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AT THE DENTIST'S OFFICE

Having moved to a new city earlier in the year, this was the girl's first visit to a new dentist. After having her teeth cleaned by the hygienist and checked by the dentist, she was led back to her mother who was sitting in the waiting room. The hygienist described the girl's successful checkup and then handed the mother a piece of paper saying, "I don't know if you are aware of it, but the water here is not fluoridated. You might think about giving your daughter fluoride tablets. Here is a prescription if you are interested."

3 Compounds

ABOUT THIS CHAPTER

The helium in a balloon, the nitrogen and oxygen gas in air, and the mercury in a thermometer are elements (contain only one type of atom). These are among the relatively few examples of pure elements that you are likely to encounter. The chances are much greater that you will come into contact with compounds (matter constructed of two or more chemically combined elements). In this chapter we will see the different ways that atoms interact with one another to form compounds, including those present in your teeth.

CHAPTER 3 OBJECTIVES

After completing this chapter, you should be able to:

1. **Define the term ion and describe the naming of monoatomic cations and anions.**
2. **Describe the naming of polyatomic cations and anions.**
3. **Explain how the electron dot structure of a representative element atom can be used to predict the charge on the monoatomic ion that it forms.**
4. **Name and write the formulas of ionic compounds and binary molecules.**
5. **Predict the number of covalent bonds that a nonmetal atom will form and draw electron dot and Lewis structures of molecules.**
6. **Define the terms formula weight and molecular weight. Use the molar mass of a compound to carry out conversions involving moles and mass.**

3.1 IONS

As we saw earlier, atomic notation allows us to determine the number of protons and neutrons in an atom's nucleus. We also saw that the number of protons and electrons are equal in a neutral atom. For example, an atom of $^{1}_{1}H$ contains 1 proton and 1 electron, an atom of $^{3}_{1}H$ has 1 proton, 2 neutrons, and 1 electron, and an atom of $^{12}_{6}C$ has 6 protons, 6 neutrons, and 6 electrons.

When an atom or a group of atoms gains or loses electrons, an **ion** results. Ions have *an unequal number of protons and electrons*. We will begin our look at ions with a few examples. Why these particular ions form will be explained later in this chapter.

If Li loses an electron, an ion with a 1+ charge, Li^{+}, is created (Figure 3.1*a*). This **monoatomic ion** (formed from a single atom) has a 1+ charge because it contains three positively charged protons and only two negatively charged electrons. (In atomic notation the *charge on a monoatomic ion is shown as a superscript to the right of the atomic symbol.*)

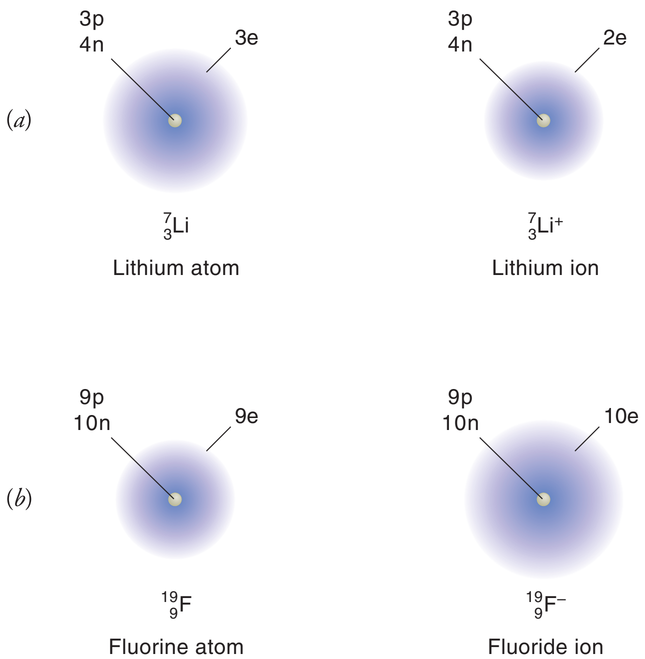
	Protons	Electrons	Charge	Name
Li	3	3	0	Lithium
Li^{+}	3	2	1+	Lithium ion

■ Ions have an unequal number of protons and electrons.

Ions that carry a *positive charge*, like Li^{+}, are called **cations**. Monoatomic cations formed from representative elements (groups 1A–8A) are given the same name as the original element, so Li^{+} is a lithium ion.

Removing two electrons from a neutral calcium atom produces a calcium ion with a 2+ charge, because it has two more protons than electrons.

	Protons	Electrons	Charge	Name
Ca	20	20	0	Calcium
Ca^{2+}	20	18	2+	Calcium ion



■ **FIGURE 3.1**

Ions (*a*) $^{7}_{3}Li$ contains 3 protons, 4 neutrons, and 3 electrons and $^{7}_{3}Li^{+}$ has 3 protons, 4 neutrons, and 2 electrons. Monoatomic cations are smaller than the atoms from which they are formed. (*b*) $^{19}_{9}F$ contains 9 protons, 10 neutrons, and 9 electrons and $^{19}_{9}F^{-}$ has 9 protons, 10 neutrons, and 10 electrons. Monoatomic anions are larger than the atoms from which they are formed.

Some transition metals (groups 1B–8B) are able to form several different cations. Iron ions, for example, can have a 2+ or a 3+ charge.

	Protons	Electrons	Charge	Name
Fe	26	26	0	Iron
Fe^{2+}	26	24	2+	Iron(II) ion
Fe^{3+}	26	23	3+	Iron(III) ion

When naming transition metal ions, a Roman numeral (in parentheses) is used to indicate charge—Fe^{2+} is iron(II) ion, Fe^{3+} is iron(III) ion, Cu^{+} is copper(I) ion, and Cu^{2+} is copper(II) ion.

SAMPLE PROBLEM 3.1

Naming monoatomic cations

Name each ion.

a. Mg^{2+} **b.** Co^{2+} **c.** Mn^{2+}

STRATEGY

There may be more to this problem than you expect. To name these monoatomic cations you must check to see if they are transition metal ions. If so, the charge on the ion should be part of the name.

SOLUTION

a. magnesium ion **b.** cobalt(II) ion **c.** manganese(II) ion

PRACTICE PROBLEM 3.1

Name each ion.

a. Sc^{3+} **b.** Ti^{2+} **c.** Sr^{2+}

In an alternative naming system used for transition metal ions, the suffix *ous* is used for the ion with the smaller charge and the suffix *ic* for that with the greater charge (Table 3.1). Using this system, Fe^{2+} is ferrous ion, Fe^{3+} is ferric ion, Cu^{+} is cuprous ion, and Cu^{2+} is cupric ion. This naming system is not as easy to use as the Roman numeral–based system, because you cannot determine the charge on the ion directly from its name. Another challenge to using this system is that, for some ions, it makes use of the older Latin names for the elements, such as *ferrum* for iron and *cuprum* for copper.

TABLE 3.1 SOME TRANSITION METAL IONS

Atom	Ion	Name	Alternative Name
Chromium	Cr^{2+}	Chromium(II) ion	Chromous ion
	Cr^{3+}	Chromium(III) ion	Chromic ion
Copper	Cu^{+}	Copper(I) ion	Cuprous ion
	Cu^{2+}	Copper(II) ion	Cupric ion
Iron	Fe^{2+}	Iron(II) ion	Ferrous ion
	Fe^{3+}	Iron(III) ion	Ferric ion
Tin	Sn^{2+}	Tin(II) ion	Stannous ion
	Sn^{4+}	Tin(IV) ion	Stannic ion

- Cations have a positive charge and anions have a negative charge.

Some atoms gain electrons to produce *negatively charged ions* called **anions**. When F gains an electron, an ion with a 1− charge, F^-, is formed. Monoatomic anions are named by changing the ending on the name of the element name to *ide*, so the anion formed from fluorine is called fluoride ion (Figure 3.1*b*).

	Protons	Electrons	Charge	Name
F	9	9	0	Fluorine
F^-	9	10	1−	Fluoride ion

When an oxygen atom gains two electrons, it is transformed into an oxide ion.

	Protons	Electrons	Charge	Name
O	8	8	0	Oxygen
O^{2-}	8	10	2−	Oxide ion

- Polyatomic ions contain two or more atoms.

Many ions are built from *two or more atoms*. These **polyatomic ions** include ammonium ion (NH_4^+) and carbonate ion (CO_3^{2-}). As is the case for monoatomic ions, the charge on polyatomic ions is due to an imbalance in the total number of protons and electrons that are present. NH_4^+ contains 1 nitrogen atom, 4 hydrogen atoms, and a total of 10 electrons.

	Protons	Electrons	Charge
NH_4^+	$7 \text{ from N} + \underbrace{4 \times 1 \text{ from each H}}_{4}$		
	$7 + 4 = 11$	10	1+

CO_3^{2-} is a combination of 1 carbon atom, 3 oxygen atoms, and a total of 32 electrons.

	Protons	Electrons	Charge
CO_3^{2-}	$6 \text{ from C} + \underbrace{3 \times 8 \text{ from each O}}_{24}$		
	$6 + 24 = 30$	32	2−

In Chapter 4 we will take a look at how the atoms in polyatomic ions are connected to one another.

Table 3.2 lists a number of the more common polyatomic ions. Some differ only in the number of hydrogen atoms that they contain, as is the case for PO_4^{3-}, HPO_4^{2-}, and $H_2PO_4^-$. The names of these ions reflect the number of hydrogen atoms that are present. Phosphate

TABLE 3.2 COMMON POLYATOMIC IONS

	Formula	Name	Formula	Name
Cations				
	H_3O^+	Hydronium ion	NH_4^+	Ammonium ion
Anions				
	OH^-	Hydroxide ion	HSO_4^-	Hydrogen sulfate (bisulfate) ion
	CO_3^{2-}	Carbonate ion	PO_4^{3-}	Phosphate ion
	HCO_3^-	Hydrogen carbonate (bicarbonate) ion	HPO_4^{2-}	Hydrogen phosphate ion
	NO_2^-	Nitrite ion	$H_2PO_4^-$	Dihydrogen phosphate ion
	NO_3^-	Nitrate ion	$Cr_2O_7^{2-}$	Dichromate ion
	SO_3^{2-}	Sulfite ion	$CH_3CO_2^-$	Acetate ion
	SO_4^{2-}	Sulfate ion	CN^-	Cyanide ion

ion (PO_4^{3-}) has no hydrogen atoms, hydrogen phosphate ion (HPO_4^{2-}) has one hydrogen atom, and dihydrogen phosphate ion ($H_2PO_4^{-}$) has two. When an element can form polyatomic ions by combining with oxygen in two different ways, the suffixes "ate" and "ite" are used to indicate the relative number of oxygen atoms: Nitrate ion (NO_3^{-}) has one more oxygen atom than nitrite ion (NO_2^{-}), and sulfate ion (SO_4^{2-}) has one more oxygen atom than sulfite ion (SO_3^{2-}).

