg
6 months to 1 year	600 mg
1 to 10 years	800 mg
11 to 24 years	1200–1500 mg
25 to 50 years	1500 mg
51 to 65	*100–1500 mg
Over age 65	1500 mg

* depending on hormone replacement therapy

Sufficient calcium for building bones is provided by a balanced diet that includes calcium-rich foods such as dairy products, certain vegetables (broccoli, kale, and turnip and collard greens), tofu, some canned fish, legumes (beans, peas, etc.) and seeds and nuts. Unfortunately, many people in the prime time of their bone-building years (pre-adolescents, adolescents, and young adults) follow diets that fall significantly short of the recommended levels of calcium for optimal bone-building. Two significant reasons for this nutritional shortfall among people in these age groups are the tendency to skip meals and the substitution of soft drinks and other non-dairy drinks in place of milk.

If a calcium supplement is needed, which type is best? Most supplements will contain calcium in one of the following three chemical forms: calcium carbonate (often from oyster shells), calcium citrate or calcium phosphate. It really makes little difference which of these three chemical forms the calcium is in, as all three are absorbed quite well by the body. The important factor in a supplement is the amount of calcium contained in each dose. This amount per dose is generally indicated on the label and typically ranges from 333 mg to 630 mg. The maximum benefit from calcium supplements is obtained when the individual dosage is 500 mg or less. So, supplements with individual dosages greater than 500 mg should be divided and taken in portions throughout the day. An additional consideration is that vitamin D is essential for maximum calcium absorption by the body. For this reason, many calcium supplements include vitamin D in their formulation, and clearly indicate this on their labels.

Osteoporosis might seem too far away in the future to concern a teenager or young adult. Nevertheless, simple lifestyle changes now such as improving the diet or taking a calcium supplement might provide substantial and much-appreciated health benefits in that (not so) distant future.

SIU Biomedical Comm./Custom Medical Stock Photo

The availability of vitamin D in a calcium supplement is usually indicated on the label.

Thus, we see that the relative masses used in the periodic table give the same results as the actual masses when the masses of atoms are compared with one another.

Actual atomic masses are given in mass units such as grams, but the relative values used in the periodic table are given in units referred to as **atomic mass units.** Until recently, the abbreviation for an atomic mass unit was amu. However, the accepted abbreviation is now u. The actual mass represented by a single atomic mass unit is $\frac{1}{12}$ the mass of a single carbon-12 atom or 1.661×10^{-24} g.

atomic mass unit (u) A unit used to express the relative masses of atoms. One u is equal to $\frac{1}{12}$ the mass of an atom of carbon-12.

The relative masses of the elements as given in the periodic table are referred to as atomic masses or atomic weights. We will use the term **atomic weights** in this book. In those cases where the naturally occurring element exists in the form of a mixture of isotopes, the recorded atomic weight is the average value for the naturally occurring isotope mixture. This idea is discussed in Section 2.5.

atomic weight The mass of an average atom of an element expressed in atomic mass units.

Chemistry and Your Health 2.1

Are You at Risk for Osteoporosis?

Osteoporosis, the abnormal thinning of bones that often accompanies aging, may lead to bone fractures, disability, and even death. While women are most susceptible, this serious condition also affects men but usually at a more advanced age than women. A number of significant risk factors have been identified, including the following:

1. A poor diet, especially one low in calcium
2. Advanced age
3. The onset of menopause (or having had ovaries removed)
4. A sedentary lifestyle
5. Smoking
6. A family history of osteoporosis or hip fracture
7. Heavy drinking
8. The long-term use of certain steroid drugs
9. Vitamin D deficiency

Ninety-nine percent of the calcium found in the body is located in the skeleton and teeth, so it is not surprising that the behavior of this metallic element in the body plays a central role in a number of these risk factors for osteoporosis. For example, it is known that in later life the body's ability to absorb calcium from food in the small intestine decreases (factor 2) and such absorption requires the presence of vitamin D (factor 9).

It is well recognized that the best insurance against developing osteoporosis in later life is to build and strengthen as much bone as possible during the first 25 to 35 years of life. Two essential components of this process are eating a healthful diet containing adequate amounts of calcium and vitamin D, and following a healthful lifestyle that includes regular weight-bearing exercise such as walking, jogging, weight lifting, stair climbing, or physical labor. Dietary calcium supplements provide an additional way to enhance calcium intake, especially for individuals who are at risk to develop osteoporosis (See At the Counter 2.1).

A bone density test provides an effective way to diagnose the presence or extent of osteoporosis in an individual. Such tests measure the absorption of X-rays by bones, and are not invasive or uncomfortable. Bone density tests were not routinely recommended in the past, but some health and wellness organizations now suggest that women who are at high risk should have such a test by the age of 50, and all women over age 65 should be routinely tested.

Sally Ho/Getty Images

An active lifestyle is an essential component of building and strengthening bones.

A convenient way to visualize the concept of relative masses, and to compare the relative masses of atoms, involves the use of the simple child's toy called a see-saw or teeter-totter shown below.

When the masses on both sides of the central pivot are equal, the see-saw will be in balance. When the masses on each side are different, the see-saw will be out of balance, and the side with the greatest mass will go down. This characteristic lends itself to comparing the masses of objects to each other without knowing their actual values in mass units such as grams.

Example 2.3

Round the atomic weights of the periodic table to the nearest whole number, and answer the following questions:

a. How many calcium atoms, Ca, would have to be put on one side of a see-saw to balance one atom of bromine, Br, that was placed on the other side?
b. How many helium atoms, He, would balance one oxygen atom on a see-saw?
c. How many hydrogen atoms, H, would balance one carbon atom, C, on a see-saw?

Solution

a. Calcium atoms have a periodic table atomic weight of 40.08u, which rounds to 40u. Bromine atoms have a periodic table atomic weight of 79.90u, which rounds to 80u. Thus, two calcium atoms would be required to balance one bromine atom as indicated below:

b. Helium atoms have a periodic table atomic weight of 4.003u, which rounds to 4u. Oxygen atoms have a periodic table atomic weight of 16.00u, which rounds to 16u. Thus, four helium atoms will balance one oxygen atom:

c. Hydrogen atoms have a periodic table atomic weight of 1.008u, which rounds to 1u. Carbon atoms have a periodic table atomic weight of 12.01u, which rounds to 12u. Twelve hydrogen atoms will balance one carbon atom:

Learning Check 2.3 Round atomic weights of the periodic table to the nearest whole number, and answer the following questions using the see-saw balance idea:

a. How many nitrogen atoms, N, will balance one atom of iron, Fe?
b. How many carbon atoms, C, will balance six helium atoms, He?
c. How many calcium atoms, Ca, will balance two argon atoms, Ar?

Molecules are made up of atoms, so the relative mass of a molecule can be calculated by adding together the atomic weights of the atoms that make up the molecule. Relative molecular masses calculated in this way are called **molecular weights** and are also given in atomic mass units.

molecular weight The relative mass of a molecule expressed in atomic mass units and calculated by adding together the atomic weights of the atoms in the molecule.

Example 2.4

Use atomic weights from the periodic table inside the front cover to determine the molecular weight of urea, CH_4N_2O, the chemical form in which much nitrogenous body waste is excreted in the urine.

Solution

According to the formula given, a urea molecule contains one carbon atom, C, two nitrogen atoms, N, four hydrogen atoms, H, and one oxygen atom, O. The molecular weight is calculated as follows:

$$\text{MW} = 1(\text{at. wt. C}) + 2(\text{at. wt. N}) + 4(\text{at. wt. H}) + 1(\text{at. wt. O})$$

$$\text{MW} = 1(12.01\text{ u}) + 2(14.01\text{ u}) + 4(1.008\text{ u}) + 1(16.00\text{ u}) = 60.062\text{ u}$$

Rounded to four significant figures, the correct answer is 60.06 u.

Learning Check 2.4

a. Use atomic weights from the periodic table to determine the molecular weight of sulfuric acid. Each molecule contains two hydrogen (H) atoms, one sulfur (S) atom, and four oxygen (O) atoms.
b. Determine the molecular weight of isopropyl alcohol (C_3H_8O), the active ingredient in most rubbing alcohol sold commercially.

2.5 Isotopes and Atomic Weights

Learning Objective

5. Use isotope percent abundances and masses to calculate atomic weights of elements.

The atomic weights discussed in Section 2.4 were defined as the relative masses of atoms of the elements. It would have been more correct to define them as the relative masses of average atoms of the elements. Why include the idea of an average atom? Remember, the mass number of an isotope is the sum of the number of protons and neutrons in the nucleus of atoms of the isotope. Also, protons and neutrons both have masses of 1 u (see Table 2.3). The masses of electrons are quite small, so the atomic weights of isotopes are very nearly equal to their mass numbers. For example, all naturally occurring phosphorus is made up of atoms containing 15 protons and 16 neutrons. The mass number of this isotope is 31, and its symbol is ${}^{31}_{15}P$. The relative mass (atomic weight) of atoms of this isotope is 30.97 u, the same as the atomic weight of the element given in the periodic table. The atomic weight of the isotope and the listed atomic weight of the element are the same because all atoms contained in the element are identical and have the same relative mass.

Naturally occurring chlorine is different. It is a mixture of two isotopes, ${}^{35}_{17}Cl$ and ${}^{37}_{17}Cl$. Chlorine-35 has a mass number of 35 and a relative mass of 34.97 u. Chlorine-37 has a mass number of 37 and a relative mass of 36.97 u. Note that the mass numbers and relative masses (atomic weights) of the atoms of an isotope are essentially identical. However, naturally occurring chlorine is a mixture containing both isotopes. Thus, the determined relative mass of chlorine atoms will be the average relative mass of the atoms found in the mixture (see Figure 2.3).

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Figure 2.3 The pieces of fruit in a bowl are somewhat like atoms of the isotopes of an element. Each piece of fruit may have the same color, taste, and texture, but it is unlikely that any two have *exactly* the same mass. The 12 oranges in the bowl weigh a total of 2.36×10^3 g. What is the average mass of each orange in the bowl?

The average mass of each particle in a group of particles is simply the total mass of the group divided by the number of particles in the group. This calculation requires that the total number of particles be known. However, the percentage of each isotope in a mixture of atoms is easier to determine than the actual number of atoms of each isotope present. The percentages can be used to calculate average masses. Remember that percent means per 100, so we use an imaginary sample of an element containing 100 atoms in the calculation. On this basis, the number of atoms of each isotope in the 100-atom sample will be the percentage of that isotope in the sample. The mass contributed to the sample by each isotope will be the product of the number of atoms of the isotope (the percentage of the isotope) and the mass of the isotope. The total mass of the sample will be the sum of the masses contributed by each isotope. This total mass divided by 100, the number of atoms in our imaginary sample, gives the mass of an average atom, which is the atomic weight of the element.

Example 2.5

Calculate the atomic weight of chlorine, given that the naturally occurring element consists of 75.53% chlorine-35 (mass = 34.97 u) and 24.47% chlorine-37 (mass = 36.97 u).

Solution

$$\text{atomic weight} = \frac{(\%\ \text{chlorine-35})(\text{mass chlorine-35}) + (\%\ \text{chlorine-37})(\text{mass chlorine-37})}{100}$$

$$= \frac{(75.53)(34.97\ \text{u}) + (24.47)(36.97\ \text{u})}{100}$$

$$= \frac{2641.28\ \text{u} + 904.66\ \text{u}}{100} = \frac{3545.94\ \text{u}}{100} = 35.4594\ \text{u}$$

$$= 35.46\ \text{u}\ (\text{rounded value})$$

This result is slightly different from the periodic table atomic weight value of 35.45 because of slight errors introduced in rounding the isotope masses to four significant figures.

Learning Check 2.5

a. Naturally occurring fluorine consists of a single isotope, fluorine-19, with a mass of 19.00 u (using four significant figures). Determine the atomic weight of fluorine, and compare your answer with the value given in the periodic table.

b. Naturally occurring magnesium has three isotopes: magnesium-24, magnesium-25, and magnesium-26. Their relative masses and percent abundances are, respectively, 23.99 u (78.70%), 24.99 (10.13%), and 25.98 (11.17%). Determine the atomic weight of magnesium, and compare it with the value given in the periodic table.

