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CHEMISTRY

AN ATOMS FIRST APPROACH

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Third Edition

CHEMISTRY
AN ATOMS FIRST APPROACH

THIRD EDITION

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***Chemistry: An Atoms First Approach,*
Third Edition**
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Donald J. DeCoste**

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Library of Congress Control Number: 2019918615

ISBN: 978-0-357-36336-2

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

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Printed in the United States of America
Print Number: 01 Print Year: 2019

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To the Professor

Features of *Chemistry: An Atoms First Approach*

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

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

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

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

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

What's New

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

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

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

Hallmarks of *Chemistry: An Atoms First Approach*

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

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

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

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

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

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

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

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

Supporting Materials

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

Acknowledgments

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

We are especially grateful to Tom Hummel, University of Illinois at Urbana–Champaign, who managed the end-of chapter problems and the solutions manuals. We greatly appreciate the outstanding work of Gretchen Adams, University of Illinois at Urbana–Champaign, who created the *Interactive Examples* and who converted many of the end-of-chapter problems to ChemWork problems. We are deeply indebted to these two colleagues who contribute in so many ways to our various projects. We owe a large debt of appreciation to Professor Tom McGrath of Baylor University for providing a great deal of useful insight as we prepared the third edition. We followed many of Professor McGrath's suggestions as we revised the text.

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

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

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To the Student

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

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

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

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

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

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

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

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

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

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

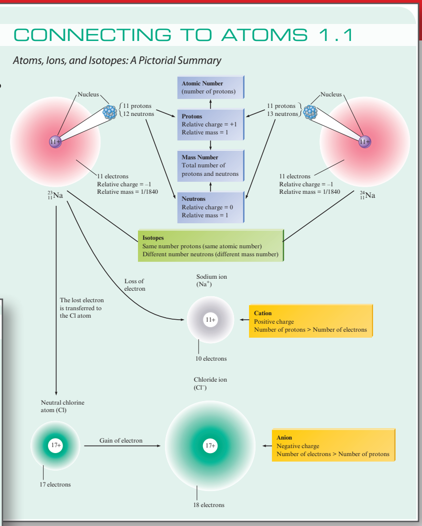
Steve and Susan Zumdahl and Don DeCoste

A GUIDE TO *Chemistry, An Atoms First Approach* THIRD EDITION

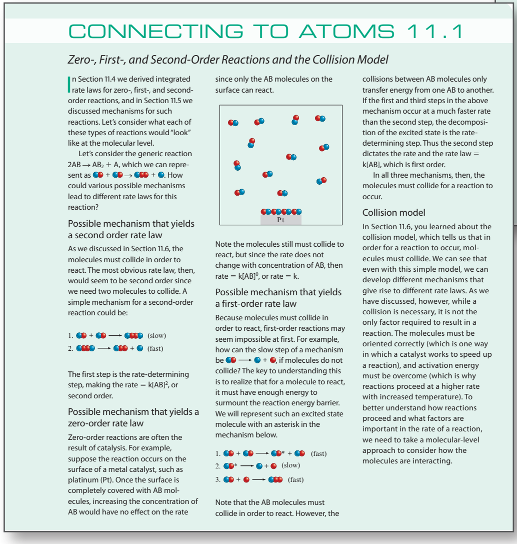
Connecting To Atoms This new boxed feature contains atoms first summaries that cover key atoms first concepts at point-of-use.

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

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


CONNECTING TO ATOMS 1.1

Atoms, Ions, and Isotopes: A Pictorial Summary


CONNECTING TO ATOMS 11.1

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

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

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

Possible mechanism that yields a second order rate law

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

1. ●● + ●● ⟶ ●●●● (slow)
2. ●●●● ⟶ ●●● + ● (fast)

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

Possible mechanism that yields a zero-order rate law

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

Note the molecules still must collide to react, but since the rate does not change with concentration of AB, then rate $= k[AB]^0$, or rate $= k$.

Possible mechanism that yields a first-order rate law

Because molecules must collide in order to react, first-order reactions may seem impossible at first. For example, how can the slow step of a mechanism be ●● ⟶ ● + ●, if molecules do not collide? The key to understanding this is to realize that for a molecule to react, it must have enough energy to surmount the reaction energy barrier. We will represent such an excited state molecule with an asterisk in the mechanism below.

1. ●● + ●● ⟶ ●●* + ●● (fast)
2. ●●* ⟶ ● + ● (slow)
3. ●● + ● ⟶ ●●● (fast)

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

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

Collision model

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


CHEMICAL CONNECTIONS 3.1

No Lead Pencils

Did you ever wonder why the part of a pencil that makes the mark is called the "lead"? Pencils have no lead in them now—and they never have. Apparently the association between writing and the element lead arose during the Roman Empire, when lead rods were used as writing utensils because they leave a gray mark on paper. Many centuries later, in 1564, a deposit of a black substance found to be very useful for writing was discovered in Borrowdale, England. This substance, originally called "black lead," was shown in 1879 by Swedish chemist Carl Scheele to be a form of carbon and was subsequently named graphite (after the Greek *graphein,* meaning "to write").

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

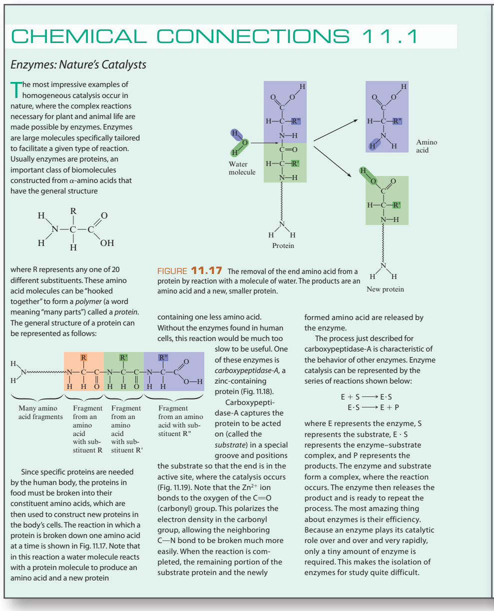
Although Borrowdale graphite was pure enough to use directly, most graphite must be mixed with other materials to be useful for writing instruments. In 1795, the French chemist Nicolas-Jaques Conté invented a process in which graphite is mixed with clay and water to produce pencil "lead," a recipe that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

The ratio of clay to graphite is adjusted to vary the hardness of the lead—the more clay in the mix, the harder the lead and the lighter the line it makes.

Pencils are made from a slat of wood with several grooves cut in it to hold the leads. A similar grooved slat is then placed on top and glued to form a "sandwich" from which individual pencils are cut, sanded smooth, and painted. Although many types of wood have been used over the years to make pencils, the current favorite is incense cedar from the Sierra Nevada Mountains of California.

Modern pencils are simple but amazing instruments. The average pencil can write approximately 45,000 words, which is equivalent to a line 35 miles long. The graphite in a pencil is easily transferred to paper because graphite contains layers of carbon atoms bound together in a "chicken-wire" structure. Although the bonding *within* each layer is very strong, the bonding *between* layers is weak, giving graphite its slippery, soft nature. In this way, graphite is much different from diamond, the other common elemental form of carbon. In diamond the carbon atoms are bound tightly in all three dimensions, making it extremely hard—the hardest natural substance.

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


CHEMICAL CONNECTIONS 11.1

Enzymes: Nature's Catalysts

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

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

Since specific proteins are needed by the human body, the proteins in food must be broken into their constituent amino acids, which are then used to construct new proteins in the body's cells. The reaction in which a protein is broken down one amino acid at a time is shown in Fig. 11.17. Note that in this reaction a water molecule reacts with a protein molecule to produce an amino acid and a new protein containing one less amino acid. Without the enzymes found in human cells, this reaction would be much too slow to be useful. One of these enzymes is *carboxypeptidase-A,* a zinc-containing protein (Fig. 11.18).

Carboxypeptidase-A captures the protein to be acted on (called the *substrate*) in a special groove and positions the substrate so that the end is in the active site, where the catalysis occurs (Fig. 11.19). Note that the Zn^{2+} ion bonds to the oxygen of the C═O (carbonyl) group. This polarizes the electron density in the carbonyl group, allowing the neighboring C—N bond to be broken much more easily. When the reaction is completed, the remaining portion of the substrate protein and the newly formed amino acid are released by the enzyme.

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

$$E + S \longrightarrow E \cdot S$$
$$E \cdot S \longrightarrow E + P$$

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

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

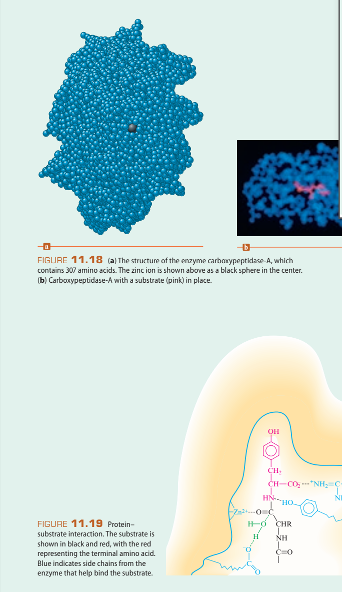


FIGURE 11.18 (**a**) The structure of the enzyme carboxypeptidase-A, which contains 307 amino acids. The zinc ion is shown above as a black sphere in the center. (**b**) Carboxypeptidase-A with a substrate (pink) in place.

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

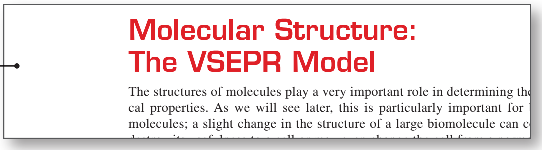
Chemical Connections describe current applications of chemistry. These special-interest boxes cover such topics as the invention of Post-it Notes, farming the wind, and the use of iron metal to clean up contaminated groundwater. Additional *Chemical Connections* are available on the student Web site.

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

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

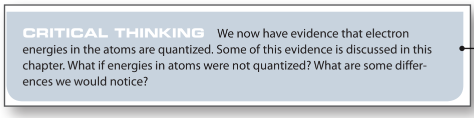
—Democritus

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



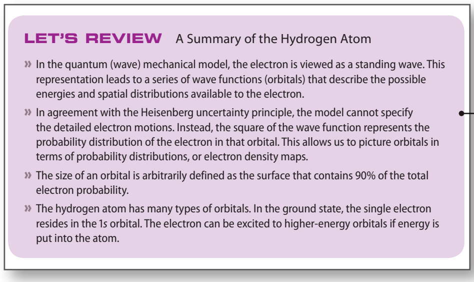
Molecular Structure: The VSEPR Model

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



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

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

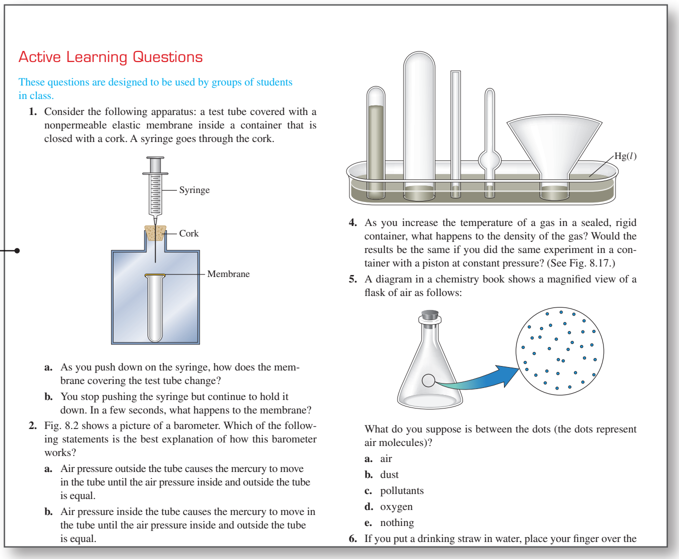


LET’S REVIEW A Summary of the Hydrogen Atom

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

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

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



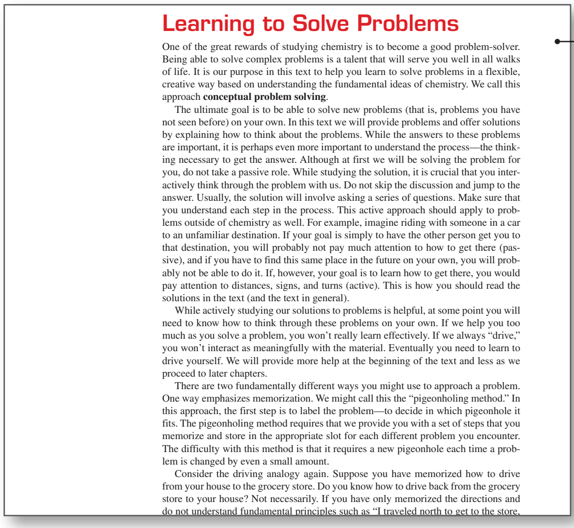
Active Learning Questions

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

1. Consider the following apparatus: a test tube covered with a nonpermeable elastic membrane inside a container that is closed with a cork. A syringe goes through the cork.
   - **a.** As you push down on the syringe, how does the membrane covering the test tube change?
   - **b.** You stop pushing the syringe but continue to hold it down. In a few seconds, what happens to the membrane?
2. Fig. 8.2 shows a picture of a barometer. Which of the following statements is the best explanation of how this barometer works?
   - **a.** Air pressure outside the tube causes the mercury to move in the tube until the air pressure inside and outside the tube is equal.
   - **b.** Air pressure inside the tube causes the mercury to move in the tube until the air pressure inside and outside the tube is equal.
4. As you increase the temperature of a gas in a sealed, rigid container, what happens to the density of the gas? Would the results be the same if you did the same experiment in a container with a piston at constant pressure? (See Fig. 8.17.)
5. A diagram in a chemistry book shows a magnified view of a flask of air as follows:

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

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



Learning to Solve Problems

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

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

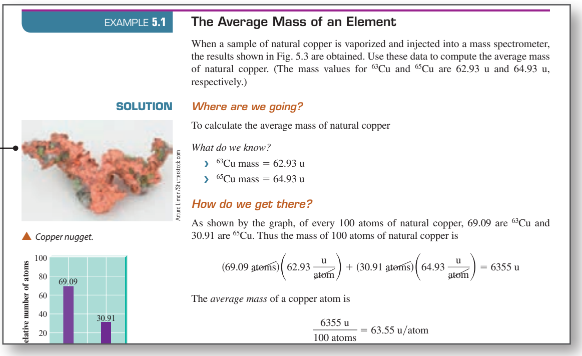
While actively studying our solutions to problems is helpful, at some point you will need to know how to think through these problems on your own. If we help you too much as you solve a problem, you won't really learn effectively. If we always "drive," you won't interact as meaningfully with the material.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

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

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

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

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



EXAMPLE 5.1 **The Average Mass of an Element**

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

SOLUTION ***Where are we going?***

To calculate the average mass of natural copper

What do we know?

