2

Atoms and the Atomic Theory

CONTENTS

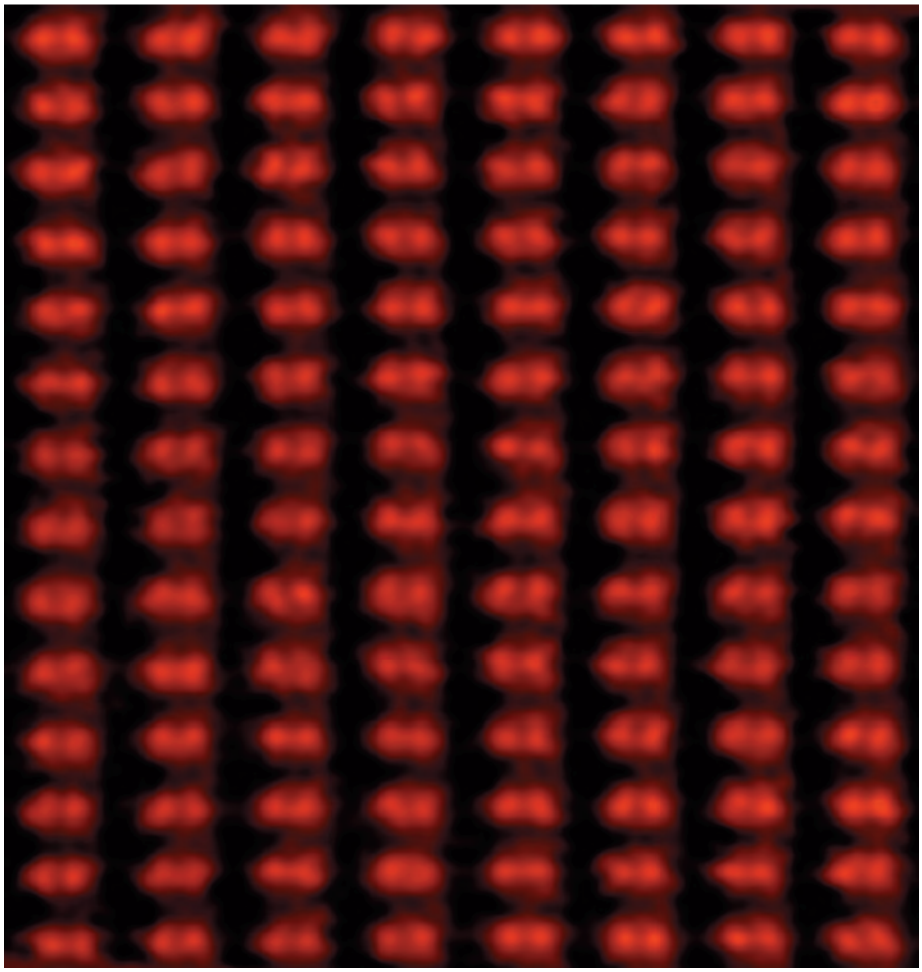


Image of silicon atoms that are only 78 pm apart; image produced by using the latest scanning transmission electron microscope (STEM). The hypothesis that all matter is made up of atoms has existed for more than 2000 years. It is only within the last few decades, however, that techniques have been developed that can render individual atoms visible.

We begin this chapter with a brief survey of early chemical discoveries, culminating in Dalton's atomic theory. This is followed by a description of the physical evidence leading to the modern picture of the *nuclear atom*, in which protons and neutrons are combined into a nucleus with electrons in space surrounding the nucleus. We will also introduce the periodic table as the primary means of organizing elements into groups with similar properties. Finally, we will introduce the concept of the mole and the Avogadro constant, which are the principal tools for counting atoms and molecules and measuring amounts of substances. We will use these tools throughout the text.

2-1 Early Chemical Discoveries and the Atomic Theory

Chemistry has been practiced for a very long time, even if its practitioners were much more interested in its applications than in its underlying principles. The blast furnace for extracting iron from iron ore appeared as early as A.D. 1300, and such important chemicals as sulfuric acid (oil of vitriol), nitric acid (aqua fortis), and sodium sulfate (Glauber's salt) were all well known and used several hundred years ago. Before the end of the eighteenth century, the principal gases of the atmosphere—nitrogen and oxygen—had been isolated, and natural laws had been proposed describing the physical behavior of gases. Yet chemistry cannot be said to have entered the modern age until the process of combustion was explained. In this section, we explore the direct link between the explanation of combustion and Dalton's atomic theory.

▲ FIGURE 2-1
Two combustion reactions
The apparent product of the combustion of the match—the ash—weighs *less* than the match. The product of the combustion of the magnesium ribbon (the "smoke") weighs *more* than the ribbon. Actually, in each case, the total mass remains unchanged. To understand this, you have to know that oxygen gas enters into both combustions and that water and carbon dioxide are also products of the combustion of the match.

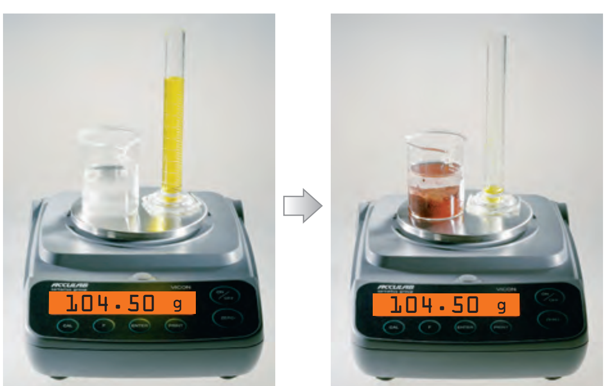
Law of Conservation of Mass

The process of combustion—burning—is so familiar that it is hard to realize what a difficult riddle it posed for early scientists. Some of the difficult-to-explain observations are described in Figure 2-1.

In 1774, Antoine Lavoisier (1743–1794) performed an experiment in which he heated a sealed glass vessel containing a sample of tin and some air. He found that the mass before heating (glass vessel + tin + air) and after heating (glass vessel + "tin calx" + remaining air) were the same. Through further experiments, he showed that the product of the reaction, tin calx (tin oxide), consisted of the original tin together with a portion of the air. Experiments like this proved to Lavoisier that oxygen from air is essential to combustion, and also led him to formulate the **law of conservation of mass**:

> The total mass of substances present after a chemical reaction is the same as the total mass of substances before the reaction.

This law is illustrated in Figure 2-2, where the reaction between silver nitrate and potassium chromate to give a red solid (silver chromate) is monitored by placing the reactants on a single-pan balance—the total mass does not change. Stated another way, the law of conservation of mass says that matter is neither created nor destroyed in a chemical reaction.



(a) (b)

◀ FIGURE 2-2
Mass is conserved during a chemical reaction
(a) Before the reaction, a beaker with a silver nitrate solution and a graduated cylinder with a potassium chromate solution are placed on a single-pan balance, which displays their combined mass—104.50 g. **(b)** When the solutions are mixed, a chemical reaction occurs that forms silver chromate (red precipitate) in a potassium nitrate solution. Note that the total mass—104.50 g—remains unchanged.

EXAMPLE 2-1 Applying the Law of Conservation of Mass

A 0.455 g sample of magnesium is allowed to burn in 2.315 g of oxygen gas. The sole product is magnesium oxide. After the reaction, no magnesium remains and the mass of unreacted oxygen is 2.015 g. What mass of magnesium oxide is produced?

Analyze

The total mass is unchanged. The total mass is the sum of the masses of the substances present initially. The mass of magnesium oxide is the total mass minus the mass of unreacted oxygen.

Solve

First, determine the total mass before the reaction.

$$\begin{aligned}\text{mass before reaction} &= 0.455\text{ g magnesium} + 2.315\text{ g oxygen}\\ &= 2.770\text{ g mass before reaction}\end{aligned}$$

The total mass after the reaction is the same as before the reaction.

$$\begin{aligned}2.770\text{ g mass after reaction} &= ?\text{ g magnesium oxide after reaction}\\ &\quad + 2.015\text{ g oxygen after reaction}\end{aligned}$$

Solve for the mass of magnesium oxide.

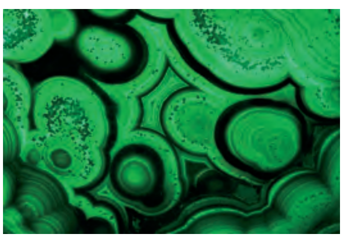
$$\begin{aligned}?\text{ g magnesium oxide after reaction} &= 2.770\text{ g mass after reaction}\\ &\quad - 2.015\text{ g oxygen after reaction}\\ &= 0.755\text{ g magnesium oxide after reaction}\end{aligned}$$

Assess

Here is another approach. The mass of oxygen that reacted is 2.315 g − 2.015 g = 0.300 g. Thus, 0.300 g oxygen combined with 0.455 g magnesium to give 0.300 g + 0.455 g = 0.755 g magnesium oxide.

PRACTICE EXAMPLE A: A 0.382 g sample of magnesium is allowed to react with 2.652 g of nitrogen gas. The sole product is magnesium nitride. After the reaction, the mass of unreacted nitrogen is 2.505 g. What mass of magnesium nitride is produced?

PRACTICE EXAMPLE B: A 7.12 g sample of magnesium is heated with 1.80 g of bromine. All the bromine is used up, and 2.07 g of magnesium bromide is the only product. What mass of magnesium remains *unreacted*?



(a)

(b)

▲ The mineral malachite **(a)** and the green patina on a copper roof **(b)** are both basic copper carbonate, just like the basic copper carbonate prepared by Proust in 1799.

2-1 CONCEPT ASSESSMENT

Jan Baptista van Helmont (1579–1644) weighed a young willow tree and the soil in which the tree was planted. Five years later he found that the mass of soil had decreased by only 0.057 kg, while that of the tree had increased by 75 kg. During that period he had added only water to the bucket in which the tree was planted. Helmont concluded that essentially all the mass gained by the tree had come from the water. Was this a valid conclusion? Explain.

Law of Constant Composition

In 1799, Joseph Proust (1754–1826) reported, "One hundred pounds of copper, dissolved in sulfuric or nitric acids and precipitated by the carbonates of soda or potash, invariably gives 180 pounds of green carbonate."* This and similar observations became the basis of the **law of constant composition**, or the **law of definite proportions**:

> All samples of a compound have the same composition—the same proportions by mass of the constituent elements.

To see how the law of constant composition works, consider the compound water. Water is made up of two atoms of hydrogen (H) for every atom of oxygen (O), a fact that can be represented symbolically by a *chemical formula*, the familiar H_2O.

*The substance Proust produced is actually a more complex substance called *basic* copper carbonate. Proust's results were valid because, like all compounds, basic copper carbonate has a constant composition.

The two samples described below have the same proportions of the two elements, expressed as percentages by mass. To determine the percent by mass of hydrogen, for example, simply divide the mass of hydrogen by the sample mass and multiply by 100%. For each sample, you will obtain the same result: 11.19% H.

Sample A and Its Composition		Sample B and Its Composition	
10.000 g		27.000 g	
1.119 g H	% H = 11.19	3.021 g H	% H = 11.19
8.881 g O	% O = 88.81	23.979 g O	% O = 88.81

EXAMPLE 2-2 Using the Law of Constant Composition

In Example 2-1 we found that when 0.455 g of magnesium reacted with 2.315 g of oxygen, 0.755 g of magnesium oxide was obtained. Determine the mass of magnesium contained in a 0.500 g sample of magnesium oxide.

Analyze

We know that 0.755 g of magnesium oxide contains 0.455 g of magnesium. According to the law of constant composition, the mass ratio 0.455 g magnesium/0.755 g magnesium oxide should exist in all samples of magnesium oxide.

Solve

Application of the law of constant composition gives

$$\frac{0.455 \text{ g magnesium}}{0.755 \text{ g magnesium oxide}} = \frac{? \text{ g magnesium}}{0.500 \text{ g magnesium oxide}}$$

Solving the expression above, we obtain

$$? \text{ g magnesium} = 0.500 \cancel{\text{g magnesium oxide}} \times \frac{0.455 \text{ g magnesium}}{0.755 \cancel{\text{g magnesium oxide}}}$$

$$= 0.301 \text{ g magnesium}$$

Assess

You can also work this problem by using mass percentages. If 0.755 g of magnesium oxide contains 0.455 g of magnesium, then magnesium oxide is $(0.455 \text{ g}/0.755 \text{ g}) \times 100\% = 60.3\%$ magnesium by mass and $(100\% - 60.3\%) = 39.7\%$ oxygen by mass. Thus, a 0.500 g sample of magnesium oxide must contain $0.500 \text{ g} \times 60.3\% = 0.301$ g of magnesium and $0.500 \text{ g} \times 39.7\% = 0.199$ g of oxygen.

PRACTICE EXAMPLE A: What masses of magnesium and oxygen must be combined to make exactly 2.000 g of magnesium oxide?

PRACTICE EXAMPLE B: What substances are present, and what are their masses, after the reaction of 10.00 g of magnesium and 10.00 g of oxygen?

2-2 CONCEPT ASSESSMENT

When 4.15 g magnesium and 82.6 g bromine react, (1) all the magnesium is used up, (2) some bromine remains unreacted, and (3) magnesium bromide is the only product. With this information alone, is it possible to deduce the mass of magnesium bromide produced? Explain.

Dalton's Atomic Theory

From 1803 to 1808, John Dalton, an English schoolteacher, used the two fundamental laws of chemical combination just described as the basis of an atomic theory. His theory involved three assumptions:

1. Each chemical element is composed of minute, indivisible particles called atoms. Atoms can be neither created nor destroyed during a chemical change.

▲ John Dalton (1766–1844), developer of the atomic theory. Dalton has not been considered a particularly good experimenter, perhaps because of his color blindness (a condition sometimes called daltonism). However, he did skillfully use the data of others in formulating his atomic theory. (The Granger Collection)

2. All atoms of an element are alike in mass (weight) and other properties, but the atoms of one element are different from those of all other elements.
3. In each of their compounds, different elements combine in a simple numerical ratio, for example, one atom of A to one of B (AB), or one atom of A to two of B (AB_2).

If atoms of an element are indestructible (assumption 1), then the *same* atoms must be present after a chemical reaction as before. The total mass remains unchanged. Dalton's theory explains the law of conservation of mass. If all atoms of an element are alike in mass (assumption 2) and if atoms unite in *fixed* numerical ratios (assumption 3), the percent composition of a compound must have a unique value, regardless of the origin of the sample analyzed. Dalton's theory also explains the law of constant composition.

Like all good theories, Dalton's atomic theory led to a prediction—the **law of multiple proportions**.

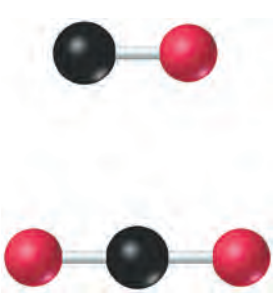
> If two elements form more than a single compound, the masses of one element combined with a fixed mass of the second are in the ratio of small whole numbers.

To illustrate, consider two oxides of carbon (an oxide is a combination of an element with oxygen). In one oxide, 1.000 g of carbon is combined with 1.333 g of oxygen, and in the other, with 2.667 g of oxygen. We see that the second oxide is richer in oxygen; in fact, it contains twice as much oxygen as the first, 2.667 g/1.333 g = 2.00. We now know that the first oxide corresponds to the formula CO and the second, CO_2 (Fig. 2-3).

KEEP IN MIND

that all we know is that the second oxide is twice as rich in oxygen as the first. If the first is CO, the possibilities for the second are CO_2, C_2O_4, C_3O_6, and so on. (See also Exercise 18.)

The characteristic relative masses of the atoms of the various elements became known as atomic weights, and throughout the nineteenth century, chemists worked at establishing reliable values of relative atomic weights. Mostly, however, chemists directed their attention to discovering new elements, synthesizing new compounds, developing techniques for analyzing materials, and in general, building up a vast body of chemical knowledge. Efforts to unravel the structure of the atom became the focus of physicists, as we see in the next several sections.

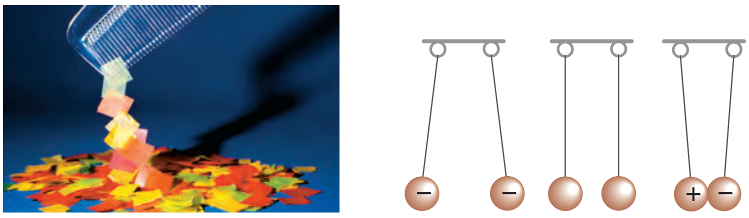


▲ FIGURE 2-3
Molecules CO and CO_2 illustrating the law of multiple proportions
The mass of carbon is the same in the two molecules, but the mass of oxygen in CO_2 is twice the mass of oxygen in CO. Thus, in accordance with the law of multiple proportions, the masses of oxygen in the two compounds, relative to a fixed mass of carbon, are in a ratio of small whole numbers, 2:1.