Most of the polyatomic ions in Table 3.2 are found in significant amounts in all living things and many are essential to survival. Hydrogen carbonate ion (HCO_3^{-}), for example, is involved in the transport of carbon dioxide from your tissues to your lungs.

3.2 | THE OCTET RULE

To understand why particular ions form, it helps to consider the arrangement of electrons about atomic nuclei. One approach is to list the distribution of electrons into the various energy levels. This was done for the elements hydrogen through calcium in Table 2.6.

Alternatively, electron dot structures can be used. In these drawings, valence electrons (electrons in an atom's highest occupied energy level) are represented by dots. For the representative elements, the number of valence electrons corresponds to the group number (Figure 3.2).

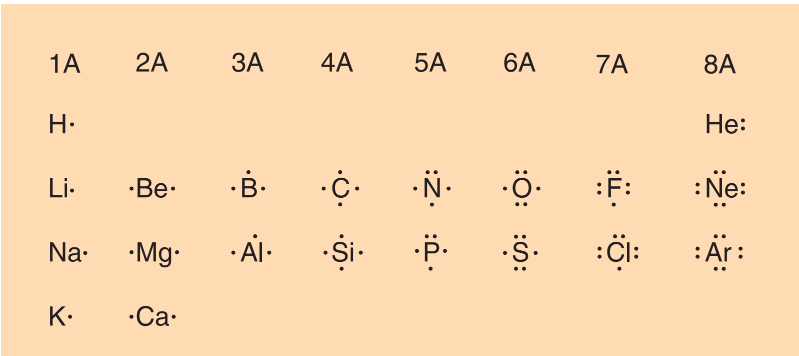
The elements helium (He), neon (Ne), argon (Ar), krypton (Kr), xenon (Xe), and radon (Rn) belong to the inert gas family, which gets its name from the fact that atoms of these elements are resistant to change and, with few exceptions, do not lose or gain electrons. This stability or resistance to change is related to the number of valence electrons held by atoms of these elements. For helium and neon, the stability may have to do with the fact that these atoms have full valence shells—a helium atom's $n = 1$ energy level is filled with two electrons and a neon atom's $n = 2$ energy level is filled with eight electrons. Atoms of the other inert gases do not have full valence shells but, like the neon atom, carry eight valence electrons (Figure 3.2).

According to the **octet rule**, *atoms gain, lose, or share valence electrons in order to end up with eight valence electrons*. The effect that this has is of providing the same stable electron arrangement as found in inert gases.

■ An octet (8 valence electrons) is a stable electron arrangement.

The octet rule helps us understand ion formation. The electron dot structure of a neutral fluorine atom, for example, has seven valence electrons. Adding one more electron gives a fluoride ion with an octet and a 1− charge. The other halogen atoms behave the same way—chlorine, bromine, and iodine each have seven valence electrons and require just one more to reach an octet (Table 3.3).

Oxygen and the other group 6A nonmetal atoms (S and Se) always form monoatomic anions with a charge of 2−. Each has six valence electrons when neutral and must gain exactly two more to achieve an octet. The group 5A nonmetals, nitrogen and phosphorus, have five valence electrons and require three more for an octet, giving ions with a charge



■ **FIGURE 3.2**
Valence electrons
Representative elements in the same group have the same number of valence electrons. In the electron dot structures used here, valence electrons are shown as dots.

TABLE 3.3 ATOMS AND IONS OF REPRESENTATIVE ELEMENTS.[a]

Electron dot structure	Group	Number of electrons per energy level $n = 1$	$n = 2$	$n = 3$	$n = 4$
Nonmetals					
:F̤̈: Fluorine atom	7A	2	**7**		
:F̤̈:$^{-}$ Fluoride ion		2	**8**		
·Ö· Oxygen atom	6A	2	**6**		
:Ö̤:$^{2-}$ Oxide ion		2	**8**		
·Ṗ̤· Phosphorus atom	5A	2	8	**5**	
:P̤̈:$^{3-}$ Phosphide ion		2	8	**8**	
Metals					
Na· Sodium atom	1A	2	8	**1**	
Na^{+} Sodium ion		2	**8**	0	
·Ca· Calcium atom	2A	2	8	8	**2**
Ca^{2+} Calcium ion		2	8	**8**	0
·Al̇· Aluminum atom	3A	2	8	**3**	
Al^{3+} Aluminum ion		2	**8**	0	

[a]The number of valence electrons is shown in bold.

of 3−. Carbon, with four valence electrons, might be expected to accept four more to make an octet. Only rarely, however, will atoms gain or lose more than three electrons, since this places too great a charge on the ion that would be produced. Carbon atoms reach an octet by other means, as we will see in Section 3.4.

■ Nonmetals gain electrons and metals lose electrons.

As the two previous paragraphs showed, nonmetal atoms gain electrons to reach an octet. In contrast, metal atoms lose electrons. The 11 electrons of a sodium atom are distributed among three energy levels: 2 electrons in the $n = 1$ energy level, 8 electrons in the $n = 2$ energy level, and 1 electron in the $n = 3$ energy level. When a sodium atom loses the valence electron held in its $n = 3$ energy level, the resulting cation has 8 electrons (an octet) in its new valence shell ($n = 2$) (Table 3.3).

The valance electron shown in the electron dot structure of a sodium atom represents the lone electron in energy level 3. When this electron is lost from a sodium ion, energy level 3 is empty. The electron dot structure of Na^+ does not show any electron dots to reinforce the idea that the original valence shell has been emptied.

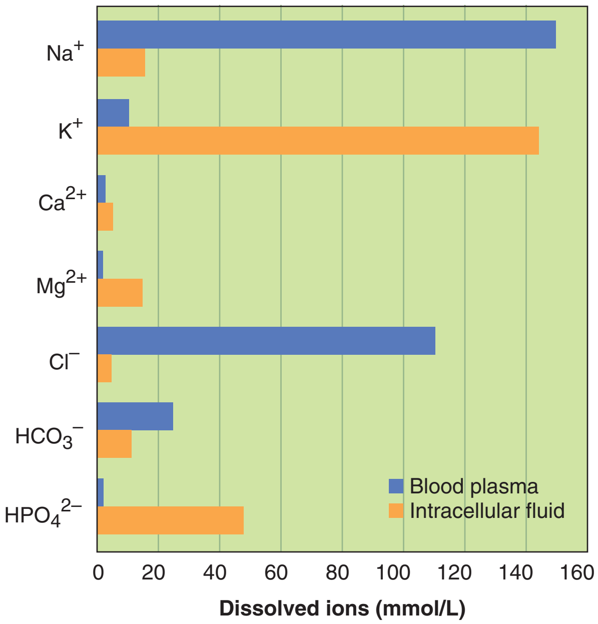
Like sodium, other group 1A atoms form cations with a 1+ charge. Group 2A metal atoms (two valence electrons) and group 3A metal atoms (three valence electrons) form cations with charges of, respectively, 2+ and 3+.

Using the octet rule and electron dot structures to predict the charge on an ion does not always work for transition metal atoms, because most transition metal atoms cannot lose the number of electrons required to reach an octet. For example, chromium (Cr), with 24 total electrons, would need to lose 6 electrons to end up with just 8 in its outermost filled shell. This does not happen because, as just mentioned, atoms rarely gain or lose more than 3 electrons. Table 3.1 shows the cations observed for some of the transition metals.

Ionophores and Biological Ion Transport

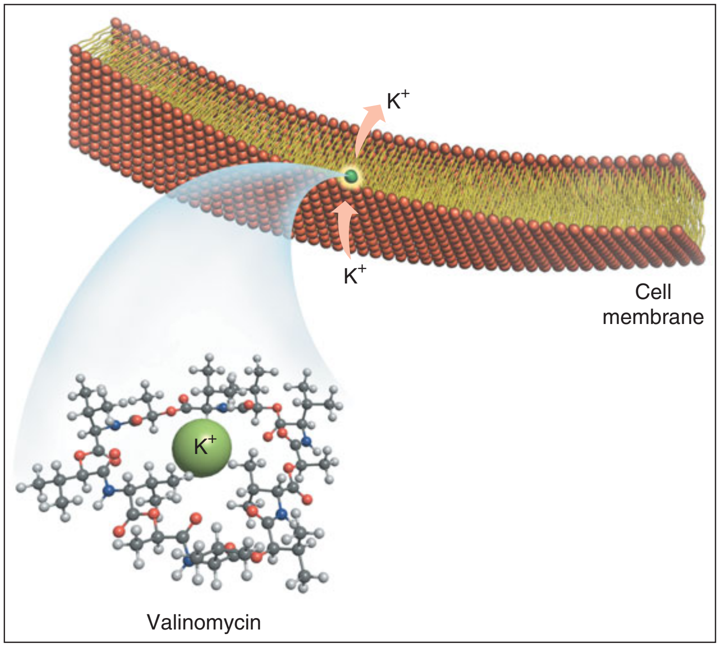
BiochemistryLink

The cell membrane that surrounds each living cell is a barrier that prevents unwanted substances from entering and desired substances from leaving. The differing amount of ions within cells (in intracellular fluid) and outside of cells (in blood plasma) can be maintained because cell membranes control the passage of ions (Figure 3.3). To defend themselves against competing microorganisms, some bacteria produce substances called ionophores. These substances transport ions across cell membranes. *Streptomyces* bacteria make valinomycin, a doughnut-shaped ionophore whose center is the correct size to hold a potassium ion (K^+) (Figure 3.4). This and other ionophores used as antibiotics (bacteria-killing drugs) destroy bacteria by transporting ions across their cell membranes, which upsets the balance of ions and disrupts key biological processes.



■ **FIGURE 3.3**

The major dissolved ions in blood plasma and in intracellular fluid Amounts of dissolved ions are expressed as mmol/L (millimoles of ion per liter of fluid).



■ **FIGURE 3.4**

Valinomycin This ionophore transports K^+ across cell membranes.