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

How do we get there?

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

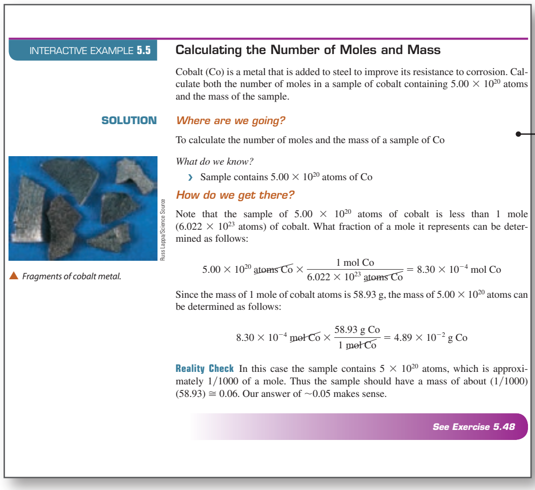
$$(69.09\ \cancel{\text{atoms}})\left(62.93\ \frac{\text{u}}{\cancel{\text{atom}}}\right) + (30.91\ \cancel{\text{atoms}})\left(64.93\ \frac{\text{u}}{\cancel{\text{atom}}}\right) = 6355\ \text{u}$$

The *average mass* of a copper atom is

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

Copper nugget.

Arturo Limon/Shutterstock.com



INTERACTIVE EXAMPLE 5.5 **Calculating the Number of Moles and Mass**

Cobalt (Co) is a metal that is added to steel to improve its resistance to corrosion. Calculate both the number of moles in a sample of cobalt containing 5.00×10^{20} atoms and the mass of the sample.

SOLUTION ***Where are we going?***

To calculate the number of moles and the mass of a sample of Co

What do we know?

- Sample contains 5.00×10^{20} atoms of Co

How do we get there?

Note that the sample of 5.00×10^{20} atoms of cobalt is less than 1 mole (6.022×10^{23} atoms) of cobalt. What fraction of a mole it represents can be determined as follows:

$$5.00 \times 10^{20}\ \cancel{\text{atoms Co}} \times \frac{1\ \text{mol Co}}{6.022 \times 10^{23}\ \cancel{\text{atoms Co}}} = 8.30 \times 10^{-4}\ \text{mol Co}$$

Since the mass of 1 mole of cobalt atoms is 58.93 g, the mass of 5.00×10^{20} atoms can be determined as follows:

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

Reality Check In this case the sample contains 5×10^{20} atoms, which is approximately 1/1000 of a mole. Thus the sample should have a mass of about (1/1000)(58.93) ≅ 0.06. Our answer of ~0.05 makes sense.

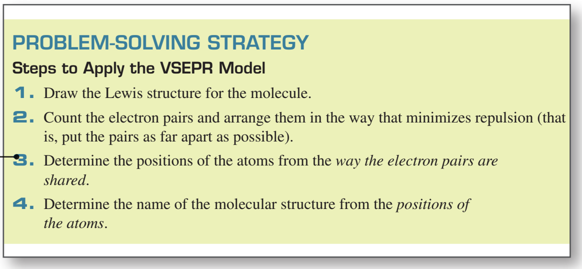
See Exercise 5.48

Fragments of cobalt metal.

Russ Lappa/Science Source

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

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



PROBLEM-SOLVING STRATEGY

Steps to Apply the VSEPR Model

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

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

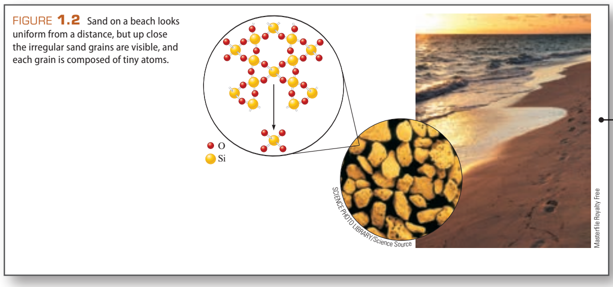


FIGURE **1.2** Sand on a beach looks uniform from a distance, but up close the irregular sand grains are visible, and each grain is composed of tiny atoms.

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

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

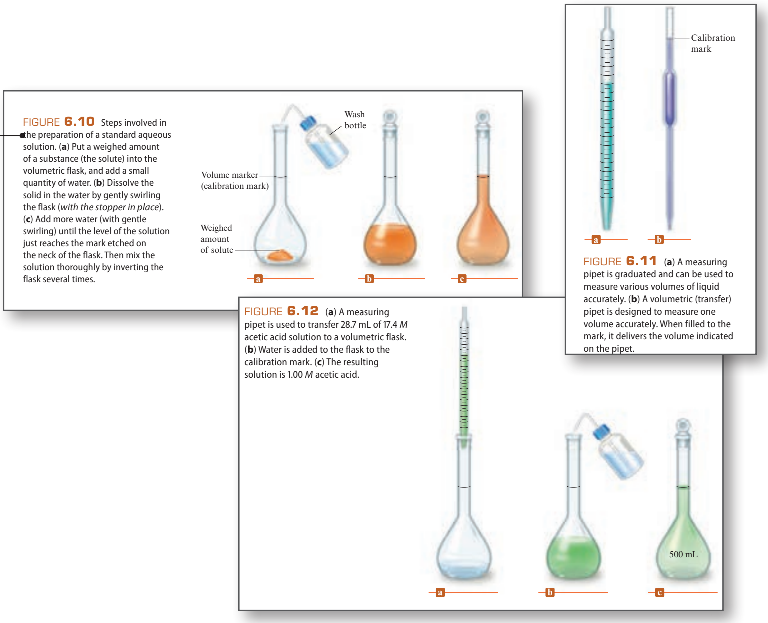


FIGURE **6.10** Steps involved in the preparation of a standard aqueous solution. (**a**) Put a weighed amount of a substance (the solute) into the volumetric flask, and add a small quantity of water. (**b**) Dissolve the solid in the water by gently swirling the flask (*with the stopper in place*).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

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

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

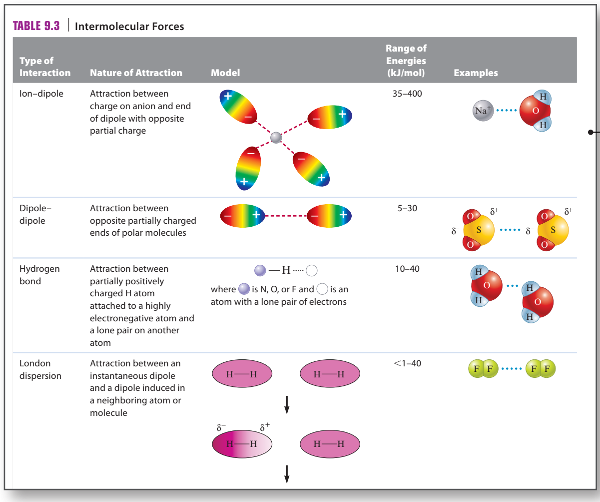
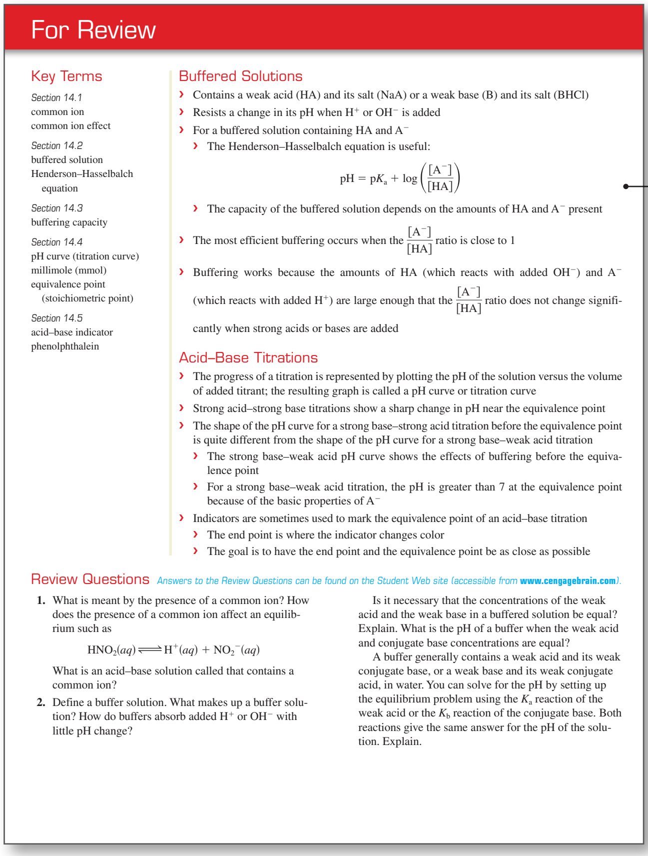


TABLE 9.3 | Intermolecular Forces

Type of Interaction	Nature of Attraction	Model	Range of Energies (kJ/mol)	Examples
Ion–dipole	Attraction between charge on anion and end of dipole with opposite partial charge		35–400	Na^+ ······ H_2O
Dipole–dipole	Attraction between opposite partially charged ends of polar molecules		5–30	δ^- SO_2 δ^+ ······ δ^- SO_2 δ^+
Hydrogen bond	Attraction between partially positively charged H atom attached to a highly electronegative atom and a lone pair on another atom	● —H ····· ○ where ● is N, O, or F and ○ is an atom with a lone pair of electrons	10–40	H_2O ··· H_2O
London dispersion	Attraction between an instantaneous dipole and a dipole induced in a neighboring atom or molecule	H—H H—H ↓ δ^- H—H δ^+ H—H ↓	<1–40	F F ······ F F

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

Comprehensive End-of-Chapter Practice and Review We offer end-of-chapter exercises for every type of student and for every kind of homework assignment.



For Review

Key Terms

Section 14.1
common ion
common ion effect

Section 14.2
buffered solution
Henderson–Hasselbalch equation

Section 14.3
buffering capacity

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

Section 14.5
acid–base indicator
phenolphthalein

Buffered Solutions

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

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

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

Acid–Base Titrations

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

Review Questions

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

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

$$HNO_2(aq) \rightleftharpoons H^+(aq) + NO_2^-(aq)$$

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

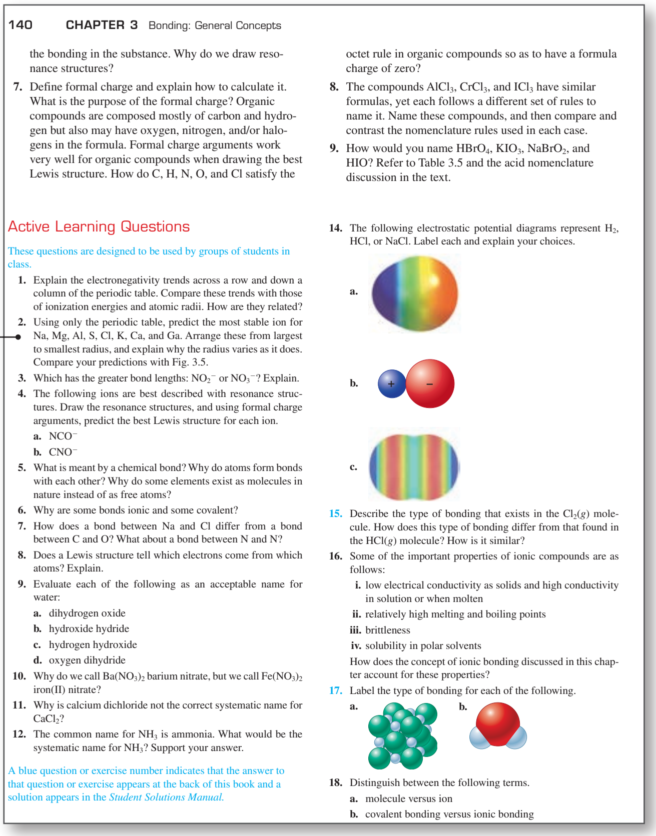
2. Define a buffer solution. What makes up a buffer solution? How do buffers absorb added H^+ or OH^- with little pH change?

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

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

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

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



140 **CHAPTER 3** Bonding: General Concepts

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

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

8. The compounds $AlCl_3$, $CrCl_3$, and ICl_3 have similar formulas, yet each follows a different set of rules to name it. Name these compounds, and then compare and contrast the nomenclature rules used in each case.

9. How would you name $HBrO_4$, KIO_3, $NaBrO_2$, and HIO? Refer to Table 3.5 and the acid nomenclature discussion in the text.

Active Learning Questions

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

1. Explain the electronegativity trends across a row and down a column of the periodic table. Compare these trends with those of ionization energies and atomic radii. How are they related?

2. Using only the periodic table, predict the most stable ion for Na, Mg, Al, S, Cl, K, Ca, and Ga. Arrange these from largest to smallest radius, and explain why the radius varies as it does. Compare your predictions with Fig. 3.5.

3. Which has the greater bond lengths: NO_2^- or NO_3^-? Explain.

4. The following ions are best described with resonance structures. Draw the resonance structures, and using formal charge arguments, predict the best Lewis structure for each ion.

- **a.** NCO^-
- **b.** CNO^-

5. What is meant by a chemical bond? Why do atoms form bonds with each other? Why do some elements exist as molecules in nature instead of as free atoms?

6. Why are some bonds ionic and some covalent?

7. How does a bond between Na and Cl differ from a bond between C and O? What about a bond between N and N?

8. Does a Lewis structure tell which electrons come from which atoms? Explain.

9. Evaluate each of the following as an acceptable name for water:

- **a.** dihydrogen oxide
- **b.** hydroxide hydride
- **c.** hydrogen hydroxide
- **d.** oxygen dihydride

10. Why do we call $Ba(NO_3)_2$ barium nitrate, but we call $Fe(NO_3)_2$ iron(II) nitrate?

11. Why is calcium dichloride not the correct systematic name for $CaCl_2$?

12. The common name for NH_3 is ammonia. What would be the systematic name for NH_3? Support your answer.

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

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

a.

b.

c.

15. Describe the type of bonding that exists in the $Cl_2(g)$ molecule. How does this type of bonding differ from that found in the $HCl(g)$ molecule? How is it similar?

16. Some of the important properties of ionic compounds are as follows:

- **i.** low electrical conductivity as solids and high conductivity in solution or when molten
- **ii.** relatively high melting and boiling points
- **iii.** brittleness
- **iv.** solubility in polar solvents

How does the concept of ionic bonding discussed in this chapter account for these properties?

