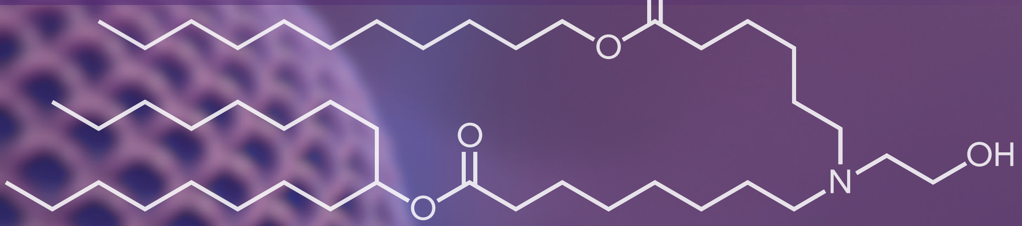
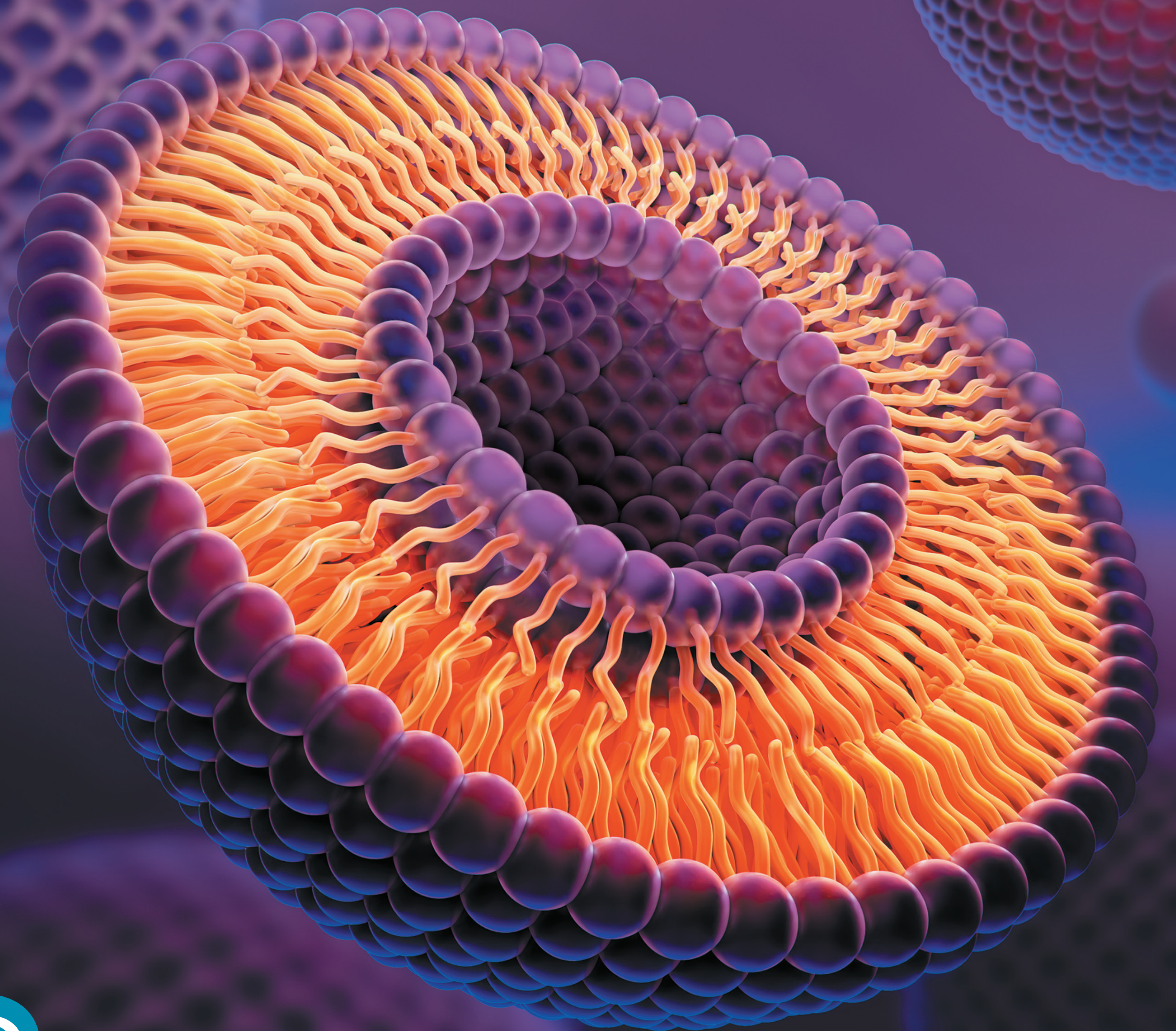
SIXTH EDITION

CHEMISTRY

A MOLECULAR APPROACH





NIVALDO J. TRO

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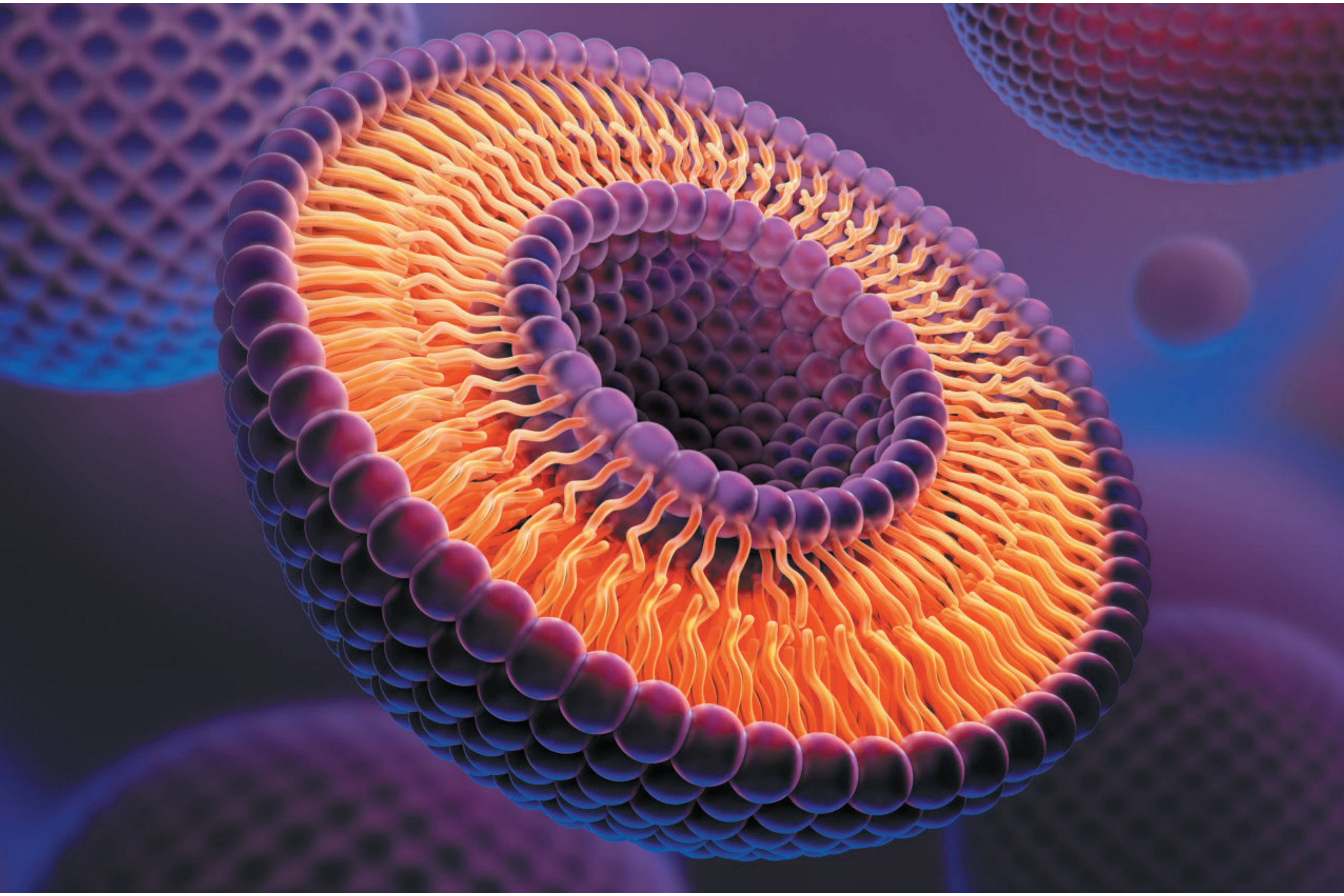

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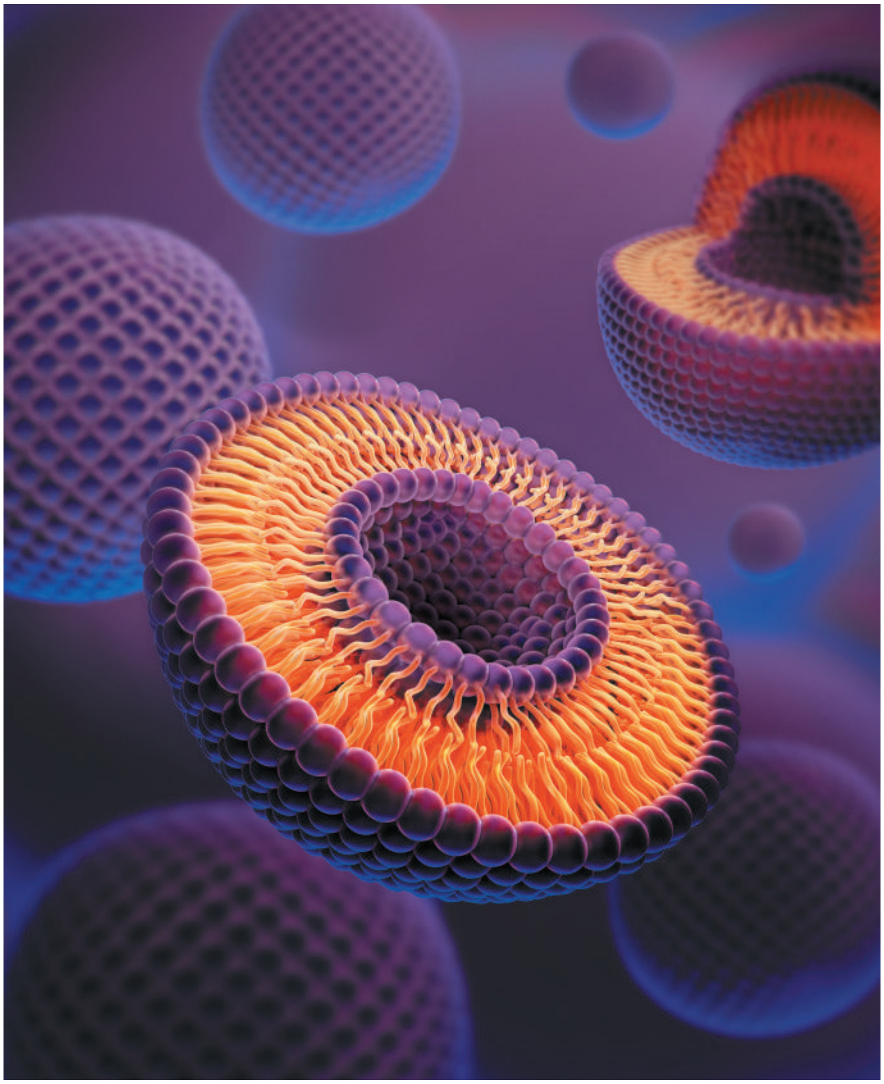
About the Author

Nivaldo Tro has been teaching college chemistry since 1990 and is currently teaching at the College of Creative Studies at the University of California, Santa Barbara, and at Santa Barbara City College. He received his Ph.D. in chemistry from Stanford University for work on developing and using optical techniques to study the adsorption and desorption of molecules to and from surfaces in ultrahigh vacuum. He then went on to the University of California at Berkeley, where he did postdoctoral research on ultrafast reaction dynamics in solution. Professor Tro has been awarded grants from the American Chemical Society Petroleum Research Fund, the Research Corporation, and the National Science Foundation to study the dynamics of various processes occurring in thin adlayer films adsorbed on dielectric surfaces. Professor Tro lives in Santa Barbara with his wife, Ann. In his leisure time, Professor Tro enjoys cycling, surfing, and being outdoors.

To Michael, Ali,
Kyle, and Kaden

About the Cover



The cover shows a lipid nanoparticle (LNP), a fat bubble that became central to the mRNA vaccines developed against COVID-19. These vaccines use lipid nanoparticles to house mRNA and transport it to cells. The molecule shown is 8-[(2-hydroxyethyl)[6-oxo-6-(undecyloxy)hexyl]amino]-octanoic acid, 1-octylnonyl ester, one of the components of the LNPs used in the Moderna vaccine. This molecule, and others like it, have a polar end and a nonpolar end, which results in the formation of the bilayers that compose the LNP shell.

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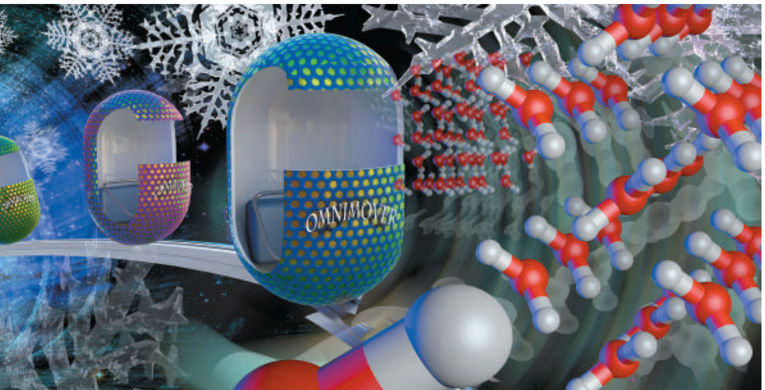

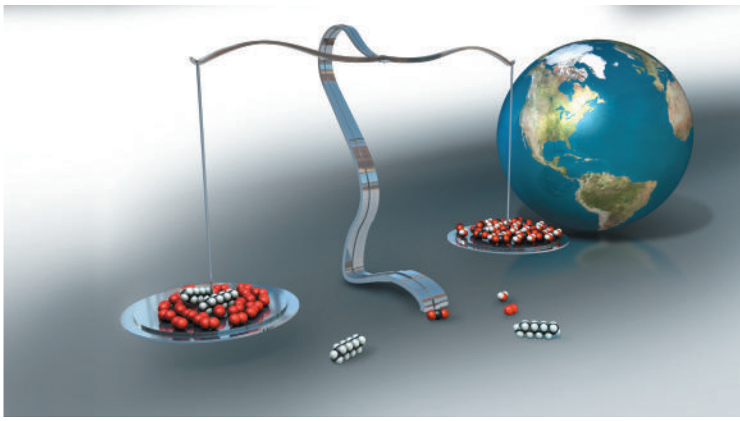

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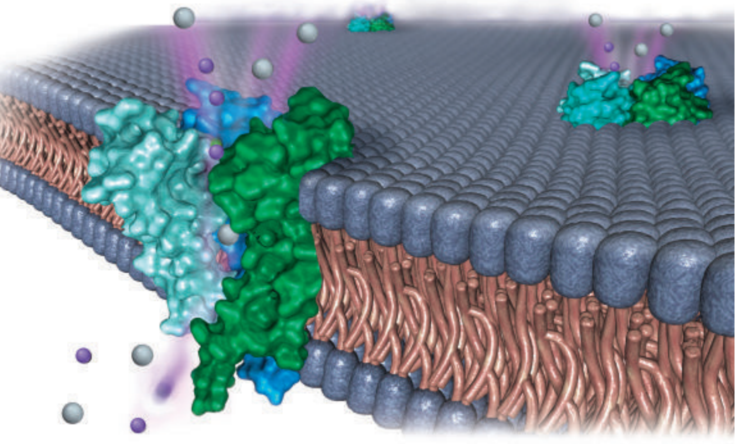

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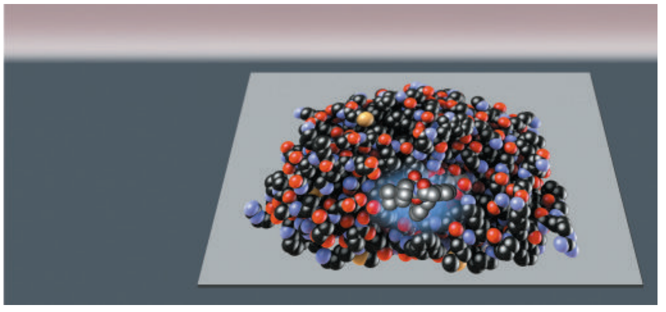

