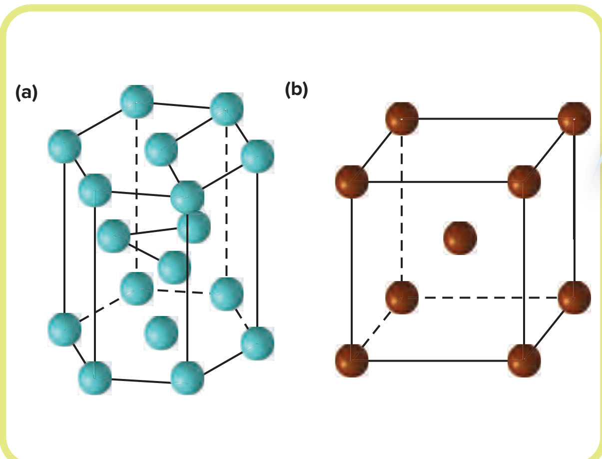
The Components of Matter



IN THIS CHAPTER . . . We examine the properties and composition of matter on the macroscopic and atomic scales. By the end of this chapter, you should be able to

- Relate the three types of matter—elements or elementary substances, compounds, and mixtures—to the simple chemical entities that they comprise—atoms, ions, and molecules; more complex entities, such as network covalent solids and macromolecules, will be discussed later
- Correlate the defining properties of the types of matter to laws discovered in the 18th century
- Describe the 19th-century atomic model that was proposed to explain these laws
- Explain how certain 20th-century experiments led to our current understanding of atomic structure and atomic mass
- Explain how the elements are organized in the periodic table, and introduce the two major ways in which elements combine
- Derive the name and formula of a compound, and calculate its mass
- Depict molecules using models used by chemists
- Classify mixtures, and explain how to separate them
- Differentiate between the components of matter

2

"Could anything at first sight seem more impractical than a body which is so small that its mass is an insignificant fraction of the mass of an atom of hydrogen?"

— J. J. Thomson.

Atoms are the building blocks of chemistry; everything is made of atoms; atoms are the tiniest particles and cannot be divided. These are facts about the atom that are learned from childhood and that are taken for granted with little thought as to how we know this or when such thoughts became prevalent. For those at the forefront of the effort to understand matter and its nature, it was a very different story. Conflicting ideas of what matter was, what differentiated one type of matter from another, and how far matter could be broken down until it became unbreakable caused huge divides among scientists. From Democritus, who first coined the term *atom* from the Greek word *atomos* meaning "that which cannot be cut," to Dalton, who proposed the first concrete framework for the properties of an atom, there was a lapse in time of over 2000 years. After much of the scientific community was finally convinced that atoms were real, it is unsurprising that the realization, first vocalized by J. J. Thomson, that atoms were in fact composed of even smaller particles came as a shock! The current model of the atom has a number of subatomic particles; in addition to the proton, neutron, and electron, atoms are known to be composed of gluons, muons, leptons, quarks, fermions, and bosons. How did scientists arrive at the current model of the atom?

This chapter briefly describes some of the important steps in the development of the model of the atom. The current model, the quantum model, accounts for known data about the atom and predicted information about recently discovered elements, thus making it a reasonable model. The observation of the Higgs boson in 2012 is one addition to our current body of scientific knowledge that supports the quantum model. This does not mean that at some time in the future, a better model may not be developed. Scientists constantly incorporate new data into existing models and revise hypotheses to improve knowledge.

Once we know how atoms of various elements differ from one another, we need to be able to express the multiple ways in which they are capable of combining. This chapter will explain the basics of chemical nomenclature and briefly discuss the periodic table as well.

Concepts and Skills to Review before Studying This Chapter

- Physical and chemical changes (Section 1.1)
- States of matter (Section 1.1)
- Attraction and repulsion between charged particles (Section 1.1)
- Meaning of a scientific model (Section 1.3)
- SI units and conversion factors (Section 1.5)
- Significant figures in calculations (Section 1.6)

2.1 Elements, Compounds, and Mixtures: An Atomic Overview

Matter can be classified into three types based on its composition: elements or elementary substances, 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 atoms with the same atomic number (number of protons)* and, therefore, cannot be broken down into a simpler type of matter

substance
A type of matter, either an element or a compound, that has a fixed composition

element
The simplest type of substance consisting of atoms with the same number of protons having unique physical and chemical properties; an element consists of only one kind of atom

by 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 colour, 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. *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 their physical state. Figure 2.1A shows atoms of an element in its gaseous state.

Most elements do not exist as individual atoms but form diatomic or polyatomic structures. These molecular forms are called **elementary substances**: a **molecule** is *an independent structure of two or more atoms bound together* (Figure 2.1B). Dioxygen or oxygen gas, for example, occurs in air as *diatomic* (two-atom) molecules.

2. *Compounds.* A **compound** consists of *two or more of the same or different elements that are bonded chemically* (Figure 2.1C). That is, the elements in a compound are not just mixed together: their atoms have joined in a chemical reaction. Many compounds, such as hydrogen, sulfur, ammonia, water, and carbon dioxide, consist of molecules. Other compounds, such as sodium sulfate (which we will discuss shortly) and silicon dioxide, do not. No matter what the compound, however, one defining feature is that *the elements are present in fixed parts*. This occurs because *each unit of the compound consists of a fixed number of atoms of each element*. For example, consider a sample of ammonia. Each ammonia molecule consists of one nitrogen atom and three hydrogen atoms:

Each ammonia molecule consists of one N atom and three 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; solid, black carbon; colourless and flammable hydrogen gas; and colourless and essential oxygen gas are very different from the compound they form—white, crystalline sodium hydrogen carbonate, or baking soda!

Unlike an element, an elementary substance or a compound *can* be broken down into simpler substances—its component elements. For example, sunlight can break a dioxygen (O_2) molecule into oxygen atoms, and 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 change.

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

elementary substance

As defined by IUPAC, the physical form of an element as it may be prepared and studied. With the exception of a very few elements that are monatomic, the majority of the elements form molecules of different types. Certain elements form diatomic molecules, such as hydrogen, oxygen, and nitrogen, while others form polyatomic species, such as phosphorus, P4, and sulfur, S8. Some elements may exist in the form of multiple elementary substances (for example, carbon exists as diamond, graphite, and fullerenes).

molecule

A structure that consists of two or more atoms bound chemically, behaving as an independent unit

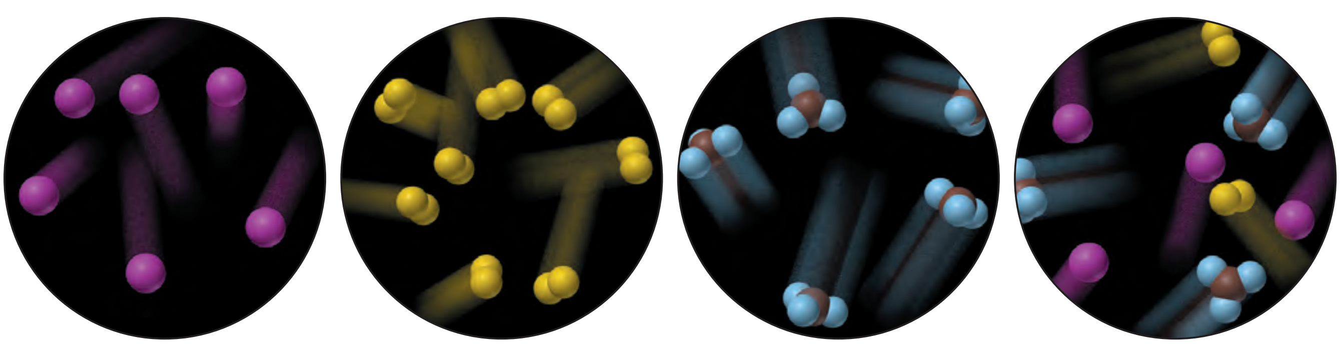
compound

A substance that consists of two or more elements chemically combined in fixed proportions

mixture

A group of two or more elements and/or compounds that are physically intermingled

FIGURE 2.1 Elements, compounds, and mixtures on the atomic scale. The samples depicted here are gases, but the three common types of matter also occur as liquids and solids. Most of the atoms and molecules are ionized when forming plasma. Not all elements form a BEC.



A Atoms of an element **B** Molecules of an element **C** Molecules of a compound **D** Mixture of two elements and a compound

TABLE 2.1 Some Properties of Sodium, Hydrogen, Carbon, Oxygen, and Sodium Hydrogen Carbonate

Property	Sodium	+	Hydrogen	+	Carbon	+	Oxygen	→	Sodium Hydrogen Carbonate (Baking Soda)
Melting point	97.8°C	−259.2°C		sublimes at 3825°C		−219°C		decomposes to sodium carbonate from 50°C	
Boiling point	881.4°C	−252.8°C		sublimes at 3825°C		−183°C		decomposes to sodium carbonate from 50°C	
Colour	silvery solid	colourless gas		black solid (graphite) white crystal (diamond)		colourless gas		white solid	
Density	0.97 g/cm^3	8.99 × 10^{-5} g/cm^3		2.2 g/cm^3 (graphite) 3.513 g/cm^3 (diamond)		1.429 kg/m^3		2.2 g/cm^3	
Behaviour in water	reacts		sparingly soluble		sparingly soluble		soluble		soluble

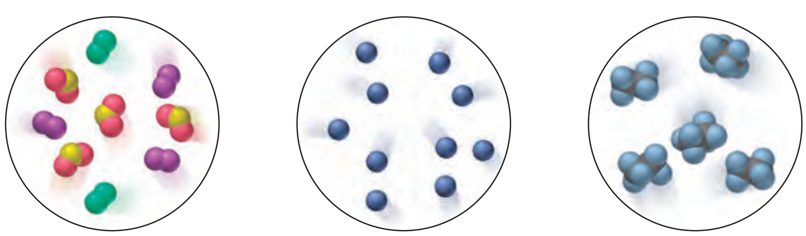
consists of the individual units that make up its component elements and/or compounds (Figure 2.1D). Thus, *a mixture retains many of the properties of its components*. Salt water, for instance, is colourless like water and tastes salty like sodium chloride. We know this because we put salt in food at home. Remember, we *never* taste anything in the lab!

Unlike compounds, mixtures can be separated into their components by *physical changes*; chemical changes are not needed. For example, the water in salt water can be boiled off, a physical process that leaves behind solid sodium chloride. Sample Problem 2.1 will help to differentiate these types of matter.

The elements, elementary substances, compounds, and mixtures shown in Figure 2.1 are depicted using specifically coloured atoms. At the end of Section 2.8 in this chapter, a diagram is provided that shows which atoms are assigned which specific colour (see Figure 2.20). The atoms of a particular element are depicted using the colour scheme provided to make it easier to consistently identify a particular element or molecule from the visual depiction. This colour scheme for atoms has been universally adopted in the field of chemistry.

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

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



(a) (b) (c)

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

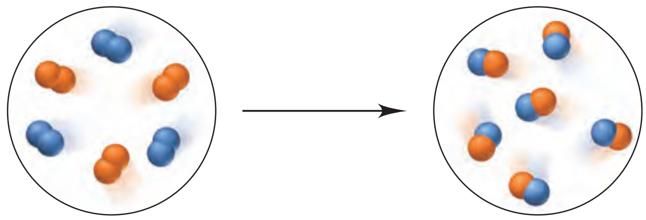
Plan We have to determine the type of matter by examining the component particles. If a sample contains only one type of particle, it is either an element or a compound; if it contains more than one type, it is a mixture. Particles of an element have only one kind of atom (one colour), and particles 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. The third type contains two red atoms for every one yellow atom, so it is a compound.

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

(c) Compound: The sample consists of molecules, each with two black atoms and six blue atoms.

Follow-Up Problem 2.1 Describe the following reaction in terms of elements, compounds, and mixtures. (See Brief Solutions to Follow-Up Problems at the end of the chapter.)



SUMMARY OF SECTION 2.1

- All matter exists as elements, elementary substances, compounds, or mixtures.
- Elements, elementary substances, and compounds are substances, which are types of matter with a fixed composition.
- An element consists of only one type of atom with a unique number of protons and occurs as a collection of individual atoms.
- An elementary substance, as defined by IUPAC, is a physical form of an element as it may be prepared and studied. With the exception of a very few elements that are monatomic, the majority of the elements form molecules of different types. Certain elements form diatomic molecules, such as hydrogen, oxygen, and nitrogen, while others form polyatomic species, such as phosphorus, P_4, and sulfur, S_8. Some elements may exist in the form of multiple elementary substances (for example, carbon exists as diamond, graphite, and fullerenes).
- A compound contains two or more elements that are chemically combined. The compound exhibits different properties from its component elements. Each unit of the compound has a fixed number of each type of atom. Only a chemical change can break a compound down 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

Originally, any model of the composition of matter had to explain the *law of mass conservation* and the *law of definite (or constant) composition*. As you will see, an atomic theory developed by John Dalton in the early 19th century explained these mass laws, as well as another law now known as the *law of multiple proportions*. However, our understanding of subatomic chemistry has given us the means to understand how compounds actually form. We can thus see that many of the earlier laws have exceptions. More compounds being found in nature and made in laboratories are proving to be the exception rather than the rule.

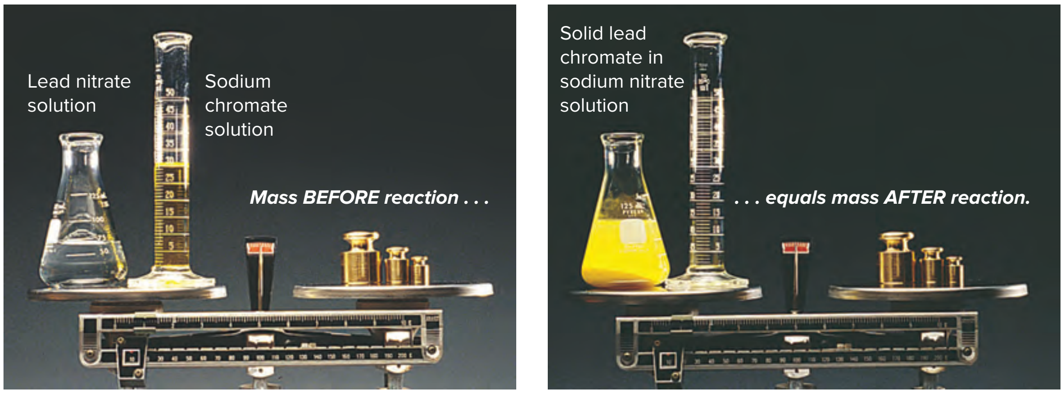


FIGURE 2.2 The law of mass conservation

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. (Lavoisier had first stated this law on the basis of his combustion experiments.) Figure 2.2 illustrates mass conservation because the lead nitrate and sodium chromate solutions (*left*) have the same mass as the solid lead chromate in sodium nitrate solution (*right*) that forms when they react.

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

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

Mass conservation means that, based on all chemical experience, *matter cannot be created or destroyed.*

To be precise, we now know, based on the work of Albert Einstein (1879–1955), that the mass before and after a reaction is not *exactly* the same. Some mass is converted to energy, or vice versa, but the difference is too small to measure, even with the best balance. For example, when 100 g of carbon burns, the carbon dioxide that is formed has mass 0.000 000 036 g (3.6×10^{-8} g) less than the carbon and oxygen that reacted. Because the energy changes of *chemical* reactions are so small, for all practical purposes, mass *is* conserved. Later in this book, you will see that energy changes in *nuclear* reactions are so large that mass changes are easy to measure.

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* (that is, multiplied by 100).

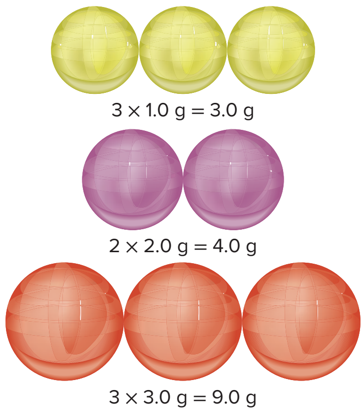
For an everyday example, consider a box that contains three types of marbles: yellow marbles have a mass of 1.0 g each, purple of 2.0 g each, and red of 3.0 g each. Each type makes up a fraction of the total mass of marbles, 16.0 g. The *mass fraction* of yellow marbles is their number times their mass divided by the total

law of mass conservation
A mass law stating that the total mass of the substances in a chemical reaction does not change during the reaction

law of definite (or constant) composition
A mass law stating that, no matter what its source, a particular compound is composed of the same elements in the same parts (fractions) by mass

fraction by mass (mass fraction)
The portion of a compound's mass that is contributed by an element; the mass of an element in a compound divided by the mass of the compound

percent by mass (mass percent, mass %)
The fraction by mass expressed as a percentage



$$\frac{3.0\text{ g}}{(3.0\text{ g} + 4.0\text{ g} + 9.0\text{ g})} = 0.19$$

$$\frac{4.0\text{ g}}{(3.0\text{ g} + 4.0\text{ g} + 9.0\text{ g})} = 0.25$$

$$\frac{9.0\text{ g}}{(3.0\text{ g} + 4.0\text{ g} + 9.0\text{ g})} = 0.56$$

mass: $\frac{3 \times 1.0\text{ g}}{16.0\text{ g}} = 0.19$. The *mass percent* (parts per 100 parts) of yellow marbles is $0.19 \times 100 = 19\%$ by mass. The purple marbles have a mass fraction of 0.25 and are 25% by mass of the total, and the red marbles have a mass fraction of 0.56 and are 56% by mass of the total.

Using this example, we can see that if there is a manufacturing defect, the three yellow marbles may not be exactly 1.0 g each. There may be small deviations in their mass. This is true for the purple and red marbles as well.

If we extend the marble analogy to a compound, each element has a *fixed* mass fraction and therefore mass percent. Calcium carbonate, the major compound in seashells, marble, and coral, 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

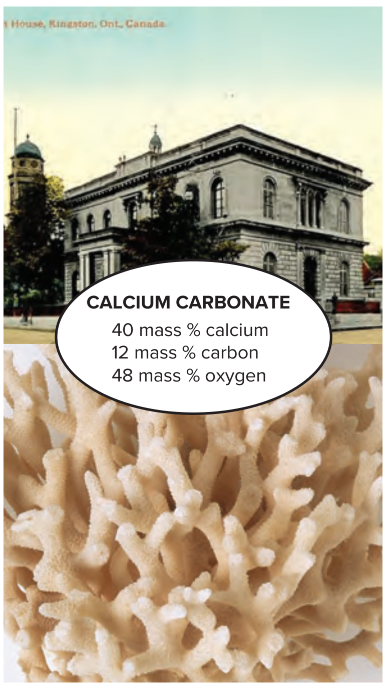


FIGURE 2.3 The law of definite composition. Calcium carbonate occurs in many forms (such as marble, *top*, and coral, *bottom*), but the mass percents of its elements are always the same.

(top): ©Historic Collection/Alamy Stock Photo; (bottom): ©Alexander Cherednichenko/Shutterstock

The masses of calcium, carbon, and oxygen in the first column are only true for a sample of calcium carbonate with a mass of 20.0 g—but *the mass fraction is fixed no matter what the size of the sample is*. The sum of the mass fractions is 1.00; or the sum of the mass percents is 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 2.3).

Again, just as with the marble analogy, there are exceptions to the rule. Many compounds will have slightly different masses depending on the actual atoms in the compound (and whether the subatomic composition of the atoms is different), based on where the compound was found. There are also many compounds that have a composition that is nonstoichiometric (see Chapter 3), but well defined. As an example, iron (II) oxide, FeO, is the commonly written compound, but it is actually not very common in nature. The more common compound is $Fe_{0.95}O$. For these reasons, while many compounds do follow the law of definite composition, many do not. We will learn more about stoichiometry in Chapter 3.

In general, since a given element constitutes the same mass fraction of a given compound, we can use this mass fraction to find the actual mass of the element in any sample of the compound:

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

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

Canada is the second-largest producer worldwide of uranium (roughly 22% of total output) and was just recently removed from first place by Kazakhstan, which now produces roughly 39% of the world output. For more information on uranium, see Chapter 25. 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. What mass of uranium is in 102 kg of pitchblende?

The McArthur River Uranium Mine in northern Saskatchewan is the world's largest.