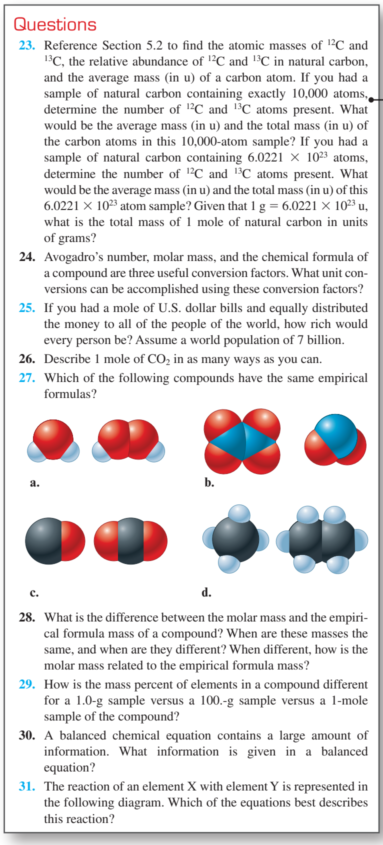
17. Label the type of bonding for each of the following.

a. **b.**

18. Distinguish between the following terms.

- **a.** molecule versus ion
- **b.** covalent bonding versus ionic bonding

Comprehensive End-of-Chapter Practice and Review



Questions

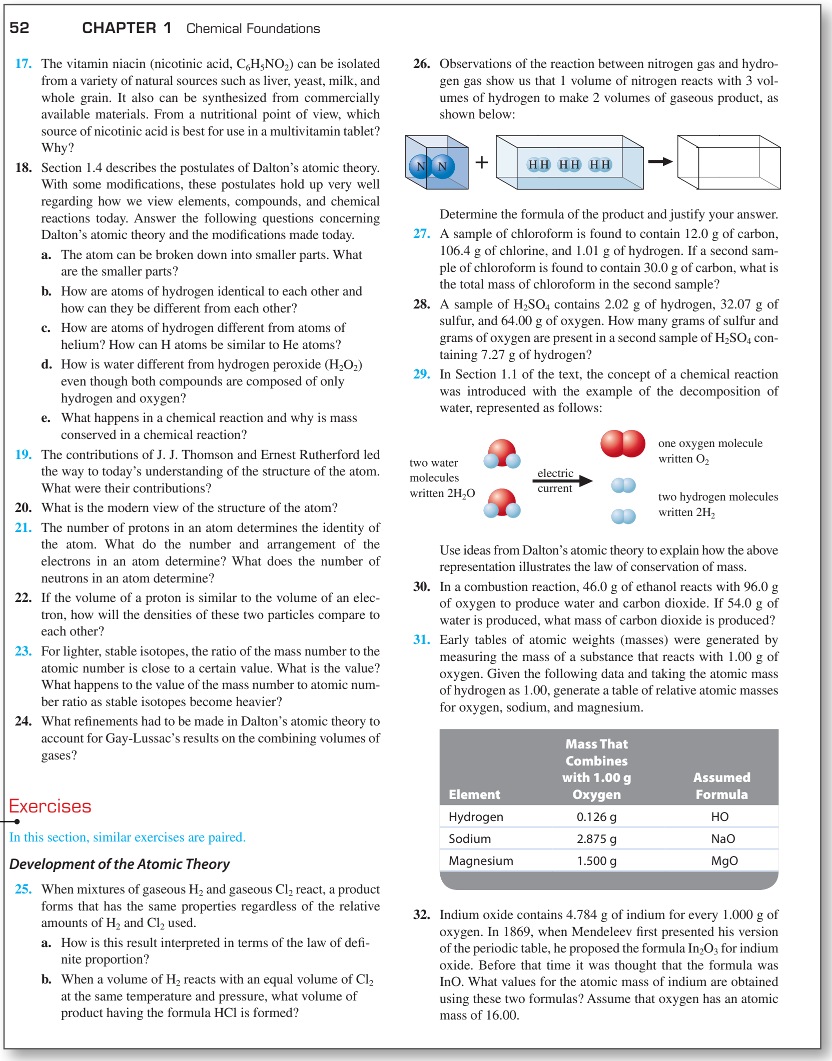
23. Reference Section 5.2 to find the atomic masses of ^{12}C and ^{13}C, the relative abundance of ^{12}C and ^{13}C in natural carbon, and the average mass (in u) of a carbon atom. If you had a sample of natural carbon containing exactly 10,000 atoms, determine the number of ^{12}C and ^{13}C atoms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

a. **b.**

c. **d.**

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

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



52 **CHAPTER 1** Chemical Foundations

17. The vitamin niacin (nicotinic acid, $C_6H_5NO_2$) can be isolated from a variety of natural sources such as liver, yeast, milk, and whole grain. It also can be synthesized from commercially available materials. From a nutritional point of view, which source of nicotinic acid is best for use in a multivitamin tablet? Why?
18. Section 1.4 describes the postulates of Dalton's atomic theory. With some modifications, these postulates hold up very well regarding how we view elements, compounds, and chemical reactions today. Answer the following questions concerning Dalton's atomic theory and the modifications made today.
    - **a.** The atom can be broken down into smaller parts. What are the smaller parts?
    - **b.** How are atoms of hydrogen identical to each other and how can they be different from each other?
    - **c.** How are atoms of hydrogen different from atoms of helium? How can H atoms be similar to He atoms?
    - **d.** How is water different from hydrogen peroxide (H_2O_2) even though both compounds are composed of only hydrogen and oxygen?
    - **e.** What happens in a chemical reaction and why is mass conserved in a chemical reaction?
19. The contributions of J. J. Thomson and Ernest Rutherford led the way to today's understanding of the structure of the atom. What were their contributions?
20. What is the modern view of the structure of the atom?
21. The number of protons in an atom determines the identity of the atom. What do the number and arrangement of the electrons in an atom determine? What does the number of neutrons in an atom determine?
22. If the volume of a proton is similar to the volume of an electron, how will the densities of these two particles compare to each other?
23. For lighter, stable isotopes, the ratio of the mass number to the atomic number is close to a certain value. What is the value? What happens to the value of the mass number to atomic number ratio as stable isotopes become heavier?
24. What refinements had to be made in Dalton's atomic theory to account for Gay-Lussac's results on the combining volumes of gases?

Exercises

In this section, similar exercises are paired.

Development of the Atomic Theory

25. When mixtures of gaseous H_2 and gaseous Cl_2 react, a product forms that has the same properties regardless of the relative amounts of H_2 and Cl_2 used.
    - **a.** How is this result interpreted in terms of the law of definite proportion?
    - **b.** When a volume of H_2 reacts with an equal volume of Cl_2 at the same temperature and pressure, what volume of product having the formula HCl is formed?
26. Observations of the reaction between nitrogen gas and hydrogen gas show us that 1 volume of nitrogen reacts with 3 volumes of hydrogen to make 2 volumes of gaseous product, as shown below:

Determine the formula of the product and justify your answer.

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

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

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

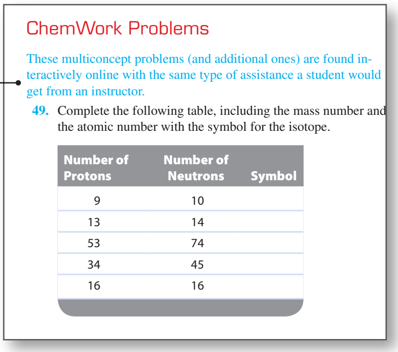
Element	Mass That Combines with 1.00 g Oxygen	Assumed Formula
Hydrogen	0.126 g	HO
Sodium	2.875 g	NaO
Magnesium	1.500 g	MgO

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

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

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

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

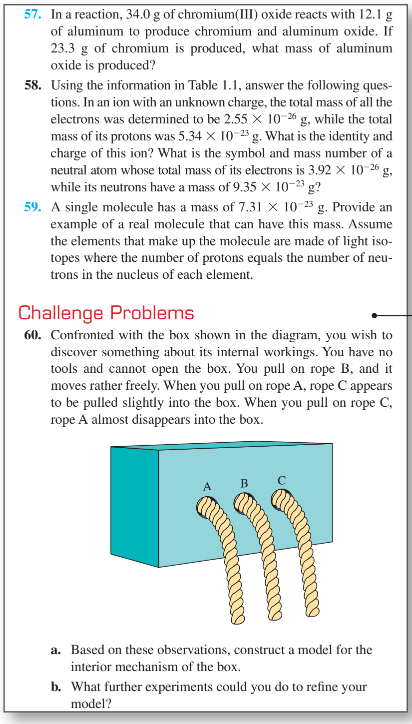


ChemWork Problems

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

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

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



57. In a reaction, 34.0 g of chromium(III) oxide reacts with 12.1 g of aluminum to produce chromium and aluminum oxide. If 23.3 g of chromium is produced, what mass of aluminum oxide is produced?

58. Using the information in Table 1.1, answer the following questions. In an ion with an unknown charge, the total mass of all the electrons was determined to be 2.55×10^{-26} g, while the total mass of its protons was 5.34×10^{-23} g. What is the identity and charge of this ion? What is the symbol and mass number of a neutral atom whose total mass of its electrons is 3.92×10^{-26} g, while its neutrons have a mass of 9.35×10^{-23} g?

59. A single molecule has a mass of 7.31×10^{-23} g. Provide an example of a real molecule that can have this mass. Assume the elements that make up the molecule are made of light isotopes where the number of protons equals the number of neutrons in the nucleus of each element.

Challenge Problems

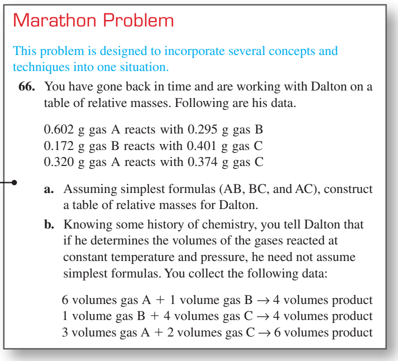
60. Confronted with the box shown in the diagram, you wish to discover something about its internal workings. You have no tools and cannot open the box. You pull on rope B, and it moves rather freely. When you pull on rope A, rope C appears to be pulled slightly into the box. When you pull on rope C,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

a. Based on these observations, construct a model for the interior mechanism of the box.

b. What further experiments could you do to refine your model?

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

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



Marathon Problem

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

66. You have gone back in time and are working with Dalton on a table of relative masses. Following are his data.

0.602 g gas A reacts with 0.295 g gas B
0.172 g gas B reacts with 0.401 g gas C
0.320 g gas A reacts with 0.374 g gas C

a. Assuming simplest formulas (AB, BC, and AC), construct a table of relative masses for Dalton.

b. Knowing some history of chemistry, you tell Dalton that if he determines the volumes of the gases reacted at constant temperature and pressure, he need not assume simplest formulas. You collect the following data:

6 volumes gas A + 1 volume gas B → 4 volumes product
1 volume gas B + 4 volumes gas C → 4 volumes product
3 volumes gas A + 2 volumes gas C → 6 volumes product

About the Authors

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REVIEW

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

Measurement and Calculations in Chemistry

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

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

R.1 Units of Measurement

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

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

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

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

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

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

TABLE R.1 | The Fundamental SI Units

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

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

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

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

TABLE R.3 | Some Examples of Commonly Used Units

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

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

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

Also, since 1 cm^3 = 1 milliliter (mL),

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

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

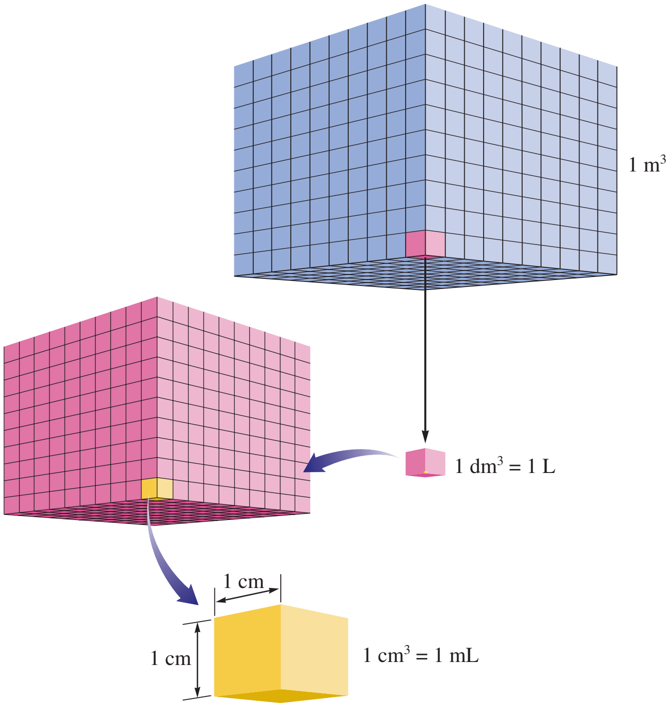
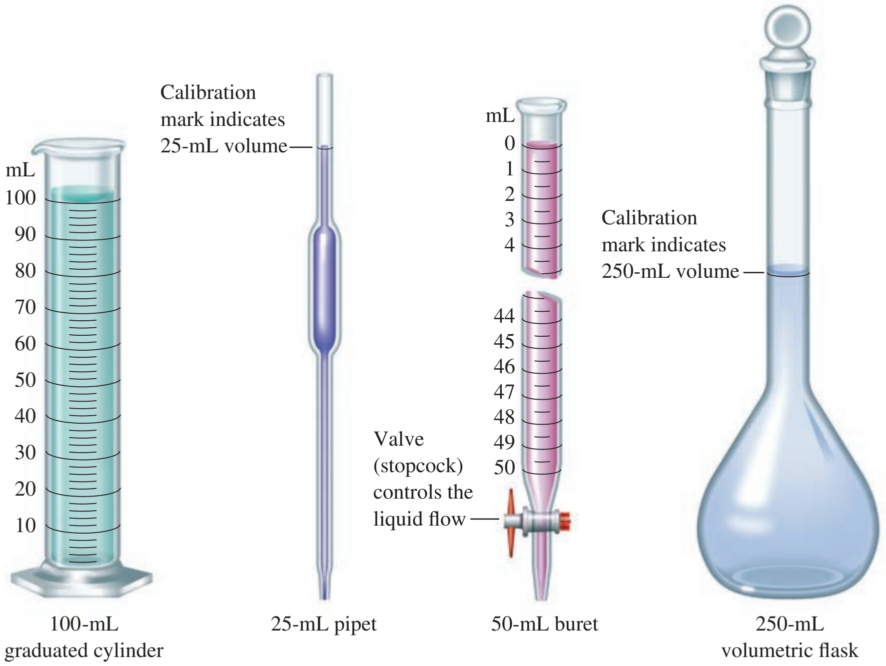


FIGURE **R.1** The largest cube has sides 1 m in length and a volume of 1 m^3. The middle-sized cube has sides 1 dm in length and a volume of 1 dm^3, or 1 L. The smallest cube has sides 1 cm in length and a volume of 1 cm^3, or 1 mL.