2.6 Avogadro's Number: The Mole

Learning Objective

6. Use the mole concept to obtain relationships between number of moles, number of grams, and number of atoms for elements, and use those relationships to obtain factors for use in factor-unit calculations.

The atomic weights of the elements given in the periodic table have far more uses than simply comparing the masses of the atoms of various elements. However, these uses are not apparent until some additional ideas are developed. According to the periodic table, the atomic weight of magnesium, Mg, is 24.31 u, and the atomic weight of carbon, C, is 12.01 u. As we have seen, this means that an average magnesium atom has about twice the mass of an average carbon atom. Modern instruments allow us to determine that the actual mass of an average magnesium atom is 4.037×10^{-23} g, and the actual mass of an average carbon atom is 1.994×10^{-23} g.

Even though the actual masses of magnesium and carbon atoms are extremely small, it should still be possible to collect together enough atoms of either element to give a sample of any desired mass. Suppose we wish to collect together enough atoms of each element to give a sample with a mass in grams equal to the atomic weight of each element. That is, we want to collect enough magnesium atoms to give a sample with a total mass of 24.31 g, and enough carbon atoms to give a sample with a total mass of 12.01 g (see Figure 2.4). How many atoms of each element will be required?

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Figure 2.4 Samples of magnesium (left) and carbon with masses of 24.31 g and 12.01 g, respectively.

We know the mass of one atom of each element, and this fact will provide us with a factor that will allow us to convert the desired sample mass into the number of atoms required. The given mass of one atom of magnesium can be written as

$$1 \text{ Mg atom} = 4.037 \times 10^{-23} \text{ g Mg}$$

This known relationship provides two factors that can be used to solve factor-unit problems:

$$\frac{1 \text{ Mg atom}}{4.037 \times 10^{-23} \text{ g Mg}} \quad \text{and} \quad \frac{4.037 \times 10^{-23} \text{ g Mg}}{1 \text{ Mg atom}}$$

Our task is to convert the desired sample mass of 24.31 g Mg into the number of Mg atoms in the sample. The first factor will cancel the units "g Mg" and will generate the units "Mg atoms."

$$(24.31 \cancel{\text{g Mg}}) \times \frac{1 \text{ Mg atom}}{4.037 \times 10^{-23} \cancel{\text{g Mg}}} = 6.022 \times 10^{23} \text{ Mg atoms}$$

In a similar way, the number of C atoms needed to produce a sample with a mass of 12.01 g is calculated:

$$(12.01 \cancel{\text{g C}}) \times \frac{1 \text{ C atom}}{1.994 \times 10^{-23} \cancel{\text{g C}}} = 6.022 \times 10^{23} \text{ C atoms}$$

The result that the same number of atoms is required for each sample might seem surprising, but it is a consequence of the sample sizes we wanted to produce. If we collected just one atom of each element, the ratio of the mass of Mg to the mass of C would be equal to the atomic weight of Mg divided by the atomic weight of C:

$$\frac{\text{Mg mass}}{\text{C mass}} = \frac{4.037 \times 10^{-23}\,\cancel{\text{g}}}{1.994 \times 10^{-23}\,\cancel{\text{g}}} = 2.024 = \frac{\text{At. wt. Mg}}{\text{At. wt. C}} = \frac{24.31\,\cancel{\text{u}}}{12.01\,\cancel{\text{u}}}$$

This result reflects the fact that magnesium atoms have about twice the mass of carbon atoms. If we collected samples of 100 atoms of each element, the ratio of the mass of the Mg sample to the mass of the C sample would still be 2.024 because each sample would have a mass 100 times greater than the mass of a single atom:

$$\frac{\text{Mg mass}}{\text{C mass}} = \frac{(100)(4.037 \times 10^{-23}\,\cancel{\text{g}})}{(100)(1.994 \times 10^{-23}\,\cancel{\text{g}})} = 2.024$$

It follows that if we collected samples containing 6.022×10^{23} atoms of each element, the ratio of the mass of the Mg sample to the mass of the C sample would still be 2.024 because each sample would have a mass that would be 6.022×10^{23} times greater than the mass of a single atom:

$$\frac{\text{Mg mass}}{\text{C mass}} = \frac{(6.022 \times 10^{23})(4.037 \times 10^{-23}\,\cancel{\text{g}})}{(6.022 \times 10^{23})(1.994 \times 10^{-23}\,\cancel{\text{g}})} = 2.024$$

These results lead to the following conclusion: Any samples of magnesium and carbon that have mass ratios equal to 2.024 will contain the same number of atoms.

Example 2.6

Show that samples of magnesium and carbon with masses of 9.663 g and 4.774 g, respectively, have a mass ratio of 2.024 and contain the same number of atoms.

Solution

The mass ratio is obtained by dividing the mass of the magnesium sample by the mass of the carbon sample:

$$\frac{\text{Mg mass}}{\text{C mass}} = \frac{9.663\text{ g}}{4.774\text{ g}} = 2.024$$

The number of atoms in each is calculated using the factors obtained earlier from the mass of one atom of each element:

$$(9.663\ \cancel{\text{g Mg}}) \times \frac{1\text{ Mg atom}}{4.037 \times 10^{-23}\ \cancel{\text{g Mg}}} = 2.394 \times 10^{23}\text{ Mg atoms}$$

$$(4.774\ \cancel{\text{g C}}) \times \frac{1\text{ C atom}}{1.994 \times 10^{-23}\ \cancel{\text{g C}}} = 2.394 \times 10^{23}\text{ C atoms}$$

Learning Check 2.6 Show that samples of magnesium and carbon with masses of 13.66 g and 6.748 g, respectively, have a mass ratio of 2.024 and contain the same number of atoms.

The preceding results for magnesium and carbon may be generalized for all the elements of the periodic table as follows. Any samples of two elements that have a mass ratio equal to the ratio of their atomic weights will contain identical numbers of atoms. In addition, we have seen that if the number of grams of sample is equal numerically to the atomic weight of an element, the number of atoms in the sample is equal to 6.022×10^{23}. Figure 2.5 illustrates these ideas for particles that are familiar to most of us.

With modern equipment, it is possible to determine the number of atoms in any size sample of an element. However, before this was possible, the practice of focusing on samples with a mass in grams equal to the atomic weights of the elements became well established. It continues today. We have learned that the number of atoms in such samples is 6.022×10^{23}.

Figure 2.5 An average jelly bean has a mass that is 1.60 times the mass of an average dry bean. Each jar contains the same number of beans. The total mass of jelly beans is 472 g. What is the total mass of the dry beans?

We saw in Section 2.4 that molecular weights, the relative masses of molecules, are calculated by adding the atomic weights of the atoms contained in the molecules. The resulting molecular weights are expressed in atomic mass units, just as are the atomic weights of the atoms. The same ideas we used to discuss the actual and relative masses of atoms can be applied to molecules. We will not go through the details but simply state the primary conclusion: A sample of compound with a mass in grams equal to the molecular weight of the compound contains 6.022×10^{23} molecules of the compound.

The number 6.022×10^{23} is called *Avogadro's number* in honor of Amadeo Avogadro (1776–1856), an Italian scientist who made important contributions to the concept of atomic weights. As we have seen, this number represents the number of atoms or molecules in a specific sample of an element or compound. Because of its importance in calculations, the number of particles represented by Avogadro's number is given a specific name; it is called a **mole,** abbreviated mol.

mole The number of particles (atoms or molecules) contained in a sample of element or compound with a mass in grams equal to the atomic or molecular weight, respectively. Numerically, 1 mol is equal to 6.022×10^{23} particles.

It is sometimes helpful to remember that the word *mole* represents a specific number, just as the word *dozen* represents 12, regardless of the objects being counted. Thus, 6.022×10^{23} people would be called 1 mol of people, just as 12 people would be called 1 dozen people. The immensity of Avogadro's number is illustrated by the results of a few calculations based on 1 mol of people. One mol of people would be enough to populate about 1×10^{14} Earths at today's level. That is 100 trillion Earths. Or, put another way, the present population of Earth is 1×10^{-12}% (0.000000000001%) of 1 mol.

In the development of these ideas to this point, we have used four significant figures for atomic weights, molecular weights, and Avogadro's number to minimize the introduction of rounding errors. However, in calculations throughout the remainder of the book, three significant figures will generally be sufficient and will be used.

In a strict sense, 1 mol is a specific number of particles. However, in chemistry it is customary to follow the useful practice of also letting 1 mol stand for the mass of a sample of element or compound that contains Avogadro's number of particles. Thus, the application of the mole concept to sulfur (at. wt. = 32.1 u) gives the following relationships:

$$1 \text{ mol S atoms} = 6.02 \times 10^{23} \text{ S atoms} = 32.1 \text{ g S}$$

When written individually, these three relationships can be used to generate six factors for use in factor-unit calculations involving sulfur:

$$\begin{aligned} 1 \text{ mol S atoms} &= 6.02 \times 10^{23} \text{ S atoms} \\ 6.02 \times 10^{23} \text{ S atoms} &= 32.1 \text{ g S} \\ 1 \text{ mol S atoms} &= 32.1 \text{ g S} \end{aligned}$$

Example 2.7

Determine the following using the factor-unit method of calculation and factors obtained from the preceding three relationships given for sulfur (S):

a. The mass in grams of 1.35 mol of S
b. The number of moles of S atoms in 98.6 g of S
c. The number of S atoms in 98.6 g of S
d. The mass in grams of one atom of S

Solution

a. The known quantity is 1.35 mol of S, and the unit of the unknown quantity is grams of S. The factor comes from the relationship 1 mol S atoms = 32.1 g S.

$$(1.35\ \cancel{\text{mol S atoms}})\left(\frac{32.1\ \text{g S}}{1\ \cancel{\text{mol S atoms}}}\right) = 43.3\ \text{g S}$$

b. The known quantity is 98.6 g of S, and the unit of the unknown quantity is moles of S atoms. The factor comes from the same relationship used in (a).

$$(98.6\ \cancel{\text{g S}})\left(\frac{1\ \text{mol S atoms}}{32.1\ \cancel{\text{g S}}}\right) = 3.07\ \text{mol S atoms}$$

c. The known quantity is, again, 98.6 g of S, and the unit of the unknown is the number of S atoms. The factor comes from the relationship 6.02×10^{23} S atoms = 32.1 g S.

$$(98.6\ \cancel{\text{g S}})\left(\frac{6.02 \times 10^{23}\ \text{S atoms}}{32.1\ \cancel{\text{g S}}}\right) = 1.85 \times 10^{24}\ \text{S atoms}$$

d. The known quantity is one S atom, and the unit of the unknown is grams of S. The factor comes from the same relationship used in (c), 6.02×10^{23} S atoms = 32.1 g S. Note that the factor is the inverse of the one used in (c) even though both came from the same relationship. Thus, we see that each relationship provides two factors.

$$(1\ \cancel{\text{S atom}})\left(\frac{32.1\ \text{g S}}{6.02 \times 10^{23}\ \cancel{\text{S atoms}}}\right) = 5.33 \times 10^{-23}\ \text{g S}$$

Learning Check 2.7 Calculate the mass of a single oxygen atom in grams. How does the ratio of this mass divided by the mass of a carbon atom given earlier (1.994×10^{-23} g) compare with the ratio of the atomic weights of oxygen and carbon given in the periodic table?

