CHAPTER

2

Atoms and Elements

These observations have tacitly led to the conclusion which seems universally adopted, that all bodies of sensible magnitude . . . are constituted of a vast number of extremely small particles, or atoms of matter . . . .
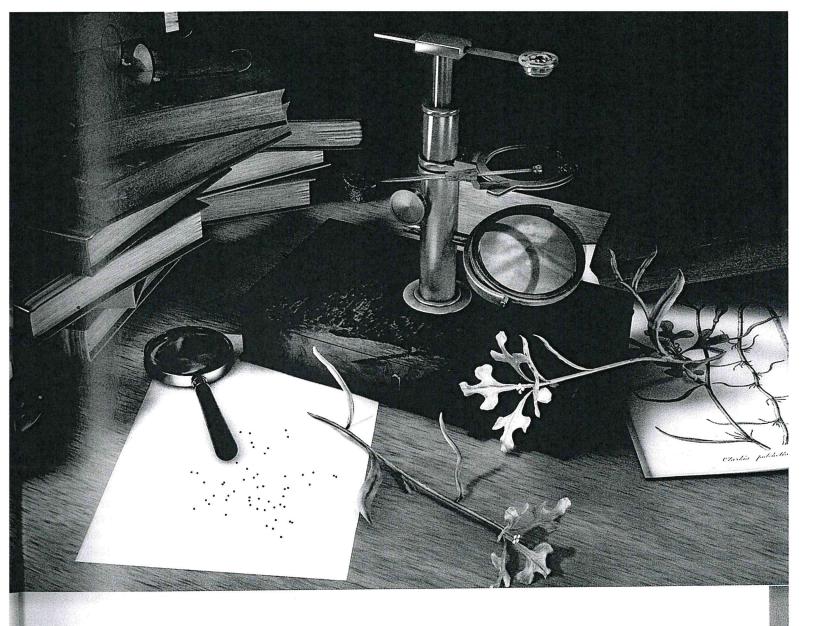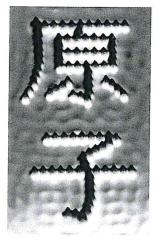
—JOHN DALTON (1766–1844)

If you cut a piece of graphite from the tip of a pencil into smaller and smaller pieces, how far could you go? Could you divide it forever? Would you eventually run into some basic particles that were no longer divisible, not because of their sheer smallness, but because of the nature of matter? This fundamental question about the nature of matter has been asked by thinkers for over two millennia. Their answers have varied over time. On the scale of everyday objects, matter appears continuous, or infinitely divisible. And until about 200 years ago, many scientists thought that matter was indeed continuous—but they were proven wrong. If you were to divide the graphite from your pencil tip into smaller and smaller pieces (far smaller than the eye could see), you would eventually end up with individual carbon atoms. The word *atom* comes from the Greek *atomos*, meaning "indivisible." You cannot divide a carbon atom into smaller pieces and still have carbon. Atoms compose all ordinary matter—if you want to understand matter, you must begin by understanding atoms.

▲ Scottish botanist Robert Brown observed random motion in tiny particles suspended in water. This motion later confirmed the particulate nature of matter.

2.1 Brownian Motion: Atoms Comfirmed

In 1827, Scottish botanist Robert Brown (1773–1858) looked through his microscope at water-suspended particles that had come from pollen grains. He noticed that the pollen particles were in continuous motion. Any one particle traveled a random, jittery path

through the liquid water. Brown initially thought that the particles might be alive and were perhaps the male sexual cells of plants (similar to sperm). However, similar particles from plants long dead exhibited the same jittery motion, and so did the dust from pulverized stones. Brown concluded that the source of the motion must not come from the particles themselves. What was causing this motion?

The definitive answer to this question did not come until 1905, when Albert Einstein (1879–1955) developed a theory that quantitatively explained what was by then called *Brownian motion*. Einstein's model explained the motion Brown had observed as a result of the molecular bombardments of the particles due to the thermal energy of the surrounding water. In other words, the water molecules in liquid water—constantly in motion due to thermal energy—were continuously battering the pollen and dust particles, causing them to oscillate and move. In Einstein's model, the jittering pollen particles are like a beach ball that is thrown into a crowd at a graduation ceremony. As the eager graduates strike the ball over and over again, the ball moves through the crowd in a jittery random path. The difference is that in the case of Brownian motion, the "crowd" is composed of molecules much too small to see.

In 1908, French physicist Jean Perrin (1870–1942) conducted experimental measurements to test Einstein's model. His measurements confirmed that Einstein's model was valid. In 1926, Perrin was awarded the Nobel Prize in Physics. During the award speech, the presenter said, "the object of the researches of Professor Jean Perrin which have gained for him the Nobel Prize in Physics for 1926 was to put a definite end to the long struggle regarding the real existence of molecules." In other words, the work of Einstein, and then Perrin, removed any lingering doubt about the particulate nature of matter.

In Einstein's day, the existence of atoms was inferred from the jittery motion first witnessed by Brown. Today, with a type of microscope called a scanning tunneling microscope (STM), we can form images of atoms themselves. In fact, STM can be used to *pick up and move* individual atoms, allowing structures and patterns to be made one atom at a time. Figure 2.1◀, for example, shows the Kanji characters for the word "atom" written with individual iron atoms on top of a copper surface. If all of the words in the books in the Library of Congress—35 million books occupying 840 miles of shelves—were written in letters the size of these Kanji characters, they would fit into an area of about five square millimeters. Scientists at IBM have also succeeded in making a short video entitled *A Boy and His Atom*, in which the main character (a boy) is animated using a few dozen atoms. In the video, which has been viewed millions of times on YouTube, the boy plays with an atom like a real boy would play with a ball.

As we discussed in Chapter 1, it was only 200 years ago that John Dalton proposed his atomic theory, and about 100 years ago that the theory was confirmed through the work of Einstein and Perrin. Yet today we can image atoms, move them, and even build tiny machines out of just a few dozen atoms (an area of research called nanotechnology). These atomic machines, and the atoms that compose them, are almost unimaginably small. To get an idea of the size of an atom, imagine picking up a grain of sand at a beach. That grain contains more atoms than you could count in a lifetime. In fact, the number of atoms in one sand grain far exceeds the number of grains on an entire beach.

In spite of their small size, atoms are the key to connecting the macroscopic and microscopic worlds. An *atom* is the smallest identifiable unit of an *element*. There are about 91 different naturally occurring elements. In addition, scientists have succeeded in making over 20 synthetic elements (elements not found in nature). In this chapter, we learn about atoms: what they are made of, how they differ from one another, and how they are structured. We also learn about the elements that are composed of these different kinds of atoms and about some of their characteristic properties. We will also discuss how the elements can be organized in a way that reveals patterns in their properties and helps us to understand what underlies those properties.

▲ **FIGURE 2.1 Imaging Atoms** The Japanese Kanji characters for "atom" written with iron atoms (red) on a copper surface (blue). The copper atoms are not as distinct as the iron atoms, but they appear as blue ripples in the background.

The exact number of naturally occurring elements is controversial because some elements that were first discovered when they were synthesized are believed to also be present in trace amounts in nature.

2.2 Early Ideas about the Building Blocks of Matter

The first people to propose that matter was composed of small, indestructible particles were Leucippus (fifth century B.C., exact dates unknown) and his student Democritus (460–370 B.C.). These Greek philosophers theorized that matter is ultimately composed of small, indivisible particles they named *atomos*. Democritus wrote, "Nothing exists except atoms and empty space; everything else is opinion." Leucippus and Democritus proposed that many different kinds of atoms exist, each different in shape and size, and that they move randomly through empty space. Other influential Greek thinkers of the time, such as Plato and Aristotle, did not embrace the atomic ideas of Leucippus and Democritus. Instead, Plato and Aristotle held that matter had no smallest parts and that different substances were composed of various proportions of fire, air, earth, and water. Since there was no experimental way to test the relative merits of the competing ideas, Aristotle's view prevailed, largely because he was so influential. The idea that matter is composed of atoms took a back seat in intellectual thought for nearly 2000 years.

In the sixteenth century, modern science began to emerge. A greater emphasis on observation led Nicolaus Copernicus (1473–1543) to publish *On the Revolution of the Heavenly Orbs* in 1543. The publication of that book—which proposed that the sun, not Earth, is at the center of the universe—marks the beginning of what we now call the *scientific revolution*. The next 200 years—and the work of scientists such as Francis Bacon (1561–1626), Johannes Kepler (1571–1630), Galileo Galilei (1564–1642), Robert Boyle (1627–1691), and Isaac Newton (1642–1727)—brought rapid advancement as the scientific approach became the established way to learn about the physical world. By the early 1800s, certain observations led the English chemist John Dalton (1766–1844) to offer convincing evidence that supported the early atomic ideas of Leucippus and Democritus. However, debate continued about whether atoms actually exist until the description of Brownian motion by Einstein in 1905 and subsequent experimental verification of the description in 1908 by Perrin (see Section 2.1).

2.3 Modern Atomic Theory and the Laws That Led to It

Key Atomic Theory

Recall the discussion of the scientific approach to knowledge from Chapter 1. The *atomic theory* (the idea that all matter is composed of atoms) grew out of observations and laws. The three most important laws that led to the development and acceptance of the atomic theory are the law of conservation of mass, the law of definite proportions, and the law of multiple proportions.

The Law of Conservation of Mass

In 1789, as we saw in Chapter 1, Antoine Lavoisier formulated the **law of conservation of mass**, which states:
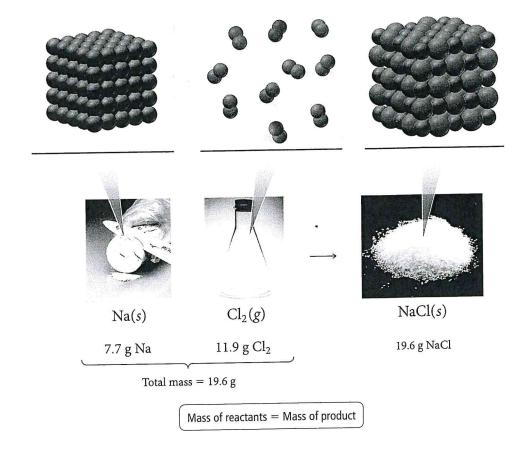
In a chemical reaction, matter is neither created nor destroyed.

In other words, when a chemical reaction occurs, the total mass of the substances involved in the reaction does not change. For example, consider the reaction between sodium and chlorine to form sodium chloride. The combined mass of the sodium and chlorine that react (the *reactants*) exactly equals the mass of the sodium chloride that forms (the *product*). This law is consistent with the idea that matter is composed of small, indestructible particles. The particles rearrange during a chemical reaction, but the amount of matter is conserved because the particles themselves are indestructible (at least by chemical means).

We will see in Chapter 20 that the law of conservation of mass is a slight oversimplification. However, the changes in mass in ordinary chemical processes are so minute that they can be ignored for all practical purposes.

Na(s)   Cl₂(g)   NaCl(s)

7.7 g Na   11.9 g Cl₂   19.6 g NaCl

Total mass = 19.6 g

Mass of reactants = Mass of product

CONCEPTUAL
CONNECTION **2.1**

The Law of Conservation of Mass   When a log completely burns in a campfire, the mass of the ash is much less than the mass of the log. What happens to the matter that composed the log?

The Law of Definite Proportions

In 1797, the French chemist Joseph Proust (1754–1826) made observations on the composition of compounds. He found that the elements composing a given compound always occur in fixed (or definite) proportions in all samples of the compound. In contrast, the components of a mixture can be present in any proportions whatsoever. Proust summarized his observations in the **law of definite proportions**:

The law of definite proportions is sometimes called the law of constant composition.

> **All samples of a given compound, regardless of their source or how they were prepared, have the same proportions of their constituent elements.**

For example, the decomposition of 18.0 g of water results in 16.0 g of oxygen and 2.0 g of hydrogen, or an oxygen-to-hydrogen mass ratio of:

$$\text{Mass ratio} = \frac{16.0\,\text{g O}}{2.0\,\text{g H}} = 8.0 \text{ or } 8{:}1$$

This ratio holds for any sample of pure water, regardless of its origin. The law of definite proportions applies to every compound. Consider ammonia, a compound composed of nitrogen and hydrogen. Ammonia contains 14.0 g of nitrogen for every 3.0 g of hydrogen, resulting in a nitrogen-to-hydrogen mass ratio of 4.7.

$$\text{Mass ratio} = \frac{14.0\,\text{g N}}{3.0\,\text{g H}} = 4.7 \text{ or } 4.7{:}1$$

Again, this ratio is the same for every sample of ammonia. The law of definite proportions also hints at the idea that matter is composed of atoms. Compounds have definite proportions of their constituent elements because the atoms that compose them, each with its own specific mass, occur in a definite ratio. Since the ratio of atoms is the same for all samples of a particular compound, the ratio of masses is also the same.

Example 2.1   Law of Definite Proportions

Two samples of carbon dioxide are decomposed into their constituent elements. One sample produces 25.6 g of oxygen and 9.60 g of carbon, and the other produces 21.6 g of oxygen and 8.10 g of carbon. Show that these results are consistent with the law of definite proportions.

SOLUTION

To show this, for both samples calculate the mass ratio of one element to the other by dividing the mass of one element by the mass of the other. For convenience, divide the larger mass by the smaller one.	For the first sample: $$\frac{\text{Mass oxygen}}{\text{Mass carbon}} = \frac{25.6}{9.60} = 2.67 \text{ or } 2.67{:}1$$ For the second sample: $$\frac{\text{Mass oxygen}}{\text{Mass carbon}} = \frac{21.6}{8.10} = 2.67 \text{ or } 2.67{:}1$$

The ratios are the same for the two samples, so these results are consistent with the law of definite proportions.

FOR PRACTICE 2.1   Two samples of carbon monoxide are decomposed into their constituent elements. One sample produces 17.2 g of oxygen and 12.9 g of carbon, and the other sample produces 10.5 g of oxygen and 7.88 g of carbon. Show that these results are consistent with the law of definite proportions.

The Law of Multiple Proportions

In 1804, John Dalton published his **law of multiple proportions**:

> **When two elements (call them A and B) form two different compounds, the masses of element B that combine with 1 g of element A can be expressed as a ratio of small whole numbers.**

Dalton suspected that matter was composed of atoms, so that when two elements A and B combine to form more than one compound, an atom of A combines with either one, two, three, or more atoms of B (AB_1, AB_2, AB_3, etc.). Therefore, the masses of B that react with a fixed mass of A are always related to one another as small whole-number ratios. Consider the compounds carbon monoxide and carbon dioxide. Carbon monoxide and carbon dioxide are two compounds composed of the same two elements: carbon and oxygen. We saw in Example 2.1 that the mass ratio of oxygen to carbon in carbon dioxide is 2.67:1; therefore, 2.67 g of oxygen reacts with 1 g of carbon. In carbon monoxide, however, the mass ratio of oxygen to carbon is 1.33:1, or 1.33 g of oxygen to every 1 g of carbon.

Carbon dioxide   Mass oxygen that combines with 1 g carbon = 2.67 g

Carbon monoxide   Mass oxygen that combines with 1 g carbon = 1.33 g

The ratio of these two masses is itself a small whole number.

$$\frac{\text{Mass oxygen to 1 g carbon in carbon dioxide}}{\text{Mass oxygen to 1 g carbon in carbon monoxide}} = \frac{2.67}{1.33} = 2$$

With the help of the molecular models in the margin, we can see why the ratio is 2:1—carbon dioxide contains two oxygen atoms to every carbon atom, while carbon monoxide contains only one. Of course, neither John Dalton nor Joseph Proust had access to any kind of modern instrumentation that could detect individual atoms—Dalton supported his atomic ideas primarily by using the *masses* of samples.

Example 2.2   Law of Multiple Proportions

Nitrogen forms several compounds with oxygen, including nitrogen dioxide and dinitrogen monoxide. Nitrogen dioxide contains 2.28 g oxygen to every 1.00 g nitrogen, while dinitrogen monoxide contains 0.570 g oxygen to every 1.00 g nitrogen. Show that these results are consistent with the law of multiple proportions.

SOLUTION

Calculate the ratio of the mass of oxygen from one compound to the mass of oxygen in the other. Always divide the larger of the two masses by the smaller one.	$\dfrac{\text{Mass oxygen to 1 g nitrogen in nitrogen dioxide}}{\text{Mass oxygen to 1 g nitrogen in dinitrogen monoxide}} = \dfrac{2.28}{0.570} = 4.00$

The ratio is a small whole number (4); these results are consistent with the law of multiple proportions.

FOR PRACTICE 2.2   Hydrogen and oxygen form both water and hydrogen peroxide. The decomposition of a sample of water forms 0.125 g hydrogen to every 1.00 g oxygen. The decomposition of a sample of hydrogen peroxide forms 0.0625 g hydrogen to every 1.00 g oxygen. Show that these results are consistent with the law of multiple proportions.

