Julia Burdge
Michelle Driessen

Introductory Chemistry

AN ATOMS FIRST APPROACH

McGraw Hill Education

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An Atoms First Approach

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INTRODUCTORY CHEMISTRY: AN ATOMS FIRST APPROACH

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To the people who will always matter the most: Katie, Beau, and Sam.

—Julia Burdge

To my family, the center of my universe and happiness, with special thanks to my husband for his support and making me the person I am today.

—Michelle Driessen

And to Robin Reed, for her timely and hilarious memes—and for her eternal good humor.

—Julia Burdge and Michelle Driessen

About the Authors

Julia Burdge holds a Ph.D. (1994) from The University of Idaho in Moscow, Idaho; and a Master's Degree from The University of South Florida. Her research interests have included synthesis and characterization of cisplatin analogues, and development of new analytical techniques and instrumentation for measuring ultra-trace levels of atmospheric sulfur compounds.

She currently holds an adjunct faculty position at The College of Western Idaho in Nampa, Idaho, where she teaches general chemistry using an atoms first approach; but spent the lion's share of her academic career at The University of Akron in Akron, Ohio, as director of the Introductory Chemistry program. In addition to directing the general chemistry program and supervising the teaching activities of graduate students, Julia established a future-faculty development program and served as a mentor for graduate students and post doctoral associates.

Julia relocated back to the Northwest to be near family. In her free time, she enjoys precious time with her three children, and with Erik Nelson, her partner and best friend.

Michelle Driessen earned a Ph.D. in 1997 from the University of Iowa in Iowa City, Iowa. Her research and dissertation focused on the thermal and photochemical reactions of small molecules at the surfaces of metal nanoparticles and high surface area oxides.

Following graduation, she held a tenure-track teaching and research position at Southwest Missouri State University for several years. A family move took her back to her home state of Minnesota where she held positions as adjunct faculty at both St. Cloud State University and the University of Minnesota. It was during these adjunct appointments that she became very interested in chemical education. Over the past several years she has transitioned the general chemistry laboratories at the University of Minnesota from verification to problem-based, and has developed both online and hybrid sections of general chemistry lecture courses. She is currently the Director of General Chemistry at the University of Minnesota where she runs the general chemistry laboratories, trains and supervises teaching assistants, and continues to experiment with active learning methods in her classroom.

Michelle and her husband love the outdoors and their rural roots. They take every opportunity to visit their family, farm, and horses in rural Minnesota.

Brief Contents

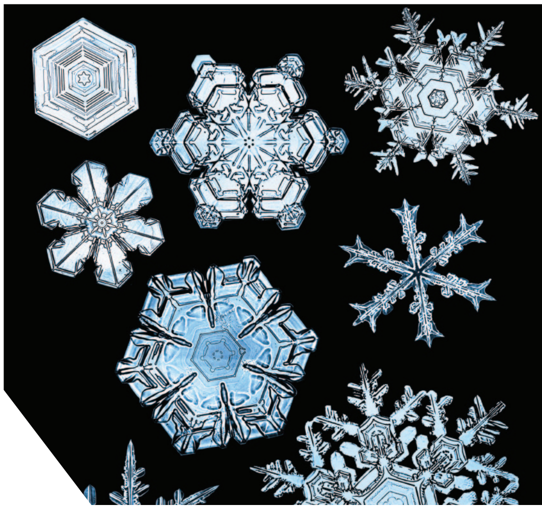
Contents

1 ATOMS AND ELEMENTS 2

2 ELECTRONS AND THE PERIODIC TABLE 30

3 COMPOUNDS AND CHEMICAL BONDS 74

6 MOLECULAR SHAPE 192

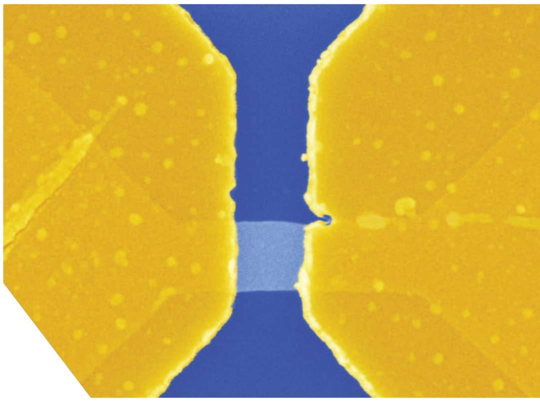


7 SOLIDS, LIQUIDS, AND PHASE CHANGES 234

Kool-Aid
DASANI
DROPS
MiO

13 EQUILIBRIUM 454

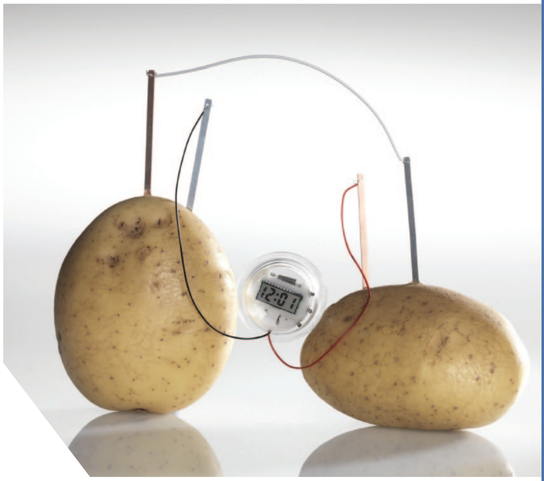
14 ORGANIC CHEMISTRY 480



15 BIOCHEMISTRY 506

16 NUCLEAR CHEMISTRY 522

17 ELECTROCHEMISTRY 538



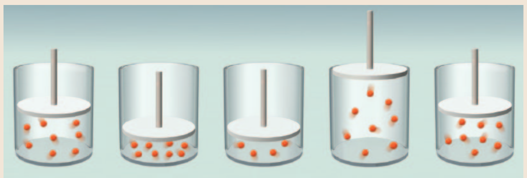
Preface

From its very origin, *Introductory Chemistry: An Atoms First Approach* by Julia Burdge and Michelle Driessen has been developed and written using an atoms first approach *specific* to introductory chemistry. It is not just a pared down version of a general chemistry text, but carefully crafted with the introductory-chemistry student in mind.