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



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

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

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

R.2 Uncertainty in Measurement

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

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

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

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

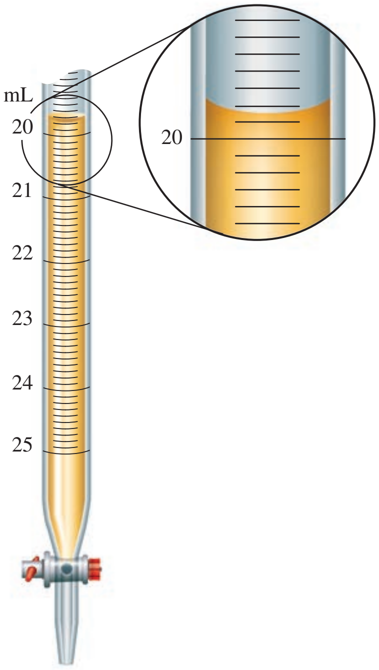


FIGURE **R.3** Measurement of volume using a buret. The volume is read at the bottom of the liquid curve (called the meniscus).

A measurement always has some degree of uncertainty.

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

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

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

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

INTERACTIVE EXAMPLE **R.1**

Uncertainty in Measurement

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

SOLUTION

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

See Question R.21

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

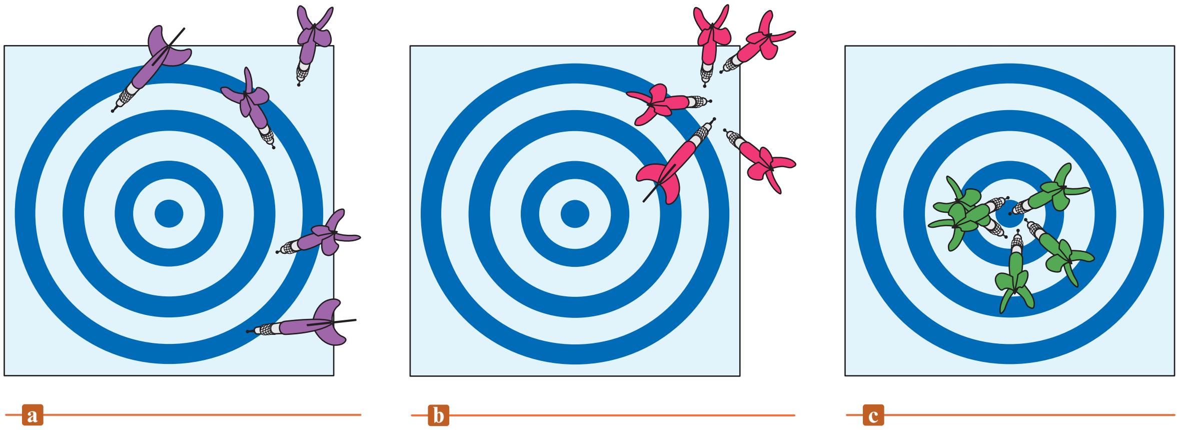


FIGURE **R.4** The results of several dart throws show the difference between precise and accurate. **(a)** Neither accurate nor precise (large random errors). **(b)** Precise but not accurate (small random errors, large systematic error). **(c)** Bull's-eye! Both precise and accurate (small random errors,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

Precision and Accuracy

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

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

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

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

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

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

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

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

EXAMPLE R.2

Precision and Accuracy

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

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

Is the graduated cylinder accurate?

SOLUTION

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

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

See Question R.2

R.3 Significant Figures and Calculations

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

Rules for Counting Significant Figures

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

Leading zeros are never significant figures.

Captive zeros are always significant figures.

Trailing zeros are sometimes significant figures.

(Box continues on the following page)

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

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

Exponential notation is reviewed in Appendix 1.1.

Note that the number 1.00 × 102 above is written in **exponential notation**. This type of notation has at least two advantages: the number of significant figures can be easily indicated, and fewer zeros are needed to write a very large or very small number. For example, the number 0.000060 is much more conveniently represented as 6.0 × 10−5. (The number has two significant figures.)

INTERACTIVE EXAMPLE **R.3**

Significant Figures

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

a. A student's extraction procedure on tea yields 0.0105 g of caffeine.

b. A chemist records a mass of 0.050080 g in an analysis.

c. In an experiment a span of time is determined to be 8.050×10^{-3} s.

SOLUTION

a. The number contains three significant figures. The zeros to the left of the 1 are leading zeros and are not significant, but the remaining zero (a captive zero) is significant.

b. The number contains five significant figures. The leading zeros (to the left of the 5) are not significant. The captive zeros between the 5 and the 8 are significant, and the trailing zero to the right of the 8 is significant because the number contains a decimal point.

c. This number has four significant figures. Both zeros are significant.

See Exercises R.15 through R.18

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

Rules for Significant Figures in Mathematical Operations

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

$$4.56 \times \underset{\substack{\uparrow \\ \text{Limiting term has} \\ \text{two significant} \\ \text{figures}}}{1.4} = 6.38 \xrightarrow{\text{Corrected}} \underset{\substack{\uparrow \\ \text{Two significant} \\ \text{figures}}}{6.4}$$

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

(Box continues on the following page)

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

$$\begin{array}{rl} 12.11 & \\ 18.0 & \leftarrow \text{Limiting term has one decimal place} \\ \underline{1.013} & \\ 31.123 & \xrightarrow{\text{Corrected}} \underset{\substack{\uparrow \\ \text{One decimal place}}}{31.1} \end{array}$$

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

For multiplication and division: significant figures are counted.

For addition and subtraction: decimal places are counted.

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

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

Rules for Rounding

1. In a series of calculations, carry the extra digits through to the final result, *then* round.

2. If the digit to be removed

  a. is less than 5, the preceding digit stays the same. For example, 1.33 rounds to 1.3.

  b. is equal to or greater than 5, the preceding digit is increased by 1. For example, 1.36 rounds to 1.4.

Rule 2 is consistent with the operation of electronic calculators.

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

$$4.3\underset{\substack{\uparrow \\ \text{Look at this number to} \\ \text{round to two significant figures.}}}{4}8$$

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

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

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

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

INTERACTIVE EXAMPLE **R.4**

Significant Figures in Mathematical Operations

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

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

b. $21 - 13.8$

c. As part of a lab assignment to determine the value of the gas constant (R), a student measured the pressure (P), volume (V), and temperature (T) for a sample of gas, where

$$R = \frac{PV}{T}$$

The following values were obtained: $P = 2.560$, $T = 275.15$, and $V = 8.8$. (Gases will be discussed in detail in Chapter 8; we will not be concerned at this time about the units for these quantities.) Calculate R to the correct number of significant figures.

SOLUTION

a. The result is 1.71×10^{-4}, which has three significant figures because the term with the least precision (1.05×10^{-3}) has three significant figures.

b. The result is 7 with no decimal point because the number with the least number of decimal places (21) has none.

When this number is rounded to two significant figures the result is 0.082.

c.

$$R = \frac{PV}{T} = \frac{(2.560)(8.8)}{275.15}$$

The correct procedure for obtaining the final result can be represented as follows:

$$\frac{(2.560)(8.8)}{275.15} = \frac{22.528}{275.15} = 0.0818753$$

$$= 0.082 = 8.2 \times 10^{-2} = R$$

The final result must be rounded to two significant figures because 8.8 (the least precise measurement) has two significant figures. To show the effects of rounding at intermediate steps, we will carry out the calculation as follows:

$$\frac{(2.560)(8.8)}{275.15} = \frac{22.528}{275.15} = \frac{\overset{\substack{\text{Rounded to two}\\\text{significant figures}\\\downarrow}}{23}}{275.15}$$

Now we proceed with the next calculation:

$$\frac{23}{275.15} = 0.0835908$$

Rounded to two significant figures, this result is

$$0.084 = 8.4 \times 10^{-2}$$

Note that intermediate rounding gives a significantly different result than was obtained by rounding only at the end. Again, we must reemphasize that in *your* calculations you should round *only at the end*. However, because rounding is carried out at intermediate steps in this text (to always show the correct number of significant figures), the final answer given in the text may differ slightly from the one you obtain (rounding only at the end).

See Exercises R.23 through R.26

There is a useful lesson to be learned from part c of Example R.4. The student measured the pressure and temperature to greater precision than the volume. A more precise value of R (one with more significant figures) could have been obtained if a more precise measurement of V had been made. As it is, the efforts expended to measure P and T very precisely were wasted. Remember that a series of measurements to obtain some final result should all be done to about the same precision.

R.4 Learning to Solve Problems Systematically

One of the challenges of mastering chemistry is learning to solve the problems that are associated with the concepts of chemistry. In this text, we will first explore the various concepts and then consider problems that test your true understanding of these concepts. It turns out that this process is excellent preparation for "real life" where you will need to think creatively about the situations you encounter in your work and your personal life. Therefore, in this text we will help you learn to approach problems in a thoughtful creative way rather than to rely on rote memorization. A little later in the text when the problems get more complicated, we will consider this process in more detail. However, at this early point in our study of chemistry we will consider some questions that will help you get started being a thoughtful problem solver:

1. What is my goal? Or you might phrase it as: ***Where am I going?***
2. Where am I starting? Or you might phrase it as: ***What do I know?***
3. How do I proceed from where I start to where I want to go? Or you might say: ***How do I get there?***

We will use these ideas as we consider unit conversions in this chapter. Then we will have much more to say about problem solving in Chapter 5, where we will start to consider more complex problems.

R.5 Dimensional Analysis

It is often necessary to convert a given result from one system of units to another. The best way to do this is by a method called the **unit factor method** or, more commonly, **dimensional analysis**. To illustrate the use of this method, we will consider several unit conversions. Some equivalents in the English and metric systems are listed in Table R.4. A more complete list of conversion factors given to more significant figures appears in Appendix 6.

TABLE R.4 | English–Metric Equivalents

Unit	Example
Length	1 m = 1.094 yd 2.54 cm = 1 in
Mass	1 kg = 2.205 lb 453.6 g = 1 lb
Volume	1 L = 1.06 qt 1 ft^3 = 28.32 L

Consider a pin measuring 2.85 cm in length. What is its length in inches? To accomplish this conversion, we must use the equivalence statement

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

If we divide both sides of this equation by 2.54 cm, we get

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

This expression is called a *unit factor.* Since 1 inch and 2.54 cm are exactly equivalent, multiplying any expression by this unit factor will not change its *value.*

The pin has a length of 2.85 cm. Multiplying this length by the appropriate unit factor gives

$$2.85 \cancel{\text{cm}} \times \frac{1 \text{ in}}{2.54 \cancel{\text{cm}}} = \frac{2.85}{2.54} \text{ in} = 1.12 \text{ in}$$

Note that the centimeter units cancel to give inches for the result. This is exactly what we wanted to accomplish. Note also that the result has three significant figures, as required by the number 2.85. Recall that the 1 and 2.54 in the conversion factor are exact numbers by definition.

INTERACTIVE EXAMPLE **R.5**

Unit Conversions I

A pencil is 7.00 in long. What is its length in centimeters?

SOLUTION

Where are we going?

To convert the length of the pencil from inches to centimeters

What do we know?

- The pencil is 7.00 in long.

How do we get there?

Since we want to convert from inches to centimeters, we need the equivalence statement 2.54 cm = 1 in. The correct unit factor in this case is $\frac{2.54\text{ cm}}{1\text{ in}}$:

$$7.00\ \cancel{\text{in}} \times \frac{2.54\text{ cm}}{1\ \cancel{\text{in}}} = (7.00)(2.54)\text{ cm} = 17.8\text{ cm}$$

Here the inch units cancel, leaving centimeters, as requested.

See Exercises R.29 and R.30

Note that two unit factors can be derived from each equivalence statement. For example, from the equivalence statement 2.54 cm = 1 in, the two unit factors are

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

Consider the direction of the required change to select the correct unit factor.

How do you choose which one to use in a given situation? Simply look at the *direction* of the required change. To change from inches to centimeters, the inches must cancel. Thus the factor 2.54 cm/1 in is used. To change from centimeters to inches, centimeters must cancel, and the factor 1 in/2.54 cm is appropriate.

PROBLEM-SOLVING STRATEGY

Converting from One Unit to Another

- To convert from one unit to another, use the equivalence statement that relates the two units.
- Derive the appropriate unit factor by looking at the direction of the required change (to cancel the unwanted units).
- Multiply the quantity to be converted by the unit factor to give the quantity with the desired units.

INTERACTIVE EXAMPLE **R.6**

Unit Conversions II

You want to order a bicycle with a 25.5-in frame, but the sizes in the catalog are given only in centimeters. What size should you order?

SOLUTION

Where are we going?

To convert from inches to centimeters

What do we know?

- The size needed is 25.5 in.

How do we get there?

Since we want to convert from inches to centimeters, we need the equivalence statement 2.54 cm = 1 in. The correct unit factor in this case is $\frac{2.54\ \text{cm}}{1\ \text{in}}$:

$$25.5\ \cancel{\text{in}} \times \frac{2.54\ \text{cm}}{1\ \cancel{\text{in}}} = 64.8\ \text{cm}$$

See Exercises R.29 and R.30

To ensure that the conversion procedure is clear, a multistep problem is considered in Example R.7.

INTERACTIVE EXAMPLE **R.7**

Unit Conversions III

A student has entered a 10.0-km run. How long is the run in miles?

SOLUTION

Where are we going?

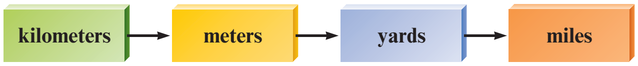
To convert from kilometers to miles

What do we know?

› The run is 10.00 km long.

How do we get there?

This conversion can be accomplished in several different ways. Since we have the equivalence statement 1 m = 1.094 yd, we will proceed by a path that uses this fact. Before we start any calculations, let us consider our strategy. We have kilometers, which we want to change to miles. We can do this by the following route:



To proceed in this way, we need the following equivalence statements:

$$1\ \text{km} = 1000\ \text{m}$$
$$1\ \text{m} = 1.094\ \text{yd}$$
$$1760\ \text{yd} = 1\ \text{mi}$$

To make sure the process is clear, we will proceed step by step:

Kilometers to Meters

$$10.0\ \cancel{\text{km}} \times \frac{1000\ \text{m}}{1\ \cancel{\text{km}}} = 1.00 \times 10^4\ \text{m}$$

Meters to Yards

$$1.00 \times 10^4\ \cancel{\text{m}} \times \frac{1.094\ \text{yd}}{1\ \cancel{\text{m}}} = 1.094 \times 10^4\ \text{yd}$$

Note that we should have only three significant figures in the result. However, since this is an intermediate result, we will carry the extra digit. Remember, round off only the final result.

In the text we round to the correct number of significant figures after each step to show the correct significant figures for each calculation. However, since you use a calculator and combine steps on it, you should round only at the end.

Yards to Miles

$$1.094 \times 10^4\ \cancel{\text{yd}} \times \frac{1\ \text{mi}}{1760\ \cancel{\text{yd}}} = 6.216\ \text{mi}$$

Note in this case that 1 mi equals exactly 1760 yd *by designation.* Thus 1760 is an exact number.