SAMPLE PROBLEM 3.2

Drawing electron dot structures of ions

Draw the electron dot structure of each atom and of the ion that it is expected to form.

a. Cs **b.** S

STRATEGY

To solve this problem you need to know how many valence electrons each atom has. You should be able to determine this by looking at a periodic table. Once you know the number of valence electrons and have drawn an electron dot structure, you need to remember that metals lose electrons and nonmetals gain electrons.

SOLUTION

a. $\mathrm{Cs}\cdot$ **b.** $\cdot\ddot{\underset{\cdot\cdot}{\mathrm{S}}}\cdot$

Cs^{+} $:\ddot{\underset{\cdot\cdot}{\mathrm{S}}}:^{2-}$

PRACTICE PROBLEM 3.2

Draw the electron dot structure of each atom and of the ion that it is expected to form.

a. Cl **b.** Se **c.** Ba

3.3 | IONIC COMPOUNDS

- Compounds contain two or more elements.

Early in Chapter 1, matter was defined as *anything that has mass and occupies space*. Chapter 2 introduced the form of matter called the element, which *contains only one type of atom*. A sample of the element iron, for example, contains only iron atoms. Because they contain only one kind of atom, elements are **pure substances**.

Compounds, *matter constructed of two or more chemically combined elements*, are another type of pure substance. Table salt, which contains sodium ions and chloride ions, is an example. A compound is not just a haphazard mixture of elements. Instead, each compound always has the same proportion of the same elements—table salt always has an equal number of sodium and chloride ions. A compound has an identity that is distinct from the identities of the elements that went into making it (Figure 3.5).

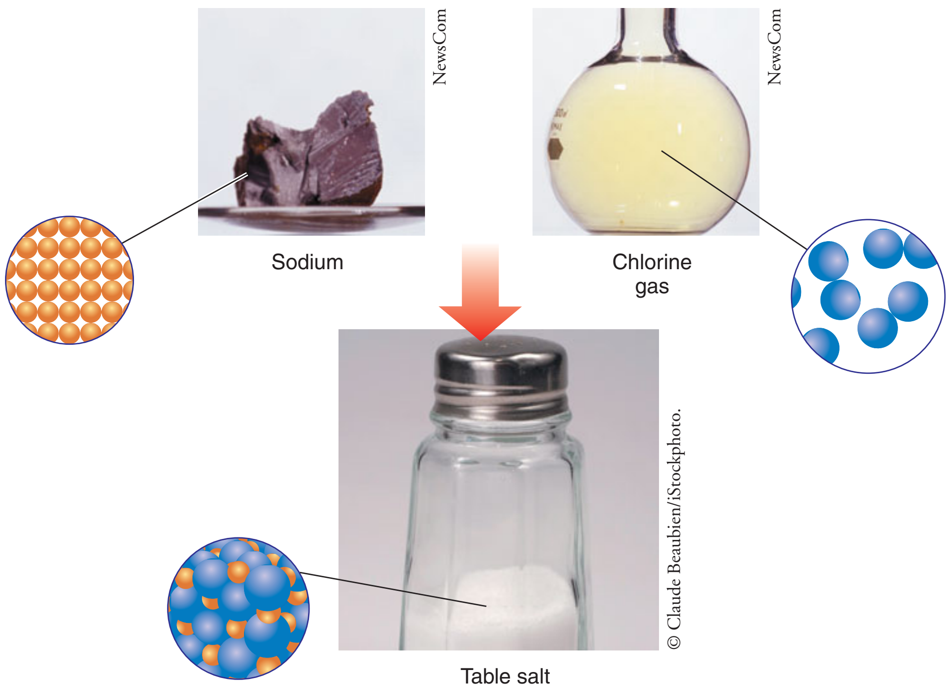


FIGURE 3.5

Compounds The elements sodium and chlorine combine to form the compound sodium chloride (table salt).

Tony Gervis/Getty Images, Inc.

■ **FIGURE 3.6**

Chemical change Iron that is exposed to air and water soon rusts. This is a chemical change.

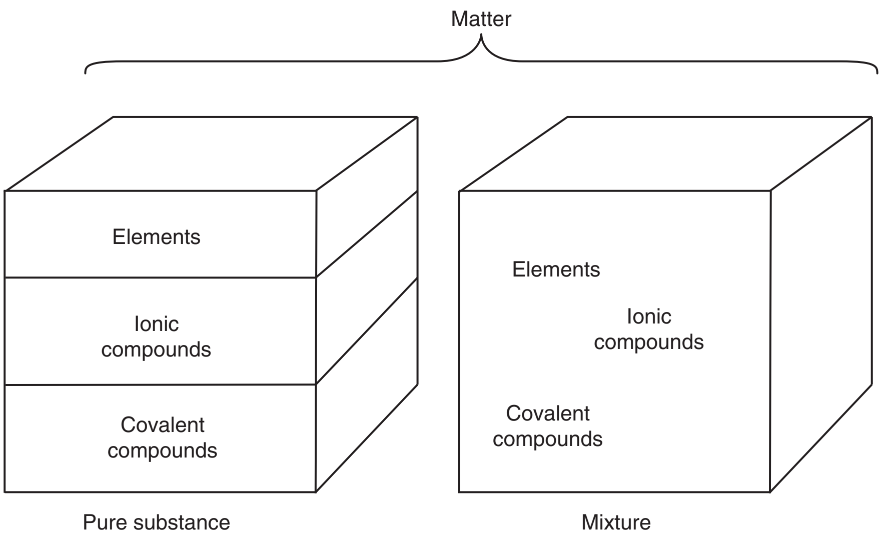
When the atoms in an element or compound combine to form new compounds, **chemical change** has taken place. For example, when a piece of iron rusts (Figure 3.6), a chemical change has occurred because iron (an element) combines with oxygen (a different element) to produce rust (a compound that is made from iron ions and oxygen ions). *The chemical changes that an element or a compound undergo* are called **chemical properties**.

Any matter that *consists only of an element or a compound* is called a pure substance (Figure 3.7). Depending on how elements combine with one another, the compound formed will be either ionic or covalent. (As you continue reading Chapter 3, you will learn the meaning of these terms.) **Mixtures** are *a combination of two or more pure substances.* Mixtures are a topic of Chapter 6.

Ionic compounds are constructed from cations (positively charged ions) and anions (negatively charged ions). The simplest ionic compounds are **binary** (contain just two elements) and contain metal cations bonded to nonmetal anions. Sodium chloride, a binary ionic compound, contains Na^+ and Cl^-, and the opposite charge on these ions provides the **ionic bond** that holds the compound together. Figure 3.8 shows the structure of sodium chloride, which consists of a **crystal lattice** (array) of alternating cations and anions. In this lattice, each ion is surrounded by others of opposite charge.

Did You Know?

In 2010, the U.S. Department of Agriculture revised its dietary guidelines related to salt (sodium chloride) consumption. It is now recommended that a person take in no more than 1500 milligrams of salt each day. This is less than the amount of salt in a McDonald's Quarter Pounder with cheese (3383 mg), a Burger King Whopper (2594 mg), a Wendy's Homestyle Chicken Fillet sandwich (2849 mg), or one Kentucky Fried Chicken original recipe chicken breast (2696 mg).



■ **FIGURE 3.7**

Pure substances and mixtures
A pure substance consists of elements or compounds. Mixtures are a combination of two or more pure substances.

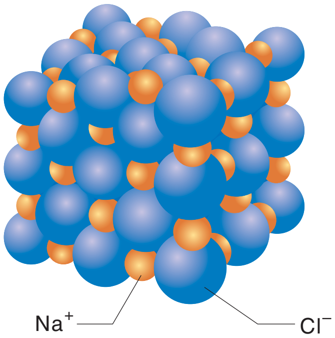


FIGURE 3.8

Sodium chloride A sodium chloride crystal is a lattice of alternating Na^+ and Cl^- ions. Each ion is surrounded by ions of opposite charge.

The formula of any ionic compound is a listing of the relative number of each type of ion that is present, with the cation listed before the anion. Sodium chloride, whose formula is NaCl, contains an equal number of Na^+ and Cl^- ions. $MgCl_2$ is a binary ionic compound that has 2 Cl^- for each Mg^{2+}. Like all ionic compounds, $MgCl_2$ is neutral—the 2+ charge on each magnesium ion is countered by two 1− charges, one from each of the chloride ions. Note that the formulas NaCl and $MgCl_2$ do not specify the charges on the ions. It is assumed that these can be determined, based on knowledge of the periodic table and the octet rule.

If asked to write the formula for the ionic compound that forms between lithium ions and sulfide ions, you could begin by determining the charge on each of the ions. A lithium atom has one valence electron. Losing this electron to reach an octet creates a lithium ion with a charge of 1+. A sulfur atom has six valence electrons. Gaining two electrons to reach an octet produces a sulfide ion with a charge of 2−. To obtain a neutral compound, each S^{2-} ion will require two Li^+ ions. The formula of the ionic compound will be Li_2S.

Things are a little more complicated when transition metals are involved, because most can appear in more than one cation form (Table 3.1). The formula of an ionic compound consisting of iron ions and sulfide ions would depend on the particular iron ion that is present. Fe^{2+} would form the compound FeS (one Fe^{2+} ion requires one S^{2-} ion), while Fe^{3+} would form Fe_2S_3 (two Fe^{3+} ions would counter the charge on three S^{2-} ions).

- Cations and anions are held to one another by ionic bonds.

SAMPLE PROBLEM 3.3

Predicting formulas of ionic compounds

Write the formula of the ionic compound that forms between

a. sodium ions and fluoride ions
b. calcium ions and oxide ions
c. aluminum ions and oxide ions
d. sodium ions and sulfide ions

STRATEGY

You must first determine the charge expected on each of the ions and then decide how many of each will be required to produce a neutral compound. Once you have done this, the formula lists the cation followed by the anion, and includes the relative numbers of each.

SOLUTION

a. NaF (used to prevent cavities)
b. CaO (lime—used in plaster, stucco, and mortar)
c. Al_2O_3 (used in the manufacture of dental cements)
d. Na_2S (used in the manufacture of rubber and in ore refining)

PRACTICE PROBLEM 3.3

Write the formula of the ionic compound that forms between

a. lithium ions and chloride ions
b. lithium ions and oxide ions
c. lithium ions and nitride ions

Ionic Compounds Containing Polyatomic Ions

Polyatomic ions also form ionic compounds. The food preservative sodium sulfite (Na_2SO_3) is the combination of sodium ions (Na^+) and sulfite ions (SO_3^{2-}). To produce a neutral ionic compound, each sulfite ion (2− charge) requires two sodium ions (1+ charge from each).