2-2 Electrons and Other Discoveries in Atomic Physics

Fortunately, we can acquire a qualitative understanding of atomic structure without having to retrace all the discoveries that preceded atomic physics. We do, however, need a few key ideas about the interrelated phenomena of electricity and magnetism, which we briefly discuss here. Electricity and magnetism were used in the experiments that led to the current theory of atomic structure.

Certain objects display a property called electric charge, which can be either positive (+) or negative (−). Positive and negative charges attract each other, while two positive or two negative charges repel each other. As we learn in this section, all objects of matter are made up of charged particles. An object having equal numbers of positively and negatively charged particles carries no net charge and is electrically neutral. If the number of positive charges exceeds the number of negative charges, the object has a net positive charge. If negative charges exceed positive charges, the object has a net negative charge. Sometimes when one substance is rubbed against another, as in combing hair, net electric charges build up on the objects, implying that rubbing separates



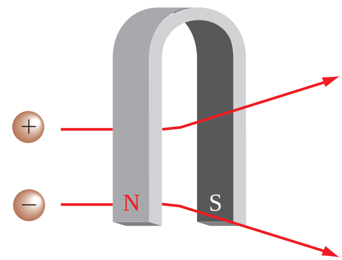
(a) (b)

▲ FIGURE 2-4
Forces between electrically charged objects
(a) Electrostatically charged comb. If you comb your hair on a dry day, a static charge develops on the comb and causes bits of paper to be attracted to the comb. **(b)** Both objects on the left carry a negative electric charge. Objects with like charge repel each other. The objects in the center lack any electric charge and exert no forces on each other. The objects on the right carry opposite charges—one positive and one negative—and attract each other.

◀ We will use *electrostatics* (charge attractions and repulsions) to explain and understand many chemical properties.

some positive and negative charges (Fig. 2-4). Moreover, when a stationary (static) positive charge builds up in one place, a negative charge of equal size appears somewhere else; charge is balanced.

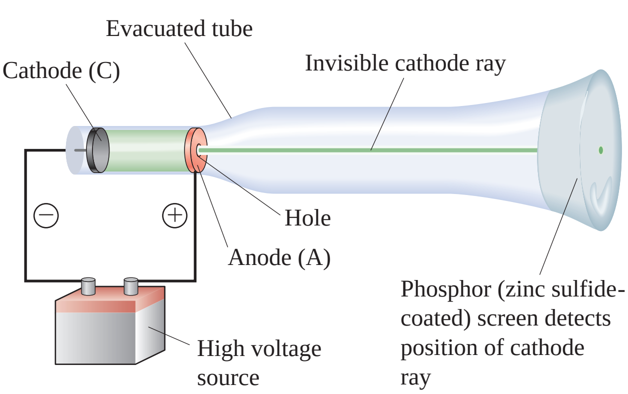
Figure 2-5 shows how charged particles behave when they move through the field of a magnet. They are deflected from their straight-line path into a curved path in a plane perpendicular to the field. Think of the field or region of influence of the magnet as represented by a series of invisible "lines of force" running from the north pole to the south pole of the magnet.



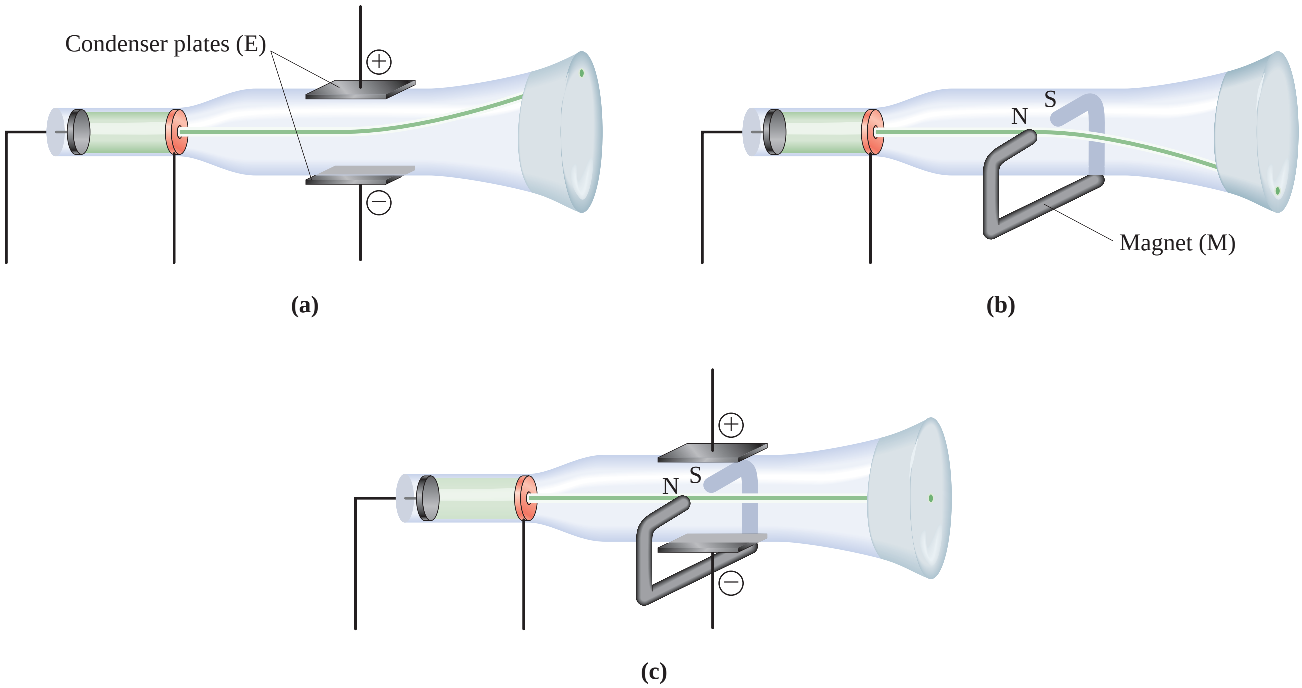
▲ FIGURE 2-5
Effect of a magnetic field on charged particles
When charged particles travel through a magnetic field so that their path is perpendicular to the field, they are deflected by the field. Negatively charged particles are deflected in one direction, and positively charged particles in the opposite direction. Several phenomena described in this section depend on this behavior.

The Discovery of Electrons

CRT, the abbreviation for cathode-ray tube, was once a familiar acronym. Before liquid crystal display (LCD) was available, the CRT was the heart of computer monitors and TV sets. The first cathode-ray tube was made by Michael Faraday (1791–1867) about 150 years ago. When he passed electricity through glass tubes from which most of the air had been evacuated, Faraday discovered **cathode rays**, a type of radiation emitted by the negative terminal, or *cathode*. The radiation crossed the evacuated tube to the positive terminal, or *anode*. Later scientists found that cathode rays travel in straight lines and have properties that are independent of the cathode material (that is, whether it is iron, platinum, and so on). The construction of a CRT is shown in Figure 2-6. The cathode rays produced in the CRT are invisible, and they can be detected only by the light emitted by materials that they strike. These materials, called *phosphors*, are painted on the end of the CRT so that the path of the cathode rays can be revealed. (*Fluorescence* is the term used to describe the emission of light by a phosphor when it is struck by



◀ FIGURE 2-6
A cathode-ray tube
The high-voltage source of electricity creates a negative charge on the electrode at the left (cathode) and a positive charge on the electrode at the right (anode). Cathode rays pass from the cathode (C) to the anode (A), which is perforated to allow the passage of a narrow beam of cathode rays. The rays are visible only through the green fluorescence that they produce on the zinc sulfide–coated screen at the end of the tube. They are invisible in other parts of the tube.



▲ FIGURE 2-7
Cathode rays and their properties
(a) Deflection of cathode rays in an electric field. The beam of cathode rays is deflected as it travels from left to right in the field of the electrically charged condenser plates (E). The deflection corresponds to that expected of negatively charged particles. **(b)** Deflection of cathode rays in a magnetic field. The beam of cathode rays is deflected as it travels from left to right in the field of the magnet (M). The deflection corresponds to that expected of negatively charged particles. **(c)** Determining the mass-to-charge ratio, m/e, for cathode rays. The cathode-ray beam strikes the end screen undeflected if the forces exerted on it by the electric and magnetic fields are counterbalanced. By knowing the strengths of the electric and magnetic fields, together with other data, a value of m/e can be obtained. Precise measurements yield a value of -5.6857×10^{-9} g per coulomb. (Because cathode rays carry a negative charge, the sign of the mass-to-charge ratio is also negative.)

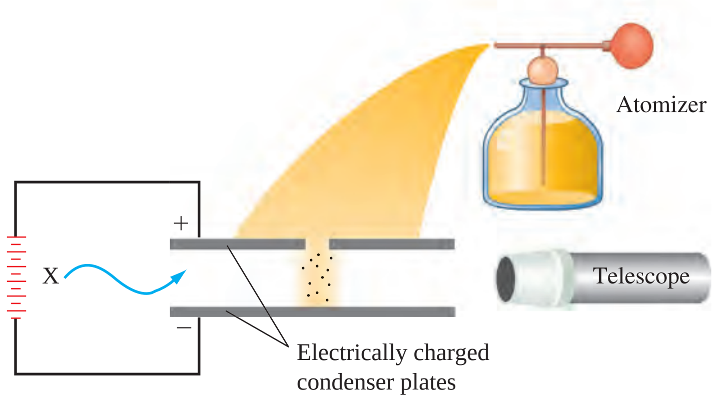
energetic radiation.) Another significant observation about cathode rays is that they are deflected by electric and magnetic fields in the manner expected for negatively charged particles (Fig. 2-7a, b).

In 1897, by the method outlined in Figure 2-7(c), J. J. Thomson (1856–1940) established the ratio of mass (m) to electric charge (e) for cathode rays, that is, m/e. Also, Thomson concluded that cathode rays are negatively charged *fundamental* particles of matter found in all atoms. (The properties of cathode rays are *independent* of the composition of the cathode.) Cathode rays subsequently became known as **electrons**, a term first proposed by George Stoney in 1874.

Robert Millikan (1868–1953) determined the electronic charge e through a series of oil-drop experiments (1906–1914), described in Figure 2-8. The currently accepted value of the electronic charge e, expressed in coulombs to five significant figures, is -1.6022×10^{-19} C. By combining this value with an accurate value of the mass-to-charge ratio for an electron, we find that the mass of an electron is 9.1094×10^{-28} g.

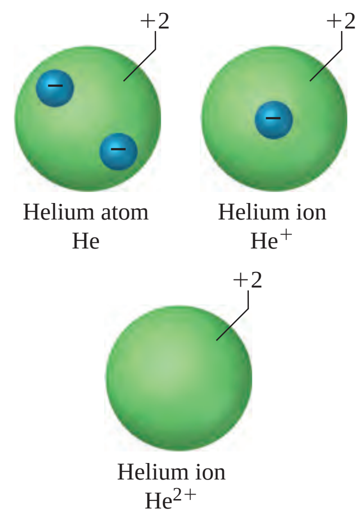
▶ The coulomb (C) is the SI unit of electric charge (see also Appendix B).

Once the electron was seen to be a fundamental particle of matter found in all atoms, atomic physicists began to speculate on how these particles were incorporated into atoms. The commonly accepted model was that proposed by J. J. Thomson. Thomson thought that the positive charge necessary to counterbalance the negative charges of electrons in a neutral atom was in the form of



◀ FIGURE 2-8
Millikan's oil-drop experiment
Ions (charged atoms or molecules) are produced by energetic radiation, such as X-rays (X). Some of these ions become attached to oil droplets, giving them a net charge. The fall of a droplet in the electric field between the condenser plates is speeded up or slowed down, depending on the magnitude and sign of the charge on the droplet. By analyzing data from a large number of droplets, Millikan concluded that the magnitude of the charge, q, on a droplet is an *integral* multiple of the electric charge, e. That is, $q = ne$ (where $n = 1, 2, 3, \ldots$).

a nebulous cloud. Electrons, he suggested, floated in a diffuse cloud of positive charge (rather like a lump of gelatin with electron "fruit" embedded in it). This model became known as the plum-pudding model because of its similarity to a popular English dessert. The plum-pudding model is illustrated in Figure 2-9 for a neutral atom and for atomic species, called *ions*, which carry a net charge.



▲ FIGURE 2-9
The plum-pudding atomic model
According to this model, a helium atom would have a +2 cloud of positive charge and two electrons (−2). If a helium atom loses one electron, it becomes charged and is called an *ion*. This ion, referred to as He^{+}, has a net charge of 1+. If the helium atom loses both electrons, the He^{2+} ion forms.

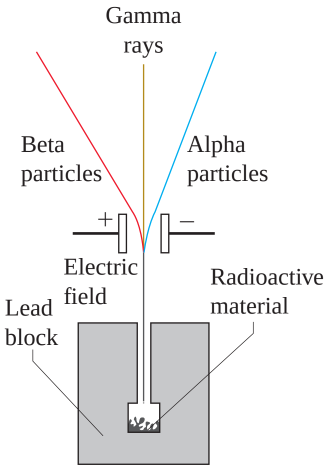
X-Rays and Radioactivity

Cathode-ray research had many important spin-offs. In particular, two natural phenomena of immense theoretical and practical significance were discovered in the course of other investigations.

In 1895, Wilhelm Roentgen (1845–1923) noticed that when cathode-ray tubes were operating, certain materials *outside* the tubes glowed or fluoresced. He showed that this fluorescence was caused by radiation emitted by the cathode-ray tubes. Because of the unknown nature of this radiation, Roentgen coined the term *X-ray*. We now recognize the X-ray as a form of high-energy electromagnetic radiation, which is discussed in Chapter 8.

Antoine Henri Becquerel (1852–1908) associated X-rays with fluorescence and wondered if naturally fluorescent materials produce X-rays. To test this idea, he wrapped a photographic plate with black paper, placed a coin on the paper, covered the coin with a uranium-containing fluorescent material, and exposed the entire assembly to sunlight. When he developed the film, a clear image of the coin could be seen. The fluorescent material had emitted radiation (presumably X-rays) that penetrated the paper and exposed the film. On one occasion, because the sky was overcast, Becquerel placed the experimental assembly inside a desk drawer for a few days while waiting for the weather to clear. On resuming the experiment, Becquerel decided to replace the original photographic film, expecting that it may have become slightly exposed. He developed the original film and found that instead of the expected feeble image, there was a very sharp one. The film had become strongly exposed because the uranium-containing material had emitted radiation continuously, even when it was not fluorescing. Becquerel had discovered **radioactivity**.

Ernest Rutherford (1871–1937) identified two types of radiation from radioactive materials, alpha (α) and beta (β). **Alpha particles** carry two fundamental units of positive charge and have essentially the same mass as helium atoms. In fact, alpha particles are identical to He^{2+} ions. **Beta particles** are negatively charged particles produced by changes occurring within the nuclei of radioactive atoms and have the same properties as electrons. A third form of radiation, which is not affected by electric or magnetic fields, was discovered in 1900 by Paul Villard. This radiation, called **gamma rays** (γ),



▲ FIGURE 2-10
Three types of radiation from radioactive materials
The radioactive material is enclosed in a lead block. All the radiation except that passing through the narrow opening is absorbed by the lead. When the escaping radiation is passed through an electric field, it splits into three beams. One beam is undeflected—these are gamma (γ) rays. A second beam is attracted to the negatively charged plate. These are the positively charged alpha (α) particles. The third beam, of negatively charged beta (β) particles, is deflected toward the positive plate.

is not made up of particles; it is electromagnetic radiation of extremely high penetrating power. These three forms of radioactivity are illustrated in Figure 2-10.

By the early 1900s, additional radioactive elements were discovered, principally by Marie and Pierre Curie. Rutherford and Frederick Soddy made another profound finding: The chemical properties of a radioactive element *change* as it undergoes radioactive decay. This observation suggests that radioactivity involves fundamental changes at the *subatomic* level—in radioactive decay, one element is changed into another, a process known as *transmutation*.

2-3 The Nuclear Atom

In 1909, Rutherford, with his assistant Hans Geiger, began a line of research using α particles as probes to study the inner structure of atoms. Based on Thomson's plum-pudding model, Rutherford expected that most particles in a beam of α particles would pass through thin sections of matter largely undeflected, but that some α particles would be slightly scattered or deflected as they encountered electrons. By studying these scattering patterns, he hoped to deduce something about the distribution of electrons in atoms.

The apparatus used for these studies is pictured in Figure 2-11. Alpha particles were detected by the flashes of light they produced when they struck a zinc sulfide screen mounted on the end of a telescope. When Geiger and Ernst Marsden, a student, bombarded very thin foils of gold with α particles, they observed the following:

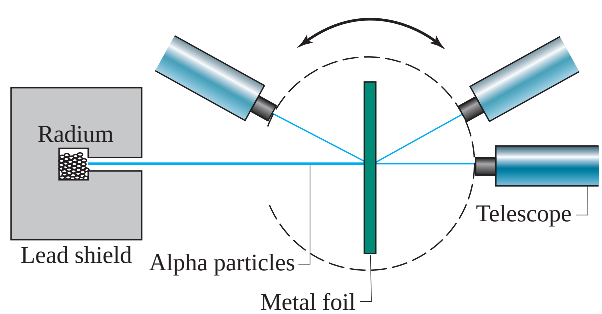
- The majority of α particles penetrated the foil undeflected.
- Some α particles experienced slight deflections.
- A few (about 1 in every 20,000) suffered rather serious deflections as they penetrated the foil.
- A similar number did not pass through the foil at all, but bounced back in the direction from which they had come.

