# 2 The Components of Matter

## Key Principles
### to focus on while studying this chapter

- A *substance* is matter with a *fixed composition.* The two types of substances are elements and compounds: an *element* consists of a single type of *atom*, which may occur separately or as *molecules;* a *compound* consists of *molecules* (or formula units) made up of two or more different atoms combined in a fixed ratio. A *mixture* consists of two or more substances intermingled physically and, thus, has a *variable composition.* A compound's properties differ from those of its components, but a mixture's properties do not. ***(Section 2.1)***
- Three *mass laws* led to an atomic theory of matter: *mass is conserved* during a chemical change; any sample of a compound has its elements in the *same proportions by mass*; in different compounds made of the same two elements, the masses of one element that combine with a given mass of the other can be expressed as a *ratio of small integers.* ***(Section 2.2)***
- According to *Dalton's atomic theory*, atoms of a given element have a unique mass and other properties. Mass is conserved during a chemical reaction because the atoms of the reacting substances are just rearranged into different substances. ***(Section 2.3)***
- Early $20^{th}$-century experiments showed that atoms *are* divisible, consisting of a positively charged *nucleus*, which contains nearly all the atom's mass but represents a tiny fraction of its volume, and negatively charged *electrons* that surround the nucleus. ***(Section 2.4)***
- Atoms consist of three types of *subatomic particles*: positively charged *protons* and uncharged *neutrons* make up the *nucleus*, and *electrons* exist outside the nucleus. *Atoms are neutral* because the number of protons equals the number of electrons. All the atoms of an element have the same number of protons *(atomic number, Z)* and thus the same chemical behavior. *Isotopes* of an element have atoms with different masses because they have different numbers of neutrons. The *atomic mass* of an element is the weighted *average* of the masses of its naturally occurring isotopes. ***(Section 2.5)***
- In the *periodic table*, the elements are arranged by increasing atomic number into a grid of horizontal rows *(periods)* and vertical columns *(groups). Metals* occupy the lower-left three-quarters of the table, and *nonmetals* are found in the upper-right corner, with *metalloids* in between. Elements in a group have similar properties. ***(Section 2.6)***
- The electrons of atoms are involved in forming compounds. In the formation of *ionic compounds*, metal atoms *transfer* electrons to nonmetal atoms, and the resulting charged particles *(ions)* attract each other into solid arrays. In the formation of *covalent compounds*, nonmetal atoms *share* electrons and usually (but not always) form individual molecules. Each compound has a unique name, formula, and mass based on its component elements. ***(Sections 2.7 and 2.8)***
- Unlike compounds, mixtures can be *separated by physical means* into their components. A *heterogeneous mixture* has a nonuniform composition with visible boundaries between the components. A *homogeneous mixture (solution)* has a uniform composition because the components (elements and/or compounds) are mixed as individual atoms, ions, or molecules. ***(Section 2.9)***

**Parts Within Parts Within Parts** Like this piece of granite, everyday matter consists of simpler parts made of even simpler parts. In this chapter, you'll learn about the components of matter and see how they combine.

## Outline

Look closely at almost any sample of matter—a rock, a piece of wood, a butterfly wing—and you'll see that it's made of smaller parts. With a microscope, you'll see still smaller parts. And, if you could zoom in a billion times closer, you'd find, on the atomic scale, the ultimate particles that make up all things.

Modern scientists are not the first to try to explain what things are made of. The philosophers of ancient Greece believed that everything was made of one or, at most, a few elemental substances (elements), whose properties gave rise to the properties of everything. But, Democritus (c. 460–370 BC), the father of atomism, took a different approach. He reasoned that if you cut a piece of, say, aluminum foil smaller and smaller, you reach a particle of aluminum too small to cut, so matter must be ultimately composed of indivisible particles with nothing but empty space between them. He called the particles *atoms* (Greek *atomos,* "uncuttable"). But, Aristotle (384–322 BC), one of the greatest philosophers of Western culture, said it was impossible for "nothing" to exist, and the concept of atoms was suppressed for 2000 years.

Finally, in the 17th century, the English scientist Robert Boyle argued that, by definition, an element is composed of "simple Bodies, not made of any other Bodies, of which all mixed Bodies are compounded, and into which they are ultimately resolved," a description remarkably close to our idea of an element, with atoms being the "simple Bodies." The next two centuries saw rapid progress in chemistry and the development of a "billiard-ball" image of the atom. Then, an early 20th-century burst of creativity led to our current model of an atom with a complex internal structure. In this chapter, we examine the properties and composition of matter on the macroscopic and atomic scales.

## CONCEPTS & SKILLS TO REVIEW before studying this chapter

- physical and chemical change (Section 1.1)
- states of matter (Section 1.1)
- attraction and repulsion between charged particles (Section 1.1)
- meaning of a scientific model (Section 1.2)
- SI units and conversion factors (Section 1.3)
- significant figures in calculations (Section 1.5)

## 2.1 • ELEMENTS, COMPOUNDS, AND MIXTURES: AN ATOMIC OVERVIEW

Matter can be classified into three types based on its composition—elements, compounds, and mixtures. Elements and compounds are the two kinds of substances: a **substance** is matter whose composition is fixed. Mixtures are not substances because they have a variable composition.

1. *Elements.* An **element** is the simplest type of matter with unique physical and chemical properties. *It consists of only one kind of atom* and, therefore, cannot be broken down into a simpler type of matter by any physical or chemical methods. Each element has a name, such as silicon, oxygen, or copper. A sample of silicon contains only silicon atoms. The *macroscopic* properties of a piece of silicon, such as color, density, and combustibility, are different from those of a piece of copper because the *submicroscopic* properties of silicon atoms are different from those of copper atoms; that is, *each element is unique because the properties of its atoms are unique.*

In nature, most elements exist as populations of atoms, either separated or in contact with each other, depending on the physical state. Figure 2.1A depicts atoms of an element in its gaseous state. Several elements occur in molecular form: a **molecule** is an independent structure of two or more atoms bound together (Figure 2.1B). Oxygen, for example, occurs in air as *diatomic* (two-atom) molecules.

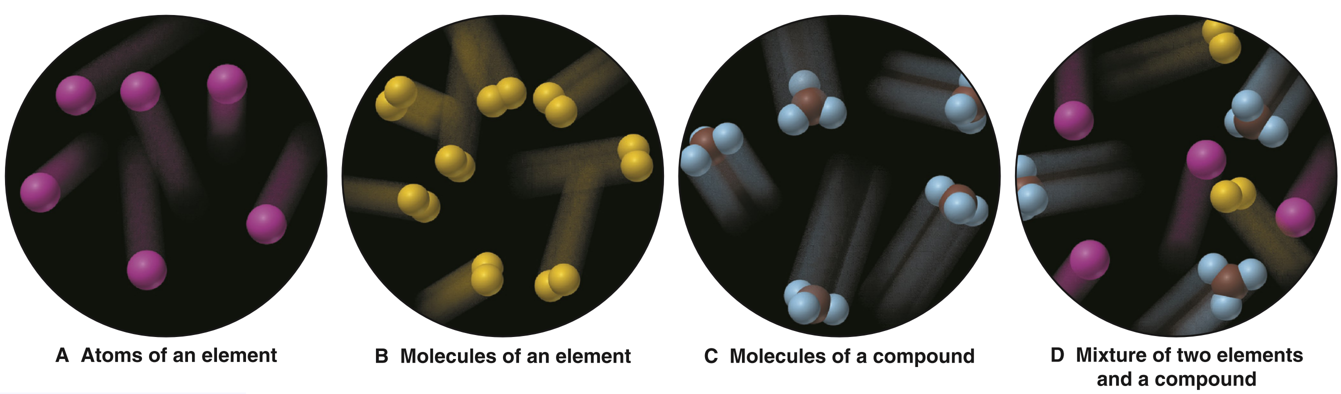


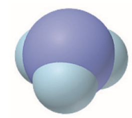

**Figure 2.1** Elements, compounds, and mixtures on the atomic scale. The samples depicted here are gases, but the three types of matter also occur as liquids and solids.

**Table 2.1 Some Properties of Sodium, Chlorine, and Sodium Chloride**

| Property | Sodium | + | Chlorine | ⟶ | Sodium Chloride | |
|---|---|---|---|---|---|---|
| Melting point | 97.8°C |  | −101°C |  | 801°C |  |
| Boiling point | 881.4°C | | −34°C | | 1413°C | |
| Color | Silvery | | Yellow-green | | Colorless (white) | |
| Density | 0.97 g/cm$^3$ | | 0.0032 g/cm$^3$ | | 2.16 g/cm$^3$ | |
| Behavior in water | Reacts | | Dissolves slightly | | Dissolves freely | |

2. *Compounds.* A **compound** consists of *two or more different elements that are bonded chemically* (Figure 2.1C). Many compounds, such as ammonia, water, and carbon dioxide, consist of molecules. But, as we'll discuss shortly, many others, like silicon dioxide and sodium sulfate, do not. No matter what the compound, however, one defining feature is that *the elements are present in fixed parts by mass* (fixed mass ratio). This is so because *each unit of the compound consists of a fixed number of atoms of each element.* For example, a sample of ammonia is 14 parts nitrogen by mass and 3 parts hydrogen by mass *because* 1 nitrogen atom has 14 times the mass of 1 hydrogen atom; thus, each ammonia molecule consists of 1 nitrogen atom and 3 hydrogen atoms:

Ammonia gas is 14 parts N by mass and 3 parts H by mass.
1 N atom has 14 times the mass of 1 H atom.
*Each ammonia molecule consists of 1 N atom and 3 H atoms.*



Another defining feature of a compound is that *its properties are different from the properties of its component elements.* Table 2.1 shows a striking example: soft, silvery sodium metal and yellow-green, poisonous chlorine gas are very different from the compound they form—white, crystalline sodium chloride, or common table salt!

Unlike an element, a compound *can* be broken down into simpler substances—its component elements. For example, an electric current breaks down molten sodium chloride into metallic sodium and chlorine gas. By definition, this breakdown is a *chemical change,* not a physical one.

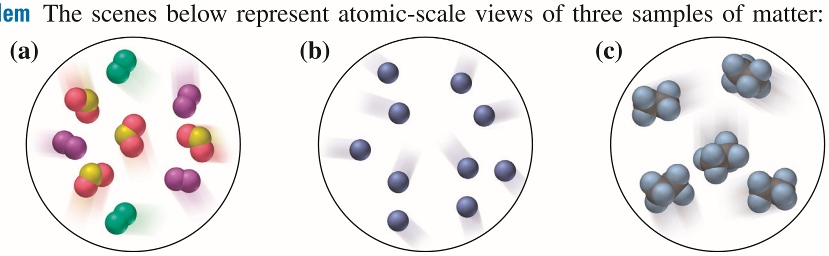
3. *Mixtures.* A **mixture** consists of two or more substances (elements and/or compounds) that are physically intermingled. Because a mixture is *not* a substance, in contrast to a compound, *the components of a mixture* ***can*** *vary in their parts by mass.* A mixture of the compounds sodium chloride and water, for example, can have many different parts by mass of salt to water. On the atomic scale, a mixture consists of the individual units that make up its component elements and/or compounds (Figure 2.1D). It makes sense, then, that *a mixture retains many of the properties of its components.* Saltwater, for instance, is colorless like water and tastes salty like sodium chloride.

Unlike compounds, mixtures can be separated into their components by *physical changes;* chemical changes are not needed. For example, the water in saltwater can be boiled off, a physical process that leaves behind solid sodium chloride. The following sample problem will help differentiate these types of matter.

## Sample Problem 2.1 Distinguishing Elements, Compounds, and Mixtures at the Atomic Scale

**Problem** The scenes below represent atomic-scale views of three samples of matter:

(a) (b) (c)



Describe each sample as an element, compound, or mixture.

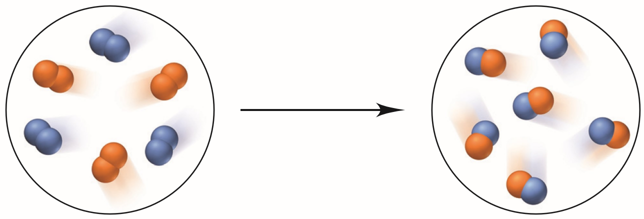
**Plan** We have to determine the type of matter by examining the component particles. If a sample contains only one type of particle, it is either an element or a compound; if it contains more than one type, it is a mixture. Particles of an element have only one kind of atom (one color), and particles of a compound have two or more kinds of atoms.

**Solution** **(a)** Mixture: there are three different types of particles. Two types contain only one kind of atom, either green or purple, so they are elements, and the third type contains two red atoms for every one yellow, so it is a compound.

**(b)** Element: the sample consists of only blue atoms.

**(c)** Compound: the sample consists of molecules that each have two black and six blue atoms.

**FOLLOW-UP PROBLEM 2.1** Describe the following reaction in terms of elements, compounds, and mixtures. (See Brief Solutions at the end of the chapter.)



## Summary of Section 2.1

- All matter exists as either elements, compounds, or mixtures.
- Every element or compound is a substance, matter with a fixed composition.
- An element consists of one type of atom and occurs as collections of individual atoms or molecules.
- A compound contains two or more elements chemically combined and exhibits different properties from its component elements. The elements occur in fixed parts by mass because each unit of the compound has a fixed number of each type of atom. Only a chemical change can break down a compound into its elements.
- A mixture consists of two or more substances mixed together, not chemically combined. The components retain their individual properties, can be present in any proportion, and can be separated by physical changes.

# 2.2 • THE OBSERVATIONS THAT LED TO AN ATOMIC VIEW OF MATTER

Any model of the composition of matter had to explain two widespread observations known as the *law of mass conservation* and the *law of definite (or constant) composition.* As you'll see, an atomic theory developed in the early 19th century explained these mass laws and another now known as the *law of multiple proportions.*

## Mass Conservation

The most fundamental chemical observation of the 18th century was the **law of mass conservation:** *the total mass of substances does not change during a chemical reaction.* The *number* of substances may change and, by definition, their properties must, but the *total amount* of matter remains constant. The great French chemist and statesman Lavoisier first stated this law on the basis of experiments in which he reacted mercury with oxygen. He found the mass of oxygen plus the mass of mercury always equaled the mass of mercuric oxide that formed.

Even in a complex biochemical change, such as the metabolism of the sugar glucose, which involves many reactions, mass is conserved:

$$\text{180 g glucose} + \text{192 g oxygen gas} \longrightarrow \text{264 g carbon dioxide} + \text{108 g water}$$
$$\text{372 g material before} \longrightarrow \text{372 g material after}$$

Mass conservation means that, based on all chemical experience, *matter cannot be created or destroyed.* (As you'll see later, mass *does* change in nuclear reactions but *not* in chemical reactions.)

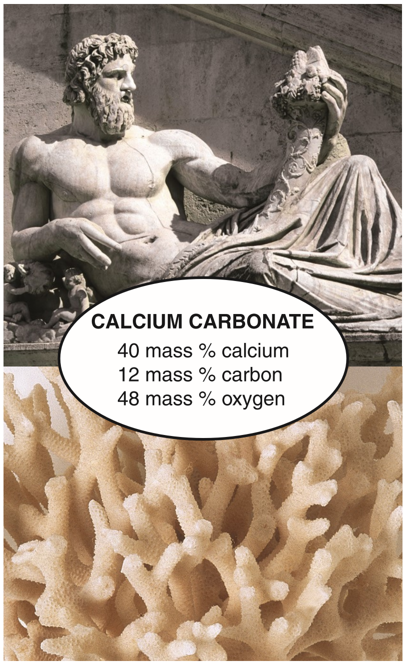



**Figure 2.2** The law of definite composition. Calcium carbonate occurs in many forms (such as marble, *top,* and coral, *bottom*), but the mass percents of its elements are always the same.

## Definite Composition

The sodium chloride in your salt shaker is the same substance whether it comes from a salt mine, a salt flat, or any other source. This fact is expressed in the **law of definite (or constant) composition,** which states that *no matter what its source, a particular compound is composed of the same elements in the same parts (fractions) by mass.* The **fraction by mass (mass fraction)** is the part of the compound's mass that each element contributes. It is obtained by dividing the mass of each element by the mass of the compound. The **percent by mass (mass percent, mass %)** is the fraction by mass expressed as a percentage (multiplied by 100).

Consider calcium carbonate, the major compound in seashells, marble, and coral. It is composed of three elements—calcium, carbon, and oxygen. The following results are obtained from a mass analysis of 20.0 g of calcium carbonate:

| Analysis by Mass (grams/20.0 g) | Mass Fraction (parts/1.00 part) | Percent by Mass (parts/100 parts) |
|---|---|---|
| 8.0 g calcium | 0.40 calcium | 40% calcium |
| 2.4 g carbon | 0.12 carbon | 12% carbon |
| 9.6 g oxygen | 0.48 oxygen | 48% oxygen |
| 20.0 g | 1.00 part by mass | 100% by mass |

The mass of each element depends on the mass of the sample—that is, more than 20.0 g of compound would contain more than 8.0 g of calcium—but *the mass fraction is fixed no matter what the size of the sample.* The sum of the mass fractions (or mass percents) equals 1.00 part (or 100%) by mass. The law of definite composition tells us that pure samples of calcium carbonate, no matter where they come from, always contain 40% calcium, 12% carbon, and 48% oxygen by mass (Figure 2.2).

Because a given element always constitutes the same mass fraction of a given compound, we can use that mass fraction to find the actual mass of the element in any sample of the compound:

$$\text{Mass of element} = \text{mass of compound} \times \frac{\text{part by mass of element}}{\text{one part by mass of compound}}$$

Or, more simply, we can skip the need to find the mass fraction first and use the results of mass analysis directly:

$$\text{Mass of element in sample} = \text{mass of compound in sample} \times \frac{\text{mass of element in compound}}{\text{mass of compound}} \qquad \textbf{(2.1)}$$

### Sample Problem 2.2 Calculating the Mass of an Element in a Compound

**Problem** Pitchblende is the most important compound of uranium. Mass analysis of an 84.2-g sample shows that it contains 71.4 g of uranium, with oxygen the only other element. How many grams of uranium are in 102 kg of pitchblende?

**Plan** We have to find the mass of uranium in a known mass (102 kg) of pitchblende, given the mass of uranium (71.4 g) in a different mass of pitchblende (84.2 g). The mass ratio of uranium to pitchblende is the same for any sample of pitchblende. Therefore, using Equation 2.1, we multiply the mass (in kg) of the pitchblende sample by the ratio of uranium to pitchblende from the mass analysis. This gives the mass (in kg) of uranium, and we convert kilograms to grams.

**Road Map**

Mass (kg) of pitchblende

↓ multiply by mass ratio of uranium to pitchblende from analysis

Mass (kg) of uranium

↓ 1 kg = 1000 g

Mass (g) of uranium

**Solution** Finding the mass (kg) of uranium in 102 kg of pitchblende:

$$\text{Mass (kg) of uranium} = \text{mass (kg) of pitchblende} \times \frac{\text{mass (kg) of uranium in pitchblende}}{\text{mass (kg) of pitchblende}}$$

$$= 102\ \cancel{\text{kg pitchblende}} \times \frac{71.4\ \text{kg uranium}}{84.2\ \cancel{\text{kg pitchblende}}} = 86.5\ \text{kg uranium}$$

Converting the mass of uranium from kg to g:

$$\text{Mass (g) of uranium} = 86.5\ \cancel{\text{kg}}\ \text{uranium} \times \frac{1000\ \text{g}}{1\ \cancel{\text{kg}}} = \boxed{8.65\times10^{4}\ \text{g uranium}}$$

**Check** The analysis showed that most of the mass of pitchblende is due to uranium, so the large mass of uranium makes sense. Rounding off to check the math gives:

$$\sim 100 \text{ kg pitchblende} \times \frac{70}{85} = 82 \text{ kg uranium}$$

**FOLLOW-UP PROBLEM 2.2** How many metric tons (t) of oxygen are combined in a sample of pitchblende that contains 2.3 t of uranium? (*Hint:* Remember that oxygen is the only other element present.) See Brief Solutions.

## Multiple Proportions

Dalton and others made an observation that applies when two elements form more than one compound, now called the **law of multiple proportions:** *if elements A and B react to form two compounds, the different masses of B that combine with a fixed mass of A can be expressed as a ratio of small whole numbers.* Consider two compounds, let's call them I and II, that carbon and oxygen form. These compounds have very different properties: the density of carbon oxide I is 1.25 g/L, whereas that of II is 1.98 g/L; I is poisonous and flammable, but II is not. Mass analysis shows that

Carbon oxide I is 57.1 mass % oxygen and 42.9 mass % carbon
Carbon oxide II is 72.7 mass % oxygen and 27.3 mass % carbon

To see the phenomenon of multiple proportions, we use the mass percents of oxygen and of carbon to find their masses in a given mass, say 100 g, of each compound. Then we divide the mass of oxygen by the mass of carbon in each compound to obtain the mass of oxygen that combines with a fixed mass of carbon:

| | Carbon Oxide I | Carbon Oxide II |
|---|---|---|
| g oxygen/100 g compound | 57.1 | 72.7 |
| g carbon/100 g compound | 42.9 | 27.3 |
| g oxygen/g carbon | $\frac{57.1}{42.9} = 1.33$ | $\frac{72.7}{27.3} = 2.66$ |

If we then divide the grams of oxygen per gram of carbon in II by that in I, we obtain a ratio of small whole numbers:

$$\frac{2.66 \text{ g oxygen/g carbon in II}}{1.33 \text{ g oxygen/g carbon in I}} = \frac{2}{1}$$

The law of multiple proportions tells us that in two compounds of the same elements, the mass fraction of one element relative to the other element changes in *increments based on ratios of small whole numbers.* In this case, the ratio is 2/1—for a given mass of carbon, compound II contains *2 times* as much oxygen as I, not 1.583 times, 1.716 times, or any other intermediate amount. In the next section, we'll explain the mass laws on the atomic scale.