The ordering of topics facilitates the conceptual development of chemistry for the novice, rather than the historical development that has been used traditionally. Its language and style are student friendly and conversational; and the importance and wonder of chemistry in everyday life are emphasized at every opportunity. Continuing in the Burdge tradition, this text employs an outstanding art program, a consistent problem-solving approach, interesting applications woven throughout the chapters, and a wide range of end-of-chapter problems.

Features

- **Logical atoms first approach,** building first an understanding of atomic structure, followed by a logical progression of atomic properties, periodic trends, and how compounds arise as a consequence of atomic properties. Following that, physical and chemical properties of compounds and chemical reactions are covered—built upon a solid foundation of how all such properties and processes are the consequence of the nature and behavior of atoms.
- **Engaging real-life examples and applications.** Each chapter contains relevant, interesting stories in Familiar Chemistry segments that illustrate the importance of chemistry to other fields of study, and how the current material applies to everyday life. Many chapters also contain brief historical profiles of some important people in chemistry and other fields of scientific endeavor.
- **Consistent problem-solving skill development.** Fostering a consistent approach to problem solving helps students learn how to approach, analyze, and solve problems. Each worked example (Sample Problem) is divided into logical steps: Strategy, Setup, Solution, and Think About It; and each is followed by three practice problems. Practice Problem A allows the student to solve a problem similar to the Sample Problem, using the same strategy and steps. Wherever possible, Practice Problem B probes understanding of the same concept(s) as the Sample Problem and Practice Problem A, but is sufficiently different that it requires a slightly different approach. Practice Problem C often uses concept art or molecular models, and probes comprehension of underlying concepts. The consistent use of this approach gives students the best chance for developing a robust set of problem-solving skills.



SAMPLE PROBLEM 8.2 Using the Ideal Gas Equation to Calculate Volume

Calculate the volume of a mole of ideal gas at room temperature (25°C) and 1.00 atm.

Strategy Convert the temperature in °C to temperature in kelvins, and use the ideal gas equation to solve for the unknown volume.

Setup The data given are $n = 1.00$ mol, $T = 298$ K, and $P = 1.00$ atm. Because the pressure is expressed in atmospheres, we use $R = 0.0821$ L · atm/K · mol to solve for volume in liters.

Solution

$$V = \frac{(1\ \text{mol})\left(0.0821\dfrac{\text{L}\cdot\text{atm}}{\text{K}\cdot\text{mol}}\right)(298\ \text{K})}{1\ \text{atm}} = 24.5\ \text{L}$$

Student Note: It is a very common mistake to fail to convert to absolute temperature when solving a gas problem. Most often, temperatures are given in degrees Celsius. The ideal gas equation only works when the temperature used is in kelvins. Remember: K = °C + 273.

THINK ABOUT IT

With the pressure held constant, we should expect the volume to increase with increased temperature. Room temperature is higher than the standard temperature for gases (0°C), so the molar volume at room temperature (25°C) should be higher than the molar volume at 0°C—and it is.

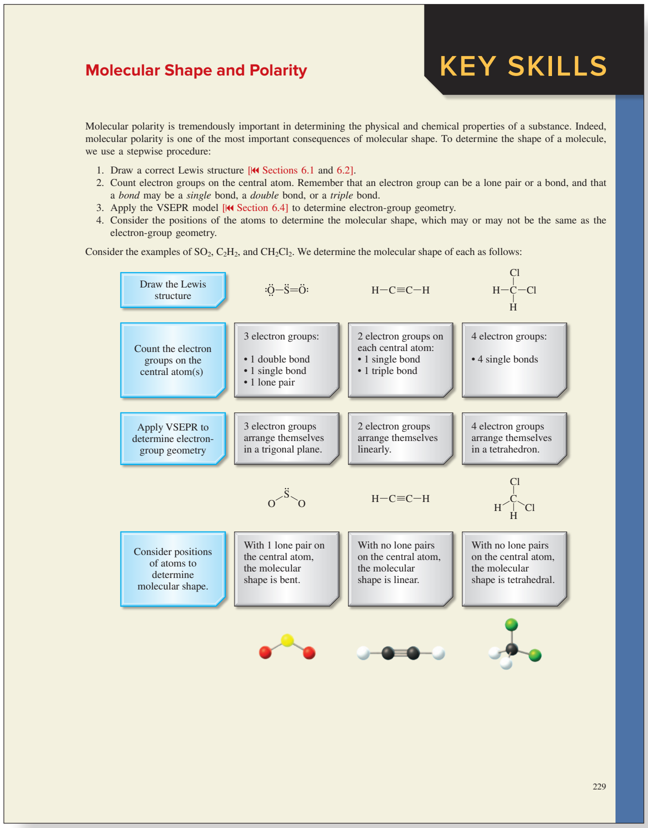
Practice Problem ATTEMPT What is the volume of 5.12 moles of an ideal gas at 32°C and 1.00 atm?

Practice Problem BUILD At what temperature (in °C) would 1 mole of ideal gas occupy 50.0 L ($P = 1.00$ atm)?

Practice Problem CONCEPTUALIZE The diagram on the left represents a sample of gas in a container with a movable piston. Which of the other diagrams [(i)–(iv)] best represents the sample (a) after the absolute temperature has been doubled; (b) after the volume has been decreased by half; and (c) after the external pressure has been doubled? (In each case, assume that the only variable that has changed is the one specified.)

