

Chemical Bonds and Compounds



Figure 5.1 (a) Allied soldiers captured by the Japanese, February 1942. (continued)

Lithium Carbonate and Bipolar Disorder

Sometimes, combinations produce surprising results. Consider the story of lithium carbonate — a simple compound used worldwide to treat bipolar disorder — and the combination of bad luck, good luck, and keen observation that led to its discovery.

In February 1942, during the heart of World War II, Japanese forces attacked the Allied stronghold at Singapore. After a week of fighting, the Allies surrendered (**Figure 5.1**). Among those captured was John Cade, a young psychiatrist serving in the Australian Army Medical Corps. He spent the next three years in a prisoner-of-war camp.

While imprisoned, Cade observed a number of prisoners who suffered from bipolar disorder: They fluctuated between wildly aggressive behavior (called the manic phase) and deep depression. Cade began to suspect that a toxic chemical caused the prisoners' erratic behavior and that their moods stabilized after the toxin was expelled through their urine.

After his release at the end of the war, Cade returned to Australia and resumed his career in psychiatry. On the side, he began exploring the ideas he had developed in captivity. He collected urine samples from bipolar patients and injected the urine into guinea pigs. Interestingly, the guinea pigs treated with urine from bipolar patients died faster than those treated with urine from healthy people. Cade delved deeper. He suspected that a compound called uric acid might be the mysterious toxin. He began to study the effects of pure uric acid and related compounds on the guinea pigs. He found that one such compound, lithium urate, reduced the toxic effects of the other compounds present. Intrigued by this result, he decided to test a simpler lithium-containing compound: lithium carbonate. When he injected guinea pigs with pure lithium carbonate, the animals became sedate.

Ultimately Cade's ideas about toxins in the urine were discarded, but the effects he observed from lithium carbonate opened a new door. Cade wondered if lithium carbonate would also sedate patients suffering from the manic phase of bipolar disorder. To see if it was safe, he first tested it on himself. Finding no long-term effects, he treated the manic

CHAPTER FIVE

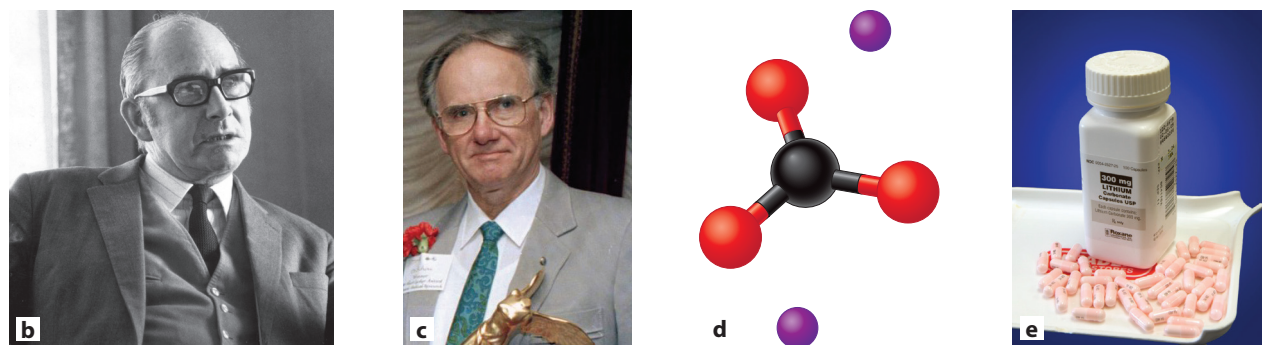



Figure 5.1 (continued) (b) John Cade discovered that lithium carbonate could treat bipolar disorder. (c) Mogens Schou carried on Cade's work, extensively studying and promoting the effects of lithium carbonate. (d) Lithium carbonate is made from lithium, carbon, and oxygen atoms. (e) Today millions of people take lithium carbonate for the treatment of bipolar disorder. © TopFoto/The Image Works; Newspix/Getty Images; AP Photo/Marty Lederhandler; Charles D. Winters/Science Source

patients in his ward with lithium carbonate. This human testing was remarkably successful, and in 1949 he published his results. In the decades that followed, another psychiatrist, Mogens Schou, extensively studied the effects of lithium carbonate. Today, lithium carbonate remains one of the most common and least expensive treatments for bipolar disorder.