Since the distance was originally given as 10.0 km, the result can have only three significant figures and should be rounded to 6.22 mi. Thus

$$10.0 \text{ km} = 6.22 \text{ mi}$$

Alternatively, we can combine the steps:

$$10.0\ \cancel{\text{km}} \times \frac{1000\ \cancel{\text{m}}}{1\ \cancel{\text{km}}} \times \frac{1.094\ \cancel{\text{yd}}}{1\ \cancel{\text{m}}} \times \frac{1 \text{ mi}}{1760\ \cancel{\text{yd}}} = 6.22 \text{ mi}$$

See Exercises R.29 and R.30

In using dimensional analysis, your verification that everything has been done correctly is that you end up with the correct units. **In doing chemistry problems, you should always include the units for the quantities used.** Always check to see that the units cancel to give the correct units for the final result. This provides a very valuable check, especially for complicated problems.

Study the procedures for unit conversions in the following examples.

INTERACTIVE EXAMPLE **R.8**

Unit Conversions IV

The speed limit on many highways in the United States is 55 mi/h. What number would be posted in kilometers per hour?

SOLUTION

Where are we going?

To convert the speed limit from 55 miles per hour to kilometers per hour

What do we know?

- The speed limit is 55 mi/h.

How do we get there?

We use the following unit factors to make the required conversion:

Result obtained by rounding only at the end of the calculation

$$\frac{55\ \cancel{\text{mi}}}{\text{h}} \times \frac{1760\ \cancel{\text{yd}}}{1\ \cancel{\text{mi}}} \times \frac{1\ \cancel{\text{m}}}{1.094\ \cancel{\text{yd}}} \times \frac{1 \text{ km}}{1000\ \cancel{\text{m}}} = 88 \text{ km/h}$$

Note that all units cancel except the desired kilometers per hour.

See Exercises R.37 through R.39

INTERACTIVE EXAMPLE **R.9**

Unit Conversions V

A Japanese car is advertised as having a gas mileage of 15 km/L. Convert this rating to miles per gallon.

SOLUTION

Where are we going?

To convert gas mileage from 15 kilometers per liter to miles per gallon

What do we know?

- The gas mileage is 15 km/L.

How do we get there?

We use the following unit factors to make the required conversion:

$$\frac{15\ \cancel{\text{km}}}{\cancel{\text{L}}} \times \frac{1000\ \cancel{\text{m}}}{1\ \cancel{\text{km}}} \times \frac{1.094\ \cancel{\text{yd}}}{1\ \cancel{\text{m}}} \times \frac{1\ \text{mi}}{1760\ \cancel{\text{yd}}} \times \frac{1\ \cancel{\text{L}}}{1.06\ \cancel{\text{qt}}} \times \frac{4\ \cancel{\text{qt}}}{1\ \text{gal}} = \overbrace{35\ \text{mi/gal}}^{\substack{\text{Result obtained by}\\ \text{rounding only at the}\\ \text{end of the calculation}}}$$

See Exercise R.40

INTERACTIVE EXAMPLE **R.10**

Unit Conversions VI

The latest model Corvette has an engine with a displacement of 6.20 L. What is the displacement in units of cubic inches?

SOLUTION

Where are we going?

To convert the engine displacement from liters to cubic inches

What do we know?

- The displacement is 6.20 L.

How do we get there?

We use the following unit factors to make the required conversion:

$$6.20\ \text{L} \times \frac{1\ \text{ft}^3}{28.32\ \text{L}} \times \frac{(12\ \text{in})^3}{(1\ \text{ft})^3} = 378\ \text{in}^3$$

Note that the unit factor for conversion of feet to inches must be cubed to accommodate the conversion of ft^3 to in^3.

See Exercise R.44

R.6 Temperature

Three systems for measuring temperature are widely used: the Celsius scale, the Kelvin scale, and the Fahrenheit scale. The first two temperature systems are used in the physical sciences, and the third is used in many of the engineering sciences. Our purpose here is to define the three temperature scales and show how conversions from one scale to another can be performed. Although these conversions can be carried out routinely on most calculators, we will consider the process in some detail here to illustrate methods of problem solving.

The three temperature scales are defined and compared in Fig. R.5. Note that the size of the temperature unit (the *degree*) is the same for the Kelvin and Celsius scales. The fundamental difference between these two temperature scales is their zero points. Conversion between these two scales simply requires an adjustment for the different zero points.

$$\text{Temperature (Kelvin)} = \text{temperature (Celsius)} + 273.15$$

$$T_\text{K} = T_\text{C} + 273.15$$

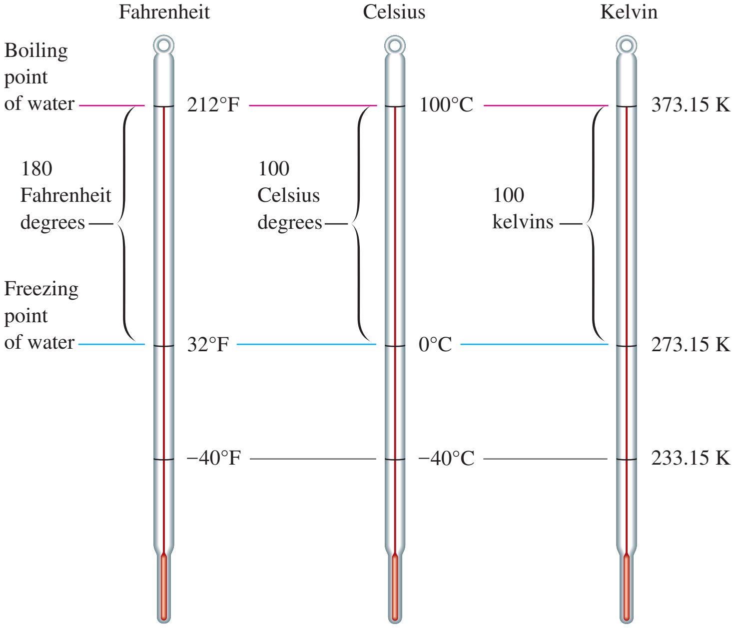


FIGURE R.5 The three major temperature scales.

or

$$\text{Temperature (Celsius)} = \text{temperature (Kelvin)} - 273.15$$

$$T_C = T_K - 273.15$$

For example, to convert 300.00 K to the Celsius scale, we do the following calculation:

$$300.00 - 273.15 = 26.85°\text{C}$$

Note that in expressing temperature in Celsius units, the designation °C is used. The degree symbol is not used when writing temperature in terms of the Kelvin scale. The unit of temperature on this scale is called a *kelvin* and is symbolized by the letter K.

Converting between the Fahrenheit and Celsius scales is somewhat more complicated because both the degree sizes and the zero points are different. Thus we need to consider two adjustments: one for degree size and one for the zero point. First, we must account for the difference in degree size. This can be done by reconsidering Fig. R.5. Notice that since 212°F = 100°C and 32°F = 0°C,

$$212 - 32 = 180 \text{ Fahrenheit degrees} = 100 - 0 = 100 \text{ Celsius degrees}$$

Thus 180° on the Fahrenheit scale is equivalent to 100° on the Celsius scale, and the unit factor is

$$\frac{180°\text{F}}{100°\text{C}} \quad \text{or} \quad \frac{9°\text{F}}{5°\text{C}}$$

or the reciprocal, depending on the direction in which we need to go.

Next, we must consider the different zero points. Since 32°F = 0°C, we obtain the corresponding Celsius temperature by first subtracting 32 from the Fahrenheit temperature to account for the different zero points. Then the unit factor is applied to adjust for the difference in the degree size. This process is summarized by the equation

$$(T_F - 32°\text{F})\frac{5°\text{C}}{9°\text{F}} = T_C \qquad \text{(R.1)}$$

where T_F and T_C represent a given temperature on the Fahrenheit and Celsius scales, respectively. In the opposite conversion, we first correct for degree size and then correct for the different zero point. This process can be summarized in the following general equation:

$$T_F = T_C \times \frac{9°F}{5°C} + 32°F \tag{R.2}$$

Equations (R.1) and (R.2) are really the same equation in different forms. See if you can obtain Equation (R.2) by starting with Equation (R.1) and rearranging.

At this point it is worthwhile to weigh the two alternatives for learning to do temperature conversions: You can simply memorize the equations, or you can take the time to learn the differences between the temperature scales and to understand the processes involved in converting from one scale to another. The latter approach may take a little more effort, but the understanding you gain will stick with you much longer than the memorized formulas. This choice also will apply to many of the other chemical concepts. Try to think things through!

INTERACTIVE EXAMPLE R.11

Temperature Conversions I

Normal body temperature is 98.6°F. Convert this temperature to the Celsius and Kelvin scales.

SOLUTION

Where are we going?

To convert the body temperature from degrees Fahrenheit to degrees Celsius and to kelvins.

What do we know?

- The body temperature is 98.6°F.


Thinkstock/Getty Images

▲ *A nurse taking the temperature of a patient.*

How do we get there?

Rather than simply using the formulas to solve this problem, we will proceed by thinking it through. The situation is diagramed in Fig. R.6. First, we want to convert 98.6°F to the Celsius scale. The number of Fahrenheit degrees between 32.0°F and 98.6°F is 66.6°F. We must convert this difference to Celsius degrees:

$$66.6°F \times \frac{5°C}{9°F} = 37.0°C$$

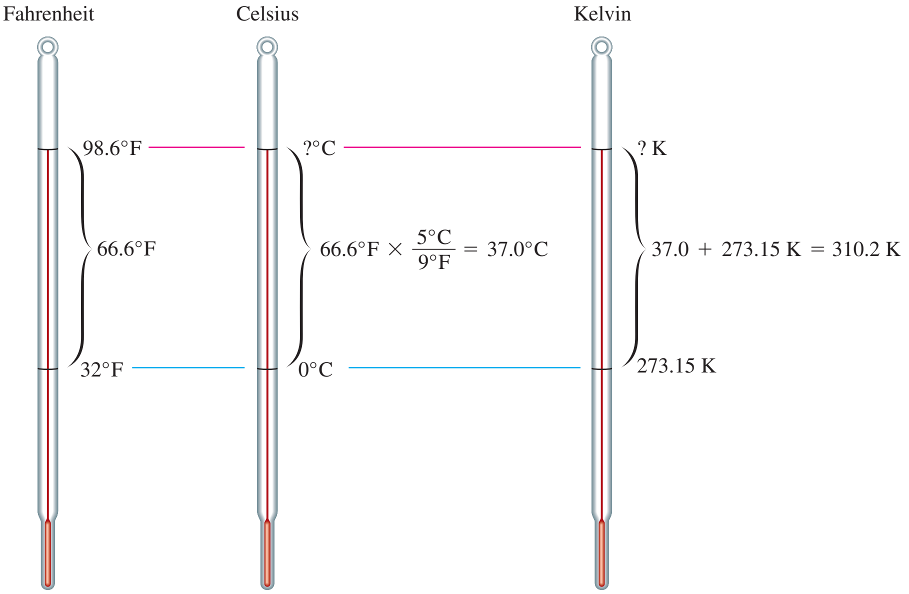


FIGURE **R.6** Normal body temperature on the Fahrenheit, Celsius, and Kelvin scales.

Thus 98.6°F corresponds to 37.0°C.

Now we can convert to the Kelvin scale:

$$T_K = T_C + 273.15 = 37.0 + 273.15 = 310.2\ \text{K}$$

Note that the final answer has only one decimal place (37.0 is limiting).

See Exercises R.45, R.47, and R.48

EXAMPLE **R.12** Temperature Conversions II

One interesting feature of the Celsius and Fahrenheit scales is that −40°C and −40°F represent the same temperature, as shown in Fig. R.5. Verify that this is true.

SOLUTION

Where are we going?

To show that −40°C = −40°F

What do we know?

- The relationship between the Celsius and Fahrenheit scales

How do we get there?

The difference between 32°F and −40°F is 72°F. The difference between 0°C and −40°C is 40°C. The ratio of these is

$$\frac{72°\text{F}}{40°\text{C}} = \frac{8 \times 9°\text{F}}{8 \times 5°\text{C}} = \frac{9°\text{F}}{5°\text{C}}$$

as required. Thus −40°C is equivalent to −40°F.

See Exercise R.49

Since, as shown in Example R.12, −40° on both the Fahrenheit and Celsius scales represents the same temperature, this point can be used as a reference point (like 0°C and 32°F) for a relationship between the two scales:

$$\frac{\text{Number of Fahrenheit degrees}}{\text{Number of Celsius degrees}} = \frac{T_F - (-40)}{T_C - (-40)} = \frac{9°\text{F}}{5°\text{C}}$$

$$\frac{T_F + 40}{T_C + 40} = \frac{9°\text{F}}{5°\text{C}} \qquad (R.3)$$

where T_F and T_C represent the same temperature (but not the same number). This equation can be used to convert Fahrenheit temperatures to Celsius, and vice versa, and may be easier to remember than Equations (R.1) and (R.2).

INTERACTIVE EXAMPLE **R.13** Temperature Conversions III

Liquid nitrogen, which is often used as a coolant for low-temperature experiments, has a boiling point of 77 K. What is this temperature on the Fahrenheit scale?

SOLUTION

Where are we going?

To convert 77 K to the Fahrenheit scale

What do we know?

- The relationship between the Kelvin and Fahrenheit scales

▲ *Liquid nitrogen is so cold that water condenses out of the surrounding air, forming a cloud as the nitrogen is poured.*

How do we get there?

We will first convert 77 K to the Celsius scale:

$$T_C = T_K - 273.15 = 77 - 273.15 = -196°\text{C}$$

To convert to the Fahrenheit scale, we will use Equation (R.3):

$$\frac{T_F + 40}{T_C + 40} = \frac{9°\text{F}}{5°\text{C}}$$

$$\frac{T_F + 40}{-196°\text{C} + 40} = \frac{T_F + 40}{-156°\text{C}} = \frac{9°\text{F}}{5°\text{C}}$$

$$T_F + 40 = \frac{9°\text{F}}{5°\text{C}}(-156°\text{C}) = -281°\text{F}$$

$$T_F = -281°\text{F} - 40 = -321°\text{F}$$

See Exercises R.45, R.47, and R.48

R.7 Density

A property of matter that is often used by chemists as an "identification tag" for a substance is **density**, the mass of substance per unit volume of the substance:

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

The density of a liquid can be determined easily by weighing an accurately known volume of liquid. This procedure is illustrated in Example R.14.

INTERACTIVE EXAMPLE **R.14**

Determining Density

A chemist, trying to identify an unknown liquid, finds that 25.00 cm^3 of the substance has a mass of 19.625 g at 20°C. The following are the names and densities of the compounds that might be the liquid:

Compound	Density in g/cm^3 at 20°C
Chloroform	1.492
Diethyl ether	0.714
Ethanol	0.789
Isopropyl alcohol	0.785
Toluene	0.867

Which of these compounds is the most likely to be the unknown liquid?

SOLUTION

Where are we going?

To calculate the density of the unknown liquid

What do we know?

- The mass of a given volume of the liquid.