(i) (ii) (iii) (iv)

- **Outstanding pedagogy for student learning.** The Checkpoints and Student Notes throughout each chapter are designed to foster frequent self-assessment and to provide timely information regarding common pitfalls, reminders of important information, and alternative approaches. Rewind and Fast Forward Buttons help to illustrate and reinforce connections between material in different chapters, and enable students to find pertinent review material easily, when necessary.
- **Key Skills pages** are reviews of specific skills that the authors know will be important to students' understanding of later chapters. These go beyond simple reviews and actually preview the importance of the skills in later chapters. They are additional opportunities for self-assessment and are meant to be revisited when the specific skills are required later in the book.



KEY SKILLS

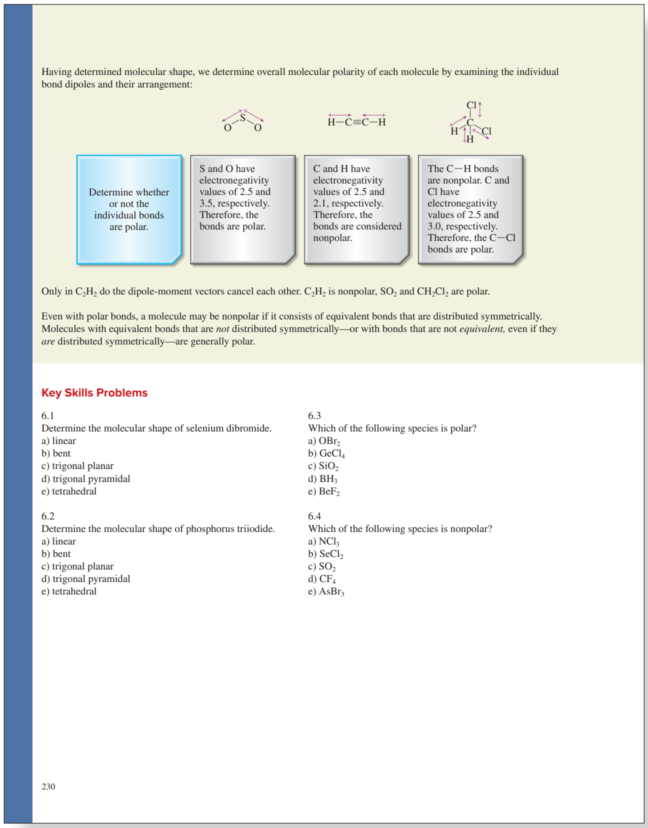
Molecular Shape and Polarity

Molecular polarity is tremendously important in determining the physical and chemical properties of a substance. Indeed, molecular polarity is one of the most important consequences of molecular shape. To determine the shape of a molecule, we use a stepwise procedure:

1. Draw a correct Lewis structure [⏮ Sections 6.1 and 6.2].
2. Count electron groups on the central atom. Remember that an electron group can be a lone pair or a bond, and that a *bond* may be a *single* bond, a *double* bond, or a *triple* bond.
3. Apply the VSEPR model [⏮ Section 6.4] to determine electron-group geometry.
4. Consider the positions of the atoms to determine the molecular shape, which may or may not be the same as the electron-group geometry.

Consider the examples of SO_2, C_2H_2, and CH_2Cl_2. We determine the molecular shape of each as follows:

229



Having determined molecular shape, we determine overall molecular polarity of each molecule by examining the individual bond dipoles and their arrangement:

Only in C_2H_2 do the dipole-moment vectors cancel each other. C_2H_2 is nonpolar, SO_2 and CH_2Cl_2 are polar.

Even with polar bonds, a molecule may be nonpolar if it consists of equivalent bonds that are distributed symmetrically. Molecules with equivalent bonds that are *not* distributed symmetrically—or with bonds that are not *equivalent,* even if they *are* distributed symmetrically—are generally polar.

Key Skills Problems

6.1
Determine the molecular shape of selenium dibromide.
a) linear
b) bent
c) trigonal planar
d) trigonal pyramidal
e) tetrahedral

6.2
Determine the molecular shape of phosphorus triiodide.
a) linear
b) bent
c) trigonal planar
d) trigonal pyramidal
e) tetrahedral

6.3
Which of the following species is polar?
a) OBr_2
b) $GeCl_4$
c) SiO_2
d) BH_3
e) BeF_2

6.4
Which of the following species is nonpolar?
a) NCl_3
b) $SeCl_2$
c) SO_2
d) CF_4
e) $AsBr_3$

230

- **Author-created online homework.** All of the online homework problems were developed entirely by co-author Michelle Driessen to ensure seamless integration with the book's content.

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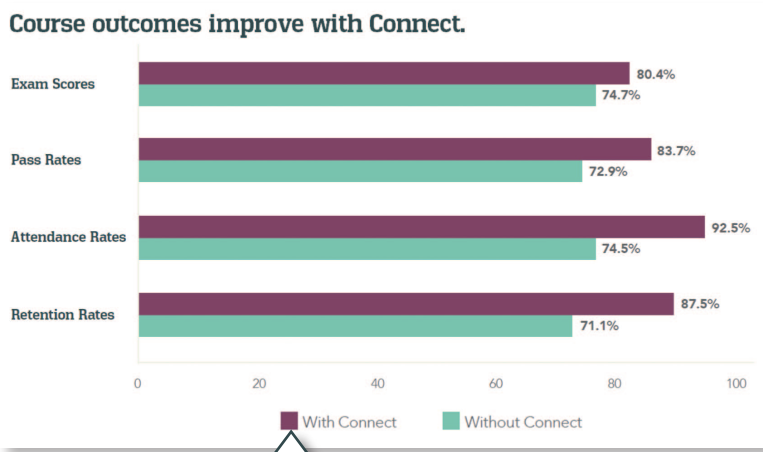