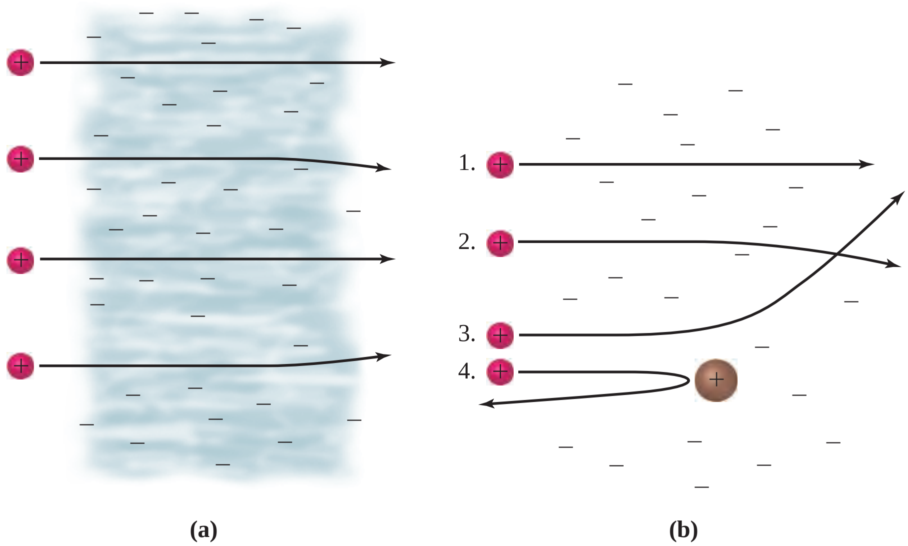
▶ Perhaps because he found it tedious to sit in the dark and count spots of light on a zinc sulfide screen, Geiger was motivated to develop an automatic radiation detector. The result was the well-known Geiger counter.

The large-angle scattering greatly puzzled Rutherford. As he commented some years later, this observation was "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." By 1911, though, Rutherford had an explanation. He based his explanation on a model of the atom known as the *nuclear atom* and having these features:

1. Most of the mass and all of the positive charge of an atom are centered in a very small region called the *nucleus*. The remainder of the atom is mostly *empty space*.
2. The magnitude of the positive charge is different for different atoms and is approximately one-half the atomic weight of the element.
3. There are as many electrons outside the nucleus as there are units of positive charge on the nucleus. The atom as a whole is electrically neutral.

▶ FIGURE 2-11
The scattering of α particles by metal foil
The telescope travels in a circular track around an evacuated chamber containing the metal foil. Most α particles pass through the metal foil undeflected, but some are deflected through large angles.





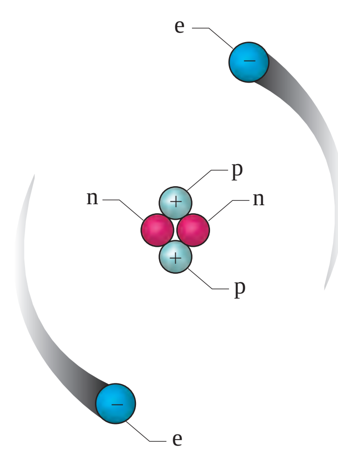
▲ FIGURE 2-12
Explaining the results of α-particle scattering experiments
(a) Rutherford's expectation was that small, positively charged α particles should pass through the nebulous, positively charged cloud of the Thomson plum-pudding model largely undeflected. Some would be slightly deflected by passing near electrons (present to neutralize the positive charge of the cloud). **(b)** Rutherford's explanation was based on a nuclear atom. With an atomic model having a small, dense, positively charged nucleus and extranuclear electrons, we would expect the four different types of paths actually observed:

1. undeflected straight-line paths exhibited by most of the α particles
2. slight deflections of α particles passing close to electrons
3. severe deflections of α particles passing close to a nucleus
4. reflections from the foil of α particles approaching a nucleus head-on

Rutherford's initial expectation and his explanation of the α-particle experiments are described in Figure 2-12.

Discovery of Protons and Neutrons

Rutherford's nuclear atom suggested the existence of positively charged fundamental particles of matter in the nuclei of atoms. Rutherford himself discovered these particles, called **protons**, in 1919 in studies involving the scattering of α particles by nitrogen atoms in air. The protons were freed as a result of collisions between α particles and the nuclei of nitrogen atoms. At about this same time, Rutherford predicted the existence in the nucleus of electrically neutral fundamental particles. In 1932, James Chadwick showed that a newly discovered penetrating radiation consisted of beams of *neutral* particles. These particles, called **neutrons**, originated from the nuclei of atoms. Thus, it has been only for about the past 75 years that we have had the atomic model suggested by Figure 2-13.



▲ FIGURE 2-13
The nuclear atom—illustrated by the helium atom
In this drawing, electrons are shown much closer to the nucleus than is the case. The actual situation is more like this: If the entire atom were represented by a room, 5 m × 5 m × 5 m, the nucleus would occupy only about as much space as the period at the end of this sentence.

2-3 CONCEPT ASSESSMENT

In light of information presented to this point in the text, explain which of the three assumptions of Dalton's atomic theory (page 37) can still be considered correct and which cannot.

TABLE 2.1 Properties of Three Fundamental Particles

	Electric Charge		Mass	
	SI (C)	Atomic	SI (g)	Atomic (u)[a]
Proton	$+1.6022 \times 10^{-19}$	$+1$	1.6726×10^{-24}	1.0073
Neutron	0	0	1.6749×10^{-24}	1.0087
Electron	-1.6022×10^{-19}	-1	9.1094×10^{-28}	0.00054858

[a]u is the SI symbol for atomic mass unit (abbreviated as amu).

▶ The masses of the proton and neutron are different in the fourth significant figure. The charges of the proton and electron, however, are believed to be exactly equal in magnitude (but opposite in sign).

Properties of Protons, Neutrons, and Electrons

The number of protons in a given atom is called the **atomic number**, or the **proton number, *Z***. The number of electrons in the atom is also equal to Z because the atom is electrically neutral. The total number of protons and neutrons in an atom is called the **mass number, *A***. The number of neutrons, the **neutron number**, is $A - Z$. An electron carries an atomic unit of negative charge, a proton carries an atomic unit of positive charge, and a neutron is electrically neutral. Table 2.1 presents the charges and masses of protons, neutrons, and electrons in two ways.

The **atomic mass unit** (described more fully on page 46) is defined as exactly 1/12 of the mass of the atom known as carbon-12 (read as carbon twelve). An atomic mass unit is abbreviated as amu and denoted by the symbol u. As we see from Table 2.1, the proton and neutron masses are just slightly greater than 1 u. By comparison, the mass of an electron is only about 1/2000th the mass of the proton or neutron.

The three subatomic particles considered in this section are the only ones involved in the phenomena of interest to us in this text. You should be aware, however, that a study of matter at its most fundamental level must consider many additional subatomic particles. The electron is believed to be a truly fundamental particle. However, modern particle physics now considers the neutron and proton to be composed of other, more-fundamental particles. This is the subject of elementary particle physics.

2-4 Chemical Elements

Now that we have acquired some fundamental ideas about atomic structure, we can more thoroughly discuss the concept of chemical elements.

All atoms of a particular element have the same atomic number, Z, and, conversely, all atoms with the same number of protons are atoms of the same element. The elements shown on the inside front cover have atomic numbers from $Z = 1$ to $Z = 111$. Each element has a name and a distinctive symbol. **Chemical symbols** are one- or two-letter abbreviations of the name (usually the English name). The first (but never the second) letter of the symbol is capitalized; for example: carbon, C; oxygen, O; neon, Ne; and silicon, Si. Some elements known since ancient times have symbols based on their Latin names, such as Fe for iron (*ferrum*) and Pb for lead (*plumbum*). The element sodium has the symbol Na, based on the Latin *natrium* for sodium carbonate. Potassium has the symbol K, based on the Latin *kalium* for potassium carbonate. The symbol for tungsten, W, is based on the German *wolfram*.

▶ Other atomic symbols not based on English names include Cu, Ag, Sn, Sb, Au, and Hg.

Elements beyond uranium ($Z = 92$) do not occur naturally and must be synthesized in particle accelerators (described in Chapter 25). Elements of the very highest atomic numbers have been produced only on a limited number of occasions, a few atoms at a time. Inevitably, controversies have arisen about which research team discovered a new element and, in fact, whether a discovery was made at all. However, international agreement has been reached on the first 111 elements, each of which now has an official name and symbol.

Isotopes

To represent the composition of any particular atom, we need to specify its number of protons (p), neutrons (n), and electrons (e). We can do this with the symbolism

$$\begin{array}{r} \text{number p} + \text{number n} \rightarrow \\ \text{number p} \rightarrow \end{array} {}^{A}_{Z}\text{E} \longleftarrow \text{symbol of element} \tag{2.1}$$

This symbolism indicates that the atom is element E and that it has atomic number Z and mass number A. For example, an atom of aluminum represented as ${}^{27}_{13}\text{Al}$ has 13 protons and 14 neutrons in its nucleus and 13 electrons outside the nucleus. (Recall that an atom has the same number of electrons as protons.)

Contrary to what Dalton thought, we now know that atoms of an element do not necessarily all have the same mass. In 1912, J. J. Thomson measured the mass-to-charge ratios of positive ions formed from neon atoms. From these ratios he deduced that about 91% of the atoms had one mass and that the remaining atoms were about 10% heavier. All neon atoms have 10 protons in their nuclei, and most have 10 neutrons as well. A very few neon atoms, however, have 11 neutrons and some have 12. We can represent these three different types of neon atoms as

◀ Because neon is the only element with $Z = 10$, the symbols ${}^{20}\text{Ne}$, ${}^{21}\text{Ne}$, and ${}^{22}\text{Ne}$ convey the same meaning as ${}^{20}_{10}\text{Ne}$, ${}^{21}_{10}\text{Ne}$, and ${}^{22}_{10}\text{Ne}$.

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

◀ Odd-numbered elements tend to have fewer isotopes than do even-numbered elements. Section 25-7 will explain why.

Atoms that have the *same* atomic number (Z) but *different* mass numbers (A) are called **isotopes**. Of all Ne atoms on Earth, 90.51% are ${}^{20}_{10}\text{Ne}$. The percentages of ${}^{21}_{10}\text{Ne}$ and ${}^{22}_{10}\text{Ne}$ are 0.27% and 9.22%, respectively. These percentages—90.51%, 0.27%, 9.22%—are the **percent natural abundances** of the three neon isotopes. Sometimes the mass numbers of isotopes are incorporated into the names of elements, such as neon-20 (neon twenty). Percent natural abundances are always based on *numbers*, not masses. Thus, 9051 of every 10,000 neon atoms are neon-20 atoms. Some elements, as they exist in nature, consist of just a single type of atom and therefore do not have naturally occurring isotopes.* Aluminum, for example, consists only of aluminum-27 atoms.

◀ Usually all the isotopes of an element share the same name and atomic symbol. The exception is hydrogen. Isotope ${}^{2}_{1}\text{H}$ is called deuterium (symbol D), and ${}^{3}_{1}\text{H}$ is tritium (T).

Ions

When atoms lose or gain electrons, for example, in the course of a chemical reaction, the species formed are called **ions** and carry net charges. Because an electron is negatively charged, adding electrons to an electrically neutral atom produces a negatively charged ion. Removing electrons results in a positively charged ion. The number of protons does not change when an atom becomes an ion. For example, ${}^{20}\text{Ne}^{+}$ and ${}^{22}\text{Ne}^{2+}$ are ions. The first one has 10 protons, 10 neutrons, and *9* electrons. The second one also has 10 protons, but 12 neutrons and *8* electrons. The charge on an ion is equal to the number of protons *minus* the number of electrons. That is

$$\begin{array}{r} \text{number p} + \text{number n} \rightarrow \\ \text{number p} \rightarrow \end{array} {}^{A}_{Z}\text{E}^{\#\pm} \longleftarrow \text{number p} - \text{number e} \tag{2.2}$$

◀ In this expression, #± indicates that the charge is written with the number *before* the + or − sign. However, when the charge is 1+ or 1−, the number 1 is not included.

Another example is the ${}^{16}\text{O}^{2-}$ ion. In this ion, there are 8 protons (atomic number 8), 8 neutrons (mass number − atomic number), and *10* electrons $(8 - 10 = -2)$.

*__Nuclide__ is the general term used to describe an atom with a particular atomic number and mass number. Although there are several elements with only one naturally occurring nuclide, it is possible to produce additional nuclides of these elements—isotopes—by artificial means (Section 25-3). The artificial isotopes are radioactive, however. In all, the number of synthetic isotopes exceeds the number of naturally occurring ones by several fold.

EXAMPLE 2-3 Relating the Numbers of Protons, Neutrons, and Electrons in Atoms and Ions

Through an appropriate symbol, indicate the number of protons, neutrons, and electrons in **(a)** an atom of barium-135 and **(b)** the double negatively charged ion of selenium-80.

Analyze

Given the name of an element, we can find the symbol and the atomic number, Z, for that element from a list of elements and a periodic table. To determine the number of protons, neutrons, and electrons, we make use of the following relationships:

$$Z = \text{number p} \qquad A = \text{number p} + \text{number n} \qquad \text{charge} = \text{number p} - \text{number e}$$

The relationships above are summarized in (2.2).

Solve

(a) We are given the name (barium) and the mass number of the atom (135). From a list of the elements and a periodic table we obtain the symbol (Ba) and the atomic number ($Z = 56$), leading to the symbolic representation

$$^{135}_{56}\text{Ba}$$

From this symbol one can deduce that the neutral atom has 56 protons; a neutron number of $A - Z = 135 - 56 = 79$ neutrons; and a number of electrons equal to Z, that is, 56 electrons.

(b) We are given the name (selenium) and the mass number of the ion (80). From a list of the elements and a periodic table we obtain the symbol (Se) and the atomic number (34). Together with the fact that the ion carries a charge of 2−, we have the data required to write the symbol

$$^{80}_{34}\text{Se}^{2-}$$

From this symbol, we can deduce that the ion has 34 protons; a neutron number of $A - Z = 80 - 34 = 46$ neutrons; and 36 electrons, leading to a net charge of $+34 - 36 = -2$.

Assess

When writing the symbol for a particular atom or ion, we often omit the atomic number. For example, for $^{135}_{56}\text{Ba}$ and $^{80}_{34}\text{Sr}^{2-}$, we often use the simpler representations ^{135}Ba and $^{80}\text{Sr}^{2-}$.

PRACTICE EXAMPLE A: Use the notation $^{A}_{Z}E$ to represent the isotope of silver having a neutron number of 62.

PRACTICE EXAMPLE B: Use the notation $^{A}_{Z}E$ to represent a tin ion having the same number of electrons as an atom of the isotope cadmium-112. Explain why there can be more than one answer.

2-4 CONCEPT ASSESSMENT

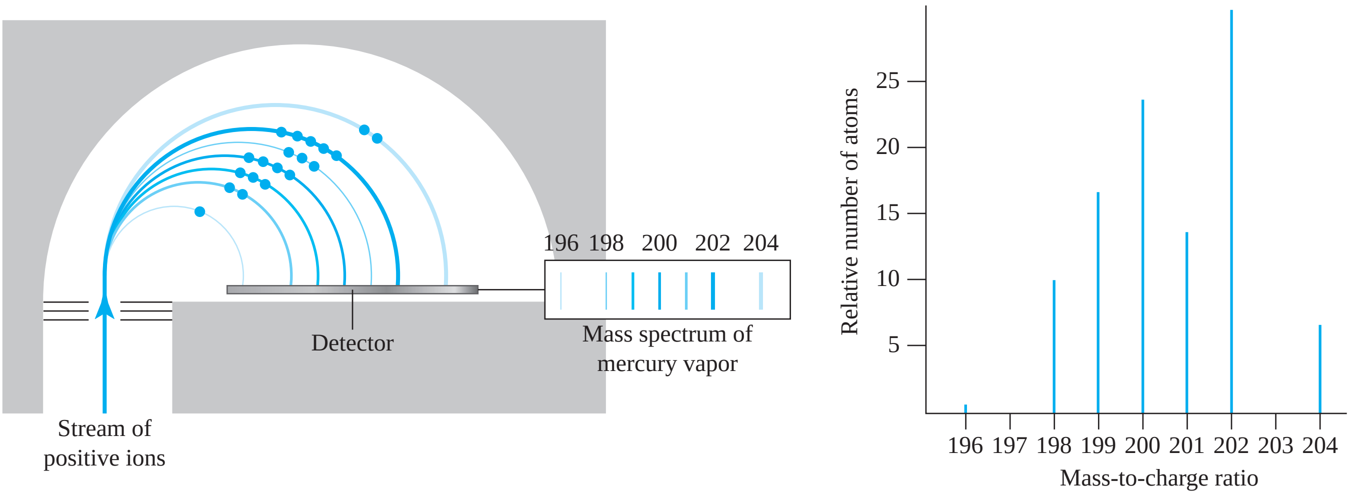
What is the single exception to the statement that all atoms comprise protons, neutrons, and electrons?