When more than one copy of a particular polyatomic ion is present in an ionic compound, the formula of the ion is usually surrounded by parentheses, as is the case for the laxative $Mg(OH)_2$ and the expectorant $(NH_4)_2CO_3$.

When an ionic compound contains polyatomic ions, interpreting its formula depends on being familiar with the formulas of the polyatomic ions involved. For example, $Mg(OH)_2$ contains Mg^{2+} and OH^- ions and $(NH_4)_2CO_3$ contains NH_4^+ and CO_3^{2-} ions. It is important to note that in ionic compounds, polyatomic ions "act as one." The compound $NaNO_2$ consists of Na^+ and NO_2^- ions, not some combination of ions formed from Na, N, and O atoms.

FIGURE 3.9

Ionic compounds

Pictured here, clockwise from the upper left, are iron(II) sulfate, iron(III) sulfate, copper(II) sulfate, copper(II) carbonate, and sodium chloride. Ionic compounds containing transition metal ions, such as the first four mentioned here, are often brightly colored.

Naming Ionic Compounds

When naming ionic compounds, the cation name is placed before the anion name. Lithium ions (Li^+) combine with bromide ions (Br^-) to form lithium bromide (LiBr) and ammonium ions (NH_4^+) combine with nitrate ions (NO_3^-) to form ammonium nitrate (NH_4NO_3).

The number of times that an ion appears in the formula of an ionic compound is *not* specified in the name, so $BaCl_2$ is called barium chloride, not barium dichloride. Similarly, Na_2SO_4 is sodium sulfate and $Mg(HCO_3)_2$ is magnesium hydrogen carbonate. It is assumed that the formula can be determined from the name, because the charges on the various ions are known. For example, calcium bromide must have the formula $CaBr_2$ because calcium ions always have a charge of 2+ and bromide ions always have a charge of 1−. This means that two bromide ions will combine with one calcium ion to create a neutral ionic compound.

Assigning names works the same way when an ionic compound contains transition metal ions (Figure 3.9). CuCl, the combination of copper(I) ion (Cu^+) and chloride ion (Cl^-), is called copper(I) chloride, and $CuCl_2$ is named copper(II) chloride. Copper(I) chloride and copper(II) chloride are also known, respectively, as cuprous chloride and cupric chloride (Table 3.1). Iron(II) hydroxide has the formula $Fe(OH)_2$—one Fe^{2+} ion requires two OH^- ions to form a neutral compound.

Ionic compounds are widely used in medicine, by industry, and around the house. Table 3.4 lists some of their common uses.

SAMPLE PROBLEM 3.4

Predicting formulas of ionic compounds that contain polyatomic ions

Write the formula of the ionic compound that forms between

a. sodium ions and cyanide ions
b. sodium ions and dichromate ions
c. calcium ions and dichromate ions
d. calcium ions and phosphate ions

STRATEGY

You must determine the charge on each ion and make sure that the formula contains enough of each to produce a neutral compound. A neutral compound will have the same total number of positive charges and negative charges.

SOLUTION

a. NaCN **b.** $Na_2Cr_2O_7$ **c.** $CaCr_2O_7$ **d.** $Ca_3(PO_4)_2$

PRACTICE PROBLEM 3.4

Write the formula of the ionic compound that forms between

a. ammonium ions and hydrogen sulfate ions
b. ammonium ions and phosphate ions
c. strontium ions and phosphate ions
d. strontium ions and sulfate ions

Did You Know?

The green color of the Statue of Liberty is due to a patina that formed on its copper surface. This patina is a mixture of ionic compounds produced when copper metal is exposed to air. The first compound to form is Cu_2O, which has a red color. Cu_2O reacts further to form the black CuO and other ionic compounds, including the green $Cu_2CO_3(OH)_2$.

TABLE | 3.4 THE USES OF SOME IONIC COMPOUNDS

Name	Formula	Use
Ammonium carbonate	$(NH_4)_2CO_3$	Smelling salts
Barium sulfate	$BaSO_4$	Compound used to help view internal organs in x-ray studies
Calcium carbonate	$CaCO_3$	Antacid
Calcium sulfate	$CaSO_4$	Plaster casts
Lithium carbonate	Li_2CO_3	Treatment for bipolar disorder
Magnesium hydroxide	$Mg(OH)_2$	Milk of magnesia
Magnesium sulfate	$MgSO_4$	Laxative
Silver nitrate	$AgNO_3$	Prevention of eye infections in newborns
Sodium bicarbonate	$NaHCO_3$	Baking soda and antacid
Sodium hydroxide	NaOH	Drain cleaner
Sodium iodide	NaI	Source of iodide ion for the thyroid
Sodium nitrate	$NaNO_3$	Food preservative
Sodium nitrite	$NaNO_2$	Meat preservative
Sodium acetate	CH_3CO_2Na[a]	Foot and hand warmers

[a]For ionic compounds involving polyatomic ions with an organic or biochemical source (such as $CH_3CO_2^-$), the formula sometimes lists the anion before the cation.

Pass the Salt, Please

HealthLink

When someone asks you to pass the salt, chances are good that they are referring to table salt, the sodium chloride that most people have on the table or in a kitchen cupboard. Table salt is usually obtained from salt mines and then refined (purified). In stores, you can typically buy either "plain" or "iodized" table salt. Iodized salt contains small amounts of various iodine-containing compounds that have been added to help people get their 0.15 mg/day recommended daily intake of this essential element (Table 2.4).

There are many other types of salt available than simple table salt. Sea salt, for example, is produced by evaporating seawater. Because it is usually sold in an unrefined form, sea salt contains sodium chloride plus smaller amounts of other elements that are present in seawater, including sulfur, magnesium, potassium, calcium, iodine, and iron.

Grey salt is sea salt from the coast of Brittany, France. Its light grey color (Figure 3.10*a*) comes from the clay of the salt flats where it is obtained. Fleur de Sel is a high-quality French sea salt that is hand harvested from evaporation ponds. Fleur de Sel is over 50 times more expensive than regular table salt.

Hawaiian sea salt is sodium chloride with added Alae (volcanic baked red clay). Fe_2O_3 in the clay produces a reddish color and the other minerals give the salt its distinctive flavor (Figure 3.10*b*). Black Indian salt (Kala Namak) is sodium chloride with small amounts of Na_2SO_4; FeS, which produces the dark color of the salt (Figure 3.10*c*); and H_2S which, along with iron(II) sulfide, gives black Indian salt an odor of sulfur. Himalayan salt blocks, which are mined in Pakistan, are sometimes used as platters to serve sushi or to cook seafood (Figure 3.10*d*).

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(*a*) (*b*) (*c*) (*d*)

FIGURE 3.10

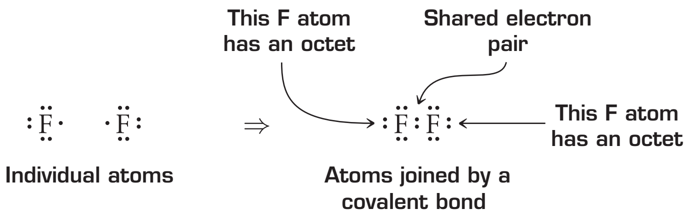
Specialty salts (*a*) Grey sea salt is obtained from salt flats. (*b*) Hawaiian sea salt is a mixture of NaCl and volcanic clay. (*c*) Black Indian salt gets its color from iron(II) sulfide. (*d*) Food can be cooked on Himalayan salt plates.

3.4 | COVALENT BONDS

To reach an octet, metals lose electrons (Na becomes Na^+) and nonmetals gain electrons (Cl becomes Cl^-). For nonmetals, a second option is available for attaining an octet—valence electrons can be shared.

An example of this is what happens when two F atoms, each of which has seven valence electrons, interact with one another. When the atoms reach an appropriate distance, one pair of electrons is shared and each atom ends up with an octet. This *shared pair of valence electrons* is called a **covalent bond**.

■ In a covalent bond a pair of valence electrons is shared between two nonmetal atoms.

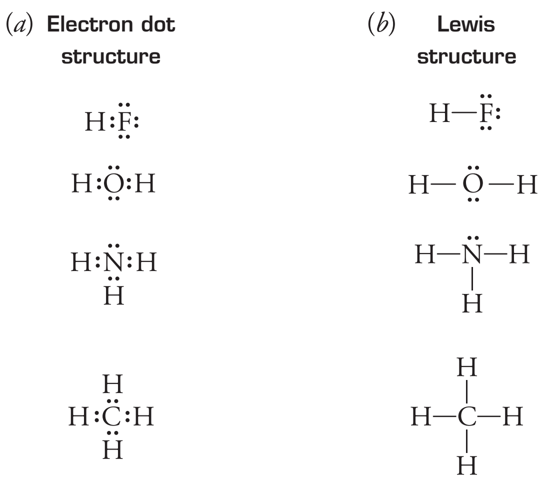


Generally, *the number of covalent bonds that a nonmetal atom forms is the same as the number of electrons that it needs to gain to reach an octet.* Fluorine atoms, with seven valence electrons, form one covalent bond because the extra electron gained by sharing is enough to complete an octet. Atoms of the other second period nonmetals, oxygen (six valence electrons), nitrogen (five valence electrons), and carbon (four valence electrons) form, respectively, two, three, and four covalent bonds. Hydrogen atoms form just one covalent bond. Figure 3.11 shows the covalent bonding that can take place between H atoms and F, O, N, and C atoms.

The drawings in Figure 3.11*a* are electron dot structures—valence electrons are shown using dots. In an alternative approach, called the **Lewis method** (Figure 3.11*b*), *each pair of shared bonding electrons is represented by a line*. In all of these drawings, the *valence electrons not involved in bonds* are called **nonbonding electrons**.

In Figure 3.11, atoms in the molecules are joined by **single bonds** (single covalent bonds), in which *one pair of electrons is shared* by two atoms. These are not the only covalent bonding patterns known, however. Under normal circumstances it is possible for two atoms to share up to three pairs of electrons. Atoms involved in a **double bond** share two pairs of electrons and atoms involved in a **triple bond** share three pairs of electrons.