The mole concept can also be applied to particles that are molecules instead of atoms. The compound carbon dioxide consists of molecules that contain one carbon atom, C, and two oxygen atoms, O. The formula for the molecule is CO_2. The molecular weight of the molecule is calculated as shown earlier by adding together the atomic weight of one carbon atom and the atomic weight of two oxygen atoms:

$$\text{MW} = 1\ (\text{at. wt. C}) + 2\ (\text{at. wt. O}) = 1\ (12.0\ \text{u}) + 2\ (16.0\ \text{u}) = 44.0\ \text{u}$$

Thus, the relative mass of a CO_2 molecule is 44.0 u, and application of the mole concept to CO_2 molecules gives the following relationships:

$$1\ \text{mol}\ CO_2\ \text{molecules} = 6.02 \times 10^{23}\ CO_2\ \text{molecules} = 44.0\ \text{g}\ CO_2$$

When written individually, these three relationships can be used to generate six factors for use in factor-unit calculations involving CO_2.

$$1\ \text{mol}\ CO_2\ \text{molecules} = 6.02 \times 10^{23}\ CO_2\ \text{molecules}$$
$$6.02 \times 10^{23}\ CO_2\ \text{molecules} = 44.0\ \text{g}\ CO_2$$
$$1\ \text{mol}\ CO_2\ \text{molecules} = 44.0\ \text{g}\ CO_2$$

Example 2.8

Determine the following using the factor-unit method of calculation and factors obtained from the preceding three relationships given for carbon dioxide, CO_2:

a. The mass in grams of 1.62 mol of CO_2
b. The number of moles of CO_2 molecules in 63.9 g of CO_2

c. The number of CO_2 molecules in 63.9 g of CO_2
d. The mass in grams of one molecule of CO_2

Solution

a. The known quantity is 1.62 mol of CO_2, and the unit of the unknown quantity is g CO_2. The factor comes from the relationship 1 mol CO_2 molecules = 44.0 g CO_2.

$$(1.62\ \cancel{\text{mol } CO_2 \text{ molecules}})\left(\frac{44.0 \text{ g } CO_2}{1\ \cancel{\text{mol } CO_2 \text{ molecules}}}\right) = 71.3 \text{ g } CO_2$$

b. The known quantity is 63.9 g of CO_2, and the unit of the unknown quantity is mol of CO_2 molecules. The factor comes from the same relationship used in (a).

$$(63.9\ \cancel{\text{g } CO_2})\left(\frac{1 \text{ mol } CO_2 \text{ molecules}}{44.0\ \cancel{\text{g } CO_2}}\right) = 1.45 \text{ mol } CO_2 \text{ molecules}$$

c. The known quantity is, again, 63.9 g of CO_2, and the unit of the unknown is the number of CO_2 molecules. The factor comes from the relationship 6.02×10^{23} CO_2 molecules = 44.0 g CO_2.

$$(63.9\ \cancel{\text{g } CO_2})\left(\frac{6.02 \times 10^{23}\ CO_2 \text{ molecules}}{44.0\ \cancel{\text{g } CO_2}}\right) = 8.74 \times 10^{23}\ CO_2 \text{ molecules}$$

d. The known quantity is one CO_2 molecule, and the unit of the unknown is g CO_2. The factor comes from the same relationship used in (c), but the factor is the inverse of the one used in (c).

$$(1\ \cancel{CO_2 \text{ molecule}})\left(\frac{44.0 \text{ g } CO_2}{6.02 \times 10^{23}\ \cancel{CO_2 \text{ molecules}}}\right) = 7.31 \times 10^{-23} \text{ g } CO_2$$

Learning Check 2.8 Calculate the molecular weight of carbon monoxide, CO, in u, then calculate the mass of a single CO molecule in grams. How does the ratio of this mass divided by the mass of a single CO_2 molecule (from Example 2.8 d) compare to the ratio of the molecular weight of CO in u divided by the molecular weight of CO_2 in u (from Example 2.8)?

2.7 The Mole and Chemical Formulas

Learning Objective

7. Use the mole concept and molecular formulas to obtain relationships between number of moles, number of grams, and number of atoms or molecules for compounds, and use those relationships to obtain factors for use in factor-unit calculations.

According to Section 2.1 and as demonstrated in Example 2.8, the formula for a compound is made up of the symbols for each element present. Subscripts following the elemental symbols indicate the number of each type of atom in the molecule represented. Thus, chemical formulas represent the numerical relationships that exist among the atoms in a compound. Application of the mole concept to the atoms making up the formulas provides additional useful information.

Consider water as an example. The formula H_2O represents a 2:1 ratio of hydrogen atoms to oxygen atoms in a water molecule. Since this ratio is fixed, the following statements can be written:

1. 2 H_2O molecules contain 4 H atoms and 2 O atoms.
2. 10 H_2O molecules contain 20 H atoms and 10 O atoms.
3. 100 H_2O molecules contain 200 H atoms and 100 O atoms.
4. 6.02×10^{23} H_2O molecules contain 12.04×10^{23} H atoms and 6.02×10^{23} O atoms.

Study Skills 2.1 Help with Mole Calculations

Problems involving the use of the mole often strike fear into the hearts of beginning chemistry students. The good news is that problems involving the use of moles, atoms, molecules, and grams are made easier by using the factor-unit method. The method focuses your attention on the goal of eliminating the unit of the known, or given, quantity and of generating the unit of the answer, or unknown, quantity. Remember, Step 1 is to write down the number and unit of the given quantity. In Step 2, write down the unit of the answer. In Step 3, multiply the known quantity by a factor whose units will cancel that of the known quantity and will generate the unit of the answer. In Step 4, obtain the answer by doing the required arithmetic using the numbers that were introduced in Steps 1–3.

The ability to write the necessary factors for use in Step 3 is essential if you are to become proficient in solving mole problems using the factor-unit method. The factors come from numerical relationships between quantities that are obtained from definitions, experimental measurements, or combinations of the two. The definition of the mole, coupled with experimentally determined atomic and molecular weights, gives the following numerical relationships:

Atom Y: $1 \text{ mol } Y \text{ atoms} = 6.02 \times 10^{23}\ Y \text{ atoms} = y \text{ g } Y$

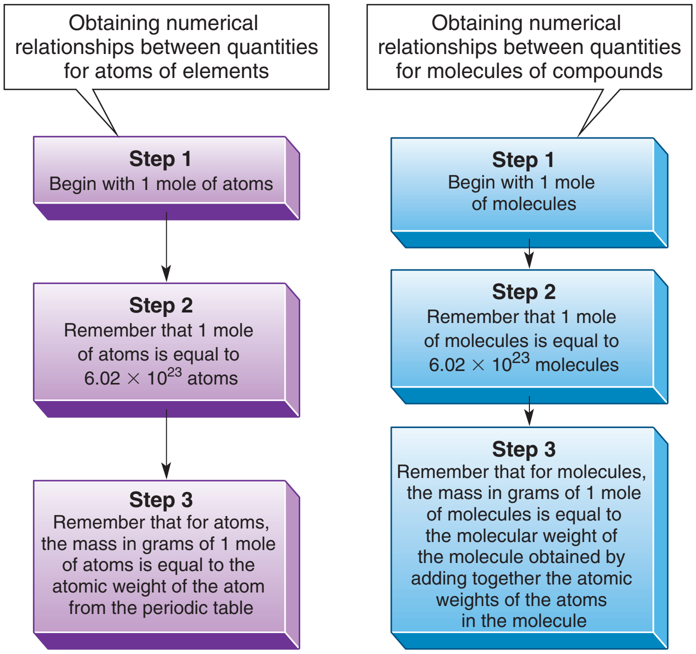
where y is the atomic weight of element Y.

Molecule Z: $1 \text{ mol } Z \text{ molecules} = 6.02 \times 10^{23}\ Z \text{ molecules} = z \text{ g } Z$

where z is the molecular weight of compound Z. Each of these sets of three related quantities will give six factors that can be used in factor-unit problems. Each factor is simply a ratio between any two of the three related quantities such as

$$\frac{1 \text{ mol } Y \text{ atoms}}{6.02 \times 10^{23}\ Y \text{ atoms}}$$

Because each ratio can be inverted, six different factors result from each set of three quantities. See if you can write the five other factors for atom Y. The steps given in the flow charts below will help you.



Statement 4 is significant because 6.02×10^{23} particles is 1 mol. Thus, Statement 4 can be changed to Statement 5:

5. 1 mol of H_2O molecules contains 2 mol of H atoms and 1 mol of O atoms.

Figure 2.6 contains another example of this concept.

Example 2.9

How many moles of ears, tails, and legs are contained in 1 mol of normal rabbits?

Solution

This example is nonchemical, but it might help you grasp the relationships that exist between the individual parts of a formula and the formula as a whole. The parts of a rabbit are related to a rabbit just as the parts of a formula are related to the entire formula.

Ears. Each rabbit has two ears and a 2:1 ratio exists between the number of ears and the number of rabbits. Therefore, 1 mol of rabbits contains 2 mol of ears.

Tails. The 1:1 ratio of tails to rabbits leads to the result that 1 mol of rabbits contains 1 mol of tails.

Legs. The 4:1 ratio of legs to rabbits leads to the result that 1 mol of rabbits contains 4 mol of legs.

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Figure 2.6 Liquid carbon disulfide (CS_2) is composed of carbon (left) and sulfur (right), elements that are solids. How many moles of sulfur atoms would be contained in 1.5 mol of CS_2 molecules?

Example 2.10

How many moles of each type of atom are contained in 1 mol of chloroform ($CHCl_3$)?

Solution

Just as each rabbit has two ears, each chloroform molecule contains one C atom, one H atom, and three Cl atoms. Therefore, 1 mol of $CHCl_3$ contains 1 mol of C atoms, 1 mol of H atoms, and 3 mol of Cl atoms.

Learning Check 2.9 How many moles of each type of atom would be contained in 0.50 mol of glucose ($C_6H_{12}O_6$)?

The usefulness of this approach can be increased by remembering and using the mass relationships of the mole concept. Thus, Statement 5 written earlier for water can be changed to Statement 6:

6. 18.0 g of water contains 2.0 g of H and 16.0 g of O.

Or, in a more concise form:

$$18.0 \text{ g } H_2O = 2.0 \text{ g H} + 16.0 \text{ g O}$$

Mass relationships such as these allow percent compositions to be calculated easily.

Example 2.11

Ammonia (NH_3) and ammonium nitrate (NH_4NO_3) are commonly used agricultural fertilizers. Which one of the two contains the higher mass percentage of nitrogen (N)?

Solution

In each case, the mass percentage of N is given by

$$\% \text{ N} = \frac{\text{part}}{\text{total}} \times 100 = \frac{\text{mass of N}}{\text{mass of compound}} \times 100$$

We will use 1 mol of each compound as a sample because the mass in grams of 1 mol of compound and the mass in grams of N in the 1 mol of compound are readily determined. One mol of NH_3 weighs 17.0 g and contains 1 mol of N atoms, which weighs 14.0 g.

$$\% \text{ N} = \frac{14.0 \text{ g}}{17.0 \text{ g}} \times 100 = 82.4\%$$

Similarly, 1 mol of NH_4NO_3 weighs 80.0 g and contains 2 mol of N atoms, which weigh 28.0 g.

$$\% \text{ N} = \frac{28.0 \text{ g}}{80.0 \text{ g}} \times 100 = 35.0\%$$

Learning Check 2.10 Determine the mass percentage of carbon in carbon dioxide (CO_2) and carbon monoxide (CO).

Concept Summary

Symbols and Formulas. Symbols based on names have been assigned to every element. Most consist of a single capital letter followed by a lowercase letter. A few consist of a single capital letter. Compounds are represented by formulas made up of elemental symbols. The number of atoms of each element in a molecule is shown by subscripts.

Objective 1, Exercise 2.4

Inside the Atom. Atoms are made up of numerous smaller particles of which the most important to chemical studies are the proton, neutron, and electron. Positively charged protons and neutral neutrons have a relative mass of 1 u each and are located in the nuclei of atoms. Negatively charged electrons with a mass of 1/1836 u are located outside the nuclei of atoms.

Objective 2, Exercises 2.10 and 2.12

Isotopes. Most elements in their natural state are made up of more than one kind of atom. These different kinds of atoms of a specific element are called isotopes and differ from one another only in the number of neutrons in their nuclei. A symbol incorporating atomic number, mass number, and elemental symbol is used to represent specific isotopes.

Objective 3, Exercises 2.16 and 2.22

Relative Masses of Atoms and Molecules. Relative masses called atomic weights have been assigned to each element and are tabulated in the periodic table. The units used are atomic mass units, abbreviated u. Relative masses for molecules, called molecular weights, are determined by adding the atomic weights of the atoms making up the molecules.

Objective 4, Exercise 2.32

Isotopes and Atomic Weights. The atomic weights measured for elements are average weights that depend on the percentages and masses of the isotopes in the naturally occurring element. If the isotope percent abundances and isotope masses are known for an element, its atomic weight can be calculated.

Objective 5, Exercise 2.38

Avogadro's Number: The Mole. Avogadro's number of the atoms of an element has a mass in grams equal to the atomic weight of the element. Avogadro's number of molecules has a mass in grams equal to the molecular weight. Avogadro's number of particles is called a mole, abbreviated mol.

Objective 6, Exercises 2.44 a & b and 2.46 a & b

The Mole and Chemical Formulas. The mole concept when applied to molecular formulas gives numerous relationships that yield useful factors for factor-unit calculations.