▶ Ordinarily we expect like-charged objects (such as protons) to repel each other. The forces holding protons and neutrons together in the nucleus are very much stronger than ordinary electrical forces (Section 25-6).

▶ This definition also establishes that one atomic mass unit (1 u) is *exactly* 1/12 the mass of a carbon-12 atom.

Isotopic Masses

We cannot determine the mass of an individual atom just by adding up the masses of its fundamental particles. When protons and neutrons combine to form a nucleus, a very small portion of their original mass is converted to energy and released. However, we cannot predict exactly how much this so-called nuclear binding energy will be. Determining the masses of individual atoms, then, is something that must be done by experiment, in the following way. By international agreement, one type of atom has been chosen and assigned a specific mass. This standard is an atom of the isotope carbon-12, which is assigned a mass of exactly 12 atomic mass units, that is, *12 u*. Next, the masses of other atoms relative to carbon-12 are determined with a **mass spectrometer**. In this device, a beam of gaseous ions passing through electric and magnetic fields separates into components of differing masses. The



▲ FIGURE 2-14
A mass spectrometer and mass spectrum
In this mass spectrometer, a gaseous sample is ionized by bombardment with electrons in the lower part of the apparatus (not shown). The positive ions thus formed are subjected to an electrical force by the electrically charged velocity selector plates and a magnetic force by a magnetic field perpendicular to the page. Only ions with a particular velocity pass through and are deflected into circular paths by the magnetic field. Ions with different masses strike the detector (here a photographic plate) in different regions. The more ions of a given type, the greater the response of the detector (intensity of line on the photographic plate). In the mass spectrum shown for mercury, the response of the ion detector (intensity of lines on photographic plate) has been converted to a scale of relative numbers of atoms. The percent natural abundances of the mercury isotopes are ^{196}Hg, 0.146%; ^{198}Hg, 10.02%; ^{199}Hg, 16.84%; ^{200}Hg, 23.13%; ^{201}Hg, 13.22%; ^{202}Hg, 29.80%; and ^{204}Hg, 6.85%.

◀ The primary standard for atomic masses has evolved over time. For example, Dalton originally assigned H a mass of 1 u. Later, chemists took naturally occurring oxygen at 16 u to be the definition of the atomic-weight scale. Concurrently, physicists defined the oxygen-16 isotope as 16 u. This resulted in conflicting values. In 1971 the adoption of carbon-12 as the universal standard resolved this disparity.

separated ions are focused on a measuring instrument, which records their presence and amounts. Figure 2-14 illustrates mass spectrometry and a typical mass spectrum.

Although mass numbers are whole numbers, the actual masses of individual atoms (in atomic mass units, u) are never whole numbers, except for carbon-12. However, they are very close in value to the corresponding mass numbers, as we can see for the isotope oxygen-16. From mass spectral data the ratio of the mass of ^{16}O to ^{12}C is found to be 1.33291. Thus, the mass of the oxygen-16 atom is

$$1.33291 \times 12\ \text{u} = 15.9949\ \text{u}$$

which is very nearly equal to the mass number of 16.

EXAMPLE 2-4 Establishing Isotopic Masses by Mass Spectrometry

With mass spectral data, the mass of an oxygen-16 atom is found to be 1.06632 times that of a nitrogen-15 atom. Given that ^{16}O has a mass of 15.9949 u (see above), what is the mass of a nitrogen-15 atom, in u?

Analyze

Given the ratio (mass of ^{16}O)/(mass of ^{15}N) = 1.06632 and the mass of ^{16}O, 15.9949 u, we solve for the mass of ^{15}N.

Solve

We know that

$$\frac{\text{mass of } ^{16}\text{O}}{\text{mass of } ^{15}\text{N}} = 1.06632$$

(continued)

We solve the expression above for the mass of ^{15}N and then substitute 15.9949 u for the mass of ^{16}O. We obtain the result

$$\text{mass of } ^{15}\text{N} = \frac{\text{mass of } ^{16}\text{O}}{1.06632} = \frac{15.9949 \text{ u}}{1.06632} = 15.0001 \text{ u}$$

Assess

The mass of ^{15}N is very nearly 15, as we should expect. If we had mistakenly multiplied instead of dividing by the ratio 1.06632, the result would have been slightly larger than 16 and clearly incorrect.

PRACTICE EXAMPLE A: What is the ratio of masses for ^{202}Hg/^{12}C, if the isotopic mass for ^{202}Hg is 201.97062 u?

PRACTICE EXAMPLE B: An isotope with atomic number 64 and mass number 158 is found to have a mass ratio relative to that of carbon-12 of 13.16034. What is the isotope, what is its atomic mass in amu, and what is its mass relative to oxygen-16?

2-5 Atomic Mass

▶ Carbon-14, used for radiocarbon dating, is formed in the upper atmosphere. The amount of carbon-14 on Earth is too small to affect the atomic mass of carbon.

In a table of atomic masses, the value listed for carbon is 12.011, yet the atomic mass standard is *exactly* 12. Why the difference? The atomic mass standard is based on a sample of carbon containing only atoms of carbon-12, whereas naturally occurring carbon contains some carbon-13 atoms as well. The existence of these two isotopes causes the observed atomic mass to be greater than 12. The **atomic mass (weight)*** of an element is the average of the isotopic masses, *weighted* according to the naturally occurring abundances of the isotopes of the element. In a weighted average, we must assign greater importance—give greater weight—to the quantity that occurs more frequently. Since carbon-12 atoms are much more abundant than carbon-13, the weighted average must lie much closer to 12 than to 13. This is the result that we get by applying the following general equation, where the right-hand side of the equation includes one term for each naturally occurring isotope.

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that the fractional abundance is the percent abundance divided by 100%. Thus, a 98.892% abundance is a 0.98892 fractional abundance.

$$\begin{array}{l}\text{at. mass}\\ \text{of an}\\ \text{element}\end{array} = \left(\begin{array}{l}\text{fractional}\\ \text{abundance of}\\ \text{isotope 1}\end{array} \times \begin{array}{l}\text{mass of}\\ \text{isotope 1}\end{array}\right) + \left(\begin{array}{l}\text{fractional}\\ \text{abundance of}\\ \text{isotope 2}\end{array} \times \begin{array}{l}\text{mass of}\\ \text{isotope 2}\end{array}\right) + \ldots \quad \textbf{(2.3)}$$

The mass spectrum of carbon shows that 98.892% of carbon atoms are carbon-12 with a mass of *exactly* 12 u and 1.108% are carbon-13 with a mass of 13.00335 u.

98,992 atoms of ^{12}C (12 u exactly)

+

1,108 atoms of ^{13}C (13.00335 u)

$\downarrow$

100,000 atoms of naturally occurring C (12.011 u)

In the following setup, we calculate the contribution of each isotope to the weighted average separately and then add these contributions together.

*Since Dalton's time, atomic masses have been called atomic weights. They still are by most chemists, yet what we are describing here is mass, not weight. Old habits die hard.

$$\begin{aligned}
\begin{array}{r}\text{contribution to}\\ \text{at. mass by } ^{12}\text{C}\end{array} &= \begin{array}{c}\text{fraction of carbon}\\ \text{atoms that are } ^{12}\text{C}\end{array} && \times \text{ mass } ^{12}\text{C atom} \\
&= 0.98892 && \times\ 12\ \text{u} && = 11.867\ \text{u} \\
\begin{array}{r}\text{contribution to}\\ \text{at. mass by } ^{13}\text{C}\end{array} &= \begin{array}{c}\text{fraction of carbon}\\ \text{atoms that are } ^{13}\text{C}\end{array} && \times \text{ mass } ^{13}\text{C atom} \\
&= 0.01108 && \times\ 13.00335\ \text{u} && = 0.1441\ \text{u} \\
\begin{array}{r}\text{at. mass of naturally}\\ \text{occurring carbon}\end{array} &= (\text{contribution by } ^{12}\text{C}) && + (\text{contribution by } ^{13}\text{C}) \\
&= 11.867\ \text{u} && + 0.1441\ \text{u} \\
&= 12.011\ \text{u}
\end{aligned}$$

To determine the atomic mass of an element having three naturally occurring isotopes, such as potassium, we would have to include three contributions in the weighted average, and so on.

The percent natural abundances of most of the elements remain very nearly constant from one sample of matter to another. For example, the proportions of ^{12}C and ^{13}C atoms are the same in samples of pure carbon (diamond), carbon dioxide gas, and a mineral form of calcium carbonate (calcite). We can treat all natural carbon-containing materials as if there were a single *hypothetical* type of carbon atom with a mass of 12.011 u. This means that once weighted-average atomic masses have been determined and tabulated, we can simply use these values in calculations requiring atomic masses.

Sometimes a qualitative understanding of the relationship between isotopic masses, percent natural abundances, and weighted-average atomic mass is all that we need, and no calculation is necessary, as illustrated in Example 2-5. Example 2-6 and the accompanying Practice Examples provide additional applications of equation (2.3).

EXAMPLE 2-5 Understanding the Meaning of a Weighted-Average Atomic Mass

The two naturally occurring isotopes of lithium, lithium-6 and lithium-7, have masses of 6.01513 u and 7.01601 u, respectively. Which of these two occurs in greater abundance?

Analyze

Look up the atomic mass of Li and compare it with the masses of ^{6}Li and ^{7}Li. If the atomic mass of Li is closer to that of ^{6}Li, then ^{6}Li is the more abundant isotope. If the atomic mass of Li is closer to that of ^{7}Li, then ^{7}Li is the more abundant isotope.

Solve

From a table of atomic masses (inside the front cover), we see that the atomic mass of lithium is 6.941 u. Because this value—a weighted-average atomic mass—is much closer to 7.01601u than to 6.01513 u, lithium-7 must be the more abundant isotope.

Assess

With the assumption that there are only two isotopes of lithium, ^{6}Li and ^{7}Li., we can actually calculate the fractional abundances. If f represents the fractional abundance of ^{6}Li, then the average atomic mass is $f \times 6.01513\ \text{u} + (1 - f) \times 7.01601\ \text{u} = 6.941$. Solving for f, we obtain $f = 0.07494$. Thus, the percent abundances of ^{6}Li and ^{7}Li in naturally occurring lithium are approximately 7.49% and 92.51%, respectively.

PRACTICE EXAMPLE A: The two naturally occurring isotopes of boron, boron-10 and boron-11, have masses of 10.012937 u and 11.009305 u, respectively. Which of these two occurs in greater abundance?

PRACTICE EXAMPLE B: Indium has two naturally occurring isotopes and a weighted atomic mass of 114.82 u. One of the isotopes has a mass of 112.9043 u. Which of the following must be the second isotope: ^{111}In, ^{112}In, ^{114}In, or ^{115}In? Which of the two naturally occurring isotopes must be the more abundant?

EXAMPLE 2-6 Relating the Masses and Natural Abundances of Isotopes to the Atomic Mass of an Element

Bromine has two naturally occurring isotopes. One of them, bromine-79, has a mass of 78.9183 u and an abundance of 50.69%. What must be the mass and percent natural abundance of the other, bromine-81?

Analyze

Although the atomic mass of Br is not given explicitly, it is a known quantity. From the inside front cover, we find that the atomic mass of Br is 79.904 u. We need to apply two key concepts: (1) the atomic mass of Br is a weighted average of the masses of ^{79}Br and ^{81}Br, and (2) the percent natural abundances of ^{79}Br and ^{81}Br must add up to 100%.

Solve

The atomic mass of Br is a weighted average of the masses of ^{79}Br and ^{81}Br:

$$\text{atomic mass} = \begin{pmatrix} \text{fraction of atoms} \\ \text{that are } ^{79}\text{Br} \times \\ \text{mass of } ^{79}\text{Br} \end{pmatrix} + \begin{pmatrix} \text{fraction of atoms} \\ \text{that are } ^{81}\text{Br} \times \\ \text{mass of } ^{81}\text{Br} \end{pmatrix}$$

Because the percent natural abundances must total 100%, the percent natural abundance of ^{81}Br is 100% − 50.69% = 49.31%. Substituting 79.904 u for the atomic mass, 78.9183 u for the mass of ^{79}Br, and the fractional abundances of the two isotopes, we obtain

$$\begin{aligned} 79.904\text{ u} &= (0.5069 \times 78.9183\text{ u}) + (0.4931 \times \text{mass of } ^{81}\text{Br}) \\ &= 40.00\text{ u} + (0.4931 \times \text{mass of } ^{81}\text{Br}) \end{aligned}$$

$$\text{mass of } ^{81}\text{Br} = \frac{79.904\text{ u} - 40.00\text{ u}}{0.4931} = 80.92\text{ u}$$

To four significant figures, the natural abundance of the bromine-81 isotope is 49.31% and its mass is 80.92 u.

Assess

We can check the final result by working the problem in reverse and using numbers that are slightly rounded. The atomic mass of Br is 50.69% × 78.92 u + 49.31% × 80.92 u ≈ $\frac{1}{2}$ (79 u + 81 u) = 80 u. The estimated atomic mass (80 u) is close to the actual atomic mass of 79.904 u.

PRACTICE EXAMPLE A: The masses and percent natural abundances of the three naturally occurring isotopes of silicon are ^{28}Si, 27.97693 u, 92.23%; ^{29}Si, 28.97649 u, 4.67%; ^{30}Si, 29.97376 u, 3.10%. Calculate the weighted-average atomic mass of silicon.

PRACTICE EXAMPLE B: Use data from Example 2-5 to determine the percent natural abundances of lithium-6 and lithium-7.

The table of atomic masses (inside the front cover) shows that some atomic masses are stated much more precisely than others. For example, the atomic mass of F is given as 18.9984 u and that of Kr is given as 83.80 u. In fact, the atomic mass of fluorine is known even more precisely (18.9984032 u); the value of 18.9984 u has been rounded off to six significant figures. Why is the atomic mass of F known so much more precisely than that of Kr? Only one type of fluorine atom occurs naturally: fluorine-19. Determining the atomic mass of fluorine means establishing the mass of this type of atom as precisely as possible. The atomic mass of krypton is known less precisely because krypton has six naturally occurring isotopes. Because the percent distribution of the isotopes of krypton differs very slightly from one sample to another, the weighted-average atomic mass of krypton cannot be stated with high precision.

2-5 CONCEPT ASSESSMENT

The value listed for chromium in the table of atomic masses inside the front cover is 51.9961 u. Should we conclude that naturally occurring chromium atoms are all of the type $^{52}_{24}$Cr? The same table lists a value of 65.39 u for zinc. Should we conclude that zinc occurs as a mixture of isotopes? Explain.

2-6 Introduction to the Periodic Table

Scientists spend a lot of time organizing information into useful patterns. Before they can organize information, however, they must possess it, and it must be correct. Botanists had enough information about plants to organize their field in the eighteenth century. Because of uncertainties in atomic masses and because many elements remained undiscovered, chemists were not able to organize the elements until a century later.

We can distinguish one element from all others by its particular set of observable physical properties. For example, sodium has a low density of 0.971 g/cm^3 and a low melting point of 97.81 °C. No other element has this same combination of density and melting point. Potassium, though, also has a low density (0.862 g/cm^3) and low melting point (63.65 °C), much like sodium. Sodium and potassium further resemble each other in that both are good conductors of heat and electricity, and both react vigorously with water to liberate hydrogen gas. Gold, conversely, has a density (19.32 g/cm^3) and melting point (1064 °C) that are very much higher than those of sodium or potassium, and gold does not react with water or even with ordinary acids. It does resemble sodium and potassium in its ability to conduct heat and electricity, however. Chlorine is very different still from sodium, potassium, and gold. It is a gas under ordinary conditions, which means that the melting point of solid chlorine is far below room temperature (−101 °C). Also, chlorine is a nonconductor of heat and electricity.

Even from these very limited data, we get an inkling of a useful classification scheme of the elements. If the scheme is to group together elements with similar properties, then sodium and potassium should appear in the same group. And if the classification scheme is in some way to distinguish between elements that are good conductors of heat and electricity and those that are not, chlorine should be set apart from sodium, potassium, and gold. The classification system we need is the one shown in Figure 2-15 (and inside the front

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that the periodic table shown in Figure 2-15 is the one currently recommended by IUPAC. Elements with atomic numbers 112 and above have been reported but not fully authenticated. In Figure 2-15, lutetium (Lu) and lawrencium (Lr) are the last members of the lanthanide and actinide series, respectively. A strong argument* has been made for placing Lu and Lr in group 3, meaning the lanthanide series would end with ytterbium (Yb) and the actinide series would end with nobelium (Nb). To date, IUPAC has not endorsed placing Lu and Lr in group 3.