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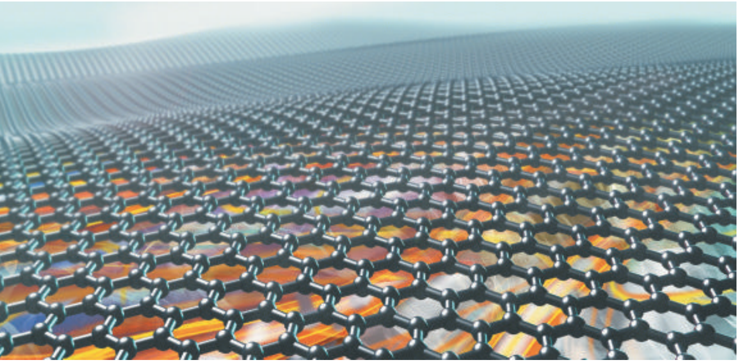


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Life of Pi

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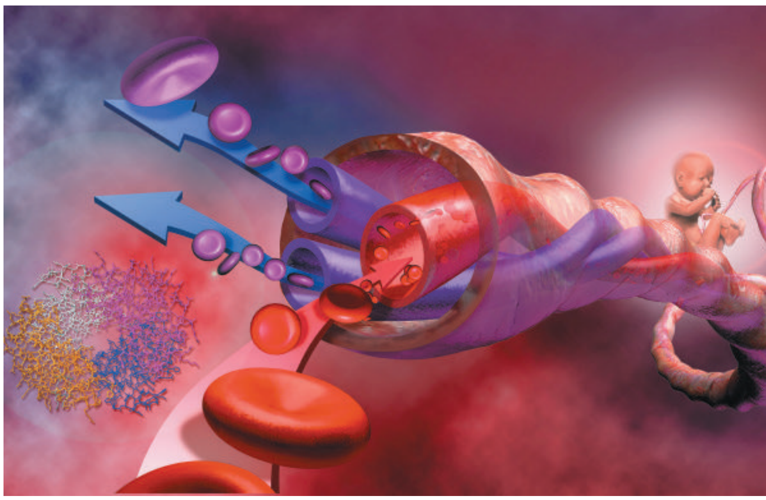

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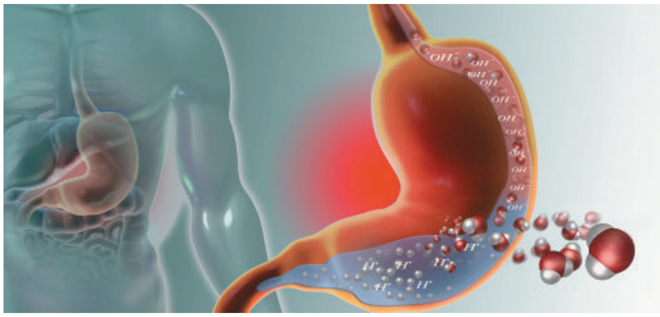

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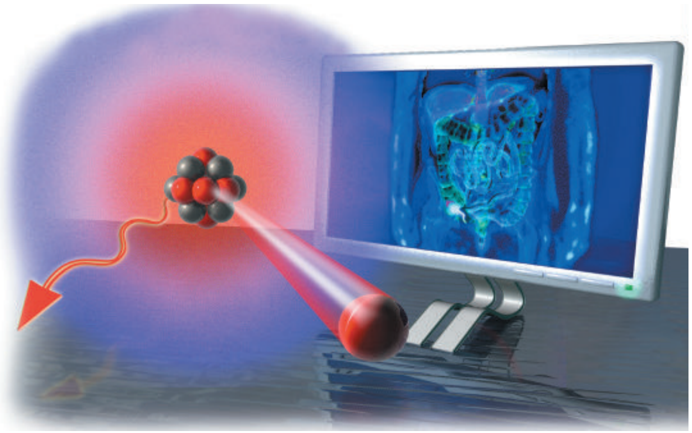

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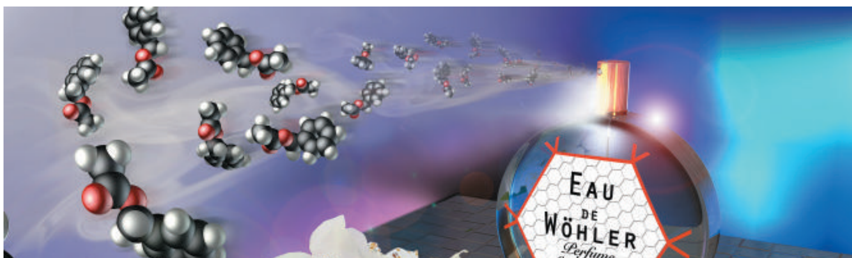

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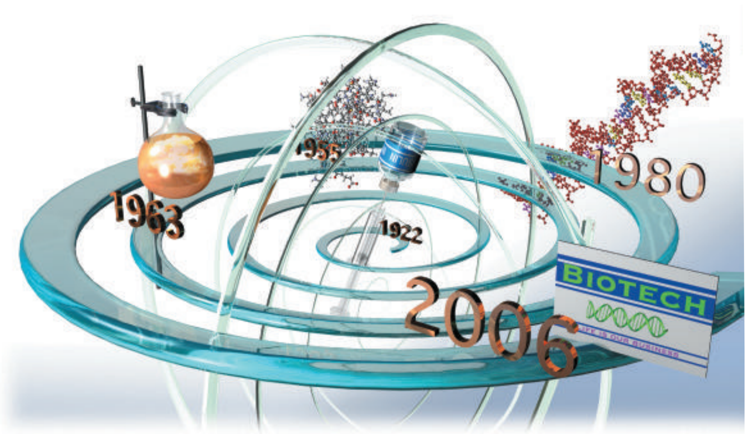

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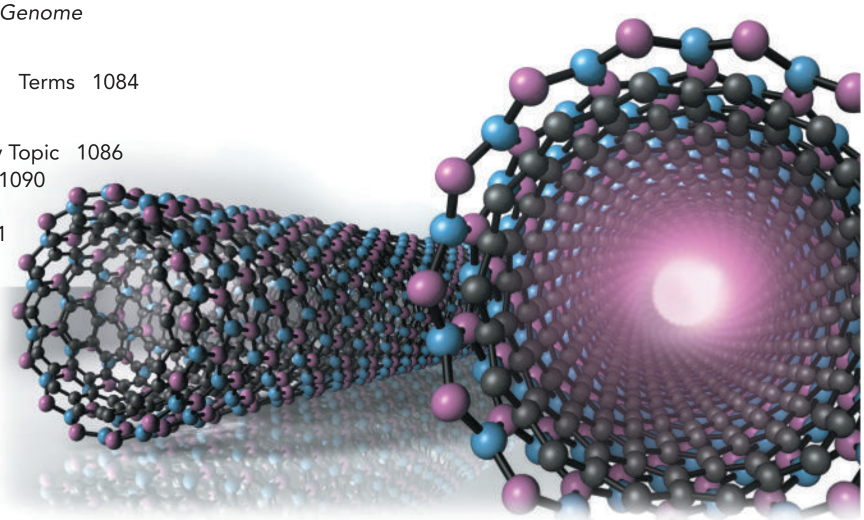

Preface

To the Student

As you begin this course, I invite you to think about your reasons for enrolling in it. Why are you taking general chemistry? More generally, why are you pursuing a college education? If you are like most college students taking general chemistry, part of your answer is probably that this course is required for your major and that you are pursuing a college education so you can get a good job some day. Although these are good reasons, I would like to suggest a better one. I think the primary reason for your education is to prepare you to *live a good life*. You should understand chemistry—not for what it can *get* you—but for what it can *do* to you. Understanding chemistry, I believe, is an important source of happiness and fulfillment. Let me explain.

Understanding chemistry helps you to live life to its fullest for two basic reasons. The first is *intrinsic*: through an understanding of chemistry, you gain a powerful appreciation for just how rich and extraordinary the world really is. The second reason is *extrinsic*: understanding chemistry makes you a more informed citizen—it allows you to engage with many of the issues of our day. In other words, understanding chemistry makes *you* a deeper and richer person and makes your country and the world a better place to live. These reasons have been the foundation of education from the very beginnings of civilization.

How does chemistry help prepare you for a rich life and conscientious citizenship? Let me explain with two examples. My first one comes from the very first page of Chapter 1 of this book. There, I ask the following question: What is the most important idea in all of scientific knowledge? My answer to that question is this: **the behavior of matter is determined by the properties of molecules and atoms**. That simple statement is the reason I love chemistry. We humans have been able to study the substances that compose the world around us and explain their behavior by reference to particles so small that they can hardly be imagined. If you have never realized the remarkable dependence of the world we *can* see on the world we *cannot*, you have missed out on a fundamental truth about our universe. To have never encountered this truth is like never having read a play by Shakespeare or seen a sculpture by Michelangelo—or, for that matter, like never having discovered that the world is round. It robs you of an amazing and unforgettable experience of the world and the human ability to understand it.

My second example demonstrates how science literacy helps you to be a better citizen. Although I am largely sympathetic to the environmental movement, a lack of science literacy within some sectors of that movement and the resulting anti-environmental backlash create confusion that impedes real progress and opens the door to what could be misinformed policies. For example, I have heard conservative pundits say that volcanoes emit more carbon dioxide—the most significant greenhouse gas—than does petroleum combustion. I have also heard a liberal environmentalist say that we have to stop using hair spray because it is causing holes in the ozone layer that will lead to global warming. Well, the claim about volcanoes emitting more carbon dioxide than petroleum combustion can be refuted by the basic tools you will learn to use in Chapter 4 of this book. We can easily show that volcanoes emit only 1/50th as much carbon dioxide as petroleum combustion. As for hair spray depleting the ozone layer and thereby leading to global warming, the chlorofluorocarbons that deplete ozone have been banned from hair spray since 1978, and ozone depletion has nothing to do with global warming anyway. People with special interests or axes to grind can conveniently distort the truth before an ill-informed public, which is why we all need to be knowledgeable.

So this is why I think you should take this course. Not just to satisfy the requirement for your major and not just to get a good job some day, but to help you to lead a fuller life and to make the world a little better for everyone. I wish you the best as you embark on the journey to understanding the world around you at the molecular level. The rewards are well worth the effort.

To the Professor

Thanks to all of you who adopted this book in its previous editions. You helped to make this book one of the most popular general chemistry textbooks in the world. I am grateful beyond words. Second, I have listened carefully to your feedback on the previous edition. The changes you see in this edition are the direct result of your input, as well as my own experience using the book in my general chemistry courses. If you have reviewed content or have contacted me directly, you will likely see your suggestions reflected in the changes I have made. Thank you.

This revision comes out in the midst of a pandemic and after a challenging year, for both the world and for higher education. Most of us taught remotely during the 2020–2021 academic year, lecturing into a Zoom meeting of often faceless students. The year demonstrated the power of technology for learning and also its shortcomings. I learned two important lessons: (1) Face-to-face teaching continues to be the superior mode for college instruction; and (2) technology can deliver learning experiences that are more valuable than passive studying. This revision focuses on lesson 2.

Technology can transform passive learning into active learning, which has been repeatedly demonstrated to be more effective. One of my main goals in this revision is to continue to expand tools to engage students in active learning. Although the term *active learning* has been applied mainly to in-class learning, the principal idea—*that we learn better when we are actively engaged*—applies to all of learning.

For this edition, I have worked with Pearson to develop 50 *Key Concept Interactives (KCIs)*. Each KCI guides the student through a key concept in that chapter. However, rather than passively reading a book, the student must continually interact with the material as they navigate through it. In the KCI on nomenclature, for example, students first learn how to categorize compounds as ionic, molecular, or acid. Before they advance, they are given a compound to categorize. If they categorize it incorrectly, they are provided with feedback to lead them in the right direction. If they categorize it correctly, they can advance. In this way, the student is completely active in the learning process, which we know from education research results in much greater retention than passive reading. In addition, each KCI has an associated follow-up question that can be assigned using Mastering Chemistry.

Other active learning tools, presented in previous editions, have been expanded in this edition. For example, the library of 3- to 6-minute *Key Concept Videos (KCVs)* now spans virtually all of the key concepts in a general chemistry course. The videos introduce a key concept and encourage active learning because they stop in the middle and pose a question that must be answered before the video continues playing. Each video also has an associated follow-up question that can be assigned using Mastering Chemistry.

Also expanded in this edition is the library of 3- to 6-minute videos called *Interactive Worked Examples (IWEs)*. Each IWE video walks a student through the solution to a chemistry problem. Like the KCV, the IWE video stops in the middle and poses a question that must be answered before the video continues playing. Each video also has an associated follow-up problem that can be assigned using Mastering Chemistry. Together, the library of KCVs and IWEs now comprises approximately 250 interactive videos.

Although we have added these active learning tools to this edition and made other changes as well, the book's goal remains the same: *to present a rigorous and accessible treatment of general chemistry in the context of relevance*. Teaching general chemistry would be much easier if all of our students had exactly the same level of preparation and ability. But alas, that is not the case. My own courses are populated with students with a range of backgrounds and abilities in chemistry. The challenge of successful teaching, in my opinion, is figuring out how to instruct and challenge the best students while not losing those with lesser backgrounds and abilities. My strategy has always been to set the bar relatively high, while at the same time providing the motivation and support necessary to reach the high bar. That is exactly the philosophy of this book. We do not have to compromise away rigor in order to make chemistry accessible to our students. In this book, I have worked hard to combine rigor with accessibility—to create a book that does not dilute the content, and yet, can be used and understood by any student willing to put in the necessary effort.

Chemistry: A Molecular Approach* is first and foremost a *student-oriented book. My main goal is to motivate students and get them to achieve at the highest possible level. As we all know, many students take general chemistry because it is a requirement; they do not see the connection between chemistry and their lives or their intended careers. *Chemistry: A Molecular Approach* strives to make those connections consistently and effectively. Unlike other books, which often teach chemistry as something that happens only in the laboratory or in industry, this book teaches chemistry in the context of relevance. It shows students *why* chemistry is important to them, to their future careers, and to their world.

Second, *Chemistry: A Molecular Approach* is a *pedagogically driven book.* In seeking to develop problem-solving skills, it applies a consistent approach (Sort, Strategize, Solve, and Check), usually in a two- or three-column format. In the two-column format, the left column shows the student how to analyze the problem and devise a solution strategy. It also lists the steps of the solution, explaining the rationale for each one, while the right column shows the implementation of each step. In the three-column format, the left column outlines the general procedure for solving an important category of problems that is then applied to two side-by-side examples. This strategy allows students to see both the general pattern and the slightly different ways in which the procedure may be applied in differing contexts. The aim is to help students understand both the *concept of the problem* (through the formulation of an explicit conceptual plan for each problem) and the *solution to the problem*.

Third, *Chemistry: A Molecular Approach* is a *visual book.* Wherever possible, I use images to deepen the student's insight into chemistry. In developing chemical principles, multipart images help show the connection between everyday processes visible to the unaided eye and what atoms and molecules are actually doing. Many of these images have three parts: macroscopic, molecular, and symbolic. This combination helps students to see the relationships between the formulas they write down on paper (symbolic), the world they see around them (macroscopic), and the atoms and molecules that compose that world (molecular). In addition, most figures are designed to teach rather than just to illustrate. They are rich with annotations and labels intended to help the student grasp the most important processes and the principles that underlie them. In this edition, the art program has been thoroughly revised in two major ways. First, navigation of the more complex figures has been reoriented to track from left to right whenever possible. Second, figure captions have been migrated into the image itself as an "author voice" that explains the image and guides the reader through it. The resulting images are rich with information but also clear and quickly understood.

Fourth, *Chemistry: A Molecular Approach* is a "big picture" book. At the beginning of each chapter, a short paragraph helps students to see the key relationships between the different topics they are learning. Through a focused and concise narrative, I strive to make the basic ideas of every chapter clear to the student. Interim summaries are provided at selected spots in the narrative, making it easier to grasp (and review) the main points of important discussions. And to make sure that students never lose sight of the forest for the trees, each chapter includes several *Conceptual Connections,* which ask them to think about concepts and solve problems without doing any math. I want students to learn the concepts, not just plug numbers into equations to churn out the right answer. This philosophy is also integral to the *Key Concept Videos,* which concisely reinforce student appreciation of the core concepts in each chapter.

***Chemistry: A Molecular Approach* is lastly a book that delivers the depth of coverage faculty want.** We do not have to cut corners and water down the material in order to get our students interested. We have to meet them where they are, challenge them to the highest level of achievement, and support them with enough pedagogy to allow them to succeed.

I hope that this book supports you in your vocation of teaching students chemistry. I am increasingly convinced of the importance of our task. Please feel free to contact me with any questions or comments about the book.

Nivaldo J. Tro
nivatro@gmail.com

Acknowledgments

The book you hold in your hands bears my name on the cover, but I am really only one member of a large team that carefully crafted this book. Most importantly, I thank my editor, Elizabeth Bell. I have known Lizzy for many years, first as my marketing manager and now as my editor. Lizzy is a constant source of good, market-informed ideas. She is creative, energetic, and always follows through on her ideas and commitments. I am fortunate to work with such a skilled and insightful colleague. Thanks, Lizzy. Thanks also to Coleen Morrison, my new developmental editor on this project. Coleen has helped me to develop the new material in this edition. She is careful, organized, and always improves my work. Thanks, Coleen, for your attention to detail and all your hard work. Thanks also to my project manager, Shercian Kinosian. She has managed the many details and moving parts of producing this book with care and precision. I appreciate her steady hand and hard work. Thanks also to my media editor, Jackie Jacob. Jackie and I have been working together for many years to produce innovative media pieces that are pedagogically sound and easy to use. She is simply the best in the business, and I am lucky to get to work with her. I am also grateful to Ian Desrosiers and Chloe Veylit who have helped tremendously with the development of the new Key Concept Interactives and have stepped in at different stages to offer their talent and expertise.