Objective 7, Exercises 2.50 b and 2.52 b

Key Terms and Concepts

Atomic mass unit (u) (2.4)
Atomic number of an atom (2.3)
Atomic weight (2.4)
Compound formula (2.1)
Elemental symbol (2.1)
Isotopes (2.3)
Mass number of an atom (2.3)
Mole (2.6)
Molecular weight (2.4)
Nucleus (2.2)

Exercises

OWL Interactive versions of these problems are assignable in OWL.

Even-numbered exercises are answered in Appendix B.

Blue-numbered exercises are more challenging.

You will find it useful to refer to Table 2.1 and the periodic table inside the front cover as you work these exercises.

Symbols and Formulas (Section 2.1)

2.1 Draw a "formula" for each of the following molecules using circular symbols of your choice to represent atoms:

- **a.** A diatomic molecule of an element
- **b.** A diatomic molecule of a compound
- **c.** A triatomic molecule of an element
- **d.** A molecule of a compound containing one atom of one element and four atoms of another element

2.2 Draw a "formula" for each of the following molecules using circular symbols of your choice to represent atoms:

- **a.** A triatomic molecule of a compound
- **b.** A molecule of a compound containing two atoms of one element and two atoms of a second element
- **c.** A molecule of a compound containing two atoms of one element, one atom of a second element, and four atoms of a third element
- **d.** A molecule containing two atoms of one element, six atoms of a second element, and one atom of a third element

2.3 Write formulas for the following molecules using elemental symbols from Table 2.1 and subscripts. Compare these formulas with those of Exercise 2.1.

- **a.** A diatomic molecule of fluorine gas
- **b.** A diatomic molecule of hydrogen chloride (one hydrogen atom and one chlorine atom)

c. A triatomic molecule of ozone (a molecular form of the element oxygen)

d. A molecule of methane (one carbon atom and four hydrogen atoms)

2.4 Write formulas for the following molecules using elemental symbols from Table 2.1 and subscripts. Compare these formulas with those of Exercise 2.2.

a. A molecule of water (two hydrogen atoms and one oxygen atom)

b. A molecule of hydrogen peroxide (two hydrogen atoms and two oxygen atoms)

c. A molecule of sulfuric acid (two hydrogen atoms, one sulfur atom, and four oxygen atoms)

d. A molecule of ethyl alcohol (two carbon atoms, six hydrogen atoms, and one oxygen atom)

2.5 Determine the number of each type of atom in molecules represented by the following formulas:

a. nitrous acid (HNO_2)

b. chlorine dioxide (ClO_2)

c. ethyl alcohol (C_2H_6O)

d. chloroform ($CHCl_3$)

2.6 Determine the number of each type of atom in molecules represented by the following formulas:

a. sulfur dioxide (SO_2)

b. butane (C_4H_{10})

c. chlorous acid ($HClO_2$)

d. boron triflouride (BF_3)

2.7 Tell what is wrong with each of the following molecular formulas and write a correct formula:

a. H3PO3 (phosphorous acid)

b. $SICI_4$ (silicon tetrachloride)

c. SOO (sulfur dioxide)

d. 2HO (hydrogen peroxide—two hydrogen atoms and two oxygen atoms)

2.8 Tell what is wrong with each of the following formulas and write a correct formula:

a. HSH (hydrogen sulfide)

b. $HCLO_2$ (chlorous acid)

c. $2HN_2$ (hydrazine—two hydrogen atoms and four nitrogen atoms)

d. C2H6 (ethane)

Inside The Atom (Section 2.2)

2.9 Determine the charge and mass (in u) of nuclei made up of the following particles:

a. 3 protons and 4 neutrons

b. 10 protons and 12 neutrons

c. 35 protons and 46 neutrons

d. 56 protons and 81 neutrons

2.10 Determine the charge and mass (in u) of nuclei made up of the following particles:

a. 5 protons and 6 neutrons

b. 11 protons and 10 neutrons

c. 36 protons and 50 neutrons

d. 50 protons and 68 neutrons

2.11 Determine the number of electrons that would have to be associated with each nucleus described in Exercise 2.9 to produce a neutral atom.

2.12 Determine the number of electrons that would have to be associated with each nucleus described in Exercise 2.10 to produce a neutral atom.

Isotopes (Section 2.3)

2.13 Determine the number of electrons and protons contained in an atom of the following elements:

a. sulfur

b. As

c. element number 24

2.14 Determine the number of electrons and protons contained in an atom of the following elements:

a. silicon

b. Sn

c. element number 74

2.15 Determine the number of protons, number of neutrons, and number of electrons in atoms of the following isotopes:

a. ${}^{25}_{12}Mg$

b. ${}^{13}_{6}C$

c. ${}^{41}_{19}K$

2.16 Determine the number of protons, number of neutrons, and number of electrons in atoms of the following isotopes:

a. ${}^{7}_{3}Li$

b. ${}^{22}_{10}Ne$

c. ${}^{44}_{20}Ca$

2.17 Write symbols like those given in Exercises 2.15 and 2.16 for the following isotopes:

a. cadmium-110

b. cobalt-60

c. uranium-235

2.18 Write symbols like those given in Exercises 2.15 and 2.16 for the following isotopes:

a. silicon-28

b. argon-40

c. strontium-88

2.19 Determine the mass number and atomic number for atoms containing the nuclei described in Exercise 2.9. Write symbols for each atom like those given in Exercises 2.15 and 2.16.

2.20 Determine the mass number and atomic number for neutral atoms containing the nuclei described in Exercise 2.10. Write symbols for each atom like those given in Exercises 2.15 and 2.16.

2.21 Write isotope symbols for neutral atoms with the following characteristics:

a. Contains 15 electrons and 16 neutrons

b. A radon atom with a mass number of 211

c. An oxygen atom that contains 10 neutrons

2.22 Write isotope symbols for neutral atoms with the following characteristics:

a. Contains 17 electrons and 20 neutrons

b. A copper atom with a mass number of 65

c. A zinc atom that contains 36 neutrons

Relative Masses of Atoms and Molecules (Section 2.4)

2.23 Write the symbols and names for two elements whose average atoms have masses that are within 0.3 u of each other. Don't look beyond element number 83.

2.24 Round atomic weights to the nearest whole number, and determine how many helium atoms would balance one carbon atom on a see-saw.

2.25 Round atomic weights to the nearest whole number, and determine how many lithium atoms would balance two nitrogen atoms on a see-saw.

2.26 What are the symbol and name for an element whose average atoms have a mass that is 77.1% of the mass of an average chromium atom?

2.27 In the first 36 elements, 6 elements have atoms whose average mass is within 0.2 u of being twice the atomic number of the element. Write the symbols and names for these 6 elements.

2.28 What are the symbol and name of the element whose average atoms have a mass very nearly half the mass of an average silicon atom?

2.29 Determine the molecular weights of the following in u:

a. oxygen (O_2)

b. carbon monoxide (CO)

c. chloric acid ($HClO_3$)

d. glycerine ($C_3H_8O_3$)

e. sulfur dioxide (SO_2)

2.30 Determine the molecular weights of the following in u:

a. nitrogen dioxide (NO_2)

b. ammonia (NH_3)

c. glucose ($C_6H_{12}O_6$)

d. ozone (O_3)

e. ethylene glycol ($C_2H_6O_2$)

2.31 The molecular weight was determined for a gas that is known to be an oxide of nitrogen. The value obtained experimentally was 43.98 u. Which of the following is most likely to be the formula of the gas? NO, N_2O, NO_2.

2.32 A flammable gas is known to contain only carbon and hydrogen. Its molecular weight is determined and found to be 28.05 u. Which of the following is the likely identity of the gas? acetylene (C_2H_2), ethylene (C_2H_4), ethane (C_2H_6).

2.33 Glycine, an amino acid found in proteins, has a molecular weight of 75.07 u and is represented by the formula $C_2H_xNO_2$. What number does x stand for in the formula?

2.34 Serine, an amino acid found in proteins, has a molecular weight of 105.10 u and is represented by the formula $C_yH_7NO_3$. What number does y stand for in the formula?

Isotopes and Atomic Weights (Section 2.5)

2.35 Naturally occurring beryllium has a single isotope. Determine the following for the naturally occurring atoms of beryllium:

a. The number of neutrons in the nucleus

b. The mass (in u) of the nucleus (to three significant figures)

2.36 Naturally occurring aluminum has a single isotope. Determine the following for the naturally occurring atoms of aluminum:

a. The number of neutrons in the nucleus

b. The mass (in u) of the nucleus (to three significant figures)

2.37 Calculate the atomic weight of lithium on the basis of the following percent composition and atomic weights of the naturally occurring isotopes. Compare the calculated value with the atomic weight listed for lithium in the periodic table.

$$\text{lithium-6} = 7.42\%\ (6.0151\ \text{u})$$
$$\text{lithium-7} = 92.58\%\ (7.0160\ \text{u})$$

2.38 Calculate the atomic weight of gallium on the basis of the following percent composition and atomic weights of the naturally occurring isotopes. Compare the calculated value with the atomic weight listed for gallium in the periodic table.

$$\text{gallium-69} = 60.40\%\ (68.9257\ \text{u})$$
$$\text{gallium-71} = 39.60\%\ (70.9249\ \text{u})$$

2.39 Calculate the atomic weight of silicon on the basis of the following percent composition and atomic weights of the naturally occurring isotopes. Compare the calculated value with the atomic weight listed for silicon in the periodic table.

$$\text{silicon-28} = 92.21\%\ (27.9769\ \text{u})$$
$$\text{silicon-29} = 4.70\%\ (28.9765\ \text{u})$$
$$\text{silicon-30} = 3.09\%\ (27.9738\ \text{u})$$

2.40 Calculate the atomic weight of copper on the basis of the following percent composition and atomic weights of the naturally occurring isotopes. Compare the calculated value with the atomic weight listed for copper in the periodic table.

$$\text{copper-63} = 69.09\%\ (62.9298\ \text{u})$$
$$\text{copper-65} = 30.91\%\ (64.9278\ \text{u})$$

Avogadro's Number: The Mole (Section 2.6)

2.41 Refer to the periodic table and determine how many grams of phosphorus contain the same number of atoms as 0.12 g of carbon.

2.42 Refer to the periodic table and determine how many grams of fluorine contain the same number of atoms as 1.60 g of oxygen.

2.43 Write three relationships (equalities) based on the mole concept for each of the following elements:

a. potassium

b. magnesium

c. tin

2.44 Write three relationships (equalities) based on the mole concept for each of the following elements:

a. phosphorus

b. aluminum

c. krypton

2.45 Use a factor derived from the relationships written in Exercise 2.43 and the factor-unit method to determine the following:

a. The number of moles of potassium atoms in a 50.0-g sample of potassium

b. The number of magnesium atoms in a 1.82-mol sample of magnesium

c. The number of tin atoms in a 200-g sample of tin

2.46 Use a factor derived from the relationships written in Exercise 2.44 and the factor-unit method to determine the following:

a. The mass in grams of one phosphorus atom

b. The number of grams of aluminum in 1.65 mol of aluminum

c. The total mass in grams of one-fourth Avogadro's number of krypton atoms

The Mole and Chemical Formulas (Section 2.7)

2.47 Refer to the periodic table and calculate the molecular weights for the compounds PH_3 and SO_2. Then, determine how many grams of PH_3 contain the same number of molecules as 6.41 g of SO_2.

2.48 Refer to the periodic table and calculate the molecular weights for the compounds BF_3 and H_2S. Then, determine how many grams of BF_3 contain the same number of molecules as 0.34 g of H_2S.

2.49 For each formula given below, write statements equivalent to Statements 1–6 (see Section 2.7):

a. methane (CH_4)

b. ammonia (NH_3)

c. chloroform ($CHCl_3$)

2.50 For each formula given below, write statements equivalent to Statements 1–6 (see Section 2.7):

a. ethyl ether ($C_4H_{10}O$)

b. fluoroacetic acid ($C_2H_3O_2F$)

c. aniline (C_6H_7N)

2.51 Answer the following questions based on information contained in the statements you wrote for Exercise 2.49.

a. How many moles of hydrogen atoms are contained in 1 mol of CH_4 molecules?

b. How many grams of nitrogen are contained in 1.00 mol of NH_3?

c. What is the mass percentage of chlorine in $CHCl_3$?