* See W. B. Jensen, *J. Chem. Educ.*, **59**, 634 (1982).

1 1A	2 2A	3 3B	4 4B	5 5B	6 6B	7 7B	8 8B	9 8B	10 8B	11 1B	12 2B	13 3A	14 4A	15 5A	16 6A	17 7A	18 8A
1 H 1.00794																	2 He 4.00260
3 Li 6.941	4 Be 9.01218											5 B 10.811	6 C 12.0107	7 N 14.0067	8 O 15.9994	9 F 18.9984	10 Ne 20.1797
11 Na 22.9898	12 Mg 24.3050											13 Al 26.9815	14 Si 28.0855	15 P 30.9738	16 S 32.065	17 Cl 35.453	18 Ar 39.948
19 K 39.0983	20 Ca 40.078	21 Sc 44.9559	22 Ti 47.867	23 V 50.9415	24 Cr 51.9961	25 Mn 54.9380	26 Fe 55.845	27 Co 58.9332	28 Ni 58.6934	29 Cu 63.546	30 Zn 65.409	31 Ga 69.723	32 Ge 72.64	33 As 74.9216	34 Se 78.96	35 Br 79.904	36 Kr 83.798
37 Rb 85.4678	38 Sr 87.62	39 Y 88.9059	40 Zr 91.224	41 Nb 92.9064	42 Mo 95.94	43 Tc (98)	44 Ru 101.07	45 Rh 102.906	46 Pd 106.42	47 Ag 107.868	48 Cd 112.411	49 In 114.818	50 Sn 118.710	51 Sb 121.760	52 Te 127.60	53 I 126.904	54 Xe 131.293
55 Cs 132.905	56 Ba 137.327	57–71 La–Lu 138.906	72 Hf 178.49	73 Ta 180.948	74 W 183.84	75 Re 186.207	76 Os 190.23	77 Ir 192.217	78 Pt 195.084	79 Au 196.967	80 Hg 200.59	81 Tl 204.383	82 Pb 207.2	83 Bi 208.980	84 Po (209)	85 At (210)	86 Rn (222)
87 Fr (223)	88 Ra (226)	89–103 Ac–Lr 227.028	104 Rf (261)	105 Db (262)	106 Sg (266)	107 Bh (264)	108 Hs (277)	109 Mt (268)	110 Ds (271)	111 Rg (272)							

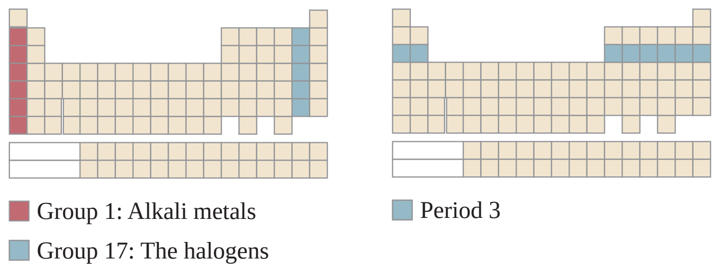
*Lanthanide series	57 La 138.905	58 Ce 140.116	59 Pr 140.908	60 Nd 144.242	61 Pm (145)	62 Sm 150.36	63 Eu 151.964	64 Gd 157.25	65 Tb 158.925	66 Dy 162.500	67 Ho 164.930	68 Er 167.259	69 Tm 168.934	70 Yb 173.04	71 Lu 174.967
†Actinide series	89 Ac (227)	90 Th 232.038	91 Pa 231.036	92 U 238.029	93 Np (237)	94 Pu (244)	95 Am (243)	96 Cm (247)	97 Bk (247)	98 Cf (251)	99 Es (252)	100 Fm (257)	101 Md (258)	102 No (259)	103 Lr (262)

◀ FIGURE 2-15
Periodic table of the elements
Atomic masses are relative to carbon-12. For certain radioactive elements, the numbers listed in parentheses are the mass numbers of the most stable isotopes. Metals are shown in tan, nonmetals in blue, and metalloids in green. The noble gases (also nonmetals) are shown in pink.

cover), known as the **periodic table** of the elements. In Chapter 9, we will describe how the periodic table was formulated, and we will also learn its theoretical basis. For the present, we will consider only a few features of the table.

▶ That elements in one group have similar properties is perhaps the most useful simplifying feature of atomic properties. Significant differences within a group do occur. The manner and reason for such differences is much of what we try to discover in studying chemistry.

Features of the Periodic Table In the periodic table, elements are listed according to increasing atomic number starting at the upper left and arranged in a series of horizontal rows. This arrangement places similar elements in *vertical* **groups**, or **families**. For example, sodium and potassium are found together in a group labeled 1 (called the *alkali metals*). We should expect other members of the group, such as cesium and rubidium, to have properties similar to sodium and potassium. Chlorine is found at the other end of the table in a group labeled 17. Some of the groups are given distinctive names, mostly related to an important property of the elements in the group. For example, the group 17 elements are called the *halogens*, a term derived from Greek, meaning "salt former."



Each element is listed in the periodic table by placing its symbol in the middle of a box in the table. The atomic number (Z) of the element is shown above the symbol, and the weighted-average atomic mass of the element is shown below its symbol. Some periodic tables provide other information, such as density and melting point, but the atomic number and atomic mass are generally sufficient for our needs. Elements with atomic masses in parentheses, such as plutonium, Pu (244), are produced synthetically, and the number shown is the mass number of the most stable isotope.

It is customary also to divide the elements into two broad categories—**metals** and **nonmetals**. In Figure 2-15, colored backgrounds are used to distinguish the metals (tan) from the nonmetals (blue and pink). Except for mercury, a liquid, metals are solids at room temperature. They are generally malleable (capable of being flattened into thin sheets), ductile (capable of being drawn into fine wires), and good conductors of heat and electricity, and have a lustrous or shiny appearance. The properties of nonmetals are generally opposite those of metals; for example, nonmetals are poor conductors of heat and electricity. Several of the nonmetals, such as nitrogen, oxygen, and chlorine, are gases at room temperature. Some, such as silicon and sulfur, are brittle solids. One—bromine—is a liquid.

▶ There is lack of agreement on just which elements to label as metalloids. However, they are generally considered either to lie adjacent to the stair-step line or to be close by.

Two other highlighted categories in Figure 2-15 are a special group of nonmetals known as the **noble gases** (pink), and a small group of elements, often called **metalloids** (green), that have some metallic and some nonmetallic properties.

The *horizontal* rows of the table are called **periods**. (The periods are numbered at the extreme left in the periodic table inside the front cover.) The first period of the table consists of just two elements, hydrogen and helium. This is followed by two periods of eight elements each, lithium through neon and sodium through argon. The fourth and fifth periods contain 18 elements each, ranging from potassium through krypton and from rubidium through xenon. The sixth period is a long one of 32 members. To fit this period in a table that is held to a maximum width of 18 members, 15 members of the period are placed at the bottom of the periodic table. This series of 15 elements start with lanthanum ($Z = 57$), and these elements are called the **lanthanides**. The seventh and final period is incomplete (some members are yet to be discovered), but it is known to be a long one. A 15-member series is also extracted from the seventh period and placed at the bottom of the table. Because the elements in this series start with actinium ($Z = 89$), they are called the **actinides**.

The labeling of the groups of the periodic table has been a matter of some debate among chemists. The 1-18 numbering system used in Figure 2-15 is the one most recently adopted. Group labels previously used in the United States consisted of a letter and a number, closely following the method adopted by Mendeleev, the developer of the periodic table. As seen in Figure 2-15, the A groups 1 and 2 are separated from the remaining A groups (3 to 8) by B groups 1 through 8. The International Union of Pure and Applied Chemistry (IUPAC) recommended the simple 1 to 18 numbering scheme in order to avoid confusion between the American number and letter system and that used in Europe, where some of the A and B designations were switched! Currently, the IUPAC system is officially recommended by the American Chemical Society (ACS) and chemical societies in other nations. Because both numbering systems are in use, we show both in Figure 2-15 and in the periodic table inside the front cover. However, except for an occasional reminder of the earlier system, we will use the IUPAC numbering system in this text.

◀ Mendeleev's arrangement of the elements in the original periodic table was based on observed chemical and physical properties of the elements and their compounds. The arrangement of the elements in the modern periodic table is based on atomic properties–atomic number and electron configuration.

Useful Relationships from the Periodic Table

The periodic table helps chemists describe and predict the properties of chemical compounds and the outcomes of chemical reactions. Throughout this text, we will use it as an aid to understanding chemical concepts. One application of the table worth mentioning here is how it can be used to predict likely charges on simple monatomic ions.

Main-group elements are those in groups 1, 2, and 13 to 18. When main-group metal atoms in groups 1 and 2 form ions, they lose the same number of electrons as the IUPAC group number. Thus, Na atoms (group 1) lose one electron to become Na^+, and Ca atoms (group 2) lose two electrons to

EXAMPLE 2-7 Describing Relationships Based on the Periodic Table

Refer to the periodic table on the inside front cover, and indicate

(a) the element that is in group 14 and the fourth period;
(b) two elements with properties similar to those of molybdenum (Mo);
(c) the ion most likely formed from a strontium atom.

Analyze

For **(a)**, the key concept is that the rows (periods) are numbered 1 through 7, starting from the top of the periodic table, and the groups are numbered 1 through 18, starting from the left side. For **(b)**, the key concept is that elements in the same group have similar properties. For **(c)**, the key concept is that main-group metal atoms in groups 1 and 2 form positive ions with charges of +1 and +2, respectively.

Solve

(a) The elements in the fourth period range from K ($Z = 19$) to Kr ($Z = 36$). Those in group 14 are C, Si, Ge, Sn, and Pb. The only element that is common to both of these groupings is Ge ($Z = 32$).
(b) Molybdenum is in group 6. Two other members of this group that should resemble it are chromium (Cr) and tungsten (W).
(c) Strontium (Sr) is in group 2. It should form the ion Sr^{2+}.

Assess

In Chapter 8, we will examine in greater detail reasons for the arrangement of the periodic table.

PRACTICE EXAMPLE A: Write a symbol for the ion most likely formed by an atom of each of the following: Li, S, Ra, F, I, and Al.

PRACTICE EXAMPLE B: Classify each of the following elements as a main-group or transition element. Also, specify whether they are metals, metalloids, or nonmetals: Na, Re, S, I, Kr, Mg, U, Si, B, Al, As, H.

become Ca^{2+}. Aluminum, in group 13, loses three electrons to form Al^{3+} (here the charge is "group number minus 10"). The few other metals in groups 13 and higher form more than one possible ion, a matter that we deal with in Chapter 9.

When nonmetal atoms form ions, they gain electrons. The number of electrons gained is normally 18 minus the IUPAC group number. Thus, an O atom gains $18 - 16 = 2$ electrons to become O^{2-}, and a Cl atom gains $18 - 17 = 1$ electron to become Cl^{-}. The "18 minus group number" rule suggests that an atom of Ne in group 18 gains no electrons: $18 - 18 = 0$. The very limited tendency of the noble gas atoms to form ions is one of several characteristics of this family of elements.

The elements in groups 3 to 12 are the **transition elements**, and because all of them are metals, they are also called the **transition metals**. Like the main-group metals, the transition metals form positive ions, but the number of electrons lost is not related in any simple way to the group number, mostly because transition metals can form two or more ions of differing charge.

2-7 The Concept of the Mole and the Avogadro Constant

Starting with Dalton, chemists have recognized the importance of relative numbers of atoms, as in the statement that *two* hydrogen atoms and *one* oxygen atom combine to form *one* molecule of water. Yet it is physically impossible to count every atom in a macroscopic sample of matter. Instead, some other measurement must be employed, which requires a relationship between the measured quantity, usually mass, and some known, but uncountable, number of atoms. Consider a practical example of mass substituting for a desired number of items. Suppose you want to nail down new floorboards on the deck of a mountain cabin, and you have calculated how many nails you will need. If you have an idea of how many nails there are in a pound, then you can buy the nails by the pound.

▶ Because the value of Avogadro's number depends, in part, on a measurement, the value has changed slightly over the years. The values recommended since 1986 by the Committee on Data for Science and Technology (CODATA) are listed below.

Year	Avogadro's Number
1986	6.0221367×10^{23}
1998	$6.02214199 \times 10^{23}$
2002	6.0221415×10^{23}
2006	$6.02214179 \times 10^{23}$

The SI quantity that describes an amount of substance by relating it to a number of particles of that substance is called the *mole* (abbreviated *mol*). A **mole** is the amount of a substance that contains the same number of elementary entities as there are atoms in exactly 12 g of pure carbon-12. The "number of elementary entities (atoms, molecules, and so on)" in a mole is the **Avogadro constant, N_A.**

$$N_A = 6.02214179 \times 10^{23}\ \text{mol}^{-1} \qquad \textbf{(2.4)}$$

The Avogadro constant consists of a number, $6.02214179 \times 10^{23}$, known as Avogadro's *number*, and a unit, mol^{-1}. The unit mol^{-1} signifies that the entities being counted are those present in 1 mole.

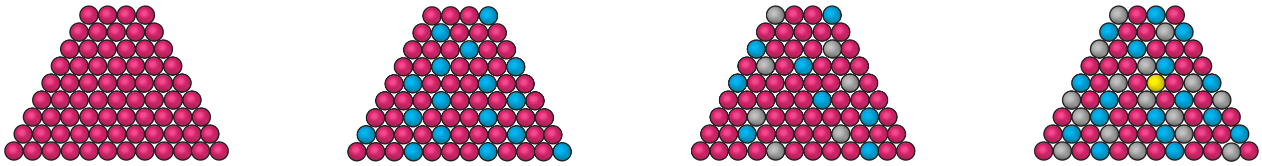
The value of Avogadro's number is based on both a definition and a measurement. A mole of carbon-12 is *defined* to be 12 g. If the mass of one carbon-12 atom is *measured* by using a mass spectrometer (see Figure 2-14), the mass would be about 1.9926×10^{-23} g. The ratio of these two masses provides an estimate of Avogadro's number. In actual fact, accurate determinations of Avogadro's number make use of other measurements, not the measurement of the mass of a single atom of carbon-12.

▶ When rounding Avogadro's number or any other accurately known value, keep one more significant figure than that of the least accurate number in the calculation to avoid rounding errors.

Often the value of N_A is rounded off to $6.022 \times 10^{23}\ \text{mol}^{-1}$, or even to $6.02 \times 10^{23}\ \text{mol}^{-1}$.

If a substance contains atoms of only a single isotope, then

$$1\ \text{mol}\ {}^{12}\text{C} = 6.02214 \times 10^{23}\ {}^{12}\text{C atoms} = 12.0000\ \text{g}$$
$$1\ \text{mol}\ {}^{16}\text{O} = 6.02214 \times 10^{23}\ {}^{16}\text{O atoms} = 15.9949\ \text{g (and so on)}$$



(a) 6.02214×10^{23} F atoms $= 18.9984$ g

(b) 6.02214×10^{23} Cl atoms $= 35.4527$ g

(c) 6.02214×10^{23} Mg atoms $= 24.3050$ g

(d) 6.02214×10^{23} Pb atoms $= 207.2$ g

▲ FIGURE 2-16
Distribution of isotopes in four elements
(a) There is only one type of fluorine atom, ^{19}F (shown in red). (b) In chlorine, 75.77% of the atoms are ^{35}Cl (red) and the remainder are ^{37}Cl (blue). (c) Magnesium has one principal isotope, ^{24}Mg (red), and two minor ones, ^{25}Mg (gray) and ^{26}Mg (blue). (d) Lead has four naturally occurring isotopes: 1.4% ^{204}Pb (yellow), 24.1% ^{206}Pb (blue), 22.1% ^{207}Pb (gray), and 52.4% ^{208}Pb (red).

Most elements are composed of mixtures of two or more isotopes so that the atoms in a sample of the element are not all of the same mass but are present in their naturally occurring proportions. Thus, in one mole of carbon, most of the atoms are carbon-12, but some are carbon-13. In one mole of oxygen, most of the atoms are oxygen-16, but some are oxygen-17 and some are oxygen-18. As a result,

$$1 \text{ mol of C} = 6.02214 \times 10^{23} \text{ C atoms} = 12.011 \text{ g}$$
$$1 \text{ mol of O} = 6.02214 \times 10^{23} \text{ O atoms} = 15.9994 \text{ g, and so on.}$$

◀ We established the atomic mass of ^{16}O relative to that of ^{12}C on page 47.

The Avogadro constant was purposely chosen so that the mass of one mole of carbon-12 atoms—exactly 12 g—would have the same *numeric* value as the mass of a single carbon-12 atom—exactly 12 u. As a result, for all other elements the numeric value of the mass in grams of one mole of atoms and the weighted-average atomic mass in atomic mass units are equal. For example, the weighted average atomic mass of lithium is 6.941 u and the mass of one mole of lithium atoms is 6.941 g. Thus, we can easily establish the mass of one mole of atoms, called the **molar mass, *M***, from a table of atomic masses.* For example, the molar mass of lithium is 6.941 g Li/mol Li. Figure 2-16 attempts to portray the distribution of isotopes of an element, and Figure 2-17 pictures one mole each of four common elements.