An oxygen atom needs two electrons to gain an octet, so a given O atom is able to form either two single bonds or one double bond. A carbon atom, which requires four electrons



■ **FIGURE 3.11**

Covalent bonds Nonmetal atoms can satisfy the octet rule by forming single covalent bonds—generally one single bond for each valence electron required to complete an octet. To represent covalent bonds, (*a*) electron dot structures use pairs of electron dots and (*b*) Lewis structures use lines.

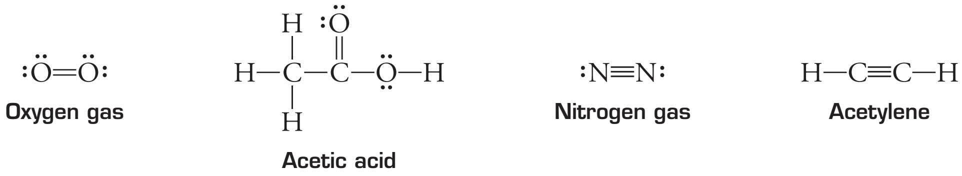


FIGURE 3.12

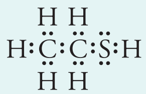
Single, double, and triple bonds To reach an octet, oxygen atoms form two covalent bonds, nitrogen atoms form three covalent bonds, and carbon atoms form four covalent bonds.

to obtain an octet, has a number of covalent bonding options. It can form four single bonds, two double bonds, or various combinations of single, double, and triple bonds, as long as the total number of bonds is four (Figure 3.12).

SAMPLE PROBLEM 3.5

Understanding electron dot and Lewis structures

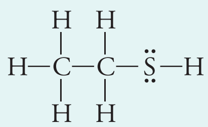
Using the Lewis method, draw ethanethiol, the compound added to natural gas to give it a detectable odor.



STRATEGY

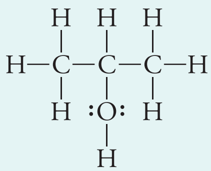
Electron dot and Lewis structures differ in that each pair of valence electrons involved in a covalent bond is shown as a pair of dots in an electron dot structure and as a line in a Lewis structure.

SOLUTION



PRACTICE PROBLEM 3.5

Draw the electron dot structure of isopropyl alcohol (rubbing alcohol).



SAMPLE PROBLEM 3.6

Evaluating Lewis structures

Hydrogen cyanide contains one atom each of H, C, and N. Which Lewis structure is the correct one for hydrogen cyanide?

H═C═N꞉ H≡C—N꞉ H—C≡N꞉

STRATEGY

As Figure 3.11 shows, each nonmetal atom is able to form a particular number of covalent bonds to reach an octet. While each of the drawings above has the same total number of electrons, only one has the expected number of bonds for H (1 bond), C (4 bonds), and N (3 bonds).

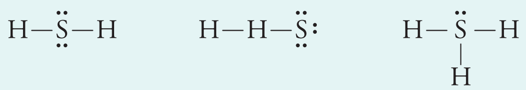
SOLUTION

The correct structure is the one with the triple bond between the C and N atoms. In this drawing the C and N atoms each have an octet of electrons and the H atom has two electrons.

$$H-C\equiv N:$$

PRACTICE PROBLEM 3.6

Which structure is expected from the combination of hydrogen and sulfur atoms?



3.5 | MOLECULES

The drawings in Figure 3.11 represent **covalent compounds** or **molecules**—*uncharged groups of atoms connected to one another by covalent bonds.* An alternative to drawing the structure of a molecule is to give its **molecular formula**, which lists the number of each type of atom that is present. For example, the molecular formulas of the molecules shown in Figure 3.11 are HF, H_2O, NH_3, and CH_4.

■ The atoms in molecules are held together by covalent bonds.

Most molecules are compounds, because they contain atoms of two or more different elements. Some molecules, however, are elements, because they contain just one type of atom. Seven elements (H_2, N_2, O_2, F_2, Cl_2, Br_2, and I_2) appear as **diatomic** (*two atom*) molecules. Oxygen is also found as the triatomic molecule called ozone (O_3).

Naming Binary Molecules

Molecules come in all shapes and sizes. How they are assigned names usually depends on the type of molecule that they are. Chapter 8, for example, will introduce a set of rules used to name organic molecules, a very large class of molecules that contain carbon atoms. Here, the relatively simple procedure used to name binary molecules—those that contain just two different elements—will be presented. Binary molecules are named by listing the elements in order of appearance in the molecular formula and changing the ending of the name of the second element to *ide*. For example, HF is hydrogen fluoride and HCl is hydrogen chloride.

■ Binary molecules contain two different elements.

When naming ionic compounds, the number of each type of ion is not specified ($CaCl_2$ is calcium chloride, not calcium dichloride) because ions always combine in fixed ratios to form a neutral compound. In binary molecules, however, atoms can sometimes combine in several different ways. Sulfur and oxygen atoms can bond to form two different molecules, SO_2 and SO_3. To distinguish such molecules by name, prefixes (Table 3.5) are added to specify the number of each type of atom that is present: SO_2 is named sulfur dioxide and SO_3 is named sulfur trioxide.

A few rules apply to using prefixes. First, names should not begin with "mono," so NO_2 is called nitrogen dioxide, not mononitrogen dioxide. Also, if adding a prefix places

TABLE 3.5 PREFIXES USED FOR NAMING BINARY MOLECULES.[a]

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

[a]Names should not begin with "mono" (CO_2 is carbon dioxide, not monocarbon dioxide). When adding a prefix places two vowels together, an "a" or "o" ending on the prefix is often dropped (CO is carbon monoxide, not carbon monooxide).

two vowels next to one another, an "a" or "o" ending on the prefix is often dropped. The molecule NO is nitrogen monoxide, not nitrogen monooxide.

Other examples of molecule names include:

- CO — carbon monoxide — (a poisonous, odorless gas)
- CO_2 — carbon dioxide — (a product of human metabolism)
- $SiCl_4$ — silicon tetrachloride — (used to prepare smoke screens in warfare)
- N_2O_5 — dinitrogen pentoxide — (used in the synthesis of certain organic compounds)

Sometimes binary molecules are better known by other names. Among these are H_2O (water, instead of dihydrogen oxide) and H_2S (hydrogen sulfide, instead of dihydrogen sulfide).

SAMPLE PROBLEM 3.7

Naming binary molecules

Name each binary molecule.

a. SiO_2 (used in glass manufacture)
b. SF_6 (used in electrical circuits)
c. P_2O_5 (a drying agent)

STRATEGY

When naming binary molecules, the element names are listed in the same order as given in the formula, the ending on the name of last element is changed to *ide*, and the number of times each appears is specified (see Table 3.5).

SOLUTION

a. silicon dioxide
b. sulfur hexafluoride
c. diphosphorus pentoxide

PRACTICE PROBLEM 3.7

Name each binary molecule.

a. $SiBr_4$ **b.** P_2O_3 **c.** P_4Se_3

Dental Fillings

HealthLink

If you need to have a tooth filled, your experience will be very different from that of your ancestors. As recently as the mid-1800s it was standard practice to press pellets of lead, tin, or gold into dental cavities. Prior to that, anything that would plug the hole in the tooth (cork, resin, and others) was used. One problem with all of these filling materials was that they were not durable and tended to break or fall out. Today, the two most commonly used dental filling materials are amalgam and tooth-colored composites.

Any mixture of mercury and one or more other metals is called an amalgam. In dental amalgam, mercury is combined with silver and lesser amounts of tin, copper, and zinc. This mixture, which is soft to begin with and can be pressed into a tooth, sets quickly to form a hard filling. Composites are a special type of plastic made from organic compounds. They are soft and pliable until being hardened by exposure to an intense blue or ultraviolet light.

There are pros and cons to choosing either type of filling. Amalgam fillings are the stronger of the two and can last for 10 to 15 years. By comparison, composite fillings last an average of 5 years. While amalgam fillings are more durable, composite fillings can actually strengthen teeth because, unlike amalgam fillings, they bond directly to tooth material. This means that less of a tooth needs to be drilled away to prepare for a composite filling than for an amalgam one. Composite fillings are more expensive, but for those concerned with the appearance of their teeth, the extra cost may be worthwhile—composite fillings match tooth color, while amalgam does not (Figure 3.13).

Currently, one of the biggest issues related to dental fillings is whether amalgam is safe to have in your mouth. The concern is that some scientific studies have shown that amalgam fillings release trace amounts of mercury, a toxic element that at high enough levels can cause sometimes fatal neurological and brain damage. Opponents of amalgam use claim that no level of mercury is safe and that anyone with amalgam fillings should have them replaced with composite ones. Those in the pro-amalgam camp say that a person's daily exposure to mercury from amalgam fillings is not a concern because it is much lower than the average daily exposure to the mercury present in food and water as a result of pollution. Which is better for you should you need a filling? That is for you and your dental professional to decide.

Ottmar Bierwagen/Spectrum Photofile

FIGURE 3.13

Dental fillings Fillings are commonly made from amalgam (top) or composites (bottom).

3.6 FORMULA WEIGHT, MOLECULAR WEIGHT, AND MOLAR MASS

Chapter 2 introduced atomic weight, the average mass of the naturally occurring atoms of an element. When dealing with an ionic compound, it can be helpful to know its **formula weight**, *the sum of the atomic weights of the elements in the formula*. Sodium chloride (NaCl) has a formula weight of 58.44 amu, which is determined by adding the atomic weights of sodium and chlorine.

$$\begin{array}{ccccc} Na^+ & & Cl^- & & NaCl \\ 22.99\text{ amu} & + & 35.45\text{ amu} & = & 58.44\text{ amu} \end{array}$$

In this calculation, the atomic weights of Na and Cl were used, even though NaCl is composed of Na^+ and Cl^- ions. Compared to the mass of the protons and neutrons that make up the nucleus of Na and Cl atoms, electron mass is negligible, so losing or gaining electrons to form ions has no effect on atomic weight—Na and Na^+ have the same atomic weight, as do Cl and Cl^-.

For more complex ionic compounds, calculating formula weight works the same way. The formula weight of copper(II) nitrate, $Cu(NO_3)_2$, is 187.52 amu.

$$\begin{array}{ccccccc} Cu^{2+} & & \multicolumn{3}{c}{\overbrace{\quad 2\ NO_3^- \quad}} & & Cu(NO_3)_2 \\ & & 2\ N & & 6\ O & & \\ 63.55\ \text{amu} & + & 2 \times 14.01\ \text{amu} & + & 6 \times 16.00\ \text{amu} & = & 187.52\ \text{amu} \end{array}$$

■ The molar mass of an ionic compound (the mass in grams of one mole) is equal to its formula weight in amu.