John Dalton and the Atomic Theory

In 1808, John Dalton explained the laws we just discussed with his **atomic theory**:

In Section 2.6, we will see that, contrary to Dalton's theory, all atoms of a given element do not have exactly the same mass.

1. Each element is composed of tiny, indestructible particles called atoms.
2. All atoms of a given element have the same mass and other properties that distinguish them from the atoms of other elements.
3. Atoms combine in simple, whole-number ratios to form compounds.
4. Atoms of one element cannot change into atoms of another element. In a chemical reaction, atoms only change the way they are *bound together* with other atoms.

Today, the evidence for the atomic theory is overwhelming. Matter is indeed composed of atoms.

Chemistry IN YOUR DAY | Atoms and Humans

You and I are composed of atoms. We get those atoms from the food we eat. Yesterday's cheeseburger contributes to today's skin, muscle, and hair. Not only are we made of atoms, but we are made of *recycled* atoms. The carbon atoms that compose our bodies were used by other living organisms before we got them. And they will be used by still others when we are done with them. In fact, it is likely that at this moment, your body contains some (over one trillion*) carbon atoms that were at one time part of your chemistry professor.

The idea that humans are composed of atoms acting in accord with the laws of chemistry and physics has significant implications and raises important questions. If atoms compose our brains, for example, do those atoms determine our thoughts and emotions? Are our feelings caused by atoms acting according to the laws of chemistry and physics?

*This calculation assumes that all of the carbon atoms metabolized by your professor over the last 40 years have been uniformly distributed into atmospheric carbon dioxide, and subsequently incorporated into the plants you have eaten.

Richard Feynman (1918–1988), a Nobel Prize–winning physicist, said that "The most important hypothesis in all of biology is that everything that animals do, atoms do. In other words, there is nothing that living things do that cannot be understood from the point of view that they are made of atoms acting according to the laws of physics." Indeed, biology has undergone a revolution in the last 50 years, mostly through investigation of the atomic and molecular basis for life. Some people have seen the atomic view of life as a devaluation of human life. We have always wanted to distinguish ourselves from everything else, and the idea that we are made of the same basic particles as all other matter takes something away from that distinction . . . or does it?

QUESTIONS

Do you find the idea that you are made of recycled atoms disturbing? Why or why not? Reductionism is the idea that complex systems can be understood by understanding their parts. Is reductionism a good way to understand humans? Is it the only way?

2.4 The Discovery of the Electron

By the end of the nineteenth century, scientists were convinced that matter is made up of atoms, the permanent, supposedly indestructible building blocks that compose everything. However, further experiments revealed that the atom itself is composed of even smaller, more fundamental particles.

Cathode Rays

In the late 1800s, an English physicist named J. J. Thomson (1856–1940), working at Cambridge University, performed experiments to probe the properties of **cathode rays**. Thomson constructed a partially evacuated glass tube called a **cathode ray tube**, shown in Figure 2.2 ▼. Thomson then applied a high electrical voltage between two electrodes at either end of the tube. He found that a beam of particles, called cathode rays, traveled from the negatively charged electrode (which is called the cathode) to the positively charged one (which is called the anode).

Thomson found that the particles that compose the cathode ray have the following properties: they travel in straight lines; they are independent of the composition of the material from which they originate (the cathode); and they carry a negative **electrical charge**. Electrical charge is a fundamental property of some of the particles that compose atoms and results in attractive and repulsive forces—called *electrostatic forces*—between those particles. The area around a charged particle where these forces exist is called an *electric field*. The characteristics of electrical charge are summarized in the figure in the margin. You have probably experienced excess electrical charge when brushing your hair on a dry day. The brushing action causes the accumulation of charged particles in your hair, which repel each other, making your hair stand on end.

Properties of Electrical Charge

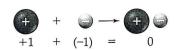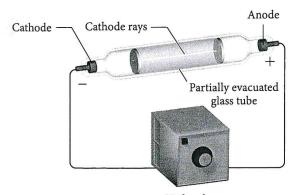Positive (red) and negative (yellow) electrical charges attract one another.

Positive charges repel one another. Negative charges repel one another.

$$+1 \quad + \quad (-1) \quad = \quad 0$$

Positive and negative charges of exactly the same magnitude sum to zero when combined.

For a full explanation of electrical voltage, see Chapter 19.

Cathode — Cathode rays — Anode

Partially evacuated glass tube

High voltage

◄ FIGURE 2.2 **Cathode Ray Tube**

J. J. Thomson measured the charge-to-mass ratio of the cathode ray particles by deflecting them using electric and magnetic fields, as shown in Figure 2.3 ▶, on the next page. The value he measured, -1.76×10^8 coulombs (C) per gram, implied that the cathode ray particle was about 2000 times lighter (less massive) than hydrogen, the lightest known atom. These results were revolutionary—the indestructible atom could apparently be chipped!
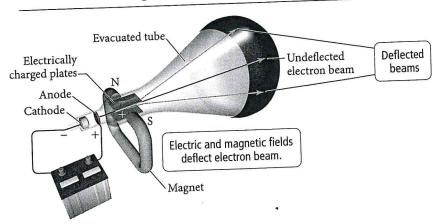
J. J. Thomson had discovered the **electron**, a negatively charged, low-mass particle present within all atoms. He wrote, "We have in the cathode rays matter in a new state, a state in which the subdivision of matter is carried very much further . . . a state in which all matter . . . is of one and the same kind; this matter being the substance from which all the chemical elements are built up."

The coulomb (C) is the SI unit for charge.

▶ FIGURE 2.3 **Thomson's Measurement of the Charge-to-Mass Ratio of the Electron** J. J. Thomson used electric and magnetic fields to deflect the electron beam in a cathode ray tube. By measuring the strengths at which the effects of the two fields (electric and magnetic) cancel exactly, leaving the beam undeflected, he was able to calculate the charge-to-mass ratio of the electron.

Charge-to-Mass Ratio of the Electron

Evacuated tube

Electrically charged plates

Anode

Cathode

N

S

+

Undeflected electron beam

Deflected beams

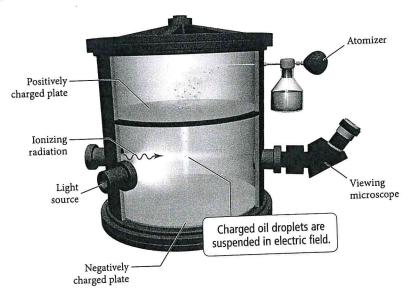Electric and magnetic fields deflect electron beam.

Magnet

Millikan's Oil Drop Experiment: The Charge of the Electron

In 1909, American physicist Robert Millikan (1868–1953), working at the University of Chicago, performed his now famous oil drop experiment in which he deduced the charge of a single electron. The apparatus for the oil drop experiment is shown in Figure 2.4▼.

In his experiment, Millikan sprayed oil into fine droplets using an atomizer. The droplets were allowed to fall under the influence of gravity through a small hole into the lower portion of the apparatus where Millikan viewed them with the aid of a light source and a viewing microscope. During their fall, the drops acquired electrons Millikan had produced by bombarding the air in the chamber with ionizing radiation (a kind of energy described in Chapter 7). The electrons imparted a negative charge to the drops. In the lower portion of the apparatus, Millikan could create an electric field between two metal plates. Since the lower plate was negatively charged, and since Millikan could vary the strength of the electric field, he could slow or even reverse the free fall of the negatively charged drops. (Remember that like charges repel each other.)

By measuring the strength of the electric field required to halt the free fall of the drops, and by figuring out the masses of the drops themselves (determined from their radii and density), Millikan calculated the charge of each drop. He then reasoned that, since each drop must contain an integral (or whole) number of electrons, the charge of each drop must be a whole-number multiple of the electron's charge. Indeed, Millikan was correct; the measured charge on any drop is always a whole-number multiple of -1.60×10^{-19} C, the fundamental charge of a single electron.

▶ FIGURE 2.4 **Millikan's Measurement of the Electron's Charge** Millikan calculated the charge on oil droplets falling in an electric field. He found that it was always a whole-number multiple of -1.60×10^{-19} C, the charge of a single electron.

Atomizer

Positively charged plate

Ionizing radiation

Light source

Viewing microscope

Charged oil droplets are suspended in electric field.

Negatively charged plate

With this number in hand, and knowing Thomson's mass-to-charge ratio for electrons, we can deduce the mass of an electron:
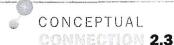
$$\text{Charge} \times \frac{\text{mass}}{\text{charge}} = \text{mass}$$

$$-1.60 \times 10^{-19}\,\mathcal{C} \times \frac{\text{g}}{-1.76 \times 10^8\,\mathcal{C}} = 9.10 \times 10^{-28}\,\text{g}$$

As Thomson had correctly determined, this mass is about 2000 times lighter than hydrogen, the lightest atom.

Why did scientists work so hard to measure the charge of the electron? Since the electron is a fundamental building block of matter, scientists want to know its properties, including its charge. The magnitude of the charge of the electron is of tremendous importance because it determines how strongly an atom holds its electrons. Imagine how matter would be different if electrons had a much smaller charge, so that atoms held them more loosely. Many atoms might not even be stable. On the other hand, imagine how matter would be different if electrons had a much greater charge, so that atoms held them more tightly. Since atoms form compounds by exchanging and sharing electrons (more on this in Chapter 3), there could be fewer compounds or maybe even none. Without the abundant diversity of compounds, life would not be possible. So, the magnitude of the charge of the electron—even though it may seem like an insignificantly small number—has great importance.

The Millikan Oil Drop Experiment Suppose that one of Millikan's oil drops has a charge of -4.8×10^{-19} C. How many excess electrons does the drop contain?

CONCEPTUAL
CONNECTION **2.3**

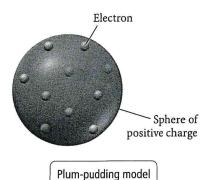
2.5 The Structure of the Atom

The discovery of negatively charged particles within atoms raised a new question. Since atoms are charge-neutral, they must contain a positive charge that neutralizes the negative charge of the electrons—but how do the positive and negative charges fit together? Are atoms just a jumble of even more fundamental particles? Are they solid spheres? Do they have some internal structure? J. J. Thomson proposed that the negatively charged electrons were small particles held within a positively charged sphere.

This model, the most popular of the time, became known as the plum-pudding model (shown at right). The model suggested by Thomson, to those of us not familiar with plum pudding (an English dessert), was like a blueberry muffin; the blueberries are the electrons, and the muffin is the positively charged sphere.

The discovery of **radioactivity**—the emission of small energetic particles from the core of certain unstable atoms—by scientists Henri Becquerel (1852–1908) and Marie Curie (1867–1934) at the end of the nineteenth century allowed researchers to experimentally probe the structure of the atom. At the time, scientists had identified three different types of radioactivity: alpha (α) particles, beta (β) particles, and gamma (γ) rays. We will discuss these and other types of radioactivity in more detail in Chapter 20. For now, just know that α particles are positively charged and that they are by far the most massive of the three.

In 1909, Ernest Rutherford (1871–1937), who had worked under Thomson and subscribed to his plum-pudding model, performed an experiment in an attempt to confirm Thomson's model. Rutherford's experiment, which employed α particles, proved Thomson wrong instead. In the experiment, Rutherford directed the positively charged α particles at an ultrathin sheet of gold foil, as shown in Figure 2.5▶, on the next page.
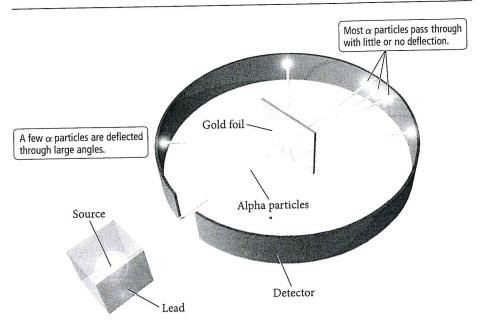
These particles were to act as probes of the gold atoms' structure. If the gold atoms were indeed like blueberry muffins or plum pudding—with their mass and charge spread throughout the entire volume of the atom—Rutherford reasoned that these speeding probes would pass right through the gold foil with minimum deflection.

Electron

Sphere of positive charge

Plum-pudding model

Alpha particles are about 7000 times more massive than electrons.

▶ FIGURE 2.5 **Rutherford's Gold Foil Experiment** Alpha particles are directed at a thin sheet of gold foil. Most of the particles pass through the foil, but a small fraction are deflected, and a few even bounce backward.
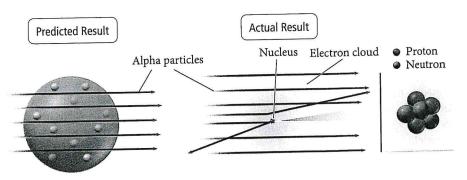
Rutherford's Gold Foil Experiment

Most α particles pass through with little or no deflection.

A few α particles are deflected through large angles.

Gold foil

Alpha particles

Source

Lead

Detector

Rutherford and his coworkers performed the experiment, but the results were not what they expected. The majority of the particles did pass directly through the foil, but some particles were deflected, and some (approximately 1 in 20,000) even bounced back. The results puzzled Rutherford, who wrote that they were "about as credible as if you had fired a 15-inch shell at a piece of tissue paper and it came back and hit you." What sort of atomic structure could explain this odd behavior? Rutherford created a new model—a modern version of which is shown in Figure 2.6 ▼ alongside the plum-pudding model—to explain his results.

Rutherford realized that to account for the deflections he observed, the mass and positive charge of an atom must be concentrated in a space much smaller than the size of the atom itself. He concluded that, in contrast to the plum-pudding model, matter must not be as uniform as it appears. It must contain large regions of empty space dotted with small regions of very dense matter. Building on this idea, he proposed the **nuclear theory** of the atom, with three basic parts:

1. Most of the atom's mass and all of its positive charge are contained in a small core called the **nucleus**.

2. Most of the volume of the atom is empty space, throughout which tiny, negatively charged electrons are dispersed.

3. There are as many negatively charged electrons outside the nucleus as there are positively charged particles (named **protons**) within the nucleus, so the atom is electrically neutral.

Predicted Result

Actual Result

Alpha particles

Nucleus   Electron cloud

● Proton
● Neutron

Plum-pudding model

Nuclear model

▶ FIGURE 2.6 **The Nuclear Atom** Rutherford's results could not be explained by the plum-pudding model. Instead, they suggest that the atom has a small, dense nucleus.

Although Rutherford's model was highly successful, scientists realized that it was incomplete. For example, hydrogen atoms contain one proton, and helium atoms contain two, yet a hydrogen atom has only one-fourth the mass of a helium atom. Why? The helium atom must contain some additional mass. Later work by Rutherford and one of his students, British scientist James Chadwick (1891–1974), demonstrated that the previously unaccounted for mass was due to **neutrons**, neutral particles within the nucleus. The mass of a neutron is similar to that of a proton, but a neutron has no electrical charge. The helium atom is four times as massive as the hydrogen atom because its nucleus contains two protons *and two neutrons* (while hydrogen contains only one proton and no neutrons).

The dense nucleus contains over 99.9% of the mass of the atom but occupies very little of its volume. For now, we can think of the electrons that surround the nucleus in analogy to the water droplets that make up a cloud—although their mass is almost negligibly small, they are dispersed over a very large volume. Consequently, an atom, like a cloud, is mostly empty space.

Rutherford's nuclear theory was a success and is still valid today. The revolutionary part of this theory is the idea that matter—at its core—is much less uniform than it appears. If the nucleus of the atom were the size of the period at the end of this sentence, the average electron would be about 10 meters away. Yet the period would contain nearly all of the atom's mass. Imagine what matter would be like if atomic structure were different. What if matter were composed of atomic nuclei piled on top of each other like marbles in a box? Such matter would be incredibly dense; a single grain of sand composed of solid atomic nuclei would have a mass of 5 million kilograms (or a weight of about 11 million pounds). Astronomers believe there are some objects in the universe composed of such matter—neutron stars.
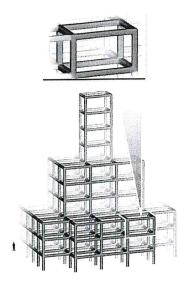
If matter really is mostly empty space, as Rutherford suggested, then why does it appear so solid? Why can we tap our knuckles on a table and feel a solid thump? Matter appears solid because the variation in its density is on such a small scale that our eyes cannot see it. Imagine a scaffolding 100 stories high and the size of a football field as shown in the margin. The volume of the scaffolding is mostly empty space. Yet if you viewed it from an airplane, it would appear as a solid mass. Matter is similar. When you tap your knuckle on the table, it is much like one giant scaffolding (your finger) crashing into another (the table). Even though they are both primarily empty space, one does not fall into the other.

2.6 Subatomic Particles: Protons, Neutrons, and Electrons in Atoms

All atoms are composed of the same subatomic particles: protons, neutrons, and electrons. Protons and neutrons, as we discussed earlier, have nearly identical masses. In SI units, the mass of the proton is 1.67262×10^{-27} kg, and the mass of the neutron is 1.67493×10^{-27} kg. A more common unit to express these masses is the **atomic mass unit (amu)**, defined as 1/12 the mass of a carbon atom that contains six protons and six neutrons. The mass of a proton or neutron is approximately 1 amu. Electrons, by contrast, have an almost negligible mass of 0.00091×10^{-27} kg or 0.00055 amu.