2.52 Answer the following questions based on information contained in the statements you wrote for Exercise 2.50.

a. How many moles of hydrogen atoms are contained in 0.50 mol of ethyl ether?

b. How many carbon atoms are contained in 0.25 mol of $C_2H_3O_2F$?

c. How many grams of hydrogen are contained in 2.00 mol of C_6H_7N?

2.53 How many moles of $C_4H_{10}O$ contain the same number of carbon atoms as 1 mol of $C_2H_3O_2F$?

2.54 How many grams of C_2H_6O contain the same number of oxygen atoms as 0.75 mol of H_2O?

2.55 Determine the mass percentage of nitrogen in N_2O and NO_2.

2.56 Determine the mass percentage of oxygen in CO and CO_2.

2.57 Any of the statements based on a mole of substance (Statements 4–6) can be used to obtain factors for problem solving by the factor-unit method. Write statements equivalent to 4, 5, and 6 for nitrophenol ($C_6H_5NO_3$). Use a single factor obtained from the statements to solve each of the following. A different factor will be needed in each case.

a. How many grams of nitrogen are contained in 70.0 g of $C_6H_5NO_3$?

b. How many moles of oxygen atoms are contained in 1.50 mol of $C_6H_5NO_3$?

c. How many atoms of carbon are contained in 9.00×10^{22} molecules of $C_6H_5NO_3$?

2.58 Any of the statements based on a mole of substance (Statements 4–6) can be used to obtain factors for problem solving by the factor-unit method. Write statements equivalent to 4, 5, and 6 for fructose ($C_6H_{12}O_6$). Use a single factor obtained from the statements to solve each of the following. A different factor will be needed in each case.

a. How many grams of oxygen are contained in 43.5 g of $C_6H_{12}O_6$?

b. How many moles of hydrogen atoms are contained in 1.50 mol of $C_6H_{12}O_6$?

c. How many atoms of carbon are contained in 7.50×10^{22} molecules of $C_6H_{12}O_6$?

2.59 Urea (CH_4N_2O) and ammonium sulfate ($N_2H_8SO_4$) are both used as agricultural fertilizers. Which one contains the higher mass percentage of nitrogen?

2.60 Two iron ores that have been used as sources of iron are magnetite (Fe_3O_4) and hematite (Fe_2O_3). Which one contains the higher mass percentage of iron?

2.61 Both calcite ($CaCO_3$) and dolomite ($CaMgC_2O_6$) are used as dietary calcium supplements. Calculate the mass percentage of calcium in each mineral.

Additional Exercises

2.62 The two major isotopes of uranium metal found in nature are U-235 and U-238. Which isotope has the greater density? Explain your reasoning.

2.63 About one billion (1.0×10^9) peas can fit into a railroad car. What percentage of one mole of peas is this number?

2.64 The mass of a single carbon-12 atom is 1.99×10^{-23} g. What is the mass in grams of a single carbon-14 atom?

2.65 One mole of water molecules, H_2O, has a mass of 18.0 g. What would be the mass in grams of one mole of heavy water molecules, D_2O, where D represents the ${}^{2}_{1}H$ isotope?

2.66 According to the caption of Figure 2.2, atoms are composed primarily of empty space (99.999% empty). What would happen to the density of matter (increase or decrease) if the electrons were actually located at the distance from the nucleus shown in the atom drawing of Figure 2.2? Explain your reasoning.

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5. *Peterson's Master the Nursing School and Allied Health Entrance Exams,* 18th Edition by Marion F. Gooding © 2008, Peterson's, a Nelnet Company.

2.67 The symbol K on the periodic table stands for

a. potassium

b. calcium

c. carbon

d. phosphorus

2.68 Which one of the following substances is a chemical compound?

a. blood

b. water

c. oxygen

d. air

2.69 Which of the following is true about compounds?

a. Compounds are pure substances that are composed of two or more elements in a fixed proportion.

b. Compounds can be broken down chemically to produce their constituent elements or other compounds.

c. Both A and B are correct.

d. Neither A nor B is correct.

2.70 ${}^{34}_{17}CI$ has

a. 17 protons, 17 electrons, and 17 neutrons.

b. 17 protons, 19 electrons, and 17 neutrons.

c. 17 protons, 18 electrons, and 17 neutrons.

d. 34 protons, 34 electrons, and 17 neutrons.

2.71 If two atoms are isotopes, they will

a. have the same number of protons and neutrons

b. have the same number of neutrons, but different numbers of protons

c. have the same number of protons, but different numbers of neutrons

d. have the same number of neutrons and electrons

2.72 Copper (Cu) has an atomic number of 29 and a mass number of 64. One copper atom, therefore, has how many protons?

a. 27

b. 29

c. 31

d. 35

2.73 Atoms are electrically neutral. This means that an atom will contain

a. more protons than neutrons

b. more electrons than protons

c. an equal number of protons and electrons

d. None of the above

2.74 The negative charged particle found within the atom is the

a. proton.

b. electron.

c. nucleus.

d. neutron.

2.75 Two atoms, L and M, are isotopes. Which of the following properties would they NOT have in common?

a. atomic number

b. atomic weight

c. chemical reactivity

d. the number of protons in the nucleus

2.76 The major portion of an atom's mass consists of:

a. neutrons and protons

b. electrons and protons

c. electrons and neutrons

d. neutrons and positrons

2.77 The mass of an atom is almost entirely contributed by its

a. nucleus

b. protons

c. electrons and protons

d. neutrons

2.78 Which of the following is the chemical symbol for the species that has 16 protons, 17 neutrons, and 18 electrons?

a. ${}^{33}_{16}S$

b. ${}^{33}_{17}Cl$

c. ${}^{35}_{17}Cl$

d. ${}^{33}_{16}S^{2-}$

2.79 An atom with an atomic number of 58 and an atomic mass of 118 has

a. 58 neutrons

b. 176 neutrons

c. 60 neutrons

d. 116 neutrons

2.80 What is the mass number of an atom with 60 protons, 60 electrons, and 75 neutrons?

a. 120

b. 135

c. 75

d. 195

2.81 Which of the following represents Avogadro's number?

a. 1.66×10^{-24}

b. 1.0×10^{24}

c. 6.022×10^{23}

d. 3.011×10^{23}

2.82 Which of the following has the greatest number of atoms?

a. 1.0 mol N

b. 1.0 g N

c. 1.0 mol NO_2

d. 0.5 mol NH_3

2.83 The formula of carbon dioxide is CO_2. Its molecular weight is 44 amu. A sample of 11 grams of CO_2 contains

a. 1.0 mole of carbon dioxide

b. 1.5 grams of carbon

c. 3.0 grams of carbon

d. 6.0 grams of oxygen

2.84 What is the molar mass of calcium oxide, CaO?

a. 56

b. 28

c. 640

d. 320

2.85 How many grams are contained in 0.200 mol of calcium phosphate, $Ca_3(PO_4)_2$?

a. 6.20

b. 62.0

c. 124

d. 31.0

2.86 How many moles are contained in a 54.0 g sample of Al?

a. 1.0

b. 2.0

c. 0.5

d. 4.0

Chemistry for Thought

2.87 **a.** Explain how atoms of different elements differ from one another.

b. Explain how atoms of different isotopes of the same element differ from one another.

2.88 The atomic weight of aluminum is 26.98 u, and the atomic weight of nickel is 58.69 u. All aluminum atoms have a mass of 26.98 u, but not a single atom of nickel has a mass of 58.69 u. Explain.

2.89 Answer the question in the caption of Figure 2.3. Would you expect any orange in the bowl to have the exact mass you calculated as an average? Explain.

2.90 Answer the question in the caption of Figure 2.5. Use your answer and the fact that an average jelly bean has a mass of 1.18 g to calculate the number of beans in each jar.

2.91 Answer the question in the caption of Figure 2.6. How many CS_2 molecules would be required to contain 0.25 mol of sulfur atoms?

2.92 In Section 2.4 it was pointed out that an atomic mass unit, u, is equal to $\frac{1}{12}$ the mass of an atom of carbon-12. Suppose one atomic mass unit was redefined as being equal to $\frac{1}{24}$ the mass of a carbon-12 atom. How would this change influence the value of the atomic weight of magnesium?

2.93 How would the change of question 2.92 influence the ratio of the atomic weight of magnesium divided by the atomic weight of hydrogen?

2.94 How would the change of question 2.92 influence the value of Avogadro's number?

3

Electronic Structure and the Periodic Law

Optometrists are health care professionals devoted to the care of eyes. An examination of the interior of a patient's eye is used to detect conditions such as a detached retina that can be corrected by using laser light produced when electrons change their locations inside atoms.

Learning Objectives

When you have completed your study of this chapter, you should be able to:

1. Locate elements in the periodic table on the basis of group and period designations. **(Section 3.1)**
2. Determine the number of electrons in designated atomic orbitals, subshells, or shells. **(Section 3.2)**
3. Determine the number of valence shell electrons and the electronic structure for atoms, and relate this information to the location of elements in the periodic table. **(Section 3.3)**
4. Determine the following for elements: the electronic configuration of atoms, the number of unpaired electrons in atoms, and the identity of atoms based on provided electronic configurations. **(Section 3.4)**
5. Determine the shell and subshell locations of the distinguishing electrons in elements, and based on their location in the periodic table, classify elements into the categories given in Figures 3.10 and 3.12. **(Section 3.5)**
6. Recognize property trends of elements within the periodic table, and use the trends to predict selected properties of the elements. **(Section 3.6)**

OWL Online homework for this chapter may be assigned in OWL.

In Chapter 1, we defined atoms as particles that represent the limit of chemical subdivision. According to this idea, atoms of a specific element cannot be divided into smaller particles or converted into atoms of another element by any physical or chemical change. Then in Chapter 2, we introduced the idea that atoms are, in fact, made up of particles that are smaller than the atoms themselves. Two of these particles, protons and neutrons, form the nuclei of atoms, whereas electrons are located outside the nuclei.

The idea that atoms are made up of subatomic particles implies that it should be possible to obtain even smaller particles from atoms. Scientists have found that it is possible. During chemical changes, electrons are transferred from one atom to another or are shared between atoms. Some details of these processes are given in Chapter 4, but it is known that they depend on the arrangements of the electrons around the nuclei of atoms. These electronic arrangements are one of the major topics of this chapter.

3.1 The Periodic Law and Table

Learning Objective

1. Locate elements in the periodic table on the basis of group and period designations.

By the early 19th century, detailed studies of the elements known at that time had produced an abundance of chemical information. Scientists looked for order in these facts, with the hope of providing a systematic approach to the study of chemistry. Two scientists independently, and almost simultaneously, made the same important contribution to this end. Julius Lothar Meyer, a German, and Dmitri Mendeleev, a Russian, each produced classification schemes for the elements in 1869. Both schemes were based on the **periodic law,** which in its present form is stated as follows: When all the elements are arranged in order of increasing atomic numbers, elements with similar chemical properties will occur at regular (periodic) intervals.

A convenient way to compactly represent such behavior is to use tables. The arrangement of the elements in a table based on the periodic law is called a *periodic table.* In a modern periodic table, such as the one inside the front cover of this book, elements with similar chemical properties are found in vertical columns called **groups** or **families.**

The groups are designated by a roman numeral and a letter at the top of each column. These group designations have not been universally accepted by chemists throughout the world. An effort has been under way since 1979 to establish a universally acceptable group designation. The simple numerical designation given in parentheses over the traditional designation appears to be the one that will be adopted. In this book, references to groups will be given using both designations, with the new one in parentheses. The horizontal rows in the table are called **periods** and are numbered from top to bottom. Thus, each element belongs to both a period and a group of the periodic table.

periodic law A statement about the behavior of the elements when they are arranged in a specific order. In its present form, it is stated as follows: Elements with similar chemical properties occur at regular (periodic) intervals when the elements are arranged in order of increasing atomic numbers.

group or family of the periodic table A vertical column of elements that have similar chemical properties.

period of the periodic table A horizontal row of elements.

Example 3.1

Identify the group and period to which each of the following belongs:

a. P **b.** Cr **c.** element number 30 **d.** element number 53

Solution

a. Phosphorus (P) is in group VA(15) and period 3.
b. Chromium (Cr) is in group VIB(6) and period 4.
c. The element with atomic number 30 is zinc (Zn), which is found in group IIB(12) and period 4.
d. Element number 53 is iodine (I), found in group VIIA(17) and period 5.