I am also grateful to Jeanne Zalesky, Editor-in-Chief for Physical Sciences. She has supported me and my projects and allowed me to succeed. Thanks also to Adam Jaworski, who oversees product management in Science at Pearson. I am grateful to have his wise and steady, yet innovative, hand at the wheel, especially during the many changes that are happening within educational publishing. I am also grateful to the Lachina Creative, Inc. team and Jerilyn DiCarlo for their creativity and hard work in crafting the cover illustration and design of this text. I also thank Francesca Monaco and her coworkers at Straive. I am a picky author and Francesca is endlessly patient and a true professional. I am also greatly indebted to my copy editor, Betty Pessagno, for her dedication and professionalism over many projects.

I am also grateful to those who have supported me personally while working on this book. First on that list is my wife, Ann. Her patience and love for me are beyond description, and without her, this book would never have been written. I am also indebted to my children, Michael, Ali, Kyle, and Kaden, whose smiling faces and love of life always inspire me. I come from a large Cuban family whose closeness and support most people would envy. Thanks to my parents, Nivaldo and Sara; my siblings, Sarita, Mary, and Jorge; my siblings-in-law, Nachy, Karen, and John; and my nephews and nieces, Germain, Danny, Lisette, Sara, and Kenny. These are the people with whom I celebrate life.

I am especially grateful to Kyle Tro, who put in many hours proofreading changes in the manuscript, working problems, and organizing appendices. Kyle, you are an amazing person—it is my privilege to have you work with me on this project.

I would like to thank all of the general chemistry students who have been in my classes throughout my 32 years as a professor. You have taught me much about teaching that is now in this book, and you are the motivation that keeps me inspired to develop the best learning tools possible.

Lastly, I am indebted to the many reviewers, listed on the following pages, whose ideas are embedded throughout this book. They have corrected me, inspired me, and sharpened my thinking on how best to teach this subject we call chemistry. I deeply appreciate their commitment to this project. I am also grateful to the accuracy reviewers who tirelessly checked page proofs for correctness.

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Peter M. Hierl, *The University of Kansas*
Paul I. Higgs, *Barry University*
Melissa Hines, *Cornell University*
Paula Hjorth-Gustin, *San Diego Mesa College*
Angela Hoffman, *University of Portland*
Norris W. Hoffman, *University of South Alabama*
Tony Holland, *Wallace Community College*
Todd A. Hopkins, *Butler University*
Byron E. Howell, *Tyler Junior College*
Tina Huang, *University of Illinois*
Amy Irwin, *Monroe Community College*
Moheb Ishak, *St. Petersburg College, St. Petersburg*
Kamal Ismail, *CUNY, Bronx Community College*
Ralph Isovitsch, *Xavier University of Louisiana*
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Milt Johnston, *University of South Florida*
Greg M. Jorgensen, *American River College*
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Jason A. Kautz, *University of Nebraska-Lincoln*
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Resa Kelly, *San Jose State University*
Chulsung Kim, *Georgia Gwinnett College*
Ellen Kime-Hunt, *Riverside Community College, Riverside Campus*
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Behnoush Memari, *Broward College*
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Tom Selegue, *Pima Community College, West*
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Donald L. Siegel, *Rutgers University, New Brunswick*
Sarah Siegel, *Gonzaga University*
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Sherril Soman-Williams, *Grand Valley State University*
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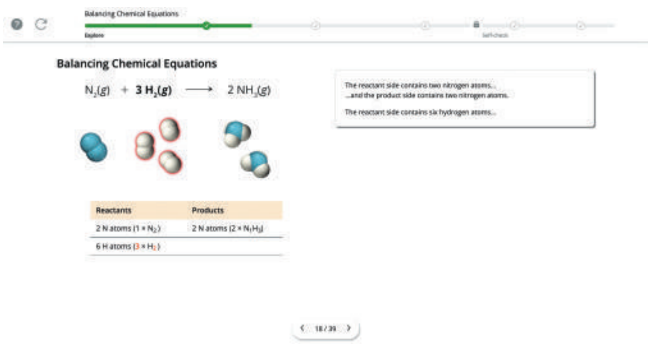
Galina G. Talanova, *Howard University*
Dennis Taylor, *Clemson University*
Robert Craig Taylor, *Oakland University*
Claire A. Tessier, *University of Akron*
Jacquelyn Thomas, *Southwestern College*
Lydia Tien, *Monroe Community College*
Tara Todd, *Vanderbilt University*
David Livingstone Toppen, *California State University Northridge*
Marcy Towns, *Purdue University*
Harold Trimm, *Broome Community College*
Frank Tsung, *Boston College*
Laura VanDorn, *University of Arizona*
Susan Varkey, *Mount Royal College*
Ramaiyer Venkatraman, *Jackson State University*
John Vincent, *University of Alabama*
John B. Vincent, *University of Alabama, Tuscaloosa*
Kent S. Voelkner, *Lake Superior College*
Sheryl K. Wallace, *South Plains College*
Wayne E. Wesolowski, *University of Arizona*
Sarah E. West, *Notre Dame University*
John Wiginton, *University of Mississippi*
Kurt J. Winkelmann, *Florida Institute of Technology*
Gary L. Wood, *Valdosta State University*
Troy D. Wood, *University of Buffalo*
Servet M. Yatin, *Quincy College*
Kazushige Yokoyama, *SUNY Geneseo*
Lin Zhu, *IUPUI*
Jim Zubricky, *University of Toledo*

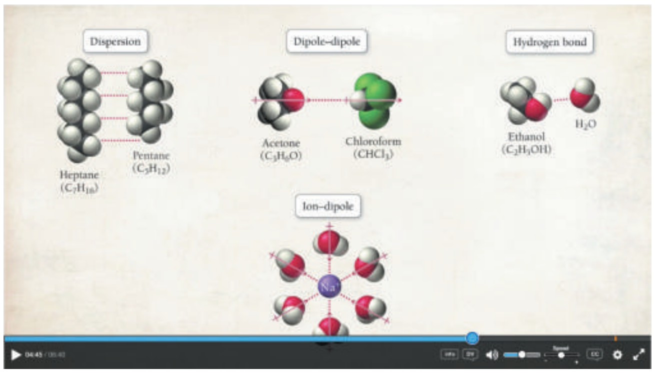
What's New in This Edition?

The book has been extensively revised and contains more small changes than can be detailed here. The most significant changes to the book and its supplements are listed below:

NEW KEY CONCEPT INTERACTIVES

50 new *Key Concept Interactives (KCIs)* have been added to the eText and are assignable in Mastering Chemistry. Each interactive guides a student through a key topic as they navigate through a series of interactive screens. As they work through the KCI, they are presented with questions that must be answered to progress. Wrong answers result in feedback to guide them toward success.



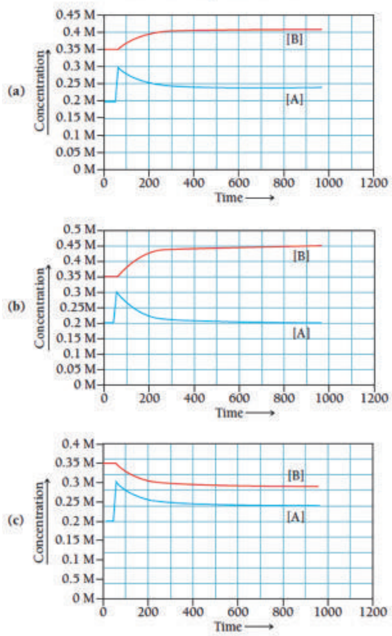


NEW INTERACTIVE VIDEOS

15 new *Key Concept Videos (KCVs)* and 25 new *Interactive Worked Examples (IWEs)* have been added to the media package that accompanies the book. All videos are available within the eText and assignable in Mastering Chemistry. *The video library now contains approximately 250 interactive videos.* These tools are designed to help professors engage their students in active learning.

NEW AND REVISED END-OF-CHAPTER PROBLEMS

176 New End-of-Chapter questions have been added throughout the book, and **378 have been revised**. Many new End-of-Chapter questions involve the interpretation of graphs and data. All new End-of-Chapter questions are assignable in Mastering Chemistry.


Le Châtelier's Principle

65. The reaction $A(g) \rightarrow B(g)$ was at equilibrium and then disturbed by adding additional A to the reaction mixture. The reaction is then allowed to come back to equilibrium. Which graph represents the concentrations of A and B during this process?
MISSED THIS? *Read Section 16.9; Watch KCV 16.9*

NEW ONLINE PROBLEM SETS

Online problem sets are web-based, online-only problems that are algorithmically randomized. They provide answer-specific feedback and will be continually updated and expanded.

REVISED DATA

All the data throughout the book have been updated to reflect the most recent measurements available. These updates include Figure 4.2: *Carbon Dioxide Concentrations in the Atmosphere*; Figure 4.3: *Global Temperature*; the unnumbered figure in Section 7.10 of *U.S. Energy Consumption by Sector*; Figure 7.12: *Energy Consumption by Source*; Table 7.6: *Changes in National Average Pollutant Levels, 1990–2019*.

DIVERSITY, EQUITY, AND INCLUSION REVIEW

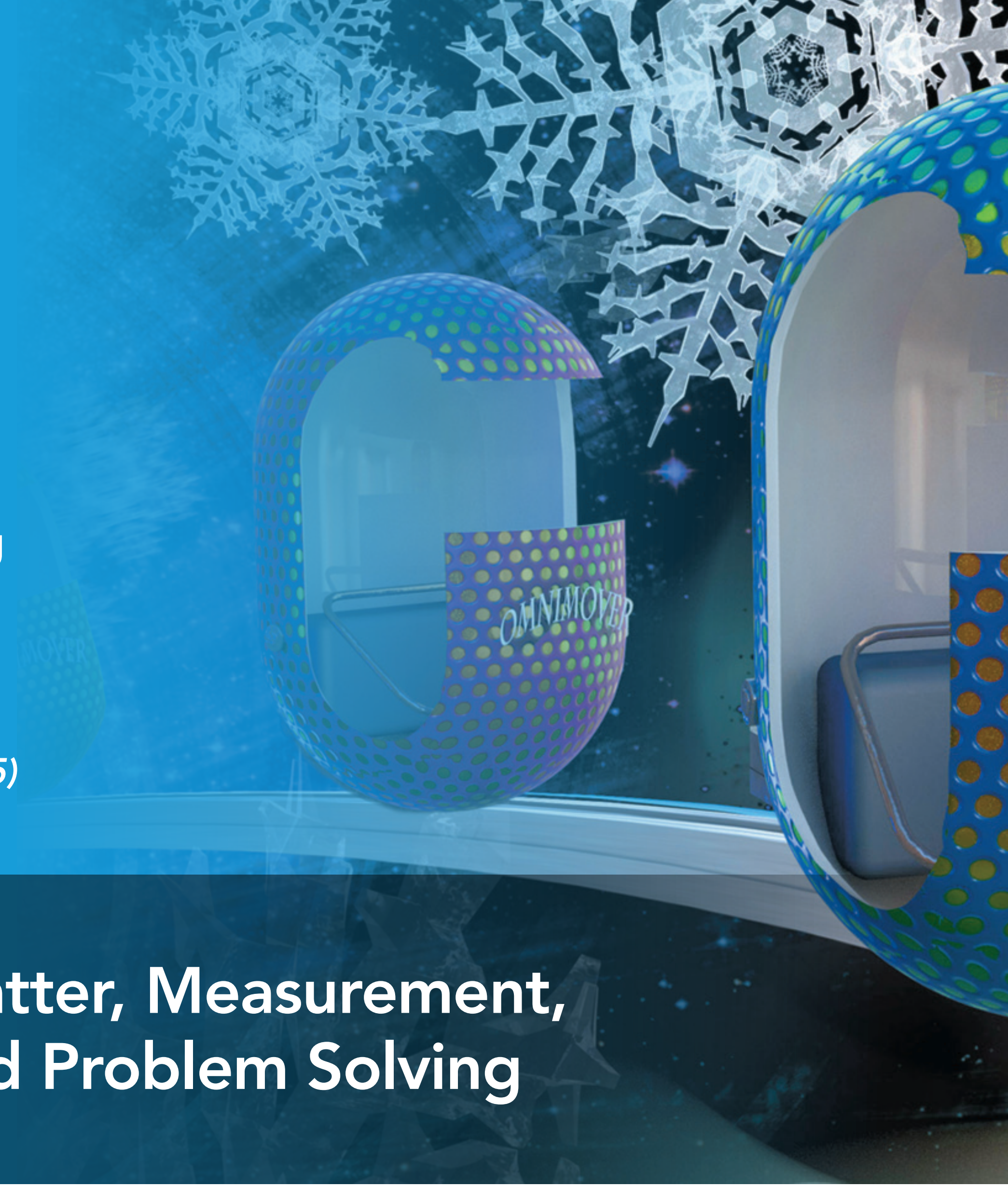
As mentioned in the Preface, the entire book went through a detailed review to ensure the content reflects the rich diversity of our learners and is inclusive of their lived experiences.

CHEMISTRY

A MOLECULAR APPROACH

The most incomprehensible thing about the universe is that it is comprehensible.

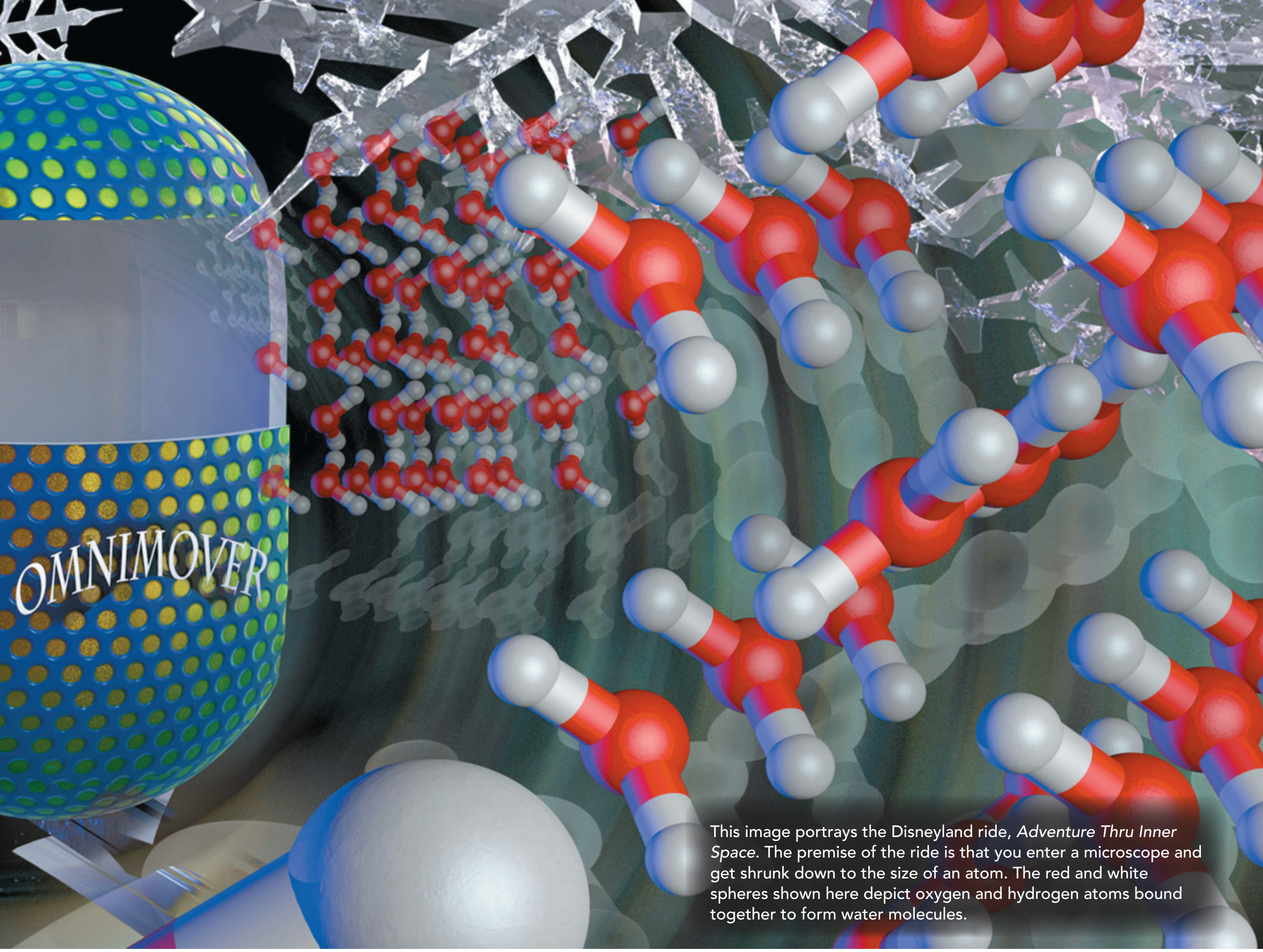
—Albert Einstein (1879–1955)



CHAPTER 1

Matter, Measurement, and Problem Solving

What do you think is the most important idea in all of human knowledge? This question has many possible answers—some practical, some philosophical, and some scientific. If we limit ourselves to scientific answers, mine would be this: **the properties of matter are determined by the properties of atoms and molecules**. Atoms and molecules determine how matter behaves—if they were different, matter would be different. The properties of water molecules determine how water behaves, the properties of sugar molecules determine how sugar behaves, and the properties of the molecules that compose our bodies determine how our bodies behave. The understanding of matter at the molecular level gives us unprecedented control over that matter. For example, our understanding of the details of the molecules that compose living organisms has revolutionized biology over the last 50 years.



