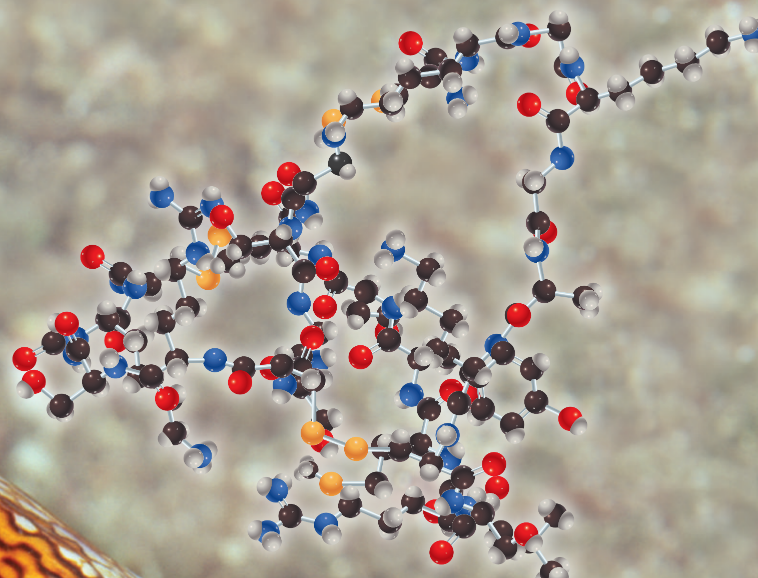
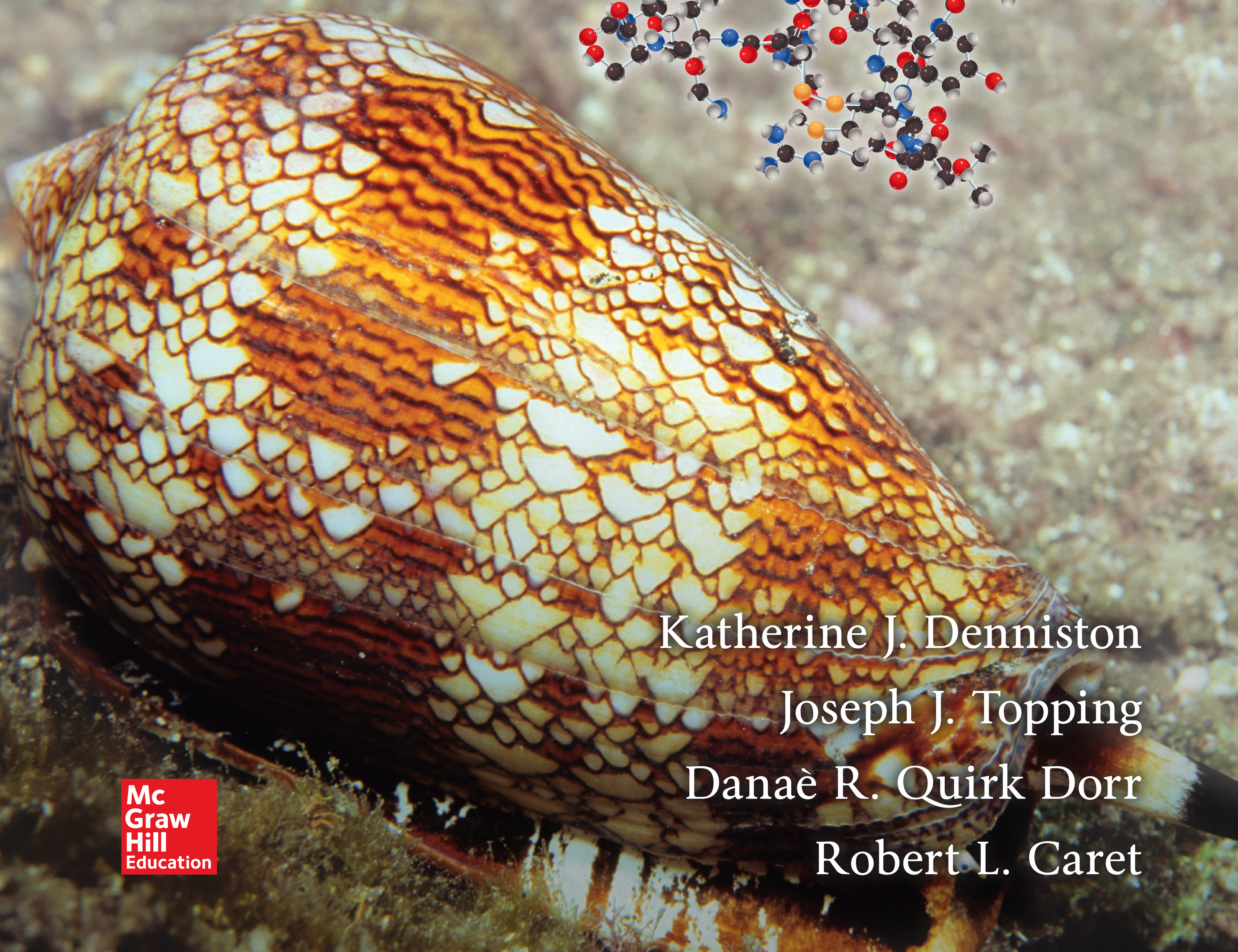
GENERAL, ORGANIC, *and* BIOCHEMISTRY

TENTH EDITION





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Mc Graw Hill Education

General, Organic, *and* Biochemistry

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Brief Contents

GENERAL CHEMISTRY

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BIOCHEMISTRY

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Contents

GENERAL CHEMISTRY

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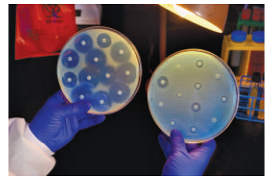

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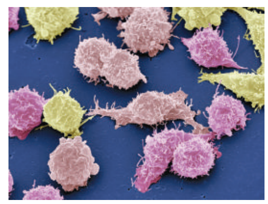
BIOCHEMISTRY



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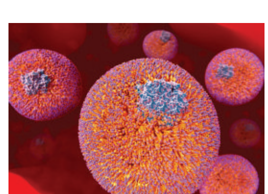

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Perspectives

A Human Perspective

A Medical Perspective

Green Chemistry

Kitchen Chemistry

Chemistry at the Crime Scene

Preface

To Our Students

Student engagement in the study of chemistry has been our primary aim since the first edition of this book. We wanted to show you that chemistry is much more than an onerous obstacle in the journey toward your career goals. Through the Perspectives boxes in each chapter, we have tried to show that chemistry is a fascinating discipline that has an enormous impact on all aspects of your life—whether chemistry in the kitchen, investigations at a crime scene, issues of environmental concern, medicine, or the chemical reactions that keep our bodies functioning.

While engagement in a subject is a good place to begin, effective study practices will ensure your success in learning the course content. In the preface of previous editions, we included suggestions for studying chemistry that included the five stages of the Study Cycle. Because education research has shown that effective use of the Study Cycle improves student performance in all subjects, we wanted to share this information with you. In this edition, we have expanded our attention to research-based learning strategies by including specific sections of the text devoted to effective study skills. In Section 1.1 you will learn about the Study Cycle, as well as some useful strategies that are specific to general chemistry. In Section 10.1, the beginning of the organic chemistry section of the course, you will be challenged to apply study strategies that are specific to that discipline. Similarly, in Section 16.1, the beginning of the biochemistry section, you will be introduced to practices and ideas that will help you master that content.

We have also introduced a new type of problem, multiple concept problems. These challenge you to apply your knowledge of many aspects of the topic to answer thought-provoking questions that will help you develop a much deeper understanding of the principles of chemistry. Research has shown that this type of deeper understanding is crucial to success in all areas of your education. It is our hope that these new elements of the text will provide you with the tools you need to successfully meet the challenges of this course.

To the Instructor

The tenth edition of *General, Organic, and Biochemistry,* like our earlier editions, has been designed to help undergraduate majors in health-related fields understand key concepts and appreciate significant connections among chemistry, health, and the treatment of disease. We have tried to strike a balance between theoretical and practical chemistry, while emphasizing material that is unique to health-related studies. We have written at a level intended for students whose professional goals do not include a mastery of chemistry, but for whom an understanding of the principles and practice of chemistry is a necessity.

Although our emphasis is the importance of chemistry to the health-related professions, we wanted this book to be appropriate for all students who need a one- or two-semester introduction to chemistry. Students learn best when they are engaged. One way to foster that engagement is to help them see clear relationships between the subject and real life. For these reasons, we have included perspectives and essays that focus on medicine and the function of the human body, as well as the environment, forensic science, and even culinary arts.

We begin that engagement with the book cover. Students may wonder why the cover of a chemistry book has a photo of a cone snail. What does an exotic marine snail have to do with the study of chemistry or the practice of medicine? They will learn that the analgesic agent Ziconotide was discovered in the venom of the cone snail in the early 1980s. The drug, sold under the name Prialt, is an unusual painkiller used only in cases of severe, chronic pain. It cannot be taken orally or intravenously, but must be administered directly into the spinal fluid. A short peptide of only twenty-five amino acids, it acts by blocking an N-type voltage-gated calcium channel, thus preventing the release of pain-causing neurochemicals in the brain and spinal fluid.

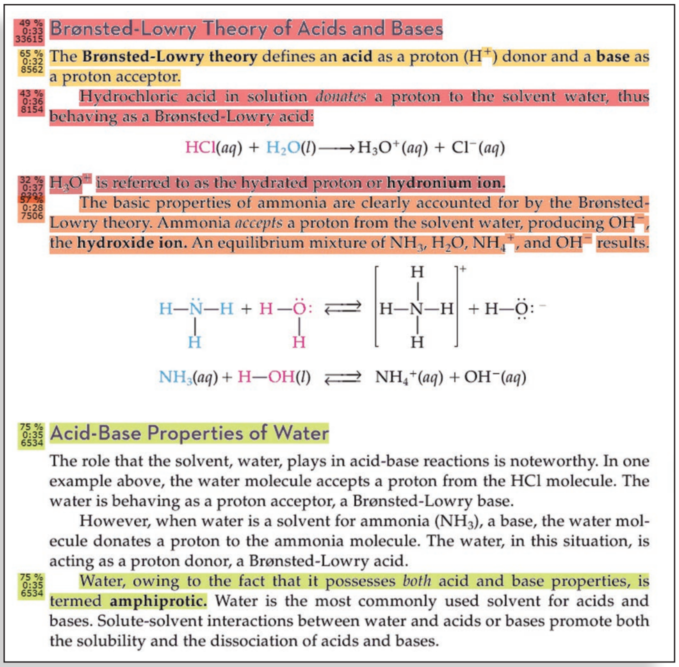
The cover sets the theme for the book: chemistry is not an abstract study, but one that has an immediate impact on our lives. We try to spark student interest with an art program that uses relevant photography, clear and focused figures, and perspectives and essays that bring life to abstract ideas. We reinforce key concepts by explaining them in a clear and concise way and encouraging students to apply the concept to solve problems. We provide guidance through the inclusion of a large number of in-chapter examples that are solved in a stepwise fashion and that provide students the opportunity to test their understanding through the practice problems that follow and the suggested end-of-chapter questions and problems that apply the same concepts.

Foundations for Our Revisions

In the preparation of each edition, we have been guided by the collective wisdom of reviewers who are expert chemists and excellent teachers. They represent experience in community colleges, liberal arts colleges, comprehensive institutions, and research universities. We have followed their recommendations, while remaining true to our overriding goal of writing a readable, student-centered text. This edition has also been designed to be amenable to a variety of teaching styles. Each feature incorporated into this edition has been carefully considered with regard to how it may be used to support student learning in both the traditional classroom and the flipped learning environment.

Also for this edition, we are very pleased to have been able to incorporate real student data points and input, derived from thousands of our LearnSmart users, to help guide our revision. LearnSmart Heat Maps provided a quick visual snapshot of usage of portions of the text and the relative difficulty students experienced in mastering the content. With these data, we were able to hone not only our text content but also the LearnSmart probes.

- If the data indicated that the subject covered was more difficult than other parts of the book, as evidenced by a high proportion of students responding incorrectly, we substantively revised or reorganized the content to be as clear and illustrative as possible.
- In some sections, the data showed that a smaller percentage of the students had difficulty learning the material. In those cases, we revised the *text* to provide a clearer presentation by rewriting the section, providing additional examples to strengthen student problem-solving skills, designing new text art or figures to assist visual learners, etc.



Brønsted-Lowry Theory of Acids and Bases

The **Brønsted-Lowry theory** defines an **acid** as a proton (H^+) donor and a **base** as a proton acceptor.

Hydrochloric acid in solution *donates* a proton to the solvent water, thus behaving as a Brønsted-Lowry acid:

$$HCl(aq) + H_2O(l) \longrightarrow H_3O^+(aq) + Cl^-(aq)$$

H_3O^+ is referred to as the hydrated proton or **hydronium ion.**

The basic properties of ammonia are clearly accounted for by the Brønsted-Lowry theory. Ammonia *accepts* a proton from the solvent water, producing OH^-, the **hydroxide ion.** An equilibrium mixture of NH_3, H_2O, NH_4^+, and OH^- results.

$$NH_3(aq) + H{-}OH(l) \rightleftharpoons NH_4^+(aq) + OH^-(aq)$$

Acid-Base Properties of Water

The role that the solvent, water, plays in acid-base reactions is noteworthy. In one example above, the water molecule accepts a proton from the HCl molecule. The water is behaving as a proton acceptor, a Brønsted-Lowry base.

However, when water is a solvent for ammonia (NH_3), a base, the water molecule donates a proton to the ammonia molecule. The water, in this situation, is acting as a proton donor, a Brønsted-Lowry acid.

Water, owing to the fact that it possesses *both* acid and base properties, is termed **amphiprotic.** Water is the most commonly used solvent for acids and bases. Solute-solvent interactions between water and acids or bases promote both the solubility and the dissociation of acids and bases.

- In other cases, one or more of the LearnSmart probes for a section was not as clear as it might be or did not appropriately reflect the content. In these cases, the *probe,* rather than the text, was edited.

The previous image is an example of one of the heat maps from Chapter 8 that was particularly useful in guiding our revisions. The highlighted sections indicate the various levels of difficulty students experienced in learning the material. This evidence informed all of the revisions described in the "New in This Edition" section of this preface.

The following is a summary of the additions and refinements that we have included in this edition.

New in This Edition

General

Chapter Introductions were rewritten and some chapter opening photos updated in order to better focus on student engagement. The new chapter introduction design leads students directly to the learning goals of the chapter.

"Strategies for Success" sections were added at the beginning of Chapters 1, 10, and 16 to provide students with tools for the most effective study methods to help them master the content and concepts most important to success in general, organic, and biochemistry. In-chapter questions and end-of-chapter problems have also been added to assess students' understanding of the tools and methods presented in the new Strategies sections.

Many updated **photos** emphasizing relevant material and applications have been added within all chapters.

The colors in the **artwork, chemical structures,** and **equations** throughout the text were revised for accessibility, emphasis, clarity, and consistency. Color has also been used in many areas to help students better understand chemical structure, stereochemistry, and reactions. The Chapter Maps were also revised as necessary to better reflect key concepts emphasized in learning goals.

A set of **Multiple Concept Problems** has been added at the end of each chapter, designed to *help students connect various concepts that are emphasized throughout each chapter.* Many other new problems have also been added, both in the text and within the end-of-chapter problem sets, increasing the variety of problems for instructors and students alike.