©Sprokop/GetStock.com

Plan We must 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 (kg) of the pitchblende sample by the ratio of uranium to pitchblende from the mass analysis. This gives the mass (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 Find the mass (kg) of uranium in 102 kg of pitchblende:

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

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

Convert the mass of uranium from kg to g:

$$\text{Mass (kg) of uranium} = 86.5 \cancel{\text{ kg}} \text{ uranium} \times \frac{1000 \text{ g}}{1 \cancel{\text{ kg}}} = \boxed{8.65 \times 10^4 \text{ g uranium}}$$

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

$$\sim 100 \text{ kg pitchblende} \times \frac{72}{84} = 100 \text{ kg pitchblende} \times \frac{6}{7} \approx 86 \text{ kg uranium}$$

Follow-Up Problem 2.2 What mass (t) of oxygen is present in a sample of pitchblende that contains 2.3 t of uranium (1 t = 1000 kg)? (*Hint*: Remember that oxygen is the only other element in pitchblende.)

Multiple Proportions

An observation that applies when two elements form more than one compound is 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 that carbon and oxygen form; let us call them I and II. These compounds have very different properties: the density of carbon oxide I is 1.25 g/L, whereas the density of carbon oxide II is 1.98 g/L; carbon oxide I is poisonous and flammable, but carbon oxide 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

law of multiple proportions
A mass law stating that if elements A and B react to form two or more compounds, the different masses of B that combine with a fixed mass of A can be expressed as a ratio of small whole numbers

To see the phenomenon of multiple proportions, we use the mass percents of oxygen and 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 the carbon oxide example, the ratio is 2/1—for a given mass of carbon, carbon oxide II contains *2 times* as much oxygen as carbon oxide I, not 1.583 times, 1.716 times, or any other intermediate amount. In Section 2.3, we will 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 samples of a given compound have 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 element can be expressed as a ratio of small whole numbers.

John Dalton (1766–1844)
©Georgios Kollidas/GeorgiosArt/iStock/Getty Images

2.3 Dalton's Atomic Theory

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

Despite having no formal education, Dalton began teaching science at age 12 and then studied colour-blindness, an affliction still known as *daltonism*. At 21, he started recording daily weather data, continuing this project for the rest of his life. His results on humidity and dew point led to a key discovery about gases (Section 4.4) and eventually to his atomic theory.

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, let us see which were original and which came from others.

atoms
The smallest particles of an element that retain the chemical nature of the element; neutral, spherical entities composed of a positively charged central nucleus surrounded by one or more negatively charged electrons

1. All matter consists of **atoms**, *tiny indivisible particles of an element that cannot be created or destroyed*. (This idea was derived 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 idea rejects the alchemical belief in the magical transmutation of elements.)
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. With our current understanding of subatomic properties, we know it is incorrect, but at the time, it was a groundbreaking idea.)
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. Again, our current understanding allows us to realize that this statement is also not always true.)

How the Atomic Theory Explains the Mass Laws

Let us 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.

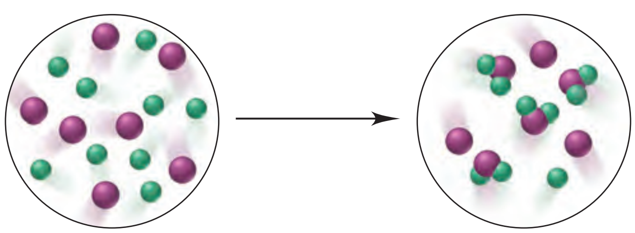
The *simplest* arrangement consistent with the mass data for carbon oxides I and II in our earlier example is one atom of oxygen combining with one atom of carbon in compound I (carbon monoxide) and two atoms of oxygen combining with one atom of carbon in compound II (carbon dioxide):

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

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

Sample Problem 2.3 Visualizing the Mass Laws

Problem The diagram below represents an atomic-scale view of a chemical reaction:

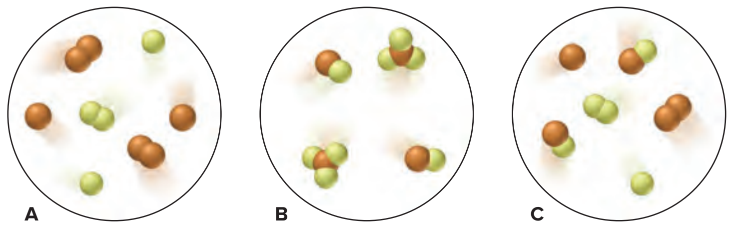


Which of the mass laws—mass conservation, definite composition, or multiple proportions—is (are) illustrated?

Plan We note the numbers, colours, and combinations of atoms (spheres) in the diagrams to see which mass laws are illustrated. 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 that always has the same atom ratio forms, 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 atom and two green atoms, so it has definite composition. Only one compound forms, so the law of multiple proportions is not illustrated.

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 each compound consists of whole atoms.

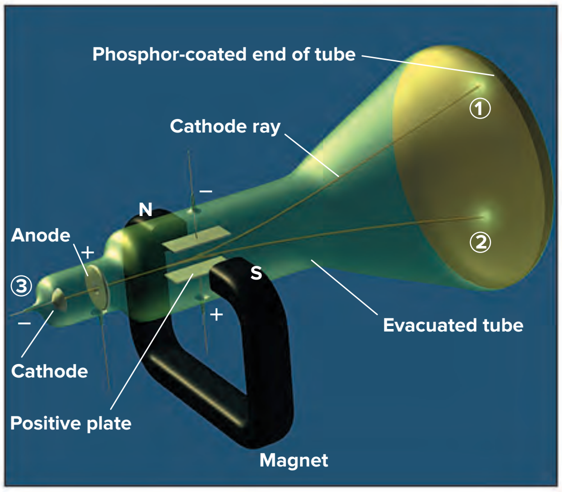
2.4 The Observations That Led to the Nuclear Atom Model

Dalton's model established that masses of reacting elements could be explained in terms of atoms. However, it did not establish 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 solid model of the atom did not predict the existence of subatomic charged particles, which were observed in later experiments that led to the discovery of *electrons* and the atomic *nucleus*. Let us 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 knew that matter and electric charge were related. When amber is rubbed with fur, or glass is rubbed with silk, positive and negative charges form—the same charges that make your hair crackle and cling to your comb on a dry day. Scientists also knew that an electric current could decompose certain compounds into their elements, but they did not know what a current was made of.



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

FIGURE 2.4 **Observations that established the properties of cathode rays**

Cathode Rays To discover its nature, some investigators tried passing electric 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 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 2.4 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*. In a neon sign, electrons collide with the gas particles in the tube, causing them to give off light.

cathode rays
Rays of light emitted by the cathode (negative electrode) in a gas discharge tube; travel in straight lines, unless deflected by magnetic or electric fields

Mass and Charge of the Electron Two classic experiments and their conclusions 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 ratio with the mass/charge ratio for the lightest charged particle in solution, Thomson estimated that the cathode ray particle weighed less than $\frac{1}{100}$ 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 2.5). 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 was within 1% of the modern value of the electron's charge, which is currently defined as $-1.602\ 176\ 634 \times 10^{-19}$ C (C stands for *coulomb*, the SI unit of charge; see Table 1.2).

   What is most interesting about Millikan's research is that it led to one of the great scientific controversies of the time. A Viennese researcher by the name of

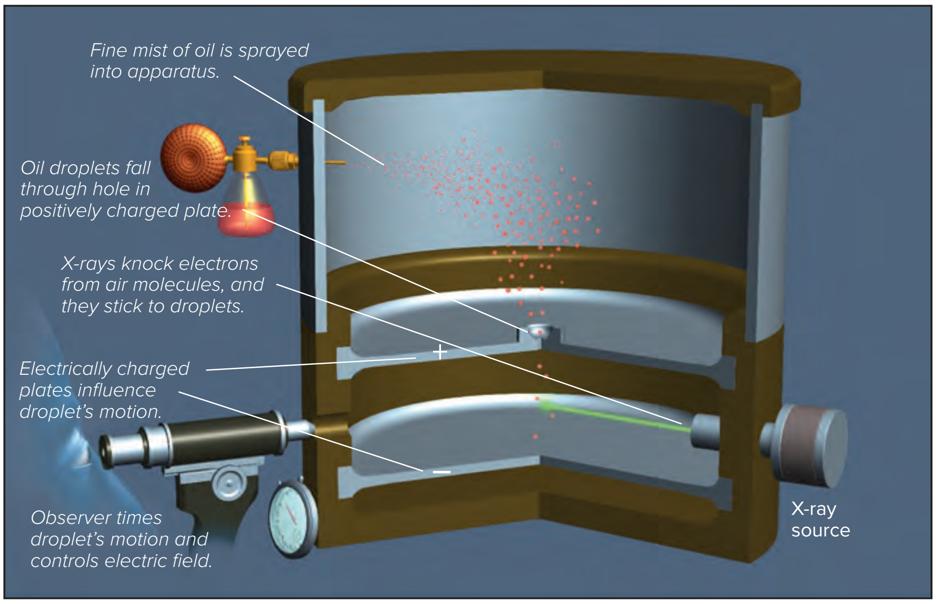


FIGURE 2.5 Millikan's oil-drop experiment for measuring the charge of an electron. The total charge on an oil droplet is some whole-number multiple of the charge of the electron.

Felix Ehrenhaft (1879–1952) was studying a similar problem. Millikan used a fixed set of guiding assumptions from which he never wavered while performing his experiments, leading him to conclude that he had determined the charge on the electron. Although Ehrenhaft performed similar experiments, he used a very different set of assumptions, which led him to quite different conclusions. Ehrenhaft was convinced that what Millikan had measured was the fractional charges on subelectrons. The Millikan–Ehrenhaft controversy lasted from 1910 to 1923 and was the subject of heated discussion involving such famous scientists as Einstein, Planck, Born, and Schrödinger. The entire situation once more came into the spotlight in 1978 when Gerald Holton (1922–) found two of Millikan's notebooks, from which it appeared that Millikan had used only part of his data to arrive at his conclusions. The issue of scientific fraud and falsification of data tainted Millikan's work for some time. It is interesting to note that Millikan always maintained a clear sense of belief in his work, and it was finally shown that, in fact, the data he had excluded had not met the strict set of guiding assumptions that the data he published had met.

As students of a largely experimental science, chemists can learn some very interesting lessons from the Millikan–Ehrenhaft controversy and the subsequent issues that arose from it. It is essential to maintain very clear notes when performing experiments in the lab. While it is critical to explain data that are published, it is sometimes equally if not more critical to explain data that are excluded. It is also important to persevere in achieving a goal and to have conviction in our work.

3. *Conclusion: calculating the mass of an electron.* The electron's mass/charge ratio (from work by Thomson and others) and the value for the electron's charge can be used to find the electron's mass, which is *extremely* small:

$$\text{Mass of electron} = \frac{\text{mass}}{\cancel{\text{charge}}} \times \cancel{\text{charge}} = \left(-5.686 \times 10^{-12}\,\frac{\text{kg}}{\cancel{\text{C}}}\right)(-1.602 \times 10^{-19}\,\cancel{\text{C}})$$
$$= 9.109 \times 10^{-31}\ \text{kg} = 9.109 \times 10^{-28}\ \text{g}$$

nucleus

The central region of the atom that, although very small with respect to the volume of the atom, is positively charged and contains essentially all of the mass of the atom

Discovery of the Atomic Nucleus

The presence of electrons in all matter presented some major questions about the structure of atoms. Matter is electrically neutral, so atoms must also be neutral. However, if atoms contain negatively charged electrons, what positive charges

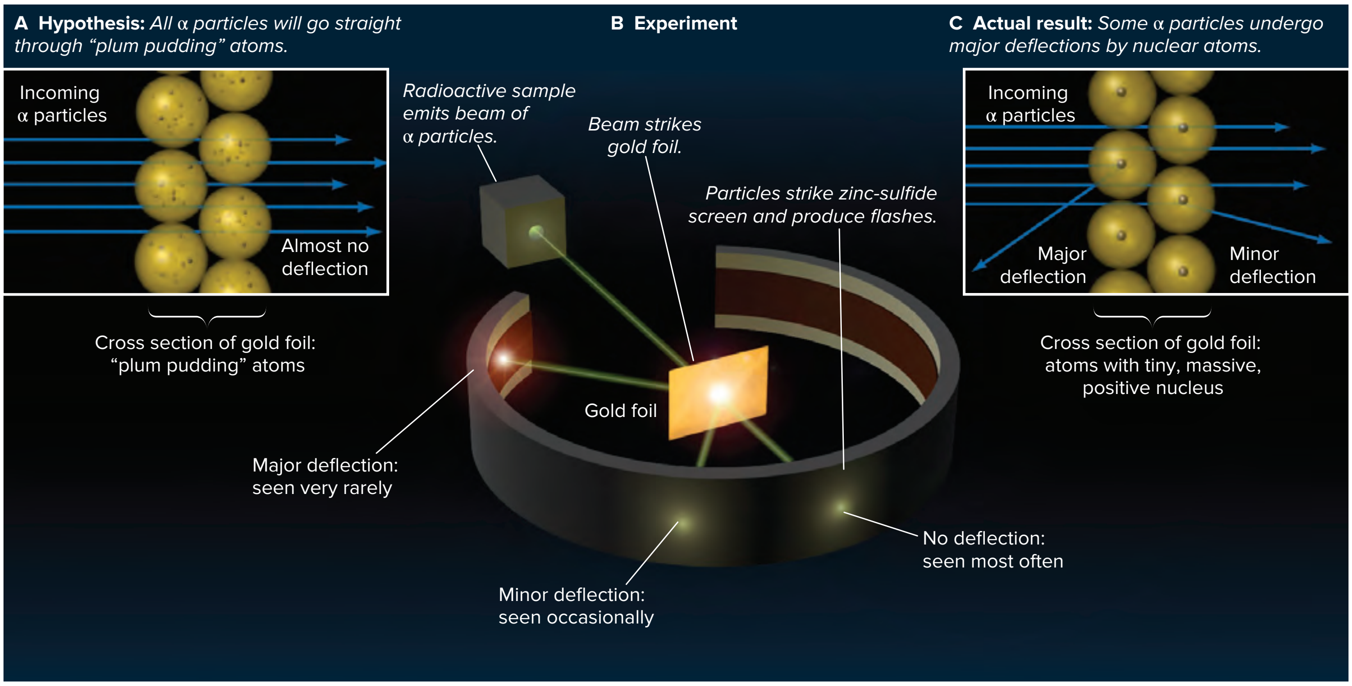


FIGURE 2.6 Rutherford's α particle scattering experiment and discovery of the atomic nucleus

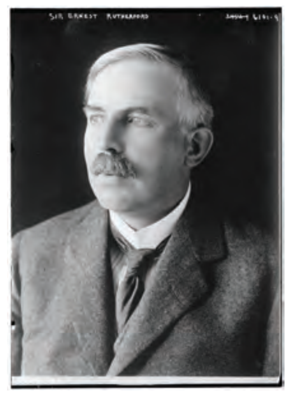
balance them? If an electron has such a tiny mass, what accounts for an atom's much larger mass? To address these issues, Thomson proposed his "plum pudding" model—a spherical atom composed of diffuse, positively charged matter with electrons embedded like "raisins in a plum pudding."

In 1910, New Zealand–born physicist Ernest Rutherford (1871–1937) tested this model and obtained an unexpected result (Figure 2.6):

1. *Experimental design*. Figure 2.6B 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.
2. *Expected results*. With Thomson's model in mind (Figure 2.6A), Rutherford expected only minor, if any, deflections of the α particles because they should act like bullets and go right through the gold atoms. After all, an electron should not deflect an α particle any more than a table-tennis ball would deflect a baseball.
3. *Actual results*. Initial results were consistent with this idea, but then the unexpected happened (Figure 2.6C). A few α particles were deflected, and 1 in 20 000 was deflected by more than 90° (it went "backwards").
4. *Rutherford's 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 mass, charge, and velocity of the α particles and the proportion of these large-angle deflections showed that
   - an atom is mostly space occupied by electrons
   - *in the centre is a tiny region*, which Rutherford called the **nucleus**, *that contains all of the positive charge and essentially all of the mass of the atom*

Rutherford proposed that positive particles lie within the nucleus and called them *protons*.

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



Ernest Rutherford (1871–1937) made enormous contributions to the fields of chemistry and physics. Born in New Zealand, Rutherford was Macdonald Professor of Physics from 1898 to 1907 at McGill University in Montréal. He received a Nobel Prize in 1908 for "his investigations into the disintegration of the elements and the chemistry of radioactive substances," work he did while at McGill. He is also the only Nobel Prize winner in Science to have performed his most famous work *after* obtaining the Nobel Prize! His astonishment at the results obtained in his famous gold foil experiment is clear in his words, "I remember two or three days later Geiger [one of his co-workers] 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."

Library of Congress Prints & Photographs Division [LC-DIG-ggbain-36570]

SUMMARY OF SECTION 2.4

- Several major discoveries at the beginning 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 the mass of the electron.
- Ernest Rutherford proposed that atoms consist of a tiny, dense, positive nucleus surrounded by electrons.

proton (p^+)
A subatomic particle that is found in the nucleus and has a unit positive charge ($1.602\ 18 \times 10^{-19}$ C)

neutron (n^0)
An uncharged subatomic particle found in the nucleus, with a mass slightly greater than the mass of a proton

electron (e^-)
A subatomic particle that possesses a unit negative charge ($-1.602\ 18 \times 10^{-19}$ C) and occupies the space around the atomic nucleus

atomic number (*Z*)
The unique number of protons in the nucleus of each atom of an element (equal to the number of electrons in the neutral atom); an integer that expresses the positive charge of a nucleus or subatomic particle in multiples of the electronic charge

mass number (*A*)
The total number of protons and neutrons in the nucleus of an atom

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. In this section, we examine that model and see how Dalton's theory stands up today.

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 2.7). The electrons move rapidly within the available volume, held there by the attraction of the nucleus. An atom's diameter ($\sim 1 \times 10^{-10}$ m) is about 20 000 times the diameter of its nucleus ($\sim 5 \times 10^{-15}$ m). The nucleus contributes 99.97% of the atom's mass, occupies only about 10^{-12} of its volume, and is incredibly dense—about 10^{14} g/mL!

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 118 known elements, of which 90 occur in nature and 28 have been synthesized by nuclear scientists.

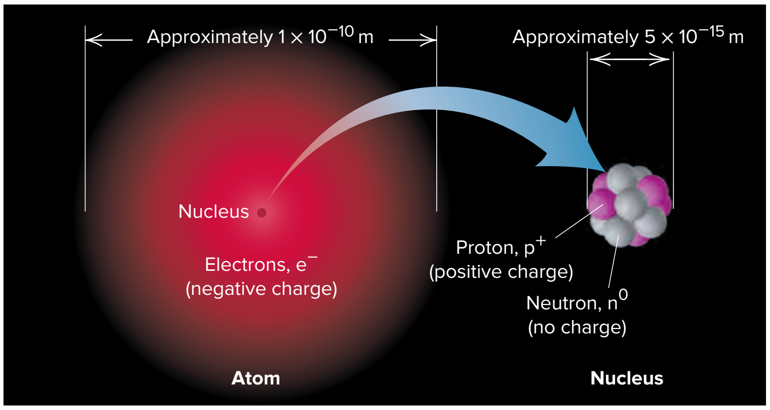


FIGURE 2.7 General features of the atom

TABLE 2.2 Properties of the Three Key Subatomic Particles

	Charge		Mass		
Name (Symbol)	**Relative**	**Absolute (C)***	**Relative (u)†**	**Absolute (g)**	**Location in Atom**
Proton (p^+)	1+	$+1.602\ 18 \times 10^{-19}$	1.007 27	$1.672\ 62 \times 10^{-24}$	Nucleus
Neutron (n^0)	0	0	1.008 66	$1.674\ 93 \times 10^{-24}$	Nucleus
Electron (e^-)	1−	$-1.602\ 18 \times 10^{-19}$	0.000 548 58	$9.109\ 39 \times 10^{-28}$	Outside nucleus

*The coulomb (C) is the SI unit of charge.
†The unified atomic mass unit (u) equals $1.660\ 54 \times 10^{-24}$ g; it is discussed later in this section.