The proton and the electron both have electrical *charge*. We know from Millikan's oil drop experiment that the electron has a charge of -1.60×10^{-19} C. In atomic (or relative) units, the electron is assigned a charge of -1 and the proton is assigned a charge of $+1$. *The charge of the proton and the charge of the electron are equal in magnitude but opposite in sign*, so that when the two particles are paired, the charges sum to zero. The neutron has no charge.

Matter is usually charge-neutral (it has no overall charge) because protons and electrons are normally present in equal numbers. When matter does acquire charge imbalances, these imbalances usually equalize quickly, often in dramatic ways. For example, the shock you receive when touching a doorknob during dry weather is the equalization
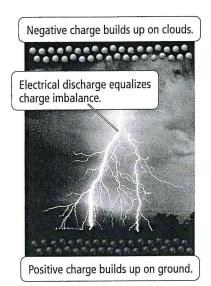
Subatomic Particles and Isotope Symbols

If a proton had the mass of a baseball, an electron would have the mass of a rice grain.

Negative charge builds up on clouds.

Electrical discharge equalizes charge imbalance.

Positive charge builds up on ground.

When the normal charge balance of matter is disturbed, as happens during an electrical storm, it quickly equalizes, often in dramatic ways.

of a charge imbalance that developed as you walked across the carpet. Lightning is equalization of charge imbalances that develop during electrical storms.

A sample of matter—even a tiny sample, such as a sand grain—composed of o protons or only electrons, would have extraordinary repulsive forces inherent withi and would be unstable. Luckily, matter is not that way. Table 2.1 summarizes the proj ties of protons, neutrons, and electrons.
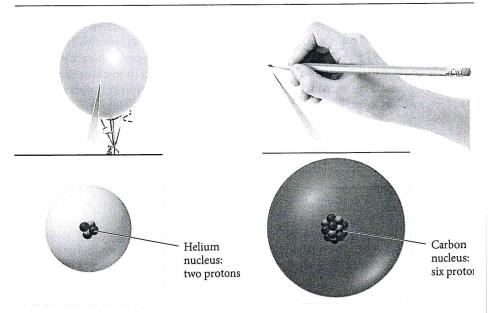
TABLE 2.1 Subatomic Particles				
	Mass (kg)	Mass (amu)	Charge (relative)	Charge (C)
Proton	1.67262×10^{-27}	1.00727	+1	$+1.60218 \times 10^{-}$
Neutron	1.67493×10^{-27}	1.00866	0	0
Electron	0.00091×10^{-27}	0.00055	−1	$-1.60218 \times 10^{-}$

Elements: Defined by Their Numbers of Protons

If all atoms are composed of the same subatomic particles, what makes the atom one element different from those of another? The answer is the *number* of these ticles. The most important number to the *identity* of an atom is the number of prot in its nucleus. *The number of protons defines the element.* For example, an atom with protons in its nucleus is a helium atom, an atom with six protons in its nucleus is a bon atom (Figure 2.7 ▼), and an atom with 92 protons in its nucleus is a uranium atom The number of protons in an atom's nucleus is its **atomic number** and is given symbol Z. The atomic numbers of known elements range from 1 to 116 (although dence for the synthesis of additional elements is currently being considered), as sho in the periodic table of the elements (Figure 2.8 ▶). In the periodic table, described more detail in Section 2.7, the elements are arranged so that those with similar pro ties are in the same column.

Each element, identified by its unique atomic number, is represented with a uni **chemical symbol**, a one- or two-letter abbreviation listed directly below its ato number on the periodic table. The chemical symbol for helium is He; for carbon, symbol is C; and for uranium, it is U. The chemical symbol and the atomic num always go together. If the atomic number is 2, the chemical symbol *must be* He. If atomic number is 6, the chemical symbol *must be* C. This is another way of saying t the number of protons defines the element.

The Number of Protons Defines the Element

▶ FIGURE 2.7 **How Elements Differ** Each element is defined by a unique atomic number (Z), the number of protons in the nucleus of every atom of that element.

Helium nucleus: two protons

Carbon nucleus: six proto

The Periodic Table

Atomic number (Z)

4
Be
beryllium

— Chemical symbol

— Name

1 **H** hydrogen																	2 **He** helium
3 **Li** lithium	4 **Be** beryllium											5 **B** boron	6 **C** carbon	7 **N** nitrogen	8 **O** oxygen	9 **F** fluorine	10 **Ne** neon
11 **Na** sodium	12 **Mg** magnesium											13 **Al** aluminum	14 **Si** silicon	15 **P** phosphorus	16 **S** sulfur	17 **Cl** chlorine	18 **Ar** argon
19 **K** potassium	20 **Ca** calcium	21 **Sc** scandium	22 **Ti** titanium	23 **V** vanadium	24 **Cr** chromium	25 **Mn** manganese	26 **Fe** iron	27 **Co** cobalt	28 **Ni** nickel	29 **Cu** copper	30 **Zn** zinc	31 **Ga** gallium	32 **Ge** germanium	33 **As** arsenic	34 **Se** selenium	35 **Br** bromine	36 **Kr** krypton
37 **Rb** rubidium	38 **Sr** strontium	39 **Y** yttrium	40 **Zr** zirconium	41 **Nb** niobium	42 **Mo** molybdenum	43 **Tc** technetium	44 **Ru** ruthenium	45 **Rh** rhodium	46 **Pd** palladium	47 **Ag** silver	48 **Cd** cadmium	49 **In** indium	50 **Sn** tin	51 **Sb** antimony	52 **Te** tellurium	53 **I** iodine	54 **Xe** xenon
55 **Cs** cesium	56 **Ba** barium	57 **La** lanthanum	72 **Hf** hafnium	73 **Ta** tantalum	74 **W** tungsten	75 **Re** rhenium	76 **Os** osmium	77 **Ir** iridium	78 **Pt** platinum	79 **Au** gold	80 **Hg** mercury	81 **Tl** thallium	82 **Pb** lead	83 **Bi** bismuth	84 **Po** polonium	85 **At** astatine	86 **Rn** radon
87 **Fr** francium	88 **Ra** radium	89 **Ac** actinium	104 **Rf** rutherfordium	105 **Db** dubnium	106 **Sg** seaborgium	107 **Bh** bohrium	108 **Hs** hassium	109 **Mt** meitnerium	110 **Ds** darmstadtium	111 **Rg** roentgenium	112 **Cn** copernicium	113 ******	114 **Fl** flerovium	115 ******	116 **Lv** livermorium	117 ******	118 ******

58 **Ce** cerium	59 **Pr** praseodymium	60 **Nd** neodymium	61 **Pm** promethium	62 **Sm** samarium	63 **Eu** europium	64 **Gd** gadolinium	65 **Tb** terbium	66 **Dy** dysprosium	67 **Ho** holmium	68 **Er** erbium	69 **Tm** thulium	70 **Yb** ytterbium	71 **Lu** lutetium
90 **Th** thorium	91 **Pa** protactinium	92 **U** uranium	93 **Np** neptunium	94 **Pu** plutonium	95 **Am** americium	96 **Cm** curium	97 **Bk** berkelium	98 **Cf** californium	99 **Es** einsteinium	100 **Fm** fermium	101 **Md** mendelevium	102 **No** nobelium	103 **Lr** lawrencium

▲ FIGURE 2.8 **The Periodic Table**   Each element is represented by its symbol and atomic number. Elements in the same column have similar properties.

Most chemical symbols are based on the English name of the element. For example, the symbol for sulfur is S; for oxygen, O; and for chlorine, Cl. Several of the oldest known elements, however, have symbols based on their original Latin names. For example, the symbol for sodium is Na from the Latin *natrium*, and the symbol for tin is Sn from the Latin *stannum*. Early scientists often gave newly discovered elements names that reflect their properties. For example, argon originates from the Greek word *argos* meaning inactive, referring to argon's chemical inertness (it does not react with other elements). Chlorine originates from the Greek word *chloros* meaning pale green, referring to chlorine's pale green color. Other elements, including helium, selenium, and mercury, are named after figures from Greek or Roman mythology or astronomical bodies. Still others (such as europium, polonium, and berkelium) are named for the places where they were discovered or where their discoverers were born. More recently, elements have been named after scientists; for example, curium for Marie Curie, einsteinium for Albert Einstein, and rutherfordium for Ernest Rutherford.

96
Cm
Curium

Isotopes: When the Number of Neutrons Varies

All atoms of a given element have the same number of protons; however, they do not necessarily have the same number of neutrons. Since neutrons have nearly the same mass as protons (1 amu), this means that—contrary to what John Dalton originally proposed in his atomic theory—all atoms of a given element *do not* have the same mass. For example, all neon atoms contain 10 protons, but they may contain 10, 11, or 12 neutrons. All three types of neon atoms exist, and each has a slightly different mass. Atoms with the same number of protons but different numbers of neutrons are

▲ Element 96 is named curium, after Marie Curie, co-discoverer of radioactivity.

isotopes. Some elements, such as beryllium (Be) and aluminum (Al), have only one naturally occurring isotope, while other elements, such as neon (Ne) and chlorine (Cl), have two or more.

The relative amount of each different isotope in a naturally occurring sample of a given element is roughly constant. For example, in any natural sample of neon atoms, 90.48% of them are the isotope with 10 neutrons, 0.27% are the isotope with 11 neutrons, and 9.25% are the isotope with 12 neutrons. These percentages are the **natural abundance** of the isotopes. Each element has its own characteristic natural abundance of isotopes. However, advances in mass spectrometry (see Section 2.8) have allowed accurate measurements that reveal small but significant variations in the natural abundance of isotopes for many elements.

The sum of the number of neutrons and protons in an atom is its **mass number** We represent mass number with the symbol A.

$$A = \text{number of protons (p)} + \text{number of neutrons (n)}$$

For neon, with 10 protons, the mass numbers of the three different naturally occurring isotopes are 20, 21, and 22, corresponding to 10, 11, and 12 neutrons, respectively.

We symbolize isotopes using the notation:

Mass number $\longrightarrow$ $^A_Z X$ $\longleftarrow$ Chemical symbol

Atomic number $\longrightarrow$

where X is the chemical symbol, A is the mass number, and Z is the atomic number Therefore, the symbols for the neon isotopes are:

$$^{20}_{10}\text{Ne} \quad ^{21}_{10}\text{Ne} \quad ^{22}_{10}\text{Ne}$$

Notice that the chemical symbol, Ne, and the atomic number, 10, are redundant: if the atomic number is 10, the symbol must be Ne. The mass numbers, however, are different for the different isotopes, reflecting the different number of neutrons in each one.

A second common notation for isotopes is the chemical symbol (or chemical name) followed by a dash and the mass number of the isotope.

Chemical symbol $\longrightarrow$ X-A $\longleftarrow$ Mass number
or name

In this notation, the neon isotopes are:

Ne-20 Ne-21 Ne-22
neon-20 neon-21 neon-22

We summarize the neon isotopes in the following table:

Symbol	Number of Protons	Number of Neutrons	A (Mass Number)	Natural Abundance (%)
Ne-20 or $^{20}_{10}$Ne	10	10	20	90.48
Ne-21 or $^{21}_{10}$Ne	10	11	21	0.27
Ne-22 or $^{22}_{10}$Ne	10	12	22	9.25

Notice that all isotopes of a given element have the same number of protons (otherwise they would be different elements). Notice also that the mass number is the sum the number of protons and the number of neutrons. The number of neutrons in an isotope is therefore the difference between the mass number and the atomic number (A–). The different isotopes of an element generally exhibit the same chemical behavior—the three isotopes of neon, for example, all exhibit chemical inertness.

Example 2.3   **Atomic Numbers, Mass Numbers, and Isotope Symbols**   ▷

(a) What are the atomic number (Z), mass number (A), and symbol of the chlorine isotope with 18 neutrons?

(b) How many protons, electrons, and neutrons are present in an atom of $^{52}_{24}Cr$?

SOLUTION

(a) Look up the atomic number (Z) for chlorine on the periodic table. The atomic number specifies the number of protons.	$Z = 17$, so chlorine has 17 protons.
The mass number (A) for an isotope is the sum of the number of protons and the number of neutrons.	A = number of protons + number of neutrons $= 17 + 18 = 35$
The symbol for an isotope is its chemical symbol with the atomic number (Z) in the lower left corner and the mass number (A) in the upper left corner.	$^{35}_{17}Cl$
(b) For any isotope (in this case $^{52}_{24}Cr$) the atomic number located at the lower left indicates the number of protons. Since this is a neutral atom, the number of electrons equals the number of protons.	Number of protons = Z = 24 Number of electrons = 24 (neutral atom)
The number of neutrons is equal to the mass number (upper left) minus the atomic number (lower left).	Number of neutrons = $52 - 24 = 28$

FOR PRACTICE 2.3

(a) What are the atomic number, mass number, and symbol for the carbon isotope with seven neutrons?

(b) How many protons and neutrons are present in an atom of $^{39}_{19}K$?

Isotopes Carbon has two naturally occurring isotopes: C-12 (natural abundance is 98.93%) and C-13 (natural abundance is 1.07%). Using circles to represent protons and squares to represent neutrons, draw the nucleus of each isotope. How many C-13 atoms are present, on average, in a 10,000-atom sample of carbon?

CONCEPTUAL
2.4

Ions: Losing and Gaining Electrons

The number of electrons in a neutral atom is equal to the number of protons in its nucleus (designated by its atomic number Z). During chemical changes, however, atoms can lose or gain electrons and become charged particles called **ions**. For example, neutral lithium (Li) atoms contain three protons and three electrons; however, in many chemical reactions lithium atoms lose one electron (e^-) to form Li^+ ions.

$$Li \longrightarrow Li^+ + 1\,e^-$$

The charge of an ion is indicated in the upper right corner of the chemical symbol. Since the Li^+ ion contains three protons and only two electrons, its charge is 1+ (ion charges are written as the magnitude first followed by the sign of the charge; for a charge of 1+, the 1 is usually dropped and the charge is written as simply +).

Ions can also be negatively charged. For example, neutral fluorine (F) atoms contain nine protons and nine electrons; however, in many chemical reactions fluorine atoms gain one electron to form F^- ions.

$$F + 1e^- \longrightarrow F^-$$

The F^- ion contains nine protons and 10 electrons, resulting in a charge of 1− (written simply as −). For many elements, such as lithium and fluorine, the ion is much more common than the neutral atom. In fact, lithium and fluorine occur in nature mostly as ions.

Positively charged ions, such as Li^+, are **cations**, and negatively charged ions, such as F^-, are **anions**. Ions behave quite differently than their corresponding atoms. Neutral sodium atoms, for example, are extremely unstable, reacting violently with

most things they contact. Sodium cations (Na^+), by contrast, are relatively inert—we eat them all the time in sodium chloride (table salt). In ordinary matter, cations and anions always occur together so that matter is charge neutral overall.
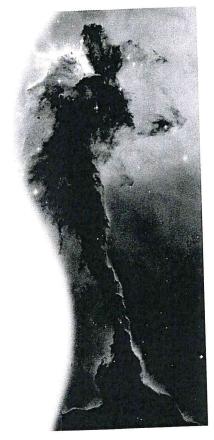
Chemistry IN YOUR DAY | Where Did Elements Come From?

We find ourselves on a planet containing many different kinds of elements. If it were otherwise, we would not exist and would not be here to reflect on why. Where did these elements come from? The story of element formation is as old as the universe itself, and we have to go back to the very beginning to tell the story.

The birth of the universe is described by the Big Bang theory, which asserts that the universe began as a hot, dense collection of matter and energy that expanded rapidly. As it expanded, it cooled, and within the first several hours, subatomic particles formed the first atomic nuclei: hydrogen and helium. These two elements were (and continue to be) the most abundant in the universe. As the universe continued expanding, some of the hydrogen and helium clumped together under the influence of gravity to form *nebulae* (clouds of gas) that eventually gave birth to stars and galaxies. These stars and galaxies became the nurseries where all other elements formed.

Stars are fueled by nuclear fusion, which we will discuss in more detail in Chapter 20. Under the conditions within the core of a star, hydrogen nuclei can combine (or fuse) to form helium. Fusion gives off enormous quantities of energy, which is why stars emit so much heat and light. The fusion of hydrogen to helium can fuel a star for billions of years.

After it burns through large quantities of hydrogen, if a star is large enough, the helium that builds up in its core can in turn fuse to form carbon. The carbon then builds up in the core and (again, if the star is large enough) can fuse to form even heavier elements. The fusion process ends with iron, which has a highly stable nucleus. By the time iron is formed, however, the star is near the end of its existence and may enter a phase of expansion, transforming into a *supernova*. Within a supernova, which is in essence a large exploding star, a shower of neutrons allows the lighter elements (which formed during the lifetime of the star through the fusion processes just described) to capture extra neutrons. These neutrons can transform into protons (through processes that we discuss in Chapter 20), contributing ultimately to the formation of elements heavier than iron, all the way up to uranium. As the supernova continues to expand, the elements present within it are blown out into space, where they can incorporate into other nebulae and perhaps even eventually form planets that orbit stars like our own sun.