Several new **Perspective** boxes to help students relate the topics from the text to real-world situations were added throughout: in Chapter 8, Human Perspective: Lithium-Ion Batteries; in Chapter 10, Human Perspective: The Father of Organic Chemistry; in Chapter 12, Kitchen Chemistry: Sugar Alcohols and the Sweet Tooth; in Chapter 13, Green Chemistry: Aldehydes, Stink Bugs, and Wine; in Chapter 15, Green Chemistry: Neoniconoids and Honey Bees; in Chapter 16, Medical Perspective: Chemistry through the Looking Glass; and in Chapter 20, Medical Perspective: CRISPR Technology and the Future of Genetics.

Chapter-Specific

Chapter 4 A new abbreviated Section 4.8, Oxidation-Reduction Reactions, now appears in this chapter, with more detailed coverage revisited in Chapter 8 Acids and Bases and Oxidation-Reduction.

Chapter 8 This chapter includes a new section, Section 8.5, Oxidation-Reduction Processes, with a new figure illustrating the relationship between a voltaic cell and an electrolytic cell and a new Human Perspective box on lithium-ion batteries, explaining why lithium is used in lightweight, rechargeable batteries and why the use of lithium in these batteries also leads to safety issues.

Chapter 12 Additional information on the physical properties of thiols is included.

Chapter 14 Section 14.1, Structure and Physical Properties, was revised to include the general structures of aliphatic and aromatic carboxylic acids, and Section 14.2, Structure and Physical Properties, was revised to include the general structures of aliphatic and aromatic esters.

Chapter 15 The information on semisynthetic penicillins was updated, and information on augmentin was added. The material on opiate biosynthesis was updated, and information on the abuse of suboxone was added to the coverage on the mutant poppy.

Chapter 17 The coverage of LDL receptor-mediated endocytosis in Section 17.5 was revised and updated, and a new table summary of the composition of lipoproteins was added.

Chapter 18 The chapter includes a new Section 18.1, Protein Functions, to help students recognize the importance of the information.

Chapter 20 Material was added to Section 20.1, The Structure of the Nucleotide, and Section 20.10 includes new information on hand-held DNA sequencers.

Chapter 21 Introductory paragraphs were added to Section 21.1 to tie in catabolism and anabolism with life and life processes. Margin notes were added to the sections on the reactions of glycolysis, and to the section on glycogenesis, to revisit the reactions of organic chemistry and to reinforce the new section on How to Succeed in Biochemistry.

Chapter 22 Section 22.1 was revised to include new content on the non-ATP related functions of mitochondria.

Applications

Each chapter contains applications that present short stories about real-world situations involving one or more topics students will encounter within the chapter. There are over 100 applications throughout the text, so students are sure to find many topics that spark their interest. Global climate change, DNA fingerprinting, the benefits of garlic, and gemstones are just a few examples of application topics.

- **Medical Perspectives** relate chemistry to a health concern or a diagnostic application.
- **Green Chemistry** explores environmental topics, including the impact of chemistry on the ecosystem and how these environmental changes affect human health.
- **Human Perspectives** delve into chemistry and society and include such topics as gender issues in science and historical viewpoints.
- **Chemistry at the Crime Scene** focuses on forensic chemistry, applying the principles of chemistry to help solve crimes.
- **Kitchen Chemistry** discusses the chemistry associated with everyday foods and cooking methods.

Learning Tools

In designing the original learning system we asked ourselves: "If we were students, what would help us organize and understand the material covered in this chapter?" Based on the feedback of reviewers and users of our text, we include a variety of learning tools:

- **Strategies for Success in Chemistry** are found at the beginning of each major unit of the course: general, organic, and biochemistry. These new sections provide students with research-based strategies for successful mastery of that content.
- **Chapter Overview** pages begin each chapter, with a chapter outline and an engaging Introduction, leading students directly to the learning goals of the chapter. Both students and professor can see, all in one place, the plan for the chapter.
- **Learning Goal Icons** mark the sections and examples in the chapter that focus on each learning goal.
- **Chapter Cross-References** help students locate pertinent background material. These references to previous chapters, sections, and perspectives are noted in the margins of the text. Marginal cross-references also alert students to upcoming topics related to the information currently being studied.
- **End-of-Chapter Questions and Problems** are arranged according to the headings in the chapter outline, with further subdivision into Foundations (basic concepts) and Applications.
- **Chapter Maps** are included just before the end-of-chapter Summaries to provide students with an overview of the chapter—showing connections among topics, how concepts are related, and outlining the chapter hierarchy.
- **Chapter Summaries** are now a bulleted list format of chapter concepts by major sections, with the integrated bold-faced **Key Terms** appearing in context. This more succinct format helps students to quickly identify and review important chapter concepts and to make connections with the incorporated Key Terms. Each Key Term is defined and listed alphabetically in the **Glossary** at the end of the book.
- **Answers to Practice Problems** are supplied in an appendix at the end of the text so that students can quickly check their understanding of important problem-solving skills and chapter concepts.
- **Summaries of Reactions** in the organic chemistry chapters highlight each major reaction type on a tan background. Major chemical reactions are summarized by equations at the end of the chapter, facilitating review.

Problem Solving and Critical Thinking

Perhaps the best preparation for a successful and productive career is the development of problem-solving and critical thinking skills. To this end, we created a variety of problems that require recall, fundamental calculations, and complex reasoning. In this edition, we have used suggestions from our reviewers, as well as from our own experience, to enhance our 2300 problems. This edition includes new problems and hundreds of example problems with step-by-step solutions.

- **In-Chapter Examples, Solutions, and Practice Problems:** Each chapter includes examples that show the student, step by step, how to properly reach the correct solution to model problems. Each example contains a practice problem, as well as a referral to further practice questions. These questions allow students to test their mastery of information and to build self-confidence. The answers to the practice problems can be found in the Answer Appendix so students can check their understanding.
- **Color-Coding System for In-Chapter Examples:** In this edition, we also introduced a color-coding and label system to help alleviate the confusion that students frequently have when trying to keep track of unit conversions. Introduced in Chapter 1, this color-coding system has been used throughout the problem-solving chapters.

$$3.01\ \cancel{\text{mol S}} \times \frac{32.06\ \text{g S}}{1\ \cancel{\text{mol S}}} = 96.5\ \text{g S}$$

Data Given × Conversion Factor = Desired Result

- **In-Chapter and End-of-Chapter Questions and Problems:** We have created a wide variety of paired concept problems. The answers to the odd-numbered questions are found in the Answer Appendix at the back of the book as reinforcement for students as they develop problem-solving skills. However, students must then be able to apply the same principles to the related even-numbered problems.
- **Multiple Concept Problems:** Each chapter includes a set of these problems intended to engage students to integrate concepts to solve more complex problems. They make a perfect complement to the classroom lecture because they provide an opportunity for in-class discussion of complex problems dealing with daily life and the health care sciences. The answers to the Multiple Concept Problems are available through the Instructor Resources in the Connect Library tab.

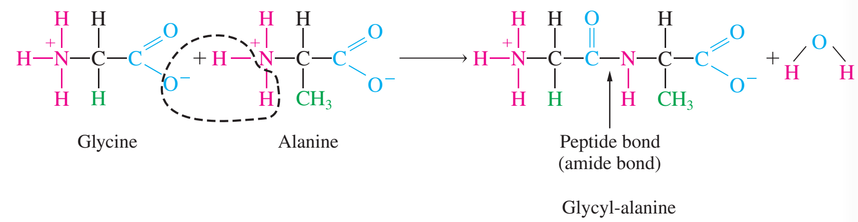
Over the course of the last ten editions, hundreds of reviewers have shared their knowledge and wisdom with us, as well as the reactions of their students to elements of this book. Their contributions, as well as our own continuing experience in the area of teaching and learning science, have resulted in a text that we are confident will provide a strong foundation in chemistry, while enhancing the learning experience of students.

The Art Program

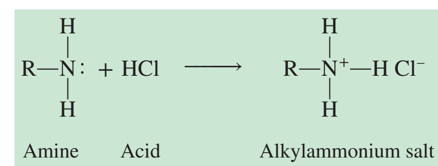
Today's students are much more visually oriented than previous generations. We have built upon this observation through the use of color, figures, and three-dimensional computer-generated models. This art program enhances the readability of the text and provides alternative pathways to learning.

- **Dynamic Illustrations:** Each chapter is amply illustrated using figures, tables, and chemical formulas. All of these illustrations are carefully annotated for clarity. To help students better understand difficult concepts, there are approximately 350 illustrations and 250 photos in the tenth edition.
- **Color-Coding Scheme:** We have color-coded equations so that chemical groups being added or removed in a reaction can be quickly recognized.
  1. **Red print** is used in chemical equations or formulas to draw the reader's eye to key elements or properties in a reaction or structure.
  2. **Blue print** is used when additional features must be highlighted.
  3. **Green background** screens denote generalized chemical and mathematical equations. In the organic chemistry chapters, the Summary of Reactions at the end of the chapter is also highlighted for ease of recognition.
  4. **Yellow backgrounds** illustrate energy, stored either in electrons or groups of atoms, in the general and biochemistry sections of the text. In the organic chemistry section of the text, yellow background screens also reveal the parent chain of an organic compound.
  5. There are situations in which it is necessary to adopt a unique color convention tailored to the material in a particular chapter. For example, in Chapter 18, the structures of amino acids require three colors to draw attention to key features of these molecules. For consistency, blue is used to denote the acid portion of an amino acid and red is used to denote the basic portion of an amino acid. Green print is used to denote the R groups.
- **Computer-Generated Models:** The ability of students to understand the geometry and three-dimensional structure of molecules is essential to the understanding of organic and biochemical reactions. Computer-generated models are used throughout the text because they are both accurate and easily visualized.

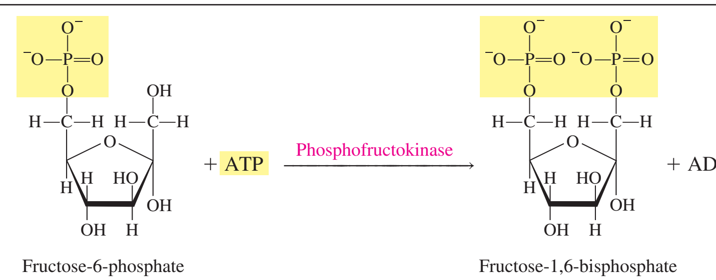
another amino acid. The reaction, shown below for the amino acids glycine and alanine, is a condensation reaction, because a water molecule is lost as the amide bond is formed.

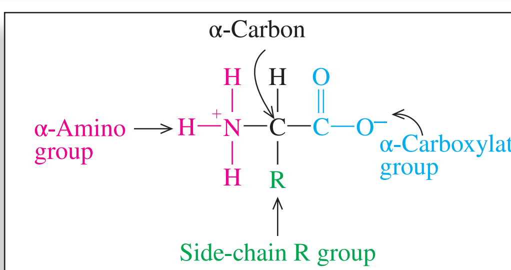


Because amines are bases, they react with acids to form alkylammonium salts.



The reaction of methylamine with hydrochloric acid shown is typical of these re
The product is an alkylammonium salt, methylammonium chloride.





For the Instructor

- **Instructor's Manual:** Written and developed for the tenth edition by the authors, this ancillary contains many useful suggestions for organizing flipped classrooms, lectures, instructional objectives, perspectives on readings from the text, answers to the even-numbered problems and the Multiple Concept problems from the text, a list of each chapter's key concepts, and more. The Instructor's Manual is available through the Instructor Resources in the Connect Library tab.
- **Laboratory Manual for General, Organic, and Biological Chemistry:** Authored by Applegate, Neely, and Sakuta to be the most current lab manual available for the GOB course, incorporating the most modern instrumentation and techniques. Illustrations and chemical structures were developed by the authors to conform to the most recent IUPAC conventions. A problem-solving methodology is also utilized throughout the laboratory exercises. There are two online virtual labs for Nuclear Chemistry and Gas Laws. This Laboratory Manual is also designed with flexibility in mind to meet the differing lengths of GOB courses and the variety of instrumentation available in GOB labs. Helpful instructor materials are also available on this companion website, including answers, solution recipes, best practices with common student issues and TA advice, sample syllabi, and a calculation sheet for the Density lab.
- **Presentation Tools:** Build instructional material wherever, whenever, and however you want with assets such as photos, artwork, and other media that can be used to create customized lectures, visually enhanced tests and quizzes, compelling course websites, or attractive printed support materials. The Presentation Tools can be accessed from the Instructor Resources in the Connect Library tab. Instructors can still access the animations from the OLC for use in their presentations.
- **More than 300 animations available through Connect, the eBook, and SmartBook:** They supplement the textbook material in much the same way as instructor demonstrations. However, they are only a few mouse-clicks away, any time, day or night. Because many students are visual learners, the animations add another dimension of learning; they bring a greater degree of reality to the written word.

For the Student

- **Student Study Guide/Solutions Manual:** A separate Student Study Guide/Solutions Manual, prepared by Danaè Quirk Dorr, is available. It contains the answers and complete solutions for the odd-numbered problems. It also offers students a variety of exercises and keys for testing their comprehension of basic, as well as difficult, concepts.
- **Schaum's Outline of General, Organic, and Biological Chemistry:** Written by George Odian and Ira Blei, this supplement provides students with more than 1400 solved problems with complete solutions. It also teaches effective problem-solving techniques.

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>
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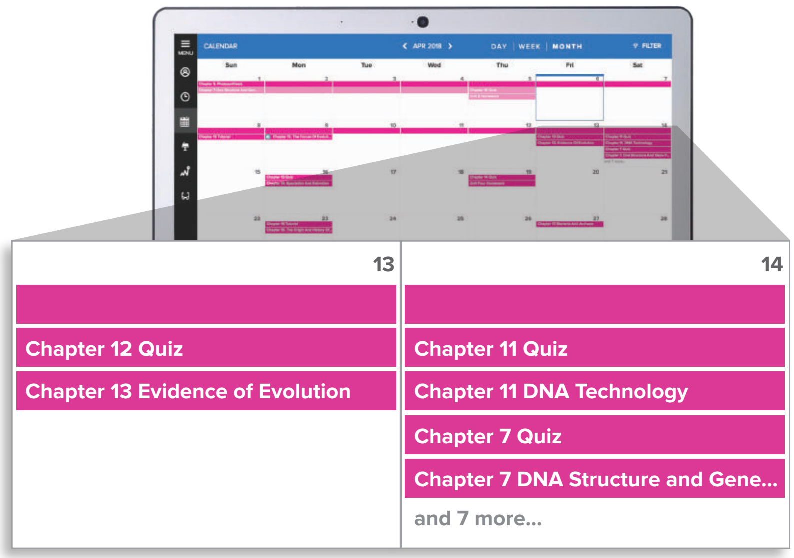

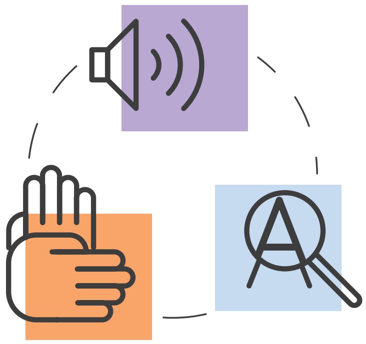

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1

METHODS AND MEASUREMENT

Chemistry

OUTLINE

INTRODUCTION

Louis Pasteur, a chemist and microbiologist, said, "Chance favors the prepared mind." In the history of science and medicine, there are many examples in which individuals made important discoveries because they recognized the value of an unexpected observation.

One such example is the use of ultraviolet (UV) light to treat infant jaundice. Infant jaundice is a condition in which the skin and the whites of the eyes appear yellow because of high levels of the bile pigment bilirubin in the blood. Bilirubin is a breakdown product of the oxygen-carrying blood protein hemoglobin. If bilirubin accumulates in the body, it can cause brain damage and death. The immature liver of the baby cannot remove the bilirubin.