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 (from the 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; for example, ${}^{A}_{Z}X$ represents element X. Since the mass number is the sum of the protons and neutrons, the number of neutrons (N) equals the mass number minus the atomic number:

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

Thus, a chlorine atom, represented by ${}^{17}_{35}Cl$, 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.

atomic symbol
A one- or two-letter abbreviation for the English, Latin, or Greek name of an element

isotopes
Atoms of a given atomic number (that is, of a specific element) that have different numbers of neutrons and, therefore, different mass numbers

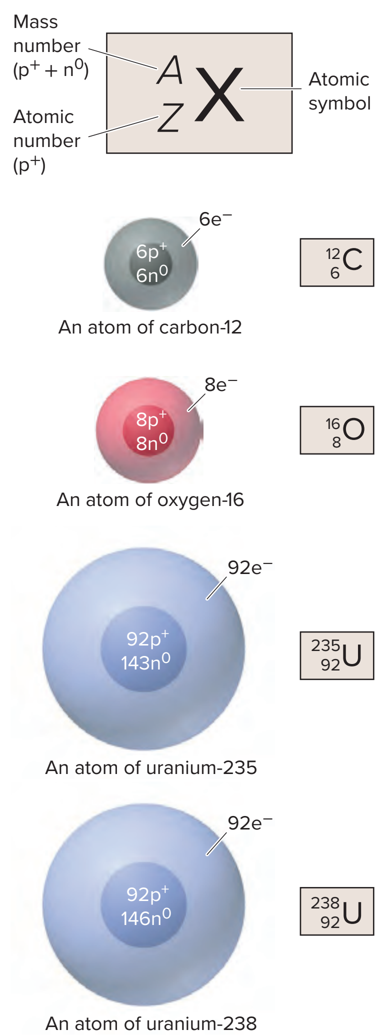


FIGURE 2.8 The ${}^{A}_{Z}X$ notations and spherical representations for four atoms. The nuclei are not drawn to scale.

Isotopes

All atoms of an element have the same atomic number but not necessarily 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 six protons and six electrons, but only 98.89% of naturally occurring carbon atoms have six neutrons ($A = 12$). A small percentage (1.11%) have seven neutrons ($A = 13$), and even fewer (less than 0.01%) have eight ($A = 14$). *A natural sample of carbon has the three natural isotopes* ${}^{12}C$, ${}^{13}C$, and ${}^{14}C$ *in the relative proportions* 98.89%, 1.11%, and less than 0.01%. Five other carbon isotopes—${}^{9}C$, ${}^{10}C$, ${}^{11}C$, ${}^{15}C$, and ${}^{16}C$—have been created in the laboratory. Figure 2.8 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 behaviour*, 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. Recall that the mass number is the sum of the protons and neutrons. From the list of elements given in the Endsheets, 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 list of elements, the atomic number of silicon is 14. Therefore,

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

Follow-Up Problem 2.4 How many protons, neutrons, and electrons are in **(a)** $^{11}_{5}Q$; **(b)** $^{41}_{20}R$; **(c)** $^{131}_{53}X$? Which elements do Q, R, and X represent?

unified atomic mass unit (u)
A unit used to specify mass on an atomic scale; defined as $\frac{1}{12}$ the mass of a carbon-12 atom

dalton (Da)
A unit of mass that is identical to the *unified atomic mass unit*

mass spectrometry
An instrumental method for measuring the relative masses of particles in a sample by creating charged particles and separating the particles according to their mass-charge ratio

isotopic mass
The mass (in unified atomic mass units, u) of an isotope relative to the mass of carbon-12

atomic mass
The average of the masses of the naturally occurring isotopes of an element weighted according to their abundances

Atomic Masses of the Elements

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 **unified atomic mass unit (u)** 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 u; in other words, a ^{12}C atom has almost 12 times the mass of a ^{1}H atom. We will continue to use the term *unified atomic mass unit* in this book, although the unit name **dalton (Da)** is also acceptable; thus, one ^{12}C atom has a mass of 12 Da, or 12 u. The unified atomic mass unit is a unit of relative mass, but it has an absolute mass of $1.660\,54 \times 10^{-24}$ g.

The isotopic makeup of an element is determined by **mass spectrometry**, *a method for measuring the relative masses and abundances of atomic-scale particles very precisely*. (Mass spectrometry is discussed in the Tools of the Laboratory section that follows.) For example, using a mass spectrometer, we measure the mass ratio of ^{28}Si to ^{12}C as

$$\frac{\text{Mass of } ^{28}\text{Si atom}}{\text{Mass of } ^{12}\text{C standard}} = 2.331\,411$$

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

$$\begin{aligned}\text{Isotopic mass of } ^{28}\text{Si} &= \text{measured mass ratio} \times \text{mass of } ^{12}\text{C}\\ &= 2.331\,411 \times 12\ \text{u} = 27.976\,93\ \text{u}\end{aligned}$$

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** 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 example, multiplying the isotopic mass of ^{28}Si by its fractional abundance gives the portion of the atomic mass of Si that is contributed by ^{28}Si:

$$\text{Portion of Si atomic mass from } ^{28}\text{Si} = 27.976\,93\ \text{u} \times 0.9223 = 25.8031\ \text{u}$$
(retaining two additional significant figures)

Similar calculations give the portions contributed by ^{29}Si ($28.976\,495\ \text{u} \times 0.0467 = 1.3532\ \text{u}$) and by ^{30}Si ($29.973\,770\ \text{u} \times 0.310 = 0.9292\ \text{u}$). Adding the three portions together (rounding to two decimal places at the end) gives the atomic mass of silicon:

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

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

TOOLS OF THE LABORATORY

Mass Spectrometry

Mass spectrometry is a powerful technique for measuring the mass and abundance of charged particles (their mass/charge ratio, m/z). The first mass spectrometer was constructed in 1912 by J. J. Thomson (Figure B2.1), who noted that ionized atoms or molecules with positive charge in a vacuum, exposed to parallel electrostatic and magnetic fields, arrived at the detector in parabolic arcs. No two ions arrived at the same spot unless they had the same m/z ratio. The fact that Thomson observed different arcs for the same ionized atoms led to the idea that different atoms of the same element could have different masses. This, in turn, led to the realization of the existence of isotopes.

Francis Aston received the Nobel Prize in Chemistry in 1922 "for his discovery, by means of his mass spectrograph, of isotopes, in a large number of nonradioactive elements, and for his enunciation of the whole-number rule." One of the first major applications of the mass spectrometer was in the Manhattan project, where it was used not only to separate radioactive isotopes but also to collect them after separation. Mass spectrometry received major interest from the petrochemical industry after World War II, in an effort to know more precisely what products were being produced during cracking processes.

How a Mass Spectrometer Works

In the figure below,

A: Could be Gas chromatograph (GC) or liquid chromatograph (LC/HPLC)

B: Could be Inductively coupled plasma (ICP) → completely destroys molecule; only atoms remain
Electron impact (EI) → medium-high fragmentation
Matrix assisted laser desorption ionization (MALDI) → low-medium fragmentation
Fast atom bombardment (FAB) → little fragmentation
Electrospray ionization (ESI) → little fragmentation

C/E: Could be Quadrupole → mass range to 1000 m/z; low resolution
Ion trap → mass range to 1000 m/z; low resolution
Time of flight (TOF) → mass range to 10 000 m/z; medium-high resolution
Electromagnet → mass range to 5000 m/z; high-resolution
Fourier transform ion cyclotron resonance (FTICR) → mass range 5000 to 10 000 m/z; extremely high resolution

D: Could be Collision with inert gas or another method of enhancing internal energy of ion of interest

F: Could be Multichannel plate (MCP)
Electron multiplier (EM)
Scintillator + Photon multiplier tube (PMT)

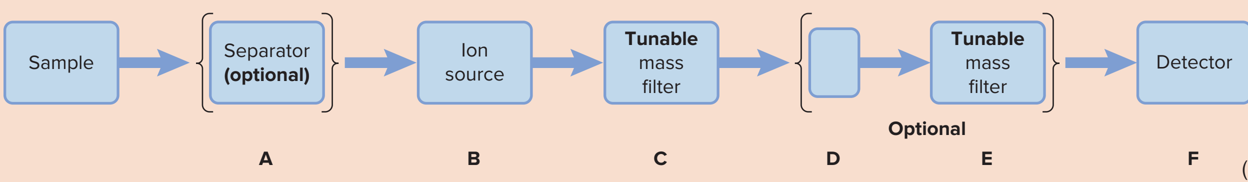
The chemical separator (if necessary) is used to separate the original mixture into the component molecules to be analyzed. The ion source then ionizes the sample. The ion source is chosen based on the type of sample to be analyzed and the extent of fragmentation desired. The method of ionization determines whether the ions have a little or a lot of internal energy remaining. If a molecule has very little internal energy after ionization, it will stay as it is, resulting in the molecular ion. However, if there is sufficient internal energy, bonds in the molecular ion can and will be broken, leading to fragment ions. These fragment ions contain information about the structure of the molecular ion. The tunable mass filter then allows the detection of a range of all of the ions generated inside the ion source by the m/z ratio. Different mass filters vary in their mass range and resolution (precision). The detector records either the charge induced or the current produced when an ion passes by or hits the surface.

FIGURE B2.1 **J. J. Thomson, the father of mass spectrometry**
©Pictorial Press Ltd./Alamy

If, for example, we had a mixture of CO (27.9949 u), N_2 (28.0061 u), and C_2H_4 (28.0313 u), it would appear that we have three molecules, all with a nominal m/z of 28, which would give a single peak using a low-resolution mass filter such as a quadrupole. If we used a high-resolution mass filter, such as a magnet or ICR, the three ions in our example would actually give three distinct peaks. A mass measured with sufficient accuracy can thus serve as proof of structure. However, even a low-resolution mass filter would allow us to distinguish the three molecules in our example based on their fragmentation patterns. (For example, fragmentation of CO would not give a peak at an m/z of 14; similarly, fragmentation of N_2 would not give a peak at an m/z of 12 or 16.) Thus, mass spectrometry is helpful for determining the structure of molecules based on exact mass and/or fragmentation patterns (Figure B2.2).

Every molecule has a characteristic fragmentation pattern, which leads to the idea of molecular fingerprints. These fragmentation fingerprints have been collected into databases and are available to identify unknown substances. Major applications include blood analysis and reference for analysis in crime labs, quality-control labs, and environmental labs.

If an instrument has only one mass analyzer, the observed fragmentation pattern is that of *everything* generated in the ion source. If there is a second mass analyzer, however, we could use the first analyzer to select one particular ion from all of those



(Continued)

generated in the ion source and the second analyzer to measure the fragmentation pattern of only the *ion* selected by the first analyzer. In order to do this, we need to put extra energy into the selected ion to make it fragment. This can be done in various ways; collision with an inert gas is commonly used. These secondary mass spectra/MS-MS spectra can be very helpful when the fragmentation pattern obtained from the ion source is ambiguous because of the presence of contaminants. MS-MS performs a function for ions that is similar to the function of the chemical separator for molecules.

Most commonly used instruments have a GC or an LC chemical separator. An EI/Quad (such as an HP benchtop GC-MS) is used for routine analysis, both qualitative and quantitative, of small organics. A (GC) EI/Magnetic instrument is used for accurate mass measurements and parts per million–level environmental analysis, while an (LC) ESI/TOF or MALDI-TOF is more suitable for the study of large molecules, such as peptides. Some of the most significant mass spectrometers used in the industry are designed, and in some cases built, in Woodbridge, Ontario. These include the SCIEX mass spectrometers and some PerkinElmer ICP mass spectrometers (SCIEX: sciex.com; PerkinElmer NexION ICP-MS: www.perkinelmer.ca/en-ca/Catalog/Product/ID/NexION350X).

Mass spectrometry is now used to measure the mass of virtually any atom, molecule, or molecular fragment. The technique is being used to study catalyst surfaces, forensic materials, fuel mixtures, medicinal agents, radiocarbon dating, and much more. In particular, proteins are a very important area of study. In fact, John B. Fenn and Koichi Tanaka shared part of the 2002 Nobel Prize in Chemistry for developing methods of studying proteins by mass spectrometry.

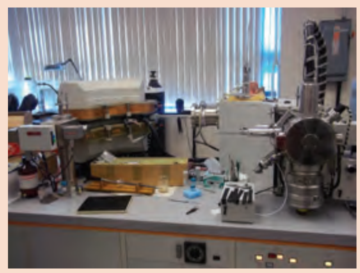
Problems

B2.1 Chlorine has two naturally occurring isotopes, ^{35}Cl (abundance 76%) and ^{37}Cl (abundance 24%), and it occurs as diatomic (two-atom) molecules. In a mass spectrum, peaks are seen for the molecule and for the separated atoms.
(a) How many peaks are in the mass spectrum?
(b) What are the m/z values of the heaviest particle and the lightest particle?

B2.2 When a sample of pure carbon is analyzed by mass spectrometry, peaks X, Y, and Z are obtained (random order). Peak Y is taller than peaks X and Z, and peak Z is taller than peak X. What is the m/z value of the isotope that is responsible for peak Z?

A

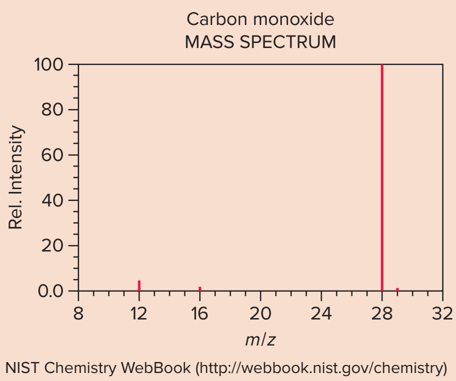
B

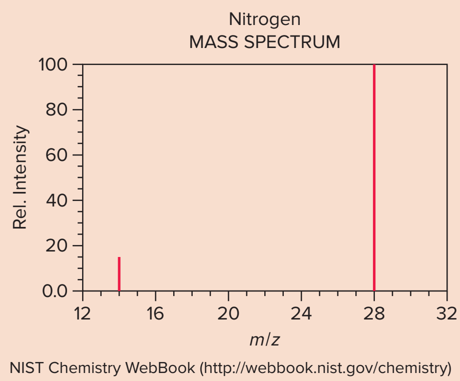


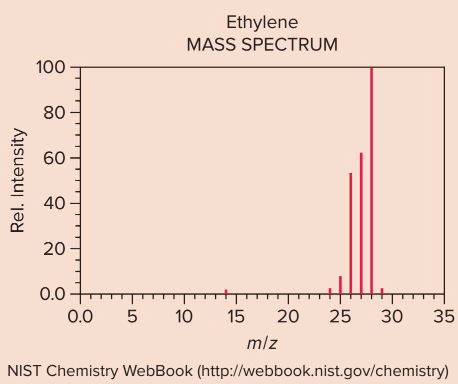
C

D

E







F

FIGURE B2.2 Mass spectrometers and spectra. A. The size of the sample quadrupoles relative to the size of a magnet. **B.** TOF tube. **C.** Magnetic Sector-Double Focusing KRATOS CONCEPT 1S. **D.** Benchtop HP GCMSD unit. **E.** Benchtop Varian 500 ITMS. **F.** Mass spectra of CO, N_2, and C_2H_4.
All photos courtesy of Rashmi Venkateswaran

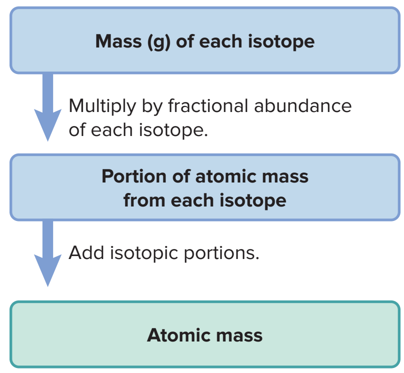
Sample Problem 2.5 Calculating the Atomic Mass of an Element

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

Isotope	Mass (u)	Abundance (%)
^{107}Ag	106.905 09	51.84
^{109}Ag	108.904 76	48.16

Plan From the mass and abundance of the two Ag isotopes, we have to find the atomic mass of Ag (the weighted average of the isotopic masses). We divide each percent abundance by 100 to get the fractional abundance and then multiply this value 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.

Road Map



Solution Find the fractional abundances:

$$\text{Fractional abundance of } ^{107}\text{Ag} = \frac{51.84}{100} = 0.5184$$

Similarly, the fractional abundance of ^{109}Ag is 0.4816.

Find the portion of the atomic mass from each isotope:

$$\begin{aligned}\text{Portion of atomic mass from } ^{107}\text{Ag} &= \text{isotopic mass} \times \text{fractional abundance}\\ &= 106.905\ 09\ \text{u} \times 0.5184 = 55.42\ \text{u}\\ \text{Portion of atomic mass from } ^{107}\text{Ag} &= 108.904\ 76\ \text{u} \times 0.4816 = 52.45\ \text{u}\end{aligned}$$

Find the atomic mass of silver:

$$\text{Atomic mass Ag} = 55.42\ \text{u} + 52.45\ \text{u} = 107.87\ \text{u}$$

Check The individual portions seem right: $\sim 100\ \text{u} \times 0.50 = 50\ \text{u}$. 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 (see Endsheets), the atomic mass is rounded to 107.9 u.

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 the following data: atomic mass of B = 10.81 u, isotopic mass of ^{10}B = 10.0129 u, and isotopic mass of ^{11}B = 11.0093 u. (*Hint*: The sum of the fractional abundances is one. 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 is the mass number (sum of protons and neutrons), and X is the atomic symbol.
- An element occurs naturally as a mixture of isotopes, which are 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. The atomic mass is determined using modern instruments, especially the mass spectrometer.

2.6 Elements: A First Look at the Periodic Table

In 1871, the Russian chemist Dmitri Mendeleev (1836–1907) published the most successful of several organizing schemes as a table of the elements. He listed the elements by increasing atomic mass and arranged them so that elements with similar

periodic table of the elements

A table in which the elements are arranged by atomic number into columns (groups) and rows (periods)

periods

The horizontal rows in the periodic table

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.

Lothar Meyer (1830–1895), a German scientist, received the degree Doctor of Medicine in 1854 and then went on to study mathematical physics. He accepted the position of Professor of Chemistry at the University of Tübingen in 1876. He worked extensively on the classification of the periodic nature of the elements. In 1862, he began writing *Die modernen Theorien der Chemie*, which was published in 1864. It was the first time that the elements had been classified into six families by valence rather than atomic mass. He went on to publish papers with other details of periodic classification in 1864 and 1870. His paper in 1869 of a periodic table containing all of the known elements of the time was unfortunately published in a German journal a few months after Mendeleev published his memorable paper on the periodic table in a Russian journal. While Meyer's paper also contained predictions of as-yet-undiscovered elements, he did not go to the extent of Mendeleev in predicting their physical and chemical properties. Nevertheless, the Royal Society honoured both Meyer and Mendeleev with the Davy Medal in 1882 for their work on the periodic table.

©FLHC Y1/Alamy Stock Photo

Organization of the Periodic Table

One common version of the modern periodic table appears in Figure 2.9 (and inside the front cover). It is formatted as follows:

1. Each element has a box that contains its atomic number, atomic symbol, and atomic mass. (The mass in parentheses is the mass number of the most stable isotope of the element.) The boxes lie, from left to right, in order of *increasing atomic number* (number of protons in the nucleus).
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 18.
3. Groups 1, 2, 13, 14, 15, 16, 17, and 18 (two on the left and six on the right) contain the *main-group elements*. Groups 3 to 12 contain the *transition elements*. Two horizontal series of *inner transition elements*, the lanthanoids and the actinoids, fit *between* the elements in group 3 and group 4 and are placed below the main body of the table.

FIGURE 2.9 The modern periodic table. As of December 1, 2018, all of the elements up to number 118 (oganesson, Og) have been identified, named, and given chemical symbols. With this confirmation, the last row, corresponding to energy level 7, is complete. Any new element will come into an entirely new energy level.