There are two ways of indicating units that occur in the denominator. For example, we can write g/cm^3 or g cm^{-3}. Although we will use the former system here, the other system is widely used.

How do we get there?

To identify the unknown substance, we must determine its density. This can be done by using the definition of density:

$$\text{Density} = \frac{\text{mass}}{\text{volume}} = \frac{19.625\ \text{g}}{25.00\ \text{cm}^3} = 0.7850\ \text{g/cm}^3$$

This density corresponds exactly to that of isopropyl alcohol, which therefore most likely is the unknown liquid. However, note that the density of ethanol is also very close. To be sure that the compound is isopropyl alcohol, we should run several more density experiments. (In the modern laboratory, many other types of tests could be done to distinguish between these two liquids.)

See Exercises R.55 and R.56

Besides being a tool for the identification of substances, density has many other uses. For example, the liquid in your car's lead storage battery (a solution of sulfuric acid) changes density because the sulfuric acid is consumed as the battery discharges. In a fully charged battery, the density of the solution is about 1.30 g/cm^3. If the density falls below 1.20 g/cm^3, the battery will have to be recharged. Density measurement is also used to determine the amount of antifreeze, and thus the level of protection against freezing, in the cooling system of a car.

The densities of various common substances are given in Table R.5.

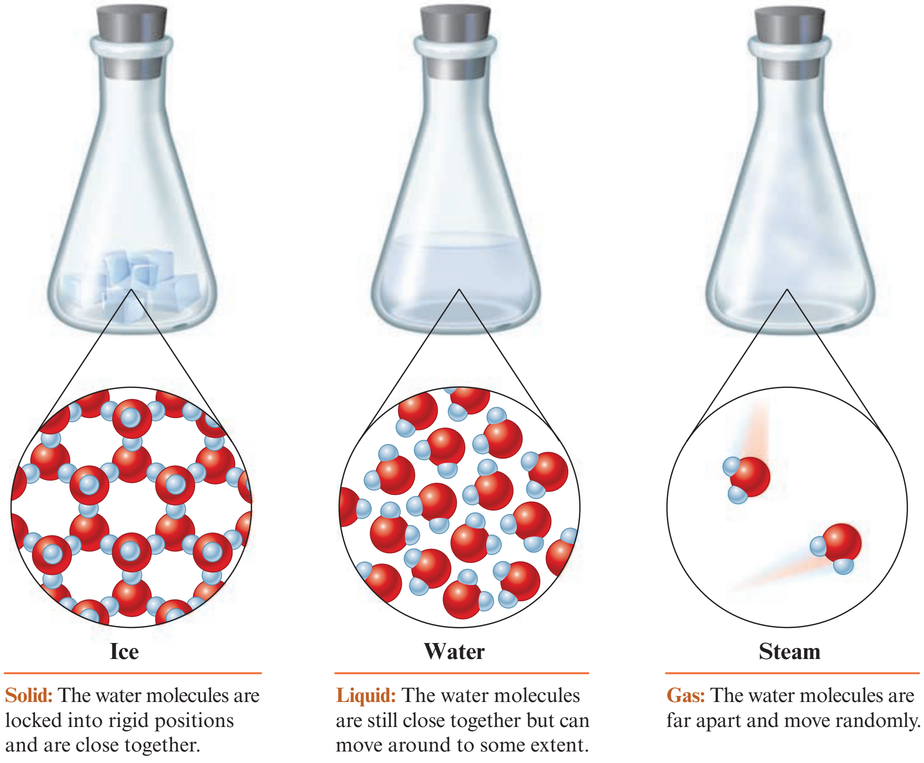
R.8 Classification of Matter

Before we can hope to understand the changes we see going on around us—the growth of plants, the rusting of steel, the aging of people, the acidification of rain—we must find out how matter is organized. **Matter**, best defined as anything occupying space

TABLE R.5 | **Densities of Various Common Substances* at 20°C**

Substance	Physical State	Density (g/cm^3)
Oxygen	Gas	0.00133
Hydrogen	Gas	0.000084
Ethanol	Liquid	0.789
Benzene	Liquid	0.880
Water	Liquid	0.9982
Magnesium	Solid	1.74
Salt (sodium chloride)	Solid	2.16
Aluminum	Solid	2.70
Iron	Solid	7.87
Copper	Solid	8.96
Silver	Solid	10.5
Lead	Solid	11.34
Mercury	Liquid	13.6
Gold	Solid	19.32

*At 1 atmosphere pressure.



Solid: The water molecules are locked into rigid positions and are close together.

Liquid: The water molecules are still close together but can move around to some extent.

Gas: The water molecules are far apart and move randomly.

FIGURE R.7 The three states of water (where red spheres represent oxygen atoms and blue spheres represent hydrogen atoms).

and having mass, is the material of the universe. Matter is complex and has many levels of organization. In this section we will introduce basic ideas about the structure of matter and its behavior.

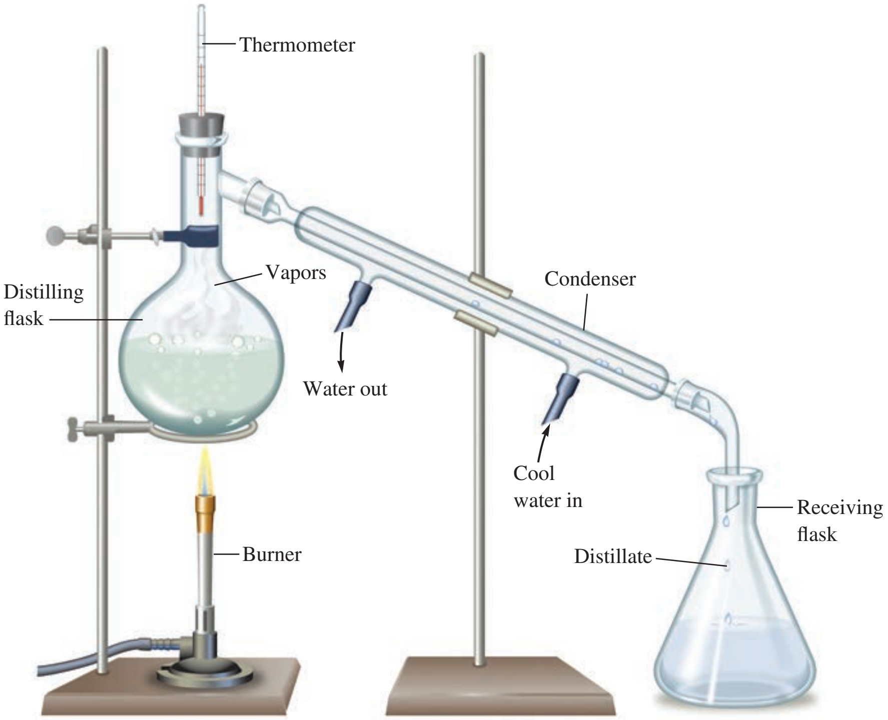
We will start by considering the definitions of the fundamental properties of matter. Matter exists in three **states**: solid, liquid, and gas. A *solid* is rigid; it has a fixed volume and shape. A *liquid* has a definite volume but no specific shape; it assumes the shape of its container. A *gas* has no fixed volume or shape; it takes on the shape and volume of its container. In contrast to liquids and solids, which are only slightly compressible, gases are highly compressible; it is relatively easy to decrease the volume of a gas. Molecular-level pictures of the three states of water are given in Fig. R.7. The different properties of ice, liquid water, and steam are determined by the different arrangements of the molecules in these substances. Table R.5 gives the states of some common substances at 20°C and 1 atmosphere pressure.

Most of the matter around us consists of **mixtures** of pure substances. Wood, gasoline, wine, soil, and air all are mixtures. The main characteristic of a mixture is that it has *variable composition.* For example, wood is a mixture of many substances, the proportions of which vary depending on the type of wood and where it grows. Mixtures can be classified as **homogeneous** (having visibly indistinguishable parts) or **heterogeneous** (having visibly distinguishable parts).

A homogeneous mixture is called a **solution**. Air is a solution consisting of a mixture of gases. Wine is a complex liquid solution. Brass is a solid solution of copper and zinc. Sand in water and iced tea with ice cubes are examples of heterogeneous mixtures. Heterogeneous mixtures usually can be separated into two or more homogeneous mixtures or pure substances (for example, the ice cubes can be separated from the tea).

Mixtures can be separated into pure substances by physical methods. A **pure substance** is one with constant composition. Water is a good illustration of these ideas. As we will discuss in detail later, pure water is composed solely of H_2O molecules, but the water found in nature (groundwater or the water in a lake or ocean) is really a mixture.

FIGURE **R.8** Simple laboratory distillation apparatus. Cool water circulates through the outer portion of the condenser, causing vapors from the distilling flask to condense into a liquid. The nonvolatile component of the mixture remains in the distilling flask.



Seawater, for example, contains large amounts of dissolved minerals. Boiling seawater produces steam, which can be condensed to pure water, leaving the minerals behind as solids. The dissolved minerals in seawater also can be separated out by freezing the mixture, since pure water freezes out. The processes of boiling and freezing are **physical changes**. When water freezes or boils, it changes its state but remains water; it is still composed of H_2O molecules. A physical change is a change in the form of a substance, not in its chemical composition. A physical change can be used to separate a mixture into pure compounds, but it will not break compounds into elements.

One of the most important methods for separating the components of a mixture is **distillation**, a process that depends on differences in the volatility (how readily substances become gases) of the components. In simple distillation, a mixture is heated in a device such as that shown in Fig. R.8. The most volatile component vaporizes at the lowest temperature, and the vapor passes through a cooled tube (a condenser), where it condenses back into its liquid state.

The term *volatile* refers to the ease with which a substance can be changed to its vapor.

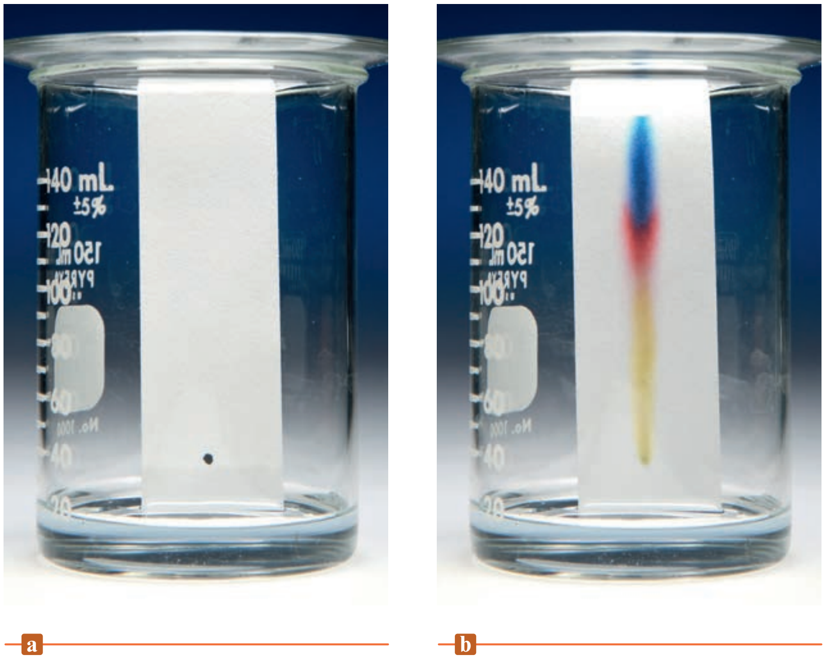
The simple, one-stage distillation apparatus shown in Fig. R.8 works very well when only one component of the mixture is volatile. For example, a mixture of water and sand is easily separated by boiling off the water. Water containing dissolved minerals behaves in much the same way. As the water is boiled off, the minerals remain behind as nonvolatile solids. Simple distillation of seawater using the sun as the heat source is an excellent way to desalinate (remove the minerals from) seawater.

However, when a mixture contains several volatile components, the one-step distillation does not give a pure substance in the receiving flask, and more elaborate methods are required.

Another method of separation is simple **filtration**, which is used when a mixture consists of a solid and a liquid. The mixture is poured onto a mesh, such as filter paper, which passes the liquid and leaves the solid behind.

A third method of separation is **chromatography**. Chromatography is the general name applied to a series of methods that use a system with two *phases* (states) of matter: a mobile phase and a stationary phase. The *stationary phase* is a solid, and the *mobile phase* is either a liquid or a gas. The separation process occurs because the

FIGURE R.9 Paper chromatography of ink. (**a**) A dot of the mixture to be separated is placed at one end of a sheet of porous paper. (**b**) The paper acts as a wick to draw up the liquid.



Photos © Charles D. Winters

components of the mixture have different affinities for the two phases and thus move through the system at different rates. A component with a high affinity for the mobile phase moves relatively quickly through the chromatographic system, whereas one with a high affinity for the solid phase moves more slowly.

One simple type of chromatography, **paper chromatography**, uses a strip of porous paper, such as filter paper, for the stationary phase. A drop of the mixture to be separated is placed on the paper, which is then dipped into a liquid (the mobile phase) that travels up the paper as though it were a wick (Fig. R.9). This method of separating a mixture is often used by biochemists, who study the chemistry of living systems.

It should be noted that when a mixture is separated, the absolute purity of the separated components is an ideal. Because water, for example, inevitably comes into contact with other materials when it is synthesized or separated from a mixture, it is never absolutely pure. With great care, however, substances can be obtained in very nearly pure form.

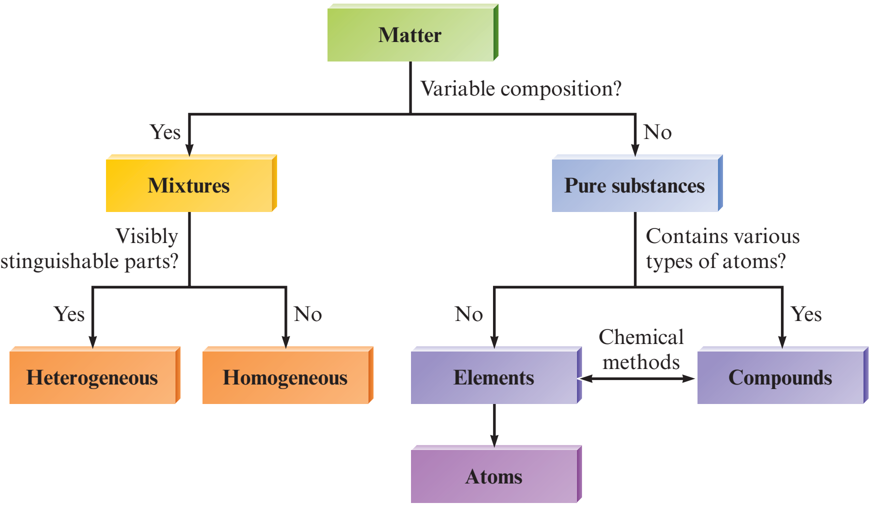
Pure substances are either compounds (combinations of elements) or free elements. A **compound** is a substance with *constant composition* that can be broken down into elements by chemical processes. An example of a chemical process is the electrolysis of water, in which an electric current is passed through water to break it down into the free elements hydrogen and oxygen. This process produces a chemical change because the water molecules have been broken down. The water is gone, and in its place we have the free elements hydrogen and oxygen. A **chemical change** is one in which a given substance becomes a new substance or substances with different properties and different composition. **Elements** are substances that cannot be decomposed into simpler substances by chemical or physical means.