## Summary of Section 2.2

- The law of mass conservation states that the total mass remains constant during a chemical reaction.
- The law of definite composition states that any sample of a given compound has the same elements present in the same parts by mass.
- The law of multiple proportions states that, in different compounds of the same elements, the masses of one element that combine with a fixed mass of the other can be expressed as a ratio of small whole numbers.

# 2.3 • DALTON'S ATOMIC THEORY

With over 200 years of hindsight, it's easy to see how the mass laws could be explained by an atomic model—matter existing in indestructible units, each with a particular mass—but it was a major breakthrough in 1808 when John Dalton (1766–1844) presented his atomic theory of matter in *A New System of Chemical Philosophy.*

## Postulates of the Atomic Theory

Dalton expressed his theory in a series of postulates. Like most great thinkers, he integrated the ideas of others into his own. As we go through the postulates, presented here in modern terms, we'll note which were original and which came from others.

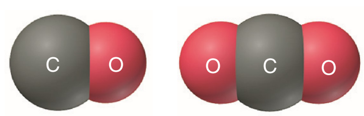
1. All matter consists of **atoms,** tiny indivisible particles of an element that cannot be created or destroyed. (This derives from the "eternal, indestructible atoms" proposed by Democritus more than 2000 years earlier and reflects mass conservation as stated by Lavoisier.)
2. Atoms of one element *cannot* be converted into atoms of another element. In chemical reactions, the atoms of the original substances recombine to form different substances. (This rejects the belief in the magical transmutation of elements that was widely held into the 17$^{th}$ century.)
3. Atoms of an element are identical in mass and other properties and are different from atoms of any other element. (This contains Dalton's major new ideas: *unique mass and properties* for the atoms of a given element.)
4. Compounds result from the chemical combination of a specific ratio of atoms of different elements. (This follows directly from the law of definite composition.)

## How the Theory Explains the Mass Laws

Let's see how Dalton's postulates explain the mass laws:

- *Mass conservation.* Atoms cannot be created or destroyed (postulate 1) or converted into other types of atoms (postulate 2). Therefore, a chemical reaction, in which atoms are combined differently, cannot possibly result in a mass change.
- *Definite composition.* A compound is a combination of a *specific* ratio of different atoms (postulate 4), each of which has a particular mass (postulate 3). Thus, each element in a compound constitutes a fixed fraction of the total mass.
- *Multiple proportions.* Atoms of an element have the same mass (postulate 3) and are indivisible (postulate 1). The masses of element B that combine with a fixed mass of element A give a small, whole-number ratio because different numbers of B atoms combine with each A atom in different compounds.

The *simplest* arrangement consistent with the mass data for carbon oxides I and II in our earlier example is that one atom of oxygen combines with one atom of carbon in compound I (carbon monoxide) and that two atoms of oxygen combine with one atom of carbon in compound II (carbon dioxide):

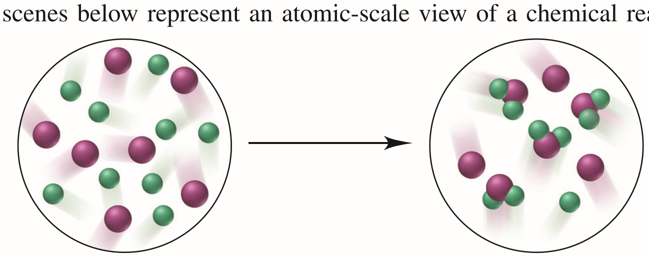



Carbon oxide I (carbon monoxide)    Carbon oxide II (carbon dioxide)

Let's work through a sample problem that reviews the mass laws.

### Sample Problem 2.3 Visualizing the Mass Laws

**Problem** The scenes below represent an atomic-scale view of a chemical reaction:

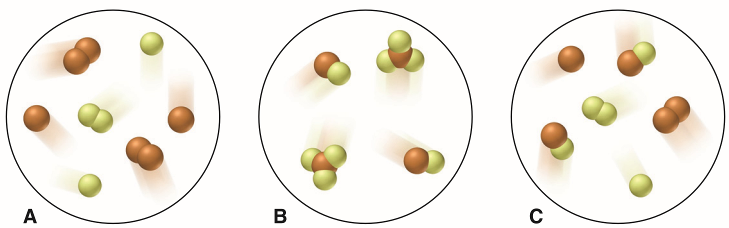


Which of the mass laws—mass conservation, definite composition, or multiple proportions—is (are) illustrated?

**Plan** From the depictions, we note the numbers, colors, and combinations of atoms (spheres) to see which mass laws pertain. If the numbers of each atom are the same before and after the reaction, the total mass did not change (mass conservation). If a compound forms that always has the same atom ratio, the elements are present in fixed parts by mass (definite composition). When the same elements form different compounds and the ratio of the atoms of one element that combine with one atom of the other element is a small whole number, the ratio of their masses is a small whole number as well (multiple proportions).

**Solution** There are seven purple and nine green atoms in each circle, so mass is conserved. The compound formed has one purple and two green atoms, so it has definite composition. Only one compound forms, so the law of multiple proportions does not pertain.

**FOLLOW-UP PROBLEM 2.3** Which sample(s) best display(s) the fact that compounds of bromine (orange) and fluorine (yellow) exhibit the law of multiple proportions? Explain.




## Summary of Section 2.3

- Dalton's atomic theory explained the mass laws by proposing that all matter consists of indivisible, unchangeable atoms of fixed, unique mass.
- Mass is conserved during a reaction because the atoms retain their identities but are combined differently.
- Each compound has a fixed mass fraction of each of its elements because it is composed of a fixed number of each type of atom.
- Different compounds of the same elements exhibit multiple proportions because they each consist of whole atoms.

# 2.4 • THE OBSERVATIONS THAT LED TO THE NUCLEAR ATOM MODEL

Dalton's model established that masses of reacting elements could be explained in terms of atoms but not why atoms bond as they do: why, for example, do two, and not three, hydrogen atoms bond with one oxygen atom in a water molecule?

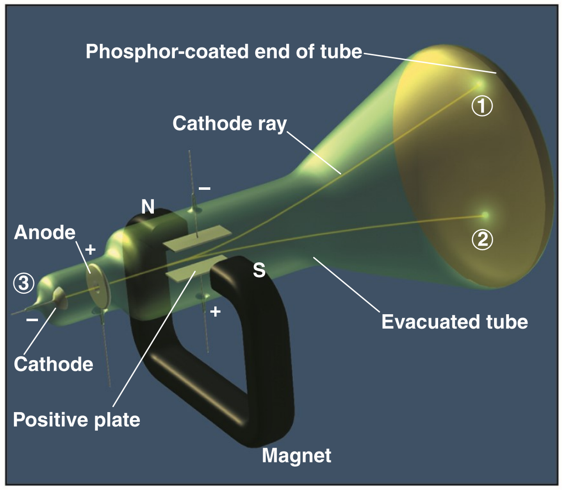
Moreover, Dalton's model of the atom, which represented it as a tiny, indivisible particle, like a minute billiard ball, did not predict the existence of subatomic charged particles. These were observed in later experiments that led to the discovery of *electrons* and the atomic *nucleus.* Let's examine some of these experiments and the more complex atomic model that emerged from them.

## Discovery of the Electron and Its Properties

For many years, scientists had known that matter and electric charge were related. When amber is rubbed with fur, or glass with silk, positive and negative charges form—the same charges that make your hair crackle and cling to your comb on a dry day. They also knew that an electric current could decompose certain compounds into their elements. But they did not know what a current was made of.

**Cathode Rays** To discover the nature of an electric current, some investigators tried passing current through nearly evacuated glass tubes fitted with metal electrodes. When the electric power source was turned on, a "ray" could be seen striking the

**Figure 2.3** Observations that established the properties of cathode rays.




| OBSERVATION | CONCLUSION |
|---|---|
| 1. Ray bends in magnetic field. | Consists of charged particles |
| 2. Ray bends toward positive plate in electric field. | Consists of negative particles |
| 3. Ray is identical for any cathode. | Particles found in all matter |

phosphor-coated end of the tube and emitting a glowing spot of light. The rays were called **cathode rays** because they originated at the negative electrode (cathode) and moved to the positive electrode (anode).

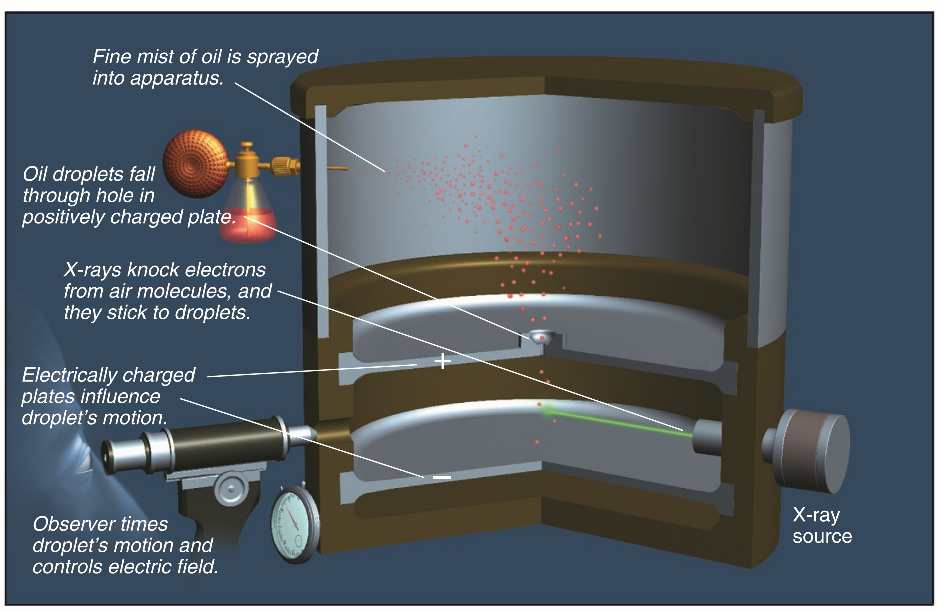
Figure 2.3 shows some properties of cathode rays based on these observations. The main conclusion was that *cathode rays consist of negatively charged particles found in all matter.* The rays appear when these particles collide with the few remaining gas molecules in the evacuated tube. Cathode ray particles were later named *electrons.*

**Mass and Charge of the Electron** Two classic experiments revealed the mass and charge of the electron:

1. *Mass/charge ratio*. In 1897, the British physicist J. J. Thomson (1856–1940) measured the ratio of the mass of a cathode ray particle to its charge. By comparing this value with the mass/charge ratio for the lightest charged particle in solution, Thomson estimated that the cathode ray particle weighed less than $\frac{1}{1000}$ as much as hydrogen, the lightest atom! He was shocked because this implied that, contrary to Dalton's atomic theory, *atoms contain even smaller particles.* Fellow scientists reacted with disbelief to Thomson's conclusion, thinking he was joking.

2. *Charge.* In 1909, the American physicist Robert Millikan (1868–1953) measured the *charge* of the electron. He did so by observing the movement of oil droplets in an apparatus that contained electrically charged plates and an x-ray source (Figure 2.4).

**Figure 2.4** Millikan's oil-drop experiment for measuring an electron's charge. The total charge on an oil droplet is some whole-number multiple of the charge of the electron.



X-rays knocked electrons from gas molecules in the air within the apparatus, and the electrons stuck to an oil droplet falling through a hole in a positively charged plate. With the electric field off, Millikan measured the mass of the droplet from its rate of fall. Then, by adjusting the field's strength, he made the droplet hang suspended in the air and, thus, measured its total charge.

After many tries, Millikan found that the total charge of the various droplets was always some *whole-number multiple of a minimum charge.* If different oil droplets picked up different numbers of electrons, he reasoned that this minimum charge must be the charge of the electron itself. Remarkably, the value that he calculated over a century ago is within 1% of the modern value of the electron's charge, $-1.602\times10^{-19}$ C (C stands for *coulomb,* the SI unit of charge).

3. *Conclusion: calculating the electron's mass.* The electron's mass/charge ratio and the value for the electron's charge can be used to find the electron's mass, which is *extremely* small:

$$\text{Mass of electron} = \frac{\text{mass}}{\cancel{\text{charge}}} \times \cancel{\text{charge}} = \left(-5.686\times10^{-12}\ \frac{\text{kg}}{\cancel{\text{C}}}\right)(-1.602\times10^{-19}\ \cancel{\text{C}})$$
$$= 9.109\times10^{-31}\ \text{kg} = 9.109\times10^{-28}\ \text{g}$$

## Discovery of the Atomic Nucleus

The presence of electrons in all matter posed some major questions about the structure of atoms. Matter is electrically neutral, so atoms must be also. But if atoms contain negatively charged electrons, what positive charges balance them? And if an electron has such a tiny mass, what accounts for an atom's much larger mass? To address these issues, Thomson proposed his "plum-pudding" model—a spherical atom composed of diffuse, positively charged matter with electrons embedded like "raisins in a plum pudding."

In 1910, New Zealand–born physicist Ernest Rutherford (1871–1937) tested this model and obtained an unexpected result (Figure 2.5):

1. *Experimental design.* Figure 2.5B shows the experimental setup, in which tiny, dense, positively charged alpha (α) particles emitted from radium are aimed at gold foil. A circular, zinc-sulfide screen registers the deflection (scattering) of the α particles by emitting light flashes when the particles strike it.

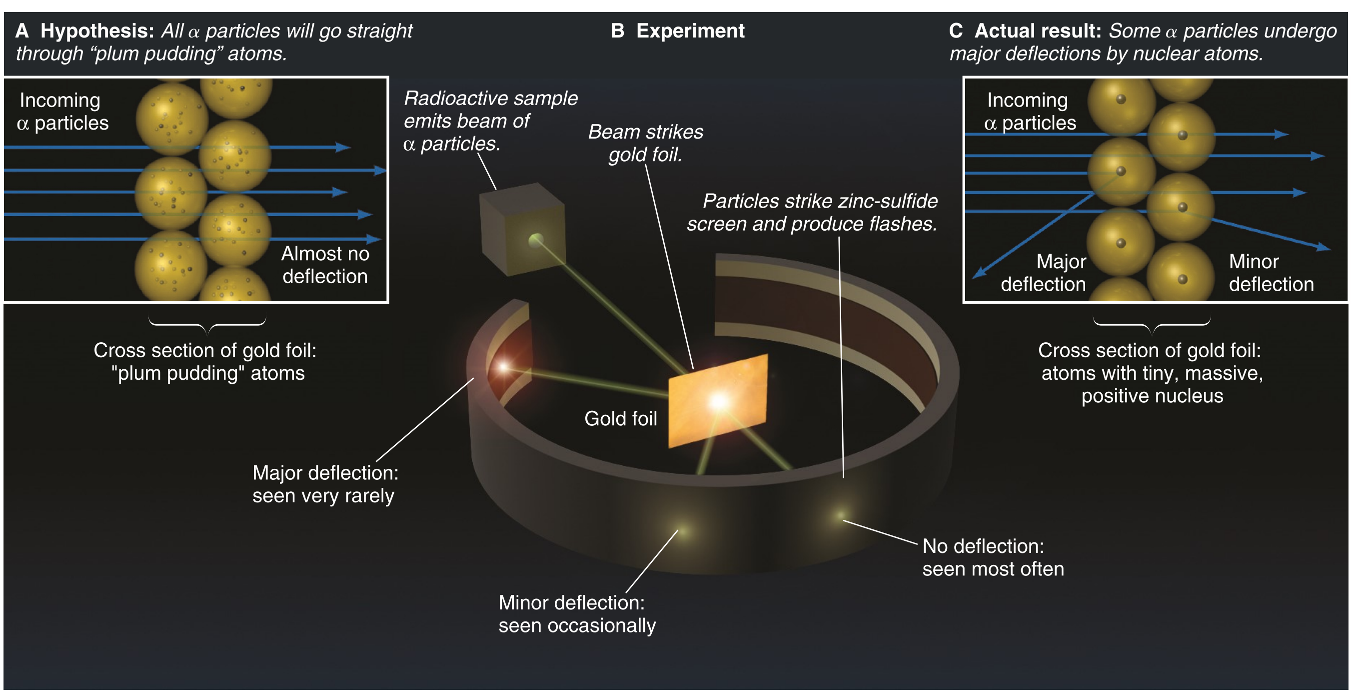



**Figure 2.5** Rutherford's α-scattering experiment and discovery of the atomic nucleus.

2. *Expected results.* With Thomson's model in mind (Figure 2.5A), Rutherford expected only minor, if any, deflections of the α particles because they should act as bullets and go right through the gold atoms. After all, an electron would not deflect an α particle any more than a Ping-Pong ball would deflect a baseball.
3. *Actual results.* Initial results were consistent with this idea, but then the unexpected happened (Figure 2.5C). As Rutherford recalled: "I remember two or three days later Geiger [one of his coworkers] coming to me in great excitement and saying, 'We have been able to get some of the α particles coming backwards . . .' It was quite the most incredible event that has ever happened to me in my life. It was almost as incredible as if you fired a 15-inch shell at a piece of tissue paper and it came back and hit you." In fact, very few α particles were deflected at all, and only 1 in 20,000 had large-angle deflections of more than 90° ("coming backwards").
4. *Conclusion.* Rutherford concluded that these few α particles were being repelled by something small, dense, and positive within the gold atoms. Calculations based on the properties of α particles and the fraction of large-angle deflections showed that
   - An atom is mostly space occupied by electrons.
   - In the center is a tiny region, which Rutherford called the **nucleus,** that contains all the positive charge and essentially all the mass of the atom.

   He proposed that positive particles lay within the nucleus and called them *protons.*

Rutherford's model explained the charged nature of matter, but it could not account for all the atom's mass. After more than 20 years, in 1932, James Chadwick (1891–1974) discovered the *neutron,* an uncharged dense particle that also resides in the nucleus.

## Summary of Section 2.4

- Several major discoveries at the turn of the 20th century resolved questions about Dalton's model and led to our current model of atomic structure.
- Cathode rays were shown to consist of negative particles (electrons) that exist in all matter. J. J. Thomson measured their mass/charge ratio and concluded that they are much smaller and lighter than atoms.
- Robert Millikan determined the charge of the electron, which he combined with other data to calculate its mass.
- Ernest Rutherford proposed that atoms consist of a tiny, massive, positive nucleus surrounded by electrons.

# 2.5 • THE ATOMIC THEORY TODAY

Dalton's model of an indivisible particle has given way to our current model of an atom with an elaborate internal architecture of subatomic particles.

## Structure of the Atom

An *atom* is an electrically neutral, spherical entity composed of a positively charged central nucleus surrounded by one or more negatively charged electrons (Figure 2.6). The electrons move rapidly within the available volume, held there by the attraction of the nucleus. An atom's diameter ($\sim 1\times10^{-10}$ m) is about 20,000 times the diameter of its nucleus ($\sim 5\times10^{-15}$ m). The nucleus contributes 99.97% of the atom's mass, occupies only about 1 quadrillionth of its volume, and is incredibly dense: about $10^{14}$ g/mL!

### THINK OF IT THIS WAY
**The Tiny, Massive Nucleus**

A few analogies will help you appreciate the incredible properties of the atomic nucleus. A nucleus the size of the period at the end of this sentence would weigh about 100 tons, as much as 50 cars. An atom the size of the Houston Astrodome would have a nucleus the size of a green pea that would contain virtually all the atom's mass. If a nucleus were actually the size shown in Figure 2.6 (about 1 cm across), the atom would be about 200 m across, or slightly more than twice the length of a football field!

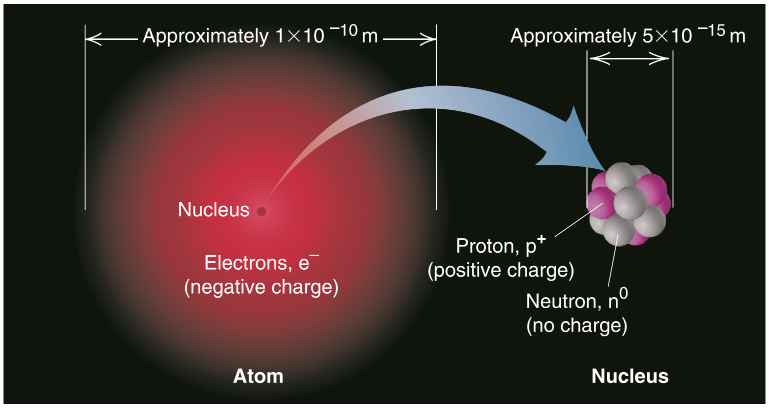



**Figure 2.6** General features of the atom.

*An atomic nucleus consists of protons and neutrons* (the only exception is the simplest hydrogen nucleus, which is a single proton). The **proton ($p^+$)** has a positive charge, and the **neutron ($n^0$)** has no charge; thus, the positive charge of the nucleus results from its protons. The *magnitudes* of the charges possessed by a proton and by an **electron ($e^-$)** are equal, but the *signs* of the charges are opposite. *An atom is neutral because the number of protons in the nucleus equals the number of electrons surrounding the nucleus.* Some properties of these three subatomic particles are listed in Table 2.2.

## Atomic Number, Mass Number, and Atomic Symbol

The **atomic number ($Z$)** of an element equals the number of protons in the nucleus of each of its atoms. *All atoms of an element have the same atomic number, and the atomic number of each element is different from that of any other element.* All carbon atoms ($Z = 6$) have 6 protons, all oxygen atoms ($Z = 8$) have 8 protons, and all uranium atoms ($Z = 92$) have 92 protons. There are currently 117 known elements, of which 90 occur in nature and 27 have been synthesized by nuclear scientists.