In 1956, an observant nurse in England noticed that when jaundiced babies were exposed to sunlight, the jaundice faded. Research based on her observation showed that the UV light changes the bilirubin into another substance, which can be excreted. To this day, jaundiced newborns undergoing phototherapy are treated with UV light. Historically, newborns were diagnosed with jaundice based only on their physical appearance. However, it has been determined that this method is not always accurate. Now it is common to use either an instrument or a blood sample to measure the amount of bilirubin present in the serum.

In this first chapter of your study of chemistry, you will learn about the scientific method: the process of developing hypotheses to explain observations and the design of experiments to test those hypotheses.

You will also see that measurement of properties of matter, and careful observation and recording of data, are essential to scientific inquiry. So too is assessment of the precision and accuracy of measurements. Measurements (data) must be reported to allow others to determine their significance. Therefore, an understanding of significant figures, and the ability to represent data in the most meaningful units, enables other scientists to interpret data and results.

Continued

The following Learning Goals of this chapter will help you develop the skills needed to represent and communicate data and results from scientific inquiry.

1 Outline a strategy for learning general chemistry.
2 Explain the relationship between chemistry, matter, and energy.
3 Discuss the approach to science, the scientific method, and distinguish among the terms *hypothesis, theory,* and *scientific law.*
4 Distinguish between *data* and *results*.
5 Describe the properties of the solid, liquid, and gaseous states.
6 Classify matter according to its composition.
7 Provide specific examples of physical and chemical properties and physical and chemical changes.
8 Distinguish between intensive and extensive properties.
9 Identify the major units of measure in the English and metric systems.
10 Report data and calculate results using scientific notation and the proper number of significant figures.
11 Distinguish between *accuracy* and *precision* and their representations: *error* and *deviation.*
12 Convert between units of the English and metric systems.
13 Know the three common temperature scales, and convert values from one scale to another.
14 Use density, mass, and volume in problem solving, and calculate the specific gravity of a substance from its density.

1.1 Strategies for Success in Chemistry

The Science of Learning Chemistry

A growing body of scientists, including neurobiologists, chemists, and educational psychologists, study the process of learning. Their research has shown that there are measurable changes in the brain as learning occurs. While the research on brain chemistry and learning continues, the results to date have taught us some very successful strategies for learning chemistry. One of the important things we have learned is that, in the same way that repetition in physical exercise builds muscle, long-term retention of facts and concepts also requires repetition. As in physical exercise, a proven plan of action is invaluable for learning. Repetition is a central component of the **Study Cycle,** Figure 1.1, a plan for learning. Following this approach can lead to success, not only in chemistry, but in any learning endeavor.

Learning General Chemistry

LEARNING GOAL

1 Outline a strategy for learning general chemistry.

The first nine chapters of this book focus on the basic principles of **general chemistry.** General chemistry incorporates concepts that connect most aspects of chemistry. The thought of mastering this information can appear to be a daunting task. As the authors, we have combined our experiences (first as students, then as instructors), along with input from dozens of fellow chemistry professors, to design a book with content and features that will support you as you learn chemistry.

We suggest several strategies that you can use to help convert the concepts in Chapters 1–9 into an organized framework that facilitates your understanding of general chemistry:

1. Several researchers have demonstrated the importance of previewing materials prior to each class. As you look through the chapter, identify the concepts that are unclear to you. It is critical to address these unclear ideas because if you don't, they will become barriers to your understanding throughout the course, not just in the chapter you are currently studying. Ask for clarification. Your instructor

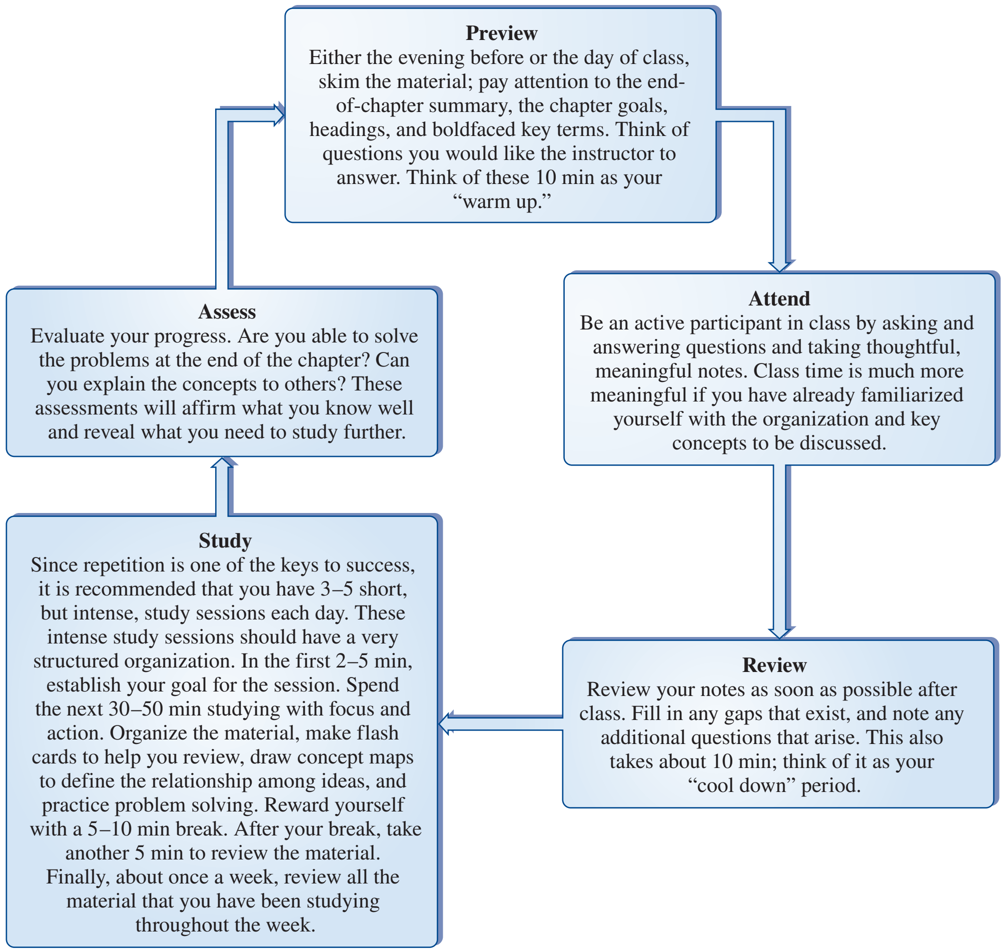
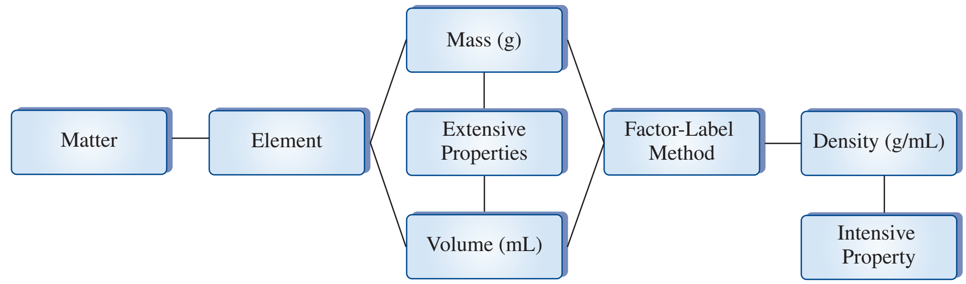


Figure 1.1 Research has shown that it can be effective for students to incorporate these five phases of the Study Cycle into their study plan.

should be a primary contact; additionally, the department or college may have a student resource center with tutors to help you.

2. Class time is another opportunity to improve your understanding. Students who actively participate in class, asking questions and participating in the discussion, gain a better understanding of the materials and achieve better grades.
3. Your class notes are another important study tool. As you review them after class, take note of questions you have and use the text to try to answer those questions.
4. You will find it very useful to design flash cards for use as a study tool for key equations, definitions, or relationships.
5. Identify big ideas. The learning goals at the beginning of each chapter are an excellent place to start. Additionally, the boldfaced terms throughout each chapter highlight the most important concepts.
6. Organize the material in a way that lends itself to processing not only individual concepts but the interrelationships that exist among these concepts. As you organize the big ideas, look for these connections. Use the chapter maps and summaries at the end of each chapter to help you visualize the organization of topics within and among the various chapters.
7. Concept maps are excellent tools to help you define and understand the relationships among ideas. For example, Chapter 1 introduces classification of matter and properties of matter. The use of "chemical arithmetic" is also presented to make

useful chemical and physical calculations. To understand these connections, you might begin with a diagram such as:



Then, next to each line you can write the relationship between these concepts. You can also continue to build upon your concept map as you continue to learn new material. The concepts and calculations introduced in Chapter 1 are used and expanded upon in subsequent chapters, enabling a fuller understanding of more complex chemical behavior.

8. Use the in-chapter and end-of-chapter questions and problems as your own personal quiz. Attempt to answer the questions and problems dealing with a certain topic; then check the answers in the textbook. Use the textbook explanations and Solutions Manual to help you determine where you may have gone wrong. Remember that numerous example problems in the chapter model solutions to the most frequently encountered situations.

Remember, these are suggestions. You may find that some work well for you and others, perhaps, not as well. The goal is active learning; you are ultimately responsible for learning the material. Preparation builds confidence; confidence is a key component of success in exams and, importantly, success in the course.

Question 1.1 Each student is a unique individual; not all students learn in the same way. Based on what you have read above, coupled with your own experience, design a learning strategy for Chapter 1 that you believe will work for you.

Question 1.2 After you have completed your reading of Chapter 1, prepare a set of flash cards that will assist you in learning important terms, definitions, and equations contained in the chapter.

1.2 The Discovery Process

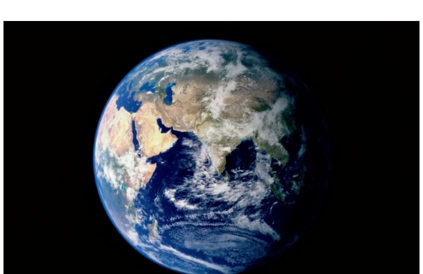
LEARNING GOAL

2 Explain the relationship between chemistry, matter, and energy.

Chemistry

Chemistry is the study of matter, its chemical and physical properties, the chemical and physical changes it undergoes, and the energy changes that accompany those processes.

Matter is anything that has mass and occupies space. The air we breathe, our bodies, our planet earth, our universe; all are made up of an immense variety and quantity of particles, collectively termed matter. Matter undergoes change. Sometimes this change occurs naturally or we change matter when we make new substances (creating drugs in a pharmaceutical laboratory). All of these changes involve **energy,** the ability to do work to accomplish some change. Hence, we may describe chemistry as a study of matter and energy and their interrelationship.



Chemistry is the study of anything that has mass and occupies space.
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Chemistry is an experimental science. A traditional image of a chemist is someone wearing a white coat and safety goggles while working in solitude in a laboratory. Although much chemistry is still accomplished in a traditional laboratory setting, over the last 40 years the boundaries of the laboratory have expanded to include the power of modern technology. For example, searching the scientific literature for information no longer involves a trip to the library as it is now done very quickly via the Internet. Computers are also invaluable in the laboratory because they control sophisticated instrumentation that measures, collects, processes, and interprets information. The behavior of matter can also be modeled using sophisticated computer programs.

Additionally, chemistry is a collaborative process. The solitary scientist, working in isolation, is a relic of the past. Complex problems dealing with topics such as the environment, disease, forensics, and DNA require input from other scientists and mathematicians who can bring a wide variety of expertise to problems that are chemical in nature.

The boundaries between the traditional sciences of chemistry, physics, and biology, as well as mathematics and computer science, have gradually faded. Medical practitioners, physicians, nurses, and medical technologists use therapies that contain elements of all these disciplines. The rapid expansion of the pharmaceutical industry is based on recognition of the relationship between the function of an organism and its basic chemical makeup. Function is a consequence of changes that chemical substances undergo.

For these reasons, an understanding of basic chemical principles is essential for anyone considering a medically related career; indeed, a worker in any science-related field will benefit from an understanding of the principles and applications of chemistry.

Investigating the causes of the rapid melting of glaciers is a global application of chemistry. How does this illustrate the interaction of matter and energy?
©Vadim Balakin/Getty Images

The Scientific Method

LEARNING GOAL

3 Discuss the approach to science, the scientific method, and distinguish among the terms *hypothesis, theory,* and *scientific law.*

The **scientific method** is a systematic approach to the discovery of new information. How do we learn about the properties of matter, the way it behaves in nature, and how it can be modified to make useful products? Chemists do this by using the scientific method to study the way in which matter changes under carefully controlled conditions.

The scientific method is not a "cookbook recipe" that, if followed faithfully, will yield new discoveries; rather, it is an organized approach to solving scientific problems. Every scientist brings his or her own curiosity, creativity, and imagination to scientific study. Yet, scientific inquiry does involve some of the "cookbook recipe" approach.

Characteristics of the scientific process include the following:

- *Observation.* The description of, for example, the color, taste, or odor of a substance is a result of observation. The measurement of the temperature of a liquid or the size or mass of a solid results from observation.
- *Formulation of a question.* Humankind's fundamental curiosity motivates questions of why and how things work.
- *Pattern recognition.* When a cause-and-effect relationship is found, it may be the basis of a generalized explanation of substances and their behavior.
- *Theory development.* When scientists observe a phenomenon, they want to explain it. The process of explaining observed behavior begins with a hypothesis. A **hypothesis** is simply an attempt to explain an observation, or series of observations. If many experiments support a hypothesis, it may attain the status of a theory. A **theory** is a hypothesis supported by extensive testing (experimentation) that explains scientific observations and data and can accurately predict new observations and data.
- *Experimentation.* Demonstrating the correctness of hypotheses and theories is at the heart of the scientific method. This is done by carrying out carefully designed experiments that will either support or disprove the hypothesis or theory. A scientific experiment produces **data.** Each piece of data is the individual result of a single measurement or observation.

  A **result** is the outcome of an experiment. Data and results may be identical, but more often, several related pieces of data are combined, and logic is used to produce a result.
- *Information summarization.* A **scientific law** is nothing more than the summary of a large quantity of information. For example, the law of conservation of matter states that matter cannot be created or destroyed, only converted from one form to another. This statement represents a massive body of chemical information gathered from experiments.

LEARNING GOAL

4 Distinguish between *data* and *results.*

LEARNING GOAL

4 Distinguish between *data* and *results*.

EXAMPLE 1.1 Distinguishing Between Data and Results

In many cases, a drug is less stable in the presence of moisture, and excess moisture can hasten the breakdown of the active ingredient, leading to loss of potency. Bupropion (Wellbutrin) is an antidepressant that is moisture sensitive. Describe an experiment that will allow for the determination of the quantity of water gained by a certain quantity of bupropion when it is exposed to air.

Solution

To do this experiment, we must first weigh the buproprion sample, and then expose it to the air for a period of time and reweigh it. The change in weight,

$$[\text{weight}_{\text{final}} - \text{weight}_{\text{initial}}] = \text{weight difference}$$

indicates the weight of water taken up by the drug formulation. The initial and final weights are individual bits of *data;* by themselves they do not answer the question, but they do provide the information necessary to calculate the answer: the results. The difference in weight and the conclusions based on the observed change in weight are the *results* of the experiment.

Note: This is actually not a very good experiment because many conditions were not measured. Measurement of the temperature, humidity of the atmosphere, and the length of time that the drug was exposed to the air would make the results less ambiguous.