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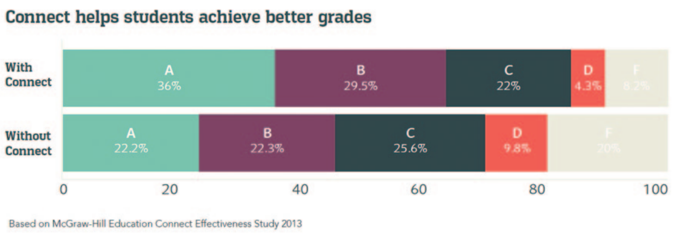

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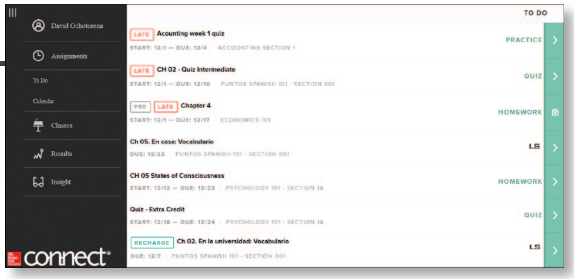

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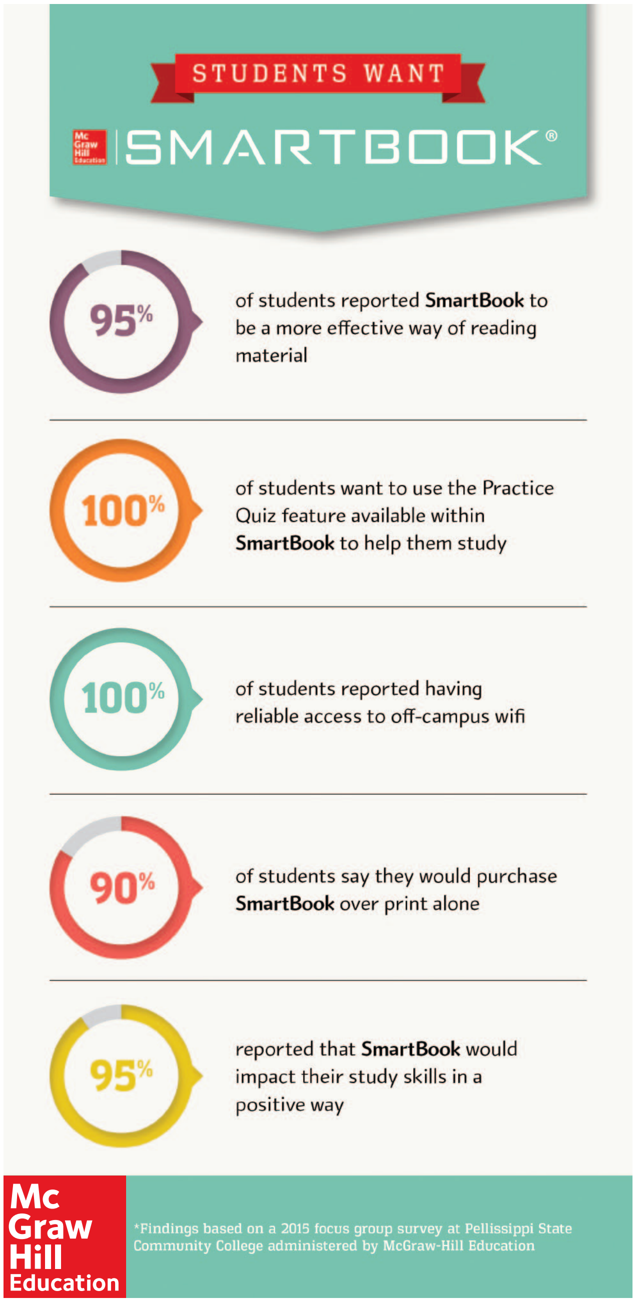

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- Lecture PowerPoint slides that facilitate classroom discussion of the concepts in the text
- Textbook images for repurposing in your personalized classroom materials
- A comprehensive bank of assignable test questions

Students can purchase a Student Solutions Manual that contains detailed solutions and explanations for the odd-numbered problems in the text.

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Julia Burdge and Michelle Driessen

Credits

CHAPTER 1

Atoms and Elements

The brilliant colors of a fireworks display result from the properties of the atoms they contain. These atoms give off specific colors when they are burned.

Credit: © Jung-Pang Wu/Getty Images

In This Chapter, You Will Learn

Some of what *chemistry* is and how it is studied using the *scientific method*. You will learn about *atomic structure* and you will become acquainted with the *periodic table*, how it is organized, and some of the information it embodies.

Things To Review Before You Begin

- Basic algebra

Have you ever wondered how an automobile airbag works? Or why iron rusts when exposed to water and air, but gold does not? Or why cookies "rise" as they bake? Or what causes the brilliant colors of fireworks displays? These phenomena, and countless others, can be explained by an understanding of the fundamental principles of *chemistry*. Whether or not we realize it, chemistry is important in every aspect of our lives. In the course of this book, you will come to understand the chemical principles responsible for many familiar observations and experiences.

1.1 The Study of Chemistry

Chemistry is the study of *matter* and the changes that matter undergoes. ***Matter,*** in turn, is anything that has mass and occupies space. ***Mass*** is one of the ways that scientists measure the *amount* of matter.

You may already be familiar with some of the terms used in chemistry—even if you have never taken a chemistry class. You have probably heard of *molecules*; and even if you don't know exactly what a *chemical formula* is, you undoubtedly know that "H_2O" is water. You may have used or at least heard the term *chemical reaction*; and you are certainly familiar with many processes that *are* chemical reactions.

Why Learn Chemistry?