This image portrays the Disneyland ride, *Adventure Thru Inner Space*. The premise of the ride is that you enter a microscope and get shrunk down to the size of an atom. The red and white spheres shown here depict oxygen and hydrogen atoms bound together to form water molecules.

1.1 Atoms and Molecules

As I sat in the "omnimover" and listened to the narrator's voice telling me that I was shrinking down to the size of an atom, I grew apprehensive but curious. Just minutes before, while waiting in line, I witnessed what appeared to be full-sized humans entering a microscope and emerging from the other end many times smaller. I was seven years old, and I was about to ride *Adventure Thru Inner Space*, a Disneyland ride (in Tomorrowland) that simulated what it would be like to shrink to the size of an atom. The ride

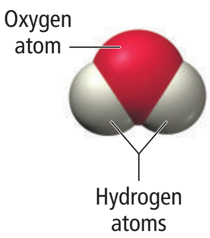
WATCH **NOW!**

KEY CONCEPT VIDEO 1.1
Atoms and Molecules

began with darkness and shaking, but then the shaking stopped and giant snowflakes appeared. The narrator explained that you were in the process of shrinking to an ever-smaller size (which explains why the snowflakes grew larger and larger). Soon, you entered the wall of the snowflake itself and began to see water molecules all around you. These also grew larger as you continued your journey into inner space and eventually ended up within the atom itself. Although this Disneyland ride bordered on being corny, and although it has since been shut down, it was my favorite ride as a young child.

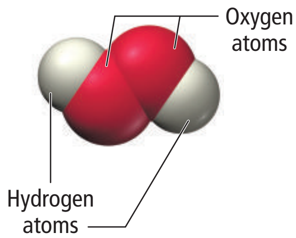
That ride sparked my interest in the world of atoms and molecules, an interest that has continued and grown to this day. I am a chemist because I am obsessed with the connection between the "stuff" around us and the atoms and molecules that compose that stuff. More specifically, I love the idea that we humans have been able to figure out the connection between the *properties of the stuff* around us and the *properties of atoms and molecules*. **Atoms** are submicroscopic particles that are the fundamental building blocks of ordinary matter. Free atoms are rare in nature; instead they bind together in specific geometrical arrangements to form **molecules**. A good example of a molecule is the water molecule, which I remember so well from the Disneyland ride.

Water molecule



A water molecule is composed of one oxygen atom bound to two hydrogen atoms in the shape shown at left. The exact properties of the water molecule—the atoms that compose it, the distances between those atoms, and the geometry of how the atoms are bound together—determine the properties of water. If the molecule were different, water would be different. For example, if water contained two oxygen atoms instead of just one, it would be a molecule like this:

Hydrogen peroxide molecule



The hydrogen peroxide we use as an antiseptic or bleaching agent is considerably diluted.

This molecule is hydrogen peroxide, which you may have encountered if you have ever bleached your hair. A hydrogen peroxide molecule is composed of *two* oxygen atoms and two hydrogen atoms. This seemingly small molecular difference results in a huge difference in the properties of water and hydrogen peroxide. Water is the familiar and stable liquid we all drink and bathe in. Hydrogen peroxide, in contrast, is an unstable liquid that, in its pure form, burns the skin on contact and is used in rocket fuel. When you pour water onto your hair, your hair simply becomes wet. However, if you put diluted hydrogen peroxide on your hair, a chemical reaction occurs that strips your hair of its color.

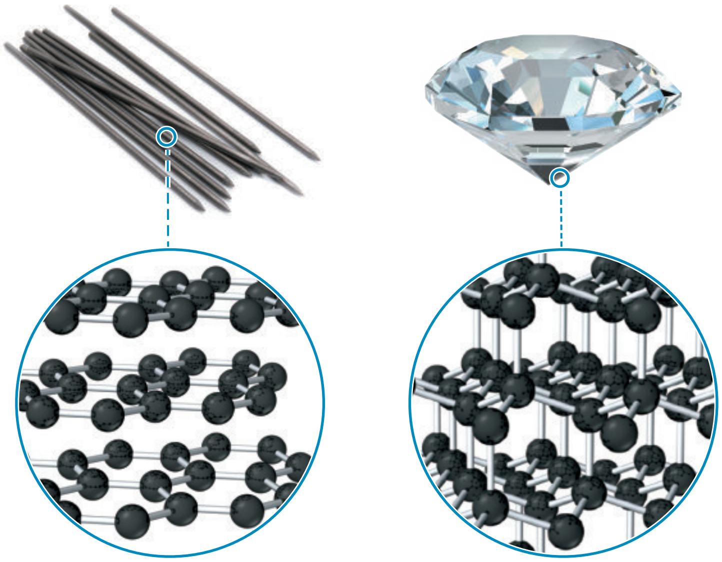
The details of how specific atoms bond to form a molecule—in a straight line, at a particular angle, in a ring, or in some other pattern—as well as the type of atoms in the molecule, determine everything about the substance that the molecule composes. If we want to understand the substances around us, we must understand the atoms and molecules that compose them—this is the central goal of chemistry. A good simple definition of **chemistry** is

Chemistry—the science that seeks to understand the behavior of matter by studying the behavior of atoms and molecules.

The term *atoms* in this definition can be interpreted loosely to include atoms that have lost or gained electrons.

Throughout this book, we explore the connection between atoms and molecules and the matter they compose. We seek to understand how differences on the atomic or molecular level affect the properties on the macroscopic level. Before we move on, let's examine one more example that demonstrates this principle. Consider the structures of graphite and diamond.

Graphite is the slippery black substance (often called pencil lead) that you have probably used in a mechanical pencil. Diamond is the brilliant gemstone found in jewelry. Graphite and diamond are both composed of exactly the same atoms—carbon atoms. The striking differences between the substances are a result of how those atoms are arranged. In graphite, the atoms are arranged in sheets. The atoms within each sheet are tightly bound to each other, but the sheets are *not* tightly bound to other sheets. Therefore the sheets can slide past each other, which is why the graphite in a pencil leaves a trail as you write. In diamond, by contrast, the carbon atoms are all bound together in a three-dimensional structure where layers are strongly bound to other layers, resulting in the strong, nearly unbreakable substance. This example illustrates how even the same atoms can compose vastly different substances when they are bound together in different patterns. Such is the atomic and molecular world—small differences in atoms and molecules can result in large differences in the substances that they compose.



Graphite structure

Diamond structure

1.2 The Scientific Approach to Knowledge

Throughout history, humans have approached knowledge about the physical world in different ways. For example, the Greek philosopher Plato (427–347 B.C.E.) thought that the best way to learn about reality was—not through the senses—but through reason. He believed that the physical world was an imperfect representation of a perfect and transcendent world (a world beyond space and time). For him, true knowledge came, not through observing the real physical world, but through reasoning and thinking about the ideal one.

Although some Greek philosophers, such as Aristotle, did use observation to attain knowledge, they did not emphasize experiment and measurement to the extent that modern science does.

The *scientific* approach to knowledge, however, is exactly the opposite of Plato's. Scientific knowledge is empirical—it is based on *observation* and *experiment*. Scientists observe and perform experiments on the physical world to learn about it. Some observations and experiments are qualitative (noting or describing how a process happens), but many are quantitative (measuring or quantifying something about the process). For example, Antoine Lavoisier (1743–1794), a French chemist who studied combustion (burning), made careful measurements of the mass of objects before and after burning them in closed containers. He noticed that there was no change in the total mass of material within the container during combustion. In doing so, Lavoisier made an important *observation* about the physical world.

▲ French chemist Antoine Lavoisier with his wife, Marie, who helped him in his work by illustrating his experiments and translating scientific articles from English. Lavoisier, who also made significant contributions to agriculture, industry, education, and government administration, was executed during the French Revolution.
(The Metropolitan Museum of Art)

Observations often lead scientists to formulate a **hypothesis**, a tentative interpretation or explanation of the observations. For example, Lavoisier explained his observations on combustion by hypothesizing that when a substance burns, it combines with a component of air. A good hypothesis is *falsifiable*, which means that it makes predictions that can be confirmed or refuted by further observations. Scientists test hypotheses by **experiments**, highly controlled procedures designed to generate observations that confirm or refute a hypothesis. The results of an experiment may support a hypothesis or prove it wrong—in which case the scientist must modify or discard the hypothesis.

In some cases, a series of similar observations leads to the development of a **scientific law**, a brief statement that summarizes past observations and predicts future ones. Lavoisier summarized his observations on combustion with the **law of conservation of mass**, which states, "In a chemical reaction, matter is neither created nor destroyed." This statement summarized his observations on chemical reactions and predicted the outcome of future observations on reactions. Laws, like hypotheses, are also subject to experiments, which can support them or prove them wrong.

Scientific laws are not *laws* in the same sense as civil or governmental laws. Nature does not follow laws in the way that we obey the laws against speeding or running a stop sign. Rather, scientific laws *describe* how nature behaves—they are generalizations about what nature does. For that reason, some people find it more appropriate to refer to them as *principles* rather than *laws*.

One or more well-established hypotheses may form the basis for a scientific **theory**. A scientific theory is a model for the way nature is and tries to explain not merely what nature does but why. As such, well-established theories are the pinnacle of scientific knowledge, often predicting behavior far beyond the observations or laws from which they were developed. A good example of a theory is the **atomic theory** proposed by English chemist John Dalton (1766–1844). Dalton explained the law of conservation of mass, as well as other laws and observations of the time, by proposing that matter is composed of small, indestructible particles called atoms. Since these particles are merely rearranged in chemical changes (and not created or destroyed), the total amount of mass remains the same. Dalton's theory is a model for the physical world—it gives us insight into how nature works and, therefore, *explains* our laws and observations.

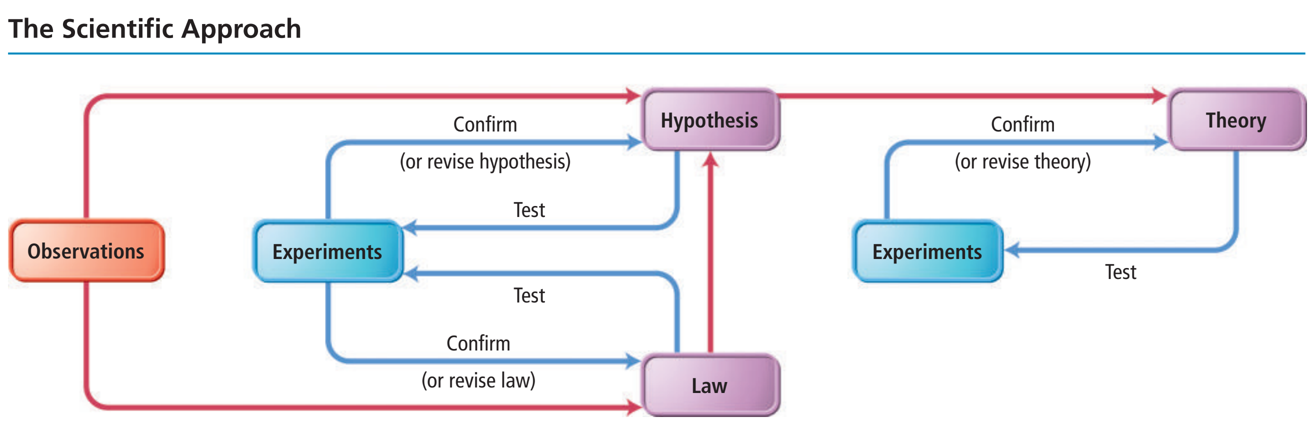
In Dalton's time, people thought atoms were indestructible. Today, because of nuclear reactions, we know that atoms can be broken apart into their smaller components.

Finally, the scientific approach returns to observation to test theories. For example, scientists can test the atomic theory by trying to isolate single atoms or by trying to image them (both of which, by the way, have already been accomplished). Theories are validated by experiments; however, theories can never be conclusively proven because some new observation or experiment always has the potential to reveal a flaw. Notice that the scientific approach to knowledge begins with observation and ends with observation. An experiment is in essence a highly controlled procedure for generating critical observations designed to test a theory or hypothesis. Each new set of observations has the potential to refine the original model. Figure 1.1▼ summarizes one way to map the scientific approach to knowledge. Scientific laws, hypotheses, and theories are all subject to continued experimentation. If a law, hypothesis, or theory is proved wrong by an experiment, it must be revised and tested with new experiments. Over time, the scientific community eliminates or corrects poor theories and laws, and valid theories and laws—those consistent with experimental results—remain.

Established theories with strong experimental support are the most powerful pieces of scientific knowledge. You may have heard the phrase "That is just a theory," as if theories are easily dismissible. Such a statement reveals a deep misunderstanding of the nature of a scientific theory. Well-established theories are as close to truth as we get in science. The idea that all matter is made of atoms is "just a theory," but it has over 200 years of experimental evidence to support it. It is a powerful piece of scientific knowledge on which many other scientific ideas are based.

One last word about the scientific approach to knowledge: some people wrongly imagine science to be a strict set of rules and procedures that automatically leads to inarguable, objective facts. This is not the case. Even our diagram of the scientific approach to knowledge is only an idealization of real science, useful to help us see the key distinctions of science. Real science requires hard work, care, creativity, and even a bit of luck.

▼ **FIGURE 1.1** **The Scientific Approach to Knowledge**



Scientific theories do not just arise out of data—individuals of genius and creativity craft theories. A great theory is not unlike a master painting, and many see a similar kind of beauty in both. (For more on this aspect of science, see the accompanying box entitled *Thomas S. Kuhn and Scientific Revolutions*.)

LAWS AND THEORIES Which statement best explains the difference between a law and a theory?

(a) A law is truth; a theory is mere speculation.
(b) A law summarizes a series of related observations; a theory gives the underlying reasons for them.
(c) A theory describes *what* nature does; a law describes *why* nature does it.

1.1 Cc Conceptual Connection

ANSWER **NOW!**

THE NATURE OF SCIENCE Thomas S. Kuhn and Scientific Revolutions

When scientists talk about science, they often talk in ways that imply that theories are "true." Further, they talk as if they arrive at theories in logical and unbiased ways. For example, a theory central to chemistry that we have discussed in this chapter is John Dalton's atomic theory—the idea that all matter is composed of atoms. Is this theory "true"? Was it reached in logical, unbiased ways? Will this theory still be around in 200 years?

The answers to these questions depend on how we view science and its development. One way to view science—let's call it the *traditional view*—is as the continual accumulation of knowledge and the building of increasingly precise theories. In this view, a scientific theory is a model of the world that reflects what is *actually in* nature. New observations and experiments result in gradual adjustments to theories. Over time, theories get better, giving us a more accurate picture of the physical world.

In the twentieth century, a different view of scientific knowledge began to develop. A book by Thomas Kuhn (1922–1996), published in 1962 and entitled *The Structure of Scientific Revolutions*, challenged the traditional view. Kuhn's ideas came from his study of the history of science, which, he argued, does not support the idea that science progresses in a smooth, cumulative way. According to Kuhn, science goes through fairly quiet periods that he called *normal science*. In these periods, scientists make their data fit the reigning theory, or paradigm. Small inconsistencies are swept aside during periods of normal science. However, when too many inconsistencies and anomalies develop, a crisis emerges. The crisis brings about a *revolution* and a new reigning theory. According to Kuhn, the new theory is usually quite different from the old one; it not only helps us to make sense of new or anomalous information, but it also enables us to see accumulated data from the past in a dramatically new way.

Kuhn further contended that theories are held for reasons that are not always logical or unbiased, and that theories are not *true* models—in the sense of a one-to-one mapping—of the physical world. Because new theories are often so different from the ones they replace, he argued, and because old theories always make good sense to those holding them, they must not be "True" with a capital *T*; otherwise "truth" would be constantly changing.

Kuhn's ideas created a controversy among scientists and science historians that continues to this day. Some, especially postmodern philosophers of science, have taken Kuhn's ideas one step further. They argue that scientific knowledge is *completely* biased and lacks any objectivity. Most scientists, including Kuhn, would disagree. Although Kuhn pointed out that scientific knowledge has *arbitrary elements*, he also said, "*Observation . . . can and must drastically restrict the range of admissible scientific belief, else there would be no science.*" In other words, saying that science contains arbitrary elements is quite different from saying that science itself is arbitrary.