Practice Problem 1.1

Describe an experiment that demonstrates that the boiling point of water changes when salt (sodium chloride) is added to the water.

▶ For Further Practice: **Questions 1.41 and 1.42.**

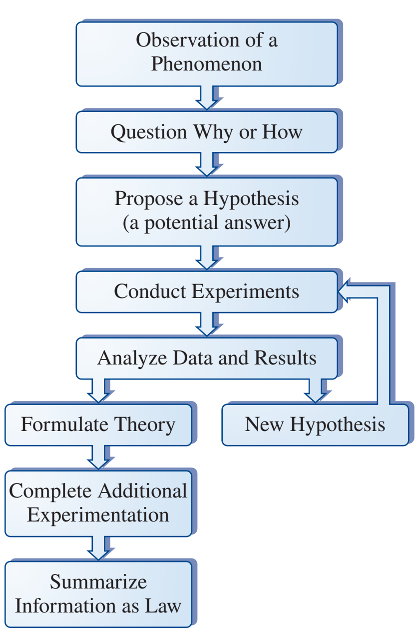


Figure 1.2 The scientific method is an organized way of doing science that incorporates a degree of trial and error. If the data analysis and results do not support the initial hypothesis, the cycle must begin again.

The scientific method involves the interactive use of hypotheses, development of theories, and thorough testing of theories using well-designed experiments. It is summarized in Figure 1.2.

Models in Chemistry

Hypotheses, theories, and laws are frequently expressed using mathematical equations. These equations may confuse all but the best of mathematicians. For this reason, a *model* of a chemical unit or system is often used to help illustrate an idea. A good model based on everyday experience, although imperfect, gives a great deal of information in a simple fashion.

Consider the fundamental unit of methane, the major component of natural gas, which is composed of one carbon atom (symbolized by C) and four hydrogen atoms (symbolized by H).

A geometrically correct model of methane can be constructed from balls and sticks. The balls represent the individual atoms of hydrogen and carbon, and the sticks correspond to the attractive forces that hold the hydrogen and carbon together. The model consists of four balls representing hydrogen symmetrically arranged around a center ball representing carbon.

A Human Perspective

The Scientific Method

The discovery of penicillin by Alexander Fleming is an example of the scientific method at work. Fleming was studying the growth of bacteria. One day, his experiment was ruined because colonies of mold were growing on his plates. From this failed experiment, Fleming made an observation that would change the practice of medicine: Bacterial colonies could not grow in the area around the mold colonies. Fleming hypothesized that the mold was making a chemical compound that inhibited the growth of the bacteria. He performed a series of experiments designed to test this hypothesis.

The success of the scientific method is critically dependent upon carefully designed experiments that will either support or disprove the hypothesis. This is what Fleming did.

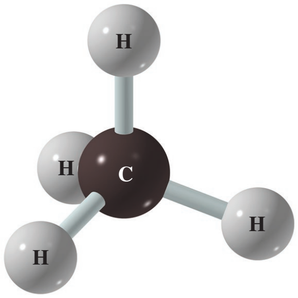
In one experiment, he used two sets of tubes containing sterile nutrient broth. To one set he added mold cells. The second set (the control tubes) remained sterile. The mold was allowed to grow for several days. Then the broth from each of the tubes (experimental and control) was passed through a filter to remove any mold cells. Next, bacteria were placed in each tube. If Fleming's hypothesis was correct, the tubes in which the mold had grown would contain the chemical that inhibits growth, and the bacteria would not grow. On the other hand, the control tubes (which were never used to grow mold) would allow bacterial growth. This is exactly what Fleming observed.

Within a few years this *antibiotic,* penicillin, was being used to treat bacterial infections in patients.

Phenoxymethylpenicillin is a form of penicillin that can be taken orally.
©Julian Claxton/Alamy Stock Photo

For Further Understanding

- What is the purpose of the control tubes used in this experiment?
- Match the features of this article with the flowchart items in Figure 1.2.



Color-coding the balls distinguishes one type of atom from another; the geometrical form of the model, all of the angles and dimensions of a tetrahedron, are the same for each methane unit found in nature. Methane is certainly not a collection of balls and sticks, but such models are valuable because they help us understand the chemical behavior of methane and other more complex substances.

The structure-properties concept has advanced so far that compounds are designed and synthesized in the laboratory with the hope that they will perform very specific functions, such as curing diseases that have been resistant to other forms of treatment. Figure 1.3 shows some of the variety of modern technology that has its roots in scientific inquiry.

Chemists and physicists have used the observed properties of matter to develop models of the individual units of matter. These models collectively make up what we now know as the atomic theory of matter, which is discussed in detail in Chapter 2.

Figure 1.3 Examples of technology originating from scientific inquiry: (a) synthesizing a new drug, (b) playing a game with virtual reality goggles, (c) using UV light to set adhesive, and (d) wireless printing from a smart phone.
(a) ©Adam Gault/AGE Fotostock; (b) ©innovatedcaptures/123RF; (c) ©Science Photo Library/Alamy Stock Photo; (d) ©Piotr Adamowicz/Shutterstock

1.3 The Classification of Matter

Matter is a large and seemingly unmanageable concept because it includes everything that has mass and occupies space. Chemistry becomes manageable as we classify matter according to its **properties**—that is, the characteristics of the matter. Matter will be classified in two ways in this section, by *state* and by *composition.*

States of Matter

We will examine each of the three states of matter in detail in Chapter 5.

LEARNING GOAL

5 Describe the properties of the solid, liquid, and gaseous states.

There are three *states of matter:* the **gaseous state,** the **liquid state,** and the **solid state.** A gas is made up of particles that are widely separated. In fact, a gas will expand to fill any container; it has no definite shape or volume. In contrast, particles of a liquid are closer together; a liquid has a definite volume but no definite shape; it takes on the shape of its container. A solid consists of particles that are close together and often have a regular and predictable pattern of particle arrangement (crystalline). The particles in a solid are much more organized than the particles in a liquid or a gas. As a result, a solid has both fixed volume and fixed shape. Attractive forces, which exist between all particles, are very pronounced in solids and much less so in gases.

Composition of Matter

LEARNING GOAL

6 Classify matter according to its composition.

We have seen that matter can be classified by its state as a solid, liquid, or gas. Another way to classify matter is by its composition. This very useful system, described in the following paragraphs and summarized in Figure 1.4, will be utilized throughout the textbook.

All matter is either a *pure substance* or a *mixture.* A **pure substance** has only one component. Pure water is a pure substance. It is made up only of particles containing two hydrogen (symbolized by H) atoms and one oxygen (symbolized by O) atom—that is, water molecules (H_2O).

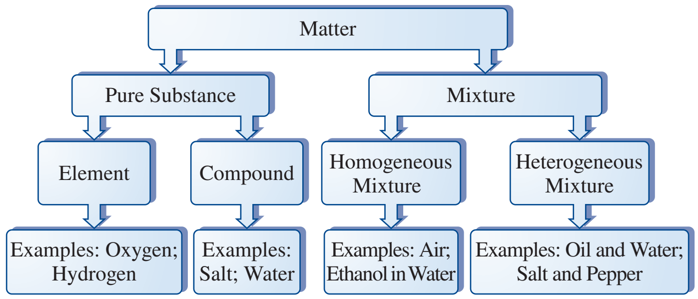


Figure 1.4 Classification of matter by composition. All matter is either a pure substance or a mixture of pure substances. Pure substances are either elements or compounds, and mixtures may be either homogeneous (uniform composition) or heterogeneous (nonuniform composition).

There are different types of pure substances. Elements and compounds are both pure substances. An **element** is a pure substance that generally cannot be changed into a simpler form of matter. Hydrogen and oxygen, for example, are elements. Alternatively, a **compound** is a substance resulting from the combination of two or more elements in a definite, reproducible way. The elements hydrogen and oxygen, as noted earlier, may combine to form the compound water, H_2O.

At present, more than 100 elements have been characterized. A complete listing of the elements and their symbols is found in Chapter 2.

A **mixture** is a combination of two or more pure substances in which each substance retains its own identity. Ethanol, the alcohol found in beer, and water can be combined in a mixture. They coexist as pure substances because they do not undergo a chemical reaction. A mixture has variable composition; there are an infinite number of combinations of quantities of ethanol and water that can be mixed. For example, the mixture may contain a small amount of ethanol and a large amount of water or vice versa. Each is, however, an ethanol-water mixture.

A mixture may be either *homogeneous* or *heterogeneous* (Figure 1.5). A **homogeneous mixture** has uniform composition. Its particles are well mixed, or thoroughly intermingled. A homogeneous mixture, such as alcohol and water, is described as a *solution*. Air, a mixture of gases, is an example of a gaseous solution. A **heterogeneous mixture** has a nonuniform composition. A mixture of salt and pepper is a good example of a heterogeneous mixture. Concrete is also composed of a heterogeneous mixture of materials (a nonuniform mixture of various types and sizes of stone and sand combined with cement).

A detailed discussion of solutions (homogeneous mixtures) and their properties is presented in Chapter 6.

(a) (b) (c)

Figure 1.5 Schematic representations of some classes of matter. (a) A pure substance, water, consists of a single component. (b) A homogeneous mixture, blue dye in water, has a uniform distribution of components. The blue spheres represent the blue dye molecules. (c) The mineral orbicular jasper is an example of a heterogeneous mixture. The lack of homogeneity is apparent from its nonuniform distribution of components. *(a) ©Image Source Plus/Alamy Stock Photo; (b) ©Image Source/Getty Images; (c) ©Danaè R. Quirk Dorr, Ph.D.*

EXAMPLE 1.2 **Classifying Matter by Composition**

Is seawater a pure substance, a homogeneous mixture, or a heterogeneous mixture?

Solution

Imagine yourself at the beach, filling a container with a sample of water from the ocean. Examine it. You would see a variety of solid particles suspended in the water: sand, green vegetation, perhaps even a small fish! Clearly, it is a mixture, and one in which the particles are not uniformly distributed throughout the water; hence, it is a heterogeneous mixture.

Practice Problem 1.2

Is each of the following materials a pure substance, a homogeneous mixture, or a heterogeneous mixture?

a. ethanol
b. blood
c. an Alka-Seltzer tablet fizzing in water
d. oxygen being delivered from a hospital oxygen tank

LEARNING GOAL 6 Classify matter according to its composition.

▶ For Further Practice: **Questions 1.53 and 1.54.**

Question 1.3 Intravenous therapy may be used to introduce a saline solution into a patient's vein. Is this solution a pure substance, a homogeneous mixture, or a heterogeneous mixture?

Question 1.4 Cloudy urine can be a symptom of a bladder infection. Classify this urine as a pure substance, a homogeneous mixture, or a heterogeneous mixture.

LEARNING GOAL

7 Provide specific examples of physical and chemical properties and physical and chemical changes.

Physical Properties and Physical Change

Water is the most common example of a substance that can exist in all three states over a reasonable temperature range (Figure 1.6). Conversion of water from one state to another constitutes a *physical change*. A **physical change** produces a recognizable difference in the appearance of a substance without causing any change in its composition or identity. For example, we can warm an ice cube and it will melt, forming liquid water. Clearly its appearance has changed; it has been transformed from the solid to the liquid state. It is, however, still water; its composition and identity remain unchanged. A physical change has occurred. We could in fact demonstrate the constancy of composition and identity by refreezing the liquid water, re-forming the ice cube. This melting and

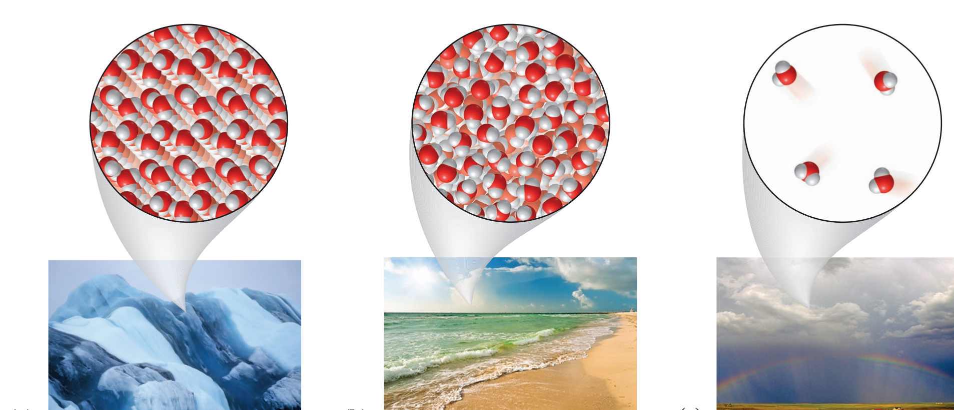


Figure 1.6 The three states of matter exhibited by water: (a) solid, as ice; (b) liquid, as ocean water; (c) gas, as humidity in the air.
(a) ©moodboard/Glow Images; (b) ©S.Borisov/Shutterstock; (c) ©WeatherVideoHD.TV

freezing cycle could be repeated over and over. This very process is a hallmark of our global weather changes. The continual interconversion of the three states of water in the environment (snow, rain, and humidity) clearly demonstrates the retention of the identity of water particles or *molecules.*

A **physical property** can be observed or measured without changing the composition or identity of a substance. As we have seen, melting ice is a physical change. We can measure the temperature when melting occurs; this is the *melting point* of water. We can also measure the *boiling point* of water, when liquid water becomes a gas. Both the melting and boiling points of water, and of any other substance, are physical properties.

A practical application of separation of materials based upon their differences in physical properties is shown in Figure 1.7.

Figure 1.7 An example of separation based on differences in physical properties. Magnetic iron is separated from nonmagnetic substances. A large-scale version of this process is important in the recycling industry.
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Chemical Properties and Chemical Change

We have noted that physical properties can be exhibited, measured, or observed without any change in identity or composition. In contrast, **chemical properties** are a consequence of change in composition and can be observed only through chemical reactions. In a **chemical reaction,** a chemical substance is converted to one or more different substances by rearranging, removing, replacing, or adding atoms. For example, the process of photosynthesis can be shown as

$$\text{Carbon dioxide} + \text{Water} \xrightarrow[\text{Chlorophyll}]{\text{Light}} \text{Sugar} + \text{Oxygen}$$

Light is the energy needed to make the reaction happen. Chlorophyll is the energy absorber, converting light energy to chemical energy.

This chemical reaction involves the conversion of carbon dioxide and water (the *reactants*) to a sugar and oxygen (the *products*). The physical properties of the reactants and products are clearly different. We know that carbon dioxide and oxygen are gases at room temperature, and water is a liquid at this temperature; the sugar is a solid white powder. A chemical property of carbon dioxide is its ability to form sugar under certain conditions. The process of formation of this sugar is the *chemical change.* The term *chemical reaction* is synonymous with **chemical change.**

EXAMPLE 1.3 Classifying Change

Can the process that takes place when an egg is fried be described as a physical or chemical change?

Solution

Examine the characteristics of the egg before and after frying. Clearly, some significant change has occurred. Furthermore, the change appears irreversible. More than a simple physical change has taken place. A chemical reaction (actually, several) must be responsible; hence, there is a chemical change.

Practice Problem 1.3

Classify each of the following as either a chemical change or a physical change:

a. water boiling to become steam
b. butter becoming rancid
c. burning wood
d. melting of ice in spring
e. decaying of leaves in winter

LEARNING GOAL 7 Provide specific examples of physical and chemical properties and physical and chemical changes.