The mass in grams of one mole of an ionic compound (its molar mass) is numerically equivalent to its formula weight (in amu). For example, $Cu(NO_3)_2$ has a formula weight of 187.52 amu, so its molar mass is 187.52 g/mol.

This relationship allows conversions of the following type to be carried out.

- The formula weight of $AgNO_3$, used to prevent eye infection in newborns, is 169.88 amu. A sample containing 0.500 mol of $AgNO_3$ has a mass of 84.9 g.

$$0.500\ \cancel{\text{mol } AgNO_3} \times \frac{169.88\ \text{g } AgNO_3}{1\ \cancel{\text{mol } AgNO_3}} = 84.9\ \text{g } AgNO_3$$

- A 7.28 g sample of $AgNO_3$ is 4.29×10^{-2} mol.

$$7.28\ \cancel{\text{g } AgNO_3} \times \frac{1\ \text{mol } AgNO_3}{169.88\ \cancel{\text{g } AgNO_3}} = 4.29 \times 10^{-2}\ \text{mol } AgNO_3$$

SAMPLE PROBLEM 3.8

Calculations involving formula weight

a. What is the formula weight of the food preservative sodium sulfite (Na_2SO_3)?
b. What is the mass of 1.50 mol of Na_2SO_3?

STRATEGY

In part a, you must add up the individual atomic weights of each element in the formula. Solving part b involves a conversion factor that uses the molar mass of Na_2SO_3.

SOLUTION

a. 126.05 amu

$$\begin{array}{ccccccc} 2\ Na^{+} & & \multicolumn{3}{c}{\overbrace{\quad SO_3^- \quad}} & & Na_2SO_3 \\ & & S & & 3\ O & & \\ 2 \times 22.99\ \text{amu} & + & 32.07\ \text{amu} & + & 3 \times 16.00\ \text{amu} & = & 126.05\ \text{amu} \end{array}$$

b. 189 g Na_2SO_3

$$1.50\ \cancel{\text{mol } Na_2SO_3} \times \frac{126.05\ \text{g } Na_2SO_3}{1\ \cancel{\text{mol } Na_2SO_3}} = 189\ \text{g } Na_2SO_3$$

PRACTICE PROBLEM 3.8

a. What is the formula weight of baking soda ($NaHCO_3$)?
b. What is the mass of 0.315 mol of baking soda?

Molecular Weight

Just as elements have an atomic weight and ionic compounds have a formula weight, molecules have a **molecular weight**—*the sum of the atomic weights of the elements in the molecular formula*. The molecular weight of water (H_2O) is 18.02 amu, which is determined from the atomic weights of hydrogen and oxygen.

$$\begin{array}{ccccc} 2\text{ H} & & \text{O} & & H_2O \\ 2 \times 1.01\text{ amu} & + & 16.00\text{ amu} & = & 18.02\text{ amu} \end{array}$$

The molecular weight of sulfur trioxide (SO_3) is 80.07 amu.

$$\begin{array}{ccccc} \text{S} & & 3\text{O} & & SO_3 \\ 32.07\text{ amu} & + & 3 \times 16.00\text{ amu} & = & 80.07\text{ amu} \end{array}$$

Since the molecular weight of sulfur trioxide is 80.07 amu, its molar mass is 80.07 g/mol and 0.0210 mol has a mass of 1.68 g.

$$0.0210 \cancel{\text{mol } SO_3} \times \frac{80.07\text{ g } SO_3}{1 \cancel{\text{mol } SO_3}} = 1.68\text{ g } SO_3$$

■ The molar mass of a molecule (the mass in grams of one mole) is equal to its molecular weight in amu.

SAMPLE PROBLEM 3.9

Calculations involving molecular weight

a. What is the molecular weight of chloroform ($CHCl_3$)?
b. What is the mass of 2.50 mol of chloroform?

STRATEGY

You can calculate the molecular weight of chloroform by adding up the atomic weights of carbon, hydrogen, and chlorine. (Remember to add in the atomic weight of chlorine three times, since it appears three times in the molecular formula.) Part b of the problem can be solved by using a conversion factor related to the molar mass of $CHCl_3$.

SOLUTION

a. 119.37 amu

$$\begin{array}{ccccccc} \text{C} & & \text{H} & & 3\text{ Cl} & & CHCl_3 \\ 12.01\text{ amu} & + & 1.01\text{ amu} & + & 3 \times 35.45 & = & 119.37\text{ amu} \end{array}$$

b. 298 g

$$2.50 \cancel{\text{mol } CHCl_3} \times \frac{119.37\text{ g } CHCl_3}{1 \cancel{\text{mol } CHCl_3}} = 298\text{ g } CHCl_3$$

PRACTICE PROBLEM 3.9

a. What is the molecular weight of glycine ($C_2H_5NO_2$), one of the amino acids used to build proteins?
b. What is the mass of 4.00 mol of glycine?
c. How many glycine molecules are present in 0.00552 g of glycine?

Nitric Oxide

HealthLink

Many people refer to the local newspaper or to the Internet for a daily report of the local Air Quality Index. This index predicts how clear or polluted the air will be on a given day. One of the air pollutants tracked by the Air Quality Index is nitrogen dioxide (NO_2) which, along with nitrogen monoxide (NO), is formed when nitrogen and oxygen in the air combine when gasoline is burned in the engines of cars and trucks. Nitrogen monoxide, known more commonly as nitric oxide, is a colorless gas that has minimal effects on human health, while nitrogen dioxide is a brown gas that can increase the risk of asthma, respiratory infection, lung tissue damage, and chronic lung disease. Sunlight helps nitric oxide combine with oxygen in the air (O_2) to form nitrogen dioxide. This is why smog, whose color comes largely from NO_2, often worsens late in the afternoon. When mixed with the moisture in clouds, NO and NO_2 react to form an acidic compound that falls with rain (acid rain). Although nitric oxide is not especially harmful, this compound has had a bad reputation because of its association with atmospheric pollution.

In nitric oxide a nitrogen atom and an oxygen atom are joined by a double bond. Unlike the other molecules shown in this chapter, one of the atoms, the nitrogen, does not have an octet. This makes the molecule unstable and readily able to react with many other compounds.

$$:\dot{N}=\ddot{O}:$$

In the past few decades, many biochemical studies involving nitric oxide have been carried out. The story behind these experiments began in the mid 19th century, soon after the discovery of the explosive called nitroglycerin. It was noticed that people who worked with nitroglycerin often experienced splitting headaches. These nitroglycerin-induced headaches were found to be the result of vasodilatation (relaxation) of the blood vessels in the brain. This led to nitroglycerin being used as a medication for treating angina, a constriction of the arteries that carry blood to the heart. It wasn't until well after a century of use that scientists discovered that the body produces nitric oxide from nitroglycerin and that it is the nitric oxide that causes coronary arteries to dilate, reducing the symptoms of angina. The 1988 Nobel Prize in Medicine was awarded to three scientists who were involved in this nitric oxide research.

It is now known that nitric oxide is produced throughout the body and that it has many functions (Figure 3.14). It serves as a neurotransmitter (a signaling molecule for the nervous system), regulates blood pressure, controls muscles that dilate arteries and blood vessels, plays a role in inflammation and shock, and is used by the immune system to help fight infections.

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FIGURE 3.14

Transfusions increase the risk of heart attack Patients who receive blood transfusions have a greater risk of heart attack and stroke than those who do not. It was recently determined that this is due to a lack of nitric oxide in stored blood. One of the important biological roles that nitric oxide has is helping with the transfer of oxygen from red blood cells to the body. Because stored red blood cells rapidly lose their nitric oxide, they become less effective at providing the body with needed oxygen.

AT THE DENTIST'S OFFICE . . . REVISITED

Tooth enamel is composed mostly of a mineral called hydroxyapatite, an ionic compound with the formula $Ca_5(OH)(PO_4)_3$. Tooth decay is what happens when enamel is damaged by the breakdown of hydroxyapatite. This demineralization takes place when tooth enamel is exposed to acids produced by the bacteria present in dental plaque.

Fluoride ion (F^-) can be used to prevent tooth decay. When children are given fluoride, it is incorporated into their develop-

© Dmitriy Shironosov/iStockphoto

ing teeth through the formation of fluorapatite, $Ca_5F(PO_4)_3$. This mineral is stronger than hydroxyapatite and is not broken down as easily by acids. Adults also benefit from the use of fluoride, because existing hydroxyapatite can be converted into the more durable fluorapatite.

Fluoride has other benefits as well. It can reverse some demineralization damage through remineralization, the formation of new fluorapatite. Some studies have also shown that fluoride reduces the ability of bacterial plaque to produce the acids that cause cavities.

A number of options are available for administering fluoride. In some areas, fluoride is naturally present in the water. In others, water is fluoridated by addition of sodium fluoride or another fluoride-containing ionic compound. Many types of toothpaste contain fluoride, and fluoride tablets are available by prescription. The fluoride rinses or gels used at the dentist office contain higher levels of this anion.

THINKING IT THROUGH

1. In some cities, there is continuing debate about the issue of fluoridating the water supply. What are the pros and cons of doing so?
2. Write the individual ions present in hydroxyapatite and give their names. Do the charges balance out to give a neutral compound?