The **mass number ($A$)** is the total number of protons and neutrons in the nucleus of an atom. Each proton and each neutron contributes one unit to the mass number. Thus, a carbon atom with 6 protons and 6 neutrons in its nucleus has a mass number of 12, and a uranium atom with 92 protons and 146 neutrons in its nucleus has a mass number of 238.

The **atomic symbol** (or *element symbol*) of an element is based on its English, Latin, or Greek name, such as C for carbon, S for sulfur, and Na for sodium (Latin *natrium*). Often written with the symbol are the atomic number ($Z$) as a left *sub*script and the mass number ($A$) as a left *super*script, so element X would be ${}^{A}_{Z}X$. Since the

**Table 2.2** Properties of the Three Key Subatomic Properties

| | Charge | | Mass | | |
|---|---|---|---|---|---|
| **Name (Symbol)** | **Relative** | **Absolute (C)*** | **Relative (amu)†** | **Absolute (g)** | **Location in Atom** |
| Proton ($p^+$) | 1+ | $+1.60218\times10^{-19}$ | 1.00727 | $1.67262\times10^{-24}$ | Nucleus |
| Neutron ($n^0$) | 0 | 0 | 1.00866 | $1.67493\times10^{-24}$ | Nucleus |
| Electron ($e^-$) | 1− | $-1.60218\times10^{-19}$ | 0.00054858 | $9.10939\times10^{-24}$ | Outside nucleus |

*The coulomb (C) is the SI unit of charge.
†The atomic mass unit (amu) equals $1.66054\times10^{-24}$ g; discussed later in this section.

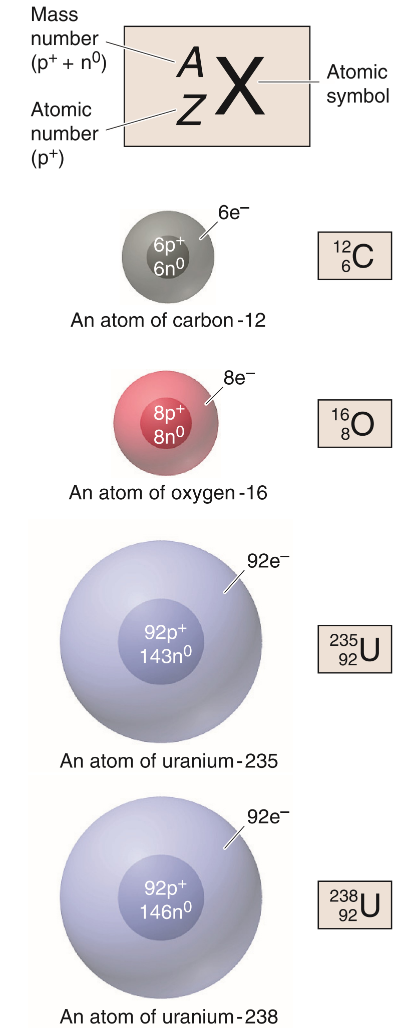



**Figure 2.7** The $^{A}_{Z}X$ notations and spherical representations for four atoms. (The nuclei are not drawn to scale.)

mass number is the sum of protons and neutrons, the number of neutrons ($N$) equals the mass number minus the atomic number:

$$\text{Number of neutrons} = \text{mass number} - \text{atomic number}, \quad \text{or} \quad N = A - Z \qquad \textbf{(2.2)}$$

Thus, a chlorine atom, symbolized $^{35}_{17}Cl$, has $A = 35$, $Z = 17$, and $N = 35 - 17 = 18$. Because each element has its own atomic number, we also know the atomic number given the symbol. For example, instead of writing $^{12}_{6}C$ for carbon with mass number 12, we can write $^{12}C$ (spoken "carbon twelve"), with $Z = 6$ understood. Another way to name this atom is carbon-12.

## Isotopes

*All atoms of an element have the same atomic number but not the same mass number.* **Isotopes** of an element are atoms that have *different numbers of neutrons* and therefore different mass numbers. For example, all carbon atoms ($Z = 6$) have 6 protons and 6 electrons, but only 98.89% of naturally occurring carbon atoms have 6 neutrons ($A = 12$). A small percentage (1.11%) have 7 neutrons ($A = 13$), and even fewer (less than 0.01%) have 8 ($A = 14$). These are carbon's three naturally occurring isotopes—$^{12}C$, $^{13}C$, and $^{14}C$. *A natural sample of carbon has these three isotopes in these relative proportions.* Five other carbon isotopes—$^{9}C$, $^{10}C$, $^{11}C$, $^{15}C$, and $^{16}C$—have been created in the laboratory. Figure 2.7 depicts the atomic number, mass number, and symbol for four atoms, two of which are isotopes of the element uranium.

The chemical properties of an element are primarily determined by the number of electrons, so *all isotopes of an element have nearly identical chemical behavior,* even though they have different masses.

### Sample Problem 2.4 Determining the Number of Subatomic Particles in the Isotopes of an Element

**Problem** Silicon (Si) is a major component of semiconductor chips. It has three naturally occurring isotopes: $^{28}Si$, $^{29}Si$, and $^{30}Si$. Determine the numbers of protons, neutrons, and electrons in each silicon isotope.

**Plan** The mass number ($A$; left superscript) of each of the three isotopes is given, which is the sum of protons and neutrons. From the elements list on this book's inside front cover, we find the atomic number ($Z$, number of protons), which equals the number of electrons. We obtain the number of neutrons by subtracting $Z$ from $A$ (Equation 2.2).

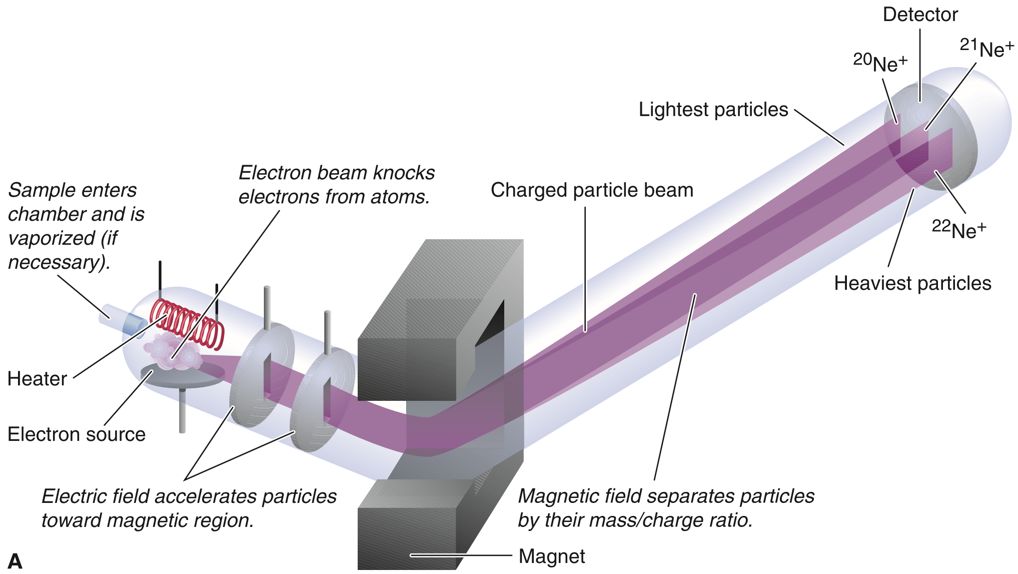
**Solution** From the elements list, the atomic number of silicon is 14. Therefore,

$^{28}Si$ has $14p^+$, $14e^-$, and $14n^0$ $(28 - 14)$
$^{29}Si$ has $14p^+$, $14e^-$, and $15n^0$ $(29 - 14)$
$^{30}Si$ has $14p^+$, $14e^-$, and $16n^0$ $(30 - 14)$

**FOLLOW-UP PROBLEM 2.4** How many protons, neutrons, and electrons are in **(a)** $^{11}_{5}Q$? **(b)** $^{41}_{20}R$? **(c)** $^{131}_{53}X$? What elements do Q, R, and X represent?

## Atomic Masses of the Elements; Mass Spectrometry

The mass of an atom is measured *relative* to the mass of an atomic standard. The modern standard is the carbon-12 atom, whose mass is defined as *exactly* 12 atomic mass units. Thus, the **atomic mass unit (amu)** is $\frac{1}{12}$ the mass of a carbon-12 atom. Based on this standard, the $^{1}H$ atom has a mass of 1.008 amu; in other words, a $^{12}C$ atom has almost 12 times the mass of an $^{1}H$ atom. We will continue to use the term *atomic mass unit* in the text, although the name of the unit has been changed to the **dalton (Da);** thus, one $^{12}C$ atom has a mass of 12 daltons (12 Da, or 12 amu). The atomic mass unit is a unit of relative mass, but it has an absolute mass of $1.66054 \times 10^{-24}$ g.



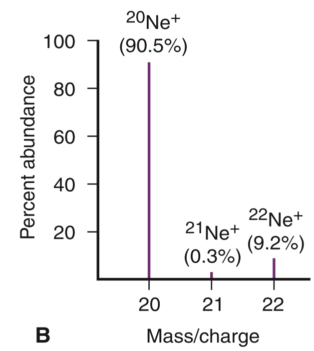




**Figure 2.8** The mass spectrometer and its data. **A,** Charged particles are separated on the basis of their *m/e* values. Ne is the sample here. **B,** The percent abundance of each Ne isotope.

The isotopic makeup of an element is determined by **mass spectrometry,** a method for measuring the relative masses and abundances of atomic-scale particles and molecules very precisely. In one type of mass spectrometer, atoms of a sample of, say, elemental neon are bombarded by a high-energy electron beam (Figure 2.8A). As a result, one electron is knocked off each Ne atom, and each resulting particle has one positive charge. Thus, its mass/charge ratio (*m/e*) equals the mass of an Ne atom divided by 1+. The *m/e* values are measured to identify the masses of different isotopes of the element. The positively charged Ne particles are attracted toward a series of negatively charged plates with slits in them, and some of the particles pass through the slits into an evacuated tube exposed to a magnetic field. As the particles zoom through this region, they are deflected (their paths are bent) according to their *m/e* values: the lightest particles are deflected most and the heaviest particles least. At the end of the tube, the particles strike a detector, which records their relative positions and abundances (Figure 2.8B). Mass spectrometry is now used to measure the mass of virtually any atom or molecule. In 2002, the Nobel Prize in chemistry was awarded for the study of proteins by mass spectrometry.

Let's see how this instrument is used. With a mass spectrometer, we measure the mass ratio of $^{28}Si$ to $^{12}C$ as

$$\frac{\text{Mass of } ^{28}\text{Si atom}}{\text{Mass of } ^{12}\text{C standard}} = 2.331411$$

From this mass ratio, we find the **isotopic mass** of the $^{28}Si$ atom, the relative mass of this silicon isotope:

$$\begin{aligned}\text{Isotopic mass of } ^{28}\text{Si} &= \text{measured mass ratio} \times \text{mass of } ^{12}\text{C} \\ &= 2.331411 \times 12 \text{ amu} = 27.97693 \text{ amu}\end{aligned}$$

Along with the isotopic mass, the mass spectrometer gives the relative abundance as a percentage (or fraction) of each isotope in a sample of the element. For example, the relative abundance of $^{28}Si$ is 92.23% (or 0.9223).

From such data, we can obtain the **atomic mass** (also called *atomic weight*) of an element, the *average* of the masses of its naturally occurring isotopes weighted according to their abundances.* Each naturally occurring isotope of an element contributes a certain portion to the atomic mass. For instance, multiplying the isotopic

*Based on mass spectrometric analysis of isotopic abundances from various sources, the International Union of Pure and Applied Chemistry has recommended that the atomic masses of 10 elements be given as *atomic mass intervals* to indicate the range from source to source; for example, boron has its lowest atomic mass (10.806) in metamorphic rock and its highest (10.821) in surface and ground water. These changes will not affect discussions throughout the text, which use the four-digit values shown in the periodic table (Figure 2.9) and the front end sheets.

mass of $^{28}Si$ by its fractional abundance gives the portion of the atomic mass of Si contributed by $^{28}Si$:

$$\text{Portion of Si atomic mass from } ^{28}\text{Si} = 27.97693 \text{ amu} \times 0.9223 = 25.8031 \text{ amu}$$
(retaining two additional significant figures)

Similar calculations give the portions contributed by $^{29}Si$ (28.976495 amu $\times$ 0.0467 = 1.3532 amu) and by $^{30}Si$ (29.973770 amu $\times$ 0.0310 = 0.9292 amu). Adding the three portions together (rounding to two decimal places at the end) gives the atomic mass of silicon:

$$\begin{aligned}\text{Atomic mass of Si} &= 25.8031 \text{ amu} + 1.3532 \text{ amu} + 0.9292 \text{ amu}\\ &= 28.0855 \text{ amu} = 28.09 \text{ amu}\end{aligned}$$

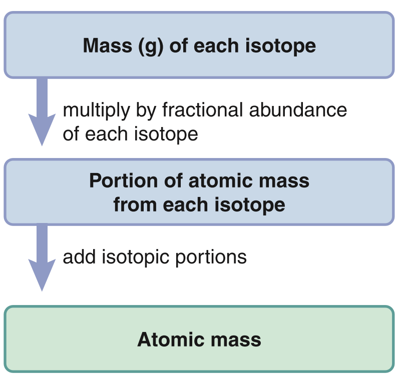
The atomic mass is an average value; thus, while no individual silicon atom has a mass of 28.09 amu, in the laboratory, we consider a sample of silicon to consist of atoms with this average mass.

## Sample Problem 2.5 Calculating the Atomic Mass of an Element

**Problem** Silver (Ag; $Z$ = 47) has 46 known isotopes, but only two occur naturally, $^{107}Ag$ and $^{109}Ag$. Given the following data, calculate the atomic mass of Ag:

| Isotope | Mass (amu) | Abundance (%) |
|---|---|---|
| $^{107}Ag$ | 106.90509 | 51.84 |
| $^{109}Ag$ | 108.90476 | 48.16 |

**Road Map**




**Plan** From the mass and abundance of the two Ag isotopes, we have to find the atomic mass of Ag (weighted average of the isotopic masses). We divide each percent abundance by 100 to get the fractional abundance and then multiply that by each isotopic mass to find the portion of the atomic mass contributed by each isotope. The sum of the isotopic portions is the atomic mass.

**Solution** Finding the fractional abundances:

$$\text{Fractional abundance of } ^{107}\text{Ag} = 51.84/100 = 0.5184; \text{ similarly, } ^{109}\text{Ag} = 0.4816$$

Finding the portion of the atomic mass from each isotope:

$$\begin{aligned}\text{Portion of atomic mass from } ^{107}\text{Ag} &= \text{isotopic mass} \times \text{fractional abundance}\\ &= 106.90509 \text{ amu} \times 0.5184 = 55.42 \text{ amu}\\ \text{Portion of atomic mass from } ^{109}\text{Ag} &= 108.90476 \text{ amu} \times 0.4816 = 52.45 \text{ amu}\end{aligned}$$

Finding the atomic mass of silver:

$$\text{Atomic mass of Ag} = 55.42 \text{ amu} + 52.45 \text{ amu} = \boxed{107.87 \text{ amu}}$$

**Check** The individual portions seem right: ~100 amu $\times$ 0.50 = 50 amu. The portions should be almost the same because the two isotopic abundances are almost the same. We rounded each portion to four significant figures because that is the number of significant figures in the abundance values. This is the correct atomic mass (to two decimal places); in the list of elements (inside front cover), it is rounded to 107.9 amu.

**FOLLOW-UP PROBLEM 2.5** Boron (B; $Z$ = 5) has two naturally occurring isotopes. Find the percent abundances of $^{10}B$ and $^{11}B$ given these data: atomic mass of B = 10.81 amu, isotopic mass of $^{10}B$ = 10.0129 amu, and isotopic mass of $^{11}B$ = 11.0093 amu. (*Hint:* The sum of the fractional abundances is 1. If $x$ = abundance of $^{10}B$, then $1 - x$ = abundance of $^{11}B$.)

## Summary of Section 2.5

- An atom has a central nucleus, which contains positively charged protons and uncharged neutrons and is surrounded by negatively charged electrons. An atom is neutral because the number of electrons equals the number of protons.

- An atom is represented by the notation $^{A}_{Z}X$, in which $Z$ is the atomic number (number of protons), $A$ the mass number (sum of protons and neutrons), and X the atomic symbol.
- An element occurs naturally as a mixture of isotopes, atoms with the same number of protons but different numbers of neutrons. Each isotope has a mass relative to the $^{12}C$ mass standard.
- The atomic mass of an element is the average of its isotopic masses weighted according to their natural abundances. It is determined using the mass spectrometer.

## 2.6 • ELEMENTS: A FIRST LOOK AT THE PERIODIC TABLE

At the end of the $18^{th}$ century, Lavoisier compiled a list of the 23 elements known at that time; by 1870, 65 were known; by 1925, 88; today, there are 117 and still counting! By the mid-$19^{th}$ century, enormous amounts of information concerning reactions, properties, and atomic masses of the elements had been accumulated. Several researchers noted recurring, or *periodic,* patterns of behavior and proposed schemes to organize the elements according to some fundamental property.

In 1871, the Russian chemist Dmitri Mendeleev (1836–1907) published the most successful of these organizing schemes as a table of the elements listed by increasing atomic mass and arranged so that elements with similar chemical properties fell in the same column. The modern **periodic table of the elements,** based on Mendeleev's version (but arranged by *atomic number,* not mass), is one of the great classifying schemes in science and an indispensable tool to chemists—and chemistry students.

**Organization of the Periodic Table** One common version of the modern periodic table appears in Figure 2.9 on the next page (and inside the front cover). It is formatted as follows:

1. Each element has a box that contains its atomic number (number of protons in the nucleus), atomic symbol, and atomic mass. (A mass in parentheses is the mass number of the most stable isotope of that element.) The boxes lie, from left to right, in order of *increasing atomic number*.
2. The boxes are arranged into a grid of **periods** (horizontal rows) and **groups** (vertical columns). Each period has a number from 1 to 7. Each group has a number from 1 to 8 *and* either the letter A or B. A newer system, with group numbers from 1 to 18 but no letters, appears in parentheses under the number-letter designations. (The text uses the number-letter system and shows the newer numbering system in parentheses.)
3. The eight A groups (two on the left and six on the right) contain the *main-group elements.* The ten B groups, located between Groups 2A(2) and 3A(13), contain the *transition elements.* Two horizontal series of *inner transition elements,* the lanthanides and the actinides, fit *between* the elements in Group 3B(3) and Group 4B(4) and are placed below the main body of the table.

**Classifying the Elements** One of the clearest ways to classify the elements is as metals, nonmetals, and metalloids. The "staircase" line that runs from the top of Group 3A(13) to the bottom of Group 6A(16) is a dividing line:

- The **metals** (three shades of blue in Figure 2.9) lie in the large lower-left portion of the table. About three-quarters of the elements are metals, including many main-group elements and all the transition and inner transition elements. They are generally shiny solids at room temperature (mercury is the only liquid) that conduct heat and electricity well. They can be tooled into sheets (are malleable) and wires (are ductile).

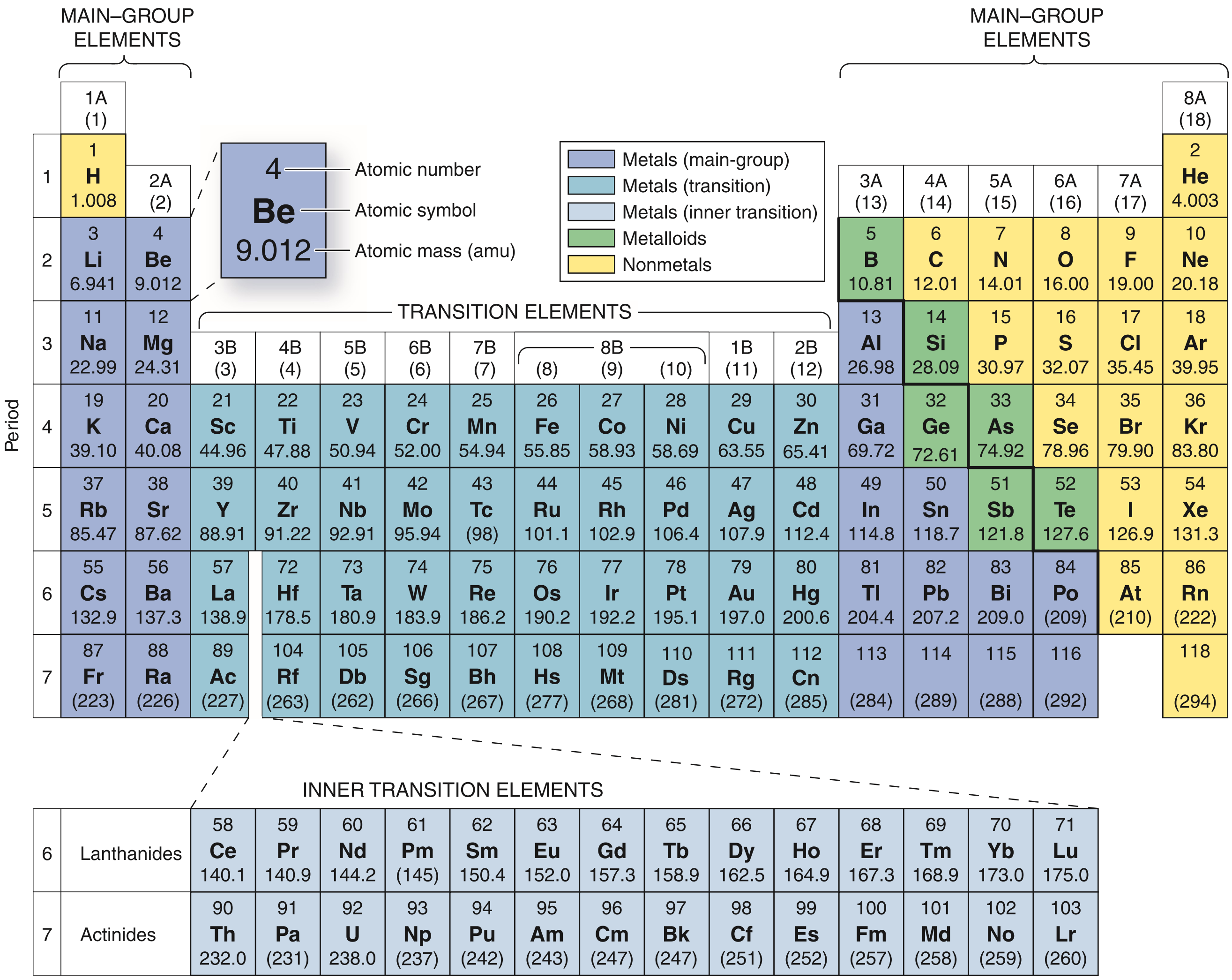



**Figure 2.9** The modern periodic table. As of early 2011, elements 113–116 and 118 had not been named, and the synthesis of element 117 had not been confirmed. (See also the footnote on p. 45.)