Learning Check 3.1 Write the symbol for the element found in the following places of the periodic table:

a. Group IVA(14) and period 4 **b.** Group VIIB(7) and period 6

It should be noted that the periodic table as given inside the front cover appears to violate the practice of arranging the elements according to increasing atomic number. Element 72 follows element 57, and 104 follows 89, whereas elements 58–71 and 90–103 are arranged in two rows at the bottom of the table. Technically, these two rows should be included in the body of the table as shown in Figure 3.1 (top). To save horizontal space, they are placed in the position shown in Figure 3.1 (bottom). This exception presents no problem, as long as it is understood.

Example 3.2

a. How many elements are found in period 6 of the periodic table?
b. How many elements are found in group VA(15) of the periodic table?

Solution

a. Period 6 includes elements 55–86, even though elements 58–71 are shown below the main table. Therefore, period 6 contains 32 elements.
b. A count shows group VA(15) contains 6 elements: N, P, As, Sb, Bi, and number 115.

Learning Check 3.2 How many elements are found in the following?

a. Period 1 of the periodic table
b. Group IIB(12) of the periodic table

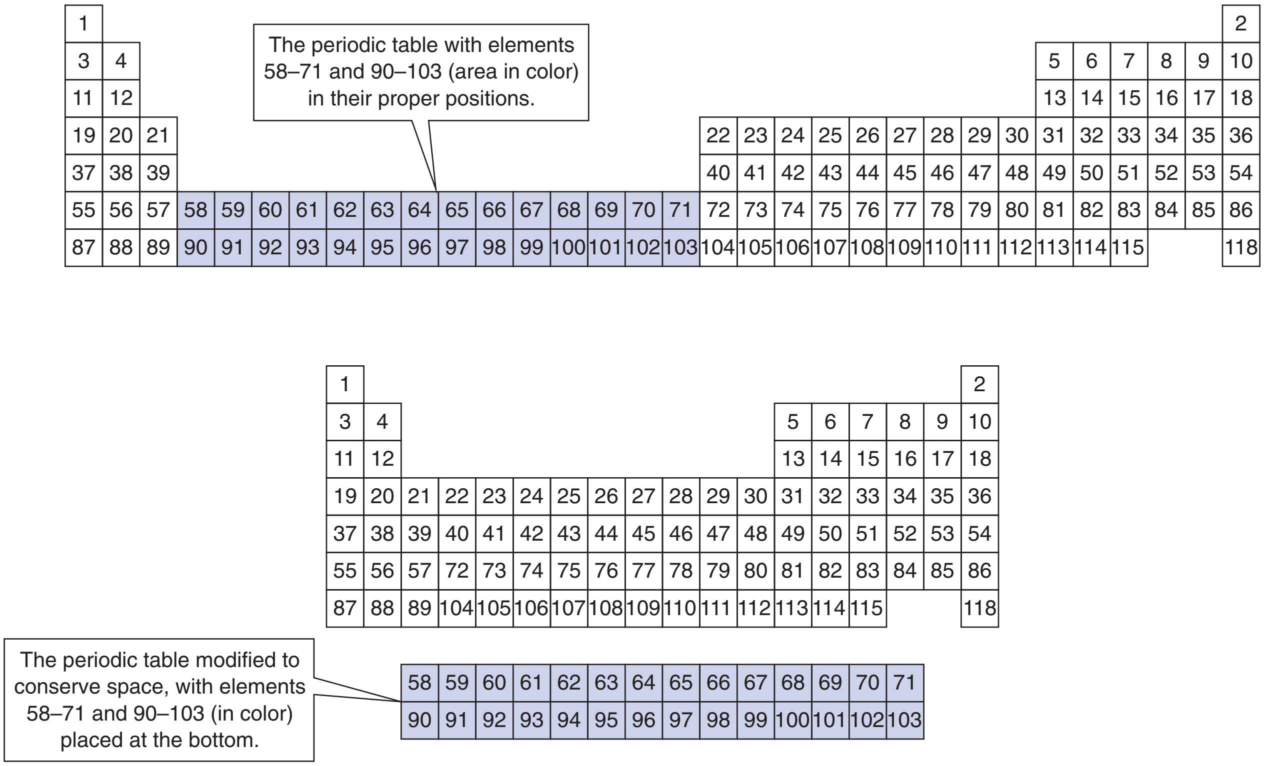

Figure 3.1 Forms of the periodic table.

At the Counter 3.1

Zinc for Colds? The Jury Is Still Out

Zinc, element number 30 of the periodic table, might be a key to relieving the cold symptoms suffered by millions every year. In preliminary studies, it has been shown that compounds of zinc have the ability to inhibit the reproduction of viruses and possibly to promote the body's production of interferon, a virus-fighter. In an attempt to determine the effectiveness of zinc compounds against the viruses that cause the common cold, 100 adult patients were studied at the Cleveland Clinic. The patients were given lozenges within 24 hours of contracting a cold. During their waking hours, the patients dissolved a lozenge in their mouths every 2 hours. This treatment was continued until the patients no longer showed any cold symptoms. Some of the patients were given lozenges that contained zinc gluconate, while others were given placebo (non-zinc-containing) lozenges.

The results of the study showed that cold symptoms lasted an average of 4.4 days in the patients who received the zinc gluconate, compared to an average duration of 7.6 days in the patients who did not receive the zinc compound. The sales of over-the-counter zinc products, especially zinc lozenges, skyrocketed following the publication of these results in 1996. However, this enthusiasm was tempered a bit by the results of another study published in 1998. In this study, the effects of zinc lozenges on the cold symptoms of 249 students in grades 1 through 12 were investigated. Researchers concluded that the zinc lozenges were not effective against cold symptoms in children and teenagers. Various explanations have been proposed for the discrepancy in the results of the two studies, but it has been generally concluded that further studies are needed to clarify what role, if any, zinc compounds may play in treating cold symptoms.

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Despite conflicting research results about their effectiveness, many brands of zinc lozenges are available for use in treating the common cold.

3.2 Electronic Arrangements in Atoms

Learning Objective

2. Determine the number of electrons in designated atomic orbitals, subshells, or shells.

In 1913 Niels Bohr, a young Danish physicist, made an important contribution to our understanding of atomic structure. He was working under the direction of Ernest Rutherford, a British scientist, who had proposed a solar system model for atoms in which negative electrons moved in circular orbits around the positive nucleus, much like the planets move around the sun. Bohr built on this model by proposing that the single electron of a hydrogen atom could occupy orbits only at specific distances from the nucleus, and thus the electron could have only specific energies (see Figure 3.2). Bohr further proposed that the electron changed orbits only by absorbing or releasing energy. The addition of energy to a hydrogen atom elevated the electron to a higher-energy orbit located farther from the nucleus. The energy released when an electron dropped from a higher- to a lower-energy orbit appeared as emitted light.

Research into the nature of atoms continued after Bohr's proposal, and in 1926 a revised model of atomic structure was proposed by Erwin Schrödinger, an Austrian physicist, who received the Nobel Prize in physics in 1933 in recognition of this achievement. According to Schrödinger's quantum mechanical model, the precise paths of electrons cannot be determined accurately, as Bohr's model required. It was found that the location and energy of electrons around a nucleus can be specified using three terms: shell, subshell, and orbital. This is somewhat like locating an individual in a city by specifying a street, building, and apartment.

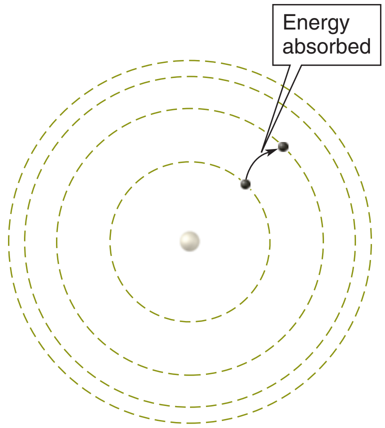


Figure 3.2 A diagram of the Bohr hydrogen atom (not drawn to scale; the orbits are actually much larger than the nucleus). The electron is elevated to a higher-energy orbit when energy is absorbed.

shell A location and energy of electrons around a nucleus that is designated by a value for n, where $n = 1, 2, 3$, etc.

subshell A component of a shell that is designated by a letter from the group *s*, *p*, *d*, and *f*.

atomic orbital A volume of space around atomic nuclei in which electrons of the same energy move. Groups of orbitals with the same n value form subshells.

The location of electrons in a **shell** is indicated by assigning a number, n, to the shell and to all the electrons within the shell. The n value of the lowest-energy shell is 1, that of the next higher energy is 2, the next is 3, and so on. Higher n values for a shell correspond to higher energies and greater distances from the nucleus for the electrons of the shell. Electrons in the third shell all have an n value of 3, all have an energy higher than the energies of electrons in shells 1 and 2, and also are located farther from the nucleus than those of shells 1 and 2.

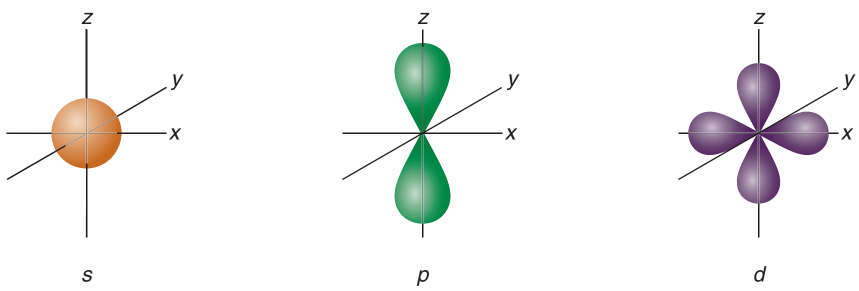
Each shell is made up of **subshells** that are designated by a letter from the group *s, p, d,* and *f.* Because all subshells are designated by one of these letters regardless of the shell in which the subshell is found, a combination of both shell number and subshell letter is used to identify subshells clearly. Thus, a *p* subshell in shell number 2 is referred to as a 2*p* subshell. The number of subshells found in a shell is the same as the value of n for the shell. Thus, shell number 2 ($n = 2$) contains two subshells. The subshells are the 2*p* mentioned earlier and a 2*s*. Electrons located in specific subshells are often referred to in terms of the same number and letter as the subshell. For example, we might refer to an atom as having three 2*p* electrons. All electrons within a specific subshell have the same energy.

The description of the location and energy of electrons moving around a nucleus is completed in the quantum mechanical model by specifying an orbital. Each subshell consists of one or more **atomic orbitals,** which are specific volumes of space around nuclei in which electrons move. These atomic orbitals must not be confused with the fixed electron orbits of the original Bohr theory; they are *not* the same. These volumes of space around nuclei have different shapes, depending on the energy of the electrons they contain (see Figure 3.3). All *s* subshells consist of a single orbital that is also designated by the letter *s* and further identified by the n value of the shell to which the subshell belongs. Thus, the 2*s* subshell mentioned earlier consists of a single 2*s* orbital. All *p* subshells consist of three *p* orbitals that also carry the n value of the shell. Thus, the 2*p* subshell of shell number 2 consists of three 2*p* orbitals. Since all the electrons in a subshell have the same energy, an electron in any one of the three 2*p* orbitals has the same energy, regardless of which orbital of the three it occupies. All *d* subshells contain five orbitals, and all *f* subshells contain seven orbitals. According to the quantum mechanical model, each orbital within a subshell can contain a maximum of two electrons.

The shapes of orbitals such as those given in Figure 3.3 must not be interpreted incorrectly. The fact that *s* orbitals are spherical in shape does not mean that the electrons move around on the spherical surface. According to the quantum mechanical model, electrons in *s* orbitals move around inside the spherical volume of the orbital in paths that cannot be determined. All that can be determined about their behavior within the orbital is the probability of finding them in a specific location, which is what Figure 3.3 depicts. Thus, if it were determined that an electron had a 2% probability of being at a specific location, it would simply mean that the electron could be found at that location within an orbital 2 times out of every 100 times we looked for it there. Similarly, electrons in *p* or *d* orbitals do not move on the surface of the dumbbell- or cloverleaf-shaped orbitals; they move *within* the three-dimensional dumbbell- or cloverleaf-shaped volumes.