▲ Stars are born in nebulae such as the Eagle Nebula (also known as M16). This image was taken by the Hubble Space Telescope and shows a gaseous pillar in a star-forming region of the Eagle Nebula.

CONCEPTUAL
CONNECTION **2.5**

The Nuclear Atom, Isotopes, and Ions Which statement is true?

(a) For a given element, the size of an isotope with more neutrons is larger than one with fewer neutrons.

(b) For a given element, the size of an atom is the same for all of the element's isotopes.

2.7 Finding Patterns: The Periodic Law and the Periodic Table

The modern periodic table grew out of the work of Dmitri Mendeleev (1834–1907), a nineteenth-century Russian chemistry professor. In his time, scientists had discovered about 65 different elements, and chemists had identified many of the

Time of Discovery

- Before 1800
- 1800–1849
- 1850–1899
- 1900–1949
- 1950–2012

▲ Many of the elements that we know today were discovered during Mendeleev's lifetime.

properties of these elements—such as their relative masses, their chemical activity, and some of their physical properties. However, no one had developed any systematic way of organizing them.

In 1869, Mendeleev noticed that certain groups of elements had similar properties. He also found that when he listed elements in order of increasing mass, these similar properties recurred in a periodic pattern (Figure 2.9▼). Mendeleev summarized these observations in the **periodic law**:

| To be periodic means to exhibit a repeating pattern.

When the elements are arranged in order of increasing mass, certain sets of properties recur periodically.

The Periodic Law

Elements with similar properties recur in a regular pattern.

▲ FIGURE 2.9 **Recurring Properties** These elements are listed in order of increasing atomic number. Elements with similar properties are represented with the same color. Notice that the colors form a repeating pattern, much like musical notes form a repeating pattern on a piano keyboard.

A Simple Periodic Table

Elements with similar properties fall into columns.

▲ FIGURE 2.10 **Making a Periodic Table** We can arrange the elements in Figure 2.9 in a table in which atomic number increases from left to right and elements with similar properties (as represented by the different colors) are aligned in columns.

Mendeleev organized the known elements in a table consisting of a series of rows in which mass increases from left to right. He arranged the rows so that elements with similar properties fall in the same vertical columns (Figure 2.10▶).

Since many elements had not yet been discovered, Mendeleev's table contained some gaps, which allowed him to predict the existence (and even the properties) of yet undiscovered elements. For example, Mendeleev predicted the existence of an element he called eka-silicon, which fell below silicon on the table and between gallium and arsenic (eka means the one beyond). In 1886, eka-silicon was discovered by German chemist Clemens Winkler (1838–1904), who named it germanium, after his home country.

СТОЛЕТИЕ ПЕРИОДИЧЕСКОГО ЗАКОНА
Д.И.МЕНДЕЛЕЕВА

ПОЧТА СССР 1969 6К

▲ Dmitri Mendeleev, a Russian chemistry professor who proposed the periodic law and arranged early versions of the periodic table, was honored on a Soviet postage stamp.

Modern Periodic Table Organization

Mendeleev's original listing evolved into the modern periodic table shown in Figure 2.11 ▼. In the modern table, elements are listed in order of increasing atomic number rather than increasing relative mass. The modern periodic table also contains more elements than Mendeleev's original table because more have been discovered since his time. Mendeleev's periodic law was based on observation. Like all scientific laws, the periodic law summarizes many observations but does not give the underlying reason for the observations—only theories do that. For now, we accept the periodic law as it is, but in Chapters 7 and 8 we will examine a powerful theory—called quantum mechanics—that explains the law and gives the underlying reasons for it.

We can broadly classify the elements in the periodic table as metals, nonmetals, or metalloids, as shown in Figure 2.11. **Metals** lie on the lower left side and middle of the periodic table and share some common properties: they are good conductors of heat and electricity, they can be pounded into flat sheets (malleability), they can be drawn into wires (ductility), they are often shiny, and they tend to lose electrons when they undergo chemical changes. Chromium, copper, strontium, and lead are typical metals.

Nonmetals lie on the upper right side of the periodic table. The dividing line between metals and nonmetals is the zigzag diagonal line running from boron to astatine. Nonmetals have varied properties—some are solids at room temperature, others are liquids or gases—but typically they tend to be poor conductors of heat and electricity, and they all tend to gain electrons when they undergo chemical changes. Oxygen, carbon, sulfur, bromine, and iodine are nonmetals.

Major Divisions of the Periodic Table

▲ FIGURE 2.11 **Metals, Nonmetals, and Metalloids** The elements in the periodic table fall into these three broad classes.

Main-group elements | Transition elements | Main-group elements

	1A 1	*Group number*												3A 13	4A 14	5A 15	6A 16	7A 17	8A 18
1	1 H	2A 2																	2 He
2	3 Li	4 Be												5 B	6 C	7 N	8 O	9 F	10 Ne
3	11 Na	12 Mg	3B 3	4B 4	5B 5	6B 6	7B 7	8	8B 9	10	1B 11	2B 12	13 Al	14 Si	15 P	16 S	17 Cl	18 Ar
4	19 K	20 Ca	21 Sc	22 Ti	23 V	24 Cr	25 Mn	26 Fe	27 Co	28 Ni	29 Cu	30 Zn	31 Ga	32 Ge	33 As	34 Se	35 Br	36 Kr
5	37 Rb	38 Sr	39 Y	40 Zr	41 Nb	42 Mo	43 Tc	44 Ru	45 Rh	46 Pd	47 Ag	48 Cd	49 In	50 Sn	51 Sb	52 Te	53 I	54 Xe
6	55 Cs	56 Ba	57 La	72 Hf	73 Ta	74 W	75 Re	76 Os	77 Ir	78 Pt	79 Au	80 Hg	81 Tl	82 Pb	83 Bi	84 Po	85 At	86 Rn
7	87 Fr	88 Ra	89 Ac	104 Rf	105 Db	106 Sg	107 Bh	108 Hs	109 Mt	110 Ds	111 Rg	112 Cn	113	114 Fl	115	116 Lv	117	118

Periods

▲ FIGURE 2.12 **The Periodic Table: Main-Group and Transition Elements** The elements in the periodic table fall into columns. The two columns at the left and the six columns at the right comprise the main-group elements. Each of these eight columns is a group or family. The properties of main-group elements can generally be predicted from their position in the periodic table. The properties of the elements in the middle of the table, known as transition elements, are less predictable.

Many of the elements that lie along the zigzag diagonal line that divides metals and nonmetals are **metalloids** and exhibit mixed properties. Several metalloids are also classified as **semiconductors** because of their intermediate (and highly temperature-dependent) electrical conductivity. Our ability to change and control the conductivity of semiconductors makes them useful to us in the manufacture of the electronic chips and circuits central to computers, cellular telephones, and many other modern devices. Good examples of metalloids are silicon, arsenic, and antimony.

We can also divide the periodic table, as shown in Figure 2.12▲, into **main-group elements**, whose properties tend to be largely predictable based on their position in the periodic table, and **transition elements** or **transition metals**, whose properties tend to be less predictable based simply on their position in the periodic table. Main-group elements are in columns labeled with a number and the letter A. Transition elements are in columns labeled with a number and the letter B. An alternative numbering system does not use letters, but only the numbers 1–18. Both numbering systems are shown in most of the periodic tables in this book. Each column within the main-group regions of the periodic table is a **family** or **group** of elements.
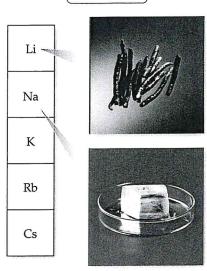
The elements within a group usually have similar properties. For example, the group 8A elements, called the **noble gases**, are mostly unreactive. The most familiar noble gas is probably helium, used to fill buoyant balloons. Helium is chemically stable—it does not combine with other elements to form compounds—and is therefore safe to put into balloons. Other noble gases are neon (often used in electronic signs), argon (a small component of our atmosphere), krypton, and xenon.
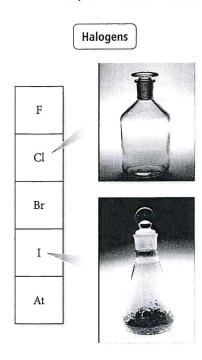
The group 1A elements, the **alkali metals**, are all reactive metals. A marble-sized piece of sodium explodes violently when dropped into water. Lithium, potassium, and rubidium are also alkali metals.

The group 2A elements, the **alkaline earth metals**, are also fairly reactive, although not quite as reactive as the alkali metals. Calcium, for example, reacts fairly vigorously when dropped into water but does not explode as dramatically as sodium. Other alkaline earth metals include magnesium (a common low-density structural metal), strontium, and barium.

Metalloids are sometimes called semimetals.

Alkali metals

Li
Na
K
Rb
Cs

Halogens

| F |
| Cl |
| Br |
| I |
| At |

The group 7A elements, the **halogens**, are very reactive nonmetals. One of the most familiar halogens is chlorine, a greenish-yellow gas with a pungent odor. Because of its reactivity, chlorine is used as a sterilizing and disinfecting agent. Other halogens include bromine, a red-brown liquid that easily evaporates into a gas; iodine, a purple solid; and fluorine, a pale-yellow gas.

Ions and the Periodic Table

In chemical reactions, metals tend to lose electrons (forming cations) and nonmetals tend to gain them (forming anions). The number of electrons lost or gained, and therefore the charge of the resulting ion, is often predictable for a given element, especially main-group elements. Main-group elements tend to form ions that have the same number of electrons as the nearest noble gas (i.e., the noble gas that has the number of electrons closest to that of the element).

- **A main-group metal tends to lose electrons, forming a cation with the same number of electrons as the nearest noble gas.**
- **A main-group nonmetal tends to gain electrons, forming an anion with the same number of electrons as the nearest noble gas.**

For example, lithium, a metal with three electrons, tends to lose one electron, forming a 1+ cation that has the same number of electrons (two) as helium. Chlorine, a nonmetal with 17 electrons, tends to gain one electron, forming a 1− anion that has the same number of electrons (18) as argon.

In general, the alkali metals (group 1A) tend to lose one electron and form 1+ ions. The alkaline earth metals (group 2A) tend to lose two electrons and form 2+ ions. The halogens (group 7A) tend to gain one electron and form 1− ions. The oxygen family nonmetals (group 6A) tend to gain two electrons and form 2− ions. More generally, for the main-group elements that form cations with predictable charge, the charge is equal to the group number. For main-group elements that form anions with predictable charge, the charge is equal to the group number minus eight. Transition elements may form various different ions with different charges. Figure 2.13▶ shows the ions formed by the main-group elements that form ions with predictable charges. In Chapters 7 and 8, we will introduce quantum-mechanical theory, which more fully explains *why* these groups form ions as they do.

Example 2.4   Predicting the Charge of Ions

Predict the charges of the monoatomic (single atom) ions formed by these main-group elements.

(a) Al   **(b)** S

SOLUTION

(a) Aluminum is a main-group metal and tends to lose electrons to form a cation with the same number of electrons as the nearest noble gas. Aluminum atoms have 13 electrons and the nearest noble gas is neon, which has 10 electrons. Aluminum therefore loses three electrons to form a cation with a 3+ charge (Al^{3+}).

(b) Sulfur is a nonmetal and tends to gain electrons to form an anion with the same number of electrons as the nearest noble gas. Sulfur atoms have 16 electrons and the nearest noble gas is argon, which has 18 electrons. Sulfur therefore gains two electrons to form an anion with a 2− charge (S^{2-}).

FOR PRACTICE 2.4   Predict the charges of the monoatomic ions formed by these main-group elements.

(a) N   **(b)** Rb

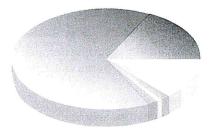**Elements That Form Ions with Predictable Charges**

▲ FIGURE 2.13 **Elements That Form Ions with Predictable Charges**

Chemistry AND **Medicine** | The Elements of Life

What kind of atoms compose us? In terms of mass, our bodies are 65% oxygen (because of the large amount of water), 18% carbon, and 10% hydrogen, with a few other elements present in smaller quantities, as shown in Figure 2.14▷ and Table 2.2. Because the atoms of different elements have different masses (more on this in Sections 2.8 and 2.9), if we consider number of atoms (instead of mass), hydrogen comes in first (because hydrogen atoms are so light), with oxygen second and carbon third.

- Oxygen: 65%
- Carbon: 18%
- Hydrogen: 10%
- Nitrogen: 3%
- Calcium: 1.5%
- Phosphorus: 1%
- Other: 1.5%

▲ FIGURE 2.14 **Elemental Composition of Humans (by Mass)**

TABLE 2.2 Approximate Percent Elemental Composition of Humans

Element	% by Mass	% by Number of Atoms
Oxygen	65	26.4
Carbon	18	9.2
Hydrogen	10	62.3
Nitrogen	3	1.4
Calcium	1.5	0.2
Phosphorus	1	0.3
Other	1.5	0.2

Much of the *chemistry* of life revolves not around oxygen or hydrogen but around carbon, an element that forms a disproportionately large number of compounds with a few other elements such as hydrogen, oxygen, and nitrogen. We will explore the chemistry of carbon—called organic chemistry—briefly in Chapter 3 and in more detail in Chapter 21.

2.8 Atomic Mass: The Average Mass of an Element's Atoms

An important part of Dalton's atomic theory is that all atoms of a given element have the same mass. In Section 2.6, we learned that because of isotopes, the atoms of a given element often have different masses, so Dalton was not completely correct. We can, however, calculate an average mass—called the **atomic mass**—for each element.

The atomic mass of each element is listed directly beneath the element's symbol in the periodic table and represents the average mass of the isotopes that compose that element, *weighted according to the natural abundance of each isotope.* For example, the

Atomic mass is sometimes called *atomic weight* or *standard atomic weight.*

17
Cl
35.45
chlorine

In this book, we use the atomic masses recommended by IUPAC for users needing an atomic mass value for an unspecified sample. Detailed studies of the atomic masses of many samples, however, have shown that atomic masses are not constants of nature because the exact isotopic abundances in any given sample depend on the history of the sample (see the Chemistry in Your Day box later in this section).

When we use percentages in calculations, we convert them to their decimal value by dividing by 100.

periodic table lists the atomic mass of chlorine as 35.45 amu. Naturally occurring chlorine consists of 75.77% chlorine-35 atoms (mass 34.97 amu) and 24.23% chlorine-37 atoms (mass 36.97 amu). We can calculate its atomic mass accordingly:

$$\text{Atomic mass} = 0.7577(34.97 \text{ amu}) + 0.2423(36.97 \text{ amu}) = 35.45 \text{ amu}$$

Notice that the atomic mass of chlorine is closer to 35 than 37. Naturally occurring chlorine contains more chlorine-35 atoms than chlorine-37 atoms, so the weighted average mass of chlorine is closer to 35 amu than to 37 amu.

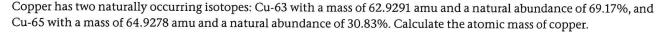In general, we calculate the atomic mass with the equation:

$$\text{Atomic mass} = \sum_n (\text{fraction of isotope } n) \times (\text{mass of isotope } n)$$

$$= (\text{fraction of isotope 1} \times \text{mass of isotope 1})$$

$$+ (\text{fraction of isotope 2} \times \text{mass of isotope 2})$$

$$+ (\text{fraction of isotope 3} \times \text{mass of isotope 3}) + \ldots$$

where the fractions of each isotope are the percent natural abundances converted to their decimal values. The concept of atomic mass is useful because it allows us to assign a characteristic mass to each element, and, as we will see shortly, it allows us to quantify the number of atoms in a sample of that element.

Example 2.5 | Atomic Mass

Copper has two naturally occurring isotopes: Cu-63 with a mass of 62.9291 amu and a natural abundance of 69.17%, and Cu-65 with a mass of 64.9278 amu and a natural abundance of 30.83%. Calculate the atomic mass of copper.

SOLUTION

Convert the percent natural abundances into decimal form by dividing by 100.	Fraction Cu-63 $= \dfrac{69.17}{100} = 0.6917$
	Fraction Cu-65 $= \dfrac{30.83}{100} = 0.3083$
Calculate the atomic mass using the equation given in the text.	Atomic mass $= 0.6917(62.9296 \text{ amu}) + 0.3083(64.9278 \text{ amu})$ $= 43.5284 \text{ amu} + 20.0172 \text{ amu} = 63.5456 = 63.55 \text{ amu}$

FOR PRACTICE 2.5 Magnesium has three naturally occurring isotopes with masses of 23.99 amu, 24.99 amu, and 25.98 amu and natural abundances of 78.99%, 10.00%, and 11.01%, respectively. Calculate the atomic mass of magnesium.