- The **nonmetals** (yellow) lie in the small upper-right portion of the table. They are generally gases or dull, brittle solids at room temperature (bromine is the only liquid) and conduct heat and electricity poorly.
- The **metalloids** (green; also called **semimetals**), which lie along the staircase line, have properties between those of metals and nonmetals.

Two major points to keep in mind:

1. In general, elements in a group have **similar** chemical properties and elements in a period have **different** chemical properties.
2. Despite this classification of three types of elements, in reality, there is a gradation in properties from left to right and top to bottom.

It is important to learn some of the group (family) names. Group 1A(1), except for hydrogen, consists of the *alkali metals,* and Group 2A(2) consists of the *alkaline earth metals.* Both groups consist of highly reactive elements. The *halogens,* Group 7A(17), are highly reactive nonmetals, whereas the *noble gases,* Group 8A(18), are relatively unreactive nonmetals. Other main groups [3A(13) to 6A(16)] are often named for the first element in the group; for example, Group 6A(16) is the *oxygen family.*

## Summary of Section 2.6

- In the periodic table, the elements are arranged by atomic number into horizontal periods and vertical groups.
- Nonmetals appear in the upper-right portion of the table, metalloids lie along a staircase line, and metals fill the rest of the table.
- Elements within a group have similar behavior, whereas elements within a period have dissimilar behavior.

# 2.7 • COMPOUNDS: INTRODUCTION TO BONDING

Only a few elements occur free in nature. The noble gases—helium (He), neon (Ne), argon (Ar), krypton (Kr), xenon (Xe), and radon (Rn)—occur in air as separate atoms. In addition to occurring in compounds, oxygen (O), nitrogen (N), and sulfur (S) occur in their most common elemental form as the molecules $O_2$, $N_2$, and $S_8$, and carbon (C) occurs in vast, nearly pure deposits of coal. And some metals—copper (Cu), silver (Ag), gold (Au), and platinum (Pt)—are also sometimes found uncombined. But, aside from these few exceptions, *the overwhelming majority of elements occur in compounds, combined with other elements.*

Elements combine in two general ways and both involve *the electrons of the atoms of interacting elements:*

1. *Transferring electrons* from one element to another to form **ionic compounds**
2. *Sharing electrons* between atoms of different elements to form **covalent compounds**

These processes generate **chemical bonds,** the forces that hold the atoms together in a compound. This section introduces compound formation, which we'll discuss in much more detail in later chapters.

## The Formation of Ionic Compounds

Ionic compounds are composed of **ions,** charged particles that form when an atom (or small group of atoms) gains or loses one or more electrons. The simplest type of ionic compound is a **binary ionic compound,** one composed of two elements. It typically forms *when a metal reacts with a nonmetal:*

- Each metal atom *loses* one or more electrons and becomes a **cation,** a positively charged ion.
- Each nonmetal atom *gains* one or more of the electrons lost by the metal atom and becomes an **anion,** a negatively charged ion.

In effect, the metal atoms *transfer electrons* to the nonmetal atoms. The resulting large numbers of cations and anions attract each other and form the ionic compound. A cation or anion derived from a single atom is called a **monatomic ion;** we'll discuss polyatomic ions, those derived from a small group of atoms, later.

**The Case of Sodium Chloride** *All ionic compounds are solid arrays of oppositely charged ions.* The formation of the binary ionic compound sodium chloride, common table salt, from its elements is depicted in Figure 2.10. In the electron transfer, a sodium atom *loses* one electron and forms a sodium cation, $Na^+$. (The charge on the ion is written as a *right superscript.*) A chlorine atom *gains* the electron and becomes a chloride anion, $Cl^-$. (The name change when the nonmetal atom becomes an anion is discussed in the next section.) The oppositely charged ions ($Na^+$ and $Cl^-$) attract each other, and the similarly charged ions ($Na^+$ and $Na^+$, or $Cl^-$ and $Cl^-$) repel each other. The resulting solid aggregation is a regular array of alternating $Na^+$ and $Cl^-$ ions that extends in all three dimensions. Even the tiniest visible grain of table salt contains an enormous number of sodium and chloride ions.

**Coulomb's Law** The strength of the ionic bonding depends to a great extent on the net strength of these attractions and repulsions and is described by *Coulomb's law,* which can be expressed as follows: *the energy of attraction (or repulsion) between two particles is directly proportional to the product of the charges and inversely proportional to the distance between them.*

$$\text{Energy} \propto \frac{\text{charge 1} \times \text{charge 2}}{\text{distance}}$$

In other words, as is summarized in Figure 2.11,

- Ions with higher charges attract (or repel) each other more strongly than ions with lower charges.
- Smaller ions attract (or repel) each other more strongly than larger ions, because their charges are closer together.

**Predicting the Number of Electrons Lost or Gained** *Ionic compounds are neutral* because they contain equal numbers of positive and negative *charges.* Thus, there are equal numbers of $Na^+$ and $Cl^-$ ions in sodium chloride, because both ions are singly charged. But there are two $Na^+$ ions for each oxide ion, $O^{2-}$, in sodium oxide because two 1+ ions balance one 2− ion.

Can we predict the number of electrons a given atom will lose or gain when it forms an ion? For A-group elements, we usually find that metal atoms lose electrons and nonmetal atoms gain electrons to *form ions with the same number of electrons as in an atom of the nearest noble gas* [Group 8A(18)]. Noble gases have a stability that is related to their number (and arrangement) of electrons. Thus, a sodium

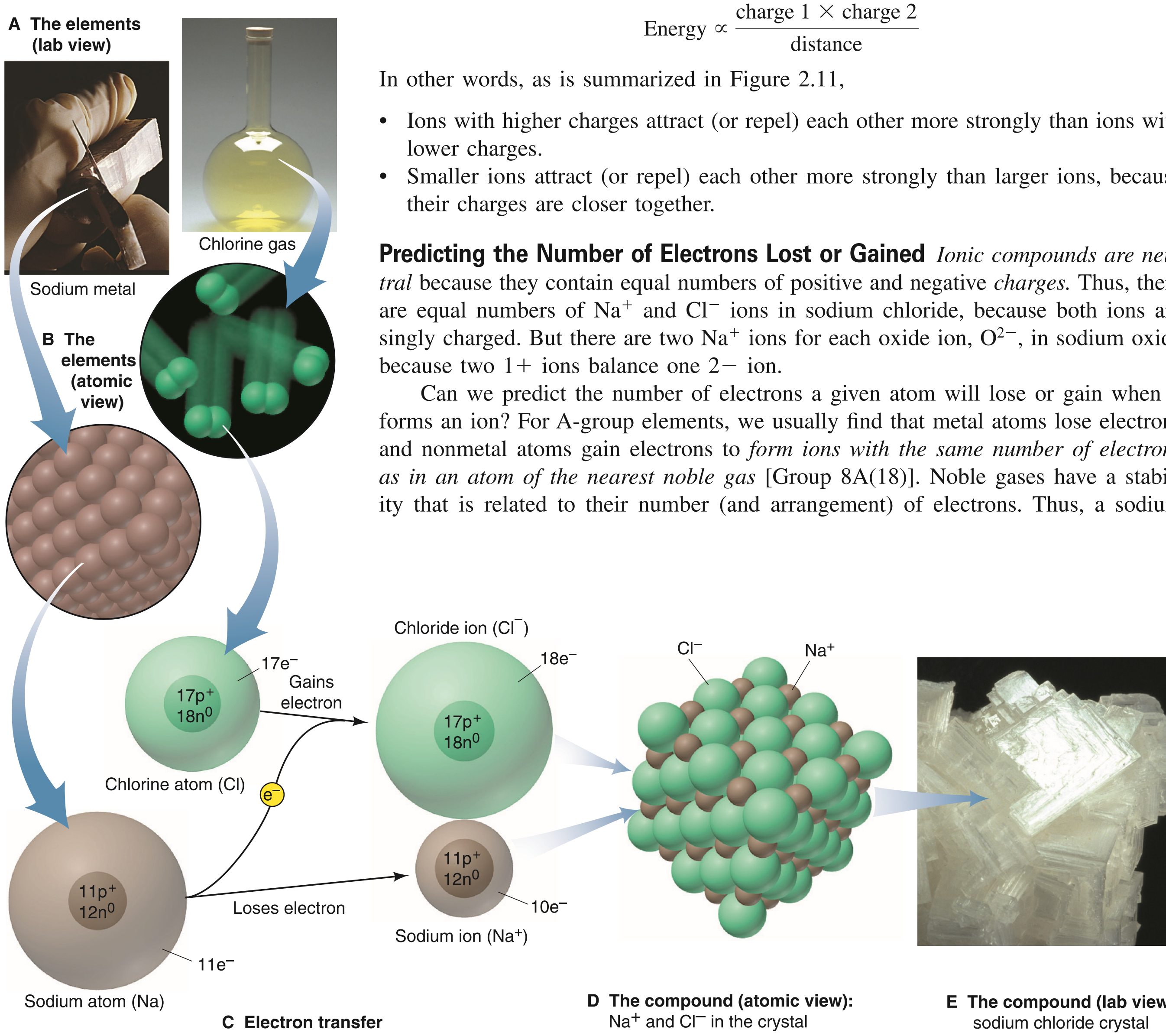



**Figure 2.10** The formation of an ionic compound. **A,** The two elements as seen in the laboratory. **B,** The elements on the atomic scale. **C,** The electron transfer from Na atom to Cl atom to form $Na^+$ and $Cl^-$ ions. **D,** Countless $Na^+$ and $Cl^-$ ions attract each other and form a regular three-dimensional array. **E,** Crystalline NaCl occurs naturally as the mineral halite.

atom (11e⁻) can attain the stability of a neon atom (10e⁻), the nearest noble gas, by losing one electron. Similarly, a chlorine atom (17e⁻) attains the stability of an argon atom (18e⁻), its nearest noble gas, by gaining one electron. Thus, in general, when an element located near a noble gas forms a monatomic ion,

- *Metals lose electrons:* elements in Group 1A(1) lose one electron, elements in Group 2A(2) lose two, and aluminum in Group 3A(13) loses three.
- *Nonmetals gain electrons:* elements in Group 7A(17) gain one electron, oxygen and sulfur in Group 6A(16) gain two, and nitrogen in Group 5A(15) gains three.

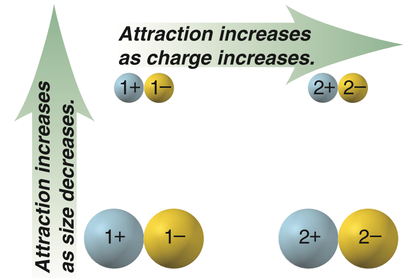



**Figure 2.11** Factors that influence the strength of ionic bonding.

## Sample Problem 2.6 Predicting the Ion an Element Forms

**Problem** What monatomic ions do the following elements form?
**(a)** Iodine ($Z$ = 53) **(b)** Calcium ($Z$ = 20) **(c)** Aluminum ($Z$ = 13)

**Plan** We use the given $Z$ value to find the element in the periodic table and see where its group lies relative to the noble gases. Elements in Groups 1A, 2A, and 3A *lose* electrons to attain the same number as the nearest noble gas and become positive ions; those in Groups 5A, 6A, and 7A *gain* electrons and become negative ions.

**Solution** **(a)** $I^-$ Iodine ($_{53}I$) is in Group 7A(17), the halogens. Like any member of this group, it gains 1 electron to attain the same number as the nearest Group 8A(18) member, in this case, $_{54}Xe$.
**(b)** $Ca^{2+}$ Calcium ($_{20}Ca$) is in Group 2A(2), the alkaline earth metals. Like any Group 2A member, it loses 2 electrons to attain the same number as the nearest noble gas, $_{18}Ar$.
**(c)** $Al^{3+}$ Aluminum ($_{13}Al$) is a metal in the boron family [Group 3A(13)] and thus loses 3 electrons to attain the same number as its nearest noble gas, $_{10}Ne$.

**FOLLOW-UP PROBLEM 2.6** What monatomic ion is formed from **(a)** $_{16}S$; **(b)** $_{37}Rb$; **(c)** $_{56}Ba$?

## The Formation of Covalent Compounds

*Covalent compounds form when elements share electrons, which usually occurs between nonmetals.* The simplest case of electron sharing occurs not in a compound but between two hydrogen atoms (H; $Z$ = 1). Imagine two separated H atoms approaching each other (Figure 2.12). As they get closer, the nucleus of each atom attracts the electron of the other atom more strongly, but repulsions between the nuclei and between the electrons are still weak. As the separated atoms begin to interpenetrate each other, these repulsions increase. At some optimum distance between the nuclei, attractions balance repulsions, and the two atoms form a **covalent bond,** a pair of electrons mutually attracted by the two nuclei. The result is a hydrogen molecule, in which each electron no longer "belongs" to a particular H atom: the two electrons are *shared* by the two nuclei. A sample of hydrogen gas consists of these diatomic molecules ($H_2$)—pairs of atoms that are chemically bound and behave as an independent unit—*not* separate H atoms. Figure 2.13 shows other nonmetals that exist as molecules at room temperature.

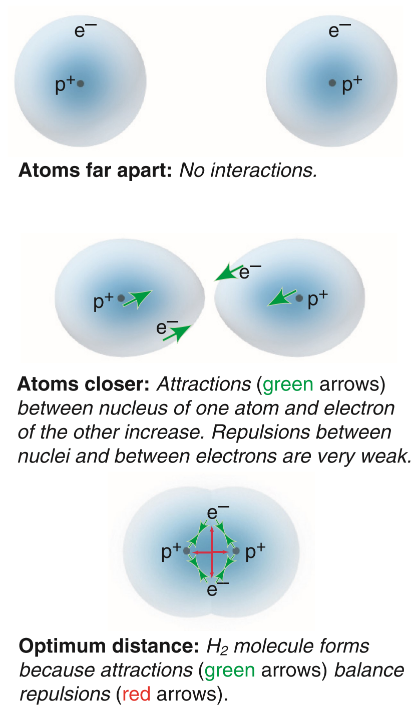



**Figure 2.12** Formation of a covalent bond between two H atoms.

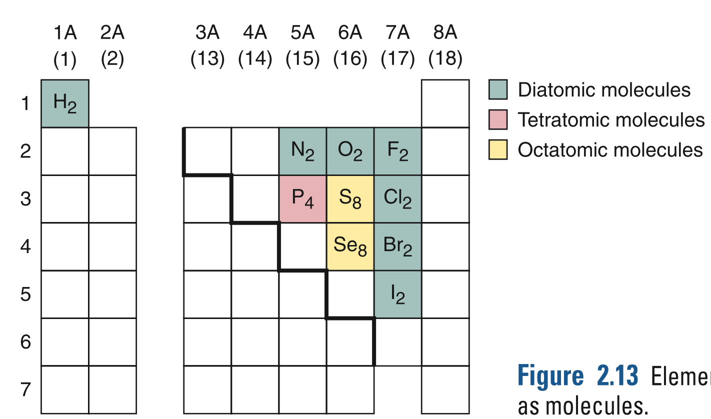
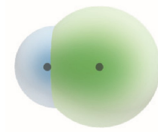
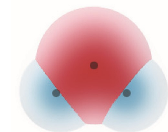



**Figure 2.13** Elements that occur as molecules.

*Atoms of different elements share electrons to form the molecules of a covalent compound.* A sample of hydrogen fluoride, for example, consists of molecules in which one H atom forms a covalent bond with one F atom; water consists of molecules in which one O atom forms covalent bonds with two H atoms:

Hydrogen fluoride, HF





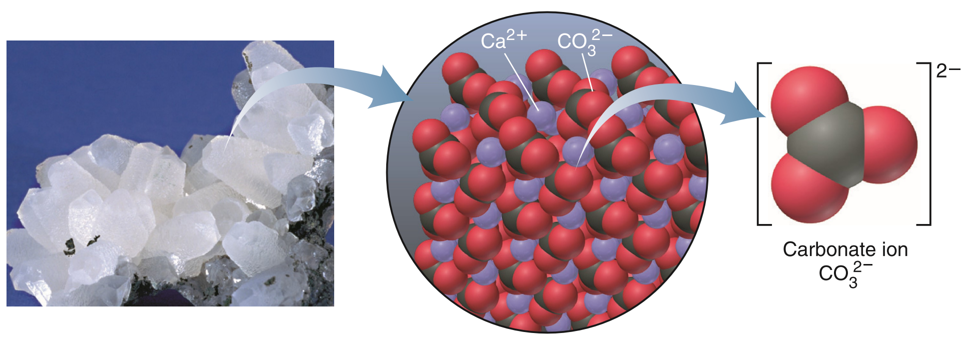
Water, $H_2O$

**Distinguishing the Entities in Covalent and Ionic Substances** There is a key distinction between the chemical entities in covalent substances and in ionic substances. *Most covalent substances consist of molecules*. A cup of water, for example, consists of individual water molecules lying near each other. In contrast, under ordinary conditions, *there are no molecules in an ionic compound*. A piece of sodium chloride, for example, is a continuous array in three dimensions of oppositely charged sodium and chloride ions, *not* a collection of individual sodium chloride "molecules."

Another key distinction between covalent and ionic substances concerns the nature of the particles attracting each other. Covalent bonding involves the mutual attraction between two (positively charged) nuclei and the two (negatively charged) electrons that reside between them. Ionic bonding involves the mutual attraction among positive and negative ions.

**Polyatomic Ions: Covalent Bonds Within Ions** Many ionic compounds contain **polyatomic ions,** which consist of two or more atoms bonded *covalently* and have a net positive or negative charge. For example, Figure 2.14 shows that a crystalline form of calcium carbonate *(left)* occurs on the atomic scale as an array *(center)* of polyatomic carbonate anions and monatomic calcium cations. The carbonate ion *(right)* consists of a carbon atom covalently bonded to three oxygen atoms, and two additional electrons give the ion its 2− charge. In many reactions, the polyatomic ion stays together as a unit.

**Figure 2.14** The carbonate ion in calcium carbonate.




## Summary of Section 2.7

- Although a few elements occur uncombined in nature, the great majority exist in compounds.
- Ionic compounds form when a metal *transfers* electrons to a nonmetal, and the resulting positive and negative ions attract each other to form a three-dimensional array. In many cases, metal atoms lose and nonmetal atoms gain enough electrons to attain the same number of electrons as in atoms of the nearest noble gas.
- Covalent compounds form when elements, usually nonmetals, *share* electrons. Each covalent bond is an electron pair mutually attracted by two atomic nuclei.
- Monatomic ions are derived from single atoms. Polyatomic ions consist of two or more covalently bonded atoms that have a net positive or negative charge due to a deficit or excess of electrons.

## 2.8 • FORMULAS, NAMES, AND MASSES OF COMPOUNDS

In a **chemical formula,** element symbols and, often, numerical subscripts show the type and number of each atom in the smallest unit of the substance. In this section, you'll learn how to write the names and formulas of ionic and simple covalent compounds, how to calculate the mass of a compound from its formula, and how to visualize molecules with three-dimensional models.

### Binary Ionic Compounds

Let's begin with two general rules:

- For *all* ionic compounds, *names and formulas give the positive ion (cation) first and the negative ion (anion) second.*
- For all *binary* ionic compounds, *the name of the cation is the name of the metal, and the name of the anion has the suffix -ide added to the root of the name of the nonmetal.*

For example, the anion formed from brom*ine* is named brom*ide* (brom+ide). Therefore, the compound formed from the metal calcium and the nonmetal bromine is named *calcium bromide.*

In general, if the metal of a binary ionic compound is a main-group element (A groups) it usually forms a single type of ion; if it is a transition element (B groups), it often forms more than one. We discuss each case in turn.

**Compounds of Elements That Form One Ion** The periodic table presents some key points about the formulas of main-group monatomic ions:

- Monatomic ions of main-group elements have the same ionic charge; the alkali metals—Li, Na, K, Rb, Cs, and Fr—form ions with a 1+ charge; the halogens—F, Cl, Br, and I—form ions with a 1− charge; and so forth.
- For cations, ion charge equals A-group number: Na is in Group 1A and forms $Na^+$, Ba is in Group 2A and forms $Ba^{2+}$. (Exceptions in Figure 2.15 are $Sn^{2+}$ and $Pb^{2+}$.)
- For anions, ion charge equals A-group number minus 8; for example, S is in Group 6A ($6 - 8 = -2$) and thus forms $S^{2-}$.