CHAPTER 3 OBJECTIVES

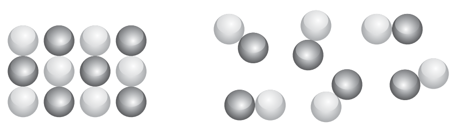
OBJECTIVE	SUMMARY	SECTION	SAMPLE AND PRACTICE PROBLEMS	END OF CHAPTER PROBLEMS
1. Define the term **ion** and describe the naming of **monoatomic cations** and **anions**.	**Ions** are charged atoms or groups of atoms. Ions with a positive charge are **cations** and those with a negative charge are **anions**. For **monoatomic ions** of representative elements, cations are named using the element name (sodium ion) and anions are named by changing the ending of the element name to "ide" (chloride ion). For transition metal cations, the charge is indicated using Roman numerals (iron (III) ion). Alternatively, the relative charge on related transition metal ions can be indicated by ending the name in "ous" or "ic" (ferrous ion and ferric ion).	3.1	3.1	3.3–3.14, 3.19–3.20
2. Describe the naming of **polyatomic** cations and anions.	While **polyatomic** ion names must be memorized, there are some useful patterns to recognize. An ion whose name ends in "ite" has one less O atom than a related one whose name ends in "ate" (nitrite ion and nitrate ion). The use of "hydrogen" indicates the difference in the relative number of H atoms for related ions (carbonate ion and hydrogen carbonate ion).	3.1		3.15–3.18

OBJECTIVE	SUMMARY	SECTION	SAMPLE AND PRACTICE PROBLEMS	END OF CHAPTER PROBLEMS
3. Explain how the electron dot structure of a representative element atom can be used to predict the charge on the monoatomic ion that it forms.	Atoms of representative elements gain or lose just enough valence electrons to reach an **octet**. For nonmetal atoms, ions are represented by adding enough electrons to electron dot structures to give 8 total (each added electron increases the negative charge by one). For metal atoms, ions are represented by removing enough electrons from electron dot structures to give 0 (each removed electron increases the positive charge by one).	3.2	3.2	3.21–3.28
4. Name and write the formulas of **ionic compounds** and **binary molecules**.	**Ionic compounds** are named by combining the name of the cation with the name of the anion (relative numbers of each are not specified). **Binary molecules** (covalent compounds that contain atoms of just two different elements) are named by listing the element names in the order that they appear in the formula, with the ending of the second element's name changed to "ide." Prefixes ("mono," "di," etc.) are used to specify relative number of each type of atom that is present.	3.3, 3.5	3.3, 3.4, 3.7	3.29–3.44, 3.53–3.68
5. Predict the number of **covalent bonds** that a nonmetal atom will form and draw electron dot and **Lewis** structures of molecules.	A nonmetal atom usually forms the same number of **covalent bonds** as the number of valence electrons it needs to reach an octet. In the electron dot structure of a molecule, **bonding** and **nonbonding electrons** are shown as dots. In **Lewis structures**, a line represents one pair of bonding electrons.	3.4, 3.5	3.5, 3.6	3.45–3.52
6. Define the terms **formula weight** and **molecular weight**. Use the molar mass of a compound to carry out conversions involving moles and mass.	The **formula weight** of an ionic compound and the **molecular weight** of a molecule is the sum of the atomic weights of the elements in its formula. The molar mass of an ionic compound or molecule (in grams) is equal to its formula weight or molecular weight (in amu). Having the unit g/mol, molar mass can be used in converting between grams and moles of compound.	3.6	3.8, 3.9	3.69–3.80

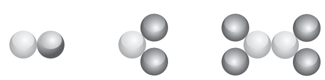
END OF CHAPTER PROBLEMS

Answers to problems whose numbers are printed in color are given in Appendix C. More challenging questions are marked with an asterisk.

3.1 Which picture represents molecules and which one represents an ionic compound?



3.2 Atoms of two nonmetal elements can combine to form the three compounds shown. Are these compounds ionic or are they binary molecules?



3.1 Ions

3.3 Give the total number of protons and electrons in each ion.
a. K^+ **b.** Mg^{2+} **c.** P^{3-}

3.4 Give the total number of protons and electrons in each ion.
a. Li^+ **b.** Al^{3+} **c.** S^{2-}

3.5 Give the total number of protons and electrons in each ion.
a. Fe^{2+} **b.** Fe^{3+} **c.** Cu^+ **d.** Cu^{2+}

3.6 Give the total number of protons and electrons in each ion.
a. Cr^{2+} **b.** Cr^{3+} **c.** Sn^{2+} **d.** Sn^{4+}

3.7 Give the total number of protons, neutrons, and electrons in each ion.
a. ${}^{63}_{29}Cu^+$ **b.** ${}^{19}_{9}F^-$ **c.** ${}^{37}_{17}Cl^-$

3.8 Give the total number of protons, neutrons, and electrons in each ion.
a. ${}^{23}_{11}Na^+$ **b.** ${}^{16}_{8}O^{2-}$ **c.** ${}^{35}_{17}Cl^-$

3.9 Identify the following monoatomic ions.
a. 15 protons and 18 total electrons
b. 20 protons and 18 total electrons
c. 7 protons and 8 valence electrons
d. 25 protons and 22 total electrons

3.10 Identify the following monoatomic ions.
a. 3 protons and 2 total electrons
b. 35 protons and 8 valence electrons
c. 47 protons and 46 total electrons
d. 28 protons and 26 total electrons

3.11 Give the name of each ion.
a. F^- **b.** O^{2-} **c.** Cl^- **d.** Br^-

3.12 Give the name of each ion.
a. Sr^{2+} **c.** S^{2-}
b. Cl^- **d.** Se^{2-}

3.13 Give a name of each ion.
a. Co^{2+} **b.** Pb^{2+} **c.** Cr^{3+} **d.** Cu^+

3.14 Give a name of each ion.
a. Co^{3+} **b.** Pb^{4+} **c.** Cr^{2+} **d.** Cu^{2+}

3.15 Give the name of each ion.
a. $CO_3{}^{2-}$ **c.** $SO_3{}^{2-}$
b. $NO_3{}^-$ **d.** $CH_3CO_2{}^-$

3.16 Give the name of each ion.
a. H_3O^+ **c.** $HPO_4{}^{2-}$
b. OH^- **d.** $H_2PO_4{}^-$

3.17 Write the formula of each ion.
a. hydrogen carbonate ion
b. nitrite ion
c. sulfate ion

3.18 Write the formula of each ion.
a. hydrogen sulfate ion
b. phosphate ion
c. dichromate ion

3.19 Draw the electron dot structure of a hydrogen cation.

3.20 Hydride ion has a 1− charge. Draw the electron dot structure of a hydride ion.

3.2 The Octet Rule

3.21 How many valence electrons must each atom gain to reach an octet?
a. O **b.** F **c.** Se **d.** Br

3.22 How many valence electrons must each atom lose to reach an octet?
a. Li **b.** Al **c.** Cs **d.** Sr

3.23 Draw the electron dot structure for the ion expected to be formed from each of the atoms in Problem 3.21. Be sure to indicate the sign (positive or negative) and magnitude of the charge on the ion.

3.24 Draw the electron dot structure for the ion expected to be formed from each of the atoms in Problem 3.22. Be sure to indicate the sign (positive or negative) and magnitude of the charge on the ion.

3.25 When a nitrogen atom is converted into an ion,
a. what is the name of the ion?
b. how many electrons does nitrogen gain?
c. the ion ends up with what charge?

3.26 When a potassium atom is converted into an ion,
a. what is the name of the ion?
b. how many electrons does potassium lose?
c. the ion ends up with what charge?

3.27 Draw the electron dot structure of each atom and of the ion that it is expected to form.
a. Na **b.** Cl **c.** Ar

3.28 Draw the electron dot structure of each atom and of the ion that it is expected to form.
a. K **b.** I **c.** Ca

3.3 Ionic Compounds

3.29 Name each ionic compound.
a. MgO **c.** CaF_2
b. Na_2SO_4 **d.** Na_2S

3.30 Name each ionic compound.
a. Li_2O **c.** Al_2O_3
b. $BaCl_2$ **d.** MgSe

3.31 Name each ionic compound.
a. $FeCl_2$ **b.** CoS **c.** $CoCl_3$ **d.** Al_2S_3

3.32 Name each ionic compound.
a. Fe_2O_3 **b.** $CrBr_3$ **c.** $CrBr_2$ **d.** PbI_2

3.33 Name each ionic compound.
a. Na_2SO_4 **c.** K_3PO_4
b. $Ca(NO_3)_2$ **d.** NH_4OH

3.34 Name each ionic compound.
a. K_2SO_3 **c.** K_2HPO_4
b. $Sr(NO_2)_2$ **d.** $(NH_4)_2CO_3$

3.35 Name each ionic compound.
a. $Mg(HSO_4)_2$ **c.** $MgSO_3$
b. $Na_2Cr_2O_7$ **d.** $Ca(HCO_3)_2$

3.36 Name each ionic compound.
a. $Al_2(CO_3)_3$ **c.** $BaSO_3$
b. NH_4CN **d.** CH_3CO_2Na

3.37 Write the formula of each ionic compound.
a. calcium hydrogen phosphate
b. copper(II) bromide
c. copper(II) sulfate
d. sodium hydrogen sulfate

3.38 Write the formula of each ionic compound.
a. iron(III) hydroxide
b. ammonium bromide
c. copper(I) sulfate
d. tin(II) oxide

3.39 Give the name of each ionic compound.
a. Li_2SO_4 (an antidepressant)
b. $Ca(H_2PO_4)_2$ (used in foods as a mineral supplement)
c. $BaCO_3$ (used as a rat poison)

3.40 Give the name of each ionic compound.
a. $AlPO_4$ (used in some dental cements)
b. $MgHPO_4$ (a laxative)
c. $SrBr_2$ (an anticonvulsant)

3.41 Write the formula of the ionic compound that forms between each pair.
a. magnesium ions and fluoride ions
b. potassium ions and bromide ions
c. potassium ions and sulfide ions
d. aluminum ions and sulfide ions

3.42 Write the formula of the ionic compound that forms between each pair.
a. copper(I) ions and chloride ions
b. copper(II) ions and hydrogen sulfate ions
c. lithium ions and iodide ions
d. cobalt(III) ions and phosphate ions

3.43 In addition to sugars, citric acid, and other ingredients, the sports drink called Powerade contains potassium phosphate and potassium dihydrogen phosphate. Write the formula of each of these ionic compounds.

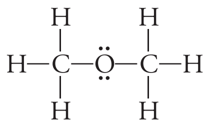
3.44 To fight tooth decay, some toothpastes contain stannous fluoride, also known as tin(II) fluoride. Write the formula of this compound.

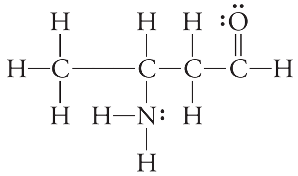
3.4 Covalent Bonds

3.45 Predict the number of covalent bonds formed by each nonmetal atom.
a. N **b.** Cl **c.** P

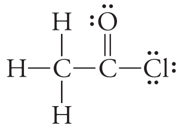
3.46 Predict the number of covalent bonds formed by each nonmetal atom.
a. C **b.** O **c.** Br

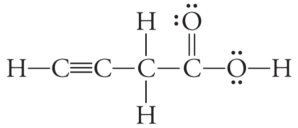
3.47 Draw the electron dot structure of each molecule.

a. 

b. 