FOR MORE PRACTICE 2.5 Gallium has two naturally occurring isotopes: Ga-69 with a mass of 68.9256 amu and a natural abundance of 60.11%, and Ga-71. Use the atomic mass of gallium from the periodic table to find the mass of Ga-71.
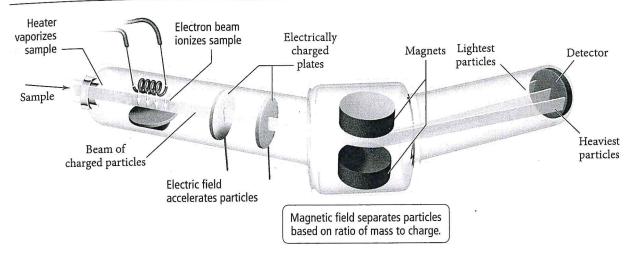
CONCEPTUAL
CONNECTION **2.6**

Atomic Mass Recall from Conceptual Connection 2.4 that carbon has two naturally occurring isotopes: C-12 (natural abundance is 98.93%; mass is 12.0000 amu) and C-13 (natural abundance is 1.07%; mass is 13.0034 amu). Without doing any calculations, determine which mass is closest to the atomic mass of carbon.
(a) 12.00 amu   (b) 12.50 amu   (c) 13.00 amu

Mass Spectrometry: Measuring the Mass of Atoms and Molecules

The masses of atoms and the percent abundances of isotopes of elements are measured using **mass spectrometry**, a technique that separates particles according to their mass. In a mass spectrometer, such as the one in Figure 2.15▶, the sample (containing
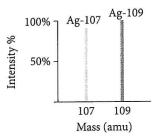
Mass Spectrometer

Heater vaporizes sample

Sample

Electron beam ionizes sample

Beam of charged particles

Electric field accelerates particles

Electrically charged plates

Magnets Lightest particles

Detector

Heaviest particles

Magnetic field separates particles based on ratio of mass to charge.

the atoms whose mass is to be measured) is injected into the instrument and vaporized. The vaporized atoms are ionized by an electron beam—the electrons in the beam collide with the atoms, removing electrons and creating positively charged ions. The ions are then accelerated into a magnetic field. When ions drift through a magnetic field, they experience a force that bends their trajectory. The amount of bending depends on the mass of the ions—the trajectories of lighter ions are bent more than those of heavier ones.

In the right side of the spectrometer in Figure 2.15, you can see three different paths, each corresponding to ions of different mass. Finally, the ions strike a detector and produce an electrical signal that is recorded. The result is the separation of the ions according to their mass, producing a mass spectrum such as the one in Figure 2.16 ▶ . The *position* of each peak on the x-axis indicates the *mass of the isotope* that was ionized, and the *intensity* (indicated by the height of the peak) indicates the *relative abundance of that isotope.*

The mass spectrum of an elemental sample can be used to determine the atomic mass of that sample of the element. For example, consider the mass spectrum of a naturally occurring sample of silver:

The two peaks correspond to the two isotopes of silver. We can determine the percent abundance of each isotope from the intensity of each line. However, the *total* intensity must be *normalized*—it must be made to equal 100%. We can accomplish this by dividing the intensity of each peak by the total intensity:

$$\text{Abundance of Ag-107} = \frac{100.0\%}{100.0\% + 92.90\%} \times 100\% = 51.84\%$$

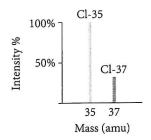$$\text{Abundance of Ag-109} = \frac{92.90\%}{100.0\% + 92.90\%} \times 100\% = 48.16\%$$

Then we can calculate the atomic mass of silver:

$$\text{Ag atomic mass} = 0.5184 \,(106.905 \text{ amu}) + 0.4816 \,(108.904 \text{ amu})$$
$$= 55.4195 \text{ amu} + 52.4482 \text{ amu} = 107.8677 = 107.87 \text{ amu}$$

Mass spectrometry can also be used on molecules. Because molecules often fragment (break apart) during ionization, the mass spectrum of a molecule usually

▲ FIGURE 2.15 **The Mass Spectrometer** Atoms are converted to positively charged ions, accelerated, and passed through a magnetic field that deflects their path. The heaviest ions undergo the least deflection.

▲ FIGURE 2.16 **The Mass Spectrum of Chlorine** The position of each peak on the x-axis indicates the mass of the isotope. The intensity (or height) of the peak indicates the relative abundance of the isotope. The intensity of the highest peak is usually set to 100% and the intensity of all other peaks is reported relative to the most intense one.

Chemistry IN YOUR DAY | Evolving Atomic Masses

At the beginning of 2011, IUPAC (International Union of Pure and Applied Chemistry) published a new periodic table with atomic masses that looked different from previous IUPAC periodic tables. For the first time, instead of listing a single atomic mass for each element, IUPAC listed upper and lower bounds for the atomic masses of several elements (see the periodic table in this box). For example, previous IUPAC periodic tables reported the atomic mass of O (rounded to four significant figures) as 16.00. However, the new periodic table reports the atomic mass as [15.99, 16.00] denoting the upper and lower bounds for the possible atomic masses of terrestrial oxygen.

Why did this happen? The changes were necessary because developments in mass spectrometry have increasingly demonstrated that the atomic masses of several elements are not constant from one sample to another because the isotopic composition is not constant from one sample to another. In other words, the isotopic composition of a sample of a given element can vary depending on the source of the sample.

For example, the lower bound for the atomic mass of oxygen (15.99 amu) comes from measurements of oxygen from Antarctic precipitation, and the upper bound (16.00 amu) comes from measurements of oxygen in marine N_2O (dinitrogen monoxide). Although we have long treated atomic masses as *constants of nature*, they are not, and the new periodic table reflects this.

So what do we do if we need an atomic mass for an element of unknown or unspecified origin? IUPAC has recommended values that apply to most samples found on Earth. The values are rounded so that atomic mass variations in samples found on Earth are plus or minus one in the last digit (just like accepted significant figure conventions). These values are adopted throughout all of the periodic tables in this book except the one shown below, which displays the upper and lower bound for those elements in which variation occurs. For further reading see IUPAC. *Pure Appl. Chem.* 2011, 83(2), 359–396.

IUPAC Periodic Table of the Elements

Notes
- IUPAC 2009 Standard atomic weights abridged to four significant digits (Table 4 published in *Pure Appl. Chem.* 83, 359-396 (2011); doi:10.1351/PAC-REP-10-09-14). The uncertainty in the last digit of the standard atomic weight value is listed in parentheses following the value. In the absence of parentheses, the uncertainty is one in that last digit. An interval in square brackets provides the lower and upper bounds of the standard atomic weight for that element. No values are listed for elements which lack isotopes with a characteristic isotopic abundance in natural terrestrial samples. See PAC for more details.

- "Aluminum" and "cesium" are commonly used alternative spellings for "aluminium" and "caesium."

- Claims for the discovery of all the remaining elements in the last row of the Table, namely elements with atomic numbers 113, 115, 117 and 118, and for which no assignments have yet been made, are being considered by a IUPAC and IUPAP Joint Working Party.

For updates to this table, see iupac.org/reports/periodic_table/. This version is dated 1 June 2012.
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contains many peaks representing the masses of different parts of the molecule, as well as a peak representing the mass of the molecule as a whole. The fragments that form upon ionization, and therefore the corresponding peaks that appear in the mass spectrum, are specific to the molecule, so a mass spectrum is like a molecular fingerprint.

Mass spectroscopy can be used to identify an unknown molecule and to determine how much of it is present in a particular sample. For example, mass spectrometry has been used to detect organic (carbon-containing) compounds present in meteorites, a puzzling observation that some scientists speculate may be evidence of life outside of our planet. Most scientists think that the carbon compounds in meteorites probably formed in the same way as the first organic molecules on Earth, indicating that the formation of organic molecules may be common in the universe.

Since the early 1990s, researchers have also successfully applied mass spectrometry to biological molecules, including proteins (the workhorse molecules in cells) and nucleic acids (the molecules that carry genetic information). For a long time, these molecules could not be analyzed by mass spectrometry because they were difficult to vaporize and ionize without being destroyed, but modern techniques have overcome this problem. A tumor, for example, can now be instantly analyzed by mass spectrometry to determine whether it contains specific proteins associated with cancer.

2.9 Molar Mass: Counting Atoms by Weighing Them

KCV **The Mole Concept**

Shrimp are normally sold by *count*, which indicates the number of shrimp per pound. For example, 41–50 count shrimp means that there are between 41 and 50 shrimp per pound. The smaller the count, the larger the shrimp. Big tiger prawns have counts as low as 10–15, which means that each shrimp can weigh up to 1/10 of a pound. One advantage of categorizing shrimp in this way is that we can count the shrimp by weighing them. For example, two pounds of 41–50 count shrimp contains between 82 and 100 shrimp.

A similar (but more precise) concept exists for atoms. Counting atoms is much more difficult than counting shrimp, yet as chemists we often need to know the number of atoms in a sample of a given mass. Why? *Because chemical processes happen between particles*. For elements, those particles are atoms. For example, when hydrogen and oxygen combine to form water, two hydrogen atoms combine with one oxygen atom to form one water molecule. If we want to know how much hydrogen to react with a given mass of oxygen to form water, we need to know the number of atoms in the given mass of oxygen. We also need to know the mass of hydrogen that contains exactly twice that number of atoms.

As another example, consider intravenous fluids—fluids that are delivered to patients by directly dripping them into veins. These fluids are saline (sodium chloride) solutions that must have a specific number of sodium and chloride ions per liter of fluid. The number of particles in the fluid directly influences the properties of the fluid. Administering an intravenous fluid with the wrong number of sodium and chloride ions could be fatal.

Atoms are far too small to count by any ordinary means. As we noted earlier, even if you could somehow count atoms, and counted them 24 hours a day for as long as you lived, you would barely begin to count the number of atoms in something as small as a sand grain. Therefore, if we want to know the number of atoms in anything of ordinary size, we count them by weighing.

The Mole: A Chemist's "Dozen"

When we count large numbers of objects, we often use units such as a dozen (12 objects) or a gross (144 objects) to organize our counting and to keep our numbers more manageable. With atoms, quadrillions of which may be in a speck of dust, we need a much larger number for this purpose. The chemist's "dozen" is the **mole** (abbreviated mol). A mole is the *amount* of material containing 6.02214×10^{23} particles.

$$1 \text{ mol} = 6.02214 \times 10^{23} \text{ particles}$$

This number is **Avogadro's number**, named after Italian physicist Amedeo Avogadro (1776–1856), and is a convenient number to use when working with atoms, molecules, and ions. In this book, we usually round Avogadro's number to four significant figures or

Twenty-two copper pennies contain approximately 1 mol of copper atoms.

Before 1982, when they became almost all zinc with only a copper coating, pennies were mostly copper.

One tablespoon of water contains approximately 1 mol of water molecules.

One tablespoon is approximately 15 mL; one mol of water occupies 18 mL.

6.022×10^{23}. Notice that the definition of the mole is an *amount* of a substance. often refer to the number of moles of substance as the *amount* of the substance.

The first thing to understand about the mole is that it can specify Avogadro's numb of anything. For example, one mole of marbles corresponds to 6.022×10^{23} marbles, a one mole of sand grains corresponds to 6.022×10^{23} sand grains. *One mole of anything* 6.022×10^{23} *units of that thing*. One mole of atoms, ions, or molecules, however, mak up objects of everyday sizes. Twenty-two copper pennies, for example, contain appro mately 1 mol of copper atoms, and 1 tablespoon of water contains approximately 1 m of water molecules.

The second, and more fundamental, thing to understand about the mole is how gets its specific value.

The value of the mole is equal to the number of atoms in exactly 12 g of pure carbon-12 (12 g C = 1 mol C atoms = 6.022×10^{23} C atoms).

The definition of the mole gives us a relationship between mass (grams of carbon) an number of atoms (Avogadro's number). This relationship, as we will see shortly, allo us to count atoms by weighing them.

Converting between Number of Moles and Number of Atoms

Converting between number of moles and number of atoms is similar to conver ing between dozens of eggs and number of eggs. For eggs, you use the conversion fa tor 1 dozen eggs = 12 eggs. For atoms, you use the conversion factor 1 mol atoms = 6.022×10^{23} atoms. The conversion factors take the following forms:

$$\frac{1 \text{ mol atoms}}{6.022 \times 10^{23} \text{ atoms}} \quad \text{or} \quad \frac{6.022 \times 10^{23} \text{ atoms}}{1 \text{ mol atoms}}$$

Example 2.6 demonstrates how to use these conversion factors in calculations.

Example 2.6 | **Converting between Number of Moles and Number of Atoms**

Calculate the number of copper atoms in 2.45 mol of copper.

SORT You are given the amount of copper in moles and asked to find the number of copper atoms.	GIVEN: 2.45 mol Cu FIND: Cu atoms
STRATEGIZE Convert between number of moles and number of atoms by using Avogadro's number as a conversion factor.	CONCEPTUAL PLAN mol Cu $\longrightarrow$ Cu atoms $\dfrac{6.022 \times 10^{23} \text{ Cu atoms}}{1 \text{ mol Cu}}$ RELATIONSHIPS USED $6.022 \times 10^{23} = 1$ mol (Avogadro's number)
SOLVE Follow the conceptual plan to solve the problem. Begin with 2.45 mol Cu and multiply by Avogadro's number to get to the number of Cu atoms.	SOLUTION $2.45 \text{ mol Cu} \times \dfrac{6.022 \times 10^{23} \text{ Cu atoms}}{1 \text{ mol Cu}} = 1.48 \times 10^{24}$ Cu atoms

CHECK Since atoms are small, it makes sense that the answer is large. The given number of moles of copper is almost 2.5, so the number of atoms is almost 2.5 times Avogadro's number.

FOR PRACTICE 2.6 A pure silver ring contains 2.80×10^{22} silver atoms. How many moles of silver atoms does it contain?

Converting between Mass and Amount (Number of Moles)

To count atoms by weighing them, we need one other conversion factor—the mass of 1 mol of atoms. For the isotope carbon-12, we know that the mass of 1 mol of atoms is exactly 12 g, which is numerically equivalent to carbon-12's atomic mass in atomic mass units. Since the masses of all other elements are defined relative to carbon-12, the same relationship holds for all elements.

The mass of one mole of atoms of an element is its **molar mass**.

> **An element's molar mass in grams per mole is numerically equal to the element's atomic mass in atomic mass units.**

For example, copper has an atomic mass of 63.55 amu and a molar mass of 63.55 g/mol. One mole of copper atoms therefore has a mass of 63.55 g. Just as the count for shrimp depends on the size of the shrimp, so the mass of 1 mol of atoms depends on the element: 1 mol of aluminum atoms (which are lighter than copper atoms) has a mass of 26.98 g, 1 mol of carbon atoms (which are even lighter than aluminum atoms) has a mass of 12.01 g, and 1 mol of helium atoms (lighter yet) has a mass of 4.003 g.
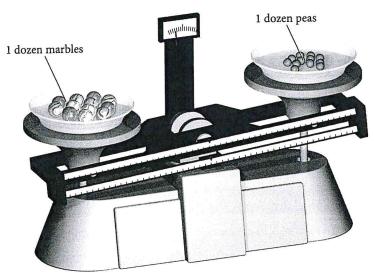
26.98 g aluminum = 1 mol aluminum = 6.022×10^{23} Al atoms $\quad$ ● Al

12.01 g carbon = 1 mol carbon = 6.022×10^{23} C atoms $\quad$ ● C

4.003 g helium = 1 mol helium = 6.022×10^{23} He atoms $\quad$ ● He

The lighter the atom, the less mass in 1 mol of atoms.

1 dozen marbles
1 dozen peas

▲ The two dishes contain the same number of objects (12), but the masses are different because peas are less massive than marbles. Similarly, a mole of light atoms will have less mass than a mole of heavier atoms.

The molar mass of any element is the conversion factor between the mass (in grams) of that element and the amount (in moles) of that element. For carbon:

$$12.01\,\text{g C} \;=\; 1\,\text{mol C} \quad \text{or} \quad \frac{12.01\,\text{g C}}{\text{mol C}} \quad \text{or} \quad \frac{1\,\text{mol C}}{12.01\,\text{g C}}$$

Example 2.7, on the next page, demonstrates how to use these conversion factors.

Example 2.7 Converting between Mass and Amount (Number of Moles)
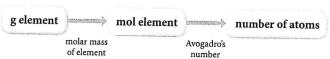
Calculate the amount of carbon (in moles) contained in a 0.0265 g pencil "lead." (Assume that the pencil lead is made of pure graphite, a form of carbon.)

SORT You are given the mass of carbon and asked to find the amount of carbon in moles.	GIVEN: 0.0265 g C FIND: mol C
STRATEGIZE Convert between mass and amount (in moles) of an element by using the molar mass of the element.	CONCEPTUAL PLAN $$\boxed{g\,C} \longrightarrow \boxed{mol\,C}$$ $$\frac{1\,mol}{12.01\,g}$$ RELATIONSHIPS USED 12.01 g C = 1 mol C (carbon molar mass)
SOLVE Follow the conceptual plan to solve the problem.	SOLUTION $$0.0265\,g\,C \times \frac{1\,mol\,C}{12.01\,g\,C} = 2.21 \times 10^{-3}\,mol\,C$$

CHECK The given mass of carbon is much less than the molar mass of carbon, so it makes sense that the answer (the amount in moles) is much less than 1 mol of carbon.