Key:

atomic number **Symbol** **name** conventional atomic weight standard atomic weight

1	2	3	4	5	6	7	8	9	10	11	12	13	14	15	16	17	18
1 **H** hydrogen 1.008 [1.0078, 1.0082]																	2 **He** helium 4.0026
3 **Li** lithium 6.94 [6.938, 6.997]	4 **Be** beryllium 9.0122											5 **B** boron 10.81 [10.806, 10.821]	6 **C** carbon 12.011 [12.009, 12.012]	7 **N** nitrogen 14.007 [14.006, 14.008]	8 **O** oxygen 15.999 [15.999, 16.000]	9 **F** fluorine 18.998	10 **Ne** neon 20.180
11 **Na** sodium 22.990	12 **Mg** magnesium 24.305 [24.304, 24.307]											13 **Al** aluminium 26.982	14 **Si** silicon 28.085 [28.084, 28.086]	15 **P** phosphorus 30.974	16 **S** sulfur 32.06 [32.059, 32.076]	17 **Cl** chlorine 35.45 [35.446, 35.457]	18 **Ar** argon 39.95 [39.792, 39.963]
19 **K** potassium 39.098	20 **Ca** calcium 40.078(4)	21 **Sc** scandium 44.956	22 **Ti** titanium 47.867	23 **V** vanadium 50.942	24 **Cr** chromium 51.996	25 **Mn** manganese 54.938	26 **Fe** iron 55.845(2)	27 **Co** cobalt 58.933	28 **Ni** nickel 58.693	29 **Cu** copper 63.546(3)	30 **Zn** zinc 65.38(2)	31 **Ga** gallium 69.723	32 **Ge** germanium 72.630(8)	33 **As** arsenic 74.922	34 **Se** selenium 78.971(8)	35 **Br** bromine 79.904 [79.901, 79.907]	36 **Kr** krypton 83.798(2)
37 **Rb** rubidium 85.468	38 **Sr** strontium 87.62	39 **Y** yttrium 88.906	40 **Zr** zirconium 91.224(2)	41 **Nb** niobium 92.906	42 **Mo** molybdenum 95.95	43 **Tc** technetium	44 **Ru** ruthenium 101.07(2)	45 **Rh** rhodium 102.91	46 **Pd** palladium 106.42	47 **Ag** silver 107.87	48 **Cd** cadmium 112.41	49 **In** indium 114.82	50 **Sn** tin 118.71	51 **Sb** antimony 121.76	52 **Te** tellurium 127.60(3)	53 **I** iodine 126.90	54 **Xe** xenon 131.29
55 **Cs** caesium 132.91	56 **Ba** barium 137.33	57–71 lanthanoids	72 **Hf** hafnium 178.49(2)	73 **Ta** tantalum 180.95	74 **W** tungsten 183.84	75 **Re** rhenium 186.21	76 **Os** osmium 190.23(3)	77 **Ir** iridium 192.22	78 **Pt** platinum 195.08	79 **Au** gold 196.97	80 **Hg** mercury 200.59	81 **Tl** thallium 204.38 [204.38, 204.39]	82 **Pb** lead 207.2	83 **Bi** bismuth 208.98	84 **Po** polonium	85 **At** astatine	86 **Rn** radon
87 **Fr** francium	88 **Ra** radium	89–103 actinoids	104 **Rf** rutherfordium	105 **Db** dubnium	106 **Sg** seaborgium	107 **Bh** bohrium	108 **Hs** hassium	109 **Mt** meitnerium	110 **Ds** darmstadtium	111 **Rg** roentgenium	112 **Cn** copernicium	113 **Nh** nihonium	114 **Fl** flerovium	115 **Mc** moscovium	116 **Lv** livermorium	117 **Ts** tennessine	118 **Og** oganesson

57 **La** lanthanum 138.91	58 **Ce** cerium 140.12	59 **Pr** praseodymium 140.91	60 **Nd** neodymium 144.24	61 **Pm** promethium	62 **Sm** samarium 150.36(2)	63 **Eu** europium 151.96	64 **Gd** gadolinium 157.25(3)	65 **Tb** terbium 158.93	66 **Dy** dysprosium 162.50	67 **Ho** holmium 164.93	68 **Er** erbium 167.26	69 **Tm** thulium 168.93	70 **Yb** ytterbium 173.05	71 **Lu** lutetium 174.97
89 **Ac** actinium	90 **Th** thorium 232.04	91 **Pa** protactinium 231.04	92 **U** uranium 238.03	93 **Np** neptunium	94 **Pu** plutonium	95 **Am** americium	96 **Cm** curium	97 **Bk** berkelium	98 **Cf** californium	99 **Es** einsteinium	100 **Fm** fermium	101 **Md** mendelevium	102 **No** nobelium	103 **Lr** lawrencium

Classifying the Elements

One of the clearest ways to classify the elements is as metals, nonmetals, and metalloids. The "imaginary staircase" line that runs from under boron in group 13, down to germanium, and then across and under antimony and polonium is a dividing line:

- The **metals** (*blue* in Figure 2.10) lie in the large lower-left portion of the table. About three quarters of the elements are metals, including many main-group elements, all of the transition and inner transition elements, the lanthanoids, and the actinoids. They are generally shiny solids at room temperature (mercury and francium are liquid) and conduct heat and electricity well. They can be tooled into sheets (are malleable) and wires (are ductile).
- 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 (semimetals)** (*green*), which lie along the imaginary staircase line, have properties between those of metals and nonmetals.

Figure 2.10 shows examples of these three classes of elements.

Keep in mind the following two major points:

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 1, except for hydrogen, consists of the *alkali metals*, and group 2 consists of the *alkaline earth metals*. The elements in both of these groups are highly reactive. Group 16, with oxygen as the first element, is called the *chalcogens*. The *halogens*, group 17, are highly reactive nonmetals, whereas the *noble gases*, group 18, are relatively unreactive nonmetals. Other main groups (13 to 15) are often named for the first element in the group; for example, group 15 is the *nitrogen family*.

Two of the major branches of chemistry have traditionally been defined by the elements that are studied. *Organic chemistry* focuses on the compounds of carbon, specifically those that contain hydrogen and often oxygen, nitrogen, and a few other

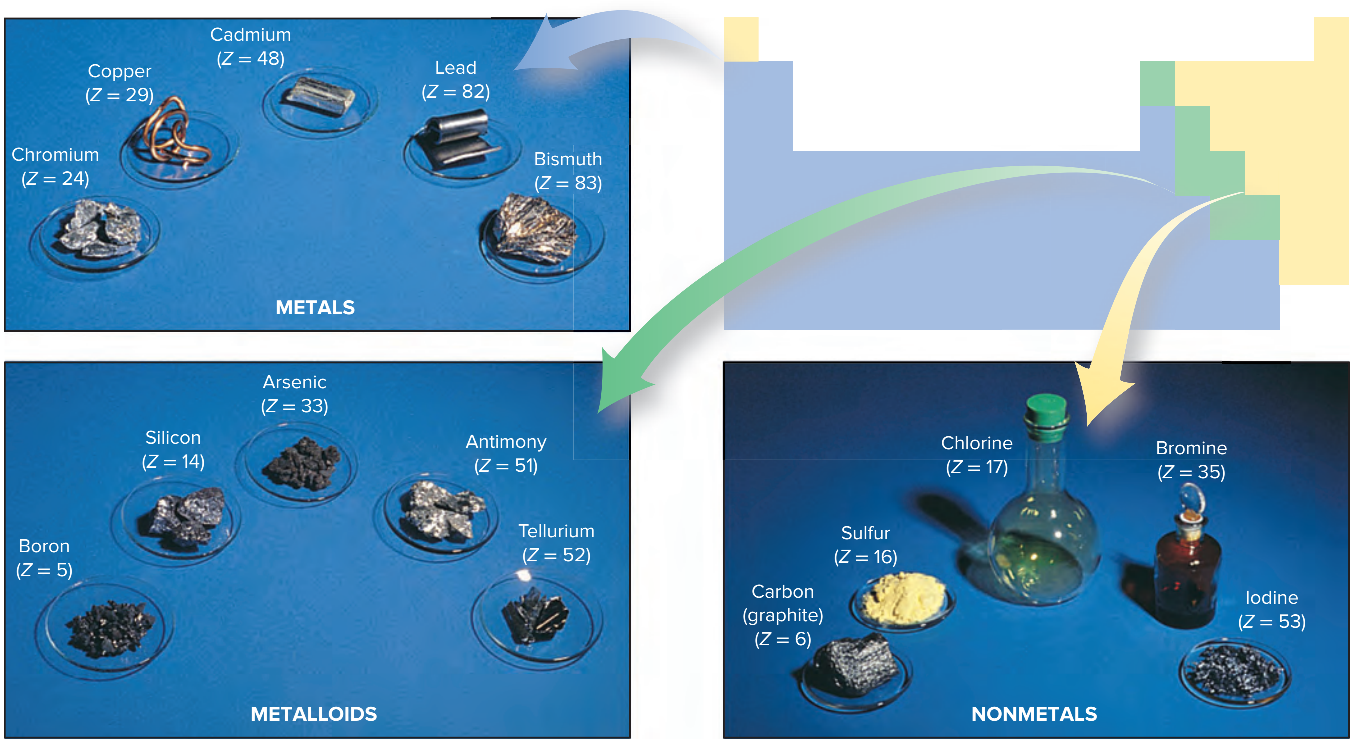
groups
The vertical columns in the periodic table; elements in a group usually have the same outer electron configuration and, thus, similar chemical behaviour

metals
Substances or mixtures that are relatively shiny and malleable and are good conductors of heat and electricity; in reactions, metals tend to transfer electrons to nonmetals and form ionic compounds

nonmetals
Elements that lack metallic properties; in reactions, nonmetals tend to share electrons with each other to form covalent compounds or accept electrons from metals to form ionic compounds

metalloids (semimetals)
Elements with properties between those of metals and nonmetals

FIGURE 2.10 Some metals, metalloids, and nonmetals



ionic compounds
Compounds that consist of oppositely charged ions

covalent compounds
Compounds that consist of atoms bonded together by shared electron pairs

elements. This branch is concerned with fuels, drugs, and dyes, to name only a few. *Inorganic chemistry*, on the other hand, focuses on the compounds of all of the other elements and is concerned with catalysts, electronic materials, metal alloys, mineral salts, and many other materials. With the explosive growth in biomedical and materials sciences, the line between these branches has all but disappeared.

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 an imaginary staircase line, and metals fill the rest of the table.
- Elements within a group have similar behaviour, whereas elements within a period have dissimilar behaviour.

2.7 Compounds: An Introduction to Bonding

Only a few elements occur unbound, or as single atoms, in nature. The noble gases—helium (He), neon (Ne), argon (Ar), krypton (Kr), xenon (Xe), and radon (Rn)—occur in air in atomic form. In addition to occurring in compounds, oxygen (O), nitrogen (N), and sulfur (S) are generally present as elementary substances. O_2 and N_2 form diatomic molecules, sulfur is often present as the polyatomic S_8, and carbon (C) occurs in vast, nearly pure deposits of coal and diamond. Certain metals—copper (Cu), silver (Ag), gold (Au), and platinum (Pt)—are sometimes found uncombined. Aside from these few exceptions, *the overwhelming majority of elements exist as compounds with other elements.*

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

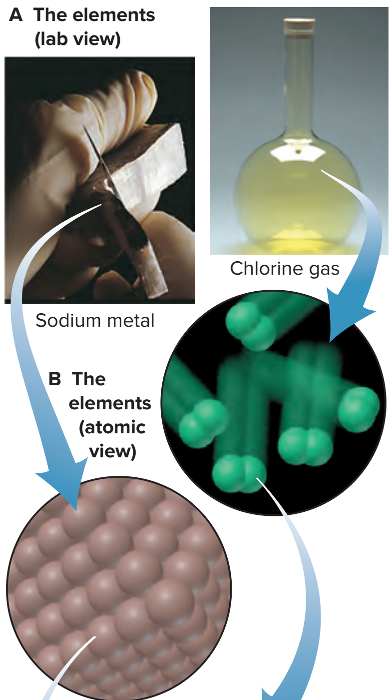
1. *Transferring electrons* from one element to another to form **ionic compounds**
2. *Sharing electrons* between atoms of different elements to form **covalent compounds**

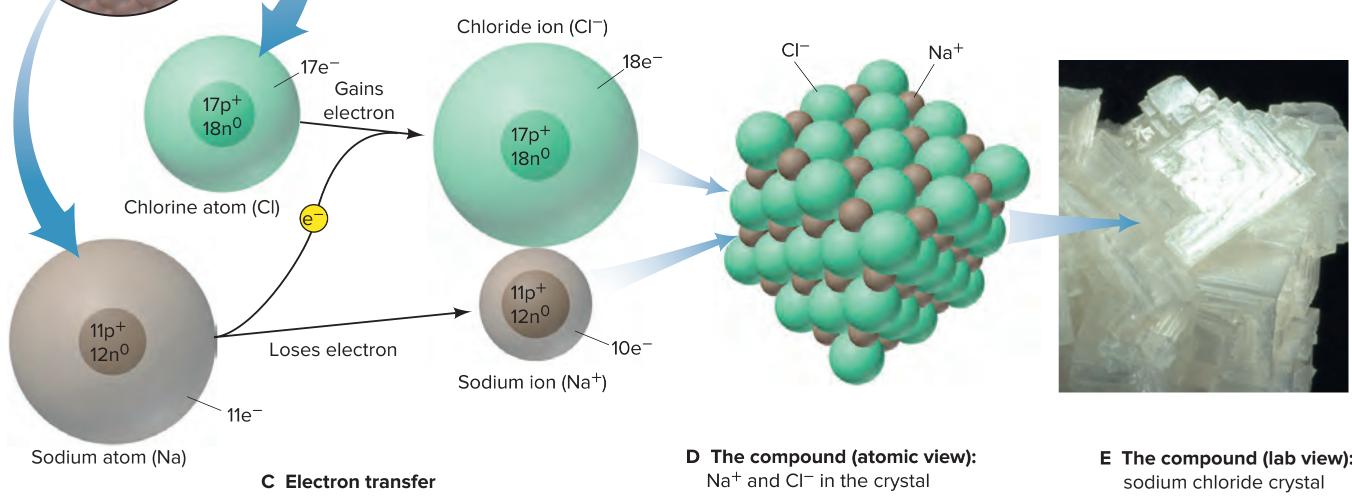
These processes generate **chemical bonds**, *the forces that hold the atoms together in a compound.* This section introduces compound formation, which we will discuss in much more detail in later chapters.

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

FIGURE 2.11 The formation of an ionic compound. A. The two elements as seen in a laboratory. **B.** The elements on the atomic scale. **C.** The electron transfer from Na atom to Cl atom forms 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.





of ionic compound is a **binary ionic compound**, *a compound 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 electrons and becomes an **anion**, *a negatively charged ion*.

Binary ionic compounds can form in many ways: one way involves a *transfer of electrons* from the metal atoms to the nonmetal atoms. These compounds may also be formed as a result of acid/base or precipitation reactions. The resulting large numbers of cations and anions attract each other and form an ionic compound. *A cation or anion derived from a single atom* is called a **monatomic ion**; we will discuss polyatomic ions, those derived from a small group of atoms, later.

The Case of Sodium Chloride *All binary 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 shown in Figure 2.11. 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 Section 2.8.) 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 can be described by *Coulomb's law: 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*:

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

This can be summarized as shown in Figure 2.12 and expressed as follows:

- 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. However, 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 that a given atom will lose or gain when it forms an ion? For elements in groups 1, 2, and 13 to 18, 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 18). Noble gases have a stability that is related to their number (and arrangement) of electrons. Thus, a sodium atom ($11e^-$) can attain the stability of a neon atom ($10e^-$), the nearest noble gas, by losing one electron. Similarly, a chlorine atom ($17e^-$) can attain the stability of an argon atom ($18e^-$), its nearest noble gas, by gaining one electron. Thus, in general, elements located near a noble gas form monatomic ions as follows:

- *Metals lose electrons*: Elements in group 1 lose one electron, elements in group 2 lose two electrons, and aluminum in group 13 loses three electrons.
- *Nonmetals gain electrons*: Elements in group 17 gain one electron, oxygen and sulfur in group 16 gain two electrons, and nitrogen in group 15 gains three electrons.

chemical bonds
The forces that hold atoms together in a molecule (or formula unit)

ions
Charged particles that form from an atom (or covalently bonded group of atoms) when it gains or loses one or more electrons

binary ionic compound
A compound that consists of the oppositely charged ions of two elements

cation
A positively charged ion

anion
A negatively charged ion

monatomic ion
An ion derived from a single atom

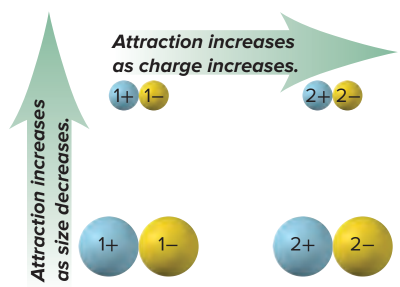


FIGURE 2.12 Factors that influence the strength of ionic bonding

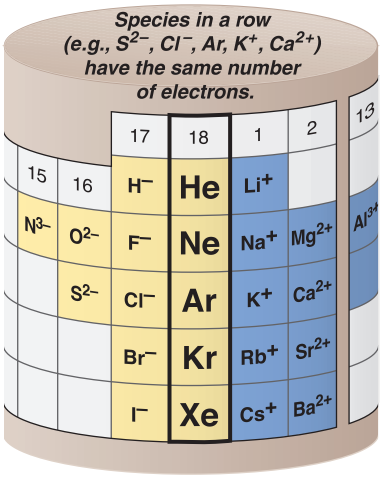


FIGURE 2.13 **The relationship between the ion formed and the nearest noble gas**

In the periodic table in Figure 2.9, the elements in group 17 appear to be "closer" to the noble gases than the elements in group 1. In truth, both groups are only one electron away from the number of electrons in the nearest noble gas. Figure 2.13 shows a periodic table of monatomic ions that is cut and rejoined as a cylinder. Note that fluorine (F; $Z = 9$) has one electron *less* than the noble gas neon (Ne; $Z = 10$), and sodium (Na; $Z = 11$) has one electron *more*; thus, fluorine and sodium form the F^- and Na^+ ions. Similarly, oxygen (O; $Z = 8$) gains two electrons and magnesium (Mg; $Z = 12$) loses two electrons to form the O^{2-} and Mg^{2+} ions and attain the same number of electrons as neon. In Figure 2.13, notice that species in a row have the same number of electrons.

Sample Problem 2.6 Predicting the Ion That an Element Forms

Problem Which monatomic ions does each element form?

(a) Iodine ($Z = 53$) **(b)** Calcium ($Z = 20$) **(c)** Aluminum ($Z = 13$)

Plan We use the given value of Z to find the element in the periodic table and see where its group lies relative to the noble gases. Elements in groups 1, 2, and 3 *lose* electrons to attain the same number of electrons as the nearest noble gas and become positive ions; those in groups 15, 16, and 17 *gain* electrons and become negative ions.

Solution **(a)** I^- Iodine ($_{53}I$) is in group 17, the halogens. Like any member of this group, it gains one electron to attain the same number of electrons as the nearest group 18 member, in this case, $_{54}Xe$.

(b) Ca^{2+} Calcium ($_{20}Ca$) is in group 2, the alkaline earth metals. Like any group 2 member, it loses two 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 13) and, thus, loses three electrons to attain the same number as its nearest noble gas, $_{10}Ne$.

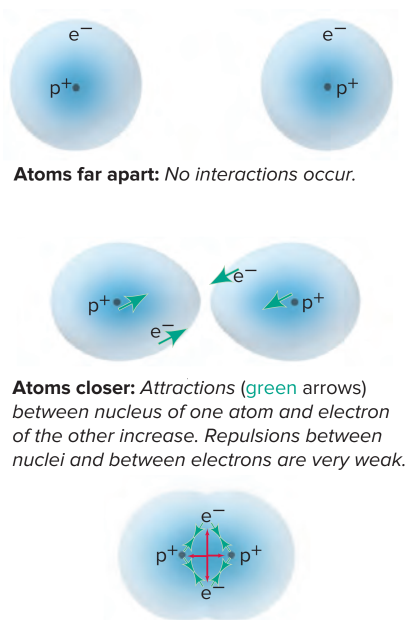
Follow-Up Problem 2.6 Which monatomic ion does each element form?
(a) $_{16}S$ **(b)** $_{37}Rb$ **(c)** $_{56}Ba$

covalent bond
A type of bond in which atoms are bonded through the sharing of electrons; the mutual attraction of the nuclei and an electron pair that holds atoms together in a molecule

The Formation of Covalent Compounds

Covalent compounds form when elements, usually nonmetals, share electrons. The simplest case of electron sharing occurs between two hydrogen atoms (H; $Z = 1$). Imagine two separated H atoms approaching each other (Figure 2.14). As they get closer, the nucleus of each atom attracts the electron of the other atom more and more strongly. As the separated atoms begin to interpenetrate, repulsions between the nuclei and between the electrons begin to increase. At some optimum distance between the nuclei, 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 2.15 shows other elementary substances that exist as molecules at room temperature.