QUESTION *In his book, Kuhn stated, "A new theory . . . is seldom or never just an increment to what is already known." From your knowledge of the history of science, can you think of any examples that support Kuhn's statement? Do you know of any instances in which a new theory or model was drastically different from the one it replaced?*

1.3 The Classification of Matter

WATCH **NOW!**

KEY CONCEPT VIDEO 1.3
Classifying Matter

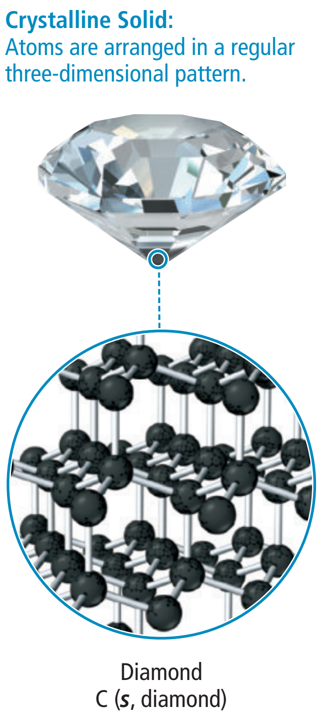
Matter is anything that occupies space and has mass. Your desk, your chair, and even your body are all composed of matter. Less obviously, the air around you is also matter—it too occupies space and has mass. We call a specific instance of matter—such as air, water, or sand—a **substance**. We classify matter according to its **state** (its physical form) and its **composition** (the basic components that make it up).

The state of matter changes from solid to liquid to gas with increasing temperature.

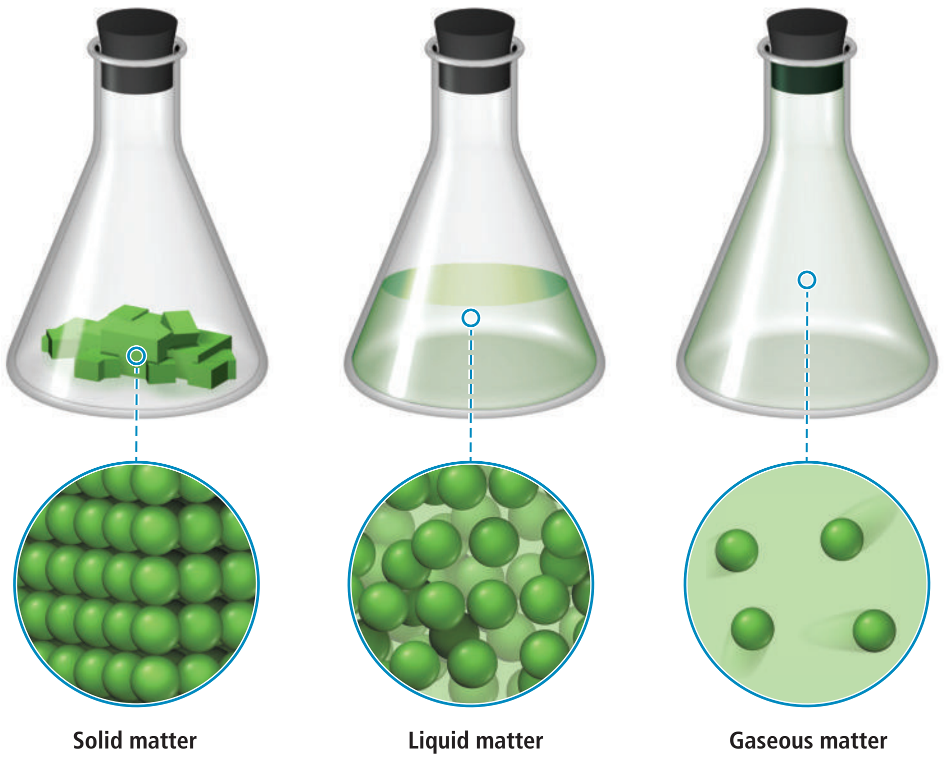
Glass and other amorphous solids can be thought of, from one point of view, as intermediate between solids and liquids. Their atoms are fixed in position at room temperature, but they have no long-range structure and do not have distinct melting points.

The States of Matter: Solid, Liquid, and Gas

Matter exists in three different states: **solid**, **liquid**, and **gas**. In *solid matter*, atoms or molecules pack closely to each other in fixed locations. Although the atoms and molecules in a solid vibrate, they do not move around or past each other. Consequently, a solid has a fixed volume and rigid shape. Ice, aluminum, and diamond are examples of solids. Solid matter may be **crystalline**, in which case its atoms or molecules are in patterns with long-range, repeating order (Figure 1.2▼), or it may be **amorphous**, in which case its atoms or molecules do not have any long-range order. Table salt and diamond are examples of *crystalline* solids; the well-ordered geometric shapes of salt and diamond crystals reflect the well-ordered geometric arrangement of their atoms (although this is not the case for *all* crystalline solids). Examples of *amorphous* solids include glass and plastic. In *liquid matter*, atoms or molecules pack about as closely as they do in solid matter, but they are free to move relative to each other, giving liquids a fixed volume but not a fixed shape. Liquids assume the shape of their containers. Water, alcohol, and gasoline are all substances that are liquids at room temperature.

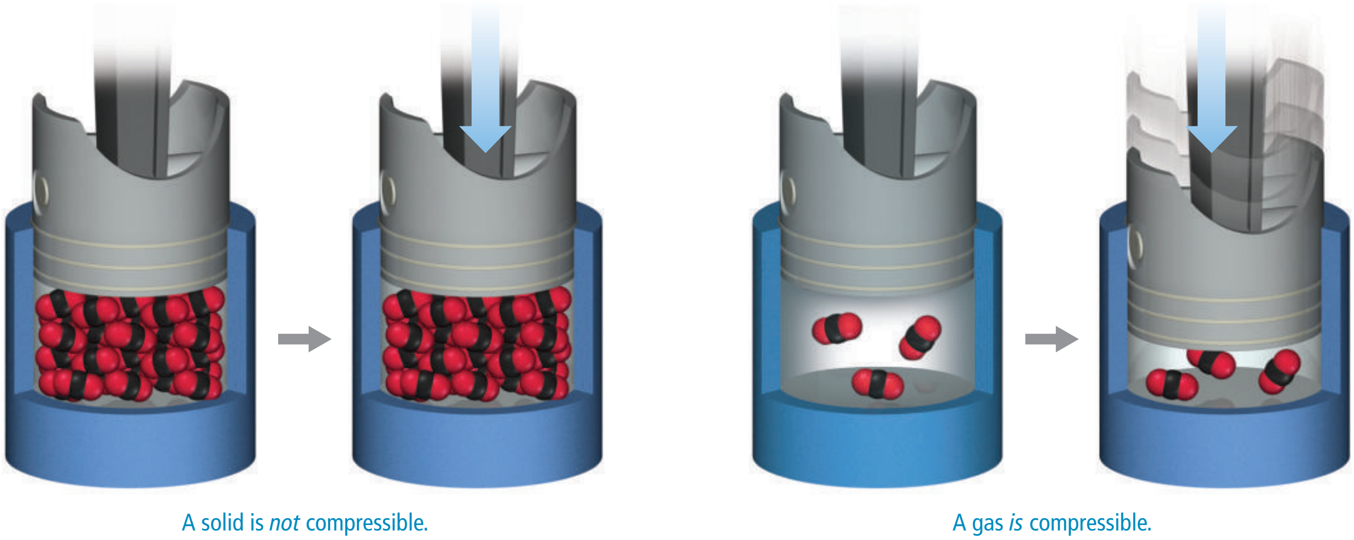


▲ **FIGURE 1.2 Crystalline Solid** Diamond (first discussed in Section 1.1) is a crystalline solid composed of carbon atoms arranged in a regular, repeating pattern.



▲ In a solid, the atoms or molecules are fixed in place and can only vibrate. In a liquid, although the atoms or molecules are closely packed, they can move past one another, allowing the liquid to flow and assume the shape of its container. In a gas, the atoms or molecules are widely spaced, making gases compressible as well as fluid (able to flow).

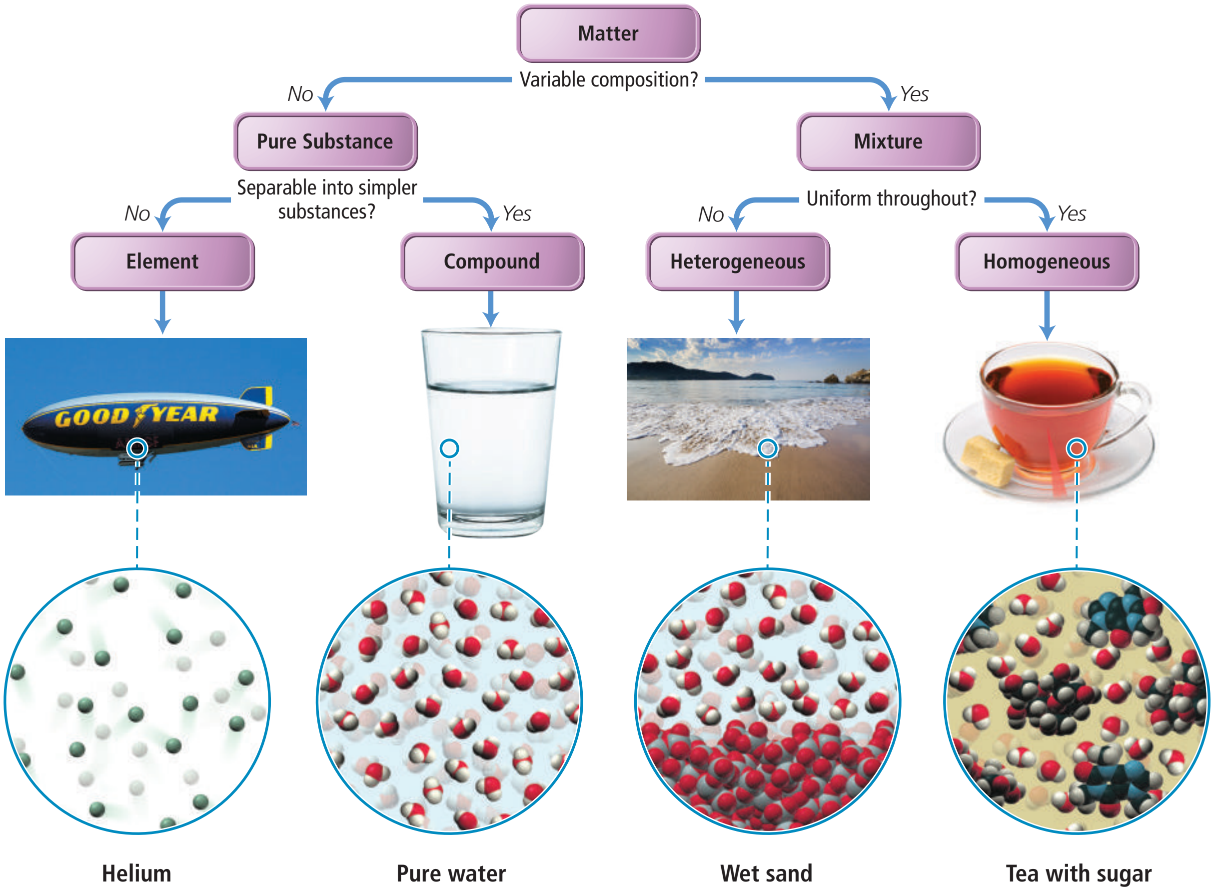
In *gaseous matter*, atoms or molecules have a lot of space between them and are free to move relative to one another, making gases *compressible* (Figure 1.3▶). When you squeeze a balloon or sit down on an air mattress, you force the atoms and molecules into a smaller space so that they are closer together. Gases always assume the shape *and* volume of their containers. Substances that are gases at room temperature include helium, nitrogen (the main component of air), and carbon dioxide.



▲ **FIGURE 1.3 The Compressibility of Gases** Gases can be compressed—squeezed into a smaller volume—because there is so much empty space between atoms or molecules in the gaseous state.

Classifying Matter by Composition: Elements, Compounds, and Mixtures

In addition to classifying matter according to its state, we classify it according to its composition, as shown in the following chart:



The first division in the classification of matter is between a *pure substance* and a *mixture*. A **pure substance** is made up of only one component, and its composition is invariant (it does not vary from one sample to another). The *components* of a pure substance can be individual atoms or groups of atoms joined together. For example, helium, water, and table salt (sodium chloride) are all pure substances. Each of these substances is made up of only one component: helium is made up of helium atoms, water is made up of water molecules, and sodium chloride is made up of sodium chloride units. The composition of a pure sample of any one of these substances is always exactly the same (because you can't vary the composition of a substance made up of only one component).

A **mixture**, by contrast, is composed of two or more components in proportions that can vary from one sample to another. For example, sweetened tea, composed primarily of water molecules and sugar molecules (with a few other substances mixed in), is a mixture. We can make tea slightly sweet (a small proportion of sugar to water) or very sweet (a large proportion of sugar to water) or any level of sweetness in between.

We categorize pure substances themselves into two types—*elements* and *compounds*—depending on whether or not they can be broken down (or decomposed) into simpler substances. Helium, which we just noted is a pure substance, is also a good example of an **element**, a substance that cannot be chemically broken down into simpler substances. Water, also a pure substance, is a good example of a **compound**, a substance composed of two or more elements (in this case, hydrogen and oxygen) in a fixed, definite proportion. On Earth, compounds are more common than pure elements because most elements combine with other elements to form compounds.

All known elements are listed in the periodic table in the inside front cover of this book.

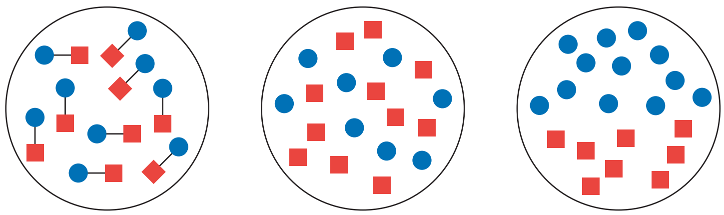
We also categorize mixtures into two types—heterogeneous and homogeneous—depending on how *uniformly* the substances within them mix. Wet sand is a **heterogeneous mixture**, one in which the composition varies from one region of the mixture to another. Sweetened tea is a **homogeneous mixture**, one with the same composition throughout. Homogeneous mixtures have uniform compositions because the atoms or molecules that compose them mix uniformly. Heterogeneous mixtures are made up of distinct regions because the atoms or molecules that compose them separate. Here again we see that the properties of matter are determined by the atoms or molecules that compose it.

Classifying a substance according to its composition is not always obvious and requires that we either know the true composition of the substance or are able to test it in a laboratory. For now, we focus on relatively common substances that you are likely to have encountered. Throughout this course, you will gain the knowledge to understand the composition of a larger variety of substances.

ANSWER **NOW!**

1.2 Cc Conceptual Connection

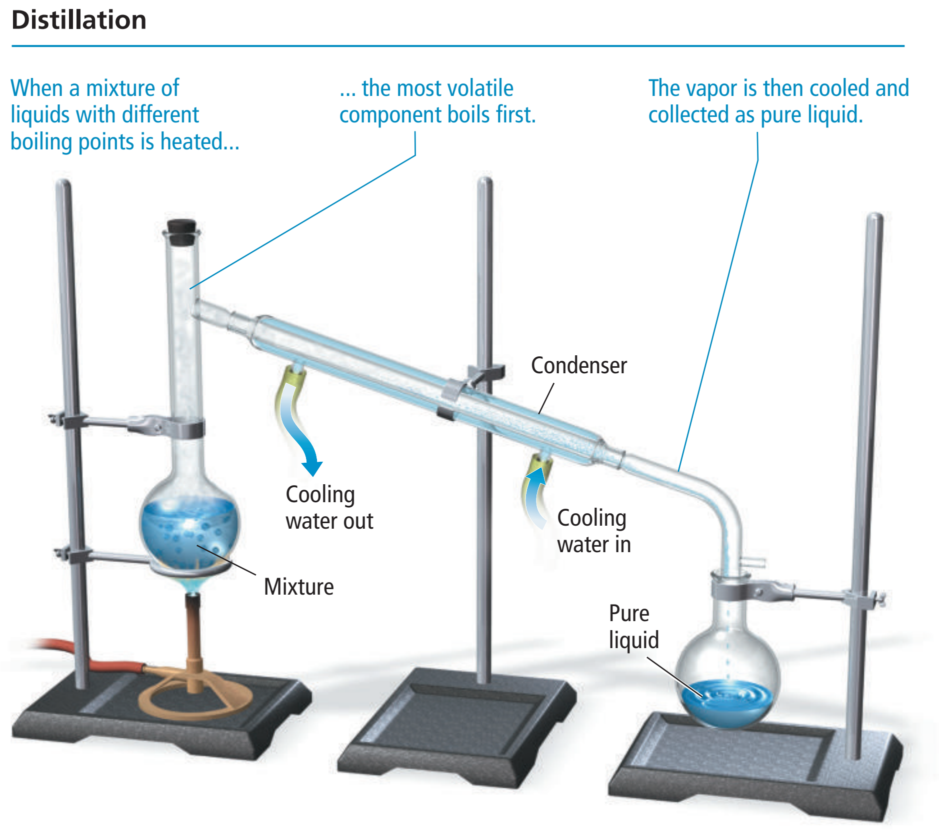
PURE SUBSTANCES AND MIXTURES In these images, a blue circle represents an atom of one type of element, and a red square represents an atom of a second type of element. Which image is a pure substance?