Chances are good that you are using this book for a chemistry class you are required to take—even though you may not be a chemistry major. Chemistry is a required part of many degree programs because of its importance in a wide variety of scientific disciplines. It sometimes is called the "central science" because knowledge of chemistry supports the understanding of other scientific fields—including physics, biology, geology, ecology, oceanography, climatology, and medicine. Whether this is the first in a series of chemistry classes you will take or the only chemistry class you will ever take, we hope that it will help you to appreciate the beauty of chemistry—and to understand its importance in our daily lives.

The Scientific Method

Scientific experiments are the key to advancing our understanding of chemistry or any science. Although different scientists may take different approaches to experimentation, we all follow a set of guidelines known as the ***scientific method.*** This helps ensure the quality and integrity of new findings that are added to the body of knowledge within a given field.

The scientific method starts with the collection of data from careful observations and/or experiments. Scientists study the data and try to identify patterns. When a pattern is found, an attempt is made to describe it with a scientific ***law.*** In this context, a *law* is simply a concise statement of the observed pattern. Scientists may then formulate a ***hypothesis,*** an attempt to explain their observations. Experiments are then designed to *test* the hypothesis. If the experiments reveal that the hypothesis is incorrect, the scientists must go back to the drawing board and come up with a different interpretation of their data, and formulate a *new* hypothesis. The new hypothesis will then be tested by experiment. When a hypothesis stands the test of extensive experimentation, it may evolve into a *scientific* ***theory*** or ***model.*** A *theory* or *model* is a unifying principle that explains a body of experimental observations and the law or laws that are based on them. Theories are used both to explain past observations and to *predict* future observations. When a theory fails to predict correctly, it must be discarded or modified to become consistent with experimental observations. Thus, by their very nature, scientific theories must be subject to change in the face of new data that do not support them.

One of the most compelling examples of the scientific method is the development of the vaccine for *smallpox*, a viral disease responsible for an estimated half a *billion* deaths during the twentieth century alone. Late in the eighteenth century, English physician Edward Jenner observed that even during smallpox outbreaks in Europe, a particular group of people, *milkmaids*, seemed not to contract it.

Law: Milkmaids are not vulnerable to the virus that causes smallpox.

Based on his observations, Jenner proposed that perhaps milkmaids, who often contracted *cowpox*, a similar but far less deadly virus from the cows they worked with, had developed a natural immunity to smallpox.

Hypothesis: Exposure to the cowpox virus causes the development of immunity to the smallpox virus.

Jenner tested his hypothesis by injecting a healthy child with the cowpox virus—and later with the smallpox virus. If his hypothesis were correct, the child would not contract smallpox—and in fact the child did *not* contract smallpox.

Theory: Because the child did not develop smallpox, immunity seemed to have resulted from exposure to cowpox.

Further experiments on many more people (mostly children and prisoners) confirmed that exposure to the cowpox virus imparted immunity to the smallpox virus.

The flowchart in Figure 1.1 illustrates the scientific method and how it guided the development of the smallpox vaccine.

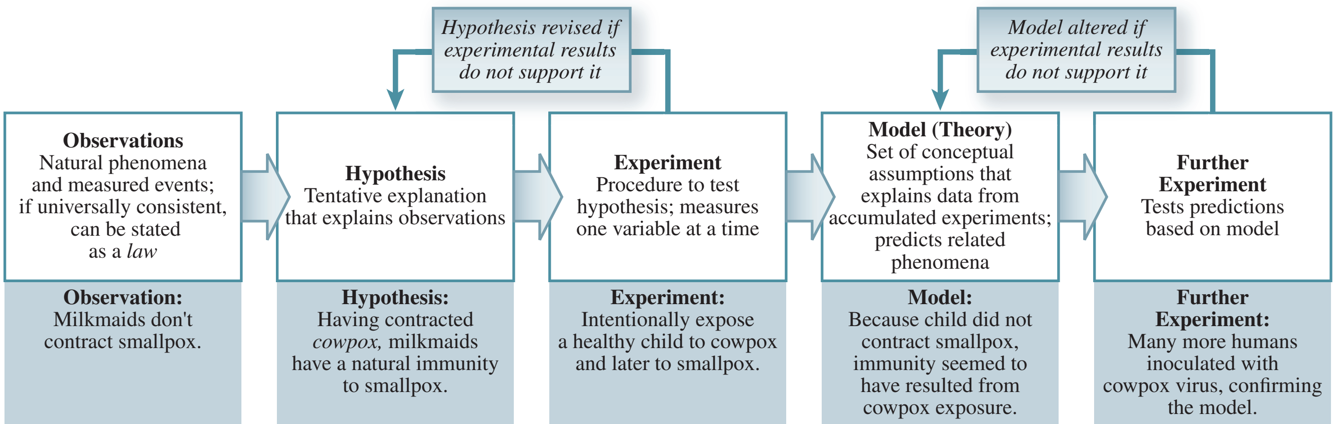

Figure 1.1 Flowchart of the scientific method and its importance to Edward Jenner's development of the smallpox vaccine.

1.2 Atoms First

Even if you have never studied chemistry before, you probably know already that atoms are the extraordinarily small building blocks that make up all matter. Specifically, an ***atom*** is the smallest quantity of matter that still retains the properties of matter. Further, an ***element*** is a substance that cannot be broken down into *simpler* substances by any means. Common examples of elements include aluminum, which we all have in our kitchens in the form of foil; carbon, which exists in several different familiar forms—including diamond and graphite (pencil "lead"); and helium, which can be used to fill balloons. The element aluminum consists entirely of *aluminum* atoms; the element carbon consists entirely of *carbon* atoms; and the element helium consists entirely of *helium* atoms. Although we can separate a sample of any element into smaller *samples* of that element, we cannot separate it into other substances.

Student Note: By contrast, consider a sample of salt water. We could divide it into smaller samples of salt water; but given the necessary equipment, we could also separate it into two different substances: water and salt. An element is different in that it is not made up of other substances. Elements are the *simplest* substances.