Atoms of elements, whether the same element or 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:



Atoms far apart: *No interactions occur.*

Atoms closer: *Attractions* (green arrows) *between nucleus of one atom and electron of the other increase. Repulsions between nuclei and between electrons are very weak.*

Optimum distance: *H_2 molecule forms because attractions* (green arrows) *balance repulsions* (red arrows).

FIGURE 2.14 **Formation of a covalent bond between two H atoms**

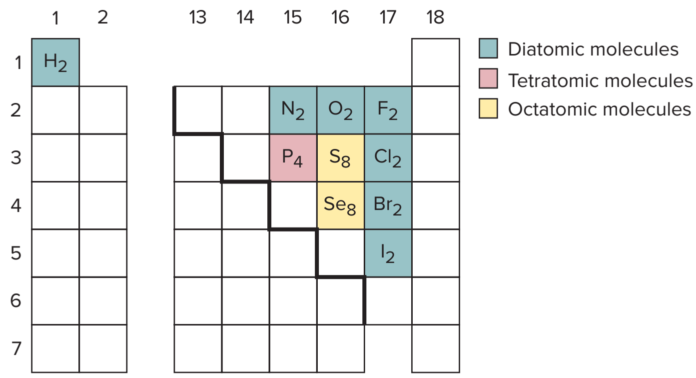


FIGURE 2.15 Elements that occur as molecules

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 in close proximity to one another. 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." Thus, we *cannot* isolate a *molecule* of sodium chloride.

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 between 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 2.16 shows that a crystalline form of calcium carbonate (*left*) occurs, on the atomic scale, as an array of polyatomic carbonate anions and monatomic calcium cations (*centre*). 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.

polyatomic ions
Ions in which two or more atoms are bonded covalently

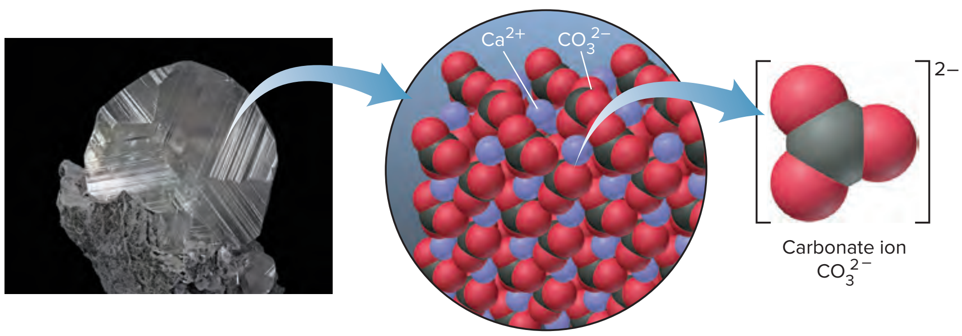


FIGURE 2.16 The carbonate ion in calcium carbonate

SUMMARY OF SECTION 2.7

- Although a few elements occur in their atomic form in nature, the great majority exist as 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.

chemical formula

A notation of atomic symbols and numerical subscripts that shows the type and number of each atom in a molecule or formula unit of a substance

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

*The ions are listed by charge; those in boldface are the most common.

2.8 Formula, Name, and Mass of a Compound

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 will learn how to write the name and formula of ionic compounds and simple covalent compounds, how to calculate the mass of a compound from its formula, and how to visualize molecules with three-dimensional models. To make learning the name and formula of a compound easier, we will rely on various rules. So, be prepared for a bit of memorization and a lot of practice!

Binary Ionic Compounds

Let us begin with two general rules:

- For *all* ionic compounds, *the name and formula 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 (groups 1 and 2), it usually forms a single type of ion; if it is a transition element (groups 3 to 12), it often forms more than one type of ion. We will discuss each case in turn. Some of the more common monatomic ions are presented in Table 2.3.

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 on.
- For cations in groups 1 and 2, the ion charge equals the group number: Na is in group 1 and forms Na^+; Ba is in group 2 and forms Ba^{2+}. (Exceptions in Figure 2.17 are Sn^{2+} and Pb^{2+}.)
- For anions, the ion charge equals the group number minus 18; for example, S is in group 16 (16 − 18 = −2) and, thus, forms S^{2-}.

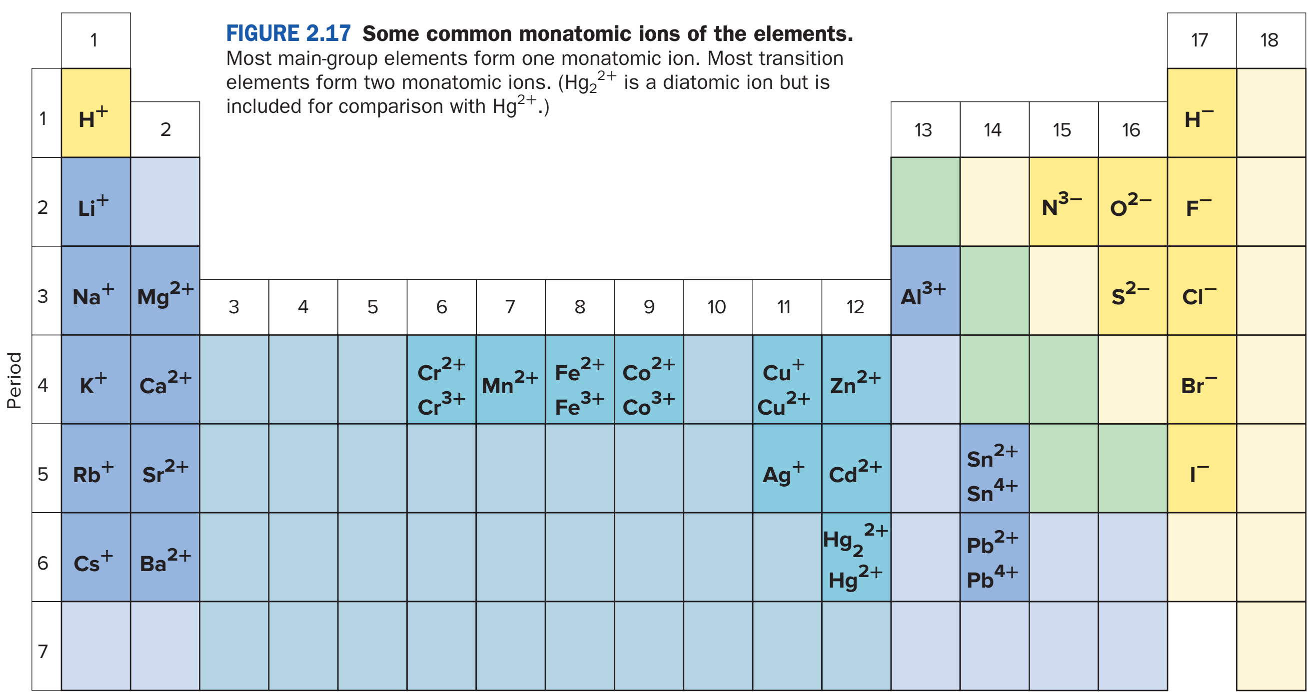


FIGURE 2.17 Some common monatomic ions of the elements. Most main-group elements form one monatomic ion. Most transition elements form two monatomic ions. (Hg_2^{2+} is a diatomic ion but is included for comparison with Hg^{2+}.)

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. For help in writing the formula of this and other compounds, remember the following rules:

formula unit
The chemical unit of a compound that contains the relative numbers of the types of atoms or ions expressed in the chemical formula for the compound

- 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:

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

- The subscripts need to be reduced 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.*

The following two sample problems apply the rules we just discussed. In Sample Problem 2.7, we name the compound from its elements. In Sample Problem 2.8, we find the formula.

Sample Problem 2.7 Naming Binary Ionic Compounds

Problem Name the ionic compound formed from each pair 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)** Magnesium is the metal; *nitr-* is the nonmetal root: magnesium nitride.

(b) Cadmium is the metal; *iod-* is the nonmetal root: cadmium iodide.

(c) Strontium is the metal; *fluor-* is the nonmetal root: strontium fluoride.

(d) Cesium is the metal; *sulf-* is the nonmetal root: cesium sulfide.

Follow-Up Problem 2.7 For the following ionic compounds, give the name of each element and its periodic table group number:

(a) Zinc oxide

(b) Silver bromide

(c) Lithium chloride

(d) Aluminum sulfide

Sample Problem 2.8 Determining the Formula of a Binary Ionic Compound

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

Plan We write a formula by finding the smallest number of each ion that gives the neutral compound. This number appears as a *right subscript* to the element symbol.

*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; in fact, reducing the subscripts for these compounds would give the incorrect formula for each, namely HgCl and NaO.

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 formula for a compound; thus, in (c), we do *not* write $Sr^{2+}F_2{}^-$.

Follow-Up Problem 2.8 Write the formulas for the compounds named in Follow-Up Problem 2.7.

Compounds with Metals That Form More than One Ion As noted earlier, many metals, particularly the transition elements (groups 3 to 12), can form more than one ion. Table 2.4 lists some examples; see Figure 2.17 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 the common names of some 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 called ferr*ic* chloride. (*Memory aid*: There is an ***o*** in -*o*us and *lower*, and an ***i*** in -*ic* and *higher*.)

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

*The ions are listed alphabetically by metal name; those in boldface are the most common.
**This ion has been included in both this table and the table for polyatomic ions (Table 2.5). It is a single element that has been included in this table for reasons of completeness for the charge on the mercury ion and in Table 2.5 because there is more than one atom in the ion.

Sample Problem 2.9 Determining the Name and Formula of Ionic Compounds of Metals That Form More than One Ion

Problem Give the systematic name for the given formula or the formula for the given name 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 name for the formula or the formula for the given name of each compound:
(a) Lead(IV) oxide **(b)** Cu_2S **(c)** $FeBr_2$ **(d)** Mercuric chloride

Compounds That Contain Polyatomic Ions

Many ionic compounds contain polyatomic ions. Table 2.5 shows some common polyatomic ions. Remember that *a polyatomic ion stays together as a charged unit.* For example, 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 a formula unit, this 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 used *only* 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*), *ions 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.

For families with two oxoanions:

- 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*.

For families with four oxoanions (a halogen bonded to O) (Figure 2.18):

- The ion with the *most* O atoms has the prefix *per-*, the nonmetal root, and the suffix *-ate.*
- The ion with *one less* 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 the *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*, and 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 centred dot in the formula and noted in the name by a Greek numerical prefix before the word *hydrate.* Table 2.6 (next page) shows these prefixes. For example, Epsom salt has seven water

TABLE 2.5 Common Polyatomic Ions*

Formula	Name
Cations	
$\mathbf{NH_4^+}$	**Ammonium**
$\mathbf{H_3O^+}$	**Hydronium**
$\mathbf{Hg_2^{2+}}$	**Mercury(I)**
Anions	
$\mathbf{CH_3COO^-}$ (or $\mathbf{C_2H_3O_2^-}$)	**Ethanoate (or 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 the most common.

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

FIGURE 2.18 Naming oxoanions. Prefixes and suffixes indicate the number of oxygen (O) atoms in the anion.

oxoanions
Anions in which an element is bonded to one or more oxygen atoms

hydrates
Compounds in which a specific number of water molecules are associated with each formula unit

TABLE 2.6 Numerical Prefixes for 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-

molecules in each formula unit: its formula is $MgSO_4 \cdot 7H_2O$, and its 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 the Name and Formula for an Ionic Compound That Contains Polyatomic Ions

Problem Give the systematic name for the given formula or the formula for the given name of each compound:

(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 name for the formula or the formula for the given name of each compound:

(a) Cupric nitrate trihydrate **(b)** Zinc hydroxide **(c)** LiCN

Sample Problem 2.11 Recognizing the Incorrect Name and Formula of an Ionic Compound

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.

Acid Names from Anion Names

Acids are an important group of hydrogen-containing compounds that have been used in chemical reactions since before alchemical times. In the laboratory, acids are typically used in water to form aqueous solutions. When naming an acid and writing its formula, 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*:

   prefix *hydro-* + nonmetal *root* + suffix *-ic* + separate word *acid*
   hydro + chlor + ic + 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 the names of the oxoanions, except for two suffix changes (Figure 2.19):
   - The *-ate* in the anion becomes *-ic* in the acid.
   - The *-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.

   (Memory aid: There is an ***o*** in -*o*us and *lo*wer, and an ***i*** in -*i*c and *hi*gher.)

No. of O atoms

	Prefix	Root	Suffix
	per	*root*	ic acid
		root	ic acid
		root	ous acid
	hypo	*root*	ous acid

FIGURE 2.19 Naming oxoacids. Prefixes and suffixes indicate the number of oxygen (O) atoms in the anion.

Sample Problem 2.12 Determining the Name and Formula for Anions and Acids

Problem Name each anion, and give the name and formula for 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 the name for the formula for 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—such as ammonia (NH_3), acetic acid (CH_3COOH), and water (H_2O)—are so familiar that we use their common names, 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*. For example, nitrogen (group 15) and fluorine (group 17) form a compound that has three fluorine atoms for every nitrogen atom. The name of the compound is nitrogen trifluoride, and the formula is NF_3.

binary covalent compounds
Compounds that consist of atoms of two elements, in which bonding occurs primarily through electron sharing

(*Exception*: When the compound contains oxygen and any of the halogens chlorine, bromine, and iodine, the halogen is named first.)

- If both elements are in the same group, the element with the higher period number is named first. Thus, the group 16 elements sulfur (period 3) and oxygen (period 2) form sulfur dioxide, SO_2.
- 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. When the second element name begins with a vowel, we usually drop the vowel attached to the prefix. For example, we say dinitrogen tetroxide, not dinitrogen tetraoxide.

Sample Problem 2.13 Determining the Name and Formula of a Binary Covalent Compound

Problem **(a)** What is the formula for carbon disulfide?

(b) What is the name of PCl_5?

(c) Each molecule in a compound consists of two N atoms and four O atoms. Give the name and formula for the compound.

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 each compound:

(a) SO_3 **(b)** SiO_2 **(c)** Dinitrogen monoxide **(d)** Selenium hexafluoride

Sample Problem 2.14 Recognizing the Incorrect Name and Formula of a Binary Covalent Compound

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.

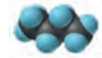
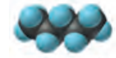
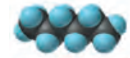
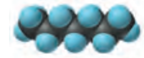
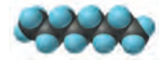
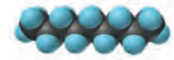
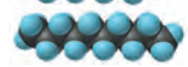
Follow-Up Problem 2.14 Explain what is wrong with the name or formula at the end of each statement, and correct it: **(a)** S_2Cl_2 is disulfurous dichloride. **(b)** Nitrogen monoxide is N_2O. **(c)** $BrCl_3$ is trichlorine bromide.

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 will lay the groundwork for naming organic compounds by focusing on the simplest ones. Rules for naming more complex organic compounds are detailed in Chapter 20.

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 constitutes 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 name, molecular formula, and space-filling model (discussed shortly) of the first 10 straight-chain alkanes. Note that the roots of the four smallest alkanes are new, but the roots of the larger alkanes 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 ($C_{10}H_{22}$)	

Masses from a Chemical Formula

In Section 2.5, we calculated the atomic mass of an element. Using the periodic table and the formula for a compound, we can calculate the *mass* of a formula unit of a compound as *the sum of the atomic masses*:

$$\text{Mass of formula unit} = \text{sum of atomic masses} \qquad (2.3)$$

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

$$\begin{aligned}\text{Mass of } H_2O &= 2 \times \text{atomic mass of H} + 1 \times \text{atomic mass of O} \\ &= 2 \times 1.008\ \text{u} + 16.00\ \text{u} = 18.02\ \text{u}\end{aligned}$$

Ionic compounds do not consist of molecules, so *the mass of a formula unit* is termed the *formula mass*, not the *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, $Ba(NO_3)_2$,

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

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 the mass of a compound.

Sample Problem 2.15 Calculating the Mass of a Compound

Problem Using the periodic table, calculate the mass of

(a) Tetraphosphorus trisulfide **(b)** Ammonium nitrate

Plan We first write the formula. Then we 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:

$$\begin{aligned}\text{Mass of } P_4S_3 &= 4 \times \text{atomic mass of P} + 3 \times \text{atomic mass of S} \\ &= 4 \times 30.97\ \text{u} + 3 \times 32.07\ \text{u} = 220.09\ \text{u}\end{aligned}$$

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

$$\begin{aligned}&\text{Mass of } NH_4NO_3 \\ &\quad = 2 \times \text{atomic mass of N} + 4 \times \text{atomic mass of H} + 3 \times \text{atomic mass of O} \\ &\quad = 2 \times 14.01\ \text{u} + 4 \times 1.008\ \text{u} + 3 \times 16.00\ \text{u} = 80.05\ \text{u}\end{aligned}$$

Check You can often find large errors by rounding atomic masses to the nearest five 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 decimal places.

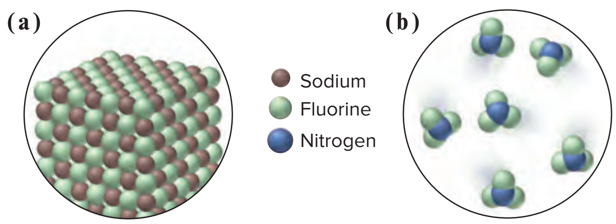
(b) $2 \times 15 + 4 \times 1 + 3 \times 15 \approx 80.05$

Follow-Up Problem 2.15 Find the molecular (or formula) mass of each compound:

(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 diagram represents a binary compound. Determine the formula, name, and mass of each compound.



Plan Each compound 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 mass.

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

$$\begin{aligned}\text{Mass of NaF} &= 1 \times \text{atomic mass of Na} + 1 \times \text{atomic mass of F}\\ &= 22.99\ \text{u} + 19.00\ \text{u} = 41.99\ \text{u}\end{aligned}$$

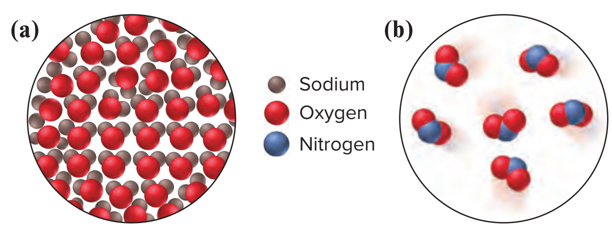
(b) There are three green spheres (fluorine) for each blue sphere (nitrogen), so the formula is NF_3. Two nonmetals form a covalent compound. Nitrogen has a lower group number, so it is named first: nitrogen trifluoride:

$$\begin{aligned}\text{Mass of } NF_3 &= 1 \times \text{atomic mass of N} + 3 \times \text{atomic mass of F}\\ &= 14.01\ \text{u} + 3 \times 19.00\ \text{u} = 71.01\ \text{u}\end{aligned}$$

Check **(a)** For binary ionic compounds, we predict ionic charges from the periodic table (see Figure 2.13). 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 diagram. **(b)** Covalent compounds often occur as individual molecules, as shown in the diagram.

Rounding gives $25 + 20 = 45$ in (a) and $15 + 3 \times 20 = 75$ in (b), so there are no large errors.

Follow-Up Problem 2.16 Each diagram below represents a binary compound. Determine the name, formula, and mass of the compound.



Representing Molecules with a Formula and a Model

To represent objects that are too small to see, chemists often use a formula and a model, of which there can be different types. For the sake of visualization and simplicity, common elements are colour coded, as shown in Figure 2.20. Each type of model conveys different information, as shown for water below:

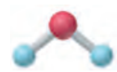
- A **molecular formula** *uses element symbols and often numerical subscripts to give the* ***actual*** *number of atoms of each element in a molecule of the compound.* (Recall that, for ionic compounds, the *formula unit* gives the *relative* number of each type of ion.) The molecular formula for water 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 the atoms in a 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 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, coloured balls represent atoms. A *ball-and-stick model* shows atoms as balls and bonds as sticks. 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 does not show the bonded atoms overlapping (see Figure 2.14) or their relative sizes, so it exaggerates the distance between them.