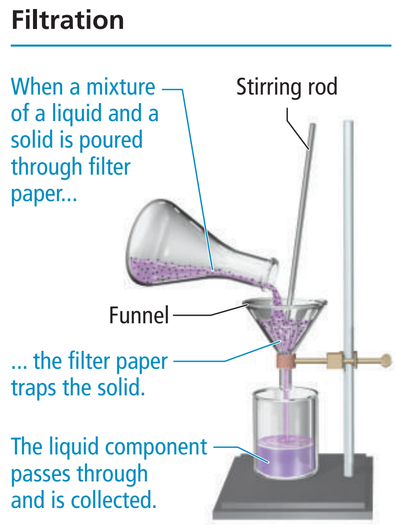
(a) **(b)** **(c)** None of these **(d)**

Separating Mixtures

Chemists often want to separate a mixture into its components. Such separations can be easy or difficult, depending on the components in the mixture. In general, mixtures are separable because the different components have different physical or chemical properties. We can use various techniques that exploit these differences to achieve

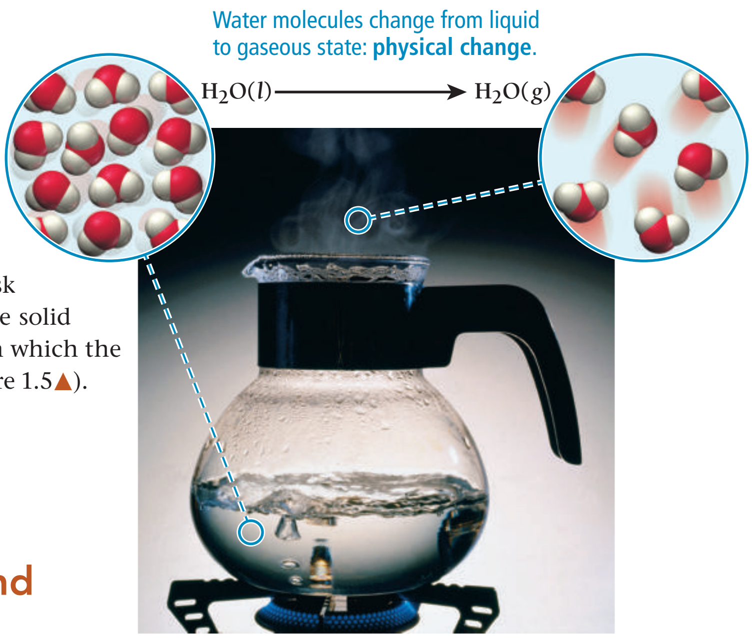


▲ **FIGURE 1.4 Separating Substances by Distillation**



▲ **FIGURE 1.5 Separating Substances by Filtration**

separation. For example, we can separate a mixture of sand and water by **decanting**—carefully pouring off—the water into another container. A homogeneous mixture of liquids can usually be separated by **distillation**, a process in which the mixture is heated to boil off the more **volatile** (easily vaporizable) liquid. The volatile liquid is then recondensed in a condenser and collected in a separate flask (Figure 1.4▲). If a mixture is composed of an insoluble solid and a liquid, we can separate the two by **filtration**, in which the mixture is poured through filter paper in a funnel (Figure 1.5▲).



▲ **FIGURE 1.6 Boiling, a Physical Change** When water boils, it turns into a gas but does not alter its chemical identity—the water molecules are the same in both the liquid and gaseous states. Boiling is a physical change, and the boiling point of water is a physical property.

1.4 Physical and Chemical Changes and Physical and Chemical Properties

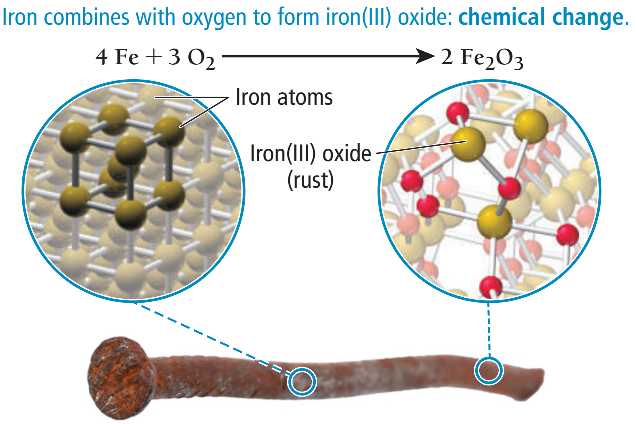
Every day we witness changes in matter: ice melts, iron rusts, gasoline burns, fruit ripens, and water evaporates. What happens to the molecules or atoms that compose these substances during such changes? The answer depends on the type of change. Changes that alter only state or appearance, but not composition, are **physical changes**. The atoms or molecules that compose a substance *do not change* their identity during a physical change. For example, when water boils, it changes its state from a liquid to a gas, but the gas remains composed of water molecules, so this is a physical change (Figure 1.6▲).

A physical change results in a different form of the same substance, while a chemical change results in a completely different substance.

In contrast, changes that alter the composition of matter are **chemical changes**. During a chemical change, atoms rearrange, transforming the original substances into different substances. For example, the rusting of iron is a chemical change. The atoms that compose iron (iron atoms) combine with oxygen molecules from air to form iron oxide, the orange substance we call rust (Figure 1.7◀). Figure 1.8▶ illustrates other examples of physical and chemical changes.

Physical and chemical changes are manifestations of physical and chemical properties. A **physical property** is a property that a substance displays without changing its composition, whereas a **chemical property** is a property that a substance displays only by changing its composition via a chemical change. The smell of gasoline is a physical property—gasoline does not change its composition when it exhibits its odor. The flammability of gasoline, in contrast, is a chemical property—gasoline does change its composition when it burns, turning into completely new substances (primarily carbon dioxide and water). Physical properties include odor, taste, color, appearance, melting point, boiling point, and density. Chemical properties include corrosiveness, flammability, acidity, toxicity, and other such characteristics.

The differences between physical and chemical changes are not always apparent. Only chemical examination can confirm whether a particular change is physical or chemical. In many cases, however, we can identify chemical and physical changes based on what we know about the changes. Changes in the state of matter, such as melting or boiling, or changes in the physical condition of matter, such as those that result from cutting or crushing, are typically physical changes. Changes involving chemical reactions—often evidenced by temperature or color changes—are chemical changes.



▲ **FIGURE 1.7 Rusting, a Chemical Change** When iron rusts, the iron atoms combine with oxygen atoms to form a different chemical substance, the compound iron(III) oxide. Rusting is a chemical change, and the tendency of iron to rust is a chemical property. A more detailed exploration of this reaction can be found in Section 20.9.

EXAMPLE 1.1 Physical and Chemical Changes and Properties

Determine whether each change is physical or chemical. What kind of property (chemical or physical) is demonstrated in each case?

(a) the evaporation of rubbing alcohol

(b) the burning of lamp oil

(c) the bleaching of hair with hydrogen peroxide

(d) the formation of frost on a cold night

SOLUTION

(a) When rubbing alcohol evaporates, it changes from liquid to gas, but it remains alcohol—this is a physical change. The volatility (the ability to evaporate easily) of alcohol is therefore a physical property.

(b) Lamp oil burns because it reacts with oxygen in air to form carbon dioxide and water—this is a chemical change. The flammability of lamp oil is therefore a chemical property.

(c) Applying hydrogen peroxide to hair changes pigment molecules in hair that give it color—this is a chemical change. The susceptibility of hair to bleaching is therefore a chemical property.

(d) Frost forms on a cold night because water vapor in air changes its state to form solid ice—this is a physical change. The temperature at which water freezes is therefore a physical property.

Answers to For Practice and For More Practice problems can be found in Appendix IV.

FOR PRACTICE 1.1 Determine whether each change is physical or chemical. What kind of property (chemical or physical) is demonstrated in each case?

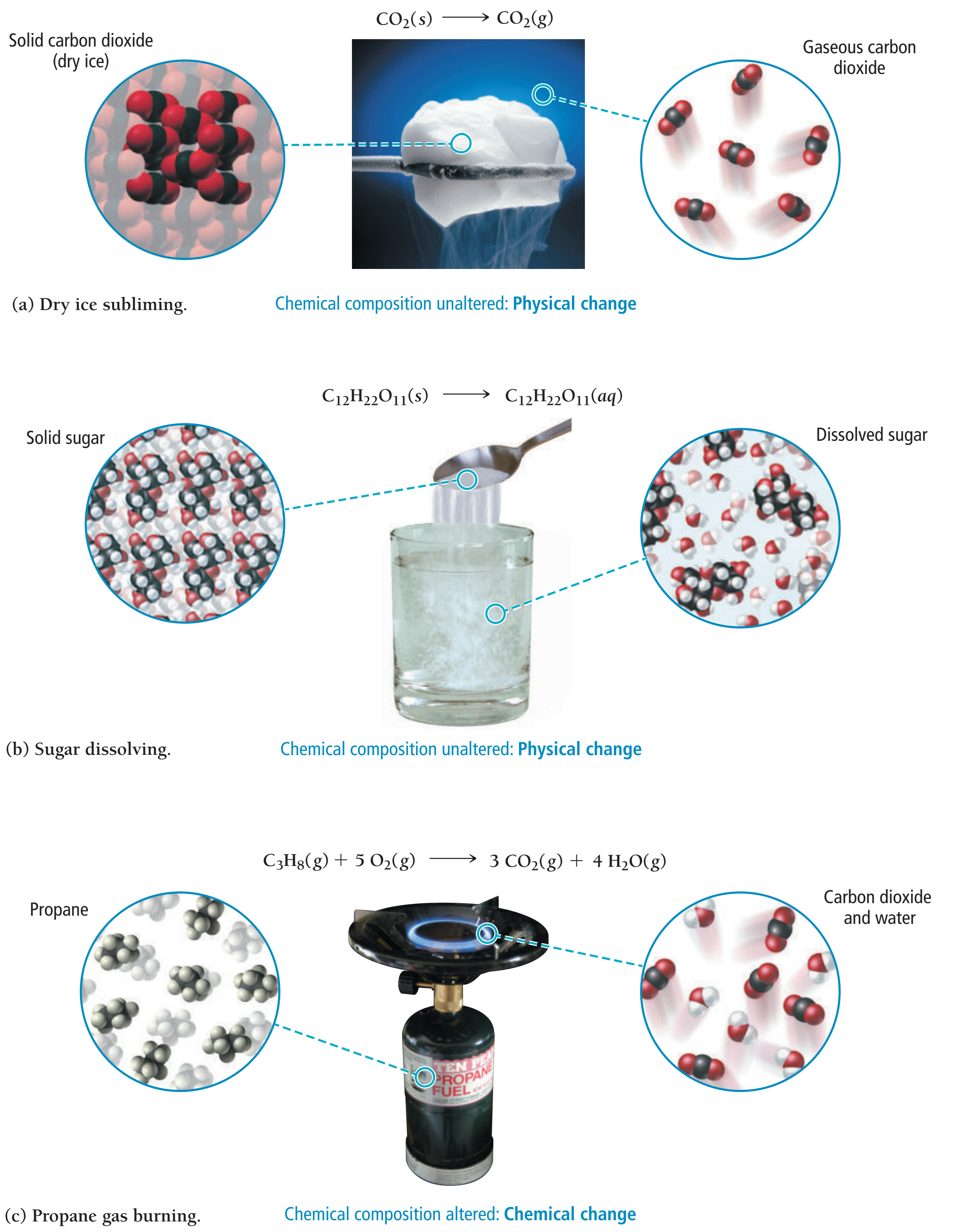
(a) A copper wire is hammered flat.

(b) A nickel dissolves in acid to form a blue-green solution.

(c) Dry ice sublimes without melting.

(d) A match ignites when struck on a flint.

Physical Change versus Chemical Change

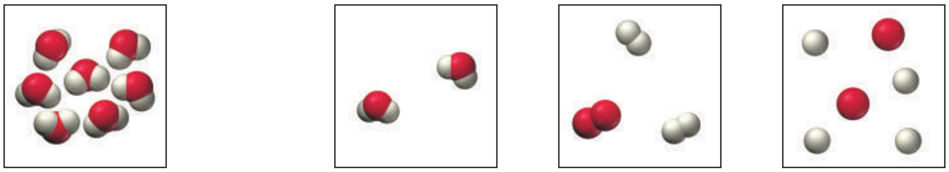


▲ **FIGURE 1.8 Physical and Chemical Changes** **(a)** The sublimation (the state change from a solid to a gas) of dry ice (solid CO_2) is a physical change. **(b)** The dissolution of sugar is a physical change. **(c)** The burning of propane is a chemical change.

ANSWER NOW!

1.3 Cc Conceptual Connection

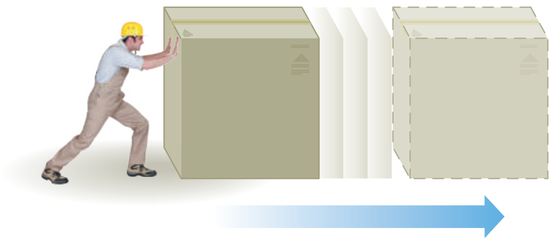
CHEMICAL AND PHYSICAL CHANGES The diagram on the left represents liquid water molecules in a pan. Which of the three diagrams (a, b, or c) best represents the water molecules after they have been vaporized by boiling?



1.5 Energy: A Fundamental Part of Physical and Chemical Change

The physical and chemical changes discussed in Section 1.4 are usually accompanied by energy changes. For example, when water evaporates from your skin (a physical change), the water molecules absorb energy from your body, making you feel cooler. When you burn natural gas on the stove (a chemical change), energy is released, heating the food you are cooking. Understanding the physical and chemical changes of matter—that is, understanding chemistry—requires that you understand energy changes and energy flow.

The scientific definition of **energy** is *the capacity to do work*. **Work** is defined as the action of a force through a distance. For instance, when you push a box across the floor or pedal your bicycle across the street, you have done work.



▲ **FIGURE 1.9 Energy Conversions** Gravitational potential energy is converted into kinetic energy when the weight is dropped. The kinetic energy is converted mostly to thermal energy when the weight strikes the ground.

The *total energy* of an object is a sum of its **kinetic energy** (the energy associated with its motion) and its **potential energy** (the energy associated with its position or composition). For example, a weight held several meters above the ground has potential energy due to its position within Earth's gravitational field (Figure 1.9 ◀). If you drop the weight, it accelerates, and its potential energy is converted to kinetic energy. When the weight hits the ground, its kinetic energy is converted primarily to **thermal energy**, the energy associated with the temperature of an object. Thermal energy is actually a type of kinetic energy because it is associated with the motion of the individual atoms or molecules that make up an object. When the weight hits the ground, its kinetic energy is essentially transferred to the atoms and molecules that compose the ground, raising the temperature of the ground ever so slightly.

The first principle to note about how energy changes as the weight falls to the ground is that *energy is neither created nor destroyed*. The potential energy of the weight becomes kinetic energy as the weight accelerates toward the ground. The kinetic energy then becomes thermal energy when the weight hits the ground. The total amount of thermal energy that is released through the process is exactly equal to the difference between the initial and final potential energy of the weight. The idea that energy is neither created nor destroyed is known as the **law of conservation of energy**. Although energy can change from one type into another, and although it can flow from one object to another, the *total quantity* of energy does not change—it remains constant.

In Chapter 21 we will discuss how energy conservation is actually part of a more general law that allows for the interconvertibility of mass and energy.

The second principle to note about the raised weight and its fall is *the tendency of systems with high potential energy to change in a way that lowers their potential energy*. For this reason, objects or systems with high potential energy tend to be *unstable*. The weight lifted several meters from the ground is unstable because it contains a significant amount of potential energy. Unless restrained, the weight will naturally fall, lowering its potential energy (due to its position in Earth's gravitational field). We can harness some of the raised weight's potential energy to do work. For example, we can attach the weight to a rope that turns a paddle wheel or spins a drill as the weight falls. After the weight falls to the ground, it contains less potential energy—it has become more *stable*.

▲ **FIGURE 1.10 Using Chemical Energy to Do Work** The compounds produced when gasoline burns have less chemical potential energy than the gasoline molecules.

Some chemical substances are like a raised weight. For example, the molecules that compose gasoline have a relatively high potential energy—energy is concentrated in them just as energy is concentrated in the raised weight. The molecules in the gasoline tend to undergo chemical changes (specifically combustion) that lower the molecules' potential energy. As the energy of the molecules is released, some of it can be harnessed to do work, such as moving a car forward (Figure 1.10▲). The molecules that result from the chemical change have less potential energy than the original molecules in gasoline and are more stable.