The energy of electrons located in an orbital within a subshell and shell is determined by two factors. As mentioned earlier, the higher the value of n, the higher the energy. In addition, if n is the same, but the subshell is different, the energy of the contained electrons increases in the order *s, p, d, f.* Thus, a 3*p* electron (an electron in a 3*p* subshell) has

Figure 3.3 Shapes of typical *s*, *p*, and *d* orbitals.



a higher energy than a 3*s* electron (an electron in a 3*s* subshell). Figure 3.4 is a diagrammatic representation of the fourth shell of an atom completely filled with electrons.

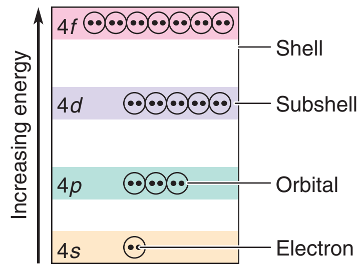


Figure 3.4 The fourth shell of an atom, completely filled with electrons.

Example 3.3

Determine the following for the third shell of an atom:

a. The number of subshells
b. The designation for each subshell
c. The number of orbitals in each subshell
d. The maximum number of electrons that can be contained in each subshell
e. The maximum number of electrons that can be contained in the shell

Solution

a. The number of subshells is the same as the number used to designate the shell. Therefore, the third shell contains three subshells.
b. Subshells increase in energy according to the order *s, p, d, f.* The subshells in the shell are therefore designated 3*s,* 3*p,* and 3*d.*
c. The number of orbitals in the subshells is 1, 3, and 5 because *s* subshells always contain a single orbital. *p* subshells always contain three orbitals, and *d* subshells always contain five orbitals.
d. Each atomic orbital can contain a maximum of two electrons, independent of the type of orbital under discussion. Therefore, the 3*s* subshell (one orbital) can hold a maximum of two electrons, the 3*p* subshell (three orbitals) a maximum of 6 electrons, and the 3*d* subshell (five orbitals) a maximum of 10 electrons.
e. The maximum number of electrons that can be contained in the shell is simply the sum of the maximum number for each subshell, $2 + 6 + 10 = 18$.

Learning Check 3.3 In each of the following, what is the maximum number of electrons that can be found?

a. a 4*p* orbital **b.** a 5*d* subshell **c.** shell number 1

It might seem to you at this point that the modifications to the Bohr theory have created a number of hard-to-remember relationships between shells, subshells, orbitals, and electrons. However, some patterns help make the relationships easy to remember. Table 3.1 summarizes them for the first four shells of an atom.

Table 3.1 Relationships Between Shells, Subshells, Orbitals, and Electrons

Shell Number (*n*)	Number of Subshells in Shell	Subshell Designation	Number of Orbitals in Subshell	Orbital Designation	Maximum Number of Electrons in Subshell	Maximum Number of Electrons in Shell
1	1	1*s*	1	1*s*	2	2
2	2	2*s*	1	2*s*	2	
		2*p*	3	2*p*	6	8
3	3	3*s*	1	3*s*	2	
		3*p*	3	3*p*	6	
		3*d*	5	3*d*	10	18
4	4	4*s*	1	4*s*	2	
		4*p*	3	4*p*	6	
		4*d*	5	4*d*	10	
		4*f*	7	4*f*	14	32

3.3 The Shell Model and Chemical Properties

Learning Objective

3. Determine the number of valence shell electrons and the electronic structure for atoms, and relate this information to the location of elements in the periodic table.

The arrangement of electrons into orbitals, subshells, and shells provides an explanation for the similarities in chemical properties of various elements. Table 3.2 gives the number of electrons in each shell for the first 20 elements of the periodic table.

In Table 3.2, notice that the third shell stops filling when 8 electrons are present, even though the shell can hold a maximum of 18 electrons. The reasons for this are discussed in Section 3.4. Also, note that all the elements in a specific group of the periodic table have the same number of electrons in the outermost occupied shell. This outermost occupied shell (the one of highest energy) is called the **valence shell.** Similarities in elemental chemical properties result from identical numbers of electrons in the valence shells of the atoms (see Figure 3.5).

valence shell The outermost (highest-energy) shell of an element that contains electrons.

Table 3.2 Electron Occupancy of Shells

Element	Belongs to Group	Symbol	Atomic Number	Electrons in Shell Number 1	2	3	4
Hydrogen	IA(1)	H	1	1			
Helium	Noble gas(18)	He	2	2			
Lithium	IA(1)	Li	3	2	1		
Beryllium	IIA(2)	Be	4	2	2		
Boron	IIIA(13)	B	5	2	3		
Carbon	IVA(14)	C	6	2	4		
Nitrogen	VA(15)	N	7	2	5		
Oxygen	VIA(16)	O	8	2	6		
Fluorine	VIIA(17)	F	9	2	7		
Neon	Noble gas(18)	Ne	10	2	8		
Sodium	IA(1)	Na	11	2	8	1	
Magnesium	IIA(2)	Mg	12	2	8	2	
Aluminum	IIIA(13)	Al	13	2	8	3	
Silicon	IVA(14)	Si	14	2	8	4	
Phosphorus	VA(15)	P	15	2	8	5	
Sulfur	VIA(16)	S	16	2	8	6	
Chlorine	VIIA(17)	Cl	17	2	8	7	
Argon	Noble gas(18)	Ar	18	2	8	8	
Potassium	IA(1)	K	19	2	8	8	1
Calcium	IIA(2)	Ca	20	2	8	8	2

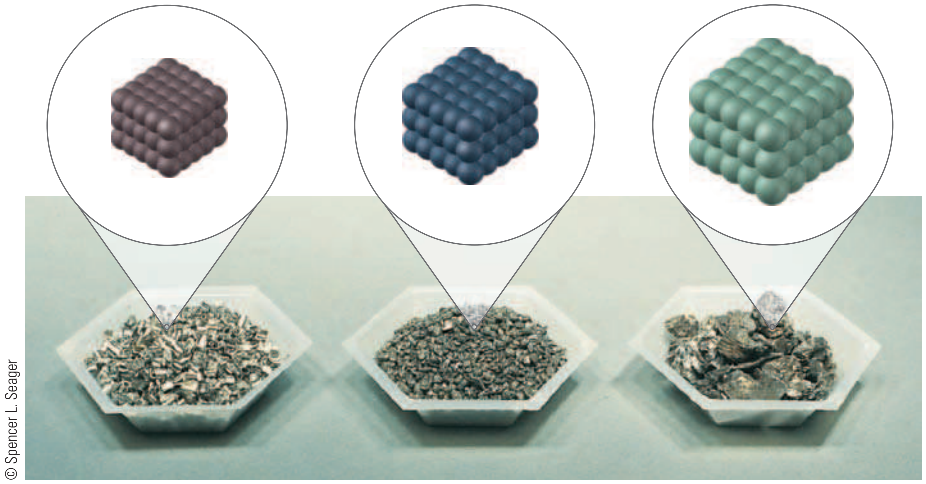


Figure 3.5 Left to right: Magnesium, calcium, and strontium, members of group IIA(2) of the periodic table, have similar chemical properties and appearances.

Example 3.4

Referring to Table 3.2, indicate the number of electrons in the valence shell of elements in groups IA(1), IIA(2), IIIA(13), and IVA(14).

Solution

According to Table 3.2, the elements in group IA(1) are hydrogen, lithium, sodium, and potassium. Each element has one electron in the valence shell. Hydrogen belongs in group IA(1) on the basis of its electronic structure, but its properties differ significantly from other group members.

The group IIA(2) elements are beryllium, magnesium, and calcium. Each has two electrons in the valence shell. The group IIIA(13) elements are boron and aluminum. Both have three electrons in the valence shell. The group IVA(14) elements are carbon and silicon; each has four valence-shell electrons.

Learning Check 3.4 Referring to Table 3.2, indicate the number of electrons in the valence shell of elements in groups VA(15), VIA(16), and VIIA(17), and the noble gases(18).

Example 3.4 and Learning Check 3.4 emphasize the fact that elements belonging to the same periodic table group have the same number of electrons in the valence shell (helium, the first element in the noble gases, is an exception). Notice also that the number of electrons in the valence shell is identical to the roman numeral that designates the group number. Elements of group IIIA(13), for example, have three electrons in the valence shell. It is also apparent that the n value for the valence shell increases by 1 with each heavier member of a group, that is, $n = 2$ for Li, $n = 3$ for Na, and $n = 4$ for K.

Example 3.5

Using the periodic table and Example 3.4 and Learning Check 3.4, determine the n value for the valence shell and the number of electrons in the valence shell for the following elements:

a. Ba **b.** Br **c.** element number 50
d. The third element of group VIA(16)

Solution

a. Ba is the fifth element of group IIA(2), and because $n = 4$ for Ca (the third element of the group), $n = 6$ for the fifth element. The number of valence-shell electrons is two, the same as the group number.
b. In a similar way, Br, the third element of group VIIA(17), has $n = 4$ and seven valence-shell electrons.
c. Element number 50 is tin (Sn), which is the fourth element of group IVA(14). Therefore, $n = 5$ and the number of valence-shell electrons is four.
d. The third element of group VIA(16) is selenium (Se), for which $n = 4$ and the number of valence-shell electrons is six.

Learning Check 3.5 How many electrons will be found in the following?

a. The valence shell of Sr
b. The valence shell of the third element in group IVA(14)
c. The valence shell of the fifteenth element in period 4

3.4 Electronic Configurations

Learning Objective

4. Determine the following for elements: the electronic configuration of atoms, the number of unpaired electrons in atoms, and the identity of atoms based on provided electronic configurations.

electronic configurations The detailed arrangement of electrons indicated by a specific notation, $1s^22s^22p^4$, etc.

Hund's rule A statement of the behavior of electrons when they occupy orbitals: Electrons will not join other electrons in an orbital if an empty orbital of the same energy is available for occupancy.

Pauli exclusion principle A statement of the behavior of electrons when they occupy orbitals: Only electrons spinning in opposite directions can simultaneously occupy the same orbital.

According to Section 3.3, similarities in chemical properties between elements are related to the number of electrons that occupy the valence shells of their atoms. We now look at the electronic arrangements of atoms in more detail. These detailed arrangements, called **electronic configurations,** indicate the number of electrons that occupy each subshell and orbital of an atom.

Imagine that electrons are added one by one to the orbitals that belong to subshells and shells associated with a nucleus. The first electron will go to as low an energy state as possible, which is represented by the $1s$ orbital of the $1s$ subshell of the first shell. The second electron will join the first and completely fill the $1s$ orbital, the $1s$ subshell, and the first shell (remember, an orbital is filled when it contains two electrons). The third electron will have to occupy the lowest-energy subshell ($2s$) of the second shell. The fourth electron will also occupy (and fill) the $2s$ subshell. The fifth electron must seek out the next–highest-energy subshell, which is the $2p$. The $2p$ subshell contains three $2p$ orbitals, so the sixth electron can either join the fifth in one of the $2p$ orbitals or go into an empty $2p$ orbital. It will go into an empty orbital in compliance with **Hund's rule,** which states: Electrons will not join other electrons in an orbital if an empty orbital of the same energy is available.

It has been found that electrons behave as if they spin on an axis, and only electrons spinning in opposite directions (indicated by $\uparrow$ and $\downarrow$) can occupy the same orbital. This principle, known as the **Pauli exclusion principle,** explains why orbitals can contain a maximum of two electrons. Hund's rule and the Pauli exclusion principle can be combined: Electrons will pair with other electrons in an orbital only if there is no empty orbital of the same energy available and if there is one electron with opposite spin already in the orbital.

The seventh added electron will occupy the last empty $2p$ orbital, and the eighth, ninth, and tenth electrons will pair up with electrons already in the $2p$ orbitals. The tenth electron fills the $2p$ subshell, thus completing the second shell. This filling order is illustrated in Figure 3.6.