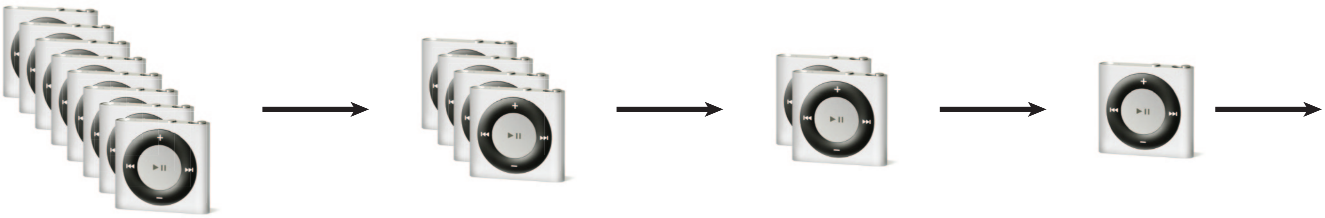

Figure 1.2 Repeatedly dividing this collection of iPods into smaller and smaller collections eventually leaves us with a single iPod, which we cannot divide further without destroying it.

Credit: © S K D/Alamy

Let's consider the example of helium. If we were to divide the helium in a balloon in half, and then divide one of the halves in half, and so on, we would eventually (after a very large number of these hypothetical divisions) be left with a sample of helium consisting of just one helium atom. This atom could not be further divided to give two smaller samples of helium. If this is difficult to imagine, think of a collection of eight identical iPods. We could divide the collection in half three times before we were left with a single iPod. Although we *could* divide the last iPod in half, neither of the resulting pieces would be an iPod! (Figure 1.2)

The notion that matter consists of tiny, indivisible pieces has been around for a very long time, first having been proposed by the philosopher Democritus in the fifth century B.C. But it was first formalized early in the nineteenth century by John Dalton (Figure 1.3). Dalton devised a theory to explain some of the most important observations made by scientists in the eighteenth century. His theory included three statements, the first of which is:

- Matter is composed of tiny, indivisible particles called atoms; all atoms of a given element are identical; and atoms of one element are different from atoms of any other element.

We will revisit this statement later in this chapter and introduce the second and third statements to complete our understanding of Dalton's theory in Chapters 3 and 10.

We know now that atoms, although very small, are not indivisible. Rather, they are made up of still smaller *sub*atomic particles. The type, number, and arrangement of subatomic particles determine the properties of atoms, which in turn determine the properties of everything we see, touch, smell, and taste.

Our goal in this book will be to understand how the nature of atoms gives rise to the properties of everything material. To accomplish this, we will take a somewhat unconventional approach. Rather than beginning with observations on the macroscopic scale and working our way backward to the atomic level of matter to explain these observations, we start by examining the structure of atoms, and the nature and arrangement of the tiny subatomic particles that atoms contain.

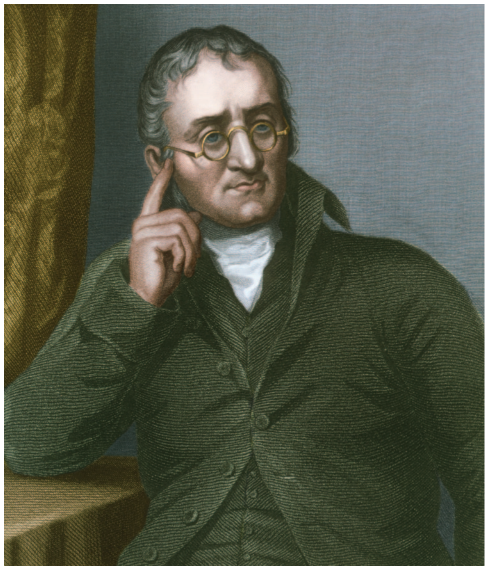

Figure 1.3 John Dalton (1766–1844) was an English chemist, mathematician, and philosopher. In addition to his atomic theory, Dalton also formulated several laws governing the behavior of gases, and gave the first detailed description of a particular type of color blindness, from which he suffered. This form of color blindness, where red and green cannot be distinguished, is known as Daltonism.

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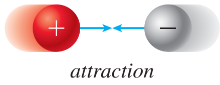
Before we begin our study of atoms, it is important for you to understand a bit about the behavior of electrically charged objects. We are all at least casually familiar with the concept of electric charge. You may have brushed your hair in very low humidity and had it stand on end; and you have certainly experienced static shocks and seen lightning. All of these phenomena result from the interactions of electric charges. The following list illustrates some of the important aspects of electric charge:

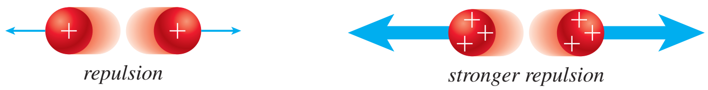
- An object that is electrically charged may have a positive (+) charge or a negative (−) charge.

- Objects with *opposite* charges (one negative and one positive) are attracted to each other. (You've heard the adage "opposites attract.")



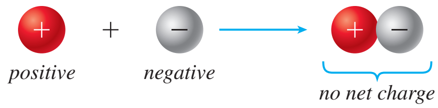
- Objects with *like* charges (either both positive or both negative) repel each other.

- Objects with larger charges interact more strongly than those with smaller charges.



- Charged objects interact more strongly when they are closer together.

- Opposite charges cancel each other.



Keeping in mind how charged objects interact will greatly facilitate your understanding of chemistry.