Chemical potential energy, such as the energy contained in the molecules that compose gasoline, arises primarily from electrostatic forces between the electrically charged particles (protons and electrons) that compose atoms and molecules. We will learn more about those particles, as well as the properties of electrical charge, in Chapter 2, but for now, know that molecules contain specific, usually complex, arrangements of these charged particles. Some of these arrangements—such as the one within the molecules that compose gasoline—have a much higher potential energy than others. When gasoline undergoes combustion, the arrangement of these particles changes, creating molecules with much lower potential energy and transferring a great deal of energy (mostly in the form of heat) to the surroundings.

Summarizing Energy:

- Energy is always conserved in a physical or chemical change; it is neither created nor destroyed.
- Systems with high potential energy tend to change in ways that lower their potential energy, transferring energy to the surroundings.

ENERGY What type of energy is chemical energy?

(a) kinetic energy
(b) thermal energy
(c) potential energy

1.4 Cc Conceptual Connection

ANSWER **NOW!**

1.6 The Units of Measurement

WATCH **NOW!**

KEY CONCEPT VIDEO 1.6
Units and Significant Figures

In 1999, NASA lost the $125 million *Mars Climate Orbiter*. The chairman of the commission that investigated the disaster concluded, "The root cause of the loss of the spacecraft was a failed translation of English units into metric units." As a result, the orbiter—which was supposed to monitor weather on Mars—descended too far into the Martian atmosphere and burned up. In chemistry as in space exploration, **units**—standard quantities used to specify measurements—are critical. If we get them wrong, the consequences can be disastrous.

▲ The $125 million *Mars Climate Orbiter* was lost in 1999 because two groups of engineers used different units.

TABLE 1.1	**SI Base Units**	
Quantity	**Unit**	**Symbol**
Length	Meter	m
Mass	Kilogram	kg
Time	Second	s
Temperature	Kelvin	K
Amount of substance	Mole	mol
Electric current	Ampere	A
Luminous intensity	Candela	cd

The two most common unit systems are the **metric system**, used in most of the world, and the **English system**, used in the United States. Scientists use the **International System of Units (SI)**, which is based on the metric system.

The abbreviation *SI* comes from the French, *Système International d'Unités*.

Standard Units

Table 1.1 shows the standard SI base units. In this chapter, we focus on the first four of these units: the *meter*, the standard unit of length; the *kilogram*, the standard unit of mass; the *second*, the standard unit of time; and the *kelvin*, the standard unit of temperature.

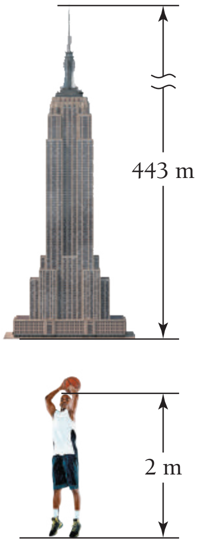
The Meter: A Measure of Length

The **meter (m)** is slightly longer than a yard (1 yard is 36 inches, while 1 meter is 39.37 inches).

Yardstick

Meterstick

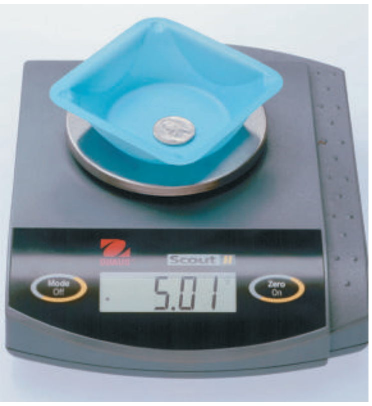
Thus, a 100-yard (yd) football field measures only 91.4 m. The meter was originally defined as 1/10,000,000 of the distance from the equator to the North Pole (through Paris). The International Bureau of Weights and Measures now defines it more precisely as the distance light travels through a vacuum in a certain period of time, 1/299,792,458 second. A tall human is about 2 m tall, and the Empire State Building stands 443 m tall (including its mast).



▲ The Empire State Building is 443 m tall. A basketball player stands about 2 m tall.

The Kilogram: A Measure of Mass

The **kilogram (kg)**, defined relative to a fundamental constant called planck's constant (*h*), is a measure of *mass*, a quantity different from *weight*. The **mass** of an object is a measure of the quantity of matter within it, while the weight of an object is a measure of the *gravitational pull* on its matter. If you could weigh yourself on the moon, for example, its weaker gravity would pull on you with less force than does Earth's gravity, resulting in a lower weight. A 130-pound (lb) person on Earth would weigh only 21.5 lb on the moon. However, a person's mass—the quantity of matter in his or her body—remains the same on every planet. One kilogram of mass is the equivalent of 2.205 lb of weight on Earth, so expressed in kilograms, a 130-lb person has a mass of approximately 59 kg and this book has a mass of about 2.5 kg. A second common unit of mass is the gram (g). One gram is 1/1000 kg. A nickel (5¢) has a mass of about 5 g.



▲ A nickel (5 cents) weighs about 5 g.

The Second: A Measure of Time

If you live in the United States, the **second (s)** is perhaps the SI unit most familiar to you. The International Bureau of Weights and Measures originally defined the second in terms of the day and the year, but a second is now defined more precisely as the duration of 9,192,631,770 periods of the radiation emitted from a certain transition in a cesium-133 atom. (We discuss transitions and the emission of radiation by atoms in Chapter 8.)

Scientists measure time on a large range of scales. The human heart beats about once every second, the age of the universe is estimated to be about 4.32×10^{17} s (13.7 billion years), and some molecular bonds break or form in time periods as short as 1×10^{-15} s.

The Kelvin: A Measure of Temperature

The **kelvin (K)** is the SI unit of **temperature**. The temperature of a sample of matter is a measure of the average kinetic energy—the energy due to motion—of the atoms or molecules that compose the matter. The molecules in a *hot* glass of water are, on average, moving faster than the molecules in a *cold* glass of water. Temperature is a measure of this molecular motion.

Temperature also determines the direction of thermal energy transfer, what we commonly call *heat*. Thermal energy transfers from hot objects to cold ones. For example, when you touch another person's warm hand (and yours is cold), thermal energy flows *from his or her hand to yours*, making your hand feel warmer. However, if you touch an ice cube, thermal energy flows *out of your hand* to the ice, cooling your hand (and possibly melting some of the ice cube).

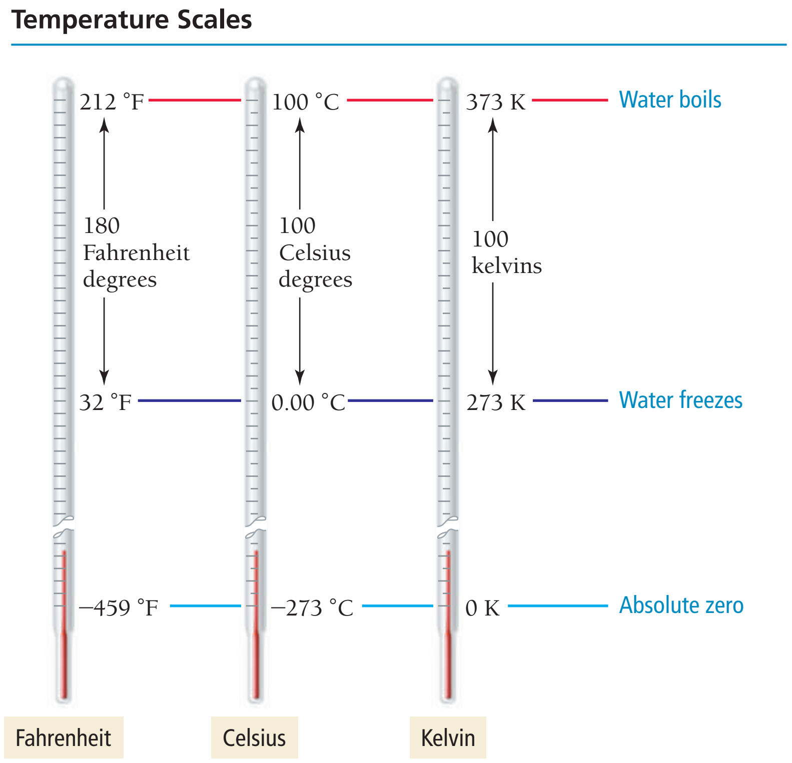
Figure 1.11▼ shows the three common temperature scales. The most common in the United States is the **Fahrenheit (°F) scale**, shown on the left in Figure 1.11. On the Fahrenheit scale, water freezes at 32 °F and boils at 212 °F (at sea level). Room temperature is approximately 72 °F. The Fahrenheit scale was originally determined by assigning 0 °F to the freezing point of a concentrated saltwater solution and 96 °F to normal body temperature.

Normal body temperature on the modern Fahrenheit scale is 98.6 °F.

Scientists and citizens of most countries other than the United States typically use the **Celsius (°C) scale**, shown in the middle in Figure 1.11. On this scale, pure water freezes at 0 °C and boils at 100 °C (at sea level). Room temperature is approximately 22 °C. The Fahrenheit scale and the Celsius scale differ both in the size of their respective degrees and the temperature each designates as "zero." Both the Fahrenheit and Celsius scales allow for negative temperatures.

The SI unit for temperature is the kelvin, shown in Figure 1.11. The **Kelvin scale** (sometimes also called the *absolute scale*) avoids negative temperatures by assigning 0 K to the coldest temperature possible, absolute zero. Absolute zero (−273.15 °C or −459 °F) is the temperature at which molecular motion virtually stops. Lower temperatures do not exist. The size of the kelvin is identical to that of the Celsius degree; the only difference is

Molecular motion does not *completely* stop at absolute zero because of the uncertainty principle in quantum mechanics, which we discuss in Chapter 8.



◀ **FIGURE 1.11 Comparison of the Fahrenheit, Celsius, and Kelvin Temperature Scales** The Fahrenheit degree is five-ninths the size of the Celsius degree and the kelvin. The zero point of the Kelvin scale is absolute zero (the lowest possible temperature), whereas the zero point of the Celsius scale is the freezing point of water.

The Celsius Temperature Scale

0 °C – Water freezes   10 °C – Brisk fall day   22 °C – Room temperature   45 °C – Summer day in Death Valley

Note that we give Kelvin temperatures in kelvins (*not* "degrees Kelvin") or K (*not* °K).

the temperature that each scale designates as zero. We can convert between the temperature scales with these formulas:

$$°C = \frac{(°F - 32)}{1.8}$$

$$K = °C + 273.15$$

ANSWER **NOW!**

1.5 Cc Conceptual Connection

TEMPERATURE SCALES Which temperature scale has no negative temperatures?

(a) Kelvin
(b) Celsius
(c) Fahrenheit

Throughout this book, we provide examples worked out in formats that are designed to help you develop problem-solving skills. The most common format uses two columns to guide you through the worked example. The left column describes the thought processes and steps used in solving the problem, while the right column shows the implementation. Example 1.2 follows this two-column format.

EXAMPLE 1.2 Converting between Temperature Scales

A sick child has a temperature of 40.00 °C. What is the child's temperature in **(a)** K and **(b)** °F?

SOLUTION

(a) Begin by finding the equation that relates the quantity that is given (°C) and the quantity you are trying to find (K).	$K = °C + 273.15$
Since this equation gives the temperature in K directly, substitute in the correct value for the temperature in °C and calculate the answer.	$K = °C + 273.15$ $K = 40.00 + 273.15 = 313.15\ K$
(b) To convert from °C to °F, first find the equation that relates these two quantities.	$°C = \frac{(°F - 32)}{1.8}$
Since this equation expresses °C in terms of °F, solve the equation for °F.	$°C = \frac{(°F - 32)}{1.8}$ $1.8(°C) = (°F - 32)$ $°F = 1.8(°C) + 32$
Now substitute °C into the equation and calculate the answer. *Note: The number of digits reported in this answer follows significant figure conventions, covered in Section 1.7.*	$°F = 1.8(°C) + 32$ $°F = 1.8(40.00\ °C) + 32 = 104.00\ °F$

FOR PRACTICE 1.2 Gallium is a solid metal at room temperature, but it will melt to a liquid in your hand. The melting point of gallium is 85.6 °F. What is this temperature on **(a)** the Celsius scale and **(b)** the Kelvin scale?

Prefix Multipliers

Scientific notation (see Appendix IA) allows us to express very large or very small quantities in a compact manner by using exponents. For example, the diameter of a hydrogen atom is 1.06×10^{-10} m. The International System of Units uses the **prefix multipliers** listed in Table 1.2 with the standard units. These multipliers change the value of the unit by powers of 10 (just like an exponent does in scientific notation). For example, the kilometer has the prefix *kilo* meaning 1000 or 10^3. Therefore,

$$1 \text{ kilometer} = 1000 \text{ meters} = 10^3 \text{ meters}$$

Similarly, the millimeter has the prefix *milli* meaning 0.001 or 10^{-3}.

$$1 \text{ millimeter} = 0.001 \text{ meters} = 10^{-3} \text{ meters}$$

TABLE 1.2 SI Prefix Multipliers

Prefix	Symbol	Multiplier	
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})
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})

When reporting a measurement, choose a prefix multiplier close to the size of the quantity you are measuring. For example, to state the diameter of a hydrogen atom, which is 1.06×10^{-10} m, use picometers (106 pm) or nanometers (0.106 nm) rather than micrometers or millimeters. Choose the prefix multiplier that is most convenient for a particular number.

PREFIX MULTIPLIERS Which prefix multiplier is most appropriate for reporting a measurement of 5.57×10^{-5} m?

(a) mega
(b) milli
(c) micro
(d) kilo

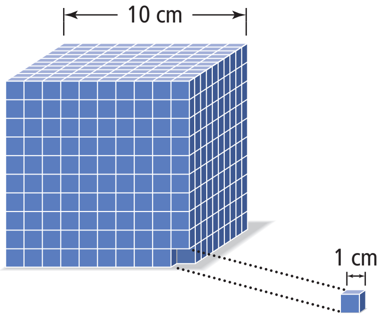
1.6 Cc Conceptual Connection

ANSWER **NOW!**

Derived Units: Volume and Density

A **derived unit** is a combination of other units. For example, the SI unit for speed is meters per second (m/s), a derived unit. Notice that this unit is formed from two other SI units—meters and seconds—put together. You are probably more familiar with speed measured in miles/hour or kilometers/hour—these are also examples of derived units. Two other common derived units are those for volume (SI base unit is m^3) and density (SI base unit is kg/m^3).

Relationship between Length and Volume



A 10-cm cube contains 1000 1-cm cubes.

▲ **FIGURE 1.12 The Relationship between Length and Volume** A cube with a 10-cm edge has a volume of $(10\ \text{cm})^3$ or $1000\ \text{cm}^3$, and a cube with a 100-cm edge has a volume of $(100\ \text{cm})^3 = 1{,}000{,}000\ \text{cm}^3$.

Volume

Volume is a measure of space. Any unit of length, when cubed (raised to the third power), becomes a unit of volume. The cubic meter (m^3), cubic centimeter (cm^3), and cubic millimeter (mm^3) are all units of volume. The cubic nature of volume is not always intuitive, and studies have shown that our brains are not naturally wired to process abstract concepts such as volume. For example, consider this question: how many small cubes measuring 1 cm on each side are required to construct a large cube measuring 10 cm on a side?

The answer to this question, as you can see by carefully examining the unit cube in Figure 1.12◀, is 1000 small cubes. When we go from a linear, one-dimensional distance to three-dimensional volume, we must raise both the linear dimension *and* its unit to the third power (not multiply by 3). Thus, the volume of a cube is equal to the length of its edge cubed.

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

Other common units of volume in chemistry are the **liter (L)** and the **milliliter (mL)**. One milliliter $(10^{-3}\ \text{L})$ is equal to $1\ \text{cm}^3$. A gallon of gasoline contains 3.785 L. Table 1.3 lists some common units—for volume and other quantities—and their equivalents.

TABLE 1.3 Some Common Units and Their Equivalents

Length	Mass	Volume
1 kilometer (km) = 0.6214 mile (mi)	1 kilogram (kg) = 2.205 pounds (lb)	1 liter (L) = 1000 mL = 1000 cm^3
1 meter (m) = 39.37 inches (in) = 1.094 yards (yd)	1 pound (lb) = 453.59 grams (g)	1 liter (L) = 1.057 quarts (qt)
1 foot (ft) = 30.48 centimeters (cm)	1 ounce (oz) = 28.35 grams (g)	1 U.S. gallon (gal) = 3.785 liters (L)
1 inch (in) = 2.54 centimeters (cm) (exact)		

Density

An old riddle asks, "Which weighs more, a ton of bricks or a ton of feathers?" The answer is neither—they both weigh the same (1 ton). If you answered bricks, you confused weight with density. The **density (*d*)** of a substance is the ratio of its mass (m) to its volume (V).