The eleventh electron, with no empty orbital available that has the same energy as a $2p$ and no partially filled orbitals, will occupy the empty lowest-energy subshell of the third shell. The filling order for electrons beyond the tenth follows the pattern given in

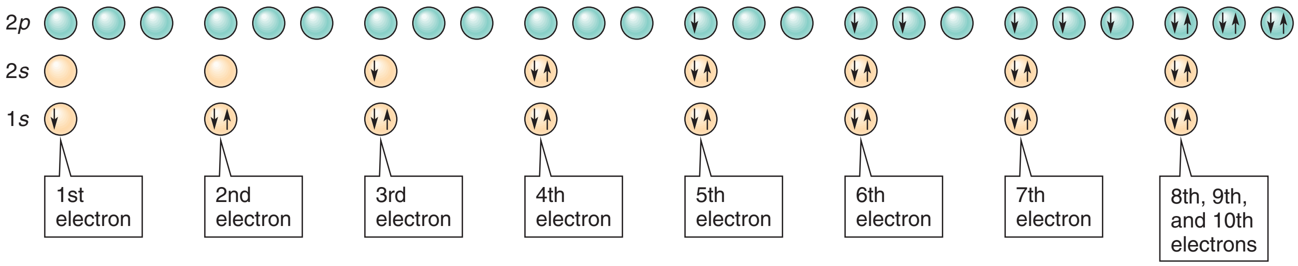


Figure 3.6 The filling order for the first 10 electrons.

Figure 3.7, where shells are indicated by large rectangles, subshells by colored rectangles, and orbitals by circles. The filling order is obtained by following the arrows.

As shown in Figure 3.7, some low-energy subshells of a specific shell have energies lower than the upper subshell of a preceding shell. For example, the 4*s* subshell has a lower energy than, and fills before, the 3*d* subshell. Figure 3.7 indicates that the third shell will not accept more than 8 electrons until the 4*s* subshell is complete. Thus, electrons 21 through 30 go into the 3*d* subshell and complete the filling of the third shell.

It is often convenient to represent the electronic configuration of an atom in a concise way. This is done by writing the subshells in the correct filling order and then indicating the number of electrons in each subshell by a superscript.

Example 3.6

Write the electronic configurations for the following, and indicate the number of unpaired electrons in each case:

a. An atom that contains 7 electrons
b. An atom that contains 17 electrons
c. An atom of element number 22
d. An atom of arsenic (As)

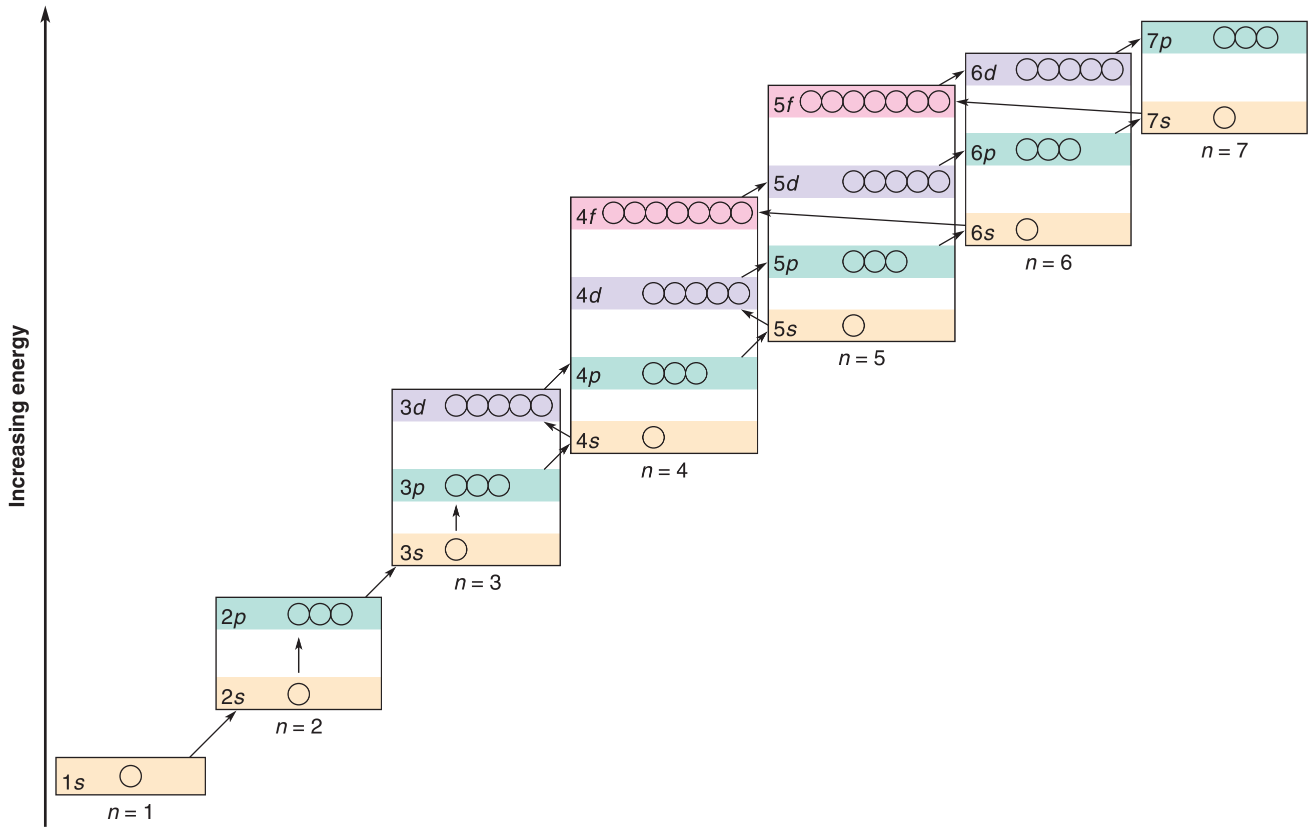


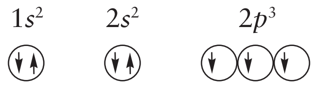
Figure 3.7 The relative energies and electron-filling order for shells and subshells.

Solution

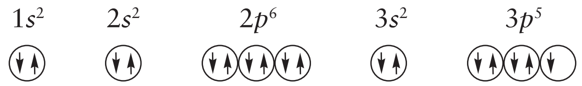
The correct filling order of subshells from Figure 3.7 is

$$1s,\ 2s,\ 2p,\ 3s,\ 3p,\ 4s,\ 3d,\ 4p,\ 5s,\ 4d,\ 5p,\ 6s,\ 4f,\ 5d,\ 6p,\ 7s,\ 5f,\ 6d,\ 7p$$

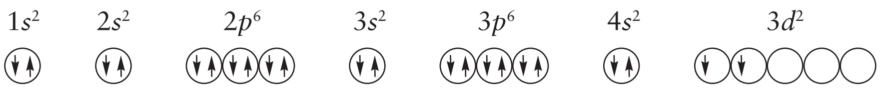
a. Even though *s* subshells can hold 2 electrons, *p* subshells 6, *d* subshells 10, and *f* subshells 14, only enough subshells to hold 7 electrons will be used. Therefore, the configuration is written as shown below, starting on the left, with circles representing orbitals and arrows representing electrons. It is apparent that no subshells beyond 2*p* are needed because that subshell contains only 3 electrons. Note that both the 1*s* and 2*s* subshells are full; the 2*p* subshell is half-full, with one electron in each of the three orbitals (Hund's rule). These three electrons are unpaired.



b. Similarly, the configuration for 17 electrons is shown below. Here the 1*s,* 2*s,* 2*p,* and 3*s* subshells are full, as are the first and second shells. The 3*p* subshell is not full and can accept 1 more electron. One electron in the 3*p* subshell is unpaired.



c. An atom of element number 22 contains 22 protons in the nucleus and must therefore contain 22 electrons. The electronic configuration is



Note here that only two of the 3*d* orbitals of the 3*d* subshell are occupied, and each of these orbitals contains a single electron. Thus, there are two unpaired electrons.

d. Arsenic (As) is element number 33 and therefore contains 33 electrons. The electronic configuration is

$$1s^22s^22p^63s^23p^64s^23d^{10}4p^3$$

This time, the circles and arrows have not been used to indicate orbitals and electrons. You should satisfy yourself that the following facts are clear: The first, second, and third shells are full. The fourth shell is partially full, with the 4*s* subshell being full and the 4*p* subshell being half-full. The 4*p* subshell contains 3 unpaired electrons.

Learning Check 3.6 Write the electronic configurations for the following, and indicate the number of unpaired electrons in each case:

a. Element number 9
b. Mg
c. The element found in group VIA(16) and period 3
d. An atom that contains 23 protons

Although Figure 3.7 gives the details of subshell-filling order, a more concise diagram is available and easy to remember. It is shown in Figure 3.8, where the subshells are first arranged as shown on the left and then diagonal arrows are drawn as shown on the right. To get the correct subshell-filling order, follow the arrows from top to bottom, going from the head of one arrow to the tail of the next.

The electronic configurations described to this point provide details of the shells, subshells, and orbitals involved but are somewhat cumbersome. In some applications, these

Chemistry Around Us 3.1
Nano World

The title given above might remind you of the title of a far out science-fiction movie. However, a quick review of Table 1.2 reminds us that *nano* is a prefix used to designate a quantity that is one-billionth (10^{-9}) the size of another quantity. Thus, a nanometer is a length that is one-billionth the length of a meter, or a length equal to the distance from one end to the other of a row of five to ten atoms laid side by side. We are living in a world that is becoming more and more influenced by the ability of scientists, engineers, and technologists to produce, manipulate, and commercially manufacture particles and other tiny objects whose sizes are measured in nanometers.

The design and creation of materials at this tiny scale is referred to as nanotechnology. Consumer products containing materials produced by nanotechnology began showing up in the mid-1990s. Two common applications at that time were the inclusion of nanometer-sized particles (called nanoparticles) in cosmetics and sunscreen products. Today, nanoparticles are found in a broad range of consumer products ranging from food packaging to sporting goods. The sport of tennis has benefited significantly from nanotechnology. One company injects nanoparticles of silicon dioxide into voids in the graphite frame of their tennis rackets. The result is a stronger frame that allows more power to be delivered to the ball with each stroke. The same company has developed a tennis ball that is coated on the inside with a very thin coating of a clay-like nanoparticle material that makes the balls retain air better, and extends their useful life. Nanoparticle-based textile treatments have revolutionized the textile industry by making possible products such as quick-drying, waterproof, wrinkle-free, and stain-resistant clothing. Other products resulting from nanotechnology include self-cleaning windows, glare-reducing and fog-resistant coatings for eyeglasses and automobile windshields, and lighter, stronger components for automobiles. Despite the fact that nanoparticles cannot be seen with the naked eye, they exert very visible influences on products we all use, and this influence will surely increase with the passage of time.

AP Photo/Tony Avelar

Both the clothing worn by these tennis players and the rackets they are using are improved by nanotechnology.

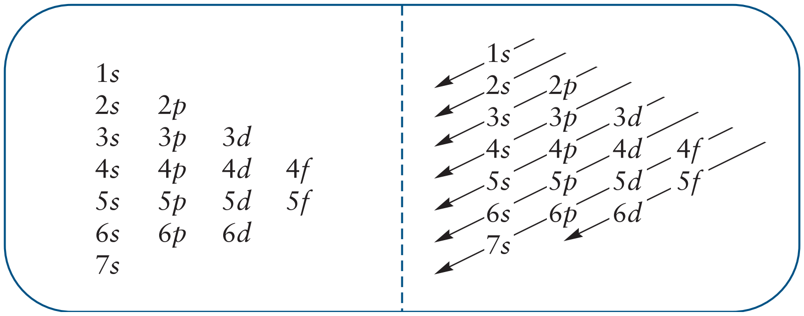


Figure 3.8 An aid to remembering subshell-filling order.

details are not needed, and simplified representations are used that emphasize the electrons in the valence shell.

We see from Table 3.2 that the noble gases neon and argon both have electronic configurations that end with a completely filled *p* subshell. This is true for all noble gases except helium, which ends with a filled 1*s* subshell. In Chapter 4, we will see that these **noble gas configurations** are important in understanding the bonding that occurs between atoms.

noble gas configuration
An electronic configuration in which the last eight electrons occupy and fill the *s* and *p* subshells of the highest-occupied shell.

Noble gas configurations can be used to write abbreviated electronic configurations. Instead of writing the configurations in their entirety, the symbols for the noble gases are used in brackets to represent the electrons found in their configurations. Electrons that are present in addition to those of the noble gases are written following the symbol. For example, the electronic configuration for sodium can be represented as $1s^22s^22p^63s^1$ or $[Ne]3s^1$.