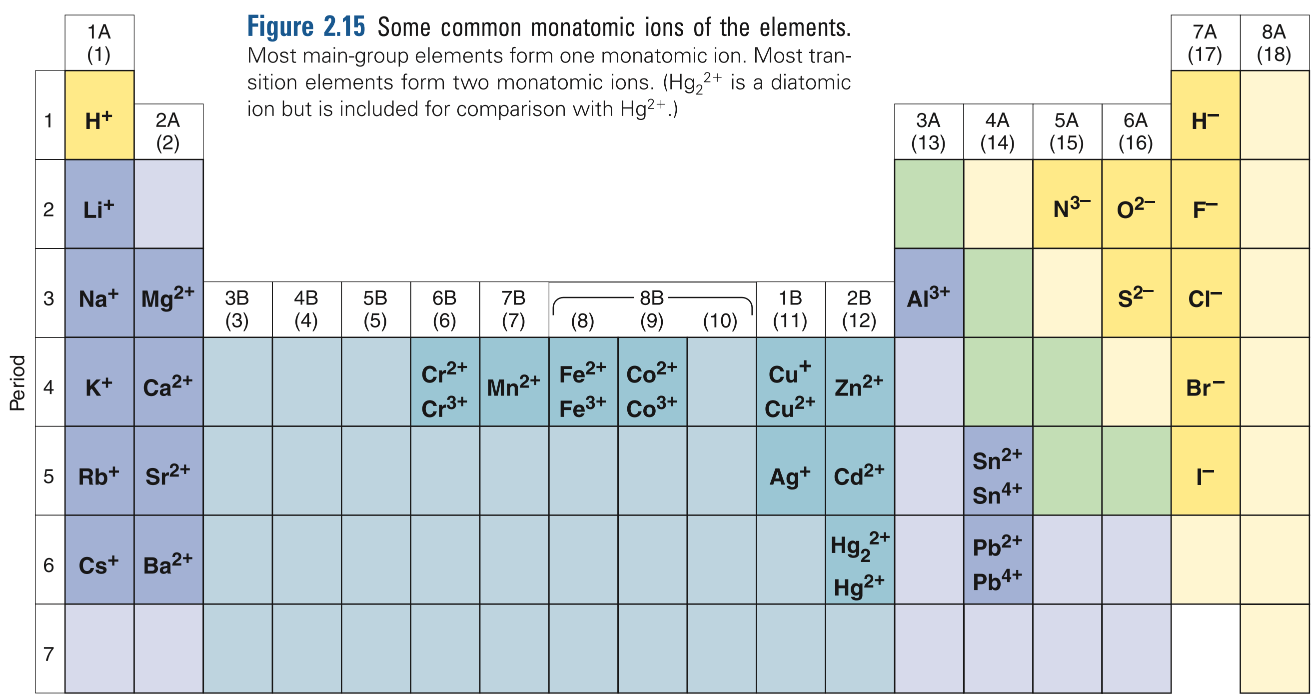



**Figure 2.15** Some common monatomic ions of the elements. Most main-group elements form one monatomic ion. Most transition elements form two monatomic ions. ($Hg_2^{2+}$ is a diatomic ion but is included for comparison with $Hg^{2+}$.)

**Table 2.3 Common Monatomic Ions***

| Charge | Formula | Name |
|---|---|---|
| **Cations** | | |
| 1+ | $H^+$ | hydrogen |
| | **$Li^+$** | **lithium** |
| | **$Na^+$** | **sodium** |
| | **$K^+$** | **potassium** |
| | $Cs^+$ | cesium |
| | **$Ag^+$** | **silver** |
| 2+ | **$Mg^{2+}$** | **magnesium** |
| | **$Ca^{2+}$** | **calcium** |
| | $Sr^{2+}$ | strontium |
| | **$Ba^{2+}$** | **barium** |
| | **$Zn^{2+}$** | **zinc** |
| | $Cd^{2+}$ | cadmium |
| 3+ | $Al^{3+}$ | aluminum |
| **Anions** | | |
| 1− | $H^-$ | hydride |
| | **$F^-$** | **fluoride** |
| | **$Cl^-$** | **chloride** |
| | **$Br^-$** | **bromide** |
| | **$I^-$** | **iodide** |
| 2− | **$O^{2-}$** | **oxide** |
| | **$S^{2-}$** | **sulfide** |
| 3− | $N^{3-}$ | nitride |

*Those in boldface are most common.

Try to memorize the A-group monatomic ions in Table 2.3 (all of these listed except $Ag^+$, $Zn^{2+}$, and $Cd^{2+}$) according to their positions in Figure 2.15. These ions have *the same number of electrons as an atom of the nearest noble gas.*

Because an ionic compound consists of an array of ions rather than separate molecules, its formula represents the **formula unit,** the *relative* numbers of cations and anions in the compound. The compound has zero net charge, so the positive charges of the cations balance the negative charges of the anions. For example, calcium bromide is composed of $Ca^{2+}$ ions and $Br^-$ ions, so two $Br^-$ balance each $Ca^{2+}$. The formula is $CaBr_2$, not $Ca_2Br$. In this and all other formulas,

- The subscript refers to the element *preceding* it.
- The *subscript 1 is understood* from the presence of the element symbol alone (that is, we do not write $Ca_1Br_2$).
- The charge (without the sign) of one ion becomes the subscript of the other:

$$Ca^{2+}\ Br^{1-} \quad \text{gives} \quad Ca_1Br_2 \quad \text{or} \quad CaBr_2$$

- Reduce the subscripts to the smallest whole numbers that retain the ratio of ions. Thus, for example, for the $Ca^{2+}$ and $O^{2-}$ ions in calcium oxide, we get $Ca_2O_2$, which we reduce to the formula CaO.*

## Sample Problem 2.7 Naming Binary Ionic Compounds

**Problem** Name the ionic compound formed from the following pairs of elements: **(a)** magnesium and nitrogen; **(b)** iodine and cadmium; **(c)** strontium and fluorine; **(d)** sulfur and cesium.

**Plan** The key to naming a binary ionic compound is to recognize which element is the metal and which is the nonmetal. When in doubt, check the periodic table. We place the cation name first, add the suffix *-ide* to the nonmetal root, and place the anion name last.

**Solution** **(a)** Magnes*ium* is the metal; *nitr-* is the nonmetal root: magnesium nitride
**(b)** Cadm*ium* is the metal; *iod-* is the nonmetal root: cadmium iodide
**(c)** Stront*ium* is the metal; *fluor-* is the nonmetal root: strontium fluoride (Note the spelling is fl*uo*ride, not fl*ou*ride.)
**(d)** Ces*ium* is the metal; *sulf-* is the nonmetal root: cesium sulfide

**FOLLOW-UP PROBLEM 2.7** For the following ionic compounds, give the name and periodic table group number of each element present: **(a)** zinc oxide; **(b)** silver bromide; **(c)** lithium chloride; **(d)** aluminum sulfide.

## Sample Problem 2.8 Determining Formulas of Binary Ionic Compounds

**Problem** Write formulas for the compounds named in Sample Problem 2.7.

**Plan** We write the formula by finding the smallest number of each ion that gives the neutral compound. These numbers appear as *right subscripts* to the element symbol.

**Solution**
**(a)** $Mg^{2+}$ and $N^{3-}$; three $Mg^{2+}$ ions (6+) balance two $N^{3-}$ ions (6−): $Mg_3N_2$
**(b)** $Cd^{2+}$ and $I^-$; one $Cd^{2+}$ ion (2+) balances two $I^-$ ions (2−): $CdI_2$
**(c)** $Sr^{2+}$ and $F^-$; one $Sr^{2+}$ ion (2+) balances two $F^-$ ions (2−): $SrF_2$
**(d)** $Cs^+$ and $S^{2-}$; two $Cs^+$ ions (2+) balance one $S^{2-}$ ion (2−): $Cs_2S$

**Comment** **1.** The subscript 1 is understood and so not written; thus, in (b), we do *not* write $Cd_1I_2$.
**2.** Ion charges do *not* appear in the compound formula; thus, in (c), we do *not* write $Sr^{2+}F_2^-$.

**FOLLOW-UP PROBLEM 2.8** Write the formulas of the compounds named in Follow-up Problem 2.7.

*Compounds of the mercury(I) ion, such as $Hg_2Cl_2$, and peroxides of the alkali metals, such as $Na_2O_2$, are the only two common exceptions to this step; in fact, reducing the subscripts for these compounds would give the incorrect formulas HgCl and NaO.

**Table 2.4 Some Metals That Form More Than One Monatomic Ion***

| Element | Ion Formula | Systematic Name | Common (Trivial) Name |
|---|---|---|---|
| Chromium | $Cr^{2+}$ | chromium(II) | chromous |
| | **$Cr^{3+}$** | **chromium(III)** | chromic |
| Cobalt | $Co^{2+}$ | cobalt(II) | |
| | $Co^{3+}$ | cobalt(III) | |
| Copper | **$Cu^{+}$** | **copper(I)** | cuprous |
| | **$Cu^{2+}$** | **copper(II)** | cupric |
| Iron | **$Fe^{2+}$** | **iron(II)** | ferrous |
| | **$Fe^{3+}$** | **iron(III)** | ferric |
| Lead | **$Pb^{2+}$** | **lead(II)** | |
| | $Pb^{4+}$ | lead(IV) | |
| Mercury | $Hg_2^{2+}$ | mercury(I) | mercurous |
| | **$Hg^{2+}$** | **mercury(II)** | mercuric |
| Tin | **$Sn^{2+}$** | **tin(II)** | stannous |
| | $Sn^{4+}$ | tin(IV) | stannic |

*Listed alphabetically by metal name; the ions in **boldface** are most common.

**Compounds with Metals That Form More Than One Ion** As noted earlier, many metals, particularly the transition elements (B groups), can form more than one ion. Table 2.4 lists some examples; see Figure 2.15 for their placement in the periodic table. Names of compounds containing these elements include a *Roman numeral within parentheses* immediately after the metal ion's name to indicate its ionic charge. For example, iron can form $Fe^{2+}$ and $Fe^{3+}$ ions. The two compounds that iron forms with chlorine are $FeCl_2$, named iron(II) chloride (spoken "iron two chloride"), and $FeCl_3$, named iron(III) chloride.

We are focusing here on systematic names, but some common (trivial) names are still used. In common names for certain metal ions, the Latin root of the metal is followed by either of two suffixes (see Table 2.4):

- The suffix *-ous* for the ion with the lower charge
- The suffix *-ic* for the ion with the higher charge

Thus, iron(II) chloride is also called ferr*ous* chloride and iron(III) chloride is ferr*ic* chloride. (Memory aid: there is an ***o*** in -***o****us* and ***l*o*w*er**, and an ***i*** in -***i****c* and *h**i**gher.*)

## Sample Problem 2.9 Determining Names and Formulas of Ionic Compounds of Metals That Form More Than One Ion

**Problem** Give the systematic names for the formulas or the formulas for the names of each compound: **(a)** tin(II) fluoride; **(b)** $CrI_3$; **(c)** ferric oxide; **(d)** CoS.

**Solution** **(a)** Tin(II) ion is $Sn^{2+}$; fluoride is $F^-$. Two $F^-$ ions balance one $Sn^{2+}$ ion: tin(II) fluoride is $SnF_2$. (The common name is stannous fluoride.)
**(b)** The anion is $I^-$, iodide, and the formula shows three $I^-$. Therefore, the cation must be $Cr^{3+}$, chromium(III) ion: $CrI_3$ is chromium(III) iodide. (The common name is chromic iodide.)
**(c)** *Ferric* is the common name for iron(III) ion, $Fe^{3+}$; oxide ion is $O^{2-}$. To balance the charges, the formula is $Fe_2O_3$. [The systematic name is iron(III) oxide.]
**(d)** The anion is sulfide, $S^{2-}$, which requires that the cation be $Co^{2+}$. The name is cobalt(II) sulfide.

**FOLLOW-UP PROBLEM 2.9** Give the systematic names for the formulas or the formulas for the names of each compound: **(a)** lead(IV) oxide; **(b)** $Cu_2S$; **(c)** $FeBr_2$; **(d)** mercuric chloride.

**Table 2.5 Common Polyatomic Ions***

| Formula | Name |
|---|---|
| **Cations** | |
| $\mathbf{NH_4^+}$ | **ammonium** |
| $\mathbf{H_3O^+}$ | **hydronium** |
| **Anions** | |
| $\mathbf{CH_3COO^-}$ (or $\mathbf{C_2H_3O_2^-}$) | **acetate** |
| $CN^-$ | cyanide |
| $\mathbf{OH^-}$ | **hydroxide** |
| $ClO^-$ | hypochlorite |
| $ClO_2^-$ | chlorite |
| $\mathbf{ClO_3^-}$ | **chlorate** |
| $\mathbf{ClO_4^-}$ | **perchlorate** |
| $NO_2^-$ | nitrite |
| $\mathbf{NO_3^-}$ | **nitrate** |
| $\mathbf{MnO_4^-}$ | **permanganate** |
| $\mathbf{CO_3^{2-}}$ | **carbonate** |
| $\mathbf{HCO_3^-}$ | **hydrogen carbonate** (or **bicarbonate**) |
| $CrO_4^{2-}$ | chromate |
| $\mathbf{Cr_2O_7^{2-}}$ | **dichromate** |
| $O_2^{2-}$ | peroxide |
| $\mathbf{PO_4^{3-}}$ | **phosphate** |
| $HPO_4^{2-}$ | hydrogen phosphate |
| $H_2PO_4^-$ | dihydrogen phosphate |
| $SO_3^{2-}$ | sulfite |
| $\mathbf{SO_4^{2-}}$ | **sulfate** |
| $HSO_4^-$ | hydrogen sulfate (or bisulfate) |

*Boldface ions are most common.

| No. of O atoms | Prefix | Root | Suffix |
|---|---|---|---|
| ↑ | per | *root* | ate |
| | | *root* | ate |
| | | *root* | ite |
| | hypo | *root* | ite |

**Figure 2.16** Naming oxoanions. Prefixes and suffixes indicate the number of O atoms in the anion.

## Compounds That Contain Polyatomic Ions

Many ionic compounds contain polyatomic ions. Table 2.5 shows some common polyatomic ions. Remember that *the polyatomic ion stays together as a charged unit.* The formula for potassium nitrate is $KNO_3$: each $K^+$ balances one $NO_3^-$. The formula for sodium carbonate is $Na_2CO_3$: two $Na^+$ balance one $CO_3^{2-}$. *When two or more of the same polyatomic ion are present in the formula unit, that ion appears in parentheses with the subscript written outside.* For example, calcium nitrate contains one $Ca^{2+}$ and two $NO_3^-$ ions and has the formula $Ca(NO_3)_2$. Parentheses and a subscript are *only* used if *more than one* of a given polyatomic ion is present; thus, sodium nitrate is $NaNO_3$, *not* $Na(NO_3)$.

**Families of Oxoanions** As Table 2.5 shows, most polyatomic ions are **oxoanions** (or *oxyanions*), those in which an element, usually a nonmetal, is bonded to one or more oxygen atoms. There are several families of two or four oxoanions that differ only in the number of oxygen atoms. The following simple naming conventions are used with these ions.

With two oxoanions in the family:

- The ion with *more* O atoms takes the nonmetal root and the suffix *-ate*.
- The ion with *fewer* O atoms takes the nonmetal root and the suffix *-ite*.

For example, $SO_4^{2-}$ is the sulf*ate* ion, and $SO_3^{2-}$ is the sulf*ite* ion; similarly, $NO_3^-$ is nitr*ate*, and $NO_2^-$ is nitr*ite*.

With four oxoanions in the family (a halogen bonded to O) (Figure 2.16):

- The ion with *most* O atoms has the prefix *per-*, the nonmetal root, and the suffix *-ate*.
- The ion with *one fewer* O atom has just the root and the suffix *-ate*.
- The ion with *two fewer* O atoms has just the root and the suffix *-ite*.
- The ion with *least (three fewer)* O atoms has the prefix *hypo-*, the root, and the suffix *-ite*.

For example, for the four chlorine oxoanions,

$ClO_4^-$ is *per*chlor*ate*, $ClO_3^-$ is chlor*ate*, $ClO_2^-$ is chlor*ite*, $ClO^-$ is *hypo*chlor*ite*

**Hydrated Ionic Compounds** Ionic compounds called **hydrates** have a specific number of water molecules in each formula unit, which is shown after a centered dot in the formula and noted in the name by a Greek numerical prefix before the word *hydrate*. Table 2.6 shows these prefixes. For example, Epsom salt has seven water molecules in each formula unit: the formula is $MgSO_4{\cdot}7H_2O$, and the name is magnesium sulfate *hepta*hydrate. Similarly, the mineral gypsum has the formula $CaSO_4{\cdot}2H_2O$ and the name calcium sulfate *di*hydrate. The water molecules, referred to as "waters of hydration," are part of the hydrate's structure. Heating can remove some or all of them, leading to a different substance. For example, when heated strongly, blue copper(II) sulfate pentahydrate ($CuSO_4{\cdot}5H_2O$) is converted to white copper(II) sulfate ($CuSO_4$).

### Sample Problem 2.10 Determining Names and Formulas of Ionic Compounds Containing Polyatomic Ions

**Problem** Give the systematic names for the formulas or the formulas for the names of the following compounds: **(a)** $Fe(ClO_4)_2$; **(b)** sodium sulfite; **(c)** $Ba(OH)_2{\cdot}8H_2O$.

**Solution** **(a)** $ClO_4^-$ is perchlorate, which has a 1− charge, so the cation must be $Fe^{2+}$. The name is iron(II) perchlorate. (The common name is ferrous perchlorate.)

**(b)** Sodium is $Na^+$; sulfite is $SO_3^{2-}$, and two $Na^+$ ions balance one $SO_3^{2-}$ ion. The formula is $Na_2SO_3$.

**(c)** $Ba^{2+}$ is barium; $OH^-$ is hydroxide. There are eight *(octa-)* water molecules in each formula unit. The name is barium hydroxide octahydrate.

**FOLLOW-UP PROBLEM 2.10** Give the systematic names for the formulas or the formulas for the names of the following compounds: **(a)** cupric nitrate trihydrate; **(b)** zinc hydroxide; **(c)** LiCN.

**Sample Problem 2.11** **Recognizing Incorrect Names and Formulas of Ionic Compounds**

**Problem** Explain what is wrong with the name or formula at the end of each statement, and correct it:
**(a)** $Ba(C_2H_3O_2)_2$ is called barium diacetate.
**(b)** Sodium sulfide has the formula $(Na)_2SO_3$.
**(c)** Iron(II) sulfate has the formula $Fe_2(SO_4)_3$.
**(d)** Cesium carbonate has the formula $Cs_2(CO_3)$.

**Solution** **(a)** The charge of the $Ba^{2+}$ ion *must* be balanced by *two* $C_2H_3O_2^-$ ions, so the prefix *di-* is unnecessary. For ionic compounds, we do not indicate the number of ions with numerical prefixes. The correct name is barium acetate.
**(b)** Two mistakes occur here. The sodium ion is monatomic, so it does *not* require parentheses. The sulfide ion is $S^{2-}$, *not* $SO_3^{2-}$ (which is sulfite). The correct formula is $Na_2S$.
**(c)** The Roman numeral refers to the charge of the ion, *not* the number of ions in the formula. $Fe^{2+}$ is the cation, so it requires one $SO_4^{2-}$ to balance its charge. The correct formula is $FeSO_4$. [$Fe_2(SO_4)_3$ is the formula for iron(III) sulfate.]
**(d)** Parentheses are *not* required when only one polyatomic ion of a kind is present. The correct formula is $Cs_2CO_3$.

**FOLLOW-UP PROBLEM 2.11** State why the formula or name at the end of each statement is incorrect, and correct it:
**(a)** Ammonium phosphate is $(NH_3)_4PO_4$.
**(b)** Aluminum hydroxide is $AlOH_3$.
**(c)** $Mg(HCO_3)_2$ is manganese(II) carbonate.
**(d)** $Cr(NO_3)_3$ is chromic(III) nitride.
**(e)** $Ca(NO_2)_2$ is cadmium nitrate.

**Table 2.6 Numerical Prefixes* for Ionic Hydrates and Binary Covalent Compounds**

| Number | Prefix |
|---|---|
| 1 | mono- |
| 2 | di- |
| 3 | tri- |
| 4 | tetra- |
| 5 | penta- |
| 6 | hexa- |
| 7 | hepta- |
| 8 | octa- |
| 9 | nona- |
| 10 | deca- |

*It is common practice to drop the final "a" in names of oxides, for example, tetroxide, *not* tetra*a*oxide.

## Acid Names from Anion Names

Acids are an important group of hydrogen-containing compounds that have been used in chemical reactions for many centuries. In the laboratory, acids are typically used in water solution. When naming them and writing their formulas, we consider acids as anions that are connected to the number of hydrogen ions ($H^+$) needed for charge neutrality. The two common types of acids are binary acids and oxoacids:

1. *Binary acid* solutions form when certain gaseous compounds dissolve in water. For example, when gaseous hydrogen chloride (HCl) dissolves in water, it forms hydrochloric acid, that is,

   Prefix *hydro-* + nonmetal *root* + suffix *-ic* + separate word *acid*
   hydro + chlor + ic + acid

   or *hydrochloric acid.* This naming pattern holds for many compounds in which hydrogen combines with an anion that has an *-ide* suffix.