▶ For Further Practice: **Questions 1.57 and 1.58.**

Question 1.5 Classify each of the following as either a chemical property or a physical property:

a. color b. flammability c. hardness

Question 1.6 Classify each of the following as either a chemical property or a physical property:

a. odor b. taste c. temperature

LEARNING GOAL

8 Distinguish between intensive and extensive properties.

The mass of a pediatric patient (in kg) is an extensive property that is commonly used to determine the proper dosage of medication [in milligrams (mg)] prescribed. Although the mass of the medication is also an extensive property, the dosage (in mg/kg) is an intensive property. This calculated dosage should be the same for every pediatric patient.

Intensive and Extensive Properties

It is important to recognize that properties can also be classified according to whether they depend on the size of the sample. Consequently, there is a fundamental difference between properties such as color and melting point and properties such as mass and volume.

An **intensive property** is a property of matter that is *independent* of the *quantity* of the substance. Boiling and melting points are intensive properties. For example, the boiling point of one single drop of water is exactly the same as the boiling point of a liter (L) of water.

An **extensive property** *depends* on the *quantity* of a substance. Mass and volume are extensive properties. There is an obvious difference between 1 gram (g) of silver and 1 kilogram (kg) of silver; the quantities and, incidentally, the monetary values, differ substantially.

EXAMPLE 1.4 Differentiating Between Intensive and Extensive Properties

Is temperature an intensive or extensive property?

Solution

Imagine two glasses, each containing 100 g of water, and each at 25°C. Now pour the contents of the two glasses into a larger glass. You would predict that the mass of the water in the larger glass would be 200 g (100 g + 100 g) because mass is an *extensive property,* dependent on quantity. However, we would expect the temperature of the water to remain the same (not 25°C + 25°C); hence, temperature is an *intensive property* . . . independent of quantity.

Practice Problem 1.4

Pure water freezes at 0°C. Is this an intensive or extensive property? Why?

LEARNING GOAL 8 Distinguish between intensive and extensive properties.

▶ For Further Practice: **Questions 1.65 and 1.66.**

Question 1.7 Label each property as intensive or extensive:

a. the length of my pencil b. the color of my pencil

Question 1.8 Label each property as intensive or extensive:

a. the shape of leaves on a tree b. the number of leaves on a tree

LEARNING GOAL

9 Identify the major units of measure in the English and metric systems.

1.4 The Units of Measurement

The study of chemistry requires the collection of data through measurement. The quantities that are most often measured include mass, length, and volume. Measurements require the determination of an amount followed by a **unit,** which defines the basic quantity being measured. A weight of 3 *ounces* (oz) is clearly quite different than 3 *pounds* (lb). A number that is not followed by the correct unit usually conveys no useful information.

The *English system of measurement* is a collection of unrelated units used in the United States in business and industry. However, it is not used in scientific work, primarily because it is difficult to convert one unit to another. In fact, the English "system" is not really a system at all; it is simply a collection of units accumulated throughout English history. Table 1.1 shows relationships among common English units of weight, length, and volume.

The United States has begun efforts to convert to the metric system. The *metric system* is truly systematic. It is composed of a set of units that are related to each other decimally; in other words, as powers of ten. Because the metric system is a decimally based system, it is inherently simpler to use and less ambiguous. Table 1.2 shows the meaning of the prefixes used in the metric system.

The metric system was originally developed in France just before the French Revolution in 1789. The more extensive version of this system is the *Systéme International,* or *S.I. system.* Although the S.I. system has been in existence for over 50 years, it has yet to gain widespread acceptance. Because the S.I. system is truly systematic, it utilizes certain units, especially for pressure, that many find unwieldy.

In this text, we will use the metric system, not the S.I. system, and we will use the English system only to the extent of converting from it to the more systematic metric system.

Now let's look at the major metric units for mass, length, volume, and time in more detail. In each case, we will compare the unit to a familiar English unit.

The photo shows 3 oz of grapes versus a 3-lb cantaloupe. Clearly units are important.

Mass

Mass describes the quantity of matter in an object. The terms *weight* and *mass,* in common usage, are often considered synonymous. They are not, in fact. **Weight** is the force of gravity on an object:

$$\text{Weight} = \text{mass} \times \text{acceleration due to gravity}$$

LEARNING GOAL

9 Identify the major units of measure in the English and metric systems.

The mathematical process of converting between units will be covered in detail in Section 1.6.

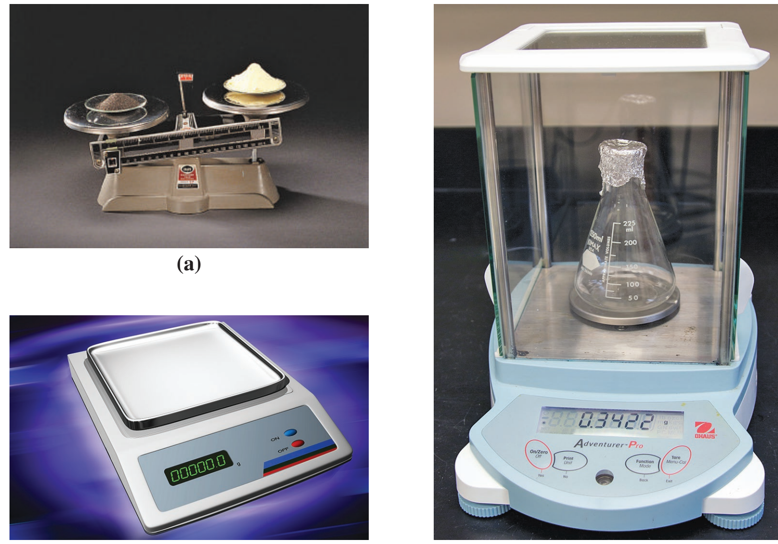
TABLE 1.1 Some Common Relationships Used in the English System

Weight	1 pound (lb) = 16 ounces (oz)
	1 ton (t) = 2000 pounds (lb)
Length	1 foot (ft) = 12 inches (in)
	1 yard (yd) = 3 feet (ft)
	1 mile (mi) = 5280 feet (ft)
Volume	1 quart (qt) = 32 fluid ounces (fl oz)
	1 quart (qt) = 2 pints (pt)
	1 gallon (gal) = 4 quarts (qt)

The table of common prefixes used in the metric system relates values to the base units. For example, it defines 1 mg as being equivalent to 10^{-3} g and 1 kg as being equivalent to 10^3 g.

TABLE 1.2 Some Common Prefixes Used in the Metric System

Prefix	Abbreviation	Meaning	Decimal Equivalent	Equality with major metric units (g, m, or L are represented by x in each)
mega	M	10^6	1,000,000.	1 Mx = 10^6x
kilo	k	10^3	1000.	1 kx = 10^3x
deka	da	10^1	10.	1 da$x$ = 10^1x
deci	d	10^{-1}	0.1	1 dx = $10^{-1}x$
centi	c	10^{-2}	0.01	1 cx = $10^{-2}x$
milli	m	10^{-3}	0.001	1 mx = $10^{-3}x$
micro	μ	10^{-6}	0.000001	1 μx = $10^{-6}x$
nano	n	10^{-9}	0.000000001	1 nx = $10^{-9}x$



(a) (b) (c)

Figure 1.8 Three common balances that are useful for the measurement of mass. (a) A two-pan comparison balance for approximate mass measurement suitable for routine work requiring accuracy to 0.1 g (or perhaps 0.01 g). (b) A top-loading single-pan electronic balance that is similar in accuracy to (a) but has the advantages of speed and ease of operation. The revolution in electronics over the past 20 years has resulted in electronic balances largely supplanting the two-pan comparison balance in routine laboratory usage. (c) An analytical balance of this type is used when the highest level of precision and accuracy is required.
(a) ©McGraw-Hill Education/Stephen Frisch; (b) ©KrishnaKumar Sivaraman/123RF; (c) ©McGraw-Hill Education/Lisa Burgess, photographer

When gravity is constant, mass and weight are directly proportional. But gravity is not constant; it varies as a function of the distance from the center of the earth. Therefore, weight cannot be used for scientific measurement because the weight of an object may vary from one place on the earth to the next.

Mass, on the other hand, is independent of gravity; it is a result of a comparison of an unknown mass with a known mass called a *standard mass.* Balances are instruments used to measure the mass of materials.

The metric unit for mass is the gram (g). A common English unit for mass is the pound (lb).

$$1 \text{ lb} = 454 \text{ g}$$

Examples of balances commonly used for the determination of mass are shown in Figure 1.8.

Length

The standard metric unit of **length,** the distance between two points, is the meter (m). A meter is close to the English yard (yd).

$$1 \text{ yd} = 0.914 \text{ m}$$

LEARNING GOAL

9 Identify the major units of measure in the English and metric systems.

Volume

The standard metric unit of **volume,** the space occupied by an object, is the liter (L). A liter is the volume occupied by 1000 g of water at 4 degrees Celsius (°C).

The English quart (qt) is similar to the liter.

$$1 \text{ qt} = 0.946 \text{ L} \qquad \text{or} \qquad 1.06 \text{ qt} = 1 \text{ L}$$

Volume can be derived using the formula

$$V = \text{length} \times \text{width} \times \text{height}$$

Therefore, volume is commonly reported with a length cubed unit. A cube with the length of each side equal to 1 m will have a volume of 1 m × 1 m × 1 m, or 1 m^3.

$$1 \text{ m}^3 = 1000 \text{ L}$$

The relationships among the units L, mL, and cm^3 are shown in Figure 1.9.

Typical laboratory devices used for volume measurement are shown in Figure 1.10. These devices are calibrated in units of milliliters (mL) or microliters (μL); 1 mL is, by definition, equal to 1 cm^3. The volumetric flask is designed to *contain* a specified volume, and the graduated cylinder, pipet, and buret *dispense* a desired volume of liquid.

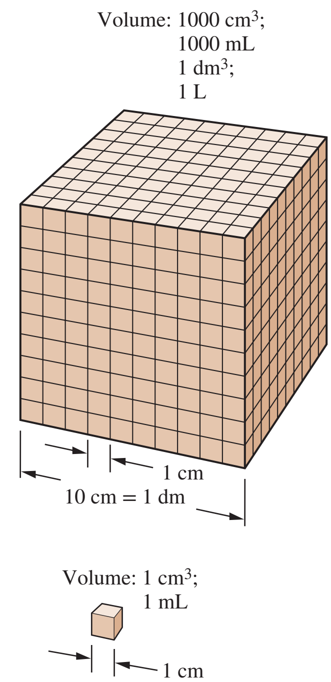


Figure 1.9 The relationships among various volume units.

Time

Time is a measurable period during which an action, process, or condition exists or continues. The standard metric unit of time is the second (s). The need for accurate measurement of time by chemists may not be as apparent as that associated with mass, length, and volume. It is necessary, however, in many applications. In fact, matter may be characterized by measuring the time required for a certain process to occur. The rate of a chemical reaction is a measure of change as a function of time.

1.5 The Numbers of Measurement

LEARNING GOAL

10 Report data and calculate results using scientific notation and the proper number of significant figures.

A measurement has two parts: a number and a unit. The English and metric units of mass, length, volume, and time were discussed in Section 1.4. In this section, we will learn to handle the numbers associated with the measurements.

Information-bearing figures in a number are termed *significant figures.* Data and results arising from a scientific experiment convey information about the way in which the experiment was conducted. The degree of uncertainty or doubt associated with a measurement or series of measurements is indicated by the number of figures used to represent the information.

Significant Figures

Consider the following situation: A student was asked to obtain the length of a section of wire. In the chemistry laboratory, several different types of measuring devices are

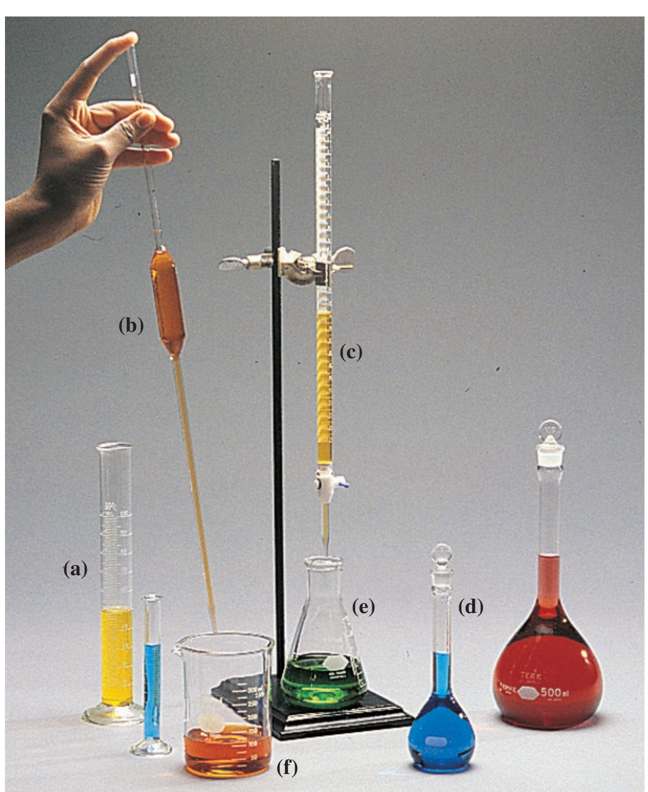
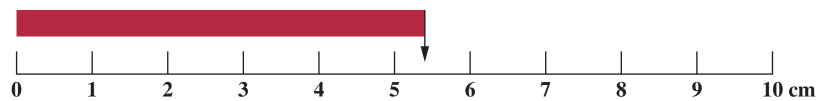
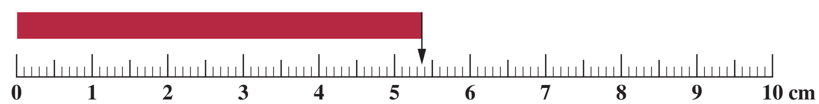


Figure 1.10 Common laboratory equipment used for the measurement of volume. Graduated (a) cylinders, (b) pipets, and (c) burets are used for the delivery of liquids. A graduated cylinder is usually used for measurement of approximate volume; it is less accurate and precise than either pipets or burets. (d) Volumetric flasks are used to contain a specific volume. (e) Erlenmeyer flasks and (f) beakers are not normally used for measuring volumes because they are less accurate than other laboratory equipment. Their volumes should never be used for precise measurements. ©*McGraw-Hill Education/Stephen Frisch, photographer*

usually available. Not knowing which was most appropriate, the student decided to measure the object using each device that was available in the laboratory. To make each measurement, the student determined the mark nearest to the end of the wire. This is depicted in the following figure; the red bar represents the wire being measured. In each case, the student estimated one additional digit by mentally subdividing the marks into ten equal divisions. The following data were obtained:



5.4 cm
(a)



5.36 cm
(b)

In case (a), we are certain that the object is at least 5 cm long and equally certain that it is *not* 6 cm long because the end of the object falls between the calibration lines 5 and 6. We can only estimate between 5 and 6, because there are no calibration indicators between 5 and 6. The end of the wire appears to be approximately four-tenths of the way between 5 and 6, hence 5.4 cm. The 5 is known with certainty, and 4 is estimated (or uncertain).

The uncertain digit results from an estimation.

In case (b), the ruler is calibrated in tenths of a centimeter. The end of the wire is at least 5.3 cm and not 5.4 cm. Estimation of the second decimal place between the two closest calibration marks leads to 5.36 cm. In this case, 5.3 is certain, and the 6 is estimated (or uncertain).

Two questions should immediately come to mind:

1. Are the two answers equivalent?
2. If not, which answer is correct?

In fact, the two answers are *not* equivalent, yet *both* are correct. How do we explain this apparent discrepancy?

The data are not equivalent because each is known to a different degree of certainty. The term **significant figures** is defined to be all digits in a number representing data or results that are known with certainty *plus one uncertain digit.* The answer 5.36 cm, containing three significant figures, specifies the length of the wire more precisely than 5.4 cm, which contains only two significant figures.