3.48 Draw the electron dot structure of each molecule.

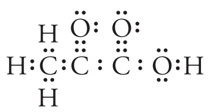
a. 

b. 

3.49 On the structure shown in Problem 3.48a, point out a pair of bonding electrons.

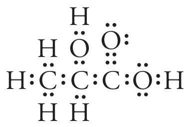
3.50 On the structure shown in Problem 3.48a, point out a pair of nonbonding electrons.

3.51 Draw the Lewis structure of pyruvic acid, a compound formed during the breakdown of sugars by the body.



Pyruvic acid

3.52 Draw the Lewis structure of lactic acid, a compound formed during anaerobic exercise.



Lactic acid

3.5 Molecules

3.53 Name each binary molecule.
a. NCl_3 **b.** PCl_3 **c.** PCl_5

3.54 Name each binary molecule.
a. CS_2 **b.** N_2O_3 **c.** NF_3

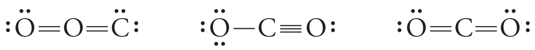
3.55 Phosphine (PH_3) is a poisonous gas that has the odor of decaying fish. Give another name for this binary molecule.

3.56 Dentists use nitrous oxide (N_2O) as an anesthetic. Give another name for this binary molecule.

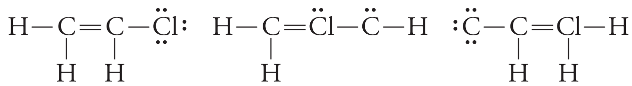
3.57 Draw the electron dot structure of the molecule formed when sufficient H atoms are added to give each atom an octet of valence electrons.
a. C **b.** N **c.** O

3.58 Draw the electron dot structure of the molecule formed when sufficient H atoms are added to give each atom an octet of valence electrons.
a. F **b.** P **c.** Br

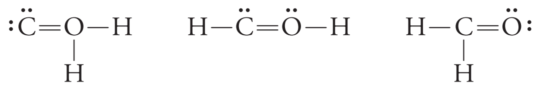
3.59 Which structure best fits the formula CO_2?



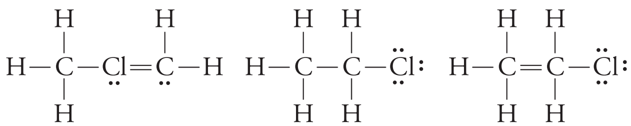
3.60 Which structure best fits the formula C_2H_3Cl?



3.61 Which structure best fits the formula CH_2O?



3.62 Which structure best fits the formula C_2H_5Cl?



3.63 Indicate whether each is an ionic compound or a binary molecule.
a. $BaCl_2$ **d.** HgO
b. OCl_2 **e.** N_2O_3
c. CS_2 **f.** Cu_2O

3.64 Indicate whether each is an ionic compound or a binary molecule.
a. SF_2 **d.** PF_5
b. MgF_2 **e.** NO_2
c. $SnCl_2$ **f.** SnO_2

3.65 Name each of the ionic compounds or binary molecules in Problem 3.63.

3.66 Name each of the ionic compounds or binary molecules Problem 3.64.

3.67 Are all diatomic molecules compounds? Explain.

3.68 Do any elements exist as binary molecules? Explain.

3.6 Formula Weight, Molecular Weight, and Molar Mass

3.69 **a.** What is the formula weight of ammonium hydroxide?
b. What is the mass of 0.950 mol of ammonium hydroxide?
c. How many moles of ammonium hydroxide are present in 0.475 g?

3.70 **a.** What is the formula weight of Li_2CO_3?
b. What is the mass of 1.33×10^{-4} mol of Li_2CO_3?
c. How many moles of CO_3^{2-} ions are present in 73.5 g of Li_2CO_3?

***3.71** **a.** What is the formula weight of sodium oxide?
b. How many oxide ions are present in 0.25 mol of sodium oxide?
c. How many sodium ions are present in 0.25 mol of sodium oxide?
d. How many oxide ions are present in 2.30 g of sodium oxide?
e. How many sodium ions are present in 2.30 g of sodium oxide?

***3.72** **a.** What is the formula weight of magnesium iodide?
b. How many magnesium ions are present in 7.5×10^{-6} mol of magnesium iodide?
c. How many iodide ions are present in 7.5×10^{-6} mol of magnesium iodide?
d. How many magnesium ions are present in 4.5 mg of magnesium iodide?
e. How many iodide ions are present in 4.5 mg of magnesium iodide?

***3.73** The food additive potassium sorbate, $K(C_6H_7O_2)$, is a mold and yeast inhibitor.
a. What is the charge on the sorbate ion?
b. How many C atoms are present in 0.0150 g of potassium sorbate?

***3.74** The food additive sodium benzoate, $C_6H_5CO_2Na$, extends the shelf life of food by preventing bacterial growth.
a. What is the charge on benzoate ion?
b. How many C atoms are present in 0.250 g of sodium benzoate?

***3.75** **a.** What is the molecular weight of CCl_4?
b. What is the mass of 61.3 mol of CCl_4?
c. How many moles of CCl_4 are present in 0.465 g of CCl_4?
d. How many molecules of CCl_4 are present in 5.50×10^{-3} g of CCl_4?

*3.76 a. What is the molecular weight of aspirin ($C_9H_8O_4$)?
b. What is the mass of 0.00225 mol of aspirin?
c. How many moles of aspirin are present in 500 mg of aspirin?
d. How many molecules of aspirin are present in 1.00 g of aspirin?

*3.77 One tablet of a particular analgesic contains 250 mg of acetaminophen ($C_8H_9NO_2$). How many acetaminophen molecules are contained in the tablet?

*3.78 A vitamin tablet contains 500 mg of vitamin C ($C_6H_8O_6$). How many vitamin C molecules are contained in the tablet?

*3.79 Isovaleric acid ($C_5H_{10}O_2$) is the molecule responsible for foot odor. To smell this compound, it must be present in the air at a minimum of 250 parts per billion. This means that a 0.50 L volume of air would hold 0.00013 g of isovaleric acid.
a. To how many moles of isovaleric acid does this correspond?
b. To how many molecules of isovaleric acid does this correspond?

*3.80 2-Isobutyl-3-methoxypyriazine, $C_9H_{14}N_2O$, has an "earthy" smell and is one of the compounds responsible for the aroma of coffee.
a. If 1.00 L of brewed coffee contains 8.3×10^{-2} mg of 2-isobutyl-3-methoxypyriazine, how many molecules of this compound are present?
b. How many molecules of the compound would be present in 1 cup of this coffee?

BiochemistryLink | IONOPHORES AND BIOLOGICAL ION TRANSPORT

3.81 Explain how ionophores act as antibiotics.

3.82 K^+ attaches more strongly to valinomycin than Na^+ because of the size of the cavity (binding site) in the center of this compound. Is this cavity too large for Na^+ or is it too small?

HealthLink | PASS THE SALT, PLEASE

3.83 Name the compounds that, in addition to NaCl, are present in Hawaiian sea salt and Black Indian salt.

3.84 What is iodized salt and why is it sold?

HealthLink | DENTAL FILLINGS

3.85 Sometimes a filling can make a tooth sensitive to temperature changes. Given that amalgam is made from metals and that composites are made from nonmetals, which type of filling would you expect is more likely to cause thermal sensitivity?

3.86 Although amalgam is sometimes referred to as a compound, it does not fit the definition of this term. Explain.

HealthLink | NITRIC OXIDE

3.87 Viagra, one of the drugs used to treat erectile dysfunction, works by enhancing the effects of nitric oxide produced by the body. Men who take nitroglycerin for angina are advised against taking Viagra. Why?

3.88 Why does the risk of heart attack and stroke increase when a person is given a blood transfusion?

Learning Group Problems

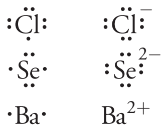
3.89 Some ionic compounds can form hydrates, in which water molecules are incorporated into their crystal structures. One example is iron(II) nitrate hexahydrate, $Fe(NO_3)_2 \cdot 6H_2O$. The dot in the formula indicates that there are 6 water molecules associated with each formula unit of iron(II) nitrate. Because water molecules are neutral, they contribute no charges to the compound.
a. What is the charge on the cation in this compound?
b. Give an alternate name for the cation.
c. What is the charge on the anion in this compound?
d. Does the compound contain any ionic bonds? Explain.
e. Does the compound contain any covalent bonds? Explain.
f. Give an alternate name for water, using the rules for naming binary compounds.
g. What is the formula weight of iron(II) nitrate hexahydrate?
h. How many moles is 4.93 g of this compound?
i. How many grams is 0.639 mol of this compound?
j. How many water molecules are in 8.58×10^{-6} mol of this compound?
k. How many iron(II) ions are in 43.8 g of this compound?

3.90 a. Name the hydrate $Cu_3(PO_4)_2 \cdot 3H_2O$. (Refer to Problem 3.89.)
b. What is the charge on the cation in this compound?
c. Give a different name for the cation than used in your answer to part a of this question.
d. What is the charge on the anion in this compound?
e. What is the formula weight of $Cu_3(PO_4)_2 \cdot 3H_2O$?
f. How many moles is 1.22 g of $Cu_3(PO_4)_2 \cdot 3H_2O$?
g. How many grams is 5.73 mol of $Cu_3(PO_4)_2 \cdot 3H_2O$?
h. How many water molecules are in 6.39×10^{-2} mol of $Cu_3(PO_4)_2 \cdot 3H_2O$?
i. How many phosphate ions are in 0.227 g of $Cu_3(PO_4)_2 \cdot 3H_2O$?

SOLUTIONS TO PRACTICE PROBLEMS

3.1 **a.** scandium(III) ion; **b.** titanium(II) ion; **c.** strontium ion

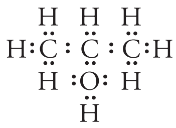
3.2 **a.** **b.** **c.**



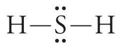
3.3 **a.** LiCl; **b.** Li_2O; **c.** Li_3N

3.4 **a.** NH_4HSO_4; **b.** $(NH_4)_3PO_4$; **c.** $Sr_3(PO_4)_2$; **d.** $SrSO_4$

3.5



3.6



3.7 **a.** silicon tetrabromide; **b.** diphosphorus trioxide; **c.** tetraphosphorus triselenide

3.8 **a.** 84.01 amu; **b.** 26.5 g

3.9 **a.** 75.08 amu; **b.** 300. g; **c.** 4.43×10^{19} molecules