2. *Oxoacid* names are similar to those of the oxoanions, except for two suffix changes:
   - *-ate* in the anion becomes *-ic* in the acid
   - *-ite* in the anion becomes *-ous* in the acid

   The oxoanion prefixes *hypo-* and *per-* are retained. Thus,

   $BrO_4^-$ is *per*brom*ate*, and $HBrO_4$ is *per*brom*ic* acid.
   $IO_2^-$ is iod*ite*, and $HIO_2$ is iod*ous* acid.

**Sample Problem 2.12** **Determining Names and Formulas of Anions and Acids**

**Problem** Name each of the following anions and give the name and formula of the acid derived from it: **(a)** $Br^-$; **(b)** $IO_3^-$; **(c)** $CN^-$; **(d)** $SO_4^{2-}$; **(e)** $NO_2^-$.

**Solution** **(a)** The anion is bromide; the acid is hydrobromic acid, HBr.
**(b)** The anion is iodate; the acid is iodic acid, $HIO_3$.
**(c)** The anion is cyanide; the acid is hydrocyanic acid, HCN.
**(d)** The anion is sulfate; the acid is sulfuric acid, $H_2SO_4$. (In this case, the suffix is added to the element name *sulfur,* not to the root, *sulf-*.)
**(e)** The anion is nitrite; the acid is nitrous acid, $HNO_2$.

**Comment** We must add *two* $H^+$ ions to the sulfate ion to obtain sulfuric acid because $SO_4^{2-}$ has a 2− charge.

**FOLLOW-UP PROBLEM 2.12** Write the formula for the name or name for the formula of each acid: **(a)** chloric acid; **(b)** HF; **(c)** acetic acid; **(d)** sulfurous acid; **(e)** HBrO.

## Binary Covalent Compounds

**Binary covalent compounds** are typically formed by the combination of two nonmetals. Some are so familiar that we use their common names, such as ammonia ($NH_3$), methane ($CH_4$), and water ($H_2O$), but most are named systematically:

- The element with the lower group number in the periodic table comes first in the name. The element with the higher group number comes second and is named with its root and the suffix *-ide*. (*Exception:* When the compound contains oxygen and any of the halogens chlorine, bromine, or iodine, the halogen is named first.)
- If both elements are in the same group, the one with the higher period number is named first.
- Covalent compounds use Greek numerical prefixes (see Table 2.6) to indicate the number of atoms of each element. The first element in the name has a prefix *only* when more than one atom of it is present; the second element *usually* has a prefix.

### Sample Problem 2.13 Determining Names and Formulas of Binary Covalent Compounds

**Problem** **(a)** What is the formula of carbon disulfide?
**(b)** What is the name of $PCl_5$?
**(c)** Give the name and formula of the compound whose molecules each consist of two N atoms and four O atoms.

**Solution** **(a)** The prefix *di-* means "two." The formula is $CS_2$.
**(b)** P is the symbol for phosphorus; there are five chlorine atoms, which is indicated by the prefix *penta-*. The name is phosphorus pentachloride.
**(c)** Nitrogen (N) comes first in the name (lower group number). The compound is dinitrogen tetroxide, $N_2O_4$.

**FOLLOW-UP PROBLEM 2.13** Give the name or formula for **(a)** $SO_3$; **(b)** $SiO_2$; **(c)** dinitrogen monoxide; **(d)** selenium hexafluoride.

### Sample Problem 2.14 Recognizing Incorrect Names and Formulas of Binary Covalent Compounds

**Problem** Explain what is wrong with the name or formula at the end of each statement, and correct it: **(a)** $SF_4$ is monosulfur pentafluoride. **(b)** Dichlorine heptoxide is $Cl_2O_6$. **(c)** $N_2O_3$ is dinitrotrioxide.

**Solution** **(a)** There are two mistakes. *Mono-* is not needed if there is only one atom of the first element, and the prefix for four is *tetra-*, not *penta-*. The correct name is sulfur tetrafluoride.
**(b)** The prefix *hepta-* indicates seven, not six. The correct formula is $Cl_2O_7$.
**(c)** The full name of the first element is needed, and a space separates the two element names. The correct name is dinitrogen trioxide.

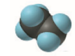
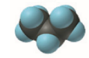
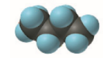
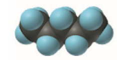
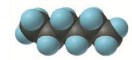
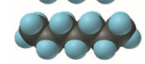
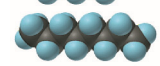
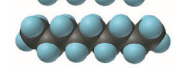
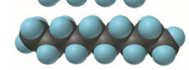
**FOLLOW-UP PROBLEM 2.14** Explain what is wrong with the name or formula at the end of each statement, and correct it: **(a)** $S_2Cl_2$ is disulfurous dichloride. **(b)** Nitrogen monoxide is $N_2O$. **(c)** $BrCl_3$ is trichlorine bromide.

## The Simplest Organic Compounds: Straight-Chain Alkanes

Organic compounds typically have complex structures that consist of chains, branches, and/or rings of carbon atoms bonded to hydrogen atoms and, often, to atoms of oxygen, nitrogen, and a few other elements. At this point, we'll lay the groundwork for naming organic compounds by focusing on the simplest ones. Rules for naming more complex ones are detailed in Chapter 15.

*Hydrocarbons,* the simplest type of organic compound, contain *only* carbon and hydrogen. *Alkanes* are the simplest type of hydrocarbon; many function as important fuels, such as methane, propane, butane, and the mixture that makes up gasoline. The simplest alkanes to name are the *straight-chain alkanes* because the carbon chains have no branches. Alkanes are named with a *root,* based on the number of C atoms in the chain, followed by the suffix *-ane.* Table 2.7 gives the names, molecular formulas, and space-filling models (discussed shortly) of the first 10 straight-chain alkanes. Note that the roots of the four smallest ones are new, but those for the larger ones are the same as the Greek prefixes shown in Table 2.6.

**Table 2.7 The First 10 Straight-Chain Alkanes**

| Name (Formula) | Model |
|---|---|
| Methane ($CH_4$) |  |
| Ethane ($C_2H_6$) |  |
| Propane ($C_3H_8$) |  |
| Butane ($C_4H_{10}$) |  |
| Pentane ($C_5H_{12}$) |  |
| Hexane ($C_6H_{14}$) |  |
| Heptane ($C_7H_{16}$) |  |
| Octane ($C_8H_{18}$) |  |
| Nonane ($C_9H_{20}$) |  |
| Decane ($C_{10}H_{22}$) |  |

## Molecular Masses from Chemical Formulas

In Section 2.5, we calculated the atomic mass of an element. Using the periodic table and the formula of a compound, we calculate the **molecular mass** (also called *molecular weight*) of a formula unit of the compound as the sum of the atomic masses:

$$\text{Molecular mass} = \text{sum of atomic masses} \quad (2.3)$$

The molecular mass of a water molecule (using atomic masses to four significant figures from the periodic table) is

$$\begin{aligned}\text{Molecular mass of } H_2O &= (2 \times \text{atomic mass of H}) + (1 \times \text{atomic mass of O}) \\ &= (2 \times 1.008 \text{ amu}) + 16.00 \text{ amu} = 18.02 \text{ amu}\end{aligned}$$

Ionic compounds don't consist of molecules, so the mass of a formula unit is termed the **formula mass** instead of *molecular mass.* To calculate the formula mass of a compound with a polyatomic ion, *the number of atoms of each element inside the parentheses is multiplied by the subscript outside the parentheses.* For barium nitrate, $Ba(NO_3)_2$,

Formula mass of $Ba(NO_3)_2$

$$\begin{aligned}&= (1 \times \text{atomic mass of Ba}) + (2 \times \text{atomic mass of N}) + (6 \times \text{atomic mass of O}) \\ &= 137.3 \text{ amu} + (2 \times 14.01 \text{ amu}) + (6 \times 16.00 \text{ amu}) = 261.3 \text{ amu}\end{aligned}$$

We can use atomic masses, not ionic masses, because electron loss equals electron gain, so electron mass is balanced. In the next two sample problems, the name or molecular depiction is used to find a compound's molecular or formula mass.

### Sample Problem 2.15 Calculating the Molecular Mass of a Compound

**Problem** Using the periodic table, calculate the molecular (or formula) mass of **(a)** tetraphosphorus trisulfide; **(b)** ammonium nitrate.

**Plan** We first write the formula, then multiply the number of atoms (or ions) of each element by its atomic mass (from the periodic table), and find the sum.

**Solution** **(a)** The formula is $P_4S_3$.

$$\begin{aligned}\text{Molecular mass} &= (4 \times \text{atomic mass of P}) + (3 \times \text{atomic mass of S}) \\ &= (4 \times 30.97 \text{ amu}) + (3 \times 32.07 \text{ amu}) = \boxed{220.09 \text{ amu}}\end{aligned}$$

**(b)** The formula is $NH_4NO_3$. We count the total number of N atoms even though they belong to different ions:

Formula mass

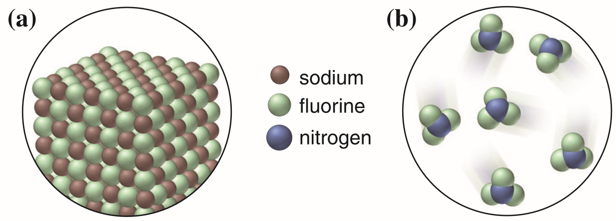
$$\begin{aligned}&= (2 \times \text{atomic mass of N}) + (4 \times \text{atomic mass of H}) + (3 \times \text{atomic mass of O}) \\ &= (2 \times 14.01 \text{ amu}) + (4 \times 1.007 \text{ amu}) + (3 \times 16.00 \text{ amu}) = \boxed{80.05 \text{ amu}}\end{aligned}$$

**Check** You can often find large errors by rounding atomic masses to the nearest 5 and adding: **(a)** $(4 \times 30) + (3 \times 30) = 210 \approx 220.09$. The sum has two decimal places because the atomic masses have two. **(b)** $(2 \times 15) + 4 + (3 \times 15) = 79 \approx 80.05$.

**FOLLOW-UP PROBLEM 2.15** What is the molecular (or formula) mass of **(a)** hydrogen peroxide; **(b)** cesium chloride; **(c)** sulfuric acid; **(d)** potassium sulfate?

## Sample Problem 2.16 Using Molecular Depictions to Determine Formula, Name, and Mass

**Problem** Each scene represents a binary compound. Determine its formula, name, and molecular (or formula) mass.




**Plan** Each of the compounds contains only two elements, so to find the formula, we find the simplest whole-number ratio of one atom to the other. From the formula, we determine the name and the molecular (or formula) mass.

**Solution** **(a)** There is one brown sphere (sodium) for each green sphere (fluorine), so the formula is NaF. A metal and nonmetal form an ionic compound, in which the metal is named first: sodium fluoride.

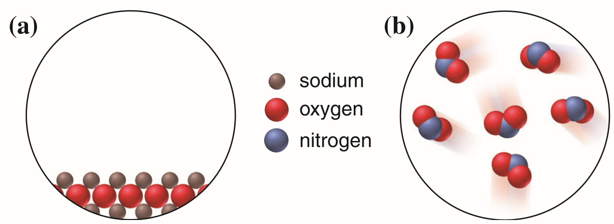
$$\text{Formula mass} = (1 \times \text{atomic mass of Na}) + (1 \times \text{atomic mass of F})$$
$$= 22.99 \text{ amu} + 19.00 \text{ amu} = 41.99 \text{ amu}$$

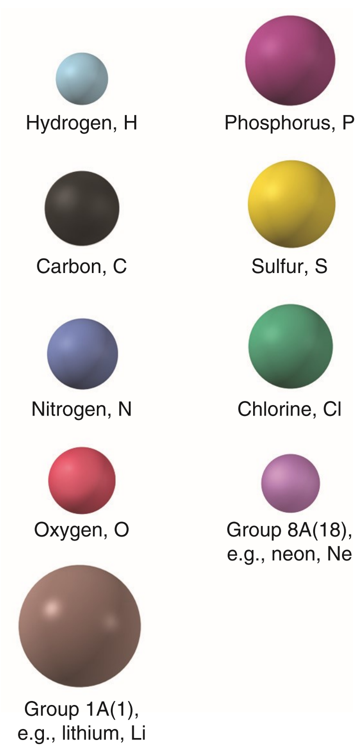
**(b)** There are three green spheres (fluorine) for each blue sphere (nitrogen), so the formula is $NF_3$. Two nonmetals form a covalent compound. Nitrogen has a lower group number, so it is named first: nitrogen trifluoride.

$$\text{Molecular mass} = (1 \times \text{atomic mass of N}) + (3 \times \text{atomic mass of F})$$
$$= 14.01 \text{ amu} + (3 \times 19.00) = 71.01 \text{ amu}$$

**Check** **(a)** For binary ionic compounds, we predict ionic charges from the periodic table. Na forms a 1+ ion, and F forms a 1− ion, so the charges balance with one $Na^+$ per $F^-$. Also, ionic compounds are solids, consistent with the picture. **(b)** Covalent compounds often occur as individual molecules, as in the picture. Rounding in (a) gives $25 + 20 = 45$; in (b), we get $15 + (3 \times 20) = 75$, so there are no large errors.

**FOLLOW-UP PROBLEM 2.16** Each scene represents a binary compound. Determine its name, formula, and molecular (or formula) mass.




## Representing Molecules with Formulas and Models

In order to represent objects too small to see, chemists employ a variety of formulas and models. Each conveys different information, as shown for water below:

- A **molecular formula** uses element symbols and, often, numerical subscripts to give the *actual* number of atoms of each element in a molecule of the compound. (Recall that, for ionic compounds, the *formula unit* gives the *relative* number of each type of ion.) The molecular formula of water is $H_2O$: there are two H atoms and one O atom in each molecule.

  $H_2O$

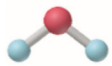
- A **structural formula** shows the relative placement and connections of atoms in the molecule. It uses symbols for the atoms *and* either a pair of dots *(electron-dot formula)* or a line *(bond-line formula)* to show the electron pairs in bonds between the atoms. In water, each H atom is bonded to the O atom, but not to the other H atom.

  H:O:H

  H–O–H

In models, colored balls represent atoms *(see margin)*.

- A *ball-and-stick model* shows atoms as balls and bonds as sticks, and the angles between the bonds are accurate. Note that water is a bent molecule (with a bond angle of 104.5°). This type of model exaggerates the distance between bonded atoms.



- A *space-filling model* is an accurately scaled-up image of the molecule, but bonds are not shown, and it can be difficult to see each atom in a complex molecule.

## Summary of Section 2.8

- An ionic compound is named with cation first and anion second. For metals that can form more than one ion, the charge is shown with a Roman numeral.
- Oxoanions have suffixes, and sometimes prefixes, attached to the root of the element name to indicate the number of oxygen atoms.
- Names of hydrates have a numerical prefix indicating the number of associated water molecules.
- Acid names are based on anion names.
- For binary covalent compounds, the first word of the name is the element farther left or lower down in the periodic table, and prefixes show the numbers of each atom.
- The molecular (or formula) mass of a compound is the sum of the atomic masses.
- Chemical formulas give the number of atoms (molecular) or the arrangement of atoms (structural) of one unit of a compound.
- Molecular models convey information about bond angles (ball-and-stick) and relative atomic sizes and distances between atoms (space-filling).

## 2.9 • CLASSIFICATION OF MIXTURES

In the natural world, *matter usually occurs as mixtures.* Air, seawater, soil, and organisms are all complex mixtures of elements and compounds. There are two broad classes of mixtures:

- A **heterogeneous mixture** has one or more visible boundaries between the components. Thus, its composition is *not* uniform, but rather varies from one region to another. Many rocks are heterogeneous, having individual grains of different minerals. In some heterogeneous mixtures, such as milk and blood, the boundaries can be seen only with a microscope.
- A **homogeneous mixture** (or **solution**) has no visible boundaries because the components are individual atoms, ions, or molecules. Thus, its composition *is* uniform. A mixture of sugar dissolved in water is homogeneous, for example, because the sugar molecules and water molecules are uniformly intermingled on the molecular level. We have no way to tell visually whether a sample of matter is a substance (element or compound) or a homogeneous mixture.

Although we usually think of solutions as liquid, they can exist in all three physical states. For example, air is a gaseous solution of mostly oxygen and nitrogen molecules, and wax is a solid solution of several fatty substances. Solutions in water, called **aqueous solutions,** are especially important in the chemistry lab and comprise a major portion of the environment and of all organisms.

Recall that mixtures differ from compounds in three major ways:

1. The proportions of the components can vary.
2. The individual properties of the components are observable.
3. The components can be separated by physical means.

The difference between a mixture and a compound is well illustrated using iron and sulfur as components (Figure 2.17). Any proportion of iron metal filings and powdered sulfur forms a mixture. The iron can be separated from the sulfur with a magnet. But if we heat the container strongly, the components combine in fixed proportions by mass to form the compound iron(II) sulfide (FeS). The magnet can no longer remove the iron because it exists as $Fe^{2+}$ ions chemically bound to $S^{2-}$ ions.

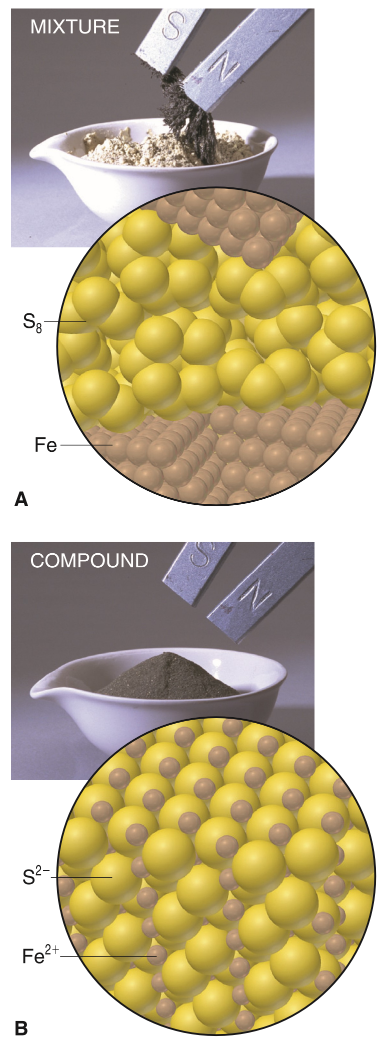



**Figure 2.17** The distinction between mixtures and compounds. **A,** A *mixture* of iron and sulfur consists of the two elements. **B,** The *compound* iron(II) sulfide consists of an array of $Fe^{2+}$ and $S^{2-}$ ions.

## An Overview of the Components of Matter

Understanding matter at the observable and atomic scales is the essence of chemistry. Figure 2.18 is a visual overview of many key terms and ideas in this chapter.

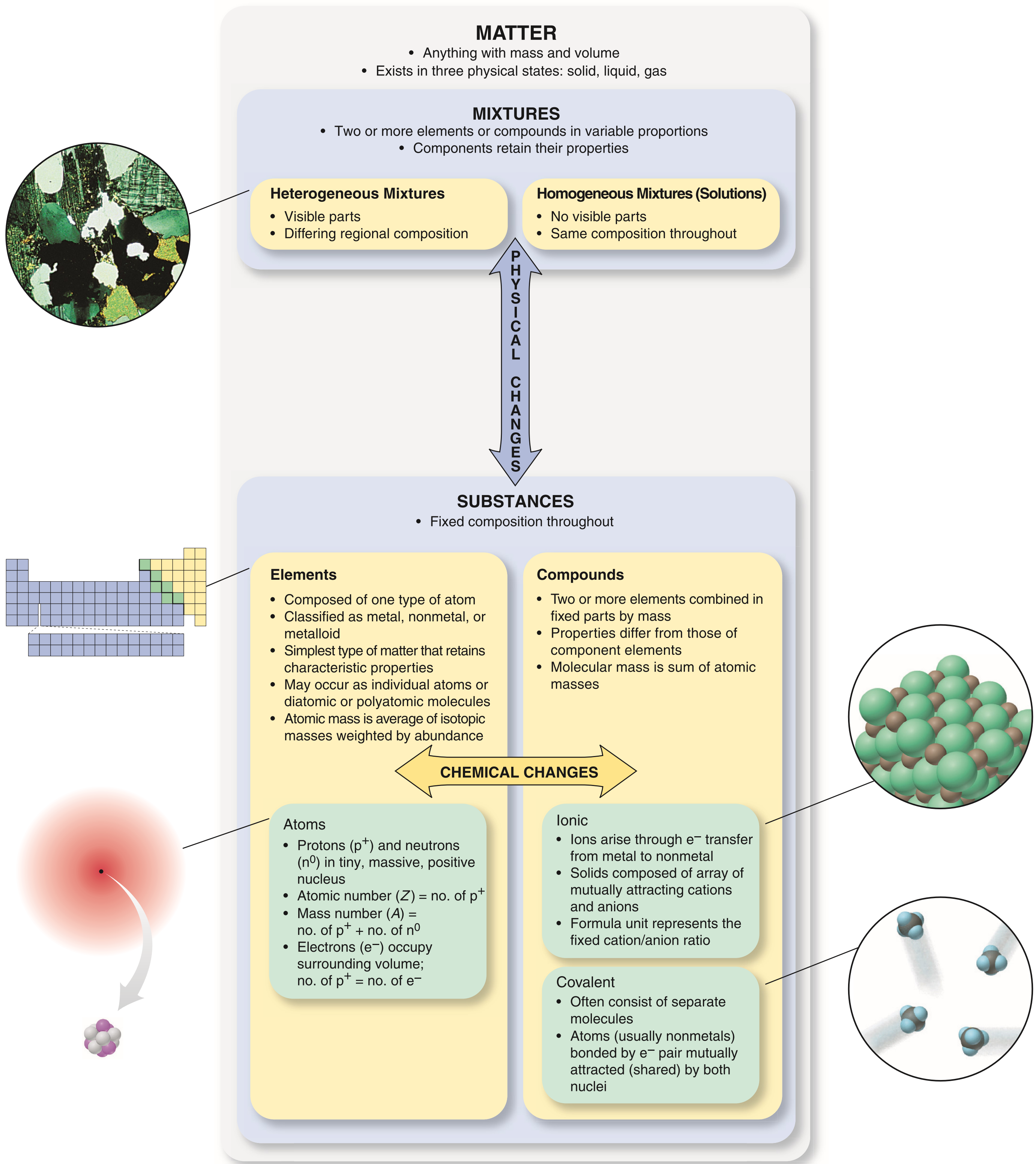



**Figure 2.18** The classification of matter from a chemical point of view.

## Summary of Section 2.9

- Heterogeneous mixtures have visible boundaries between the components.
- Homogeneous mixtures (solutions) have no visible boundaries because mixing occurs at the molecular level. They can occur in any physical state.
- Components of mixtures (unlike those of compounds) can have variable proportions, can be separated physically, and retain their properties.