- A *space-filling model* is an accurately scaled up image of the molecule, so it shows the relative sizes of the atoms, the relative distances between the nuclei (centres of the spheres), and the angles between the bonds. However, the bonds are not shown, and it can be difficult to see each atom in a complex molecule.

Every molecule is minute, but the range of molecular sizes, and thus molecular masses, is enormous. Table 2.8 (next page) shows some diatomic and small polyatomic molecules, as well as two extremely large molecules, called *macromolecules*, deoxyribonucleic acid (DNA) and nylon.

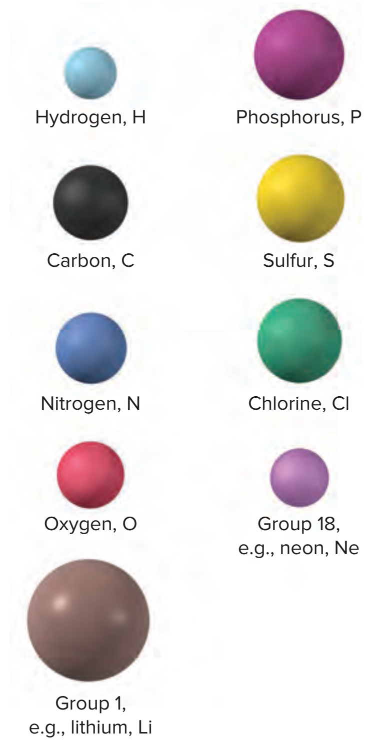


FIGURE 2.20 Although the elements themselves do not have colour, when drawing molecules, standard colours have been adopted for the most common elements used in structures. This allows for the immediate identification of a molecule based on its structure and the atoms it contains.

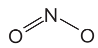
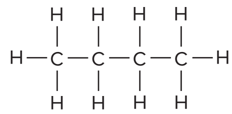
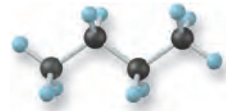
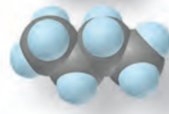
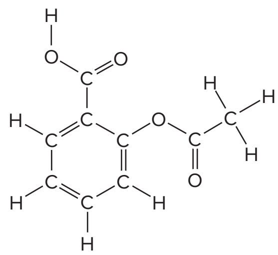
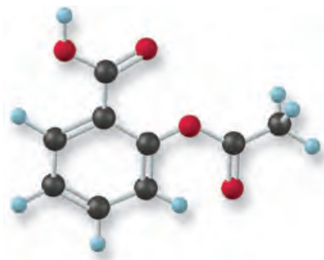
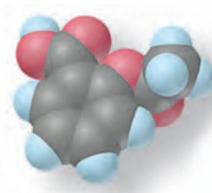
molecular formula
A formula that shows the actual number of atoms of each element in a molecule of a compound; it is always a multiple of the empirical formula

structural formula
A formula that shows the actual number of atoms, their relative placement, and the bonds between them

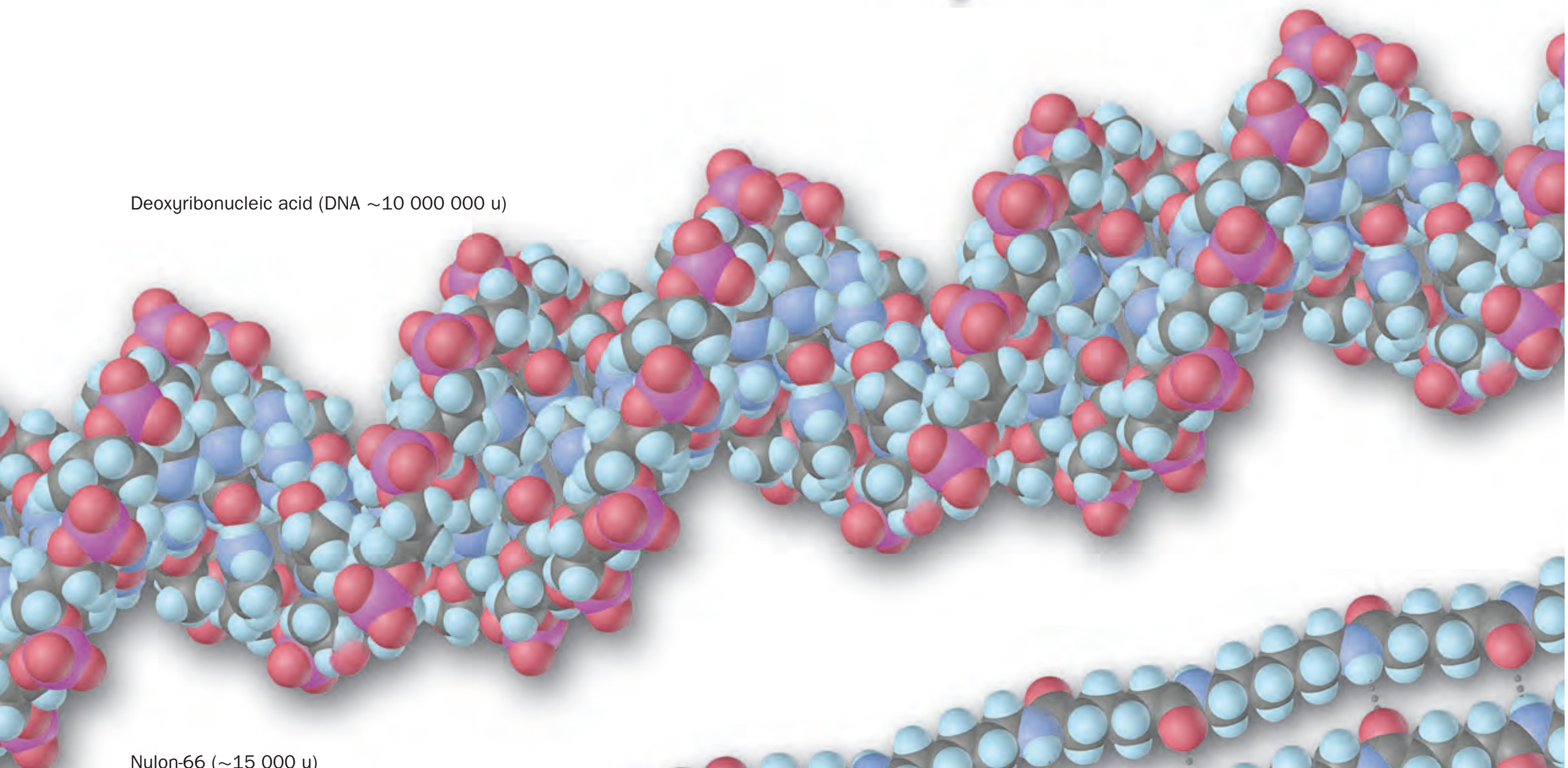
SUMMARY OF SECTION 2.8

- An ionic compound is named with the cation first and the anion second. If a metal can form more than one ion, the charge is shown with a roman numeral.
- The name of an oxoanion has a suffix, and sometimes a prefix, attached to the root of the element name to indicate the number of oxygen atoms.
- The name of a hydrate has a numerical prefix indicating the number of associated water molecules.
- Acid names are based on anion names.
- In the name of a binary covalent compound, the first word is the element farther left or lower down in the periodic table, and a prefix shows the number of each atom.
- The mass of a compound is the sum of the atomic masses.
- A chemical formula gives the number of atoms (molecular) or the arrangement of atoms (structural) of one unit of the compound.
- Molecular models convey information about bond angles (ball-and-stick) and relative atomic sizes and distances between atoms (space-filling).

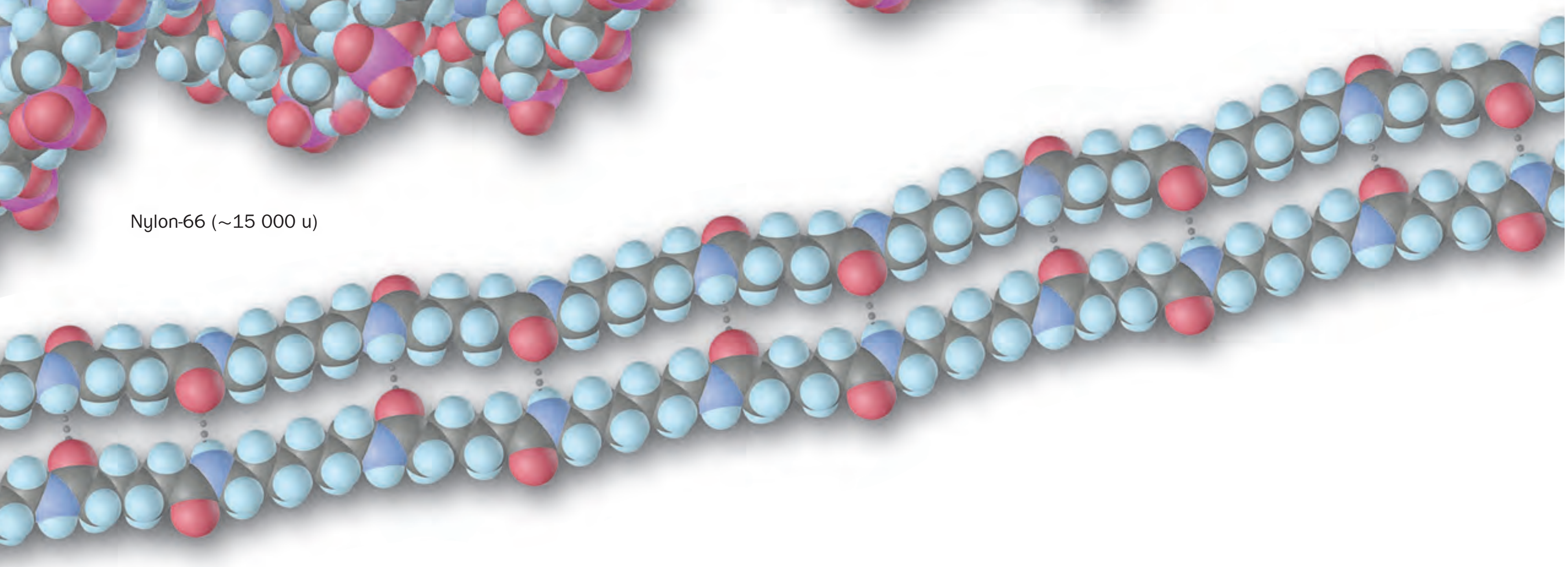
TABLE 2.8 Representing Molecules

Name	Molecular Formula (Mass of molecule, in u)	Bond-Line Formula	Ball-and-Stick Model	Space-Filling Model
Carbon monoxide	CO (28.01)	C≡O		
Nitrogen dioxide	NO_2 (46.01)			
Butane	C_4H_{10} (58.12)			
Aspirin (acetylsalicylic acid)	$C_9H_8O_4$ (180.15)			

Deoxyribonucleic acid (DNA ~10 000 000 u)



Nylon-66 (~15 000 u)



2.9 Mixtures: Classification and Separation

In the natural world, *matter usually occurs as mixtures.* A sample of clean air, for example, consists of many elements and compounds physically mixed together, including O_2, N_2, CO_2, the noble gases (group 18), and water vapour (H_2O). The oceans are complex mixtures of dissolved ions and covalent substances, including Na^+, Mg^{2+}, Cl^-, SO_4^{2-}, O_2, CO_2, and, of course, H_2O. Rocks and soils are mixtures of numerous compounds, including calcium carbonate ($CaCO_3$), silicon dioxide (SiO_2), aluminum oxide (Al_2O_3), and iron(III) oxide (Fe_2O_3). Living things contain thousands of substances: carbohydrates, lipids, proteins, nucleic acids, and many simpler ionic and covalent 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 only be seen with a microscope.
- A **homogeneous mixture (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.

heterogeneous mixture
A mixture that has one or more visible boundaries among its components

homogeneous mixture (solution)
A mixture that has no visible boundaries among its components

aqueous solutions
Solutions in which water is the solvent

Although we usually think of solutions as liquid, they can exist in all of the common 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 constitute a major portion of the environment *and* a major portion 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 2.21). Any proportion of iron metal filings and powdered sulfur forms a mixture. The components can be separated with a magnet because iron metal is magnetic. If we heat the container strongly, however, the components 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.

Chemists have devised many techniques for separating a mixture into its components. Some of the common techniques are described in the Tools of the Laboratory section (page 77).

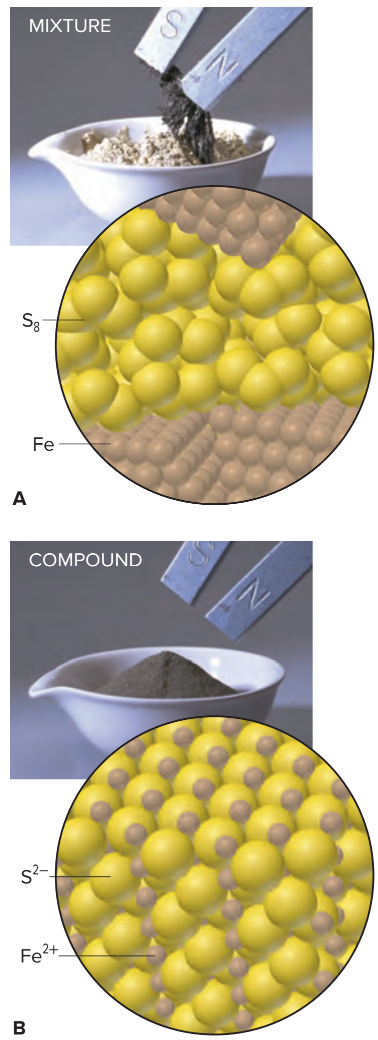


FIGURE 2.21 The distinction between mixtures and compounds. **A.** A *mixture* of iron and sulfur consists of the two elements. **B.** The *compound* iron(II) sulfide consists of an array of Fe^{2+} and S^{2-} ions.

An Overview of the Components of Matter

Understanding matter at both the observable scale and the atomic scale is the essence of chemistry. Figure 2.22 (next page) is a visual overview of many key terms and ideas in this chapter.

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 components of compounds) can have variable proportions, can be separated physically, and retain their properties.
- Separation methods are based on differences in physical properties and include filtration (particle size), crystallization (solubility), distillation (volatility), and chromatography (solubility).

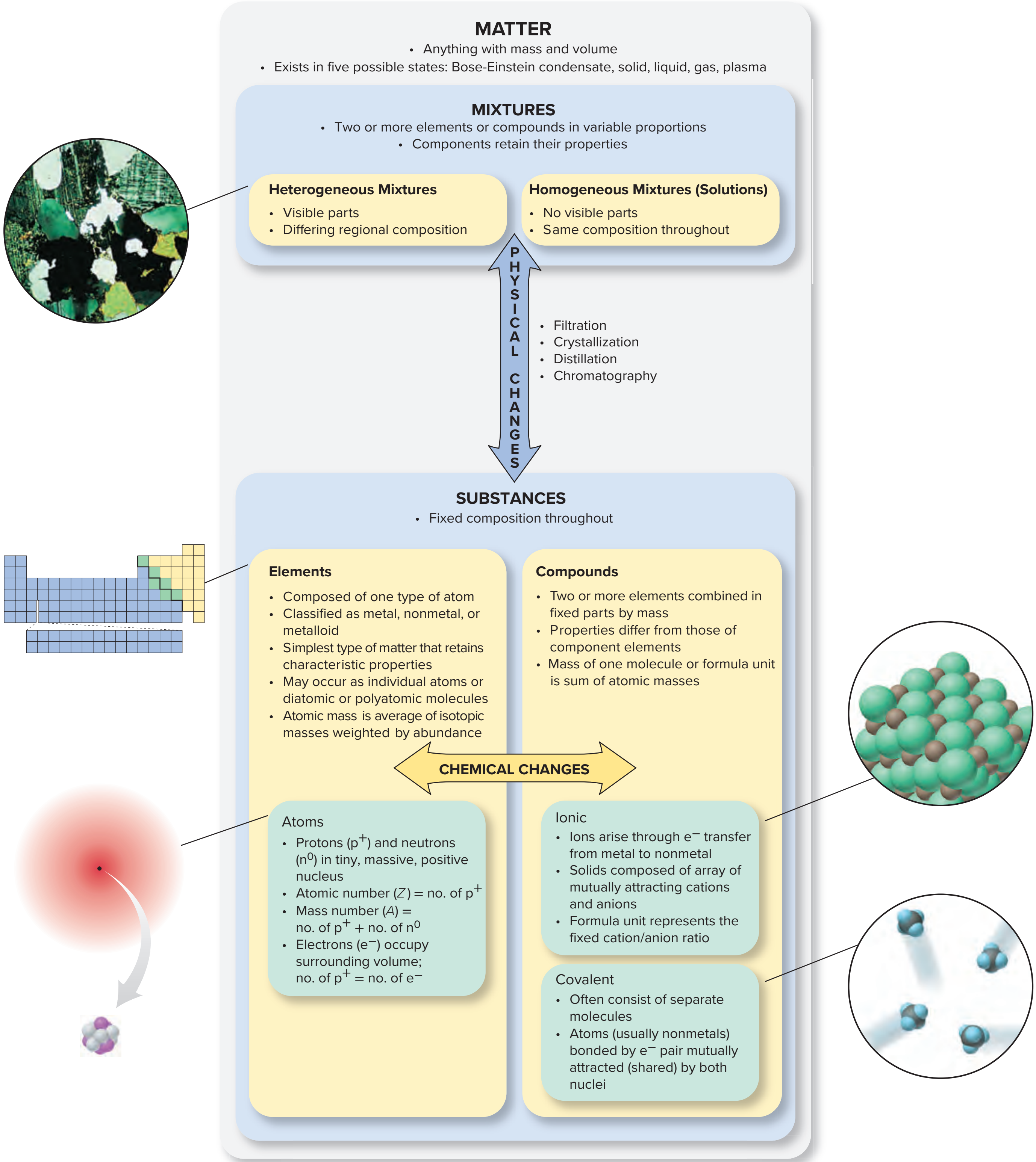


FIGURE 2.22 **The classification of matter from a chemical point of view**

Some of the most challenging laboratory techniques involve separating mixtures and purifying the components. All of the techniques described here depend on the *physical properties* of the substances in the mixture; no chemical changes occur.

filtration
A method of separating the components of a mixture on the basis of differences in particle size

crystallization
A technique used to separate and purify the components of a mixture through differences in solubility, causing a component to come out of solution as crystals

distillation
A separation technique in which a more volatile component of a mixture vaporizes and condenses separately from the less volatile components

Filtration is based on *differences in particle size* and is often used to separate a solid from a liquid. The liquid flows through the tiny holes in filter paper, and the solid is retained. In vacuum filtration, reduced pressure speeds the flow of the liquid through the filter. Filtration is used to purify tap water.

Crystallization is based on *differences in solubility*. The *solubility* of a substance is the amount that dissolves in a fixed volume of solvent at a given temperature. Since solubility often increases with temperature, the impure solid is dissolved in hot solvent. When the solution cools, the purified compound crystallizes. A key component of computer chips is purified by a type of crystallization.

Distillation separates components through *differences in* **volatility**, *the tendency of a substance to become a gas*. *Simple* distillation separates components with *large* differences in volatility, such as water from dissolved ionic compounds (Figure B2.3). As the mixture boils, the vapour is richer in the more volatile component, in this case water, which is condensed and collected separately. *Fractional* distillation uses many vaporization-condensation steps to separate components with small volatility differences, such as those in petroleum (discussed in Chapter 12).