The m in the equation for density is in italic type, meaning that it stands for mass rather than for meters. In general, the symbols for units such as meters (m), seconds (s), or kelvins (K) appear in regular type while those for variables such as mass (m), volume (V), and time (t) appear in italics.

$$\text{Density} = \frac{\text{mass}}{\text{volume}} \quad \text{or} \quad d = \frac{m}{V}$$

Density is a characteristic physical property of substances (see Table 1.4) that depends on temperature. Density is an example of an **intensive property**, one that is *independent* of the amount of the substance. The density of aluminum, for example, is the same whether you have a gram or a kilogram. Intensive properties are often used to identify substances because these properties depend only on the type of substance, not on the amount of it. For example, from Table 1.4 you can see that pure gold has a density of $19.3\ \text{g/cm}^3$. One way to determine whether a substance is pure gold is to determine its density and compare it to $19.3\ \text{g/cm}^3$. Mass, in contrast, is an **extensive property**, one that depends on the amount of the substance. If you know only the mass of a sample of gold, that information alone will not allow you to identify it as gold.

The units of density are those of mass divided by volume. Although the SI-derived unit for density is kg/m^3, the density of liquids and solids is most often expressed in g/cm^3 or g/mL. (Remember that cm^3 and mL are equivalent units: $1\ \text{cm}^3 = 1\ \text{mL}$.) Aluminum is among the least dense metals with a density of $2.7\ \text{g/cm}^3$, while platinum is one of the densest metals with a density of $21.4\ \text{g/cm}^3$.

Calculating Density

We can calculate the density of a substance by dividing the mass of a given amount of the substance by its volume. For example, suppose a small nugget we suspect to be gold has a mass of 22.5 g and a volume of 2.38 cm^3. To find its density, we divide the mass by the volume:

$$d = \frac{m}{V} = \frac{22.5\text{ g}}{2.38\text{ cm}^3} = 9.45\text{ g/cm}^3$$

In this case, the density reveals that the nugget is not pure gold because the density of gold is 19.3 g/cm^3.

EXAMPLE 1.3 Calculating Density

A person receives a platinum ring from their partner. Before the wedding, they notice that the ring feels a little light for its size, and so they decide to determine its density. They place the ring on a balance and find that it has a mass of 3.15 g. They then find that the ring displaces 0.233 cm^3 of water. Is the ring made of platinum? (Note: The volume of irregularly shaped objects is often measured by the displacement of water. To use this method, the object is placed in water and the change in volume of the water is measured. This increase in the total volume represents the volume of water *displaced* by the object and is equal to the volume of the object.)

Set up the problem by writing the important information that is *given* as well as the information that you are asked to *find*. In this case, we need to find the density of the ring and compare it to that of platinum. *Note: In Section 1.8, we discuss this standardized way of setting up problems.*	**GIVEN:** $m = 3.15\text{ g}$ $V = 0.233\text{ cm}^3$ **FIND:** Density in g/cm^3
Next, write down the equation that defines density.	**EQUATION** $d = \frac{m}{V}$
Solve the problem by substituting the correct values of mass and volume into the expression for density.	**SOLUTION** $d = \frac{m}{V} = \frac{3.15\text{ g}}{0.233\text{ cm}^3} = 13.5\text{ g/cm}^3$

The density of the ring is much too low to be platinum (platinum's density is 21.4 g/cm^3), and the ring is therefore a fake.

FOR PRACTICE 1.3 The person in Example 1.3 is shocked that the ring is fake and returns it. They buy a new ring that has a mass of 4.53 g and a volume of 0.212 cm^3. Is the new ring genuine?

FOR MORE PRACTICE 1.3 A metal cube has an edge that is 11.4 mm long and a mass of 6.67 g. Calculate the density of the metal and use Table 1.4 to determine the likely identity of the metal.

TABLE 1.4 The Density of Some Common Substances at 20 °C

Substance	Density (g/cm^3)
Charcoal (from oak)	0.57
Ethanol	0.789
Ice	0.917 (at 0 °C)
Water	1.00 (at 4 °C)
Sugar (sucrose)	1.58
Table salt (sodium chloride)	2.16
Glass	2.6
Aluminum	2.70
Titanium	4.51
Iron	7.86
Copper	8.96
Lead	11.4
Mercury	13.55
Gold	19.3
Platinum	21.4

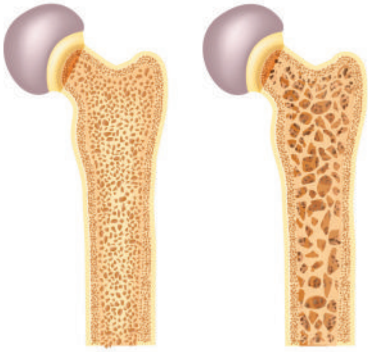
DENSITY The density of copper decreases as temperature increases (as does the density of most substances). Which change occurs in a sample of copper when it is warmed from room temperature to 95 °C?

(a) The sample becomes lighter.
(b) The sample becomes heavier.
(c) The sample expands.
(d) The sample contracts.

1.7 Cc Conceptual Connection

ANSWER NOW!

CHEMISTRY AND MEDICINE Bone Density



Osteoporosis—which means *porous bone*—is a condition in which bone density becomes low. The healthy bones of a young adult have a density of about 1.0 g/cm^3. Patients suffering from osteoporosis, however, can have bone densities as low as 0.22 g/cm^3. These low densities indicate that the bones have deteriorated and weakened, resulting in increased susceptibility to fractures, especially hip fractures. Patients suffering from osteoporosis can also experience height loss and disfiguration such as dowager's hump, a condition in which the patient becomes hunched over due to compression of the vertebrae. Osteoporosis is most common in postmenopausal women, but it can also occur in people who have certain diseases, such as insulin-dependent diabetes, or who take certain medications, such as prednisone. Osteoporosis is usually diagnosed and monitored with hip X-rays. Low-density bones absorb less of the X-rays than do high-density bones, producing characteristic differences in the X-ray image. Treatments for osteoporosis include calcium and vitamin D supplements, drugs that prevent bone weakening, exercise and strength training, and, in extreme cases, hip-replacement surgery.

QUESTION *Suppose you find a large animal bone in the woods, too large to fit in a beaker or flask. How might you approximate its density?*

◀ Top: Severe osteoporosis can necessitate surgery to implant an artificial hip joint, seen in this X-ray image. Bottom: Views of the bone matrix in a normal bone (left) and one weakened by osteoporosis (right).

1.7 The Reliability of a Measurement

Carbon monoxide is a colorless gas emitted by motor vehicles and found in polluted air. The table shown here lists carbon monoxide concentrations in Los Angeles County as reported by the U.S. Environmental Protection Agency (EPA) over the period 1990–2017:

Year	Carbon Monoxide Concentration (ppm)*
1990	9.9
2000	7.4
2010	1.9
2020	2.2

*Second maximum, 8-hour average; ppm = parts per million (Pasadena Site 06-037-2005)

The first thing you should notice about these values is that they decrease over time. For this decrease, we can thank the Clean Air Act and its amendments, which have resulted in more efficient engines and specially blended fuels and consequently in cleaner air in all major U.S. cities over the last 30 years. The second thing you might notice is the number of digits to which the measurements are reported. The number of digits in a reported measurement indicates the certainty associated with that measurement. A less certain measurement of carbon monoxide levels might be reported as follows:

Year	Carbon Monoxide Concentration (ppm)
1990	10
2000	7
2010	2
2020	2

Notice that the first set of data is reported to the nearest 0.1 ppm, while the second set is reported to the nearest 1 ppm. Scientists report measured quantities in an agreed-upon standard way. The number of reported digits reflects the certainty in the measurement: more digits, more certainty; fewer digits, less certainty. Numbers are usually written so that the uncertainty is in the last reported digit. (We assume that uncertainty to be ± 1 in the last digit unless otherwise indicated.) By reporting the 2010 carbon monoxide concentration as 1.9 ppm, the scientists mean 1.9 ± 0.1 ppm. The carbon monoxide concentration is between 1.8 and 2.0 ppm—it might be 2.0 ppm, for example, but it could not be 3.0 ppm. In contrast, if the reported value was 2 ppm (as in the second set of measurements), this would mean 2 ± 1 ppm, or between 1 and 3 ppm. In general,

Scientific measurements are reported so that every digit is certain except the last, which is estimated.

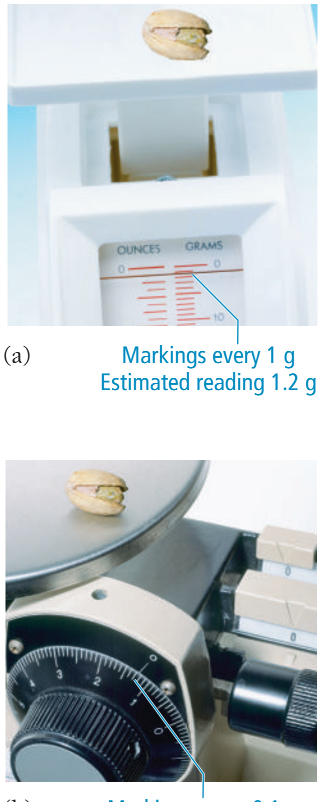
For example, consider the following reported number:
The first three digits are certain; the last digit is estimated.

5.213

certain   estimated

The number of digits reported in a measurement depends on the measuring device. Consider weighing a pistachio nut on two different balances (Figure 1.13▶). The balance on the top has marks every 1 g, while the balance on the bottom has marks every 0.1 g. For the balance on the top, we mentally divide the space between the 1- and 2-g marks into ten equal spaces and estimate that the pointer is at about 1.2 g. We then write the measurement as 1.2 g, indicating that we are sure of the "1" but we have estimated the ".2." The balance on the bottom, with marks every *tenth* of a gram, requires that we write the result with more digits. The pointer is between the 1.2-g mark and the 1.3-g mark. We again divide the space between the two marks into ten equal spaces and estimate the third digit. For the figure shown, we report 1.27 g.

Estimation in Weighing



(a) Markings every 1 g
Estimated reading 1.2 g

(b) Markings every 0.1 g
Estimated reading 1.27 g

▲ **FIGURE 1.13 Estimation in Weighing** **(a)** This scale has markings every 1 g, so we mentally divide the space into ten equal spaces to estimate the last digit. This reading is 1.2 g. **(b)** Because this balance has markings every 0.1 g, we estimate to the hundredths place. This reading is 1.27 g.

WATCH **NOW!**
INTERACTIVE WORKED EXAMPLE 1.4

EXAMPLE 1.4 Reporting the Correct Number of Digits

The graduated cylinder shown to the right has markings every 0.1 mL. Report the volume (which is read at the bottom of the meniscus) to the correct number of digits. (Note: The meniscus is the crescent-shaped surface at the top of a column of liquid.)

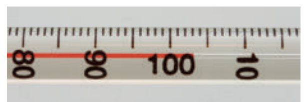
Meniscus

SOLUTION

Since the bottom of the meniscus is between the 4.5-mL and 4.6-mL markings, mentally divide the space between the markings into ten equal spaces and estimate the next digit. In this case, you should report the result as 4.57 mL.

What if you estimated a little differently and wrote 4.56 mL? In general, a one-unit difference in the last digit is acceptable because the last digit is estimated and different people might estimate it slightly differently. However, if you wrote 4.63 mL, you would have misreported the measurement.

FOR PRACTICE 1.4 Record the temperature on the thermometer shown at the right to the correct number of digits.



Counting Significant Figures

The precision of a measurement—which depends on the instrument used to make the measurement—must be preserved, not only when recording the measurement, but also when performing calculations that use the measurement. We can accomplish the preservation of this precision by using *significant figures*. In any reported measurement, the non–place-holding digits—those that are not simply marking the decimal place—are called **significant figures** (or **significant digits**). *The greater the number of significant figures, the greater the certainty of the measurement*. For example, the number 23.5 has three significant figures while the number 23.56 has four. To determine the number of significant figures in a number containing zeroes, we distinguish between zeroes that are significant and those that simply mark the decimal place. For example, in the number 0.0008, the leading zeroes (zeroes to the left of the first nonzero digit) mark the decimal place but *do not* add to the certainty of the measurement and are therefore not significant; this number has only one significant figure. In contrast, the trailing zeroes (zeroes at the end of a number) in the number 0.000800 *do add* to the certainty of the measurement and are therefore counted as significant; this number has three significant figures.

HOW TO: *Determine the Number of Significant Figures in a Given Value*

Significant Figure Rules	Examples	
1. All nonzero digits are significant.	28.03	0.0540
2. Interior zeroes (zeroes between two nonzero digits) are significant.	408	7.0301
3. Leading zeroes (zeroes to the left of the first nonzero digit) are not significant. They only serve to locate the decimal point.	0.0032	0.00006
	not significant	
4. Trailing zeroes (zeroes at the end of a number) are categorized as follows:		
• Trailing zeroes after a decimal point are always significant.	45.000	3.5600
• Trailing zeroes before a decimal point (and after a nonzero number) are always significant.	140.00	2500.55
• Trailing zeroes before an *implied* decimal point are ambiguous and should be avoided by using scientific notation.	1200	ambiguous
	1.2×10^3	2 significant figures
	1.20×10^3	3 significant figures
	1.200×10^3	4 significant figures
• Some textbooks put a decimal point after one or more trailing zeroes if the zeroes are to be considered significant. We avoid that practice in this book, but you should be aware of it.	1200.	4 significant figures (common in some textbooks)

Exact Numbers

Exact numbers have no uncertainty and thus do not limit the number of significant figures in any calculation. We can regard an exact number as having an unlimited number of significant figures. Exact numbers originate from three sources:

- From the accurate counting of discrete objects. For example, 3 atoms means 3.00000 . . . atoms.
- From defined quantities, such as the number of centimeters in 1 m. Because 100 cm is defined as 1 m,

$$100\text{ cm} = 1\text{ m means } 100.00000\ldots\text{ cm} = 1.0000000\ldots\text{ m}$$

- From integral numbers that are part of an equation. For example, in the equation $radius = \frac{diameter}{2}$, the number 2 is exact and therefore has an unlimited number of significant figures.

WATCH **NOW!**

INTERACTIVE WORKED EXAMPLE 1.5

EXAMPLE 1.5 Determining the Number of Significant Figures in a Number

How many significant figures are in each number?

(a) 0.04450 m **(b)** 5.0003 km **(c)** 10 dm = 1 m **(d)** 1.000×10^5 s **(e)** 0.00002 mm **(f)** 10,000 m

SOLUTION

(a) 0.04450 m	*Four significant figures*. The two 4s and the 5 are significant (Rule 1). The trailing zero is after a decimal point and is therefore significant (Rule 4). The leading zeroes only mark the decimal place and are therefore not significant (Rule 3).
(b) 5.0003 km	*Five significant figures*. The 5 and 3 are significant (Rule 1), as are the three interior zeroes (Rule 2).
(c) 10 dm = 1 m	*Unlimited significant figures*. Defined quantities have an unlimited number of significant figures.
(d) 1.000×10^5 s	*Four significant figures*. The 1 is significant (Rule 1). The trailing zeroes are after a decimal point and therefore significant (Rule 4).
(e) 0.00002 mm	*One significant figure*. The 2 is significant (Rule 1). The leading zeroes only mark the decimal place and are therefore not significant (Rule 3).
(f) 10,000 m	*Ambiguous*. The 1 is significant (Rule 1), but the trailing zeroes occur before an implied decimal point and are therefore ambiguous (Rule 4). Without more information, you would assume one significant figure. It is better to write this as 1×10^5 to indicate one significant figure or as 1.0000×10^5 to indicate five (Rule 4).

FOR PRACTICE 1.5 How many significant figures are in each number?

(a) 554 km **(b)** 7 pennies **(c)** 1.01×10^5 m **(d)** 0.00099 s **(e)** 1.4500 km **(f)** 21,000 m

Significant Figures in Calculations

When we use measured quantities in calculations, the results of the calculation must reflect the precision of the measured quantities. We should not lose or gain precision during mathematical operations. Follow these rules when carrying significant figures through calculations.

WATCH **NOW!**

KEY CONCEPT VIDEO 1.7

Significant Figures in Calculations

HOW TO: *Determine Significant Figures in Calculated Quantities*

Rules for Calculations

1. In multiplication or division, the result carries the same number of significant figures as the factor with the fewest significant figures.

2. In addition or subtraction, the result carries the same number of decimal places as the quantity with the fewest decimal places.

Examples

$$\underset{\text{(4 sig. figures)}}{1.052} \times \underset{\text{(5 sig. figures)}}{12.054} \times \underset{\text{(2 sig. figures)}}{0.53} = 6.7208 = \underset{\text{(2 sig. figures)}}{6.7}$$

$$\underset{\text{(5 sig. figures)}}{2.0035} \div \underset{\text{(3 sig. figures)}}{3.20} = 0.626094 = \underset{\text{(3 sig. figures)}}{0.626}$$

$$\begin{array}{r} 2.345 \\ 0.07 \\ 2.9975 \\ \hline 5.4125 \end{array} = 5.41 \qquad \begin{array}{r} 5.9 \\ -0.221 \\ \hline 5.679 \end{array} = 5.7$$

In addition and subtraction, it is helpful to draw a line next to the number with the fewest decimal places. This line determines the number of decimal places in the answer.

—Continued on the next page