The uncertain digit represents the degree of doubt in a single measurement.

Both answers are correct because each is consistent with the measuring device used to generate the data. An answer of 5.36 cm obtained from a measurement using ruler (a) would be *incorrect* because the measuring device is not capable of that precise specification. On the other hand, a value of 5.4 cm obtained from ruler (b) would be erroneous as well; in that case, the measuring device is capable of generating a higher level of certainty (more significant digits) than is actually reported.

In summary, the number of significant figures associated with a measurement is determined by the measuring device. Conversely, the number of significant figures reported is an indication of the precision of the measurement itself.

Recognition of Significant Figures

Only *significant* digits should be reported as data or results. However, are all digits, as written, significant digits? Let's look at a few examples illustrating the rules that are used to represent data and results with the proper number of significant digits.

- All nonzero digits are significant.
  7.314 has *four* significant figures.
- The number of significant digits is independent of the position of the decimal point.
  73.14 has *four* significant figures, as does 7.314.
- Zeros located between nonzero digits are significant.
  60.052 has *five* significant figures.
- Zeros at the end of a number (often referred to as trailing zeros) are significant or not significant depending upon the existence of a decimal point in the number.
  - If there *is* a decimal point, any trailing zeros are significant.
    4.70 has *three* significant figures.
    1000. has *four* significant figures because the decimal point is included.
  - If the number *does not* contain a decimal point, trailing zeros are not significant.
    1000 has *one* significant figure.
- Zeros to the left of the first nonzero integer are not significant; they serve only to locate the position of the decimal point.
  0.0032 has *two* significant figures.

Question 1.9 How many significant figures are contained in each of the following numbers?

a. 7.26 c. 700.2 e. 0.0720
b. 726 d. 7.0 f. 720

Question 1.10 How many significant figures are contained in each of the following numbers?

a. 0.042 c. 24.0 e. 204
b. 4.20 d. 240 f. 2.04

Scientific Notation

LEARNING GOAL

10 Report data and calculate results using scientific notation and the proper number of significant figures.

It is often difficult to express very large numbers to the proper number of significant figures using conventional notation. The solution to this problem lies in the use of **scientific notation,** a system that represents numbers in powers of ten.

The conversion is illustrated as:

$$6200 = 6.2 \times 1000 = 6.2 \times 10^3$$

If we wish to express 6200 with three significant figures, we can write it as:

$$6.20 \times 10^3$$

The trailing zero becomes significant with the existence of the decimal point in the number. Note also that the exponent of 3 has no bearing on the number of significant figures. The value of 6.20×10^{14} also contains three significant figures.

Scientific notation is also referred to as exponential notation. When a number is not written in scientific notation, it is said to be in standard form.

■ **RULE:** To convert a number greater than one to scientific notation, the original decimal point is moved x places to the left, and the resulting number is multiplied by 10^x. The exponent (x) is a *positive* number equal to the number of places the original decimal point was moved.

By convention, in the exponential form, we represent the number with one digit to the left of the decimal point.

Scientific notation is also useful in representing numbers less than one. The conversion is illustrated as:

Scientific notation is written in the format: $y \times 10^x$, in which y represents a number between 1 and 10, and x represents a positive or negative whole number.

$$0.0062 = 6.2 \times \frac{1}{1000} = 6.2 \times \frac{1}{10^3} = 6.2 \times 10^{-3}$$

RULE: To convert a number less than one to scientific notation, the original decimal point is moved x places to the right, and the resulting number is multiplied by 10^{-x}. The exponent $(-x)$ is a *negative* number equal to the number of places the original decimal point was moved.

When a number is exceedingly large or small, scientific notation must be used to enter the number into a calculator. For example, the mass of a single helium atom is a rather cumbersome number as written:

$$0.000000000000000000000006692 \text{ g}$$

Most calculators only allow for the input of nine digits. Scientific notation would express this number as 6.692×10^{-24} g.

Question 1.11 Represent each of the following numbers in scientific notation, showing only significant digits:

a. 0.0024 c. 224 e. 72.420
b. 0.0180 d. 673,000 f. 0.83

Question 1.12 Represent each of the following numbers in scientific notation, showing only significant digits:

a. 48.20 c. 0.126 e. 0.0520
b. 480.0 d. 9,200 f. 822

LEARNING GOAL

11 Distinguish between *accuracy* and *precision* and their representations: *error* and *deviation*.

Accuracy and Precision

The terms *accuracy* and *precision* are often used interchangeably in everyday conversation. However, they have very different meanings when discussing scientific measurement.

Accuracy is the degree of agreement between the true value and the measured value. The measured value may be a single number (such as the mass of an object) or the average value of a series of replicate measurements of the same quantity (reweighing the same object several times). We represent accuracy in terms of **error,** the numerical difference between the measured and true value.

Error is an unavoidable consequence of most laboratory measurements (except counted numbers, discussed later in this section), but not for the reasons you might expect. Spills and contamination are certainly problems in a laboratory, but proper training and a great deal of practice eliminate most of these human errors. Still, errors, *systematic* and *random,* remain.

Systematic errors cause results to be generally higher than the true value or generally lower than the true value. An example would be something as simple as dust on a balance pan, causing each measurement to be higher than the true value. The causes of systematic error can often be discovered and removed. Even after correcting for systematic error, we are still left with random error. Random error is an unavoidable, intrinsic consequence of measurement. Replicate measurements of the same quantity will produce some results greater than the true value and some less than the true value.

When possible, we prefer to make as many replicate measurements of the same quantity to "cancel out" the high (+) and low (−) fluctuations.

Precision is a measure of the agreement within a set of replicate measurements. Just as accuracy is measured in terms of error, precision is represented by **deviation,** the amount of variation present in a set of replicate measurements.

It is important to recognize that accuracy and precision are not the same thing. It is possible to have one without the other. However, when scientific measurements are carefully made, the two most often go hand in hand; high-quality data are characterized by high levels of precision and accuracy.

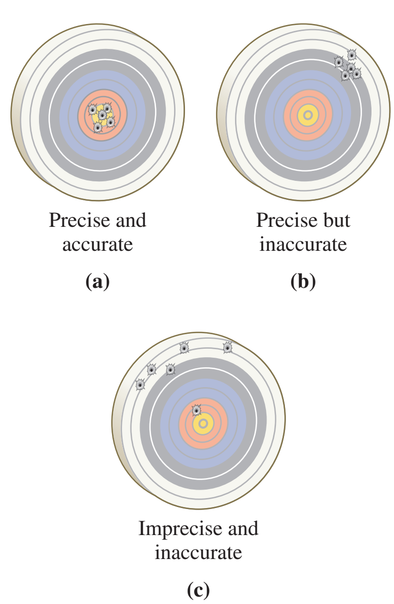


Figure 1.11 An illustration of precision and accuracy in replicate experiments.

In Figure 1.11, bull's-eye (a) shows the goal of all experimentation: accuracy *and* precision. Bull's-eye (b) shows the results to be repeatable (good precision); however, some error in the experimental procedure has caused the results to center on an incorrect value. This error is systematic, occurring in each replicate measurement. Occasionally, an experiment may show "accidental" accuracy. The precision is poor, but the average of replicate measurements leads to a correct value. We don't want to rely on accidental success; the experiment should be repeated until the precision inspires faith in the accuracy of the method. Modern measuring devices in chemistry, equipped with powerful computers with immense storage capacity, are capable of making literally thousands of individual replicate measurements to enhance the quality of the result. In bull's-eye (c), we see a representation of poor precision and poor accuracy. Often, poor precision is accompanied by poor accuracy.

In summary, the presence of error and deviation in most measurements is the real basis for significant figures: all the *certain* digits plus one *uncertain* digit.

Exact (Counted) and Inexact Numbers

Inexact numbers, by definition, have uncertainty (the degree of doubt in the final significant digit). *Exact numbers,* on the other hand, have no uncertainty. Exact numbers may arise from a definition; there are *exactly* 60 min in 1 h or there are exactly 1000 mL in 1 L.

Exact numbers are a consequence of counting. Counting the number of dimes in your pocket or the number of letters in the alphabet are examples. The fact that exact numbers have no uncertainty means that they do not limit the number of significant figures in the result of a calculation.

For example, we may wish to determine the mass of three bolts purchased from the hardware store. Each bolt has a mass of 12.97 g. The total mass is determined by:

$$3 \times 12.97 \text{ g} = 38.91 \text{ g}$$

The number of significant figures in the result is governed by the data (mass of the bolt) and not by the counted (*exact*) number of bolts.

In Section 1.6, we will learn how to convert between units. A good rule of thumb to follow when performing these types of calculations is to use the measured quantity, *not* the conversion factor in order to determine the number of significant figures in the answer.

Rounding Numbers

The use of a calculator generally produces more digits for a result than are justified by the rules of significant figures on the basis of the data input. For example, your calculator may show:

$$3.84 \times 6.72 = 25.8048$$

The most correct answer would be 25.8, dropping 048. A convenient way to show this is:

$$3.84 \times 6.72 = 25.8048 \approx 25.8$$

The rule for multiplication and significant figures dictates three significant figures in the answer.

A number of acceptable conventions for rounding exist. Throughout this book, we will use the following:

■ **RULE:** When the number to be dropped is less than five, the preceding number is not changed. When the number to be dropped is five or larger, the preceding number is increased by one unit.

Question 1.13 Round each of the following numbers to three significant figures.
a. 61.40 b. 6.171 c. 0.066494 d. 63.669 e. 8.7715

Question 1.14 Round each of the following numbers to two significant figures.
a. 6.2262 b. 3895 c. 6.885 d. 2.2247 e. 0.0004109

LEARNING GOAL

10 Report data and calculate results using scientific notation and the proper number of significant figures.

Significant Figures in Calculation of Results

Addition and Subtraction

If we combine the following numbers:

$$\begin{array}{r} 37.68 \\ 108.428 \\ 6.71864 \end{array}$$

our calculator will show a final result of

$$152.82664$$

Remember the distinction between the words *zero* and *nothing*. *Zero* is one of the ten digits and conveys as much information as 1, 2, and so forth. *Nothing* implies no information; the digits in the positions indicated by *x* could be 0, 1, 2, or any other.

Clearly the answer, with eight digits, defines the total much more accurately than *any* of the individual quantities being combined. This cannot be correct; *the answer cannot have greater significance than any of the quantities that produced the answer.* We rewrite the problem:

$$\begin{array}{r|l} 37.68 & xxx \\ 108.42 & 8xx \\ +\quad 6.71 & 864 \\ \hline \underline{152.82} & 664 \end{array} \quad \text{(should be 152.83)}$$

See the rules for rounding discussed earlier in this section.

where x = no information; x may be any integer from 0 to 9. Adding 4 to two unknown numbers (in the rightmost column) produces no information. Similar logic prevails for the previous two columns. Thus, five digits remain, all of which are significant. Conventional rules for rounding would dictate a final answer of 152.83.

Question 1.15 Report the result of each of the following to the proper number of significant figures:

a. $4.26 + 3.831 =$ b. $8.321 - 2.4 =$ c. $16.262 + 4.33 - 0.40 =$

Question 1.16 Report the result of each of the following to the proper number of significant figures:

a. $7.939 + 6.26 =$ b. $2.4 - 8.321 =$ c. $2.333 + 1.56 - 0.29 =$

Adding numbers that are in scientific notation requires a bit more consideration. The numbers must either be converted to standard form or converted to numbers that have the same exponents. Example 1.5 demonstrates this point.

EXAMPLE 1.5 Determining Significant Figures When Adding Numbers in Scientific Notation

Report the result of the following addition to the proper number of significant figures and in scientific notation.

$$9.47 \times 10^{-6} + 9.3 \times 10^{-5}$$

Solution

There are two strategies that may be used in order to arrive at the correct answer.

First solution strategy.

When both numbers are converted to standard form, they can be added together. The initial answer is not the correct answer because it does not have the proper number of significant figures.

$$
\begin{array}{r}
0.000009|47 \\
+\ 0.000093|xx \\
\hline
0.000102|47
\end{array}
$$

After rounding, the answer 0.000102 can then be converted to the final answer, which in scientific notation is 1.02×10^{-4}.

Second solution strategy.
When both numbers have the same power of 10 exponent, they can be added together. In this example, 9.47×10^{-6} is converted to 0.947×10^{-5}.

$$9.47 \times 10^{-6} = 0.947 \times 10^{-5}$$

$$
\begin{array}{r}
0.9|47 \times 10^{-5} \\
+\ \ 9.3|xx \times 10^{-5} \\
\hline
10.2|47 \times 10^{-5}
\end{array}
$$

As in the first solution strategy, the initial answer is rounded to the proper number of significant figures, 10.2×10^{-5}, which is written in scientific notation as 1.02×10^{-4}.

Practice Problem 1.5

Report the result of the addition of $6.72 \times 10^5 + 7.4 \times 10^4$ to the proper number of significant figures and in scientific notation.

LEARNING GOAL 10 Report data and calculate results using scientific notation and the proper number of significant figures.

▶ For Further Practice: **Questions 1.85 b, d; 1.86 d, e.**

Question 1.17 Report the result of the following addition to the proper number of significant figures and in scientific notation.

$8.23 \times 10^{-4} + 6.1 \times 10^{-5}$

Question 1.18 Report the result of the following addition to the proper number of significant figures and in scientific notation.

$4.80 \times 10^8 + 9.149 \times 10^2$

Multiplication and Division

In the preceding discussion of addition and subtraction, the position of the decimal point in the quantities being combined had a bearing on the number of significant figures in the answer. In multiplication and division, this is not the case. The decimal point position is irrelevant when determining the number of significant figures in the answer. It is the number of significant figures in the data that is important. Consider

$$\frac{4.237 \times 1.21 \times 10^{-3} \times 0.00273}{11.125} = 1.26 \times 10^{-6}$$

The answer is limited to three significant figures; the answer can have *only* three significant figures because two numbers in the calculation, 1.21×10^{-3} and 0.00273, have three significant figures and “limit” the answer. *The answer can be no more precise than the least precise number from which the answer is derived.* The *least precise number* is the number with the fewest significant figures.

LEARNING GOAL

10 Report data and calculate results using scientific notation and the proper number of significant figures.

EXAMPLE 1.6 **Determining Significant Figures When Multiplying Numbers in Scientific Notation**

Report the result of the following operation to the proper number of significant figures and in scientific notation.

$$\frac{2.44 \times 10^4}{91}$$

Solution

Often problems that combine multiplication and addition can be broken into parts. This allows each part to be solved in a stepwise fashion.

Step 1. The numerator operation can be completed.

$2.44 \times 10^4 = 24{,}400$

Step 2. This value can now be divided by the value in the denominator.

$$\frac{24{,}400}{91} = 268$$

Step 3. The answer is limited to two significant figures. This is because of the two numbers in the calculation, 2.244 and 91, the number 91 has fewer significant figures and limits the number of significant figures in the answer. The answer in scientific notation is 2.7×10^2.

Practice Problem 1.6

Report the result of the following operation to the proper number of significant figures and in scientific notation.

$$\frac{837}{1.8 \times 10^{-2}}$$

LEARNING GOAL 10 Report data and calculate results using scientific notation and the proper number of significant figures.