FOR PRACTICE 2.7 Calculate the amount of copper (in moles) in a 35.8 g pure copper sheet.

FOR MORE PRACTICE 2.7 Calculate the mass (in grams) of 0.473 mol of titanium.

We now have all the tools to count the number of atoms in a sample of an element by weighing it. First, we obtain the mass of the sample. Then we convert it to an amount in moles using the element's molar mass. Finally, we convert it to number of atoms using Avogadro's number. The conceptual plan for these kinds of calculations takes the following form:

$$\boxed{g\,element} \longrightarrow \boxed{mol\,element} \longrightarrow \boxed{number\,of\,atoms}$$

molar mass   Avogadro's
of element   number

Examples 2.8 and 2.9 demonstrate these conversions.

Example 2.8 The Mole Concept—Converting between Mass and Number of Atoms

How many copper atoms are in a copper penny with a mass of 3.10 g? (Assume that the penny is composed of pure copper.)

SORT You are given the mass of copper and asked to find the number of copper atoms.	GIVEN: 3.10 g Cu FIND: Cu atoms
STRATEGIZE Convert between the mass of an element in grams and the number of atoms of the element by first converting to moles (using the molar mass of the element) and then to number of atoms (using Avogadro's number).	CONCEPTUAL PLAN $$\boxed{g\,Cu} \longrightarrow \boxed{mol\,Cu} \longrightarrow \boxed{number\,of\,Cu\,atoms}$$ $$\frac{1\,mol\,Cu}{63.55\,g\,Cu} \qquad \frac{6.022 \times 10^{23}\,Cu\,atoms}{1\,mol\,Cu}$$ RELATIONSHIPS USED 63.55 g Cu = 1 mol Cu (molar mass of copper) 6.022×10^{23} = 1 mol (Avogadro's number)

| SOLVE Follow the conceptual plan to solve the problem. Begin with 3.10 g Cu and multiply by the appropriate conversion factors to arrive at the number of Cu atoms. | SOLUTION $$3.10 \text{ g Cu} \times \frac{1 \text{ mol C}}{63.55 \text{ g Cu}} \times \frac{6.022 \times 10^{23} \text{ Cu atoms}}{1 \text{ mol Cu}} = 2.94 \times 10^{22} \text{ Cu atoms}$$ |

CHECK The answer (the number of copper atoms) is less than 6.022×10^{23} (1 mol). This is consistent with the given mass of copper atoms, which is less than the molar mass of copper.

FOR PRACTICE 2.8   How many carbon atoms are there in a 1.3-carat diamond? Diamonds are a form of pure carbon. (1 carat = 0.20 g)

FOR MORE PRACTICE 2.8   Calculate the mass of 2.25×10^{22} tungsten atoms.

Notice that numbers with large exponents, such as 6.022×10^{23}, are almost unbelievably large. Twenty-two copper pennies contain 6.022×10^{23} or 1 mol of copper atoms, but 6.022×10^{23} pennies would cover Earth's entire surface to a depth of 300 m. Even objects small by everyday standards occupy a huge space when we have a mole of them. For example, a grain of sand has a mass of less than 1 mg and a diameter of less than 0.1 mm, yet 1 mol of sand grains would cover the state of Texas to a depth of several feet. For every increase of 1 in the exponent of a number, the number increases by a factor of 10, so 10^{23} is incredibly large. Of course 1 mol has to be a large number if it is to have practical value because atoms are so small.

Example 2.9   The Mole Concept

An aluminum sphere contains 8.55×10^{22} aluminum atoms. What is the sphere's radius in centimeters? The density of aluminum is 2.70 g/cm^3.

SORT You are given the number of aluminum atoms in a sphere and the density of aluminum. You are asked to find the radius of the sphere.	GIVEN: 8.55×10^{22} Al atoms $d = 2.70 \text{ g/cm}^3$ FIND: radius (r) of sphere
STRATEGIZE The heart of this problem is density, which relates mass to volume; though you aren't given the mass directly, you are given the number of atoms, which you can use to find mass. 1. Convert from number of atoms to number of moles using Avogadro's number as a conversion factor. 2. Convert from number of moles to mass using molar mass as a conversion factor. 3. Convert from mass to volume (in cm^3) using density as a conversion factor. 4. Once you calculate the volume, find the radius from the volume using the formula for the volume of a sphere.	CONCEPTUAL PLAN Number of Al atoms → mol Al → g Al → V (in cm^3) $\dfrac{1 \text{ mol Al}}{6.022 \times 10^{23} \text{ Al atoms}}$ $\dfrac{26.98 \text{ g Al}}{1 \text{ mol Al}}$ $\dfrac{1 \text{ cm}^3}{2.70 \text{ g Al}}$ V (in cm^3) → r $V = \dfrac{4}{3}\pi r^3$ RELATIONSHIPS AND EQUATIONS USED $6.022 \times 10^{23} = 1$ mol (Avogadro's number) $26.98 \text{ g Al} = 1$ mol Al (molar mass of aluminum) 2.70 g/cm^3 (density of aluminum) $V = \dfrac{4}{3}\pi r^3$ (volume of a sphere)

—Continued

SOLVE	SOLUTION
Finally, follow the conceptual plan to solve the problem. Begin with 8.55×10^{22} Al atoms and multiply by the appropriate conversion factors to arrive at volume in cm^3.	$8.55 \times 10^{22} \text{ Al atoms} \times \dfrac{1 \text{ mol Al}}{6.022 \times 10^{23} \text{ Al atoms}}$ $\times \dfrac{26.98 \text{ g Al}}{1 \text{ mol Al}} \times \dfrac{1 \text{ cm}^3}{2.70 \text{ g Al}} = 1.4\underline{1}87 \text{ cm}^3$
Then solve the equation for the volume of a sphere for r and substitute the volume to calculate r.	$V = \dfrac{4}{3}\pi r^3$ $r = \sqrt[3]{\dfrac{3V}{4\pi}} = \sqrt[3]{\dfrac{3(1.4\underline{1}87 \text{ cm}^3)}{4\pi}} = 0.697 \text{ cm}$

CHECK The units of the answer (cm) are correct. The magnitude cannot be estimated accurately, but a radius of about one-half of a centimeter is reasonable for just over one-tenth of a mole of aluminum atoms.

FOR PRACTICE 2.9 A titanium cube contains 2.86×10^{23} atoms. What is the edge length of the cube? The density of titanium is 4.50 g/cm^3.

FOR MORE PRACTICE 2.9 Find the number of atoms in a copper rod with a length of 9.85 cm and a radius of 1.05 cm. The density of copper is 8.96 g/cm^3.

CONCEPTUAL
CONNECTION **2.7**

Avogadro's Number Why is Avogadro's number defined as 6.022×10^{23} and not a simpler round number such as 1.00×10^{23}?

CONCEPTUAL
CONNECTION **2.8**

The Mole Without doing any calculations, determine which sample contains the most atoms.
  (a) a 1-g sample of copper
  (b) a 1-g sample of carbon
  (c) a 10-g sample of uranium

Self-Assessment Quiz

Q1. Two samples of a compound containing elements A and B are decomposed. The first sample produces 15 g of A and 35 g of B. The second sample produces 25 g of A and what mass of B?

  a) 11 g   b) 58 g   c) 21 g   d) 45 g

Q2. A compound containing only carbon and hydrogen has a carbon-to-hydrogen mass ratio of 11.89. Which carbon-to-hydrogen mass ratio is possible for another compound composed only of carbon and hydrogen?

  a) 2.50   b) 3.97   c) 4.66   d) 7.89

Q3. Which idea came out of Rutherford's gold foil experiment?

  a) Atoms contain protons and neutrons.
  b) Matter is composed of atoms.
  c) Elements have isotopes.
  d) Atoms are mostly empty space.

Q4. A student re-creates the Millikan oil drop experiment and tabulates the relative charges of the oil drops in terms of a constant, α.

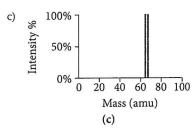| Drop #1	α
Drop #2	$\dfrac{3}{2}\alpha$
Drop #3	$\dfrac{5}{2}\alpha$
Drop #4	3α
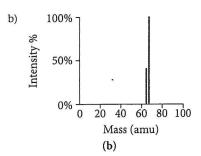
What charge for the electron (in terms of α) is consistent with these data?

  a) $\dfrac{1}{2}\alpha$ b) α c) $\dfrac{3}{2}\alpha$ d) 2α
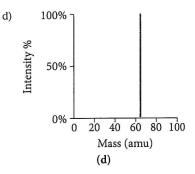
Q5. Determine the number of protons and neutrons in the isotope Fe-58.

a) 26 protons and 58 neutrons
b) 32 protons and 26 neutrons
c) 26 protons and 32 neutrons
d) 58 protons and 58 neutrons

Q6. An isotope of an element contains 82 protons and 122 neutrons. What is the symbol for the isotope?

a) $^{204}_{82}Pb$ b) $^{122}_{82}Pb$ c) $^{122}_{40}Zr$ d) $^{204}_{40}Zr$

Q7. Determine the number of electrons in the Cr^{3+} ion.

a) 24 electrons   b) 27 electrons
c) 3 electrons   d) 21 electrons

Q8. Which pair of elements do you expect to be most similar in their chemical properties?

a) K and Fe   b) O and Si
c) Ne and N   d) Br and I

Q9. Which element is *not* a main-group element?

a) Se   b) Mo   c) Sr   d) Ba

Q10. What is the charge of the ion most commonly formed by S?

a) 2+   b) +   c) −   d) 2−

Q11. A naturally occurring sample of an element contains only two isotopes. The first isotope has a mass of 68.9255 amu and a natural abundance of 60.11%. The second isotope has a mass of 70.9247 amu. Find the atomic mass of the element.

a) 70.13 amu   b) 69.72 amu
c) 84.06 amu   d) 69.93 amu

Q12. Which sample contains the greatest number of atoms?

a) 14 g C   b) 49 g Cr   c) 102 g Ag   d) 202 g Pb

Q13. Determine the number of atoms in 1.85 mL of mercury. (The density of mercury is 13.5 g/mL.)

a) 3.02×10^{27} atoms   b) 4.11×10^{20} atoms
c) 7.50×10^{22} atoms   d) 1.50×10^{25} atoms

Q14. A 20.0 g sample of an element contains 4.95×10^{23} atoms. Identify the element.

a) Cr   b) O   c) Mg   d) Fe

Q15. Copper has two naturally occurring isotopes with masses 62.94 amu and 64.93 amu and has an atomic mass of 63.55 amu. Which mass spectrum is most likely to correspond to a naturally occurring sample of copper?

a)

(a)

b)

(b)

c)

(c)

d)

(d)

Answers: 1. (b) **2.** (b) **3.** (d) **4.** (a) **5.** (c) **6.** (a) **7.** (d) **8.** (d) **9.** (b) **10.** (d) **11.** (b) **12.** (a) **13.** (c) **14.** (c) **15.** (a)

Chapter 2 IN REVIEW

Key Terms

Section 2.3
law of conservation of mass (47)
law of definite proportions (48)
law of multiple proportions (49)
atomic theory (50)

Section 2.4
cathode rays (51)
cathode ray tube (51)
electrical charge (51)
electron (51)

Section 2.5
radioactivity (53)
nuclear theory (54)
nucleus (54)
proton (54)
neutron (55)

Key Concepts

Brownian Motion (2.1)

▷ Brownian motion is the erratic, jittery motion of small particles that was first observed by Robert Brown in 1827. The description of Brownian motion by Einstein in 1905 and confirmation by Perrin in 1908 removed any lingering doubt about the particulate nature of matter.

The Atomic Theory (2.2, 2.3)

▷ Each element is composed of indestructible particles called atoms.
▷ All atoms of a given element have the same mass and other properties.
▷ Atoms combine in simple, whole-number ratios to form compounds.
▷ Atoms of one element cannot change into atoms of another element. In a chemical reaction, atoms change the way that they are bound together with other atoms to form a new substance.

The Electron (2.4)

▷ J. J. Thomson discovered the electron in the late 1800s through experiments with cathode rays. He deduced that electrons are negatively charged, and he measured their charge-to-mass ratio.
▷ Robert Millikan measured the charge of the electron, which—in conjunction with Thomson's results—led to the calculation of the mass of an electron.

The Nuclear Atom (2.5)

▷ In 1909, Ernest Rutherford probed the inner structure of the atom by working with a form of radioactivity called alpha radiation and developed the nuclear theory of the atom.
▷ Nuclear theory states that the atom is mainly empty space, with most of its mass concentrated in a tiny region called the nucleus and most of its volume occupied by relatively light electrons.

Subatomic Particles (2.6)

▷ Atoms are composed of three fundamental particles: the proton (1 amu, +1 charge), the neutron (1 amu, 0 charge), and the electron (~0 amu, −1 charge).

▷ The number of protons in the nucleus of the atom is its atomic number (Z) and defines the element.
▷ The sum of the number of protons and neutrons is the mass number (A).
▷ Atoms of an element that have different numbers of neutrons (and therefore different mass numbers) are isotopes.
▷ Atoms that lose or gain electrons become charged and a ions. Cations are positively charged and anions are negative charged.

The Periodic Table (2.7)

▷ The periodic table tabulates all known elements in order of i creasing atomic number.
▷ The periodic table is arranged so that similar elements are grouped together in columns.
▷ Elements on the left side and in the center of the periodic table are metals and tend to lose electrons in chemical changes.
▷ Elements on the upper right side of the periodic table are non-metals and tend to gain electrons in chemical changes.
▷ Elements located on the boundary between metals and nonmeta are metalloids.

Atomic Mass and the Mole (2.8, 2.9)

▷ The atomic mass of an element, listed directly below its symb in the periodic table, is a weighted average of the masses of t naturally occurring isotopes of the element.
▷ One mole of an element is the amount of that element that cor tains Avogadro's number (6.022×10^{23}) of atoms.
▷ Any sample of an element with a mass (in grams) that equals atomic mass contains one mole of the element. For example, t atomic mass of carbon is 12.011 amu; therefore, 12.011 g of carb contains 1 mol of carbon atoms.

Key Equations and Relationships

Relationship between Mass Number (A), Number of Protons (p), and Number of Neutrons (n) (**2.6**)

$$A = \text{number of protons (p)} + \text{number of neutrons (n)}$$

Atomic Mass (**2.8**)

$$\text{Atomic mass} = \sum_{n}(\text{fraction of isotope } n) \times (\text{mass of isotope } n)$$

Avogadro's Number (**2.9**)

$$1 \text{ mol} = 6.0221421 \times 10^{23} \text{ particles}$$

Key Learning Outcomes

CHAPTER OBJECTIVES	ASSESSMENT
Using the Law of Definite Proportions (**2.3**)	Example 2.1 For Practice 2.1 Exercises 31, 32
Using the Law of Multiple Proportions (**2.3**)	Example 2.2 For Practice 2.2 Exercises 35–38
Working with Atomic Numbers, Mass Numbers, and Isotope Symbols (**2.6**)	Example 2.3 For Practice 2.3 Exercises 51–58
Predicting the Charge of Ions (**2.7**)	Example 2.4 For Practice 2.4 Exercises 59–62
Calculating Atomic Mass (**2.8**)	Example 2.5 For Practice 2.5 For More Practice 2.5 Exercises 71, 72, 74–77
Converting between Moles and Number of Atoms (**2.9**)	Example 2.6 For Practice 2.6 Exercises 81, 82
Converting between Mass and Amount (in Moles) (**2.9**)	Example 2.7 For Practice 2.7 For More Practice 2.7 Exercises 83, 84
Using the Mole Concept (**2.9**)	Examples 2.8, 2.9 For Practice 2.8, 2.9 For More Practice 2.8, 2.9 Exercises 85–94, 112, 113

EXERCISES

Review Questions

1. What is Brownian motion? How is it related to the development of the idea that matter is particulate?

2. Summarize the history of the atomic idea. How was Dalton able to convince others to accept an idea that had been controversial for 2000 years?

3. State and explain the law of conservation of mass.

4. State and explain the law of definite proportions.

5. State and explain the law of multiple proportions. How is the law of multiple proportions different from the law of definite proportions?

6. What are the main ideas in Dalton's atomic theory? How do they help explain the laws of conservation of mass, of constant composition, and of definite proportions?

7. How and by whom was the electron discovered? What basic properties of the electron were reported with its discovery?

8. Explain Millikan's oil drop experiment and how it led to the measurement of the electron's charge. Why is the magnitude of the charge of the electron so important?

9. Describe the plum-pudding model of the atom.

10. Describe Rutherford's gold foil experiment. How did the experiment prove that the plum-pudding model of the atom was wrong?