1.3 Subatomic Particles and the Nuclear Model of the Atom

Experiments conducted late in the nineteenth century indicated that atoms, which had been considered the smallest possible pieces of matter, contained even *smaller* particles. The first of these experiments were done by J. J. Thomson, an English physicist. The experiments revealed that a wide variety of different materials could all be made to emit a stream of tiny, negatively charged particles—that we now know as ***electrons.*** Thomson reasoned that because all atoms appeared to contain these negative particles but were themselves electrically *neutral*, they must also contain something *positively* charged. This gave rise to a model of the atom as a sphere of positive charge, throughout which negatively charged electrons were uniformly distributed (Figure 1.4). This model was known as the "plum-pudding" model—named after a then-popular English dessert. Thomson's plum-pudding model, which was generally accepted for a number of years, was an early attempt to describe the internal structure of atoms. Although it was generally accepted for a number of years, this model ultimately was proven wrong by subsequent experiments.

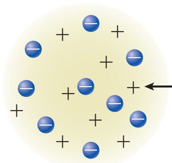


Figure 1.4 Thomson's experiments indicated that atoms contained negatively charged particles, which he envisioned as uniformly distributed in a sphere of positive charge.

Working with Thomson, New Zealand physicist Ernest Rutherford (one of Thomson's own students) devised an experiment to test the plum-pudding model of atomic structure. By that time, Rutherford had already established the existence of another subatomic particle known as an ***alpha particle,*** which is emitted by some *radioactive* substances. Alpha particles are positively charged, and are thousands of times more massive than electrons. In his most famous experiment, Rutherford directed a stream of alpha particles at a thin gold foil. A schematic of the experimental setup is shown in Figure 1.5. If Thomson's model of the atom were correct, nearly all of the alpha particles would pass directly through the foil—although a small number would be deflected slightly by virtue of passing very close to electrons. Rutherford surrounded the gold foil target with a detector that produced a tiny flash of light each time an alpha particle collided with it. This allowed Rutherford to determine the paths taken by alpha particles. Figure 1.6 illustrates the expected experimental result.

The actual experimental result was very different from what had been expected. Although most of the alpha particles did pass directly through the gold foil, some were deflected at much larger angles than had been anticipated. Some even bounced off the foil back toward the source—a result that Rutherford found absolutely shocking. He knew that alpha particles could only be deflected at such large angles, and occasionally bounce back in the direction of their source, if they encountered something within the gold atoms that was (1) positively charged, and (2) much larger than themselves. Figure 1.7 illustrates the actual result of Rutherford's experiment.

This experimental result gave rise to a new model of the internal structure of atoms. Rutherford proposed that atoms are mostly empty space, but that each has a tiny, dense core that contains *all* of its positive charge and *nearly* all of its mass. This core is called the atomic ***nucleus.***

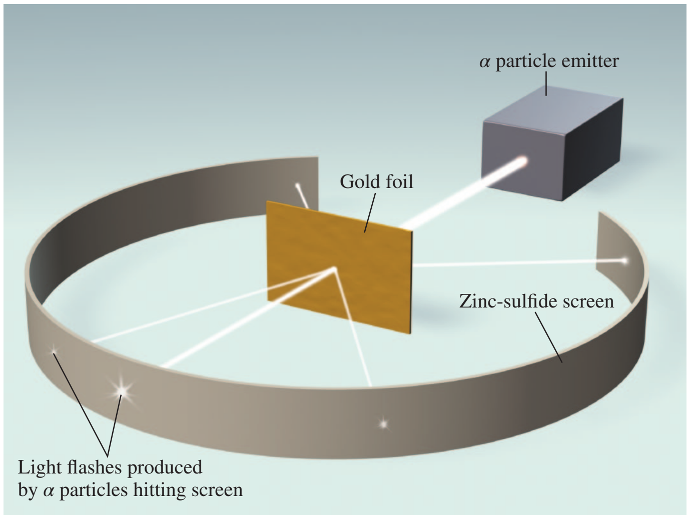


Figure 1.5 Rutherford's experiment directed a stream of positively charged alpha particles at a gold foil. The nearly circular detector emitted a flash of light when struck by an alpha particle.

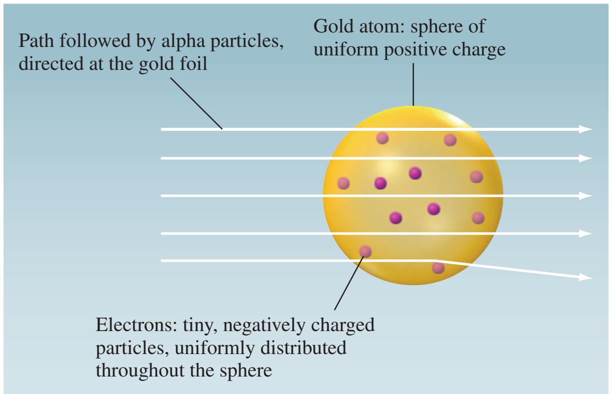


Figure 1.6 Rutherford's gold foil experiment was designed to test Thomson's plum-pudding model of the atom, which depicted the atom as negatively charged electrons uniformly distributed in a sphere of positive charge. If the model had been correct, the alpha particles would have passed directly through the foil, with a few being deflected slightly by interaction with electrons. (Remember that a positively charged object and a negatively charged object are attracted to each other. A positively charged alpha particle could be pulled slightly off course if it passed very close to one of the negatively charged electrons.)

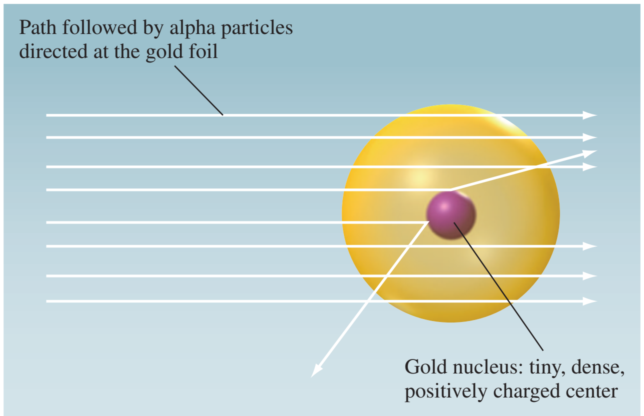


Figure 1.7 The actual result of Rutherford's gold foil experiment. Positively charged alpha particles were directed at a gold foil. Most passed through undeflected, but a few were deflected at angles much greater than expected—some even bounced back toward the source. This indicated that as they passed through the gold atoms, they encountered something positively charged and significantly more massive than themselves.