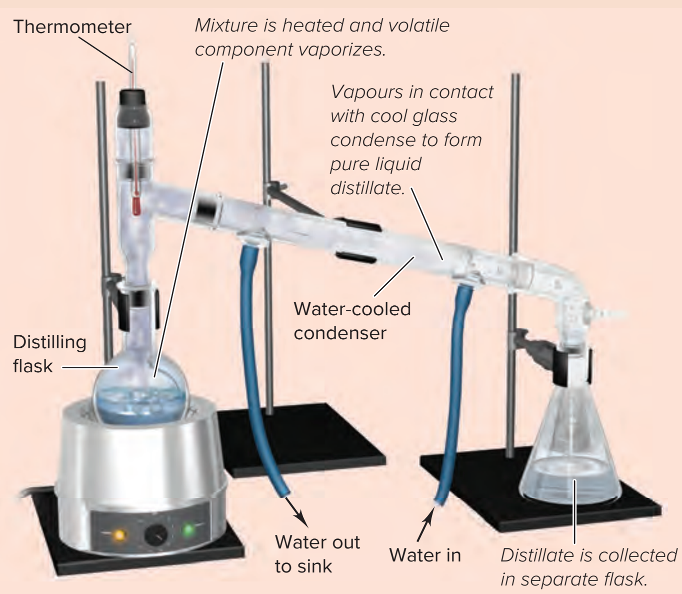


FIGURE B2.3 **Distillation**

Extraction is used to separate the components of a mixture based on *differences* in **solubility (s)**. The mixture from which one component is to be separated is placed in a mixture of two immiscible solvents. The solvents are chosen so that the desired component is soluble in one of the solvents and the other (undesired) components are not. Shaking the mixture and solvents in a *separation or extraction funnel* allows the desired component to be transferred to the solvent in which it is more soluble. At times, the desired component may be slightly soluble in one solvent and more soluble in the other. The amount of the desired component that is transferred to the preferential solvent can be calculated using the *distribution coefficient*, $K_D = (m_1/V_1)/(m_2/V_2)$, where m_1 represents the amount of the desired component in the volume of one of the solvents, V_1, and m_2 represents the amount of the desired component in the volume of the other solvent, V_2. Generally, extraction is used to separate organic substances from inorganic substances using water and an organic solvent.

volatility
The tendency of a substance to become a gas

extraction
A separation technique used to isolate the components of a mixture based on differences in solubility

solubility (s)
The maximum amount of solute that dissolves in a fixed quantity of a particular solvent at a specified temperature

Chromatography, originally used to separate coloured compounds in organic mixtures (from the Greek *chromos*, meaning "colour"), is now one of the most commonly used techniques to separate mixtures of chemical

FIGURE B2.4 **Laboratory and large-scale column chromatography**
(left): Courtesy of the Food and Drug Administration; (right): ©BSIP SA/Alamy Stock Photo

(Continued)

and/or biological compounds, both coloured and colourless. Chromatography is based on the principle that mixtures are separated based on the differences in the way that the components are distributed between two phases: the mobile phase and the stationary phase. The mixture is dissolved in the mobile phase, and the solution is passed through the stationary phase. Depending on the affinity that each component has for each phase, and the differences in the polarities of the components and the phases, the movement of the components varies. Even a small difference in affinities becomes magnified as the components travel the length of the stationary phase. The net result is the effective separation of the components of the mixture.

Many types of chromatography are available today. The type of chromatography used is based on the nature of the components (solid, liquid, gas, aqueous, and so on), as well as the actual components of the mixture. Some examples of chromatographic methods are gas chromatography (GC), liquid chromatography (LC; Figure B2.4), gas liquid chromatography (GLC; Figure B2.5), high-performance liquid chromatography (HPLC), thin-layer chromatography (TLC), affinity chromatography, supercritical fluid chromatography, ion-exchange chromatography, size-exclusion chromatography, reversed-phase chromatography, fast protein liquid chromatography (FPLC), and the latest and best-performing version of high-performance countercurrent chromatography (HPCCC). You will encounter chromatography again in Chapter 22.

Problem

B2.3 Name the technique(s) and briefly describe the steps for separating each of the following mixtures into pure components:

(a) Table salt and pepper
(b) Drinking water contaminated with soot
(c) Crushed ice and crushed glass
(d) Table sugar dissolved in ethanol
(e) Two pigments (chlorophyll *a* and chlorophyll *b*) from spinach leaves

chromatography
A separation technique in which a mixture is dissolved in a fluid (gas or liquid) and the components are separated through differences in adsorption to (or solubility in) a solid surface (or viscous liquid)

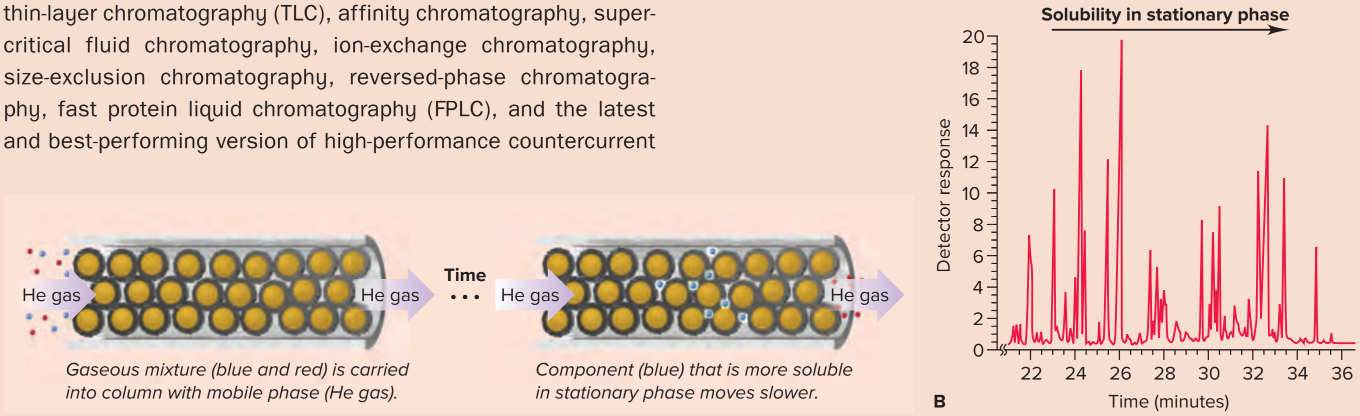


FIGURE B2.5 Principle of gas-liquid chromatography (GLC). A. The stationary phase is shown as a viscous liquid (*grey circles*) coating the solid beads (*yellow*) of an inert packing. **B.** Typical spectrum obtained using GLC.

CHAPTER REVIEW GUIDE

Learning Objectives

Relevant section (§) and/or sample problem (SP) numbers appear in parentheses.

Concepts

1. Define the characteristics of the three types of matter—element, compound, and mixture—on the macroscopic and atomic scales. (§2.1)
2. Discuss the significance of the three mass laws—mass conservation, definite composition, and multiple proportions—and identify their key characteristics. (§2.2)
3. Summarize the postulates of Dalton's atomic theory and how it explains the mass laws. (§2.3)
4. Compare and contrast the major contributions to our understanding of atomic structure of experiments by Thomson, Millikan, and Rutherford. (§2.4)
5. Characterize the structure of the atom, the main features of the subatomic particles, and the importance of isotopes. (§2.5)
6. Explain the format of the periodic table, and identify the general location and characteristics of metals, metalloids, and nonmetals. (§2.6)

7. Compare and contrast the essential features of ionic and covalent compounds and the distinction between them. (§2.7)
8. Name different types of compounds (ionic, molecular, acid, simple organic) (§2.8)
9. Categorize the types of mixtures and their properties. (§2.9)

Skills

1. Distinguish between elements, compounds, and mixtures on the atomic scale. (SP 2.1)
2. Apply the idea of the mass ratio of element to compound to find the mass of an element in a compound. (SP 2.2)
3. Visualize the mass laws. (SP 2.3)
4. Express the subatomic makeup of an isotope using atomic notation. (SP 2.4)
5. Calculate an atomic mass from isotopic composition. (SP 2.5)
6. Predict the monatomic ion formed from a main-group element. (SP 2.6)
7. Name and write the formula for an ionic compound formed from the ions in Tables 2.3, 2.4, and 2.5. (SPs 2.7, 2.8, 2.9, 2.10, 2.11, 2.12, 2.16)
8. Name and write the formula for a binary covalent compound. (SPs 2.13, 2.14, 2.16)
9. Calculate the mass of a compound (for a molecule or a formula unit). (SPs 2.15, 2.16)

Key Terms

Section 2.1
substance
element
elementary substance
molecule
compound
mixture

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

Section 2.3
atoms

Section 2.4
cathode rays
nucleus

Section 2.5
proton (p^+)
neutron (n^0)
electron (e^-)
atomic number (Z)
mass number (A)
atomic symbol
isotopes
unified atomic mass unit (u)
dalton (Da)
mass spectrometry
isotopic mass
atomic mass

Section 2.6
periodic table of the elements
periods
groups
metals
nonmetals
metalloids (semimetals)

Section 2.7
ionic compounds
covalent compounds
chemical bonds
ions
binary ionic compound
cation
anion
monatomic ion
covalent bond
polyatomic ions

Section 2.8
chemical formula
formula unit
oxoanions
hydrates
binary covalent compounds
molecular formula
structural formula

Section 2.9
heterogeneous mixture
homogeneous mixture (solution)
aqueous solutions
filtration
crystallization
distillation
volatility
extraction
solubility (s)
chromatography

Key Equations and Relationships

2.1 Finding the mass of an element in a given mass of compound:

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

2.2 Calculating the number of neutrons in an atom:

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

or

$$N = A - Z$$

2.3 Determining the mass of a formula unit of a compound:

$$\text{Mass of formula unit} = \text{sum of atomic masses}$$

Brief Solutions to Follow-Up Problems

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

2.2 Mass (t) of pitchblende

$$= 2.3\ \text{t}\ \cancel{\text{uranium}} \times \frac{84.2\ \cancel{\text{g}}\ \text{pitchblende}}{71.4\ \cancel{\text{g uranium}}} = 2.7\underline{1}\ \text{t pitchblende}$$

Mass (t) of oxygen

$$= 2.7\underline{1}\ \text{t pitchblende} - 2.3\ \text{t uranium} = 0.4\ \text{t oxygen}$$

2.3 Sample B best shows the law of multiple proportions. Two bromine-fluorine compounds appear. In one compound, there are three fluorine atoms for each bromine atom; in the other, there is one fluorine atom for each bromine atom. 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 $10.0129x + 11.0093(1 - x) = 10.81$

$$0.9964x = 0.199$$

$$x = 0.20 \text{ and } 1 - x = 0.80$$

% abundance of ^{10}B = 20.%

% abundance of ^{11}B = 80.%

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

2.7 (a) Zinc (group 12) and oxygen (group 16)
(b) Silver (group 11) and bromine (group 17)
(c) Lithium (group 1) and chlorine (group 17)
(d) Aluminum (group 13) and sulfur (group 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) Ammonium is NH_4^+, and phosphate is PO_4^{3-}. The correct formula is $(NH_4)_3PO_4$.
(b) Parentheses are needed around the polyatomic ion OH^-. The correct formula is $Al(OH)_3$.
(c) Mg^{2+} is magnesium and can have only a 2+ charge, so it does not need (II); HCO_3^- is hydrogen carbonate (or bicarbonate). The correct name is magnesium hydrogen carbonate (or magnesium bicarbonate).
(d) The *-ic* ending is not used with roman numerals; NO_3^- is nitr*ate*. The correct name is chromium(III) nitrate.
(e) Ca^{2+} is calcium and NO_2^- is nitr*ite*. The correct name is calcium nitrite.

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) The *-ous* suffix is not used. The correct name is disulfur dichloride.
(b) The name indicates one nitrogen. The correct formula is NO.
(c) Br is in a higher period in group 17, so it is named first. The correct name is bromine trichloride.

2.15 (a) H_2O_2, 34.02 u; (b) CsCl, 168.35 u; (c) H_2SO_4, 98.09 u; (d) K_2SO_4, 174.27 u

2.16 (a) Na_2O; this is an ionic compound, so the name is sodium oxide.

Formula mass = 2 × atomic mass of Na + 1 × atomic mass of O
$= 2 \times 22.99\ u + 16.00\ u = 61.98\ u$

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

Mass of NO_2 = 1 × atomic mass of N + 2 × atomic mass of O
$= 14.01\ u + 2 \times 16.00\ u = 46.01\ u$

PROBLEMS

Problems with **red** numbers are answered in Appendix G and worked in detail in the Student Solutions Manual. Problem sections match those in this book and provide the numbers of relevant sample problems. Most offer Concept Review Questions, Skill-Building Exercises (grouped in pairs covering the same concept), and Problems in Context. The Comprehensive Problems are based on material from any section or previous chapter.

Elements, Compounds, and Mixtures: An Atomic Overview

(Sample Problem 2.1)

Concept Review Questions

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 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, a compound, or a mixture, and explain your answers.

2.5 Explain the following statement: The smallest particles unique to an element may be atoms or molecules.

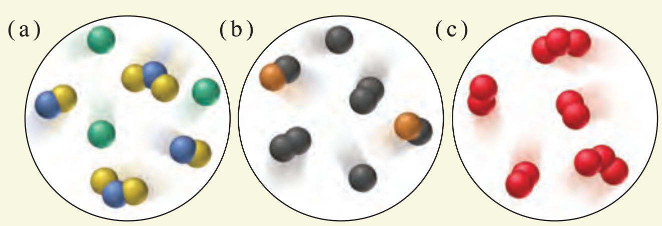
2.6 Explain the following statement: The smallest particles unique to a compound cannot be atoms.

2.7 Can the relative amounts of the components of a mixture vary? Can the relative amounts of the components of a compound vary? Explain.

Problems in Context

2.8 The tap water found in many areas of Canada leaves white deposits when it evaporates. Is this tap water a mixture or a compound? Explain.

2.9 The following diagrams represent mixtures. Describe each diagram in terms of the number(s) of elements and/or compounds present.



2.10 Samples of illicit "street" drugs often contain an inactive component, such as ascorbic acid (vitamin C). After obtaining a sample of cocaine, government chemists calculate the mass of vitamin C per gram of sample and use it to track the cocaine's distribution. For example, if different samples of cocaine, obtained on the streets of New York, Vancouver, and Paris, all contain 0.6384 g of vitamin C per gram of sample, they very likely came from a common source. Do these street samples consist of a compound, an element, or a mixture? Explain.

The Observations That Led to an Atomic View of Matter

(Sample Problem 2.2)

Concept Review Questions

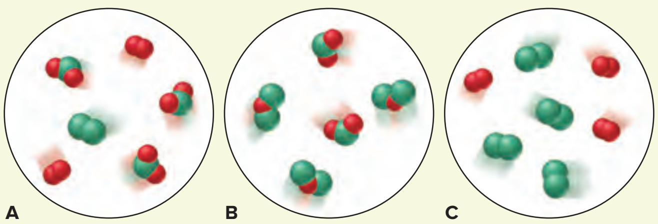
2.11 Why was it necessary for separation techniques and methods of chemical analysis to be developed before the laws of definite composition and multiple proportions could be formulated?

2.12 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?

2.13 In our modern view of matter and energy, is the law of mass conservation still relevant to chemical reactions? Explain.

2.14 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 a sample from Poland.
(b) A glass bulb 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.

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



A B C

2.16 (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.

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

Skill-Building Exercises (grouped in similar pairs)

2.18 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 colourless 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 colourless gas.

2.19 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; 0.80 g of copper remains.

2.20 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 (a) the mass of fluorine in the sample; (b) the mass fractions of calcium and fluorine in fluorite; (c) the mass percents of calcium and fluorine in fluorite.

2.21 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 (a) the mass of sulfur in the sample; (b) the mass fractions of lead and sulfur in galena; (c) the mass percents of lead and sulfur in galena.

2.22 Magnesium oxide (MgO) forms when the metal burns in air.
(a) If 1.25 g of MgO contains 0.754 g of Mg, what is the mass ratio of magnesium to oxide?
(b) What mass of Mg is in 534 g of MgO?

2.23 Zinc sulfide (ZnS) occurs in the zinc blende crystal structure.
(a) If 2.54 g of ZnS contains 1.70 g of Zn, what is the mass ratio of zinc to sulfide?
(b) What mass, in kilograms, of Zn is in 3.82 kg of ZnS?

2.24 A compound of copper and sulfur contains 88.39 g of metal and 44.61 g of nonmetal. What mass of copper is in 5264 kg of the compound? What mass of sulfur?

2.25 A compound of iodine and cesium contains 63.94 g of metal and 61.06 g of nonmetal. What mass of cesium is in 38.77 g of the compound? What mass of iodine?

2.26 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

2.27 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

Problems in Context

2.28 Dolomite is a carbonate of magnesium and calcium. Analysis shows that 7.81 g of dolomite contains 1.70 g of calcium. Calculate the mass percent of calcium in dolomite. On the basis of the mass percent of calcium, and neglecting all other factors, which is the richer source of calcium, dolomite or fluorite (see Problem 2.20)?

2.29 The mass percent of sulfur in a sample of coal is a key factor in the environmental impact of the coal. 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)

Concept Review Questions

2.30 Which of Dalton's postulates about atoms are inconsistent with later observations? Do these inconsistencies mean that Dalton was wrong? Is Dalton's model still useful? Explain.

2.31 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

Concept Review Questions

2.32 Thomson was able to determine the mass/charge ratio of the electron, but not its mass. How did Millikan's experiment allow the electron's mass to be determined?

2.33 The following charges on individual oil droplets were obtained during an experiment similar to Millikan's: -3.204×10^{-19} C; -4.806×10^{-19} C; -8.010×10^{-19} C; -1.442×10^{-18} C. Determine a charge for the electron (in C, coulomb), and explain your .

2.34 Describe Thomson's model of the atom. How might it account for the production of cathode rays?

2.35 When Rutherford's co-workers 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)

Concept Review Questions

2.36 Define *atomic number* and *mass number*. Which can vary without changing the identity of the element?

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

2.38 Even though several elements have only one naturally occurring isotope and all atomic nuclei have whole numbers of protons and neutrons, no atomic mass is a whole number. Use data from Table 2.2 to explain this fact.

Skill-Building Exercises (grouped in similar pairs)

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

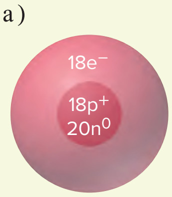
2.40 Chlorine has two naturally occurring isotopes: ^{35}Cl and ^{37}Cl. What is the mass number of each isotope? How many protons, neutrons, and electrons are present in each?

2.41 Do both atoms in each pair 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
In which pair(s) do the atoms have the same Z value? N value? A value?

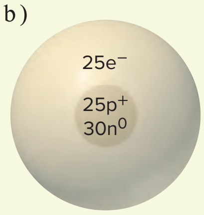
2.42 Do both atoms in each pair have the same number of protons? Neutrons? Electrons?
(a) $^{3}_{1}$H and $^{3}_{2}$He (b) $^{14}_{6}$C and $^{15}_{7}$N (c) $^{19}_{9}$F and $^{18}_{9}$F
In which pair(s) do the atoms have the same Z value? N value? A value?

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

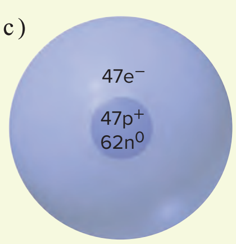
(a)



(b)

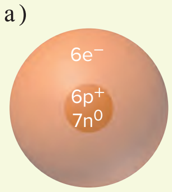


(c)

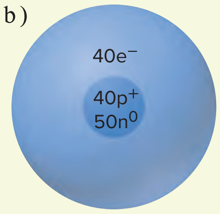


2.44 Write the $^{A}_{Z}$X notation for each atomic depiction:

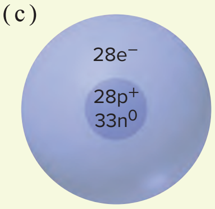
(a)



(b)



(c)



2.45 Draw atomic depictions similar to those in Problem 2.43 for these atoms: (a) $^{49}_{22}$Ti; (b) $^{79}_{34}$Se; (c) $^{11}_{5}$B.

2.46 Draw atomic depictions similar to those in Problem 2.43 for these atoms: (a) $^{207}_{22}$Pb; (b) $^{9}_{4}$Be; (c) $^{75}_{33}$As.

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

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

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

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

Elements: A First Look at the Periodic Table

Concept Review Questions

2.51 How can iodine ($Z = 53$) have a higher atomic number and yet a lower atomic mass than tellurium ($Z = 52$)?

2.52 Correct each statement:
(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.

2.53 What class of elements lies along the imaginary staircase line in the periodic table? How do the properties of these elements compare with the properties of metals and nonmetals?

2.54 What are some characteristic properties of elements to the left of the elements along the imaginary staircase? To the right of the elements along the imaginary staircase?

2.55 All of the elements in groups 1 and 17 are quite reactive. What is a major difference between them?

Skill-Building Exercises (grouped in similar pairs)

2.56 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) $Z = 42$

2.57 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) $Z = 13$

2.58 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 14 are ______ and ______.
(c) Group 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 ______.

2.59 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 symbol and atomic number of the first group 16 element displaying a metallic nature are ______ and ______.
(d) The symbol and number of protons of the period 4 alkali metal atom are ______ and ______.