▶ For Further Practice: **1.85 a, c, e and 1.86 a, b, c.**

Question 1.19 Report the result of each of the following operations using the proper number of significant figures:

a. $63.8 \times 0.80 =$

b. $\dfrac{63.8}{0.80} =$

c. $\dfrac{16.4 \times 78.11}{22.1} =$

d. $\dfrac{42.2}{21.38 \times 2.3} =$

e. $\dfrac{4.38 \times 10^8}{0.9462} =$

f. $\dfrac{6.1 \times 10^{-4}}{0.3025} =$

Question 1.20 Report the result of each of the following operations using the proper number of significant figures:

a. $\dfrac{27.2 \times 15.63}{1.84} =$

b. $\dfrac{13.6}{18.02 \times 1.6} =$

c. $\dfrac{4.79 \times 10^5}{0.7911} =$

d. $3.58 \times 4.0 =$

e. $\dfrac{3.58}{4.0} =$

f. $\dfrac{11.4 \times 10^{-4}}{0.45} =$

1.6 Unit Conversion

To convert from one unit to another, we must have a *conversion factor* or series of conversion factors that relate two units. The proper use of these conversion factors is called the *factor-label method* or *dimensional analysis.* This method is used for two kinds of conversions: to convert from one unit to another within the *same system* or to convert units from *one system to another.*

LEARNING GOAL

12 Convert between units of the English and metric systems.

Conversion of Units Within the Same System

Based on the information presented in Table 1.1, Section 1.4, we know that in the English system,

$$1 \text{ gal} = 4 \text{ qt}$$

Dividing both sides of the equation by the same term does not change its identity. These ratios are equivalent to unity (1); therefore,

$$\frac{1 \text{ gal}}{1 \text{ gal}} = \frac{4 \text{ qt}}{1 \text{ gal}} = 1$$

Multiplying any other expression by either of these ratios will not change the value of the term because multiplication of any number by 1 produces the original value. However, the units will change.

The speed of an automobile is indicated in both English (mi/h) and metric (km/h) units.
©Ryan McGinnis/Getty Images

Factor-Label Method

When the expressions are written as ratios, they can be used as conversion factors in the factor-label method. For example, if you were to convert 12 gal to quarts, you must decide which conversion factor to use,

$$\frac{1 \text{ gal}}{4 \text{ qt}} \quad \text{or} \quad \frac{4 \text{ qt}}{1 \text{ gal}}$$

Since you are converting 12 gal (Data Given) to qt (Desired Result), it is important to choose a conversion factor with gal in the denominator and qt in the numerator. That way, when the initial quantity (12 gal) is multiplied by the conversion factor, the original unit (gal) will cancel, leaving you with the unit qt in the answer.

$$12 \cancel{\text{gal}} \times \frac{4 \text{ qt}}{1 \cancel{\text{gal}}} = 48 \text{ qt}$$

$$\text{Data Given} \times \text{Conversion Factor} = \text{Desired Result}$$

If the incorrect ratio was selected as a conversion factor, the answer would be incorrect.

$$12 \text{ gal} \times \frac{1 \text{ gal}}{4 \text{ qt}} = \frac{3 \text{ gal}^2}{\text{qt}} \quad \text{(incorrect units)}$$

Therefore, the factor-label method is a self-indicating system. The product will only have the correct units if the conversion factor is set up properly.

The factor-label method is also useful when more than one conversion factor is needed to convert the data given to the desired result. The use of a series of conversion factors is illustrated in Example 1.7.

Conversion factors are used to relate units through the process of the factor-label method (dimensional analysis).

EXAMPLE 1.7 Using English System Conversion Factors

Convert 3.28×10^4 ounces to tons.

Solution

Using the equalities provided in Table 1.1, the data given in ounces (oz) can be directly converted to a bridging data result in pounds (lb), so lb can be converted to the desired result in tons (t). The possible conversion factors are

$$\frac{1 \text{ lb}}{16 \text{ oz}} = \frac{16 \text{ oz}}{1 \text{ lb}} \quad \text{and} \quad \frac{2000 \text{ lb}}{1 \text{ t}} = \frac{1 \text{ t}}{2000 \text{ lb}}$$

Continued

Step 1. Since the initial value is in oz, the conversion factor with oz in the denominator should be used first.

$$3.28 \times 10^4\ \cancel{\text{oz}} \times \frac{1\ \text{lb}}{16\ \cancel{\text{oz}}} = 2.05 \times 10^3\ \text{lb}$$

Data Given × Conversion Factor = Initial Data Result

If the other conversion factor relating oz and lb was used, the resulting units would have been oz^2/lb, and the answer would have been incorrect.

Step 2. Now that 3.28×10^4 oz has been converted to 2.05×10^3 lb, the conversion factor relating lb to t is used. The conversion factor with lb in the denominator and t in the numerator is the only one that leads to the correct answer.

$$2.05 \times 10^3\ \cancel{\text{lb}} \times \frac{1\ \text{t}}{2000\ \cancel{\text{lb}}} = 1.03\ \text{t}$$

Initial Data Result × Conversion Factor = Desired Result

This calculation may also be done in a single step by arranging the conversion factors in a chain:

$$3.28 \times 10^4\ \cancel{\text{oz}} \times \frac{1\ \cancel{\text{lb}}}{16\ \cancel{\text{oz}}} \times \frac{1\ \text{t}}{2000\ \cancel{\text{lb}}} = 1.03\ \text{t}$$

Data Given × Conversion Factor × Conversion Factor = Desired Result

Helpful Hint: After the conversion factors have been selected and set up in the solution to the problem, it is important to also cancel the units that can be canceled. This process will allow for you to verify that you have set up the problem correctly. In addition, the unit ton represents a significantly larger quantity than the unit ounce. Therefore, one would expect a small number of tons to equal a large number of ounces.

Practice Problem 1.7

Convert 360 ft to mi.

LEARNING GOAL 12 Convert between units of the English and metric systems.

▶ For Further Practice: **Questions 1.97 a, b and 1.98 a, b.**

Table 1.2 is located in Section 1.4.

Conversion of units within the metric system may be accomplished by using the factor-label method as well. Unit prefixes that dictate the conversion factor facilitate unit conversion (refer to Table 1.2). Example 1.8 demonstrates this process.

EXAMPLE 1.8 Using Metric System Conversion Factors

Convert 0.0047 kg to mg.

Solution

Using the equalities provided in Table 1.2, the data given in kg can be directly converted to a bridging data result in g, so g can be converted to the desired result in mg. The possible conversion factors are

$$\frac{10^3\ \text{g}}{1\ \text{kg}} = \frac{1\ \text{kg}}{10^3\ \text{g}} \quad \text{and} \quad \frac{10^{-3}\ \text{g}}{1\ \text{mg}} = \frac{1\ \text{mg}}{10^{-3}\ \text{g}}$$

Step 1. Since the initial value is in kg, the conversion factor with kg in the denominator should be used first.

$$0.0047\ \cancel{\text{kg}} \times \frac{10^3\ \text{g}}{1\ \cancel{\text{kg}}} = 4.7\ \text{g}$$

Data Given × Conversion Factor = Initial Data Result

If the other conversion factor relating kg and g was used, the resulting units would have been kg^2/g, and the answer would have been incorrect.

Step 2. Now that 0.0047 kg has been converted to 4.7 g, the conversion factor relating g to mg is used. The conversion factor with g in the denominator and mg in the numerator is the only one that leads to the correct answer.

$$4.7\ \cancel{g} \times \frac{1\ \text{mg}}{10^{-3}\ \cancel{g}} = 4.7 \times 10^{3}\ \text{mg}$$

Initial Data Result × Conversion Factor = Desired Result

This calculation may also be done in a single step by arranging the conversion factors in a chain:

$$0.0047\ \cancel{\text{kg}} \times \frac{10^{3}\ \cancel{g}}{1\ \cancel{\text{kg}}} \times \frac{1\ \text{mg}}{10^{-3}\ \cancel{g}} = 4.7 \times 10^{3}\ \text{mg}$$

Data Given × Conversion Factor × Conversion Factor = Desired Result

Helpful Hint: After the conversion factors have been selected and set up in the solution to the problem, it is important to also cancel the units that can be canceled. This process will allow for you to verify that you have set up the problem correctly. In addition, the unit mg represents a significantly smaller quantity than the unit kg. Therefore, one would expect a large number of mg to equal a small number of kg.

Practice Problem 1.8

Convert:

a. 750 cm to mm
b. 1.5×10^{8} µL to cL
c. 0.00055 Mg to kg

LEARNING GOAL 12 Convert between units of the English and metric systems.

▶ For Further Practice: **Questions 1.99 and 1.100.**

Conversion of Units Between Systems

The conversion of a quantity expressed in a unit of one system to an equivalent quantity in the other system (English to metric or metric to English) requires the use of a relating unit, a conversion factor that relates the two systems. Examples are shown in Table 1.3.

The conversion may be represented as a three-step process:

Step 1. Conversion from the unit given in the problem to a relating unit.

Data Given × Conversion Factor = Relating Unit

Step 2. Conversion to the other system using the relating unit.

Relating Unit × Conversion Factor = Initial Data Result

Step 3. Conversion within the desired system to unit required by the problem.

Initial Data Result × Conversion Factor = Desired Result

Example 1.9 demonstrates a conversion from the English system to the metric system.

TABLE 1.3 Relationships Between Common English and Metric Units

Quantity	English		Metric
Mass	1 pound	=	454 grams
Length	1 yard	=	0.914 meter
Volume	1 quart	=	0.946 liter

EXAMPLE 1.9 Using Both English and Metric System Conversion Factors

Convert 4.00 oz to kg.

Solution

Based on the English system relationships provided in Tables 1.1 and 1.3, the data given in oz should be converted to a relating unit in lb, so lb can be converted to g. Then, using the prefix equalities in Table 1.2, the initial data result in g can be converted to the desired result in kg. The possible conversion factors are:

$$\frac{16\ \text{oz}}{1\ \text{lb}} = \frac{1\ \text{lb}}{16\ \text{oz}} \quad \text{and} \quad \frac{454\ \text{g}}{1\ \text{lb}} = \frac{1\ \text{lb}}{454\ \text{g}} \quad \text{and} \quad \frac{10^3\ \text{g}}{1\ \text{kg}} = \frac{1\ \text{kg}}{10^3\ \text{g}}$$

Step 1. Since the initial value is in oz, the conversion factor with oz in the denominator should be used first because it relates the data given to a relating unit.

$$4.00\ \cancel{\text{oz}} \times \frac{1\ \text{lb}}{16\ \cancel{\text{oz}}} = 0.250\ \text{lb}$$

Data Given × Conversion Factor = Relating Unit

If the other conversion factor relating oz and lb was used, the resulting units would have been oz^2/lb, and the answer would have been incorrect.

Step 2. Now that 4.00 oz has been converted to 0.250 lb, the conversion factor relating lb to g is used. The conversion factor with lb in the denominator and g in the numerator is the only one that leads to the correct answer.

$$0.250\ \cancel{\text{lb}} \times \frac{454\ \text{g}}{1\ \cancel{\text{lb}}} = 114\ \text{g}$$

Relating Unit × Conversion Factor = Initial Data Result

Step 3. In the final step of this conversion, the conversion is within the desired system of units required by the problem. The conversion factor relating g and kg with g in the denominator and kg in the numerator is the only one that leads to the correct answer.

$$114\ \cancel{\text{g}} \times \frac{1\ \text{kg}}{10^3\ \cancel{\text{g}}} = 0.114\ \text{kg}$$

Initial Data Result × Conversion Factor = Desired Result

This calculation may also be done in a single step by arranging the conversion factors in a chain:

$$4.00\ \cancel{\text{oz}} \times \frac{1\ \cancel{\text{lb}}}{16\ \cancel{\text{oz}}} \times \frac{454\ \cancel{\text{g}}}{1\ \cancel{\text{lb}}} \times \frac{1\ \text{kg}}{10^3\ \cancel{\text{g}}} = 0.114\ \text{kg}$$

Data Given × Conversion Factor × Conversion Factor × Conversion Factor = Desired Result

Helpful Hint: After the conversion factors have been selected and set up in the solution to the problem, it is important to also cancel the units that can be canceled. This process will allow for you to verify that you have set up the problem correctly. In addition, the unit oz represents a smaller quantity than the unit kg. Therefore, one would expect a large number of oz to equal a small number of kg.

Practice Problem 1.9

Convert:

a. 0.50 in to m
b. 0.75 qt to L
c. 56.8 g to oz
d. 0.50 in to cm
e. 0.75 qt to mL
f. 56.8 mg to oz

LEARNING GOAL 12 Convert between units of the English and metric systems.

▶ For Further Practice: **Questions 1.101 a, b and 1.102 a, b.**

A Medical Perspective

Curiosity and the Science That Leads to Discovery

Curiosity is one of the most important human traits. Small children constantly ask, "Why?" As we get older, our questions become more complex, but the curiosity remains. Curiosity is also the basis of the scientific method. A scientist observes an event, wonders why it happened, and sets out to answer the question. Dr. Eric Wieschaus's story provides an example of curiosity that led to the discovery of gene pathways that are currently the target of new medicines.

As a child, Dr. Wieschaus dreamed of being an artist, but during the summer following his junior year of high school, he took part in a science program and found his place in the laboratory. When he was a sophomore in college, he accepted a job preparing fly food in a *Drosophila* (fruit fly) lab. Later, while learning about mitosis (cell division) in his embryology course, he became excited about the process of embryonic development. He was fascinated watching how a fertilized frog egg underwent cell division with little cellular growth or differentiation until it formed an embryo. Then, when the embryo grew, the cells in the various locations within the embryo developed differently. As a direct result of his observations, he became determined to understand why certain embryonic cells developed the way they did.

Throughout graduate school, his interest in solving this mystery continued. In his search for the answer, he devised different types of experiments in order to collect data that could explain what caused certain embryonic cells to differentiate into their various shapes, sizes, and positions within the growing embryo. It is these cellular differentiations that determine which cells may become tissues, organs, muscles, or nerves. Although many of his experiments failed, some of the experiments that he completed using normal embryonic cells provided data that led to his next series of experiments in which he used mutated embryos.

After graduate school, Dr. Wieschaus and his colleague, Dr. Christiane Nüsslein-Volhard, used a trial-and-error approach to determine which of the fly's 20,000 genes were essential to embryonic development. They used a chemical to create random mutations in the flies. The mutated flies were bred, and the fly families were analyzed under a microscope. Although the fly embryos are only 0.18 mm in length, the average adult female fly is 2.5 mm long. This allowed for the physical characteristics that resulted from mutated genes to be observed. An artist at heart, Dr. Wieschaus enjoyed this visual work.

Each day was exciting because he knew that at any moment, he could find the answer that he had been seeking for so long. After many years, the team was able to find the genes that controlled the cellular development process within *Drosophila*. The hedgehog gene was one of several genes they identified. It controls a pathway that provides cells with the information they need to develop.

Discoveries about the *Drosophila* embryo have led to advances in medicine.
©janeff/Getty Images

Although it was initially discovered while *Drosophila* embryos were being studied, the hedgehog gene has roles in other adult animals. It has been found that if the hedgehog pathway becomes impaired in humans, basal cell carcinoma (BCC), the most common form of skin cancer, develops. The curiosity that led to the hedgehog gene also led to the discovery of an entirely new type of cancer drug, the first Food and Drug Administration (FDA) approved drug for patients with advanced BCC. This and other types of gene-controlled pathways are allowing for the creation of drugs that target specific diseases. Since these drugs can be designed to be selective, they should also have fewer side effects.

The curiosity that enabled Dr. Wieschaus to advance the field of medicine also catalyzed the development of chemistry. We will see the product of this fundamental human characteristic as we study the work of many extraordinary chemists throughout this textbook.

For Further Understanding

- What is the length of the fruit fly embryo in cm?
- What is the length of the fruit fly embryo in inches?