◀ The weighted-average atomic mass of carbon was calculated on page 48.

KEEP IN MIND

that molar mass has the unit g/mol.

◀ FIGURE 2-17
One mole of an element
The watch glasses contain one mole of copper atoms (left) and one mole of sulfur atoms (right). The beaker contains one mole of mercury atoms as liquid mercury, and the balloon, of which only a small portion is visible here, contains one mole of helium atoms in the gaseous state.

*Atomic mass (atomic weight) values in tables are often written without units, especially if they are referred to as *relative* atomic masses. This simply means that the values listed are in relation to *exactly* 12 (rather than 12 u) for carbon-12. We will use the atomic mass unit (u) when referring to atomic masses (atomic weights). Most chemists do.

2-6 CONCEPT ASSESSMENT

Dividing the molar mass of gold by the Avogadro constant yields the mass of any individual atom of naturally occurring gold. In contrast, no naturally occurring atom of silver has the mass obtained by dividing the molar mass of silver by the Avogadro constant. How can this be?

Thinking About Avogadro's Number

Avogadro's number (6.02214×10^{23}) is an enormously large number and practically inconceivable in terms of ordinary experience. Suppose we were counting garden peas instead of atoms. If the typical pea had a volume of about 0.1 cm^3, the required pile of peas would cover the United States to a depth of about 6 km (4 mi). Or imagine that grains of wheat could be counted at the rate of 100 per minute. A given individual might be able to count out about 4 billion grains in a lifetime. Even so, if all the people currently on Earth were to spend their lives counting grains of wheat, they could not reach Avogadro's number. In fact, if all the people who ever lived on Earth had spent their lifetimes counting grains of wheat, the total would still be far less than Avogadro's number. (And Avogadro's number of wheat grains is far more wheat than has been produced in human history.) Now consider a much more efficient counting device, a modern personal computer; it is capable of counting at a rate of about 1 billion units per second. The task of counting out Avogadro's number would still take about 20 million years!

▶ You will have an opportunity to calculate a value of N_A at several points in the text, starting with Exercise 113 in Chapter 3.

Avogadro's number is clearly not a useful number for counting ordinary objects. However, when this inconceivably large number is used to count inconceivably small objects, such as atoms and molecules, the result is a quantity of material that is easily within our grasp, essentially a "handful."

2-8 Using the Mole Concept in Calculations

Throughout the text, the mole concept will provide conversion factors for problem-solving situations. With each new situation, we will explore how the mole concept applies. For now, we will deal with the relationship between numbers of atoms and the mole. Consider the statement: 1 mol S = 6.022×10^{23} S atoms = 32.07 g S. This allows us to write the conversion factors

$$\frac{1 \text{ mol S}}{6.022 \times 10^{23} \text{ S atoms}} \quad \text{and} \quad \frac{32.07 \text{ g S}}{1 \text{ mol S}}$$

In calculations requiring the Avogadro constant, students often ask when to multiply and when to divide by N_A. One answer is always to use the constant in a way that gives the proper cancellation of units. Another answer is to think in terms of the expected result. In calculating a number of atoms, we expect the answer to be a very large number and certainly *never* smaller than one. The number of moles of atoms, conversely, is generally a number of more modest size and will often be less than one.

In the following examples, we use atomic masses and the Avogadro constant in calculations to determine the number of atoms present in a given sample. Atomic masses and the Avogadro constant are known rather precisely, and students often wonder how many significant figures to carry in atomic masses or the Avogadro constant when performing calculations. Here is a useful rule of thumb.

▲ FIGURE 2-18
Measurement of 7.65×10^{22} S atoms (0.127 mol S)—Example 2-8 illustrated
The balance was set to zero (tared) when just the weighing boat was present. The sample of sulfur weighs 4.07 g.

To ensure the maximum precision allowable, carry at least one more significant figure in well-known physical constants than in other measured quantities.

For example, in calculating the mass of 0.600 mol of sulfur, we should use the atomic mass of S with *at least* four significant figures. The answer 0.600 mol S × 32.07 g S/mol S = 19.2 g S is a more precise response than 0.600 mol S × 32.1 g S/mol S = 19.3 g S.

EXAMPLE 2-8 Relating Number of Atoms, Amount in Moles, and Mass in Grams

In the sample of sulfur weighing 4.07 g pictured in Figure 2-18, **(a)** how many moles of sulfur are present, and **(b)** what is the total number of sulfur atoms in the sample?

Analyze

For **(a)**, the conversion pathway is g S → mol S. To carry out this conversion, we multiply 4.07 g S by the conversion factor (1 mol S/32.07 g S). The conversion factor is the molar mass inverted. For **(b)**, the conversion pathway is mol S → atoms S. To carry out this conversion, we multiply the quantity in moles from part **(a)** by the conversion factor (6.022×10^{23} atoms S/1 mol S).

Solve

(a) For the conversion g S → mol S, using $(1/M)$ as a conversion factor achieves the proper cancellation of units.

$$? \text{ mol S} = 4.07 \cancel{\text{g S}} \times \frac{1 \text{ mol S}}{32.07 \cancel{\text{g S}}} = 0.127 \text{ mol S}$$

(b) The conversion mol S → atoms S is carried out using the Avogadro constant as a conversion factor.

$$? \text{ atoms S} = 0.127 \cancel{\text{mol S}} \times \frac{6.022 \times 10^{23} \text{ atoms S}}{1 \cancel{\text{mol S}}} = 7.65 \times 10^{22} \text{ atoms S}$$

Assess

By including units in our calculations, we can check that proper cancellation of units occurs. Also, if our only concern is to calculate the number of sulfur atoms in the sample, the calculations carried out in parts **(a)** and **(b)** could be combined into a single calculation, as shown below.

$$? \text{ atoms S} = 4.07 \cancel{\text{g S}} \times \frac{1 \cancel{\text{mol S}}}{32.07 \cancel{\text{g S}}} \times \frac{6.022 \times 10^{23} \text{ atoms S}}{1 \cancel{\text{mol S}}} = 7.65 \times 10^{22} \text{ atoms S}$$

With a single line calculation, we do not have to write down an intermediate result and we avoid round-off errors.

PRACTICE EXAMPLE A: What is the mass of 2.35×10^{24} atoms of Cu?

PRACTICE EXAMPLE B: How many lead-206 atoms are present in a 22.6 g sample of lead metal? [*Hint:* See Figure 2-16.]

Example 2-9 is perhaps the most representative use of the mole concept. Here it is part of a larger problem that requires other unrelated conversion factors as well. One approach is to outline a conversion pathway to get from the given to the desired information.

EXAMPLE 2-9 Combining Several Factors in a Calculation—Molar Mass, the Avogadro Constant, Percent Abundances

Potassium-40 is one of the few naturally occurring radioactive isotopes of elements of low atomic number. Its percent natural abundance among K isotopes is 0.012%. How many ^{40}K atoms are present in 225 mL of whole milk containing 1.65 mg K/mL?

Analyze

Ultimately we need to complete the conversion mL milk → atoms ^{40}K. There is no single conversion factor that allows us to complete this conversion in one step, so we anticipate having to complete several steps or conversions. We are told the milk contains 1.65 mg K/mL = 1.65×10^{-3} g K/mL, and this information can be used to

(continued)

carry out the conversion mL milk $\rightarrow$ g K. We can carry out the conversions g K $\rightarrow$ mol K $\rightarrow$ atoms K by using conversion factors based on the molar mass of K and the Avogadro constant. The final conversion, atoms K $\rightarrow$ atoms ^{40}K, can be carried out by using a conversion factor based on the percent natural abundance of ^{40}K. A complete conversion pathway is shown below:

$$\text{mL milk} \rightarrow \text{mg K} \rightarrow \text{g K} \rightarrow \text{mol K} \rightarrow \text{atoms K} \rightarrow \text{atoms }^{40}\text{K}$$

Solve

The required conversions can be carried out in a stepwise fashion, or they can be combined into a single line calculation. Let's use a stepwise approach. First, we convert from mL milk to g K.

$$?\ \text{g K} = 225\ \text{mL milk} \times \frac{1.65\ \text{mg K}}{1\ \text{mL milk}} \times \frac{1\ \text{g K}}{1000\ \text{mg K}} = 0.371\ \text{g K}$$

Next, we convert from g K to mol K,

$$?\ \text{mol K} = 0.371\ \text{g K} \times \frac{1\ \text{mol K}}{39.10\ \text{g K}} = 9.49 \times 10^{-3}\ \text{mol K}$$

and then we convert from mol K to atoms K.

$$?\ \text{atoms K} = 9.49 \times 10^{-3}\ \text{mol K} \times \frac{6.022 \times 10^{23}\ \text{atoms K}}{1\ \text{mol K}} = 5.71 \times 10^{21}\ \text{atoms K}$$

Finally, we convert from atoms K to atoms ^{40}K.

$$?\ \text{atoms }^{40}\text{K} = 5.71 \times 10^{21}\ \text{atoms K} \times \frac{0.012\ \text{atoms }^{40}\text{K}}{100\ \text{atoms K}} = 6.9 \times 10^{17}\ \text{atoms }^{40}\text{K}$$

Assess

The final answer is rounded to two significant figures because the least precisely known quantity in the calculation, the percent natural abundance of ^{40}K, has two significant figures. It is possible to combine the steps above into a single line calculation.

$$?\ \text{atoms }^{40}\text{K} = 225\ \text{mL milk} \times \frac{1.65\ \text{mg K}}{1\ \text{mL milk}} \times \frac{1\ \text{g K}}{1000\ \text{mg K}} \times \frac{1\ \text{mol K}}{39.10\ \text{g K}}$$

$$\times \frac{6.022 \times 10^{23}\ \text{atoms K}}{1\ \text{mol K}} \times \frac{0.012\ \text{atoms }^{40}\text{K}}{100\ \text{atoms K}}$$

$$= 6.9 \times 10^{17}\ \text{atoms }^{40}\text{K}$$

PRACTICE EXAMPLE A: How many Pb atoms are present in a small piece of lead with a volume of 0.105 cm^3? The density of Pb $= 11.34\ g/cm^3$.

PRACTICE EXAMPLE B: Rhenium-187 is a radioactive isotope that can be used to determine the age of meteorites. A 0.100 mg sample of Re contains 2.02×10^{17} atoms of ^{187}Re. What is the percent abundance of rhenium-187 in the sample?

MasteringCHEMISTRY www.masteringchemistry.com

For a discussion of the Occurrence and Abundances of the Elements, go to Focus On 2-1 on the MasteringChemistry site.

Summary

2-1 Early Chemical Discoveries and the Atomic Theory—Modern chemistry began with eighteenth-century discoveries leading to the formulation of two basic laws of chemical combination, the **law of conservation of mass** and the **law of constant composition (definite proportions)**. These discoveries led to Dalton's atomic theory—that matter is composed of indestructible particles called atoms, that the atoms of an element are identical to one another but different from atoms of all other elements, and that chemical compounds are combinations of atoms of different elements. Based on this theory, Dalton proposed still another law of chemical combination, the **law of multiple proportions**.

2-2 Electrons and Other Discoveries in Atomic Physics—The first clues to the structures of atoms came through the discovery and characterization of **cathode rays** (**electrons**). Key experiments were those that established the mass-to-charge ratio (Fig. 2-7) and then the charge on an electron (Fig. 2-8). Two important accidental discoveries made in the course of cathode-ray research were of X-rays and **radioactivity**. The principal types of radiation emitted by radioactive substances are **alpha** ($\boldsymbol{\alpha}$) **particles**, **beta** ($\boldsymbol{\beta}$) **particles**, and **gamma** ($\boldsymbol{\gamma}$) **rays** (Fig. 2-10).

2-3 The Nuclear Atom—Studies on the scattering of α particles by thin metal foils (Fig. 2-11) led to the concept of the nuclear atom—a tiny, but massive, positively charged nucleus surrounded by lightweight, negatively charged electrons (Fig. 2-12). A more complete description of the nucleus was made possible by the discovery of **protons** and **neutrons**. An individual atom is characterized in terms of its **atomic number** (**proton number**) $\boldsymbol{Z}$ and **mass number**, $\boldsymbol{A}$. The difference, $A - Z$, is the **neutron number**. The masses of individual atoms and their component parts are expressed in **atomic mass units** (**u**).

2-4 Chemical Elements—All elements from $Z = 1$ to $Z = 111$ have been characterized and given a name and **chemical symbol**. Knowledge of the several elements following $Z = 111$ is more tenuous. **Nuclide** is the term used to describe an atom with a particular atomic number and a particular mass number. Atoms of the same element that differ in mass number are called **isotopes**. The **percent natural abundance** of an isotope and the precise mass of its atoms can be established with a **mass spectrometer** (Fig. 2-14). A special symbolism (expression 2.2) is used to represent the composition of an atom or an **ion** derived from the atom.

2-5 Atomic Mass—The **atomic mass** (**weight**) of an element is a weighted average based on an assigned value of exactly 12 u for the isotope carbon-12. This weighted average is calculated from the experimentally determined atomic masses and percent abundances of the naturally occurring isotopes of the element through expression (2.3).

2-6 Introduction to the Periodic Table—The **periodic table** (Fig. 2-15) is an arrangement of the elements in horizontal rows called **periods** and vertical columns called **groups** or **families**. Each group consists of elements with similar physical and chemical properties. The elements can also be subdivided into broad categories. One categorization is that of **metals**, **nonmetals**, **metalloids**, and **noble gases**. Another is that of **main-group elements** and **transition elements** (**transition metals**). Included among the transition elements are the two subcategories **lanthanides** and **actinides**. The table has many uses, as will be seen throughout the text. Emphasis in this chapter is on the periodic table as an aid in writing symbols for simple ions.

2-7 The Concept of the Mole and Avogadro Constant—The **Avogadro constant**, $N_A = 6.02214 \times 10^{23}\ \text{mol}^{-1}$, represents the number of carbon-12 atoms in *exactly* 12 g of carbon-12. More generally, it is the number of elementary entities (for example, atoms or molecules) present in an amount known as one **mole** of substance. The mass of one mole of atoms of an element is called its **molar mass**, $\boldsymbol{M}$.

2-8 Using the Mole Concept in Calculations—Molar mass and the Avogadro constant are used in a variety of calculations involving the mass, amount (in moles), and number of atoms in a sample of an element. Other conversion factors may also be involved in these calculations. The mole concept is encountered in ever broader contexts throughout the text.

Integrative Example

A stainless steel ball bearing has a radius of 6.35 mm and a density of 7.75 g/cm^3. Iron is the principal element in steel. Carbon is a key minor element. The ball bearing contains 0.25% carbon, by mass. Given that the percent natural abundance of ^{13}C is 1.108%, how many ^{13}C atoms are present in the ball bearing?

Analyze

The goal is to determine the number of carbon-13 atoms found in a ball bearing with a particular composition. The critical point in this problem is recognizing that we can relate number of atoms to mass by using molar mass and Avogadro's constant. The first step is to use the radius of the ball bearing to determine its volume. The second step is to determine the mass of carbon present by using the density of steel along with the percent composition. The third step uses the molar mass of carbon to convert grams of carbon to moles of carbon; Avogadro's constant is then used to convert moles of carbon to the number of carbon atoms. In the final step, the natural abundance of carbon-13 atoms is used to find the number of carbon-13 atoms in the total number of carbon atoms in the ball bearing.

Solve

The ball-bearing volume in cubic centimeters is found by applying the formula for the volume of a sphere, $V = 4/3\,\pi r^3$. Remember to convert the given radius from millimeters to centimeters, so that the volume will be in cubic centimeters.

$$V = \frac{4\pi}{3}\left[6.35\text{ mm} \times \frac{1\text{ cm}}{10\text{ mm}}\right]^3 = 1.07\text{ cm}^3$$

The product of the volume of the ball bearing and the density of steel equals the mass. The mass of the ball bearing multiplied by the percent carbon in the steel gives the mass of carbon present.

$$?\text{ g C} = 1.07\text{ cm}^3 \times \frac{7.75\text{ g steel}}{1\text{ cm}^3\text{ steel}} \times \frac{0.25\text{ g C}}{100\text{ g steel}} = 0.021\text{ g C}$$

The mass of carbon is first converted to moles of carbon by using the inverse of the molar mass of carbon. Avogadro's constant is then used to convert moles of carbon to atoms of carbon.

$$?\text{ C atoms} = 0.021\text{ g C} \times \frac{1\text{ mol C}}{12.011\text{ g C}} \times \frac{6.022 \times 10^{23}\text{ C atoms}}{1\text{ mol C}}$$
$$= 1.1 \times 10^{21}\text{ C atoms}$$

The number of ^{13}C atoms is determined by using the percent natural abundance of carbon-13.

$$?\ ^{13}\text{C atoms} = 1.1 \times 10^{21}\text{ C atoms} \times \frac{1.108\ ^{13}\text{C atoms}}{100\text{ C atoms}}$$
$$= 1.2 \times 10^{19}\ ^{13}\text{C atoms}$$

Assess

The number of carbon-13 atoms is smaller than the number of carbon atoms, which it should be, given that the natural abundance of carbon-13 is just 1.108%. To avoid mistakes, every quantity should be clearly labeled with its appropriate unit so that units cancel properly. Two points made by this problem are, first, that the relatively small ball bearing contains a large number of carbon-13 atoms even though carbon-13's abundance is only 1.108% of all carbon atoms. Second, the size of any atom must be very small.