# CHAPTER REVIEW GUIDE

The following sections provide many aids to help you study this chapter. (Numbers in parentheses refer to pages, unless noted otherwise.)

## Learning Objectives

These are concepts and skills to review after studying this chapter.

Related section (§), sample problem (SP), and upcoming end-of-chapter problem (EP) numbers are listed in parentheses.

1. Define the characteristics of the three types of matter—element, compound, and mixture—on the macroscopic and atomic levels (§2.1) (SP 2.1) (EPs 2.1–2.5)
2. Understand the laws of mass conservation, definite composition, and multiple proportions; use the mass ratio of element-to-compound to find the mass of an element in a compound (§2.2) (SP 2.2) (EPs 2.6–2.20, 2.80)
3. Understand Dalton's atomic theory and how it explains the mass laws (§2.3) (SP 2.3) (EP 2.21)
4. Describe the results of the key experiments by Thomson, Millikan, and Rutherford concerning atomic structure (§2.4) (EPs 2.22–2.24)
5. Explain the structure of the atom, the main features of the subatomic particles, and the significance of isotopes; use atomic notation to express the subatomic makeup of an isotope; calculate the atomic mass of an element from its isotopic composition (§2.5) (SPs 2.4, 2.5) (EPs 2.25–2.37)
6. Describe the format of the periodic table and the general location and characteristics of metals, metalloids, and nonmetals (§2.6) (EPs 2.38–2.44)
7. Explain the essential features of ionic and covalent compounds and distinguish between them; predict the monatomic ion formed from a main-group element (§2.7) (SP 2.6) (EPs 2.45–2.57)
8. Name, write the formula, and calculate the molecular (or formula) mass of ionic and binary covalent compounds (§2.8) (SPs 2.7–2.16) (EPs 2.58–2.79, 2.81)
9. Describe the types of mixtures and their properties (§2.9) (EPs 2.82–2.86)

## Key Terms

These important terms appear in boldface in the chapter and are defined again in the Glossary.

**Section 2.1**
substance (33)
element (33)
molecule (33)
compound (34)
mixture (34)

**Section 2.2**
law of mass conservation (35)
law of definite (or constant) composition (36)
fraction by mass (mass fraction) (36)
percent by mass (mass percent, mass %) (36)
law of multiple proportions (37)

**Section 2.3**
atom (38)

**Section 2.4**
cathode ray (40)
nucleus (42)

**Section 2.5**
proton ($p^+$) (43)
neutron ($n^0$) (43)
electron ($e^-$) (43)
atomic number ($Z$) (43)
mass number ($A$) (43)
atomic symbol (43)
isotope (44)
atomic mass unit (amu) (44)
dalton (Da) (44)
mass spectrometry (45)
isotopic mass (45)
atomic mass (45)

**Section 2.6**
periodic table of the elements (47)
period (47)
group (47)
metal (47)
nonmetal (48)
metalloid (semimetal) (48)

**Section 2.7**
ionic compound (49)
covalent compound (49)
chemical bond (49)
ion (49)
binary ionic compound (49)
cation (49)
anion (49)
monatomic ion (49)
covalent bond (51)
polyatomic ion (52)

**Section 2.8**
chemical formula (53)
formula unit (54)
oxoanion (56)
hydrate (56)
binary covalent compound (58)
molecular mass (59)
formula mass (59)
molecular formula (60)
structural formula (60)

**Section 2.9**
heterogeneous mixture (61)
homogeneous mixture (solution) (61)
aqueous solution (61)

## Key Equations and Relationships

Numbered and screened concepts are listed for you to refer to or memorize.

**2.1** Finding the mass of an element in a given mass of compound (36):

$$\text{Mass of element in sample} = \text{mass of compound in sample} \times \frac{\text{mass of element}}{\text{mass of compound}}$$

**2.2** Calculating the number of neutrons in an atom (44):

$$\text{Number of neutrons} = \text{mass number} - \text{atomic number}$$

or

$$N = A - Z$$

**2.3** Determining the molecular mass of a formula unit of a compound (59):

$$\text{Molecular mass} = \text{sum of atomic masses}$$

## BRIEF SOLUTIONS TO FOLLOW-UP PROBLEMS

Compare your own solutions to these calculation steps and answers.

**2.1** There are two types of particles reacting (left circle), one with two blue atoms and the other with two orange; the depiction shows a mixture of two elements. In the product (right circle), all the particles have one blue atom and one orange; this is a compound.

**2.2** Mass (t) of pitchblende

$$= 2.3\ \cancel{\text{t uranium}} \times \frac{84.2\ \text{t pitchblende}}{71.4\ \cancel{\text{t uranium}}} = 2.7\ \text{t pitchblende}$$

Mass (t) of oxygen

$$= 2.7\ \cancel{\text{t pitchblende}} \times \frac{(84.2 - 71.4\ \text{t oxygen})}{84.2\ \cancel{\text{t pitchblende}}} = 0.41\ \text{t oxygen}$$

**2.3** Sample B. Two bromine-fluorine compounds appear. In one, there are three fluorine atoms for each bromine; in the other, there is one fluorine for each bromine. Therefore, in the two compounds, the ratio of fluorines combining with one bromine is 3/1.

**2.4** (a) $5p^+$, $6n^0$, $5e^-$; Q = B
(b) $20p^+$, $21n^0$, $20e^-$; R = Ca
(c) $53p^+$, $78n^0$, $53e^-$; X = I

**2.5** $10.0129x + [11.0093(1 - x)] = 10.81$; $0.9964x = 0.1993$; $x = 0.2000$ and $1 - x = 0.8000$; % abundance of $^{10}B = 20.00\%$; % abundance of $^{11}B = 80.00\%$

**2.6** (a) $S^{2-}$; (b) $Rb^+$; (c) $Ba^{2+}$

**2.7** (a) Zinc [Group 2B(12)] and oxygen [Group 6A(16)]
(b) Silver [Group 1B(11)] and bromine [Group 7A(17)]
(c) Lithium [Group 1A(1)] and chlorine [Group 7A(17)]
(d) Aluminum [Group 3A(13)] and sulfur [Group 6A(16)]

**2.8** (a) $ZnO$; (b) $AgBr$; (c) $LiCl$; (d) $Al_2S_3$

**2.9** (a) $PbO_2$; (b) copper(I) sulfide (cuprous sulfide); (c) iron(II) bromide (ferrous bromide); (d) $HgCl_2$

**2.10** (a) $Cu(NO_3)_2{\cdot}3H_2O$; (b) $Zn(OH)_2$; (c) lithium cyanide

**2.11** (a) $(NH_4)_3PO_4$; ammonium is $NH_4^+$ and phosphate is $PO_4^{3-}$.
(b) $Al(OH)_3$; parentheses are needed around the polyatomic ion $OH^-$.
(c) Magnesium hydrogen carbonate; $Mg^{2+}$ is magnesium and can have only a 2+ charge, so the Roman numeral II is not needed; $HCO_3^-$ is hydrogen carbonate (or bicarbonate).
(d) Chromium(III) nitrate; the *-ic* ending is not used with Roman numerals; $NO_3^-$ is nitr*ate*.
(e) Calcium nitrite; $Ca^{2+}$ is calcium and $NO_2^-$ is nitr*ite*.

**2.12** (a) $HClO_3$; (b) hydrofluoric acid; (c) $CH_3COOH$ (or $HC_2H_3O_2$); (d) $H_2SO_3$; (e) hypobromous acid

**2.13** (a) Sulfur trioxide; (b) silicon dioxide; (c) $N_2O$; (d) $SeF_6$

**2.14** (a) Disulfur dichloride; the *-ous* suffix is not used.
(b) NO; the name indicates one nitrogen.
(c) Bromine trichloride; Br is in a higher period in Group 7A(17), so it is named first.

**2.15** (a) $H_2O_2$, 34.02 amu; (b) CsCl, 168.4 amu; (c) $H_2SO_4$, 98.09 amu; (d) $K_2SO_4$, 174.27 amu

**2.16** (a) $Na_2O$. This is an ionic compound, so the name is sodium oxide.

Formula mass

$$= (2 \times \text{atomic mass of Na}) + (1 \times \text{atomic mass of O})$$
$$= (2 \times 22.99\ \text{amu}) + 16.00\ \text{amu} = 61.98\ \text{amu}$$

(b) $NO_2$. This is a covalent compound, and N has the lower group number, so the name is nitrogen dioxide.

Molecular mass

$$= (1 \times \text{atomic mass of N}) + (2 \times \text{atomic mass of O})$$
$$= 14.01\ \text{amu} + (2 \times 16.00\ \text{amu}) = 46.01\ \text{amu}$$

# PROBLEMS

Problems with **colored** numbers are answered in Appendix E. Sections match the text and provide the numbers of relevant sample problems. Bracketed problems are grouped in pairs (indicated by a short rule) that cover the same concept. Comprehensive Problems are based on material from any section or previous chapter.

## Elements, Compounds, and Mixtures: An Atomic Overview

(Sample Problem 2.1)

**2.1** What is the key difference between an element and a compound?

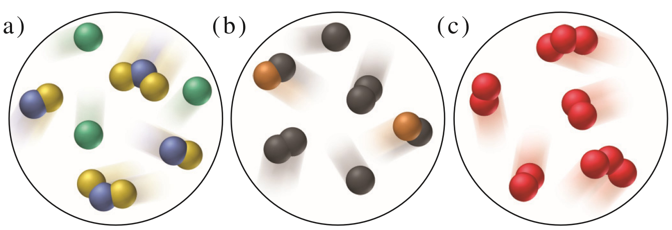
**2.2** List two differences between a compound and a mixture.

**2.3** Which of the following are pure substances? Explain.
(a) Calcium chloride, used to melt ice on roads, consists of two elements, calcium and chlorine, in a fixed mass ratio.
(b) Sulfur consists of sulfur atoms combined into octatomic molecules.
(c) Baking powder, a leavening agent, contains 26% to 30% sodium hydrogen carbonate and 30% to 35% calcium dihydrogen phosphate by mass.
(d) Cytosine, a component of DNA, consists of H, C, N, and O atoms bonded in a specific arrangement.

**2.4** Classify each substance in Problem 2.3 as an element, compound, or mixture, and explain your answers.

**2.5** Each scene below represents a mixture. Describe each one in terms of the number(s) of elements and/or compounds present.

(a) (b) (c)



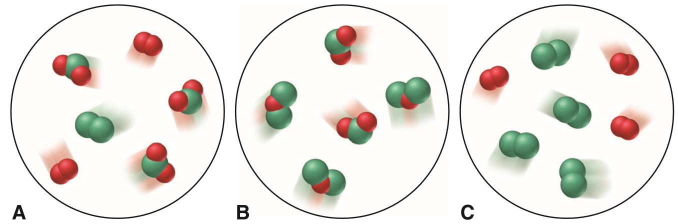
## The Observations That Led to an Atomic View of Matter

(Sample Problem 2.2)

**2.6** To which classes of matter—element, compound, and/or mixture—do the following apply: (a) law of mass conservation; (b) law of definite composition; (c) law of multiple proportions?

**2.7** Identify the mass law that each of the following observations demonstrates, and explain your reasoning:
(a) A sample of potassium chloride from Chile contains the same percent by mass of potassium as one from Poland.
(b) A flashbulb contains magnesium and oxygen before use and magnesium oxide afterward, but its mass does not change.
(c) Arsenic and oxygen form one compound that is 65.2 mass % arsenic and another that is 75.8 mass % arsenic.

**2.8** Which of the following scenes illustrate(s) the fact that compounds of chlorine (green) and oxygen (red) exhibit the law of multiple proportions? Name the compounds.



A B C

**2.9** (a) Does the percent by mass of each element in a compound depend on the amount of compound? Explain.
(b) Does the mass of each element in a compound depend on the amount of compound? Explain.

**2.10** Does the percent by mass of each element in a compound depend on the amount of that element used to make the compound? Explain.

**2.11** State the mass law(s) demonstrated by the following experimental results, and explain your reasoning:

Experiment 1: A student heats 1.00 g of a blue compound and obtains 0.64 g of a white compound and 0.36 g of a colorless gas.

Experiment 2: A second student heats 3.25 g of the same blue compound and obtains 2.08 g of a white compound and 1.17 g of a colorless gas.

**2.12** State the mass law(s) demonstrated by the following experimental results, and explain your reasoning:

Experiment 1: A student heats 1.27 g of copper and 3.50 g of iodine to produce 3.81 g of a white compound; 0.96 g of iodine remains.

Experiment 2: A second student heats 2.55 g of copper and 3.50 g of iodine to form 5.25 g of a white compound, and 0.80 g of copper remains.

**2.13** Fluorite, a mineral of calcium, is a compound of the metal with fluorine. Analysis shows that a 2.76-g sample of fluorite contains 1.42 g of calcium. Calculate the (a) mass of fluorine in the sample; (b) mass fractions of calcium and fluorine in fluorite; (c) mass percents of calcium and fluorine in fluorite.

**2.14** Galena, a mineral of lead, is a compound of the metal with sulfur. Analysis shows that a 2.34-g sample of galena contains 2.03 g of lead. Calculate the (a) mass of sulfur in the sample; (b) mass fractions of lead and sulfur in galena; (c) mass percents of lead and sulfur in galena.

**2.15** A compound of copper and sulfur contains 88.39 g of metal and 44.61 g of nonmetal. How many grams of copper are in 5264 kg of compound? How many grams of sulfur?

**2.16** A compound of iodine and cesium contains 63.94 g of metal and 61.06 g of nonmetal. How many grams of cesium are in 38.77 g of compound? How many grams of iodine?

**2.17** Show, with calculations, how the following data illustrate the law of multiple proportions:
Compound 1: 47.5 mass % sulfur and 52.5 mass % chlorine
Compound 2: 31.1 mass % sulfur and 68.9 mass % chlorine

**2.18** Show, with calculations, how the following data illustrate the law of multiple proportions:
Compound 1: 77.6 mass % xenon and 22.4 mass % fluorine
Compound 2: 63.3 mass % xenon and 36.7 mass % fluorine

**2.19** Dolomite is a carbonate of magnesium and calcium. Analysis shows that 7.81 g of dolomite contains 1.70 g of Ca. Calculate the mass percent of Ca in dolomite. On the basis of the mass percent of Ca, and neglecting all other factors, which is the richer source of Ca, dolomite or fluorite (see Problem 2.13)?

**2.20** The mass percent of sulfur in a sample of coal is a key factor in the environmental impact of the coal because the sulfur combines with oxygen when the coal is burned and the oxide can then be incorporated into acid rain. Which of the following coals would have the smallest environmental impact?

| | Mass (g) of Sample | Mass (g) of Sulfur in Sample |
|---|---|---|
| Coal A | 378 | 11.3 |
| Coal B | 495 | 19.0 |
| Coal C | 675 | 20.6 |

## Dalton's Atomic Theory

(Sample Problem 2.3)

**2.21** Use Dalton's theory to explain why potassium nitrate from India or Italy has the same mass percents of K, N, and O.

## The Observations That Led to the Nuclear Atom Model

**2.22** Thomson was able to determine the mass/charge ratio of the electron but not its mass. How did Millikan's experiment allow determination of the electron's mass?

**2.23** The following charges on individual oil droplets were obtained during an experiment similar to Millikan's. Determine a charge for the electron (in C, coulombs), and explain your answer: $-3.204\times10^{-19}$ C; $-4.806\times10^{-19}$ C; $-8.010\times10^{-19}$ C; $-1.442\times10^{-18}$ C.

**2.24** When Rutherford's coworkers bombarded gold foil with α particles, they obtained results that overturned the existing (Thomson) model of the atom. Explain.

## The Atomic Theory Today

(Sample Problems 2.4 and 2.5)

**2.25** Choose the correct answer. The difference between the mass number of an isotope and its atomic number is (a) directly related to the identity of the element; (b) the number of electrons; (c) the number of neutrons; (d) the number of isotopes.

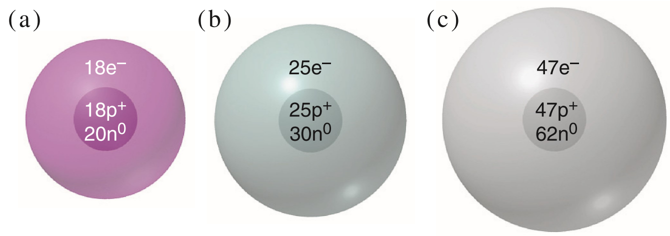
**2.26** Argon has three naturally occurring isotopes, $^{36}$Ar, $^{38}$Ar, and $^{40}$Ar. What is the mass number of each? How many protons, neutrons, and electrons are present in each?

**2.27** Chlorine has two naturally occurring isotopes, $^{35}$Cl and $^{37}$Cl. What is the mass number of each isotope? How many protons, neutrons, and electrons are present in each?

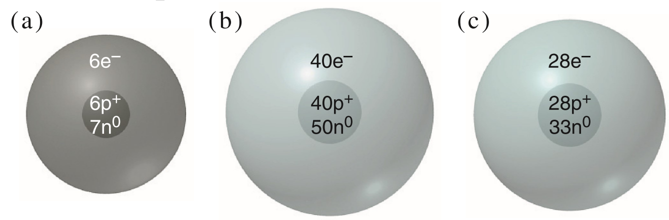
**2.28** Do both members of the following pairs have the same number of protons? Neutrons? Electrons?
(a) $^{16}_{8}$O and $^{17}_{8}$O (b) $^{40}_{18}$Ar and $^{41}_{19}$K (c) $^{60}_{27}$Co and $^{60}_{28}$Ni
Which pair(s) consist(s) of atoms with the same $Z$ value? $N$ value? $A$ value?

**2.29** Do both members of the following pairs have the same number of protons? Neutrons? Electrons?
(a) $^{3}_{1}$H and $^{3}_{2}$He (b) $^{14}_{6}$C and $^{15}_{7}$N (c) $^{19}_{9}$F and $^{18}_{9}$F
Which pair(s) consist(s) of atoms with the same $Z$ value? $N$ value? $A$ value?

**2.30** Write the $^{A}_{Z}$X notation for each atomic depiction:




**2.31** Write the $^{A}_{Z}$X notation for each atomic depiction:




**2.32** Draw atomic depictions similar to those in Problem 2.30 for (a) $^{48}_{22}$Ti; (b) $^{79}_{34}$Se; (c) $^{11}_{5}$B.

**2.33** Draw atomic depictions similar to those in Problem 2.30 for (a) $^{207}_{82}$Pb; $^{9}_{4}$Be; $^{75}_{33}$As.

**2.34** Gallium has two naturally occurring isotopes, $^{69}$Ga (isotopic mass = 68.9256 amu, abundance = 60.11%) and $^{71}$Ga (isotopic mass = 70.9247 amu, abundance = 39.89%). Calculate the atomic mass of gallium.

**2.35** Magnesium has three naturally occurring isotopes, $^{24}$Mg (isotopic mass = 23.9850 amu, abundance = 78.99%), $^{25}$Mg (isotopic mass = 24.9858 amu, abundance = 10.00%), and $^{26}$Mg (isotopic mass = 25.9826 amu, abundance = 11.01%). Calculate the atomic mass of magnesium.

**2.36** Chlorine has two naturally occurring isotopes, $^{35}$Cl (isotopic mass = 34.9689 amu) and $^{37}$Cl (isotopic mass = 36.9659 amu). If chlorine has an atomic mass of 35.4527 amu, what is the percent abundance of each isotope?

**2.37** Copper has two naturally occurring isotopes, $^{63}$Cu (isotopic mass = 62.9396 amu) and $^{65}$Cu (isotopic mass = 64.9278 amu). If copper has an atomic mass of 63.546 amu, what is the percent abundance of each isotope?

## Elements: A First Look at the Periodic Table

**2.38** Correct each of the following statements:
(a) In the modern periodic table, the elements are arranged in order of increasing atomic mass.
(b) Elements in a period have similar chemical properties.
(c) Elements can be classified as either metalloids or nonmetals.

**2.39** What class of elements lies along the “staircase” line in the periodic table? How do the properties of these elements compare with those of metals and nonmetals?

**2.40** What are some characteristic properties of elements to the left of the elements along the “staircase”? To the right?