When a unit is raised to a power, the corresponding conversion factor must also be raised to that power. This ensures that the units cancel properly. Example 1.10 demonstrates how to convert units that are squared or cubed.

EXAMPLE 1.10 Using Conversion Factors Involving Exponents

Convert 1.5 m^2 to cm^2.

Solution

This problem is similar to the conversion problems performed in the previous examples. However, in solving this problem using the factor-label method, the unit exponents must be included.

Using the metric system equalities provided in Table 1.2, the data given in m^2 can be directly converted to the desired result in cm^2. The possible conversion factors are

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

Since the initial value is in m^2, the conversion factor with cm in the denominator should be used. If the incorrect conversion factor was used, the units would not cancel and the result would be m^4 in the numerator and cm^2 in the denominator.

$$1.5\ \cancel{\text{m}^2} \times \frac{1\ \text{cm}}{10^{-2}\ \cancel{\text{m}}} \times \frac{1\ \text{cm}}{10^{-2}\ \cancel{\text{m}}} = 1.5 \times 10^4\ \text{cm}^2$$

$$\text{Data Given} \times (\text{Conversion Factor})^2 = \text{Desired Result}$$

Helpful Hint: When converting a value with a squared unit, the impact of the conversion factor is much greater than if the unit had no exponent. Without the squared unit, the two numbers would be different by a factor of 100; whereas in this example, the two numbers are different by a factor of 10,000.

Practice Problem 1.10

Convert:

a. 1.5 cm^2 to m^2

b. 3.6 m^2 to cm^2

LEARNING GOAL 12 Convert between units of the English and metric systems.

▶ For Further Practice: **Question 1.103.**

Sometimes the unit to be converted is in the denominator. Be sure to set up your conversion factor accordingly. Example 1.11 demonstrates this process.

EXAMPLE 1.11 Converting Units in the Denominator

The density of air is 1.29 g/L. What is the value in g/mL? (Note: Density will be discussed in more detail in Section 1.7.)

Solution

This problem requires the use of one conversion factor. According to the metric system equalities relating mL and L provided in Table 1.2, the possible conversion factors are

$$\frac{10^{-3}\ \text{L}}{1\ \text{mL}} = \frac{1\ \text{mL}}{10^{-3}\ \text{L}}$$

Since the density given has L in the denominator, the conversion factor with L in the numerator should be used. If the incorrect conversion factor was used, the units would not cancel and the result would be g and mL in the numerator and L^2 in the denominator.

$$\frac{1.29\ \text{g}}{\cancel{\text{L}}} \times \frac{10^{-3}\ \cancel{\text{L}}}{1\ \text{mL}} = 1.29 \times 10^{-3}\ \frac{\text{g}}{\text{mL}}$$

$$\text{Data Given} \times \text{Conversion Factor} = \text{Desired Result}$$

Helpful Hint: If the incorrect conversion factor was used, the units would not cancel and the result would be g and mL in the numerator and L^2 in the denominator.

Practice Problem 1.11

Convert 0.791 g/mL to kg/L.

LEARNING GOAL 12 Convert between units of the English and metric systems.

▶ For Further Practice: **Question 1.104.**

It is difficult to overstate the importance of paying careful attention to units and unit conversions. Just one example of the tremendous cost that can result from a "small error" is the loss of a 125-million-dollar Mars-orbiting satellite because of failure to convert from English to metric units during one phase of its construction. As a consequence of this error, the satellite established an orbit too close to Mars and burned up in the Martian atmosphere along with 125 million dollars of the National Aeronautics and Space Administration (NASA) budget.

1.7 Additional Experimental Quantities

In Section 1.4, we introduced the experimental quantities of mass, length, volume, and time. We will now introduce other commonly measured and derived quantities.

LEARNING GOAL

13 Know the three common temperature scales, and convert values from one scale to another.

Temperature

Temperature is the degree of "hotness" of an object. This may not sound like a very "scientific" definition, and, in a sense, it is not. Intuitively, we know the difference between a "hot" and a "cold" object, but developing a precise definition to explain this is not easy. We may think of the temperature of an object as a measure of the amount of heat in the object. However, this is not strictly true. An object increases in temperature because its heat content has increased and vice versa; however, the relationship between heat content and temperature depends on the quantity and composition of the material.

Many substances, such as mercury, expand as their temperature increases, and this expansion provides us with a way to measure temperature and temperature changes. If the mercury is contained within a sealed tube, as it is in a thermometer, the height of the mercury is proportional to the temperature. A mercury thermometer may be calibrated, or scaled, in different units, just as a ruler can be. Three common temperature scales are *Fahrenheit* (*°F*), *Celsius* (*°C*), and *Kelvin* (*K*). Two convenient reference temperatures that are used to calibrate a thermometer are the freezing and boiling temperatures of water. Figure 1.12 shows the relationship between the scales and these reference temperatures.

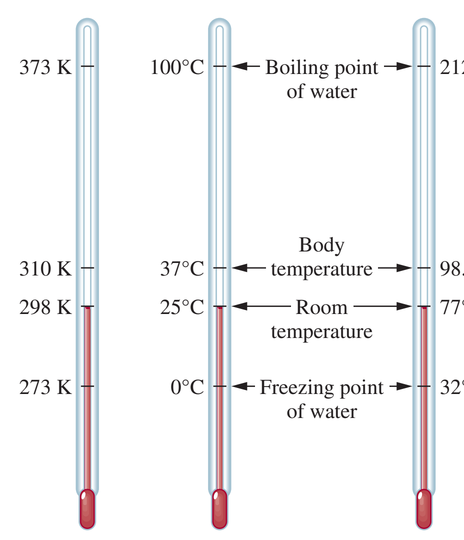


Figure 1.12 The freezing point and boiling point of water, body temperature, and room temperature expressed in the three common units of temperature.

Although Fahrenheit temperature is most familiar to us, Celsius and Kelvin temperatures are used exclusively in scientific measurements. It is often necessary to convert a temperature reading from one scale to another. To convert from Fahrenheit to Celsius, we use the following formula:

$$T_{°C} = \frac{T_{°F} - 32}{1.8}$$

To convert from Celsius to Fahrenheit, we solve this formula for °F, resulting in

$$T_{°F} = (1.8 \times T_{°C}) + 32$$

To convert from Celsius to Kelvin, we use the formula

$$T_K = T_{°C} + 273.15$$

The Kelvin symbol does not have a degree sign. The degree sign implies a value that is *relative* to some standard. Kelvin is an *absolute* scale.

EXAMPLE 1.12 **Converting from Fahrenheit to Celsius and Kelvin**

Normal body temperature is 98.6°F. Calculate the corresponding temperature in both degrees Celsius and Kelvin units and report the answer to the appropriate number of significant figures.

Solution

Using the expression relating °C and °F,

$$T_{°C} = \frac{T_{°F} - 32}{1.8}$$

Substituting the information provided,

$$= \frac{98.6 - 32}{1.8} = \frac{66.6}{1.8}$$

results in:

$$= 37.0°C$$

To calculate the corresponding temperature in Kelvin units, use the expression relating K and °C.

$$T_K = T_{°C} + 273.15$$

Substituting the value obtained in the first part,

$$= 37.0 + 273.15$$

results in:

$$= 310.2 \text{ K}$$

According to Figure 1.12, these three temperatures are at the same place on each thermometer. Therefore, 98.6°F, 37.0°C, and 310.2 K are equivalent.

Practice Problem 1.12

a. The freezing temperature of water is 32°F. Calculate the freezing temperature of water in Celsius units and Kelvin units.

b. When a patient is ill, his or her temperature may increase to 104°F. Calculate the temperature of this patient in Celsius units and Kelvin units.

LEARNING GOAL 13 Know the three common temperature scales, and convert values from one scale to another.

▶ For Further Practice: **Questions 1.121 and 1.122.**

Water in the environment (lakes, oceans, and streams) has a powerful effect on the climate because of its ability to store large quantities of energy. In summer, water stores heat energy and moderates temperatures of the surrounding area. In winter, some of this stored energy is released to the air as the water temperature falls; this prevents the surroundings from experiencing extreme changes in temperature.

Energy

Energy, the ability to do work, may be categorized as either **kinetic energy,** the energy of motion, or **potential energy,** the energy of position. Kinetic energy may be considered as energy in action; potential energy is stored energy. All energy is either kinetic or potential.

Another useful way of classifying energy is by form. The principal forms of energy include light, heat, electrical, mechanical, nuclear, and chemical energy. All of these forms of energy share the following set of characteristics:

- Energy cannot be created or destroyed.
- Energy may be converted from one form to another.
- Conversion of energy from one form to another always occurs with less than 100% efficiency. Energy is not lost (remember, energy cannot be destroyed) but, rather, is not useful. We use gasoline to move our cars from place to place; however, much of the energy stored in the gasoline is released as heat.
- All chemical reactions involve either a “gain” or a “loss” of energy.

Energy absorbed or liberated in chemical reactions is usually in the form of heat energy. Heat energy may be represented in units of *calories* (cal) or *joules* (J), their relationship being

$$1 \text{ cal} = 4.18 \text{ J}$$

One calorie is defined as the amount of heat energy required to increase the temperature of 1 g of water 1°C.

Heat energy measurement is a quantitative measure of heat content. It is an extensive property, dependent upon the quantity of material. Temperature, as we have mentioned, is an intensive property, independent of quantity.

Not all substances have the same capacity for holding heat; 1 g of iron and 1 g of water, even if they are at the same temperature, do *not* contain the same amount of heat energy. One gram of iron will absorb and store 0.108 cal of heat energy when the temperature is raised 1°C. In contrast, 1 g of water will absorb almost ten times as much energy, 1.00 cal, when the temperature is increased an equivalent amount.

Units for other forms of energy will be introduced in later chapters.

The *kilocalorie* (kcal) is the familiar nutritional calorie. It is also known as the large Calorie (Cal); note that in this term the Cal is uppercase to distinguish it from the normal calorie. The large Calorie is 1000 normal calories. Refer to A Human Perspective: Food Calories for more information.

Question 1.21 Convert 595 cal to units of J.

Question 1.22 Convert 2.00×10^2 J to units of cal.

Concentration

Concentration is a measure of the number or mass of particles of a substance that are contained in a specified volume. Examples include:

- The concentration of oxygen in the air
- Pollen counts, given during the hay fever seasons, which are simply the number of grains of pollen contained in a measured volume of air
- The amount of an illegal drug in a certain volume of blood, indicating the extent of drug abuse
- The proper dose of an antibiotic, based on a patient's weight

We will describe many situations in which concentration is used to predict useful information about chemical reactions (Chapters 6–8, for example). In Chapter 6, we calculate a numerical value for concentration from experimental data.

LEARNING GOAL

14 Use density, mass, and volume in problem solving, and calculate the specific gravity of a substance from its density.

Density and Specific Gravity

Both mass and volume are functions of the *amount* of material present (extensive properties). **Density,** the ratio of mass to volume,

$$\text{Density } (d) = \frac{\text{mass}}{\text{volume}} = \frac{m}{V}$$

is *independent* of the amount of material (intensive property). Density is a useful way to characterize or identify a substance because each substance has a unique density (Figure 1.13).

Figure 1.13 Density (mass/volume) is a unique property of a material. A mixture of water and oil is shown, with lithium—the least dense—floating on the oil. The oil, with a density greater than lithium, but less than water, floats on the interface between lithium and water.
©McGraw-Hill Education/Stephen Frisch, photographer

In density calculations, mass is usually represented in g, and volume is given in either mL, cm^3, or cc:

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

The unit of density would therefore be g/mL, g/cm^3, or g/cc. It is important to recognize that because the units of density are a ratio of mass to volume, density can be used as a conversion factor when the factor-label method is used to solve for either mass or volume from density data.

A 1-mL sample of air and 1-mL sample of iron have different masses. There is much more mass in 1 mL of iron; its density is greater. Density measurements were used to distinguish between real gold and "fool's gold" during the gold rush era. Today, the measurement of the density of a substance is still a valuable analytical technique. The densities of a number of common substances are shown in Table 1.4.

A Human Perspective

Food Calories

The body gets its energy through the processes known collectively as metabolism, which will be discussed in detail in Chapters 21–23. The primary energy sources for the body are carbohydrates, fats, and proteins, which we obtain from the foods we eat. The amount of energy available from a given foodstuff is related to the Calories (Cal) available in the food. Calories are a measure of energy that can be derived from food. One (food) Calorie (symbolized by Cal) equals 1000 (metric) calories (symbolized by cal):

$$1 \text{ Cal} = 1000 \text{ cal} = 1 \text{ kcal}$$

The energy available in food can be measured by completely burning the food; in other words, using the food as fuel. The energy given off in the form of heat is directly related to the amount of chemical energy, the energy stored in chemical bonds, that is available in the food. Food provides energy to the body through various metabolic pathways.

The classes of food molecules are not equally energy-rich. When oxidized via metabolic pathways, carbohydrates and proteins provide the cell with 4 Cal/g, whereas fats generate approximately 9 Cal/g.

In addition, as with all processes, not all the available energy can be efficiently extracted from the food; a certain percentage is always released to the surroundings as heat. The average person requires between 2000 and 3000 Cal/day to maintain normal body functions such as the regulation of body temperature and muscle movement. If a person takes in more Cal than the body uses, the person will gain weight. Conversely, if a person uses more Cal than are ingested, the individual will lose weight.

Excess Cal are stored in the form of fat, the form that provides the greatest amount of energy per g. Too many Cal lead to too much fat. Similarly, a lack of Cal (in the form of food) forces the body to raid its storehouse, the fat. Weight is lost in this process as the fat is consumed. Unfortunately, it always seems easier to add fat to the storehouse than to remove it.

The "rule of thumb" is that 3500 Cal are equivalent to approximately 1 lb of body fat. You have to take in 3500 Cal more than you use to gain 1 lb, and you have to expend 3500 Cal more than you normally use to lose 1 lb. If you eat as few as 100 Cal/day beyond your body's needs, you could gain about 10–11 lb per year (yr):

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$$\frac{100 \cancel{\text{Cal}}}{\cancel{\text{day}}} \times \frac{365 \cancel{\text{day}}}{1 \text{ yr}} \times \frac{1 \text{ lb}}{3500 \cancel{\text{Cal}}} = \frac{10.4 \text{ lb}}{\text{yr}}$$

A frequently recommended procedure for increasing the rate of weight loss involves a combination of dieting (taking in fewer Cal) and exercise. Running, swimming, jogging, and cycling are particularly efficient forms of exercise. Running burns 0.11 Cal/min for every lb of body weight, and swimming burns approximately 0.05 Cal/min for every lb of body weight.

For Further Understanding

- Sarah runs 1 h each day, and Nancy swims 2 h each day. Assuming that Sarah and Nancy are the same weight, which girl burns more calories in 1 week?
- Would you expect a runner to burn more calories in summer or winter? Why?

TABLE 1.4 Densities of Some Common Materials

Substance	Density (g/mL)	Substance	Density (g/mL)
Air	0.00129 (at 0°C)	Mercury	13.6
Ammonia	0.000771 (at 0°C)	Methanol	0.792
Benzene	0.879	Milk	1.028–1.035
Blood	1.060	Oxygen	0.00143 (at 0°C)
Bone	1.7–2.0	Rubber	0.9–1.1
Carbon dioxide	0.001963 (at 0°C)	Turpentine	0.87
Ethanol	0.789	Urine	1.010–1.030
Gasoline	0.66–0.69	Water	1.000 (at 4°C)
Gold	19.3	Water	0.998 (at 20°C)
Hydrogen	0.000090 (at 0°C)	Wood (balsa, least dense; ebony and teak, most dense)	0.3–0.98
Kerosene	0.82		
Lead	11.3		