PRACTICE EXAMPLE A: Calculate the number of ^{63}Cu atoms in a cubic crystal of copper that measures 25 nm on edge. The density of copper is 8.92 g/cm^3 and the percent natural abundance of ^{63}Cu is 69.17%.

PRACTICE EXAMPLE B: The United States Food and Drug Administration (USFDA) suggests a daily value of 18 mg Fe for adults and for children over four years of age. The label on a particular brand of cereal states that one serving (55 g) of dry cereal contains 45% of the daily value of Fe. Given that the percent natural abundance of ^{58}Fe is 0.282%, how many full servings of dry cereal must be eaten to consume exactly one mole of ^{58}Fe? The atomic weight of ^{58}Fe is 57.9333 u. Is it possible for a person to consume this much cereal in a lifetime, assuming that one full serving of cereal is eaten every day?

Exercises

Law of Conservation of Mass

1. When an iron object rusts, its mass increases. When a match burns, its mass decreases. Do these observations violate the law of conservation of mass? Explain.
2. When a strip of magnesium metal is burned in air (recall Figure 2-1), it produces a white powder that weighs more than the original metal. When a strip of magnesium is burned in a flashbulb, the bulb weighs the same before and after it is flashed. Explain the difference in these observations.
3. A 0.406 g sample of magnesium reacts with oxygen, producing 0.674 g of magnesium oxide as the only product. What mass of oxygen was consumed in the reaction?
4. A 1.446 g sample of potassium reacts with 8.178 g of chlorine to produce potassium chloride as the only product. After the reaction, 3.960 g of chlorine remains unreacted. What mass of potassium chloride was formed?
5. When a solid mixture consisting of 10.500 g calcium hydroxide and 11.125 g ammonium chloride is strongly heated, gaseous products are evolved and 14.336 g of a solid residue remains. The gases are passed into 62.316 g water, and the mass of the resulting solution is 69.605 g. Within the limits of experimental error, show that these data conform to the law of conservation of mass.
6. Within the limits of experimental error, show that the law of conservation of mass was obeyed in the following experiment: 10.00 g calcium carbonate (found in limestone) was dissolved in 100.0 mL hydrochloric acid ($d = 1.148$ g/mL). The products were 120.40 g solution (a mixture of hydrochloric acid and calcium chloride) and 2.22 L carbon dioxide gas ($d = 1.9769$ g/L).

Law of Constant Composition

7. In Example 2-1, we established that the mass ratio of magnesium to magnesium oxide is 0.455 g magnesium/ 0.755 g magnesium oxide.
(a) What is the ratio of oxygen to magnesium oxide, by mass?
(b) What is the mass ratio of oxygen to magnesium in magnesium oxide?
(c) What is the percent by mass of magnesium in magnesium oxide?
8. Samples of pure carbon weighing 3.62, 5.91, and 7.07 g were burned in an excess of air. The masses of carbon dioxide obtained (the sole product in each case) were 13.26, 21.66, and 25.91 g, respectively
(a) Do these data establish that carbon dioxide has a fixed composition?
(b) What is the composition of carbon dioxide, expressed in % C and % O, by mass?
9. In one experiment, 2.18 g sodium was allowed to react with 16.12 g chlorine. All the sodium was used up, and 5.54 g sodium chloride (salt) was produced. In a second experiment, 2.10 g chlorine was allowed to react with 10.00 g sodium. All the chlorine was used up, and 3.46 g sodium chloride was produced. Show that these results are consistent with the law of constant composition.
10. When 3.06 g hydrogen was allowed to react with an excess of oxygen, 27.35 g water was obtained. In a second experiment, a sample of water was decomposed by electrolysis, resulting in 1.45 g hydrogen and 11.51 g oxygen. Are these results consistent with the law of constant composition? Demonstrate why or why not.
11. In one experiment, the burning of 0.312 g sulfur produced 0.623 g sulfur dioxide as the sole product of the reaction. In a second experiment, 0.842 g sulfur dioxide was obtained. What mass of sulfur must have been burned in the second experiment?
12. In one experiment, the reaction of 1.00 g mercury and an excess of sulfur yielded 1.16 g of a sulfide of mercury as the sole product. In a second experiment, the same sulfide was produced in the reaction of 1.50 g mercury and 1.00 g sulfur
(a) What mass of the sulfide of mercury was produced in the second experiment?
(b) What mass of which element (mercury or sulfur) remained *unreacted* in the second experiment?

Law of Multiple Proportions

13. Sulfur forms two compounds with oxygen. In the first compound, 1.000 g sulfur is combined with 0.998 g oxygen, and in the second, 1.000 g sulfur is combined with 1.497 g oxygen. Show that these results are consistent with Dalton's law of multiple proportions.
14. Phosphorus forms two compounds with chlorine. In the first compound, 1.000 g of phosphorus is combined with 3.433 g chlorine, and in the second, 2.500 g phosphorus is combined with 14.308 g chlorine. Show that these results are consistent with Dalton's law of multiple proportions.
15. The following data were obtained for compounds of nitrogen and hydrogen:

Compound	Mass of Nitrogen, g	Mass of Hydrogen, g
A	0.500	0.108
B	1.000	0.0720
C	0.750	0.108

(a) Show that these data are consistent with the law of multiple proportions.
(b) If the formula of compound B is N_2H_2, what are the formulas of compounds A and C?

16. The following data were obtained for compounds of iodine and fluorine:

Compound	Mass of Iodine, g	Mass of Fluorine, g
A	1.000	0.1497
B	0.500	0.2246
C	0.750	0.5614
D	1.000	1.0480

(a) Show that these data are consistent with the law of multiple proportions.
(b) If the formula for compound A is IF, what are the formulas for compounds B, C, and D?
17. There are two oxides of copper. One oxide has 20% oxygen, by mass. The second oxide has a *smaller* percent of oxygen than the first. What is the probable percent of oxygen in the second oxide?
18. The two oxides of carbon described on page 38 were CO and CO_2. Another oxide of carbon has 1.106 g of oxygen in a 2.350 g sample. In what ratio are carbon and oxygen atoms combined in molecules of this third oxide? Explain.

Fundamental Charges and Mass-to-Charge Ratios

19. The following observations were made for a series of five oil drops in an experiment similar to Millikan's (see Figure 2-8). Drop 1 carried a charge of 1.28×10^{-18} C; drops 2 and 3 each carried $\frac{1}{2}$ the charge of drop 1; drop 4 carried $\frac{1}{8}$ the charge of drop 1; drop 5 had a charge four times that of drop 1. Are these data consistent with the value of the electronic charge given in the text? Could Millikan have inferred the charge on the electron from this particular series of data? Explain.
20. In an experiment similar to that described in Exercise 19, drop 1 carried a charge of 6.41×10^{-19} C; drop 2 had $\frac{1}{2}$ the charge of drop 1; drop 3 had twice the charge

of drop 1; drop 4 had a charge of 1.44×10^{-18} C; and drop 5 had $\frac{1}{3}$ the charge of drop 4. Are these data consistent with the value of the electronic charge given in the text? Could Millikan have inferred the charge on the electron from this particular series of data? Explain.

21. Use data from Table 2.1 to verify that
(a) the mass of electrons is about 1/2000 that of H atoms;
(b) the mass-to-charge ratio (m/e) for positive ions is considerably larger than that for electrons.

22. Determine the approximate value of m/e in coulombs per gram for the ions $^{127}_{53}I^-$ and $^{32}_{16}S^{2-}$. Why are these values only approximate?

Atomic Number, Mass Number, and Isotopes

23. The following radioactive isotopes have applications in medicine. Write their symbols in the form $^{A}_{Z}E$. **(a)** cobalt-60; **(b)** phosphorus-32; **(c)** iron-59; **(d)** radium-226.

24. For the isotope ^{202}Hg, express the percentage of the fundamental particles in the nucleus that are neutrons.

25. Complete the following table. What minimum amount of information is required to completely characterize an atom or ion?
[*Hint:* Not all rows can be completed.]

Name	Symbol	Number Protons	Number Electrons	Number Neutrons	Mass Number
Sodium	$^{23}_{11}Na$	11	11	12	23
Silicon	—	—	—	14	—
—	—	37	—	—	85
—	^{40}K	—	—	—	—
—	—	—	33	42	—
—	$^{20}Ne^{2+}$	—	—	—	—
—	—	—	—	—	80
—	—	—	—	126	—

26. Arrange the following species in order of increasing **(a)** number of electrons; **(b)** number of neutrons; **(c)** mass.

$$^{112}_{50}Sn \quad ^{40}_{18}Ar \quad ^{122}_{52}Te \quad ^{59}_{29}Cu \quad ^{120}_{48}Cd \quad ^{58}_{27}Co \quad ^{39}_{19}K$$

27. For the atom ^{108}Pd with mass 107.90389 u, determine
(a) the numbers of protons, neutrons, and electrons in the atom;
(b) the ratio of the mass of this atom to that of an atom of $^{12}_{6}C$.

28. For the ion $^{228}Ra^{2+}$ with a mass of 228.030 u, determine
(a) the numbers of protons, neutrons, and electrons in the ion;
(b) the ratio of the mass of this ion to that of an atom of ^{16}O (refer to page 47).

29. An isotope of silver has a mass that is 6.68374 times that of oxygen-16. What is the mass in u of this isotope? (Refer to page 47.)

30. The ratio of the masses of the two naturally occurring isotopes of indium is 1.0177:1. The heavier of the two isotopes has 7.1838 times the mass of ^{16}O. What are the masses in u of the two isotopes? (Refer to page 47.)

31. The following data on isotopic masses are from a chemical handbook. What is the ratio of each of these masses to that of $^{12}_{6}C$? **(a)** $^{35}_{17}Cl$, 34.96885 u; **(b)** $^{26}_{12}Mg$, 25.98259 u; **(c)** $^{222}_{86}Rn$, 222.0175 u.

32. The following ratios of masses were obtained with a mass spectrometer: $^{19}_{9}F/^{12}_{6}C = 1.5832$; $^{35}_{17}Cl/^{19}_{9}F = 1.8406$; $^{81}_{35}Br/^{35}_{17}Cl = 2.3140$. Determine the mass of a $^{81}_{35}Br$ atom in amu.

33. Which of the following species has
(a) equal numbers of neutrons and electrons;
(b) protons, neutrons, and electrons in the ratio 9:11:8;
(c) a number of neutrons equal to the number of protons plus one-half the number of electrons?

$$^{24}Mg^{2+},\ ^{47}Cr,\ ^{60}Co^{3+},\ ^{35}Cl^-,\ ^{124}Sn^{2+},\ ^{226}Th,\ ^{90}Sr$$

34. Given the same species as listed in Exercise 33, which has
(a) equal numbers of neutrons and protons;
(b) protons contributing more than 50% of the mass;
(c) about 50% more neutrons than protons?

35. An isotope with mass number 44 has four more neutrons than protons. This is an isotope of what element?

36. Identify the isotope X that has one more neutron than protons and a mass number equal to nine times the charge on the ion X^{3+}.

37. Iodine has many radioactive isotopes. Iodine-123 is a radioactive isotope used for obtaining images of the thyroid gland. Iodine-123 is administered to patients in the form of sodium iodide capsules that contain $^{123}I^-$ ions. Determine the number of neutrons, protons, and electrons in a single $^{123}I^-$ ion.

38. Iodine-131 is a radioactive isotope that has important medical uses. Small doses of iodine-131 are used for treating hyperthyroidism (overactive thyroid) and larger doses are used for treating thyroid cancer. Iodine-131 is administered to patients in the form of sodium iodide capsules that contain $^{131}I^-$ ions. Determine the number of neutrons, protons, and electrons in a single $^{131}I^-$ ion.

39. Americium-241 is a radioactive isotope that is used in high-precision gas and smoke detectors. How many neutrons, protons, and electrons are there in an atom of americium-241?

40. Some foods are made safer to eat by being exposed to gamma rays from radioactive isotopes, such as cobalt-60. The energy from the gamma rays kills bacteria in the food. How many neutrons, protons, and electrons are there in an atom of cobalt-60?

Atomic Mass Units, Atomic Masses

41. Which statement is probably true concerning the masses of *individual* chlorine atoms: *All have*, *some have*, or *none has* a mass of 35.4527 u? Explain.
42. The mass of a carbon-12 atom is taken to be exactly 12 u. Are there likely to be any other atoms with an *exact* integral (whole number) mass, expressed in u? Explain.
43. There are three naturally occurring isotopes of magnesium. Their masses and percent natural abundances are 23.985042 u, 78.99%; 24.985837 u, 10.00%; and 25.982593 u, 11.01%. Calculate the weighted-average atomic mass of magnesium.
44. There are four naturally occurring isotopes of chromium. Their masses and percent natural abundances are 49.9461 u, 4.35%; 51.9405 u, 83.79%; 52.9407 u, 9.50%; and 53.9389 u, 2.36%. Calculate the weighted-average atomic mass of chromium.
45. The two naturally occurring isotopes of silver have the following abundances: ^{107}Ag, 51.84%; ^{109}Ag, 48.16%. The mass of ^{107}Ag is 106.905092 u. What is the mass of ^{109}Ag?
46. Bromine has two naturally occurring isotopes. One of them, bromine-79, has a mass of 78.918336 u and a natural abundance of 50.69%. What must be the mass and percent natural abundance of the other isotope, bromine-81?
47. The three naturally occurring isotopes of potassium are ^{39}K, 38.963707 u; ^{40}K, 39.963999 u; and ^{41}K. The percent natural abundances of ^{39}K and ^{41}K are 93.2581% and 6.7302%, respectively. Determine the isotopic mass of ^{41}K.
48. What are the percent natural abundances of the two naturally occurring isotopes of boron, ^{10}B and ^{11}B? These isotopes have masses of 10.012937 u and 11.009305 u, respectively.

Mass Spectrometry

49. A mass spectrum of germanium displayed peaks at mass numbers 70, 72, 73, 74, and 76, with relative heights of 20.5, 27.4, 7.8, 36.5, and 7.8, respectively.
(a) In the manner of Figure 2-14, sketch this mass spectrum.
(b) Estimate the weighted-average atomic mass of germanium, and state why this result is only approximately correct.
50. Hydrogen and chlorine atoms react to form simple diatomic molecules in a 1:1 ratio, that is, HCl. The natural abundances of the chlorine isotopes are 75.77% ^{35}Cl and 24.23% ^{37}Cl. The natural abundances of ^{2}H and ^{3}H are 0.015% and less than 0.001%, respectively.
(a) How many different HCl molecules are possible, and what are their mass numbers (that is, the sum of the mass numbers of the H and Cl atoms)?
(b) Which is the most abundant of the possible HCl molecules? Which is the second most abundant?

The Periodic Table

51. Refer to the periodic table inside the front cover and identify
(a) the element that is in group 14 and the fourth period
(b) one element similar to and one unlike sulfur
(c) the alkali metal in the fifth period
(d) the halogen element in the sixth period
52. Refer to the periodic table inside the front cover and identify
(a) the element that is in group 11 and the sixth period
(b) an element with atomic number greater than 50 that has properties similar to the element with atomic number 18
(c) the group number of an element E that forms an ion E^{2-}
(d) an element M that you would expect to form the ion M^{3+}
53. Assuming that the seventh period of the periodic table has 32 members, what should be the atomic number of **(a)** the noble gas following radon (Rn); **(b)** the alkali metal following francium (Fr)?
54. Find the several pairs of elements that are "out of order" in terms of increasing atomic mass and explain why the reverse order is necessary.

The Avogadro Constant and the Mole

55. What is the total number of atoms in **(a)** 15.8 mol Fe; **(b)** 0.000467 mol Ag; **(c)** 8.5×10^{-11} mol Na?
56. *Without doing detailed calculations*, indicate which of the following quantities contains the greatest number of atoms: 6.022×10^{23} Ni atoms, 25.0 g nitrogen, 52.0 g Cr, 10.0 cm^3 Fe ($d = 7.86$ g/cm^3). Explain your reasoning.
57. Determine
(a) the number of moles of Zn in a 415.0 g sample of zinc metal

(b) the number of Cr atoms in 147.4 kg chromium
(c) the mass of a one-trillion-atom (1.0×10^{12}) sample of metallic gold
(d) the mass of one fluorine atom

58. Determine
(a) the number of Kr atoms in a 5.25-mg sample of krypton
(b) the molar mass, M, and identity of an element if the mass of a 2.80×10^{22}-atom sample of the element is 2.09 g
(c) the mass of a sample of phosphorus that contains the same number of atoms as 44.75 g of magnesium

59. How many Cu atoms are present in a piece of sterling-silver jewelry weighing 33.24 g? (Sterling silver is a silver–copper alloy containing 92.5% Ag by mass.)