Kristen Brochmann/Fundamental Photographs

▲ *The element mercury (top left) combines with the element iodine (top right) to form the compound mercuric iodide (bottom). This is an example of a chemical change.*

We have seen that the matter around us has various levels of organization. The most fundamental substances we have discussed so far are elements. As we will see in later chapters, elements also have structure: They are composed of atoms, which in turn are composed of nuclei and electrons. Even the nucleus has structure: It is composed of protons and neutrons. And even these can be broken down further, into elementary particles called *quarks.* However, we need not concern ourselves with such details at this point. Fig. R.10 summarizes our discussion of the organization of matter.

FIGURE **R.10** The organization of matter.



R.9 Energy

Although the concept of energy is quite familiar, energy itself is rather difficult to define precisely. For our purposes we will define **energy** as the *capacity to do work or to produce heat*.

The total energy content of the universe is constant.

One of the most important characteristics of energy is that it is conserved. The **law of conservation of energy** states that *energy can be converted from one form to another but can be neither created nor destroyed*. That is, the energy of the universe is constant. Energy can be classified as either potential or kinetic energy. **Potential energy** is energy due to position or composition. For example, water behind a dam has potential energy that can be converted to work when the water flows down through turbines, thereby creating electricity. Attractive and repulsive forces also lead to potential energy. The energy released when gasoline is burned results from differences in attractive forces (bond energies) between the nuclei and electrons in the reactants and products. The **kinetic energy** of an object is energy due to the motion of the object and depends on the mass of the object m and its velocity v: $\text{KE} = \frac{1}{2}mv^2$.

Energy can be converted from one form to another. For example, consider the two balls in Fig. R.11(a). Ball A, because of its higher position initially, has more potential energy than ball B. When A is released, it moves down the hill and strikes B. Eventually, the arrangement shown in Fig. R.11(b) is achieved. What has happened in going from the initial to the final arrangement? The potential energy of A has decreased, but since energy is conserved, all the energy lost by A must be accounted for. How is this energy distributed?

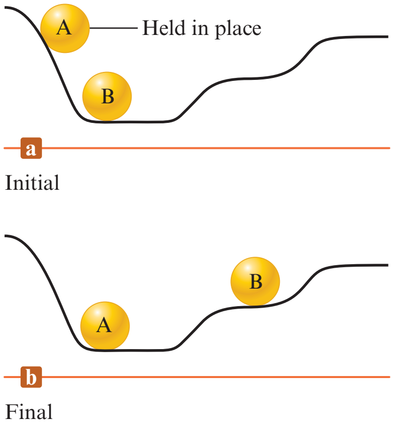


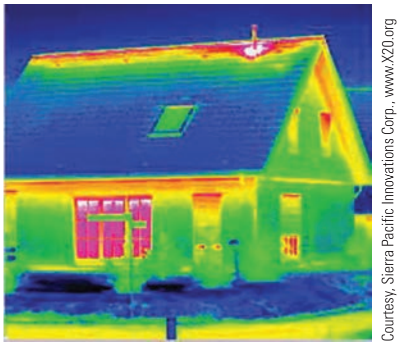
FIGURE **R.11** (**a**) In the initial positions, ball A has a higher potential energy than ball B. (**b**) After A has rolled down the hill, the potential energy lost by A has been converted to random motions of the components of the hill (frictional heating) and to the increase in the potential energy of B.

Initially, the potential energy of A is changed to kinetic energy as the ball rolls down the hill. Part of this kinetic energy is then transferred to B, causing it to be raised to a higher final position. Thus the potential energy of B has been increased. However, since the final position of B is lower than the original position of A, some of the energy is still unaccounted for. Both balls in their final positions are at rest, so the missing energy cannot be due to their motions. What has happened to the remaining energy?

The answer lies in the interaction between the hill's surface and the ball. As ball A rolls down the hill, some of its kinetic energy is transferred to the surface of the hill as heat. This transfer of energy is called *frictional heating*. The temperature of the hill increases very slightly as the ball rolls down.

Heat involves a transfer of energy.

Before we proceed further, it is important to recognize that heat and temperature are decidedly different. As we will see in detail in Chapter 8, *temperature* is a property that reflects the random motions of the particles in a particular substance. **Heat**, on the other hand, involves the *transfer* of energy between two objects due to a temperature difference. Heat is not a substance contained by an object, although we often talk of heat as if this were true.


Courtesy, Sierra Pacific Innovations Corp., www.X20.org

▲ *This infrared photo of a house shows where energy leaks occur. The more red the color, the more energy (heat) is leaving the house.*

Note that in going from the initial to the final arrangements in Fig. R.11, ball B gains potential energy because work was done by ball A on B. **Work** is defined as force acting over a distance. Work is required to raise B from its original position to its final one. Part of the original energy stored as potential energy in A has been transferred through work to B, thereby increasing B's potential energy. Thus there are two ways to transfer energy: through work and through heat.

In rolling to the bottom of the hill as shown as in Fig. R.11, ball A will always lose the same amount of potential energy. However, the way that this energy transfer is divided between work and heat depends on the specific conditions—the **pathway**. For example, the surface of the hill might be so rough that the energy of A is expended completely through frictional heating; A is moving so slowly when it hits B that it cannot move B to the next level. In this case, no work is done. Regardless of the condition of the hill's surface, the *total energy* transferred will be constant. However, the amounts of heat and work will differ. Energy change is independent of the pathway; however, work and heat are both dependent on the pathway.

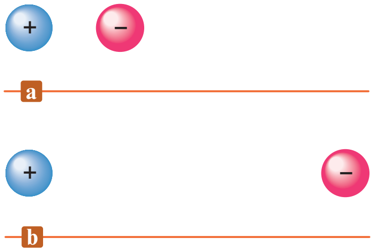


FIGURE **R.12** Two positions of opposite charges.

Another type of PE important in chemical processes is that due to the attraction of opposite charges. For example, consider two relative positions of a positively charged proton and a negatively charged electron, shown in Fig. R.12. The potential energies of the two cases are different. Does the situation shown in Fig. R.12(a) have a higher or lower PE than that shown in Fig. R.12(b)? We can answer this question by devising an imaginary tiny machine, as illustrated in Fig. R.13. Work is done by this machine as the negative charge moves toward the positive charge. Which situation [Fig. R.12(a) or Fig. R.12(b)] would produce the most work (the greatest lifting of the weight)? We can see that the situation in Fig. R.12(b), with the charges furthest apart, can do the most work because the negative charge moves the greatest distance and thus lifts the weight the greatest distance. Thus, because Fig. R.12(b) has more stored energy than Fig. R.12(a), Fig. R.12(b) has a higher PE. In general, the farther apart two opposite charges are, the greater the stored energy and the greater the PE.

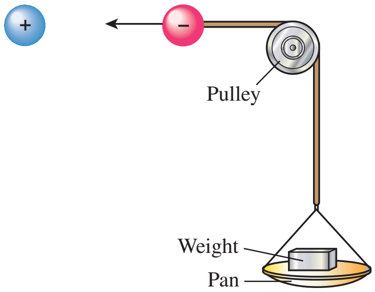


FIGURE **R.13** Work is done as the charges move together and the weight is lifted.

The significance of charged-based PE in chemistry has to do with the energies of electrons in atoms and the energies of chemical reactions, as we will see in Chapters 4 and 7, respectively.

The fundamental unit of energy in science is the joule (pronounced jewel). A **joule** is defined as a kilogram meter squared per second squared:

$$1 \text{ joule} = \frac{\text{kg}\cdot\text{m}^2}{\text{s}^2}.$$

Often the kilojoule (kJ) is used to describe energies.

$$1 \text{ kJ} = 10^3 \text{ J}$$

Chemical Energy

The ideas we have just illustrated using mechanical examples also apply to chemical systems. The combustion of methane (CH_4), for example, is used to heat many homes in the United States by reacting it with the oxygen in the air. As you recall from your previous chemistry course, we can represent this reaction as follows:

$$CH_4(g) + 2O_2(g) \longrightarrow CO_2(g) + 2H_2O(g) + \text{energy (heat)}$$

To discuss this reaction, we divide the universe into two parts: the system and the surroundings. The **system** is the part of the universe on which we wish to focus attention; the **surroundings** include everything else in the universe. In this case we define the system as the reactants and products of the reaction. The surroundings consist of the reaction container (a furnace, for example), the room, and anything else other than the reactants and products.

When a reaction results in the evolution of heat, it is said to be **exothermic** (*exo-* is a prefix meaning "out of"); that is, energy flows *out of the system*. For example, in the combustion of methane, energy flows out of the system as heat. Reactions that absorb

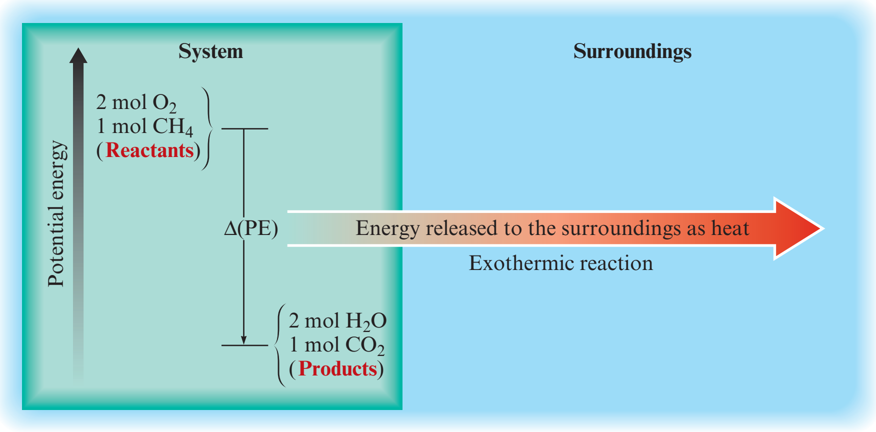


FIGURE **R.14** The combustion of methane releases the quantity of energy Δ(PE) to the surroundings via heat flow. This is an exothermic process.

energy from the surroundings are said to be **endothermic**. When the heat flow is *into a system*, the process is endothermic. For example, the formation of nitric oxide from nitrogen and oxygen is endothermic:

$$N_2(g) + O_2(g) + \text{energy (heat)} \longrightarrow 2NO(g)$$

Where does the energy, released as heat, come from in an exothermic reaction? The answer lies in the difference in potential energies between the products and the reactants. Which has lower potential energy, the reactants or the products? We know that total energy is conserved and that energy flows from the system into the surroundings in an exothermic reaction. This means that *the energy gained by the surroundings must be equal to the energy lost by the system*. In the combustion of methane, the energy content of the system decreases, which means that the product molecules possess less potential energy than do the reactant molecules. The heat flow into the surroundings results from a lowering of the potential energy of the reaction system. This always holds true. *In any exothermic reaction, some of the potential energy is being converted to thermal energy (random kinetic energy) via heat*.

The energy diagram for the combustion of methane is shown in Fig. R.14, where Δ(PE) represents the *change* in potential energy that occurs in the chemical reaction with the bonds of the reactants. In other words, this quantity represents the difference between the energy required to break up the reactant molecules and the energy released when reactant molecules are formed. In an exothermic process, the forces that hold the product molecules together are stronger (on average) than those of the reactants. That is, more energy is released by forming *the new bonds in* the products than is consumed in breaking up the reactants. The net result is that the quantity of energy Δ(PE) is transferred to the surroundings through heat.

For an endothermic reaction, the situation is reversed, as shown in Fig. R.15. Energy that flows into the system as heat is used to increase the potential energy of the system.

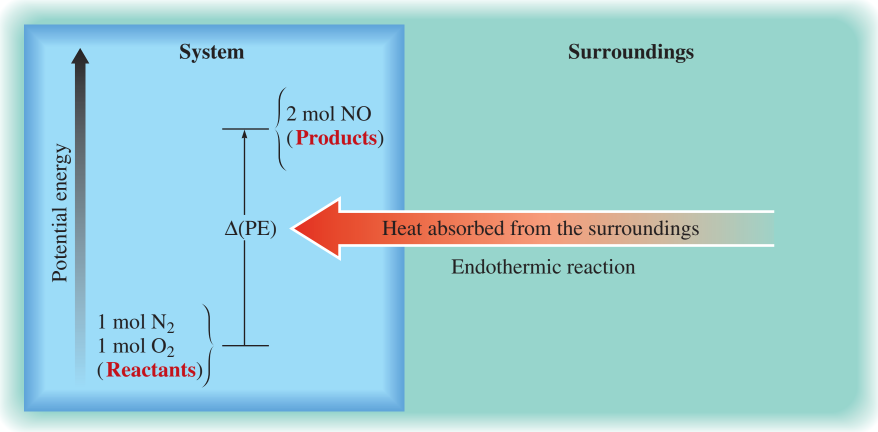


FIGURE **R.15** The energy diagram for the reaction of nitrogen and oxygen to form nitric oxide. This is an endothermic process: Heat [equal in magnitude to Δ(PE)] flows into the system from the surroundings.

In this case the products have higher potential energy (weaker forces between the atoms on average) than the reactants.

In the previous paragraphs we have introduced the term ΔE, which represents the change in energy that occurs in a given process. It is important to understand the properties of ΔE. The number represented by ΔE describes the magnitude of the energy transferred. The sign of ΔE tells which direction the energy is moved. A positive sign means that the energy of the system increases. Thus in this case the sign of ΔE is positive. Energy is transferred into the system. On the other hand, a negative sign for ΔE means that energy leaves the system. The energy of the system is decreased.

We will have much more to say about energy in Chapter 7.

R.10 The Mole

Have you ever wondered how many water molecules are in a glass of water (H_2O)? How can we find out? The molecules are so tiny we can't see them with the naked eye, so we can't count them directly. As we will see in Section 1.1, recent technology, such as the scanning tunneling microscope, allows us to "see" individual atoms and molecules. However, this technology is very recent. How did we count atoms and molecules before this? The answer is we count atoms and molecules by weighing samples containing them.

How do we count by weighing? Consider a sample of jelly beans. What do we need to know about these jelly beans to count them by weighing samples of them? The answer is we need to know the average mass of a jelly bean. For example, if the average mass is 5.0 g, then a 500 g sample would contain 100 jelly beans.

We count atoms and molecules by weighing in the same way we count jelly beans. If we know the average mass of a water molecule, we can determine the number of water molecules in a glass of water by weighing the sample of water in the glass.

We won't detail here how we determine the mass of a water molecule (we will do that later). At this point we ask you to trust us that we have determined the average masses of all the atoms and molecules and can use these masses to count atoms and molecules by weighing samples of them.