**2.41** Give the name, atomic symbol, and group number of the element with each $Z$ value, and classify it as a metal, metalloid, or nonmetal:
(a) $Z = 32$ (b) $Z = 15$ (c) $Z = 2$
(d) $Z = 3$ (e) $Z = 42$

**2.42** Give the name, atomic symbol, and group number of the element with each $Z$ value, and classify it as a metal, metalloid, or nonmetal:
(a) $Z = 33$ (b) $Z = 20$ (c) $Z = 35$
(d) $Z = 19$ (e) $Z = 13$

**2.43** Fill in the blanks:
(a) The symbol and atomic number of the heaviest alkaline earth metal are ______ and ______.
(b) The symbol and atomic number of the lightest metalloid in Group 4A(14) are ______ and ______.
(c) Group 1B(11) consists of the *coinage metals*. The symbol and atomic mass of the coinage metal whose atoms have the fewest electrons are ______ and ______.
(d) The symbol and atomic mass of the halogen in Period 4 are ______ and ______.

**2.44** Fill in the blanks:
(a) The symbol and atomic number of the heaviest nonradioactive noble gas are ______ and ______.
(b) The symbol and group number of the Period 5 transition element whose atoms have the fewest protons are ______ and ______.
(c) The elements in Group 6A(16) are sometimes called the *chalcogens*. The symbol and atomic number of the first metallic chalcogen are ______ and ______.
(d) The symbol and number of protons of the Period 4 alkali metal atom are ______ and ______.

## Compounds: Introduction to Bonding

(Sample Problem 2.6)

**2.45** Describe the type and nature of the bonding that occurs between reactive metals and nonmetals.

**2.46** Describe the type and nature of the bonding that often occurs between two nonmetals.

**2.47** Given that the ions in LiF and in MgO are of similar size, which compound has stronger ionic bonding? Use Coulomb’s law in your explanation.

**2.48** Describe the formation of solid magnesium chloride ($MgCl_2$) from large numbers of magnesium and chlorine atoms.

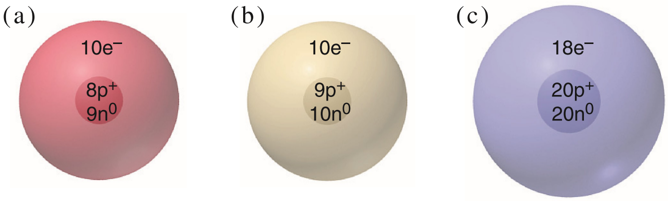
**2.49** Does potassium nitrate ($KNO_3$) incorporate ionic bonding, covalent bonding, or both? Explain.

**2.50** What monatomic ions do potassium ($Z = 19$) and iodine ($Z = 53$) form?

**2.51** What monatomic ions do barium ($Z = 56$) and selenium ($Z = 34$) form?

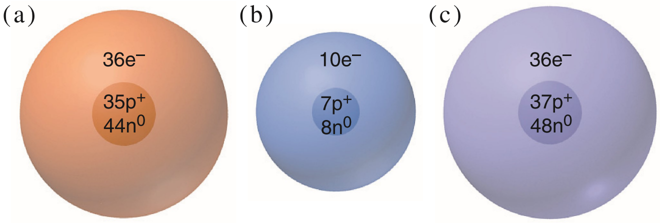
**2.52** For each ionic depiction, give the name of the parent atom, its mass number, and its group and period numbers:

(a) (b) (c)




**62.53** For each ionic depiction, give the name of the parent atom, its mass number, and its group and period numbers:

(a) (b) (c)




**2.54** An ionic compound forms when lithium ($Z = 3$) reacts with oxygen ($Z = 8$). If a sample of the compound contains $8.4\times10^{21}$ lithium ions, how many oxide ions does it contain?

**2.55** An ionic compound forms when calcium ($Z = 20$) reacts with iodine ($Z = 53$). If a sample of the compound contains $7.4\times10^{21}$ calcium ions, how many iodide ions does it contain?

**2.56** The radii of the sodium and potassium ions are 102 pm and 138 pm, respectively. Which compound has stronger ionic attractions, sodium chloride or potassium chloride?

**2.57** The radii of the lithium and magnesium ions are 76 pm and 72 pm, respectively. Which compound has stronger ionic attractions, lithium oxide or magnesium oxide?

## Formulas, Names, and Masses of Compounds

(Sample Problems 2.7 to 2.16)

**2.58** How is a structural formula similar to a molecular formula? How is it different?

**2.59** Consider a mixture of 10 billion $O_2$ molecules and 10 billion $H_2$ molecules. In what way is this mixture similar to a sample containing 10 billion hydrogen peroxide ($H_2O_2$) molecules? In what way is it different?

**2.60** Write a formula for each of the following compounds:
(a) Hydrazine, a rocket fuel, consists of two nitrogen atoms and four hydrogen atoms.
(b) Glucose, a sugar, consists of six carbon atoms, twelve hydrogen atoms, and six oxygen atoms.

**2.61** Write a formula for each of the following compounds:
(a) Ethylene glycol, car antifreeze, consists of two carbon atoms, six hydrogen atoms, and two oxygen atoms.
(b) Peroxodisulfuric acid, a compound used to make bleaching agents, consists of two hydrogen atoms, two sulfur atoms, and eight oxygen atoms.

**2.62** Give the name and formula of the compound formed from the following elements:
(a) Sodium and nitrogen
(b) Oxygen and strontium
(c) Aluminum and chlorine

**2.63** Give the name and formula of the compound formed from the following elements:
(a) Cesium and bromine
(b) Sulfur and barium
(c) Calcium and fluorine

**2.64** Give the name and formula of the compound formed from the following elements:
(a) $_{12}L$ and $_{9}M$ (b) $_{30}L$ and $_{16}M$ (c) $_{17}L$ and $_{38}M$

**2.65** Give the name and formula of the compound formed from the following elements:
(a) $_{37}Q$ and $_{35}R$ (b) $_{8}Q$ and $_{13}R$ (c) $_{20}Q$ and $_{53}R$

**2.66** Give the systematic names for the formulas or the formulas for the names: (a) tin(IV) chloride; (b) $FeBr_3$; (c) cuprous bromide; (d) $Mn_2O_3$.

**2.67** Give the systematic names for the formulas or the formulas for the names: (a) $Na_2HPO_4$; (b) potassium carbonate dihydrate; (c) $NaNO_2$; (d) ammonium perchlorate.

**2.68** Correct each of the following formulas:
(a) Barium oxide is $BaO_2$
(b) Iron(II) nitrate is $Fe(NO_3)_3$
(c) Magnesium sulfide is $MnSO_3$

**2.69** Correct each of the following names:
(a) CuI is cobalt(II) iodide
(b) $Fe(HSO_4)_3$ is iron(II) sulfate
(c) $MgCr_2O_7$ is magnesium dichromium heptoxide

**2.70** Give the name and formula for the acid derived from each of the following anions:
(a) hydrogen sulfate (b) $IO_3^-$ (c) cyanide (d) $HS^-$

**2.71** Give the name and formula for the acid derived from each of the following anions:
(a) perchlorate (b) $NO_3^-$ (c) bromite (d) $F^-$

**2.72** Give the name and formula of the compound whose molecules consist of two sulfur atoms and four fluorine atoms.

**2.73** Give the name and formula of the compound whose molecules consist of two chlorine atoms and one oxygen atom.

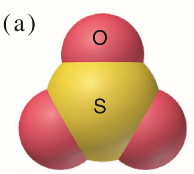
**2.74** Give the number of atoms of the specified element in a formula unit of each of the following compounds, and calculate the molecular (formula) mass:
(a) Oxygen in aluminum sulfate, $Al_2(SO_4)_3$
(b) Hydrogen in ammonium hydrogen phosphate, $(NH_4)_2HPO_4$
(c) Oxygen in the mineral azurite, $Cu_3(OH)_2(CO_3)_2$

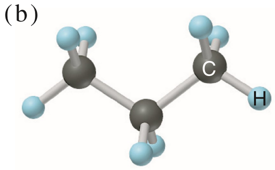
**2.75** Give the number of atoms of the specified element in a formula unit of each of the following compounds, and calculate the molecular (formula) mass:
(a) Hydrogen in ammonium benzoate, $C_6H_5COONH_4$
(b) Nitrogen in hydrazinium sulfate, $N_2H_6SO_4$
(c) Oxygen in the mineral leadhillite, $Pb_4SO_4(CO_3)_2(OH)_2$

**2.76** Write the formula of each compound, and determine its molecular (formula) mass: (a) ammonium sulfate; (b) sodium dihydrogen phosphate; (c) potassium bicarbonate.

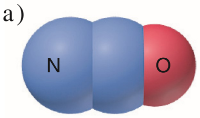
**2.77** Write the formula of each compound, and determine its molecular (formula) mass: (a) sodium dichromate; (b) ammonium perchlorate; (c) magnesium nitrite trihydrate.

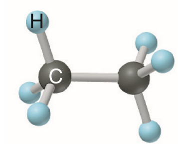
**2.78** Give the formula, name, and molecular mass of the following molecules:

(a)



(b)



**2.79** Give the formula, name, and molecular mass of the following molecules:

(a)



(b)



**2.80** You are working in the laboratory preparing sodium chloride. Consider the following results for three preparations of the compound:

Case 1: 39.34 g Na + 60.66 g $Cl_2$ $\longrightarrow$ 100.00 g NaCl
Case 2: 39.34 g Na + 70.00 g $Cl_2$ $\longrightarrow$ 100.00 g NaCl + 9.34 g $Cl_2$
Case 3: 50.00 g Na + 50.00 g $Cl_2$ $\longrightarrow$ 82.43 g NaCl + 17.57 g Na

Explain these results in terms of the laws of conservation of mass and definite composition.

**2.81** Before the use of systematic names, many compounds had common names. Give the systematic name for each of the following:
(a) Blue vitriol, $CuSO_4 \cdot 5H_2O$
(b) Slaked lime, $Ca(OH)_2$
(c) Oil of vitriol, $H_2SO_4$
(d) Washing soda, $Na_2CO_3$
(e) Muriatic acid, HCl
(f) Epsom salt, $MgSO_4 \cdot 7H_2O$
(g) Chalk, $CaCO_3$
(h) Dry ice, $CO_2$
(i) Baking soda, $NaHCO_3$
(j) Lye, NaOH

## Classification of Mixtures

**2.82** In what main way is separating the components of a mixture different from separating the components of a compound?

**2.83** What is the difference between a homogeneous and a heterogeneous mixture?

**2.84** Is a solution a homogeneous or a heterogeneous mixture? Give an example of an aqueous solution.

**2.85** Classify each of the following as a compound, a homogeneous mixture, or a heterogeneous mixture: (a) distilled water; (b) gasoline; (c) beach sand; (d) wine; (e) air.

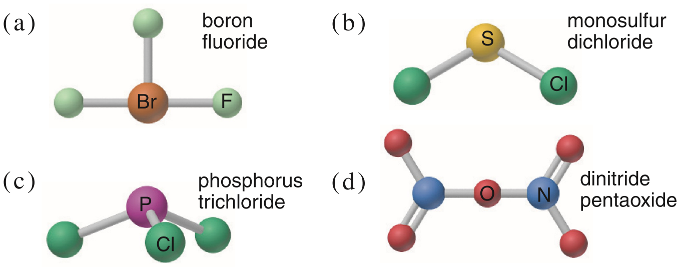
**2.86** Classify each of the following as a compound, a homogeneous mixture, or a heterogeneous mixture: (a) orange juice; (b) vegetable soup; (c) cement; (d) calcium sulfate; (e) tea.

## Comprehensive Problems

**2.87** Helium is the lightest noble gas and the second most abundant element (after hydrogen) in the universe.
(a) The radius of a helium atom is $3.1 \times 10^{-11}$ m; the radius of its nucleus is $2.5 \times 10^{-15}$ m. What fraction of the spherical atomic volume is occupied by the nucleus ($V$ of a sphere $= \frac{4}{3}\pi r^3$)?

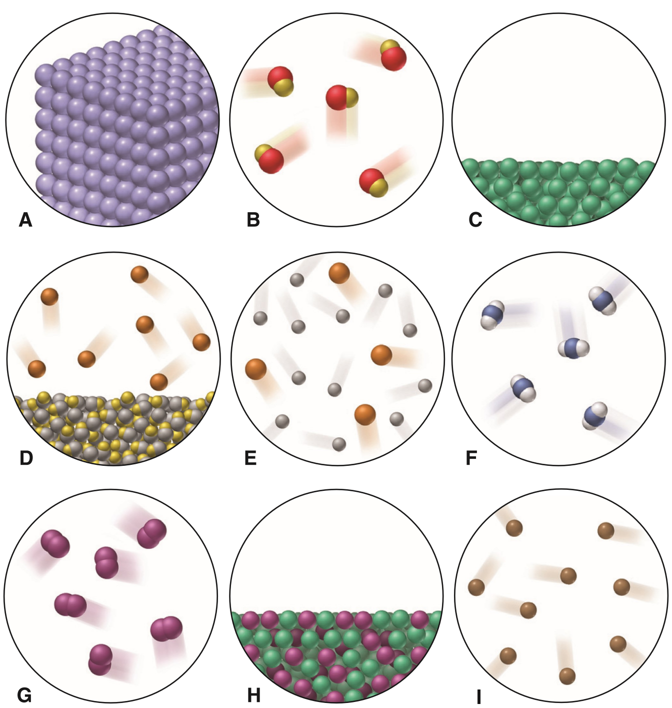
(b) The mass of a helium-4 atom is $6.64648 \times 10^{-24}$ g, and each of its two electrons has a mass of $9.10939 \times 10^{-28}$ g. What fraction of this atom's mass is contributed by its nucleus?

**2.88** Give the molecular mass of each compound depicted below, and provide a correct name for any that are named incorrectly.



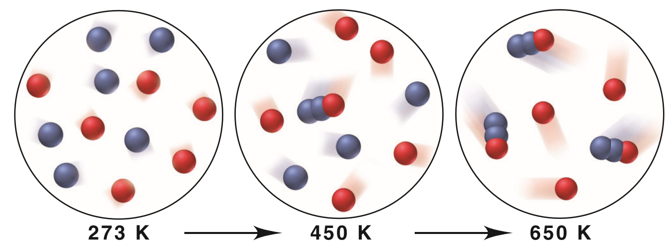

**2.89** Nitrogen forms more oxides than any other element. The percents by mass of N in three different nitrogen oxides are (I) 46.69%; (II) 36.85%; (III) 25.94%. For each compound, determine (a) the simplest whole-number ratio of N to O, and (b) the number of grams of oxygen per 1.00 g of nitrogen.

**2.90** Scenes A–I depict various types of matter on the atomic scale. Choose the correct scene(s) for each of the following:
(a) A mixture that fills its container
(b) A substance that cannot be broken down into simpler ones
(c) An element with a very high resistance to flow
(d) A homogeneous mixture
(e) An element that conforms to the walls of its container and displays an upper surface
(f) A gas consisting of diatomic particles
(g) A gas that can be broken down into simpler substances
(h) A substance with a 2/1 ratio of its component atoms
(i) Matter that can be separated into its component substances by physical means
(j) A heterogeneous mixture
(k) Matter that obeys the law of definite composition



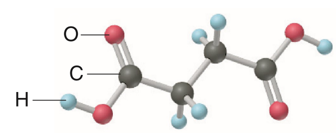

**2.91** The seven most abundant ions in seawater make up more than 99% by mass of the dissolved compounds. Here are their abundances in units of mg ion/kg seawater: chloride 18,980; sodium 10,560; sulfate 2650; magnesium 1270; calcium 400; potassium 380; hydrogen carbonate 140.
(a) What is the mass % of each ion in seawater?
(b) What percent of the total mass of ions is sodium ion?
(c) How does the total mass % of alkaline earth metal ions compare with the total mass % of alkali metal ions?
(d) Which make up the larger mass fraction of dissolved components, anions or cations?

**2.92** The scenes below represent a mixture of two monatomic gases undergoing a reaction when heated. Which mass law(s) is (are) illustrated by this change?




**2.93** When barium (Ba) reacts with sulfur (S) to form barium sulfide (BaS), each Ba atom reacts with an S atom. If 2.50 cm$^3$ of Ba reacts with 1.75 cm$^3$ of S, are there enough Ba atoms to react with the S atoms ($d$ of Ba = 3.51 g/cm$^3$; $d$ of S = 2.07 g/cm$^3$)?

**2.94** Succinic acid *(below)* is an important metabolite in biological energy production. Give the molecular formula, molecular mass, and the mass percent of each element in succinic acid.

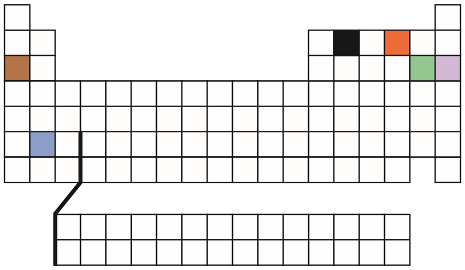



**2.95** Fluoride ion is poisonous in relatively low amounts: 0.2 g of $F^-$ per 70 kg of body weight can cause death. Nevertheless, in order to prevent tooth decay, $F^-$ ions are added to drinking water at a concentration of 1 mg of $F^-$ ion per L of water. How many liters of fluoridated drinking water would a 70-kg person have to consume in one day to reach this toxic level? How many kilograms of sodium fluoride would be needed to treat a $8.50 \times 10^7$-gal reservoir?

**2.96** Antimony has many uses, for example, in infrared devices and as part of an alloy in lead storage batteries. The element has two naturally occurring isotopes, one with mass 120.904 amu, the other with mass 122.904 amu. (a) Write the ${}^{A}_{Z}X$ notation for each isotope. (b) Use the atomic mass of antimony from the periodic table to calculate the natural abundance of each isotope.

**2.97** Dinitrogen monoxide ($N_2O$; nitrous oxide) is a greenhouse gas that enters the atmosphere principally from natural fertilizer breakdown. Some studies have shown that the isotope ratios of $^{15}N$ to $^{14}N$ and of $^{18}O$ to $^{16}O$ in $N_2O$ depend on the source, which can thus be determined by measuring the relative abundance of molecular masses in a sample of $N_2O$.
(a) What different molecular masses are possible for $N_2O$?
(b) The percent abundance of $^{14}N$ is 99.6%, and that of $^{16}O$ is 99.8%. Which molecular mass of $N_2O$ is least common, and which is most common?

**2.98** Use the box color(s) in the periodic table below to identify the element(s) described by each of the following:



(a) Four elements that are nonmetals
(b) Two elements that are metals
(c) Three elements that are gases at room temperature
(d) Three elements that are solid at room temperature
(e) One pair of elements likely to form a covalent compound
(f) Another pair of elements likely to form a covalent compound
(g) One pair of elements likely to form an ionic compound with formula MX
(h) Another pair of elements likely to form an ionic compound with formula MX
(i) Two elements likely to form an ionic compound with formula $M_2X$
(j) Two elements likely to form an ionic compound with formula $MX_2$
(k) An element that forms no compounds
(l) A pair of elements whose compounds exhibit the law of multiple proportions

**2.99** Dimercaprol ($HSCH_2CHSHCH_2OH$) is a complexing agent developed during World War I as an antidote to arsenic-based poison gas and used today to treat heavy-metal poisoning. Such an agent binds and removes the toxic element from the body. (a) If each molecule of dimercaprol binds one arsenic (As) atom, how many atoms of As can be removed by 250. mg of dimercaprol? (b) If one molecule binds one metal atom, calculate the mass % of each of the following metals in a metal-dimercaprol complex: mercury, thallium, chromium.

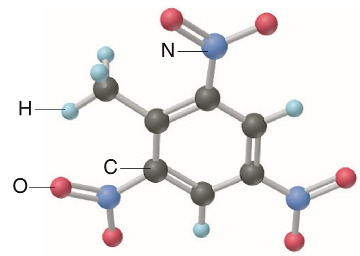
**2.100** From the following ions and their radii (in pm), choose a pair that gives the strongest ionic bonding and a pair that gives the weakest: $Mg^{2+}$, 72; $K^+$, 138; $Rb^+$, 152; $Ba^{2+}$, 135; $Cl^-$, 181; $O^{2-}$, 140; $I^-$, 220.

**2.101** A rock is 5.0% by mass fayalite ($Fe_2SiO_4$), 7.0% by mass forsterite ($Mg_2SiO_4$), and the remainder silicon dioxide. What is the mass percent of each element in the rock?

**2.102** The two isotopes of potassium with significant abundance in nature are $^{39}K$ (isotopic mass 38.9637 amu, 93.258%) and $^{41}K$ (isotopic mass 40.9618 amu, 6.730%). Fluorine has only one naturally occurring isotope, $^{19}F$ (isotopic mass 18.9984 amu). Calculate the formula mass of potassium fluoride.

**2.103** Nitrogen monoxide (NO) is a bioactive molecule in blood. Low NO concentrations cause respiratory distress and the formation of blood clots. Doctors prescribe nitroglycerin, $C_3H_5N_3O_9$, and isoamyl nitrate, $(CH_3)_2CHCH_2CH_2ONO_2$, to increase NO. If each compound releases one molecule of NO per atom of N it contains, calculate the mass percent of NO in each.

**2.104** TNT (trinitrotoluene; *below*) is used as an explosive in construction. Calculate the mass of each element in 1.00 lb of TNT.




**2.105** The anticancer drug Platinol (Cisplatin), $Pt(NH_3)_2Cl_2$, reacts with the cancer cell's DNA and interferes with its growth. (a) What is the mass % of platinum (Pt) in Platinol? (b) If Pt costs \$32/g, how many grams of Platinol can be made for \$1.00 million (assume that the cost of Pt determines the cost of the drug)?

**2.106** Which of the following steps in an overall process involve(s) a physical change and which involve(s) a chemical change?