60. How many atoms are present in a 75.0 cm^3 sample of plumber's solder, a lead–tin alloy containing 67% Pb by mass and having a density of 9.4 g/cm^3?

61. How many ^{204}Pb atoms are present in a piece of lead weighing 215 mg? The percent natural abundance of ^{204}Pb is 1.4%.

62. A particular lead–cadmium alloy is 8.0% cadmium by mass. What mass of this alloy, in grams, must you weigh out to obtain a sample containing 6.50×10^{23} Cd atoms?

63. Medical experts generally believe a level of 30 μg Pb per deciliter of blood poses a significant health risk (1 dL = 0.1 L). Express this level **(a)** in the unit mol Pb/L blood; **(b)** as the number of Pb atoms per milliliter blood.

64. During a severe episode of air pollution, the concentration of lead in the air was observed to be 3.01 μg Pb/m^3. How many Pb atoms would be present in a 0.500 L sample of this air (the approximate lung capacity of a human adult)?

65. *Without doing detailed calculations*, determine which of the following samples has the greatest number of atoms:
(a) a cube of iron with a length of 10.0 cm ($d = 7.86\ g/cm^3$)
(b) 1.00 kg of hydrogen contained in a 10,000 L balloon
(c) a mound of sulfur weighing 20.0 kg
(d) a 76 lb sample of liquid mercury ($d = 13.5\ g/mL$)

66. *Without doing detailed calculations*, determine which of the following samples occupies the largest volume:
(a) 25.5 mol of sodium metal ($d = 0.971\ g/cm^3$)
(b) 0.725 L of liquid bromine ($d = 3.12\ g/mL$)
(c) 1.25×10^{25} atoms of chromium metal ($d = 9.4\ g/cm^3$)
(d) 2.15 kg of plumber's solder ($d = 9.4\ g/cm^3$), a lead–tin alloy with a 2:1 atom ratio of lead to tin

Integrative and Advanced Exercises

67. A solution was prepared by dissolving 2.50 g potassium chlorate (a substance used in fireworks and flares) in 100.0 mL water at 40 °C. When the solution was cooled to 20 °C, its volume was still found to be 100.0 mL, but some of the potassium chlorate had crystallized (deposited from the solution as a solid). At 40 °C, the density of water is 0.9922 g/mL, and at 20 °C, the potassium chlorate solution had a density of 1.0085 g/mL.
(a) Estimate, to two significant figures, the mass of potassium perchlorate that crystallized.
(b) Why can't the answer in **(a)** be given more precisely?

68. William Prout (1815) proposed that all other atoms are built up of hydrogen atoms, suggesting that all elements should have integral atomic masses based on an atomic mass of one for hydrogen. This hypothesis appeared discredited by the discovery of atomic masses, such as 24.3 u for magnesium and 35.5 u for chlorine. In terms of modern knowledge, explain why Prout's hypothesis is actually quite reasonable.

69. Fluorine has a single atomic species, ^{19}F. Determine the atomic mass of ^{19}F by summing the masses of its protons, neutrons, and electrons, and compare your results with the value listed on the inside front cover. Explain why the agreement is poor.

70. Use 1×10^{-13} cm as the approximate diameter of the spherical nucleus of the hydrogen-1 atom, together with data from Table 2.1, to estimate the density of matter in a proton.

71. Use fundamental definitions and statements from Chapters 1 and 2 to establish the fact that $6.022 \times 10^{23}\ u = 1.000\ g$.

72. In each case, identify the element in question.
(a) The mass number of an atom is 234 and the atom has 60.0% more neutrons than protons.
(b) An ion with a 2+ charge has 10.0% more protons than electrons.
(c) An ion with a mass number of 110 and a 2+ charge has 25.0% more neutrons than electrons.

73. Determine the only possible 2+ ion for which the following two conditions are both satisfied:
- The net ionic charge is *one-tenth* the nuclear charge.
- The number of neutrons is *four* more than the number of electrons.

74. Determine the only possible isotope (E) for which the following conditions are met:
- The mass number of E is 2.50 times its atomic number.
- The atomic number of E is equal to the mass number of another isotope (Y). In turn, isotope Y has a neutron number that is 1.33 times the atomic number of Y and equal to the neutron number of selenium-82.

75. Suppose we redefined the atomic mass scale by arbitrarily assigning to the naturally occurring *mixture* of chlorine isotopes an atomic mass of 35.00000 u.
(a) What would be the atomic masses of helium, sodium, and iodine on this new atomic mass scale?

(b) Why do these three elements have nearly integral (whole-number) atomic masses based on carbon-12, but not based on naturally occurring chlorine?

76. The two naturally occurring isotopes of nitrogen have masses of 14.0031 and 15.0001 u, respectively. Determine the percentage of ^{15}N atoms in naturally occurring nitrogen.

77. The masses of the naturally occurring mercury isotopes are ^{196}Hg, 195.9658 u; ^{198}Hg, 197.9668 u; ^{199}Hg, 198.9683 u; ^{200}Hg, 199.9683 u; ^{201}Hg, 200.9703 u; ^{202}Hg, 201.9706 u; and ^{204}Hg, 203.9735 u. Use these data, together with data from Figure 2-14, to calculate the weighted-average atomic mass of mercury.

78. Germanium has three major naturally occurring isotopes: ^{70}Ge (69.92425 u, 20.85%), ^{72}Ge (71.92208 u, 27.54%), ^{74}Ge (73.92118 u, 36.29%). There are also two minor isotopes: ^{73}Ge (72.92346 u) and ^{76}Ge (75.92140 u). Calculate the percent natural abundances of the two minor isotopes. Comment on the precision of these calculations.

79. From the densities of the lines in the mass spectrum of krypton gas, the following observations were made:
- Somewhat more than 50% of the atoms were krypton-84.
- The numbers of krypton-82 and krypton-83 atoms were essentially equal.
- The number of krypton-86 atoms was 1.50 times as great as the number of krypton-82 atoms.
- The number of krypton-80 atoms was 19.6% of the number of krypton-82 atoms.
- The number of krypton-78 atoms was 3.0% of the number of krypton-82 atoms.

The masses of the isotopes are

^{78}Kr, 77.9204 u	^{80}Kr, 79.9164 u	^{82}Kr, 81.9135 u
^{83}Kr, 82.9141 u	^{84}Kr, 83.9115 u	^{86}Kr, 85.9106 u

The weighted-average atomic mass of Kr is 83.80. Use these data to calculate the percent natural abundances of the krypton isotopes.

80. The two naturally occurring isotopes of chlorine are ^{35}Cl (34.9689 u, 75.77%) and ^{37}Cl (36.9658 u, 24.23%). The two naturally occurring isotopes of bromine are ^{79}Br (78.9183 u, 50.69%) and ^{81}Br (80.9163 u, 49.31%). Chlorine and bromine combine to form bromine monochloride, BrCl. Sketch a mass spectrum for BrCl with the relative number of molecules plotted against molecular mass (similar to Figure 2-14b).

81. How many atoms are present in a 1.00 m length of 20-gauge copper wire? A 20-gauge wire has a diameter of 0.03196 in., and the density of copper is 8.92 g/cm^3.

82. Monel metal is a corrosion-resistant copper–nickel alloy used in the electronics industry. A particular alloy with a density of 8.80 g/cm^3 and containing 0.022% Si by mass is used to make a rectangular plate 15.0 cm long, 12.5 cm wide, 3.00 mm thick, and has a 2.50 cm diameter hole drilled through its center. How many silicon-30 atoms are found in this plate? The mass of a silicon-30 atom is 29.97376 u, and the percent natural abundance of silicon-30 is 3.10%.

83. Deuterium, ^2H (2.0140 u), is sometimes used to replace the principal hydrogen isotope ^1H in chemical studies. The percent natural abundance of deuterium is 0.015%. If it can be done with 100% efficiency, what mass of naturally occurring hydrogen gas would have to be processed to obtain a sample containing 2.50×10^{21} ^2H atoms?

84. An alloy that melts at about the boiling point of water has Bi, Pb, and Sn atoms in the ratio 10:6:5, respectively. What mass of alloy contains a total of one mole of atoms?

85. A particular silver solder (used in the electronics industry to join electrical components) is to have the *atom* ratio of 5.00 Ag/4.00 Cu/1.00 Zn. What masses of the three metals must be melted together to prepare 1.00 kg of the solder?

86. A low-melting Sn–Pb–Cd alloy called *eutectic alloy* is analyzed. The *mole* ratio of tin to lead is 2.73:1.00, and the *mass* ratio of lead to cadmium is 1.78:1.00. What is the mass percent composition of this alloy?

87. In an experiment, 125 cm^3 of zinc and 125 cm^3 of iodine are mixed together and the iodine is completely converted to 164 cm^3 of zinc iodide. What volume of zinc remains unreacted? The densities of zinc, iodine, and zinc iodide are 7.13 g/cm^3, 4.93 g/cm^3, and 4.74 g/cm^3, respectively.

88. Atoms are spherical and so when silver atoms pack together to form silver metal, they cannot fill all the available space. In a sample of silver metal, approximately 26.0% of the sample is empty space. Given that the density of silver metal is 10.5 g/cm^3, what is the radius of a silver atom? Express your answer in picometers.

Feature Problems

89. The data Lavoisier obtained in the experiment described on page 35 are as follows:

Before heating: glass vessel + tin + air

= 13 onces, 2 gros, 2.50 grains

After heating: glass vessel + tin calx + remaining air

= 13 onces, 2 gros, 5.62 grains

How closely did Lavoisier's results conform to the law of conservation of mass? (1 livre = 16 onces; 1 once = 8 gros; 1 gros = 72 grains. In modern terms, 1 livre = 30.59 g.)

90. Some of Millikan's oil-drop data are shown on the next page. The measured quantities were not actual charges on oil drops but were proportional to these charges.

Show that these data are consistent with the idea of a fundamental electronic charge.

Observation	Measured Quantity	Observation	Measured Quantity
1	19.66	8	53.91
2	24.60	9	59.12
3	29.62	10	63.68
4	34.47	11	68.65
5	39.38	12	78.34
6	44.42	13	83.22
7	49.41		

91. Before 1961, the standard for atomic masses was the isotope ^{16}O, to which physicists assigned a value of exactly 16. At the same time, chemists assigned a value of exactly 16 to the naturally occurring mixture of the isotopes ^{16}O, ^{17}O, and ^{18}O. Would you expect atomic masses listed in a 50-year-old text to be the same, generally higher, or generally lower than in this text? Explain.

92. The German chemist Fritz Haber proposed paying off the reparations imposed against Germany after World War I by extracting gold from seawater. Given that (1) the amount of the reparations was $28.8 billion dollars, (2) the value of gold at the time was about $21.25 per troy ounce (1 troy ounce = 31.103 g), and (3) gold occurs in seawater to the extent of 4.67×10^{17} atoms per ton of seawater (1 ton = 2000 lb), how many cubic kilometers of seawater would have had to be processed to obtain the required amount of gold? Assume that the density of seawater is 1.03 g/cm^3. (Haber's scheme proved to be commercially infeasible, and the reparations were never fully paid.)

93. Mass spectrometry is one of the most versatile and powerful tools in chemical analysis because of its capacity to discriminate between atoms of different masses. When a sample containing a mixture of isotopes is introduced into a mass spectrometer, the ratio of the peaks observed reflects the ratio of the percent natural abundances of the isotopes. This ratio provides an internal standard from which the amount of a certain isotope present in a sample can be determined. This is accomplished by deliberately introducing a known quantity of a particular isotope into the sample to be analyzed. A comparison of the new isotope ratio to the first ratio allows the determination of the amount of the isotope present in the original sample.

An analysis was done on a rock sample to determine its rubidium content. The rubidium content of a portion of rock weighing 0.350 g was extracted, and to the extracted sample was added an additional 29.45 μg of ^{87}Rb. The mass spectrum of this spiked sample showed a ^{87}Rb peak that was 1.12 times as high as the peak for ^{85}Rb. Assuming that the two isotopes react identically, what is the Rb content of the rock (expressed in parts per million by mass)? The natural abundances and isotopic masses are shown in the table.

Isotope	% Natural Abundance	Atomic Mass, u
^{87}Rb	27.83	86.909
^{85}Rb	72.17	84.912

Self-Assessment Exercises

94. In your own words, define or explain these terms or symbols: **(a)** $^{A}_{Z}E$; **(b)** β particle; **(c)** isotope; **(d)** ^{16}O; **(e)** molar mass.

95. Briefly describe
(a) the law of conservation of mass
(b) Rutherford's nuclear atom
(c) weighted-average atomic mass
(d) a mass spectrum

96. Explain the important distinctions between each pair of terms:
(a) cathode rays and X-rays
(b) protons and neutrons
(c) nuclear charge and ionic charge
(d) periods and groups of the periodic table
(e) metal and nonmetal
(f) the Avogadro constant and the mole

97. When 10.0 g zinc and 8.0 g sulfur are allowed to react, all the zinc is consumed, 15.0 g zinc sulfide is produced, and the mass of unreacted sulfur remaining is
(a) 2.0 g
(b) 3.0 g
(c) 5.0 g
(d) impossible to predict from this information alone

98. One oxide of rubidium has 0.187 g O per gram of Rb. A possible O:Rb mass ratio for a second oxide of rubidium is **(a)** 16:85.5; **(b)** 8:42.7; **(c)** 1:2.674; **(d)** any of these.

99. Cathode rays
(a) may be positively or negatively charged
(b) are a form of electromagnetic radiation similar to visible light
(c) have properties identical to β particles
(d) have masses that depend on the cathode that emits them

100. The scattering of α particles by thin metal foils established that
(a) the mass of an atom is concentrated in a positively charged nucleus
(b) electrons are fundamental particles of all matter
(c) all electrons carry the same charge
(d) atoms are electrically neutral

101. Which of the following have the same charge and approximately the same mass?
(a) an electron and a proton; **(b)** a proton and a neutron; **(c)** a hydrogen atom and a proton; **(d)** a neutron and a hydrogen atom; **(e)** an electron and an H^- ion.

102. What is the correct symbol for the species that contains 18 neutrons, 17 protons, and 16 electrons?

103. The properties of magnesium will most resemble those of which of the following? **(a)** cesium; **(b)** sodium; **(c)** aluminum; **(d)** calcium; **(e)** manganese.

104. Which group in the main group of elements contains **(a)** no metals or metalloids? **(b)** only one metal or metalloid? **(c)** only one nonmetal? **(d)** only nonmetals?

105. The two species that have the same number of electrons as ^{32}S are **(a)** ^{32}Cl; **(b)** $^{34}S^+$; **(c)** $^{33}P^+$; **(d)** $^{28}Si^{2-}$; **(e)** $^{35}S^{2-}$; **(f)** $^{40}Ar^{2+}$; **(g)** $^{40}Ca^{2+}$.

106. To four significant figures, all of the following masses are possible for an individual titanium atom except one. The exception is **(a)** 45.95 u; **(b)** 46.95 u; **(c)** 47.88 u; **(d)** 47.95 u; **(e)** 48.95 u; **(f)** 49.94 u.

107. The mass of the isotope $^{84}_{36}Xe$ is 83.9115 u. If the atomic mass scale were redefined so that $^{84}_{36}Xe = 84$ u, *exactly*, the mass of the $^{12}_{6}C$ isotope would be **(a)** 11.9115 u; **(b)** 11.9874 u; **(c)** 12 u exactly; **(d)** 12.0127 u; **(e)** 12.0885 u.

108. A 5.585-kg sample of iron (Fe) contains
(a) 10.0 mol Fe
(b) twice as many atoms as does 600.6 g C
(c) 10 times as many atoms as does 52.00 g Cr
(d) 6.022×10^{24} atoms

109. There are three common iron-oxygen compounds. The one with the greatest proportion of iron has one Fe atom for every O atom and the formula FeO. A second compound has 2.327 g Fe per 1.000 g O, and the third has 2.618 g Fe per 1.000 g O. What are the formulas of these other two iron-oxygen compounds?

110. The four naturally occurring isotopes of strontium have the atomic masses 83.9134 u; 85.9093 u; 86.9089 u; and 87.9056 u. The percent natural abundance of the lightest isotope is 0.56% and of the heaviest, 82.58%. Estimate the percent natural abundances of the other two. Why is this result only a rough approximation?

111. Gold is present in seawater to the extent of 0.15 mg/ton. Assume the density of the seawater is 1.03 g/mL and determine how many Au atoms could conceivably be extracted from 0.250 L of seawater ($1 \text{ ton} = 2.000 \times 10^3$ lb; 1 kg = 2.205 lb).

112. Appendix E describes a useful study aid known as concept mapping. Using the method presented in Appendix E, construct a concept map illustrating the different concepts in Sections 2-7 and 2-8.