When we count the water molecules in a glass of water (by weighing the water) we find that the number present is unimaginably large. The number of molecules in a glass of water is much larger than the age of the earth in seconds (4.32×10^{17} seconds) and larger than the number of milliliters of water in the earth's oceans (1.3×10^{24} mL). Therefore, because of the huge numbers of atoms and molecules in normal-sized samples of matter, we need to invent a unit for describing the number of atoms and molecules present. This unit has to be very large to be convenient. The unit of one dozen (12) works fine for eggs, but it wouldn't help much for describing numbers of atoms or molecules. The unit we have chosen is called the **mole**, which for our purposes we will define as 6.022×10^{23}. Thus, a glass of water that contains 9.0×10^{24} water molecules contains

$$9.0 \times 10^{24}\ \text{H}_2\text{O molecules} \times \frac{1\ \text{mol}}{6.022 \times 10^{23}\ \text{molecules}} = 15\ \text{mol H}_2\text{O molecules.}$$

Note that 15 moles is much more convenient than 9.0×10^{24} water molecules.

We will discuss the mole and its use in chemistry in much more detail in Chapter 5. For the present we simply need to know that the mole is a unit for counting atoms and molecules. For example, when energy terms are given, they are usually given per mole. Examples are the bond energies we will consider in Chapter 3. The bond energy for the hydrogen molecule (H_2) is 432 kJ/mol (where mol is the abbreviation for mole). This means that 432 kJ of energy is required to break one mole (6.022×10^{23}) of H—H bonds. Just remember that when you encounter "per mole" on a unit, it means that 6.022×10^{23} events are represented by the quantity in question.

CHEMICAL CONNECTIONS R.1

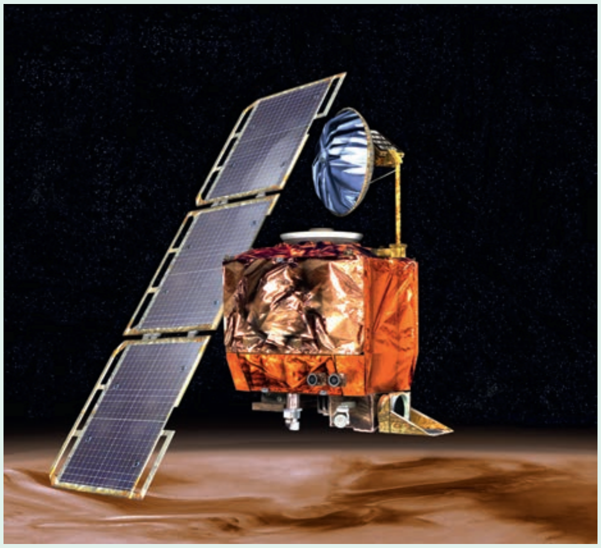
Critical Units!

How important are conversions from one unit to another? If you ask the National Aeronautics and Space Administration (NASA), very important! In 1999, NASA lost a $125 million Mars Climate Orbiter because of a failure to convert from English to metric units.

The problem arose because two teams working on the Mars mission were using different sets of units. NASA's scientists at the Jet Propulsion Laboratory in Pasadena, California, assumed that the thrust data for the rockets on the Orbiter they received from Lockheed Martin Astronautics in Denver, which built the spacecraft, were in metric units. In reality, the units were English. As a result the Orbiter dipped 100 kilometers lower into the Mars atmosphere than planned, and the friction from the atmosphere caused the craft to burn up.

NASA's mistake refueled the controversy over whether Congress should require the United States to switch to the metric system. About 95% of the world now uses the metric system, and the United States is slowly switching from English to metric. For example, the automobile industry has adopted metric fasteners, and we buy our soda in two-liter bottles.

Units can be very important. In fact, they can mean the difference between life and death on some occasions. In 1983, for example, a Canadian jetliner almost ran out of fuel when someone pumped 22,300 pounds of fuel into the aircraft instead of 22,300 kilograms. Remember to watch your units!



NASA/JPL

Artist's conception of the lost Mars Climate Orbiter.

For Review

Key Terms
measurement
number
unit

Section R.1
SI system
mass
weight

Section R.2
uncertainty
significant figures
accuracy
precision
random error
systematic error

Section R.3
exponential notation

Quantitative Observations Are Called Measurements.
- Consist of a number and a unit
- Involve some uncertainty
- Uncertainty is indicated by using significant figures
  - Rules to determine significant figures
  - Calculations using significant figures
- Preferred system is SI

Solving Problems Systematically
Use these questions:
- ***Where am I going?***
- ***What do I know?***
- ***How do I get there?***

Section R.5
unit factor method
dimensional analysis

Section R.7
density

Section R.8
matter
states (of matter)
mixture
homogeneous mixture
heterogeneous mixture
solution
pure substance
physical change
distillation
filtration
chromatography
paper chromatography
compound
chemical change
element

Section R.9
energy
law of conservation of energy
potential energy
kinetic energy
heat
work
pathway
joule
system
surroundings
exothermic
endothermic

Section R.10
mole

Temperature Conversions

- $T_K = T_C + 273$
- $T_C = (T_F - 32°\text{F})\left(\dfrac{5°\text{C}}{9°\text{F}}\right)$
- $T_F = T_C\left(\dfrac{9°\text{F}}{5°\text{C}}\right) + 32°\text{F}$

Density

- $\text{Density} = \dfrac{\text{mass}}{\text{volume}}$

Matter Can Exist in Three States:

- Solid
- Liquid
- Gas

Mixtures Can Be Separated by Methods Involving Only Physical Changes:

- Distillation
- Filtration
- Chromatography

Compounds Can Be Decomposed to Elements Only Through Chemical Changes.

Energy

- The ability to do work or release heat
  - Kinetic—energy of motion
  - Potential—stored energy of position
  - Joule—SI unit

Mole

- 6.022×10^{23} units

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

Questions

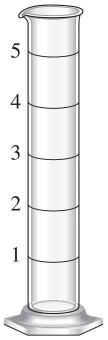
1. What is the difference between random error and systematic error?
2. To determine the volume of a cube, a student measured one of the dimensions of the cube several times. If the true dimension of the cube is 10.62 cm, give an example of four sets of measurements that would illustrate the following.
   a. imprecise and inaccurate data
   b. precise but inaccurate data
   c. precise and accurate data

   Give a possible explanation as to why data can be imprecise or inaccurate. What is wrong with saying a set of measurements is imprecise but accurate?
3. A student performed an analysis of a sample for its calcium content and got the following results:

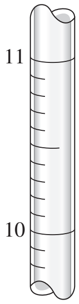
   14.92%   14.91%   14.88%   14.91%

   The actual amount of calcium in the sample is 15.70%. What conclusions can you draw about the accuracy and precision of these results?

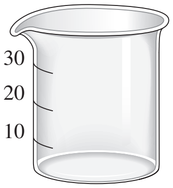
4. For each of the following pieces of glassware, provide a sample measurement and discuss the number of significant figures and uncertainty.



a.



b.



c.

5. What are significant figures? Show how to indicate the number one thousand to 1 significant figure, 2 significant figures, 3 significant figures, and 4 significant figures. Why is the answer, to the correct number of significant figures, not 1.0 for the following calculation?

$$\frac{1.5 - 1.0}{0.50} =$$

6. Compare and contrast the multiplication/division significant figure rule to the significant figure rule applied for addition/subtraction mathematical operations.

7. A rule of thumb in designing experiments is to avoid using a result that is the small difference between two large measured quantities. In terms of uncertainties in measurement, why is this good advice?

8. Explain how density can be used as a conversion factor to convert the volume of an object to the mass of the object, and vice versa.

9. When the temperature in degrees Fahrenheit (T_F) is plotted versus the temperature in degrees Celsius (T_C), a straight-line plot results. A straight-line plot also results when T_C is plotted versus T_K (the temperature in kelvins). Reference Appendix A1.3 and determine the slope and y-intercept of each of these two plots.

10. On which temperature scale (°F, °C, or K) does 1 degree represent the smallest change in temperature?

11. Give four examples illustrating each of the following terms.

- **a.** homogeneous mixture
- **b.** heterogeneous mixture
- **c.** compound
- **d.** element
- **e.** physical change
- **f.** chemical change

12. Use molecular-level (microscopic) drawings for each of the following.

- **a.** Show the differences between a gaseous mixture that is a homogeneous mixture of two different compounds, and a gaseous mixture that is a homogeneous mixture of a compound and an element.
- **b.** Show the differences among a gaseous element, a liquid element, and a solid element.

13. What is the law of conservation of energy? Differentiate between kinetic energy and potential energy.

14. Explain the concept of counting by weighing. What is a mole and why do we use the mole unit?

Exercises

In this section, similar exercises are paired.

Significant Figures and Unit Conversions

15. Which of the following are exact numbers?

- **a.** There are *100* cm in 1 m.
- **b.** One meter equals *1.094* yards.
- **c.** We can use the equation

$$^\circ F = \frac{9}{5}{}^\circ C + 32$$

to convert from Celsius to Fahrenheit temperature. Are the numbers $\frac{9}{5}$ and *32* exact or inexact?

- **d.** $\pi = 3.1415927$.

16. Indicate the number of significant figures in each of the following:

- **a.** This book contains more than 1000 pages.
- **b.** A mile is about 5300 ft.
- **c.** A liter is equivalent to 1.059 qt.
- **d.** The population of the United States in 2015 is 3.1×10^2 million.
- **e.** A kilogram is 1000 g.
- **f.** The Boeing 747 cruises at around 600 mi/h.

17. How many significant figures are there in each of the following values?

- **a.** 6.07×10^{-15}
- **b.** 0.003840
- **c.** 17.00
- **d.** 8×10^8
- **e.** 463.8052
- **f.** 300
- **g.** 301
- **h.** 300.

18. How many significant figures are in each of the following?

- **a.** 100
- **b.** 1.0×10^2
- **c.** 1.00×10^3
- **d.** 100.
- **e.** 0.0048
- **f.** 0.00480
- **g.** 4.80×10^{-3}
- **h.** 4.800×10^{-3}

19. Round off each of the following numbers to the indicated number of significant digits and write the answer in standard scientific notation.

- **a.** 0.00034159 to three digits
- **b.** 103.351×10^2 to four digits
- **c.** 17.9915 to five digits
- **d.** 3.365×10^5 to three digits

20. Use exponential notation to express the number 385,500 to

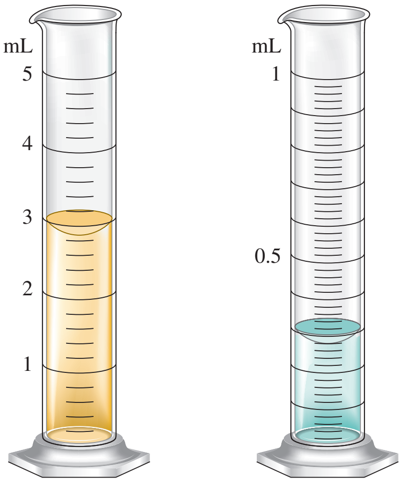
a. one significant figure.

b. two significant figures.

c. three significant figures.

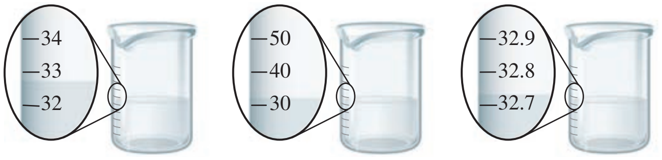
d. five significant figures.

21. You have water in each graduated cylinder shown:



You then add both samples to a beaker. How would you write the number describing the total volume? What limits the precision of this number?

22. The beakers shown below have different precisions.



a. Label the amount of water in each of the three beakers to the correct number of significant figures.

b. Is it possible for each of the three beakers to contain the exact same amount of water? If no, why not? If yes, did you report the volumes as the same in part a? Explain.

c. Suppose you pour the water from these three beakers into one container. What should be the volume in the container reported to the correct number of significant figures?

23. Evaluate each of the following and write the answer to the appropriate number of significant figures.

a. $212.2 + 26.7 + 402.09$

b. $1.0028 + 0.221 + 0.10337$

c. $52.331 + 26.01 - 0.9981$

d. $2.01 \times 10^2 + 3.014 \times 10^3$

e. $7.255 - 6.8350$

24. Perform the following mathematical operations, and express each result to the correct number of significant figures.

a. $\dfrac{0.102 \times 0.0821 \times 273}{1.01}$

b. $0.14 \times 6.022 \times 10^{23}$

c. $4.0 \times 10^4 \times 5.021 \times 10^{-3} \times 7.34993 \times 10^2$

d. $\dfrac{2.00 \times 10^6}{3.00 \times 10^{-7}}$

25. Perform the following mathematical operations and express the result to the correct number of significant figures.

a. $\dfrac{2.526}{3.1} + \dfrac{0.470}{0.623} + \dfrac{80.705}{0.4326}$

b. $(6.404 \times 2.91)/(18.7 - 17.1)$

c. $6.071 \times 10^{-5} - 8.2 \times 10^{-6} - 0.521 \times 10^{-4}$

d. $(3.8 \times 10^{-12} + 4.0 \times 10^{-13})/(4 \times 10^{12} + 6.3 \times 10^{13})$

e. $\dfrac{9.5 + 4.1 + 2.8 + 3.175}{4}$

(Assume that this operation is taking the average of four numbers. Thus 4 in the denominator is exact.)

f. $\dfrac{8.925 - 8.905}{8.925} \times 100$

(This type of calculation is done many times in calculating a percentage error. Assume that this example is such a calculation; thus 100 can be considered to be an exact number.)

26. Perform the following mathematical operations, and express the result to the correct number of significant figures.

a. $6.022 \times 10^{23} \times 1.05 \times 10^2$

b. $\dfrac{6.6262 \times 10^{-34} \times 2.998 \times 10^8}{2.54 \times 10^{-9}}$

c. $1.285 \times 10^{-2} + 1.24 \times 10^{-3} + 1.879 \times 10^{-1}$

d. $\dfrac{(1.00866 - 1.00728)}{6.02205 \times 10^{23}}$

e. $\dfrac{9.875 \times 10^2 - 9.795 \times 10^2}{9.875 \times 10^2} \times 100$ (100 is exact)

f. $\dfrac{9.42 \times 10^2 + 8.234 \times 10^2 + 1.625 \times 10^3}{3}$ (3 is exact)

27. Perform each of the following conversions.

a. 8.43 cm to millimeters

b. 2.41×10^2 cm to meters

c. 294.5 nm to centimeters

d. 1.445×10^4 m to kilometers

e. 235.3 m to millimeters

f. 903.3 nm to micrometers

28. a. How many kilograms are in one teragram?

b. How many nanometers are in 6.50×10^2 terameters?

c. How many kilograms are in 25 femtograms?

d. How many liters are in 8.0 cubic decimeters?

e. How many microliters are in one milliliter?

f. How many picograms are in one microgram?