Student Note: An alpha particle is the combination of *two* protons and *two* neutrons.

Subsequent experiments supported Rutherford's nuclear model of the atom; and we now know that all atomic *nuclei* (the plural of *nucleus*) contain positively charged particles called ***protons.*** And with the exception of *hydrogen*, the lightest element, atomic nuclei also contain electrically *neutral* particles called ***neutrons.*** Together, the protons and neutrons in an atom account for nearly all of its mass, but only a tiny fraction of its volume. The nucleus is surrounded by a "cloud" of electrons—and just as Rutherford proposed, atoms are mostly empty space. Figure 1.8 illustrates the nuclear model of the atom.

Of the three subatomic particles in our model of the atom, the electron is the smallest and lightest. Protons and neutrons have very similar masses, and each is nearly

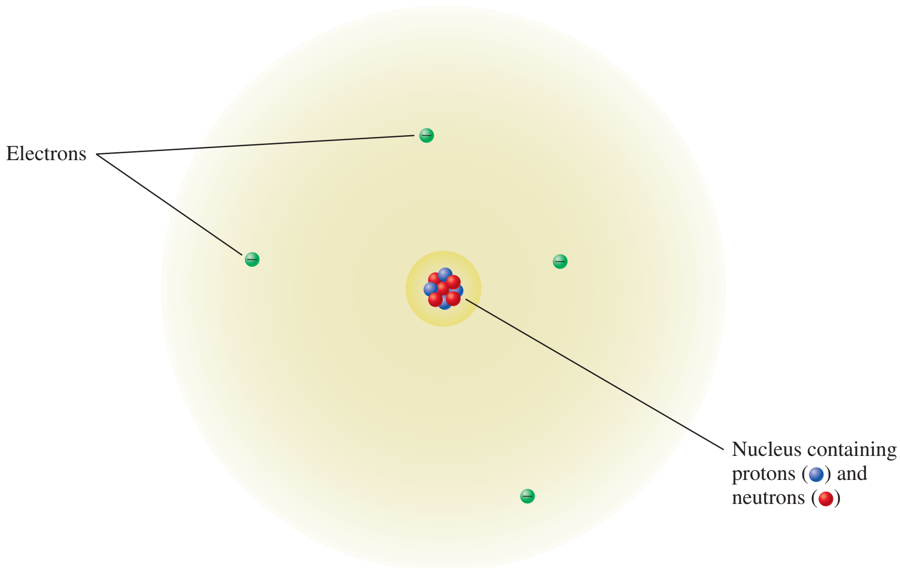


Figure 1.8 Nuclear model of the atom. Protons (blue) and neutrons (red) are contained within the nucleus, a tiny space at the center of the atom. The rest of the volume of the atom is nearly empty, but is occupied by the atom's electrons. This illustration exaggerates the size of the nucleus relative to the size of the atom. If the picture were actually done to scale, and the nucleus were the size shown here (1 centimeter), the atom would be on the order of 100 meters across—about the length of a football field.

2000 times as heavy as an electron. Further, because protons are positively charged and electrons are negatively charged, combination of equal numbers of each results in complete cancellation of the charges. The number of electrons is equal to the number of protons in a neutral atom. Because neutrons are electrically neutral, they do not contribute to an atom's overall charge.

Sample Problem 1.1 lets you practice identifying which combinations of subatomic particles constitute a neutral atom.

SAMPLE PROBLEM 1.1 Identifying a Neutral Atom Using Numbers of Subatomic Particles

The following table contains data sets that indicate numbers of subatomic particles. Which of the sets of data represent neutral atoms? For those that do not represent neutral atoms, determine what the charge is—based on the numbers of subatomic particles.

	neutrons	protons	electrons
(a)	5	10	5
(b)	11	12	12
(c)	8	9	9
(d)	20	21	20

Strategy You have learned that the charge on a proton is +1 and the charge on an electron is −1. Neutrons have no charge. The overall charge is the sum of charges of the protons and electrons, and a neutral atom has no charge. Therefore, a set of data in which the number of protons is equal to the number of electrons represents a neutral atom.

Setup Data sets (b) and (c) each contain equal numbers of protons and electrons. Data sets (a) and (d) do not.

Solution The data in sets (b) and (c) represent neutral atoms. Those in (a) and (d) represent charges species. The charge on the species represented by data set (a) is +5: 10 protons (+1 each) and 5 electrons (−1 each). The charge on the species represented by data set (d) is +1: 21 protons (+1 each) and 20 electrons (−1 each).

THINK ABOUT IT

By summing the charges of protons and electrons, we can determine the overall charge on a species. Note that the number of neutrons is not a factor in determining overall charge because neutrons have no charge.

Practice Problem ATTEMPT Which of the following data sets represent neutral atoms? For those that do not represent neutral atoms, determine the charge.

	neutrons	protons	electrons
(a)	31	31	30
(b)	24	22	24
(c)	12	11	11
(d)	6	5	5

Practice Problem BUILD Fill in the appropriate missing numbers in the following table:

	overall charge	protons	electrons
(a)	+2	23	
(b)	−3		42
(c)	0	53	
(d)		16	18

Practice Problem CONCEPTUALIZE Determine which of the following pictures represents a neutral atom. For any that does not represent a neutral atom, determine the overall charge. (Protons are blue, neutrons are red, and electrons are green.)