11. Describe Rutherford's nuclear model of the atom. What was revolutionary about his model?

12. If matter is mostly empty space, as suggested by Rutherford, then why does it appear so solid?

13. List the three subatomic particles that compose atoms and give the basic properties (mass and charge) of each.

14. What defines an element?

15. Explain the difference between Z (the atomic number) and A (the mass number).

16. Where do elements get their names?

17. What are isotopes? What is percent natural abundance of isotopes?

18. Describe the two different notations used to specify isotopes and give an example of each.

19. What is an ion? A cation? An anion?

20. State the periodic law. How did the periodic law lead to the periodic table?

21. Describe the characteristic properties of metals, nonmetals, and metalloids.

22. List the characteristic properties of each group.
a. noble gases
b. alkali metals
c. alkaline earth metals
d. halogens

23. How do you predict the charges of ions formed by main-group elements?

24. What is atomic mass? How is it calculated?

25. Explain how a mass spectrometer works.

26. What kind of information can be determined from a mass spectrum?

27. What is a mole? How is the mole concept useful in chemical calculations?

28. Why is the mass corresponding to a mole of one element different from the mass corresponding to a mole of another element?

Problems by Topic

Note: Answers to all odd-numbered Problems, numbered in blue, can be found in Appendix III. Exercises in the Problems by Topic section are paired, with each odd-numbered problem followed by a similar even-numbered problem. Exercises in the Cumulative Problems section are also paired, but somewhat more loosely. (Challenge Problems and Conceptual Problems, because of their nature, are unpaired.)

The Laws of Conservation of Mass, Definite Proportions, and Multiple Proportions

29. A hydrogen-filled balloon is ignited and 1.50 g of hydrogen is reacted with 12.0 g of oxygen. How many grams of water vapor form? (Assume that water vapor is the only product.)

30. An automobile gasoline tank holds 21 kg of gasoline. When the gasoline burns, 84 kg of oxygen is consumed, and carbon dioxide and water are produced. What is the total combined mass of carbon dioxide and water that is produced?

31. Two samples of carbon tetrachloride are decomposed into their constituent elements. One sample produces 38.9 g of carbon and 448 g of chlorine, and the other sample produces 14.8 g of carbon and 134 g of chlorine. Are these results consistent with the law of definite proportions? Explain your answer.

32. Two samples of sodium chloride are decomposed into their constituent elements. One sample produces 6.98 g of sodium and 10.7 g of chlorine, and the other sample produces 11.2 g of sodium and 17.3 g of chlorine. Are these results consistent with the law of definite proportions? Explain your answer.

33. The mass ratio of sodium to fluorine in sodium fluoride is 1.21:1. A sample of sodium fluoride produces 28.8 g of sodium upon decomposition. How much fluorine (in grams) forms?

34. Upon decomposition, one sample of magnesium fluoride produces 1.65 kg of magnesium and 2.57 kg of fluorine. A second sample produces 1.32 kg of magnesium. How much fluorine (in grams) does the second sample produce?

35. Two different compounds containing osmium and oxygen have the following masses of oxygen per gram of osmium: 0.168 and 0.3369 g. Show that these amounts are consistent with the law of multiple proportions.

36. Palladium forms three different compounds with sulfur. The mass of sulfur per gram of palladium in each compound is listed here. Show that these masses are consistent with the law of multiple proportions.

Compound	Grams S per Gram Pd
A	0.603
B	0.301
C	0.151

37. Sulfur and oxygen form both sulfur dioxide and sulfur trioxide. When samples of these are decomposed, the sulfur dioxide produces 3.49 g oxygen and 3.50 g sulfur, while the sulfur trioxide produces 6.75 g oxygen and 4.50 g sulfur. Calculate the mass of oxygen per gram of sulfur for each sample and show that these results are consistent with the law of multiple proportions.

38. Sulfur and fluorine form several different compounds including sulfur hexafluoride and sulfur tetrafluoride. Decomposition of a sample of sulfur hexafluoride produces 4.45 g of fluorine and 1.25 g of sulfur, while decomposition of a sample of sulfur tetrafluoride produces 4.43 g of fluorine and 1.87 g of sulfur. Calculate the mass of fluorine per gram of sulfur for each sample and show that these results are consistent with the law of multiple proportions.

Atomic Theory, Nuclear Theory, and Subatomic Particles

39. Which statements are *consistent* with Dalton's atomic theory as it was originally stated? Why?
a. Sulfur and oxygen atoms have the same mass.
b. All cobalt atoms are identical.
c. Potassium and chlorine atoms combine in a 1:1 ratio to form potassium chloride.
d. Lead atoms can be converted into gold.

40. Which statements are *inconsistent* with Dalton's atomic theory as it was originally stated? Why?
a. All carbon atoms are identical.
b. An oxygen atom combines with 1.5 hydrogen atoms to form a water molecule.
c. Two oxygen atoms combine with a carbon atom to form a carbon dioxide molecule.
d. The formation of a compound often involves the destruction of one or more atoms.

41. Which statements are *consistent* with Rutherford's nuclear theory as it was originally stated? Why?
a. The volume of an atom is mostly empty space.
b. The nucleus of an atom is small compared to the size of the atom.
c. Neutral lithium atoms contain more neutrons than protons.
d. Neutral lithium atoms contain more protons than electrons.

42. Which statements are *inconsistent* with Rutherford's nuclear theory as it was originally stated? Why?
a. Since electrons are smaller than protons, and since a hydrogen atom contains only one proton and one electron, it must follow that the volume of a hydrogen atom is mostly due to the proton.
b. A nitrogen atom has seven protons in its nucleus and seven electrons outside of its nucleus.
c. A phosphorus atom has 15 protons in its nucleus and 15 electrons outside of its nucleus.
d. The majority of the mass of a fluorine atom is due to its nine electrons.

43. A chemist in an imaginary universe, where electrons have a different charge than they do in our universe, performs the Millikan oil drop experiment to measure the electron's charge. The charges of several drops are recorded here. What is the charge of the electron in this imaginary universe?

Drop #	Charge
A	-6.9×10^{-19} C
B	-9.2×10^{-19} C
C	-11.5×10^{-19} C
D	-4.6×10^{-19} C

44. Imagine a unit of charge called the *zorg*. A chemist performs the oil drop experiment and measures the charge of each drop in zorgs. Based on the results shown here, what is the charge of the electron in zorgs (z)? How many electrons are in each drop?

Drop #	Charge
A	-4.8×10^{-9} z
B	-9.6×10^{-9} z
C	-6.4×10^{-9} z
D	-12.8×10^{-9} z

45. On a dry day, your body can accumulate static charge from walking across a carpet or from brushing your hair. If your body develops a charge of -15 μC (microcoulombs), how many excess electrons has it acquired? What is their collective mass?

46. How many electrons are necessary to produce a charge of -1.0 C? What is the mass of this many electrons?

47. Which statements about subatomic particles are true?
    a. If an atom has an equal number of protons and electrons, it will be charge-neutral.
    b. Electrons are attracted to protons.
    c. Electrons are much lighter than neutrons.
    d. Protons have twice the mass of neutrons.

48. Which statements about subatomic particles are false?
    a. Protons and electrons have charges of the same magnitude but opposite signs.
    b. Protons have about the same mass as neutrons.
    c. Some atoms don't have any protons.
    d. Protons and neutrons have charges of the same magnitude but opposite signs.

49. How many electrons does it take to equal the mass of a proton?

50. A helium nucleus has two protons and two neutrons. How many electrons does it take to equal the mass of a helium nucleus?

Isotopes and Ions

51. Write isotopic symbols in the form X-*A* (e.g., C-13) for each isotope.
    a. the silver isotope with 60 neutrons
    b. the silver isotope with 62 neutrons
    c. the uranium isotope with 146 neutrons
    d. the hydrogen isotope with one neutron

52. Write isotopic symbols in the form $^A_Z X$ for each isotope.
    a. the copper isotope with 34 neutrons
    b. the copper isotope with 36 neutrons
    c. the potassium isotope with 21 neutrons
    d. the argon isotope with 22 neutrons

53. Determine the number of protons and the number of neutrons in each isotope.
    a. $^{14}_{7}N$ b. $^{23}_{11}Na$ c. $^{222}_{86}Rn$ d. $^{208}_{82}Pb$

54. Determine the number of protons and the number of neutrons in each isotope.
    a. $^{40}_{19}K$ b. $^{226}_{88}Ra$ c. $^{99}_{43}Tc$ d. $^{33}_{15}P$

55. The amount of carbon-14 in ancient artifacts and fossils is often used to establish their age. Determine the number of protons and the number of neutrons in carbon-14 and write its symbol in the form $^A_Z X$.

56. Uranium-235 is used in nuclear fission. Determine the number of protons and the number of neutrons in uranium-235 and write its symbol in the form $^A_Z X$.

57. Determine the number of protons and the number of electrons in each ion.
    a. Ni^{2+} b. S^{2-} c. Br^- d. Cr^{3+}

58. Determine the number of protons and the number of electrons in each ion.
    a. Al^{3+} b. Se^{2-} c. Ga^{3+} d. Sr^{2+}

59. Predict the charge of the ion formed by each element.
    a. O   b. K   c. Al   d. Rb

60. Predict the charge of the ion formed by each element.
    a. Mg   b. N   c. F   d. Na

61. Fill in the blanks to complete the table.

Symbol	Ion Formed	Number of Electrons in Ion	Number of Protons in Ion
Ca	Ca^{2+}	____	____
____	Be^{2+}	2	____
Se	____	____	34
In	____	____	49

62. Fill in the blanks to complete the table.

Symbol	Ion Formed	Number of Electrons in Ion	Number of Protons in Ion
Cl	____	____	17
Te	____	54	____
Br	Br^-	____	____
____	Sr^{2+}	____	38

The Periodic Table and Atomic Mass

63. Write the name of each element and classify it as a metal, nonmetal, or metalloid.
    a. K   b. Ba   c. I   d. O   e. Sb

64. Write the symbol for each element and classify it as a metal, nonmetal, or metalloid.
    a. gold
    b. fluorine
    c. sodium
    d. tin
    e. argon

65. Determine whether or not each element is a main-group element.
    a. tellurium
    b. potassium
    c. vanadium
    d. manganese

66. Determine whether or not each element is a transition element.
   a. Cr
   b. Br
   c. Mo
   d. Cs

67. Classify each element as an alkali metal, alkaline earth metal, halogen, or noble gas.
   a. sodium
   b. iodine
   c. calcium
   d. barium
   e. krypton

68. Classify each element as an alkali metal, alkaline earth metal, halogen, or noble gas.
   a. F
   b. Sr
   c. K
   d. Ne
   e. At

69. Which pair of elements do you expect to be most similar? Why?
   a. N and Ni
   b. Mo and Sn
   c. Na and Mg
   d. Cl and F
   e. Si and P

70. Which pair of elements do you expect to be most similar? Why?
   a. nitrogen and oxygen
   b. titanium and gallium
   c. lithium and sodium
   d. germanium and arsenic
   e. argon and bromine

71. Gallium has two naturally occurring isotopes with the following masses and natural abundances:

Isotope	Mass (amu)	Abundance (%)
Ga-69	68.92558	60.108
Ga-71	70.92470	39.892

Sketch the mass spectrum of gallium.

72. Magnesium has three naturally occurring isotopes with the following masses and natural abundances:

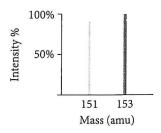| Isotope	Mass (amu)	Abundance (%)
Mg-24	23.9850	78.99
Mg-25	24.9858	10.00
Mg-26	25.9826	11.01

Sketch the mass spectrum of magnesium.

73. The atomic mass of fluorine is 18.998 amu, and its mass spectrum shows a large peak at this mass. The atomic mass of chlorine is 35.45 amu, yet the mass spectrum of chlorine does not show a peak at this mass. Explain the difference.

74. The atomic mass of copper is 63.546 amu. Do any copper isotopes have a mass of 63.546 amu? Explain.

75. An element has two naturally occurring isotopes. Isotope 1 has a mass of 120.9038 amu and a relative abundance of 57.4%, and isotope 2 has a mass of 122.9042 amu. Find the atomic mass of this element and identify it.
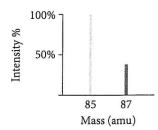
76. An element has four naturally occurring isotopes with the masses and natural abundances given here. Find the atomic mass of the element and identify it.

Isotope	Mass (amu)	Abundance (%)
1	135.90714	0.19
2	137.90599	0.25
3	139.90543	88.43
4	141.90924	11.13

77. Bromine has two naturally occurring isotopes (Br-79 and Br-81) and has an atomic mass of 79.904 amu. The mass of Br-81 is 80.9163 amu, and its natural abundance is 49.31%. Calculate the mass and natural abundance of Br-79.

78. Silicon has three naturally occurring isotopes (Si-28, Si-29, and Si-30). The mass and natural abundance of Si-28 are 27.9769 amu and 92.2%, respectively. The mass and natural abundance of Si-29 are 28.9765 amu and 4.67%, respectively. Find the mass and natural abundance of Si-30.

79. Use the mass spectrum of europium to determine the atomic mass of europium.

80. Use the mass spectrum of rubidium to determine the atomic mass of rubidium.

The Mole Concept

81. How many sulfur atoms are there in 5.52 mol of sulfur?

82. How many moles of aluminum do 3.7×10^{24} aluminum atoms represent?

83. What is the amount, in moles, of each elemental sample?
   a. 11.8 g Ar   b. 3.55 g Zn
   c. 26.1 g Ta   d. 0.211 g Li

84. What is the mass, in grams, of each elemental sample?
   a. 2.3×10^{-3} mol Sb
   b. 0.0355 mol Ba
   c. 43.9 mol Xe
   d. 1.3 mol W

85. How many silver atoms are there in 3.78 g of silver?

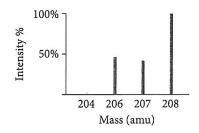**86.** What is the mass of 4.91×10^{21} platinum atoms?

87. Calculate the number of atoms in each sample.
    a. 5.18 g P   b. 2.26 g Hg
    c. 1.87 g Bi   d. 0.082 g Sr

88. Calculate the number of atoms in each sample.
    a. 14.955 g Cr   b. 39.733 g S
    c. 12.899 g Pt   d. 97.552 g Sn

89. Calculate the mass, in grams, of each sample.
    a. 1.1×10^{23} gold atoms
    b. 2.82×10^{22} helium atoms
    c. 1.8×10^{23} lead atoms
    d. 7.9×10^{21} uranium atoms

90. Calculate the mass, in kg, of each sample.
    a. 7.55×10^{26} cadmium atoms
    b. 8.15×10^{27} nickel atoms
    c. 1.22×10^{27} manganese atoms
    d. 5.48×10^{29} lithium atoms

91. How many carbon atoms are there in a diamond (pure carbon) with a mass of 52 mg?

92. How many helium atoms are there in a helium blimp containing 536 kg of helium?

93. Calculate the average mass, in grams, of one platinum atom.

94. Using scanning tunneling microscopy, scientists at IBM wrote the initials of their company with 35 individual xenon atoms (as shown below). Calculate the total mass of these letters in grams.
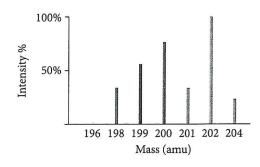

Cumulative Problems

95. A 7.83 g sample of HCN contains 0.290 g of H and 4.06 g of N. Find the mass of carbon in a sample of HCN with a mass of 3.37 g.

96. The ratio of sulfur to oxygen by mass in SO_2 is 1.0:1.0.
    a. Find the ratio of sulfur to oxygen by mass in SO_3.
    b. Find the ratio of sulfur to oxygen by mass in S_2O.

97. The ratio of oxygen to carbon by mass in carbon monoxide is 1.33:1.00. Find the formula of an oxide of carbon in which the ratio by mass of oxygen to carbon is 2.00:1.00.

98. The ratio of the mass of a nitrogen atom to the mass of an atom of ^{12}C is 7:6, and the ratio of the mass of nitrogen to oxygen in N_2O is 7:4. Find the mass of 1 mol of oxygen atoms.

99. An α particle, $^{4}He^{2+}$, has a mass of 4.00151 amu. Find the value of its charge-to-mass ratio in C/kg.

100. Naturally occurring iodine has an atomic mass of 126.9045 amu. A 12.3849 g sample of iodine is accidentally contaminated with an additional 1.00070 g of ^{129}I, a synthetic radioisotope of iodine used in the treatment of certain diseases of the thyroid gland. The mass of ^{129}I is 128.9050 amu. Find the apparent "atomic mass" of the contaminated iodine.

101. Use the mass spectrum of lead to estimate the atomic mass of lead. Estimate the mass and percent intensity values from the graph to three significant figures.

102. Use the mass spectrum of mercury to estimate the atomic mass of mercury. Estimate the masses and percent intensity values from the graph to three significant figures.

103. Nuclei with the same number of *neutrons* but different mass numbers are called *isotones*. Write the symbols of four isotones of ^{236}Th.