Compounds: An Introduction to Bonding

(Sample Problem 2.6)

Concept Review Questions

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

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

2.62 How can ionic compounds be neutral if they consist of positive and negative ions?

2.63 Given that the ions in LiF and the ions in MgO are similar sizes, which compound has stronger ionic bonding? Use Coulomb's law in your explanation.

2.64 Are molecules present in a sample of BaF_2? Explain.

2.65 Are ions present in a sample of P_4O_6? Explain.

2.66 The monatomic ions of groups 1 and 17 are all singly charged. In what major way do they differ? Why?

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

2.68 Describe the formation of solid potassium sulfide (K_2S) from large numbers of potassium and sulfur atoms.

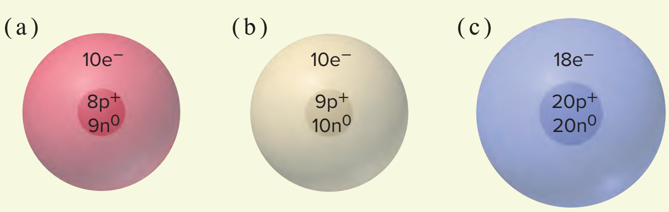
2.69 Does potassium nitrate (KNO_3) incorporate ionic bonding, covalent bonding, or both? Explain.

Skill-Building Exercises (grouped in similar pairs)

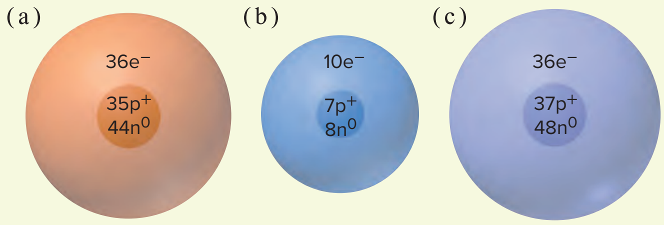
2.70 What monatomic ions do potassium ($Z = 19$) and iodine ($Z = 53$) form?

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

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



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



2.74 An ionic compound forms when lithium ($Z = 3$) reacts with oxygen ($Z = 8$). If a sample of the compound contains 8.4×10^{21} lithium ions, how many oxide ions does it contain?

2.75 An ionic compound forms when calcium ($Z = 20$) reacts with iodine ($Z = 53$). If a sample of the compound contains 7.4×10^{21} calcium ions, how many iodide ions does it contain?

2.76 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?

2.77 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?

Formula, Name, and Mass of a Compound

(Sample Problems 2.7 to 2.16)

Concept Review Questions

2.78 What information about the relative numbers of ions and the mass percents of the elements is in the formula MgF_2?

2.79 How is a structural formula similar to a molecular formula? How is it different?

2.80 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 that contains 10 billion hydrogen peroxide (H_2O_2) molecules? In what way is it different?

2.81 For what type(s) of compound(s) do we use roman numerals in the names?

2.82 For what type(s) of compound(s) do we use Greek numerical prefixes in the names?

2.83 For what type of compound are we unable to write a molecular formula?

Skill-Building Exercises (grouped in similar pairs)

2.84 Give the name and formula for the compound formed from each pair of elements:
(a) Sodium and nitrogen (b) Oxygen and strontium
(c) Aluminum and chlorine

2.85 Give the name and formula for the compound formed from each pair of elements:
(a) Cesium and bromine (b) Sulfur and barium
(c) Calcium and fluorine

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

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

2.88 Give the systematic name for the formula or the formula for the name:
(a) Tin(IV) chloride (b) $FeBr_3$
(c) Cuprous bromide (d) Mn_2O_3

2.89 Give the systematic name for the formula or the formula for the name:
(a) Na_2HPO_4 (b) Potassium carbonate dihydrate
(c) $NaNO_2$ (d) Ammonium perchlorate

2.90 Give the systematic name for the formula or the formula for the name:
(a) CoO (b) Mercury(I) chloride
(c) $Pb(C_2H_3O_2)_2{\cdot}3H_2O$ (d) Chromic oxide

2.91 Give the systematic name for the formula or the formula for the name:
(a) $Sn(SO_3)_2$ (b) Potassium dichromate
(c) $FeCO_3$ (d) Copper(II) nitrate

2.92 Correct each incorrect formula:
(a) Barium oxide is BaO_2. (b) Iron(II) nitrate is $Fe(NO_3)_3$.
(c) Magnesium sulfide is $MnSO_3$.

2.93 Correct each name:
(a) CuI is cobalt (II) iodide.
(b) $Fe(HSO_4)_3$ is iron(II) sulfate.
(c) $MgCr_2O_7$ is magnesium dichromium heptaoxide.

2.94 Give the name and formula for the acid derived from each anion:
(a) Hydrogen sulfate (b) IO_3^- (c) Cyanide (d) HS^-

2.95 Give the name and formula for the acid derived from each anion:
(a) Perchlorate (b) NO_3^- (c) Bromite (d) F^-

2.96 Many chemical names are similar at first glance. Give the formula for the different species in each set:
(a) Ammonium ion and ammonia
(b) Magnesium sulfide, magnesium sulfite, and magnesium sulfate
(c) Hydrochloric acid, chloric acid, and chlorous acid
(d) Cuprous bromide and cupric bromide

2.97 Give the formula for the different compounds in each set:
(a) Lead(II) oxide and lead(IV) oxide
(b) Lithium nitride, lithium nitrite, and lithium nitrate
(c) Strontium hydride and strontium hydroxide
(d) Magnesium oxide and manganese(II) oxide

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

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

2.100 Correct the name to match the formula for each compound:
(a) Calcium(II) dichloride, $CaCl_2$ (b) Copper(II) oxide, Cu_2O
(c) Stannous tetrafluoride, SnF_4 (d) Hydrogen chloride acid, HCl

2.101 Correct the formula to match the name of each compound:
(a) Iron(III) oxide, Fe_3O_4 (b) Chloric acid, HCl
(c) Mercuric oxide, Hg_2O (d) Dichlorine heptaoxide, Cl_2O_6

2.102 Give the number of atoms of the specified element in a formula unit of each compound, and calculate the molecular or 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, $Cu_3(OH)_2(CO_3)_2$

2.103 Give the number of atoms of the specified element in a formula unit of each compound, and calculate the molecular or 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, $Pb_4SO_4(CO_3)_2(OH)_2$

2.104 Write the formula for each compound, and determine its molecular or formula mass:
(a) Ammonium sulfate (b) Sodium dihydrogen phosphate
(c) Potassium bicarbonate

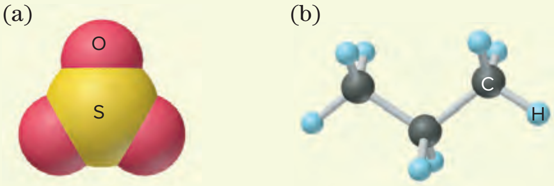
2.105 Write the formula for each compound, and determine its molecular or formula mass:
(a) Sodium dichromate (b) Ammonium perchlorate
(c) Magnesium nitrite trihydrate

2.106 Calculate the molecular or formula mass of each compound:
(a) Dinitrogen pentoxide (b) Lead(II) nitrate
(c) Calcium peroxide

2.107 Calculate the molecular or formula mass of each compound:
(a) Iron(II) acetate tetrahydrate (b) Sulfur tetrachloride
(c) Potassium permanganate

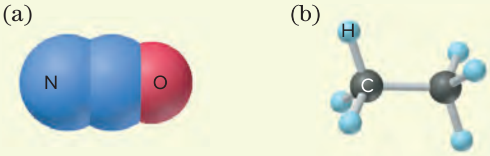
2.108 Give the formula, name, and molecular mass of each molecule:

(a) (b)



2.109 Give the formula, name, and molecular mass of each molecule:

(a) (b)

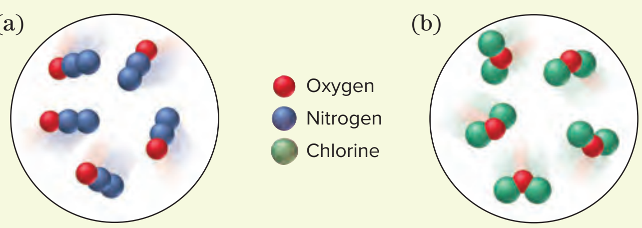


Problems in Context

2.110 Before the use of systematic names, many compounds had common names. Give the systematic name for each compound:
(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

2.111 Each circle contains a representation of a binary compound. Determine the name, formula, and molecular (formula) mass of the compound:

(a) (b)



Mixtures: Classification and Separation

Concept Review Questions

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

2.113 What is the difference between a homogeneous mixture and a heterogeneous mixture?

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

Skill-Building Exercises (grouped in similar pairs)

2.115 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

2.116 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

Problems in Context

2.117 Which separation method is operating in each procedure? (a) Pouring a mixture of cooked pasta and boiling water into a colander; (b) Removing coloured impurities from raw sugar to make refined sugar.

2.118 A quality-control laboratory analyzes a product mixture using gas-liquid chromatography. The separation of components is more than adequate, but the process takes too long. Suggest two ways, other than changing the stationary phase, to shorten the analysis time.

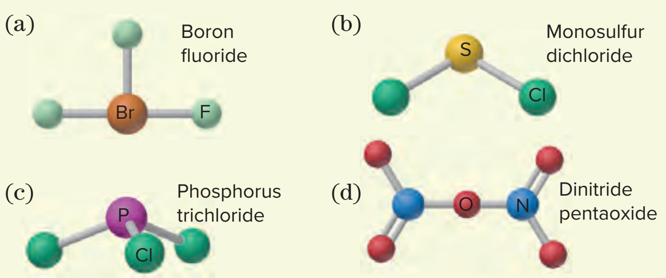
Comprehensive Problems

2.119 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 3.1×10^{-11} m; the radius of its nucleus is 2.5×10^{-15} m. What fraction of its spherical atomic volume is occupied by its nucleus? (V of sphere $= \frac{4}{3}\pi r^3$.)
(b) The mass of a helium-4 atom is $6.646\,48 \times 10^{-24}$ g, and each of its two electrons has a mass of $9.109\,39 \times 10^{-28}$ g. What fraction of this atom's mass is contributed by its nucleus?

2.120 From the following ions (with their radii in pm), choose the pair that forms the strongest ionic bond and the pair that forms the weakest ionic bond:

Ion:	Mg^{2+}	K^+	Rb^+	Ba^{2+}	Cl^-	O^{2-}	I^-
Radius:	72	138	152	135	181	140	220

2.121 Give the molecular mass of each structure depicted below, and provide a correct name for any that are named incorrectly:



2.122 Polyatomic ions are named by patterns that apply to elements in a given group. Using the periodic table and Table 2.5, give the name of each ion:
(a) SeO_4^{2-} (b) AsO_4^{3-} (c) BrO_2^- (d) $HSeO_4^-$ (e) TeO_3^{2-}

2.123 Ammonium dihydrogen phosphate, formed from the reaction of phosphoric acid with ammonia, is used as a crop fertilizer as well as a component of some fire extinguishers.
(a) What are the mass percents of N and P in the compound?
(b) What mass of ammonia is contained in 100. g of the compound?

2.124 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%; and (III) 25.94%. For each oxide, determine (a) the simplest whole-number ratio of N to O and (b) the mass of oxygen, in grams per 1.00 g of nitrogen.

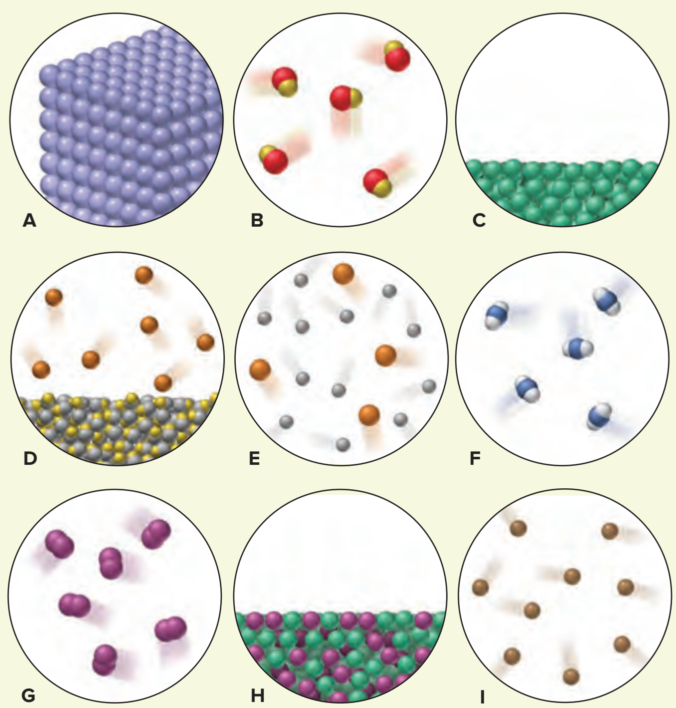
2.125 The number of atoms in 1 dm^3 of aluminum is nearly the same as the number of atoms in 1 dm^3 of lead, but the densities of these metals are very different (see Table 1.4). Explain.

2.126 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 $\rightarrow$ 100.00 g NaCl
Case 2: 39.34 g Na + 70.00 g Cl_2 $\rightarrow$ 100.00 g NaCl + 9.34 g Cl_2
Case 3: 50.00 g Na + 50.00 g Cl_2 $\rightarrow$ 82.43 g NaCl + 17.57 g Na

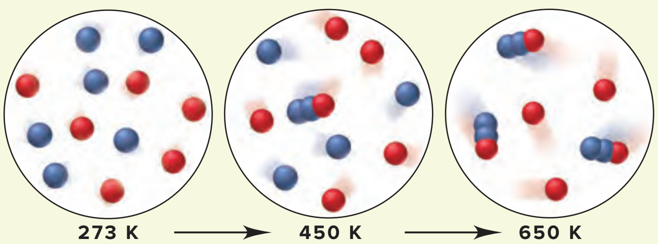
Explain these results in terms of the laws of mass conservation and definite composition.

2.127 Diagrams A to I depict various types of matter on the atomic scale. Choose the correct diagram(s) for each description:
(a) A mixture that fills its container
(b) A substance that cannot be broken down into simpler substances
(c) An element that has 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 that consists of diatomic particles
(g) A gas that can be broken down into simpler substances
(h) A substance that has 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

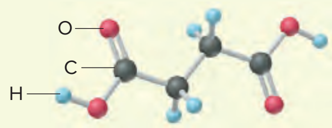


2.128 The seven most abundant ions in seawater make up more than 99% by mass of the dissolved compounds. They have the following abundances, in units of mg ion/kg seawater:
chloride 18 980; sodium 10 560; sulfate 2650; magnesium 1270; calcium 400; potassium 380; hydrogen carbonate 140.
(a) What is the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

2.129 The diagram below represents a mixture of two monatomic gases undergoing a reaction when heated. Which mass law(s) is (are) illustrated by this reaction?



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



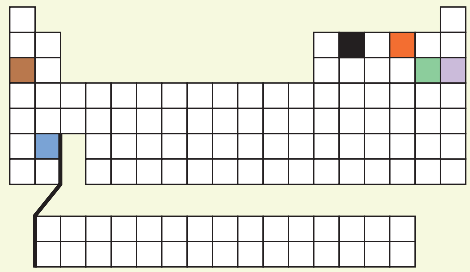
2.131 Fluoride ion is poisonous in relatively low amounts: 0.2 g of F^- per 70 kg of body mass can cause death. Nevertheless, in order to prevent tooth decay, F^- ions are added to drinking water at a concentration of 1 mg of F^- ions per litre of water. What volume (L) of fluoridated drinking water would a 70 kg person have to consume in one day to reach this toxic level? What mass (kg) of sodium fluoride would be needed to treat an 8.50×10^7 L reservoir?

2.132 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 u and the other with mass 122.904 u.
(a) Write the notation for each isotope.
(b) Use the atomic mass of antimony from the periodic table to calculate the natural abundance of each isotope.

2.133 Dinitrogen monoxide (N_2O; nitrous oxide) is a greenhouse gas that enters the atmosphere principally from the breakdown of natural fertilizer. 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?

2.134 Nuclei differ in their stability, and some are so unstable that they undergo radioactive decay. The ratio of the number of neutrons to the number of protons (N/Z) in a nucleus correlates with its stability.
(a) Calculate the N/Z ratio for (i) ^{144}Sm; (ii) ^{56}Fe; (iii) ^{20}Ne; and (iv) ^{107}Ag.
(b) The radioactive isotope ^{238}U decays in a series of nuclear reactions that includes another uranium isotope, ^{234}U, and three lead isotopes: ^{214}Pb, ^{210}Pb, and ^{206}Pb. How many neutrons, protons, and electrons are in each of these five isotopes?

2.135 Use the box colour(s) in the periodic table below to identify the element(s) described by each of the following:



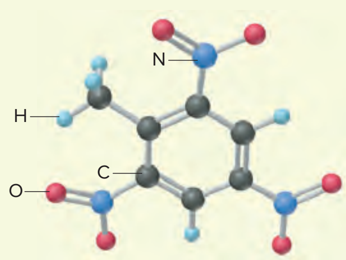
(a) Four elements that are nonmetals
(b) Two elements that are metals
(c) Three elements that are gases at room temperature
(d) Three elements that are solids at room temperature
(e) One pair of elements likely to form a covalent compound
(f) Another pair of elements likely to form a covalent compound
(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

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

2.137 Boron trifluoride is used as a catalyst in the synthesis of organic compounds. When this compound is analyzed by mass spectrometry (see the first Tools of the Laboratory in this chapter), several different 1+ ions form, including ions representing the whole molecule as well as molecular fragments formed by the loss of one, two, and three F atoms. Given that boron has two naturally occurring isotopes, ^{10}B and ^{11}B, and fluorine has one, ^{19}F, calculate the masses of all of the possible 1+ ions.

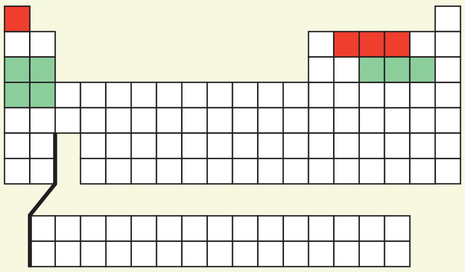
2.138 Nitrogen monoxide (NO) is a bioactive molecule in blood. Low NO concentrations cause respiratory distress and the formation of blood clots. Doctors prescribe nitroglycerin, $C_3H_5N_3O_9$, and isoamyl nitrate, $(CH_3)_2CHCH_2CH_2ONO_2$, to increase NO. If each compound releases one molecule of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

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

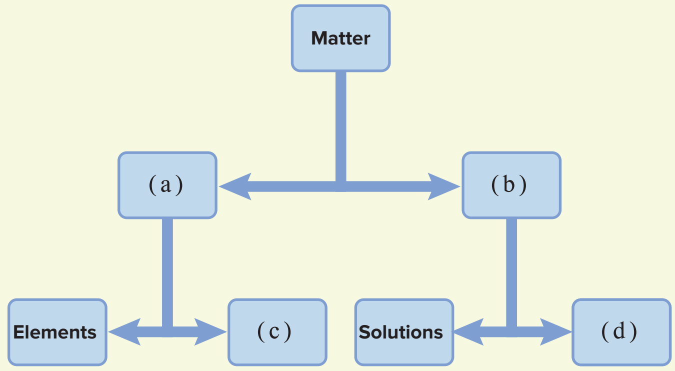


2.140 The anticancer drug Platinol (cisplatin), $Pt(NH_3)_2Cl_2$, reacts with a cancer cell's DNA and interferes with its growth. (a) What is the mass percent of platinum (Pt) in Platinol? (b) If Pt costs \$32/g, what mass of Platinol can be made for \$1.00 million? (Assume that the cost of Pt determines the cost of the drug.)

2.141 In the periodic table below, give the name, symbol, atomic number, atomic mass, period number, and group number of (a) the *building-block elements (red)*, which occur in nearly every biological molecule, and (b) the *macronutrients (green)*, which are either essential ions in cell fluids or part of many biomolecules.



2.142 The following block diagram classifies the components of matter on the macroscopic scale. Identify blocks (a) to (d).



2.143 Which of the following steps in an overall process involve(s) a physical change? Which involve(s) a chemical change?