104. Fill in the blanks to complete the table.

Symbol	Z	A	Number of p	Number of e⁻	Number of n	Charge
Si	14	___	___	14	14	___
S²⁻	___	32	___	___	___	2−
Cu²⁺	___	___	___	___	34	2+
___	15	___	___	15	16	___

105. Fill in the blanks to complete the table.

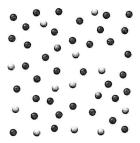| Symbol	Z	A	Number of p	Number of e⁻	Number of n	Charge
___	8	___	___	___	8	2−
Ca²⁺	20	___	___	___	20	___
Mg²⁺	___	25	___	___	13	2+
N³⁻	___	14	___	10	___	___

106. Neutron stars are composed of solid nuclear matter, primarily neutrons. Assume the radius of a neutron is approximately 1.0×10^{-13} cm. Calculate the density of a neutron. [*Hint*: For a sphere $V = (4/3)\pi r^3$.] Assuming that a neutron star has the same density as a neutron, calculate the mass (in kg) of a small piece of a neutron star the size of a spherical pebble with a radius of 0.10 mm.

107. Carbon-12 contains six protons and six neutrons. The radius of the nucleus is approximately 2.7 fm (femtometers) and the radius of the atom is approximately 70 pm (picometers). Calculate the volume of the nucleus and the volume of the atom. What percentage of the carbon atom's volume is occupied by the nucleus? (Assume two significant figures.)

108. A penny has a thickness of approximately 1.0 mm. If you stacked Avogadro's number of pennies one on top of the other on Earth's surface, how far would the stack extend (in km)? [For comparison, the sun is about 150 million km from Earth and the nearest star (Proxima Centauri) is about 40 trillion km from Earth.]

109. Consider the stack of pennies in the previous problem. How much money (in dollars) would this represent? If this money were equally distributed among the world's population of 7.0 billion people, how much would each person receive? Would each person be a millionaire? A billionaire? A trillionaire?

110. The mass of an average blueberry is 0.75 g and the mass of an automobile is 2.0×10^3 kg. Find the number of automobiles whose total mass is the same as 1.0 mol of blueberries.

111. Suppose that atomic masses were based on the assignment of a mass of 12.000 g to 1 mol of carbon, rather than 1 mol of ^{12}C. What would the atomic mass of oxygen be? (The atomic masses of carbon and oxygen based on the assignment of 12.000 g to 1 mol of ^{12}C are 12.011 amu and 15.9994 amu, respectively.)

112. A pure titanium cube has an edge length of 2.78 in. How many titanium atoms does it contain? Titanium has a density of 4.50 g/cm^3.

113. A pure copper sphere has a radius of 0.935 in. How many copper atoms does it contain? [The volume of a sphere is $(4/3)\pi r^3$ and the density of copper is 8.96 g/cm^3.]

114. What is the radius (in cm) of a pure copper sphere that contains 1.14×10^{24} copper atoms? [The volume of a sphere is $(4/3)\pi r^3$ and the density of copper is 8.96 g/cm^3.]

115. What is the edge length (in cm) of a titanium cube that contains 2.55×10^{24} titanium atoms? The density of titanium 4.50 g/cm^3.

116. Boron has only two naturally occurring isotopes. The mass boron-10 is 10.01294 amu and the mass of boron-11 is 11.009 amu. Calculate the relative abundances of the two isotopes.

117. Lithium has only two naturally occurring isotopes. The m of lithium-6 is 6.01512 amu and the mass of lithium is 7.01601 amu. Calculate the relative abundances of the t isotopes.

118. Common brass is a copper and zinc alloy containing 37.0% z by mass and having a density of 8.48 g/cm^3. A fitting compo of common brass has a total volume of 112.5 cm^3. How ma atoms (copper and zinc) does the fitting contain?

119. A 67.2 g sample of a gold and palladium alloy conta 2.49×10^{23} atoms. What is the composition (by mass) of t alloy?

120. Naturally occurring chlorine is composed of two isotop 75.76% Cl-35 (mass 34.9688 amu) and 24.24% Cl-37 (m 36.9659 amu). Naturally occurring oxygen is composed of th isotopes: 99.757% O-16 (mass 15.9949 amu), 0.038% O- (mass 16.9991 amu), and 0.205% O-18 (mass 17.9991 amu). T compound dichlorine monoxide is composed of two chlori atoms and one oxygen atom bonded together to form the Cl molecule. How many Cl$_2$O molecules of different masses na rally exist? Give the masses of the three most abundant Cl molecules.

121. Silver is composed of two naturally occurring isotopes: Ag-1 (51.839%) and Ag-109. The ratio of the masses of the two is topes is 1.0187. What is the mass of Ag-107?

122. The U.S. Environmental Protection Agency (EPA) sets limits healthful levels of air pollutants. The maximum level that t EPA considers safe for lead air pollution is 1.5 μg/m^3. If yo lungs were filled with air containing this level of lead, h many lead atoms would be in your lungs? (Assume a total lu volume of 5.50 L.)

123. Pure gold is usually too soft for jewelry, so it is often alloy with other metals. How many gold atoms are in an 0.25 ounce, 18 K gold bracelet? (18 K gold is 75% gold by mass.)

Challenge Problems

124. In Section 2.9, it was stated that 1 mol of sand grains would cover the state of Texas to several feet. Estimate how many feet by assuming that the sand grains are roughly cube-shaped, each one with an edge length of 0.10 mm. Texas has a land area of 268,601 square miles.

125. Use the concepts in this chapter to obtain an estimate for the number of atoms in the universe. Make the following assumptions: **(a)** All of the atoms in the universe are hydrogen atoms in stars. (This is not a ridiculous assumption because over three-fourths of the atoms in the universe are in fact hydrogen. Gas and dust between the stars represent only about 15% of the visible matter of our galaxy, and planets compose a far tinier fraction.) **(b)** The sun is a typical star composed of pure hydrogen with a density of 1.4 g/cm^3 and a radius of 7×10^8 m.

(c) Each of the roughly 100 billion stars in the Milky Way ga axy contains the same number of atoms as our sun. **(d)** Each the 10 billion galaxies in the visible universe contains the san number of atoms as our Milky Way galaxy.

126. On the previous page is a representation of 50 atoms of a fictitious element called westmontium (Wt). The red spheres represent Wt-296, the blue spheres Wt-297, and the green spheres Wt-298.

    a. Assuming that the sample is statistically representative of a naturally occurring sample, calculate the percent natural abundance of each Wt isotope.

    b. Draw the mass spectrum for a naturally occurring sample of Wt.

    c. The mass of each Wt isotope is measured relative to C-12 and tabulated. Use the mass of C-12 to convert each of the masses to amu and calculate the atomic mass of Wt.

Isotope	Mass
Wt-296	$24.6630 \times$ Mass(^{12}C)
Wt-297	$24.7490 \times$ Mass(^{12}C)
Wt-298	$24.8312 \times$ Mass(^{12}C)

127. The ratio of oxygen to nitrogen by mass in NO_2 is 2.29. The ratio of fluorine to nitrogen by mass in NF_3 is 4.07. Find the ratio of oxygen to fluorine by mass in OF_2.

128. Naturally occurring cobalt consists of only one isotope, ^{59}Co, whose relative atomic mass is 58.9332. A synthetic radioactive isotope of cobalt, ^{60}Co, has a relative atomic mass of 59.9338 and is used in radiation therapy for cancer. A 1.5886 g sample of cobalt has an apparent "atomic mass" of 58.9901. Find the mass of ^{60}Co in this sample.

129. A 7.36 g sample of copper is contaminated with an additional 0.51 g of zinc. Suppose an atomic mass measurement was performed on this sample. What would be the measured atomic mass?

130. The ratio of the mass of O to the mass of N in N_2O_3 is 12:7. Another binary compound of nitrogen has a ratio of O to N of 16:7. What is its formula? What is the ratio of O to N in the next member of this series of compounds?

131. Naturally occurring magnesium has an atomic mass of 24.312 and consists of three isotopes. The major isotope is ^{24}Mg, natural abundance 78.99%, relative atomic mass 23.98504. The next most abundant isotope is ^{26}Mg, relative atomic mass 25.98259. The third most abundant isotope is ^{25}Mg, whose natural abundance is in the ratio of 0.9083 to that of ^{26}Mg. Find the relative atomic mass of ^{25}Mg.

Conceptual Problems

132. Which answer is an example of the law of multiple proportions? Explain.

    a. Two different samples of water are found to have the same ratio of hydrogen to oxygen.

    b. When hydrogen and oxygen react, the mass of water formed is exactly equal to the mass of hydrogen and oxygen that reacted.

    c. The mass ratio of oxygen to hydrogen in water is 8:1. The mass ratio of oxygen to hydrogen in hydrogen peroxide (a compound that only contains hydrogen and oxygen) is 16:1.

133. Lithium has two naturally occurring isotopes: Li-6 (natural abundance 7.5%) and Li-7 (natural abundance 92.5%). Using circles to represent protons and squares to represent neutrons, draw the nucleus of each isotope. How many Li-6 atoms are present, on average, in a 1000-atom sample of lithium?

134. As we saw in the previous problem, lithium has two naturally occurring isotopes: Li-6 (natural abundance 7.5%; mass 6.0151 amu) and Li-7 (natural abundance 92.5%; mass

7.0160 amu). Without doing any calculations, determine which mass is closest to the atomic mass of Li.

    a. 6.00 amu   **b.** 6.50 amu   **c.** 7.00 amu

135. The mole is defined as the amount of a substance containing the same number of particles as exactly 12 g of C-12. The amu is defined as 1/12 of the mass of an atom of C-12. Why is it important that both of these definitions reference the same isotope? What would be the result, for example, of defining the mole with respect to C-12, but the amu with respect to Ne-20?

136. Without doing any calculations, determine which of the samples contains the greatest amount of the element in moles. Which contains the greatest mass of the element?

    a. 55.0 g Cr   **b.** 45.0 g Ti   **c.** 60.0 g Zn

137. The atomic radii of the isotopes of an element are identical to one another. However, the atomic radii of the ions of an element are significantly different from the atomic radii of the neutral atom of the element. Explain.

Active Classroom Learning

Questions for Group Work

Discuss these questions with the group and record your consensus answer.

138. The table shown here includes data similar to those used by Mendeleev when he created the periodic table. On a small card, write the symbol, atomic mass, and a stable compound formed by each element. Without consulting a periodic table, arrange the cards so that atomic mass increases from left to right and elements with similar properties are above and below each other. Copy the periodic table you have invented onto a piece of paper. There is one element missing. Predict its mass and a stable compound it might form.

Element	Atomic Mass	Stable Compound
Be	9	$BeCl_2$
S	32	H_2S
F	19	F_2
Ca	40	$CaCl_2$
Li	7	$LiCl$
Si	28	SiH_4
Cl	35.4	Cl_2
B	10.8	BH_3

—Continued

Element	Atomic Mass	Stable Compound
Ge	72.6	GeH_4
N	14	NF_3
O	16	H_2O
Ga	69.7	GaH_3
As	75	AsF_3
C	12	CH_4
K	39	KCl
Mg	24.3	$MgCl_2$
Se	79	H_2Se
Al	27	AlH_3
Br	80	Br_2
Na	23	$NaCl$

139. In a naturally occurring sample, 19.8% of boron atoms ha five neutrons and 80.2% have six neutrons. What is the ma number of each boron isotope? Sketch a sample of 10 atoms th is nearly representative of a natural sample. What is the avera mass of the atoms in your drawing? What is the atomic mass boron? (Boron-10 has a mass of 10.01294 amu and boron-11 h a mass of 11.00931 amu.)

140. In complete sentences, describe the similarities and differenc between:
a. different isotopes of an element
b. a neutral atom and an ion of the same element

141. Calculate the mass in grams of one mole of each of the followir (the mass of a single item is given in parentheses): electro $(9.10938 \times 10^{-28}$ g), protons $(1.67262 \times 10^{-24}$ g), neutro $(1.67493 \times 10^{-24}$ g), atoms of carbon-12 $(1.992646 \times 10^{-23}$ and doughnuts (74 g). Compare the mass of one mole carbon-12 atoms to the sum of the masses of the particles th it contains. If the doughnut mentioned in this question we made entirely of carbon, how many atoms would it contain?

Data Interpretation and Analysis

Heavy Metals in Recycled Paper Packages

142. Demand for recycled paper has increased as consumers have become more aware of the environmental issues surrounding waste disposal. Paper is a natural raw material made from renewable wood and plants. Recycled paper is made from waste paper and paperboard. Both new paper and recycled paper contain traces of heavy metals. However, some types of recycled paper contain more heavy metals than new paper due, in part, to the inks used for printing or adding color to the original paper. Metals can migrate from the paper packaging and containers used for food to the food itself. This metal migration is controlled and monitored to reduce the exposure of humans to heavy metals.

Food agencies in each country determine the limit of cadmium, zinc, nickel, and copper in packaging materials. Most countries impose a limit of heavy metals in recycled paper not to exceed 100.0 ppm or 100.0 mg/kg. The limit for lead in egg-, fruit-, or vegetable-packaging is lower—not to exceed—20.0 mg/kg. The table in Figure a ▼ lists the results of the analysis of samples of recycled paper produced by a manufacturer during a three-month period.

	Zn	Pb	Cd	Ni	Cu
March	2.7	3.5	0.129	1.7	32.3
April	3.2	2.8	0.092	1.6	29.2
May	3.7	2.4	0.096	1.5	19.9

▲ FIGURE a **Heavy Metal Values in Three Samples of Recycled Paper (mg/kg ± 0.1)**

The goal of the manufacturer is to reduce the amount heavy metals—especially lead—in the recycled paper it produ es and sells. From March to May, the manufacturer varied tl methods of production each month to determine which metl od would produce paper with the lowest metal content. Eac month, lab technicians cut small samples of recycled paper wil an area of 1.000 dm² and a thickness of 0.0500 cm. The techr cians then prepared the samples for analysis. The average densi of the samples is 800.0 ± 0.0245 kg/m³.

Use the information provided in the figure to answer tl following questions:

a. Did the company reduce the amount of lead in its product between March and May?

b. Which metal was not reduced?

c. Each month's sample represents a different manufacturing process. Which process would you recommend the manufacturer choose to continue to use over the next year while additional processes are tested? Why?

d. What is the total amount of the five metals for the April sample (in mg/kg)?

e. What is the total mass (in mg) of the five metals for the Api sample found in the 1.000 dm² × 0.0500 cm sample?

f. Lead has four stable isotopes: ^{204}Pb, ^{206}Pb, ^{207}Pb, ^{208}Pb witl % abundances of 1.40, 24.10, 22.10, and 52.40, respectively Determine the mass (in mg) of ^{208}Pb in the April sample.

g. Sketch the mass spectrum for lead.

Answers to Conceptual Connections

The Law of Conservation of Mass

2.1 Most of the matter that composed the log undergoes a chemical change by reacting with oxygen molecules in the air. The products of the reaction (mostly carbon dioxide and water) are released as gases into the air.

The Laws of Definite and Multiple Proportions

2.2 The law of definite proportions applies to two or more samples of the *same compound* and states that the ratio of one element to the other is always the same. The law of multiple proportions applies to two *different compounds* containing the same two elements (A and B) and states that the masses of B that combine with 1 g of A are always related to each other as a small whole-number ratio.

The Millikan Oil Drop Experiment

2.3 The drop contains three excess electrons $(3 \times (-6 \times 10^{-19} \text{ C}) = -4.8 \times 10^{-19} \text{ C})$.

Isotopes

2.4

C-12 nucleus   C-13 nucleus

A 10,000-atom sample of carbon, on average, contains 107 C-13 atoms.

The Nuclear Atom, Isotopes, and Ions

2.5 (b) The number of neutrons in the nucleus of an atom does not affect the atom's size because the nucleus is miniscule compared to the atom itself.

Atomic Mass

2.6 (a) Since 98.93% of the atoms are C-12, we would expect the atomic mass to be very close to the mass of the C-12 isotope.

Avogadro's Number

2.7 Avogadro's number is defined with respect to carbon-12—it is the number equal to the number of atoms in exactly 12 g of carbon-12. If Avogadro's number was defined as 1.00×10^{23} (a nice round number), it would correspond to 1.99 g of carbon-12 atoms (an inconvenient number). Avogadro's number is defined with respect to carbon-12 because, as you recall from Section 2.6, the amu (the basic mass unit used for all atoms) is defined relative to carbon-12. Therefore, the mass in grams of 1 mol of *any* element is equal to its atomic mass. As we have seen, these two definitions together make it possible to determine the number of atoms in a known mass of any element.

The Mole

2.8 (b) The carbon sample contains more atoms than the copper sample because carbon has a lower molar mass than copper. Carbon atoms are lighter than copper atoms, so a 1-g sample of carbon contains more atoms than a 1-g sample of copper. The carbon sample also contains more atoms than the uranium sample because even though the uranium sample has 10 times the mass of the carbon sample, a uranium atom is more than 10 times as massive (238 g/mol for uranium versus 12 g/mol for carbon).