Like most science — and most other human activities — this story is messy. Cade's ideas were conceived in the harshest of circumstances. His initial ideas were incorrect. And by today's standards, Cade's experiments seem reckless. But his careful observations, both in the prison camp and the laboratory, led him to insights that changed the way we treat mental illness. Out of all the messy pieces, a beautiful discovery emerged.

In this chapter, we'll begin to study how chemical bonds bring atoms together to form compounds. Just as a complete story can be much different from the pieces that comprise it, compounds behave much differently from the elements they are composed of. As atoms combine to form compounds, new and intriguing properties emerge — properties that create new materials, new medicines, and new opportunities. 

→ Intended Learning Outcomes

After completing this chapter and working the practice problems, you should be able to:

5.1 Lewis Symbols and the Octet Rule

- Use the periodic table to identify the number of valence electrons in an atom.
- Represent valence electrons using Lewis dot symbols.

5.2 Ions

- Describe and predict the formation of main-group ions using the octet rule.
- Identify common monatomic and polyatomic ions by name, symbol or formula, and charge.

5.3 Ionic Bonds and Compounds

- Predict ionic formulas based on cation and anion charges.
- Broadly describe the arrangement of ions in an ionic solid.
- Convert between the name and formula for an ionic compound.

5.4 Covalent Bonding

- Describe how nonmetals fulfill the octet rule through covalent bonds.
- Differentiate between empirical and molecular formulas.
- Name binary covalent compounds.

5.5 Distinguishing Ionic and Covalent Compounds

- Distinguish ionic and covalent compounds based on their chemical formulas.

5.6 Aqueous Solutions: How Ionic and Covalent Compounds Differ

- Contrast the behavior of ionic compounds and covalent compounds in aqueous solutions.

5.7 Acids — An Introduction

- Describe the ionization of acids in aqueous solution.
- Name binary acids and oxyacids.

5.1 Lewis Symbols and the Octet Rule

In Chapter 4, we saw that families of atoms such as the alkali metals, the halogens, and the noble gases exhibit similar behaviors because they have similar electronic configurations. In this chapter, we will explore how these electronic configurations lead to the formation of chemical bonds.

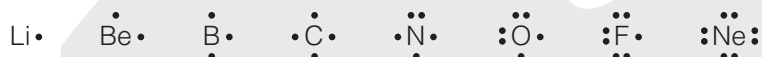
Chemical bonding involves changes in an atom's outer or *valence electrons*. Recall that valence electrons are the electrons in the highest-occupied energy level of an atom. Because of the sublevel filling sequence, the valence level involves only the *s* and *p* sublevels. Since two electrons can fit in an *s* sublevel, and six electrons can fit in the *p* sublevel, up to eight electrons can occupy the valence level. For main-group elements, we can quickly determine the number of valence electrons from the periodic table: The column (group) number for the main groups is also the number of valence electrons (**Figure 5.2**). For example, nitrogen (N) is in group 5A, so it has five valence electrons. Neon (Ne) is in group 8A, so it has eight valence electrons.

Group	1A	2A		3A	4A	5A	6A	7A	8A
Valence electrons	1	2		3	4	5	6	7	8
Configuration	s^1	s^2		s^2p^1	s^2p^2	s^2p^3	s^2p^4	s^2p^5	s^2p^6
	H								He
	Li	Be		B	C	N	O	F	Ne
	Na	Mg		Al	Si	P	S	Cl	Ar

The valence level holds up to eight electrons. ■

Figure 5.2 The main-group numbers (1A–8A) also indicate the number of electrons in each atom's valence level.

To visualize chemical bonding, it is often helpful to draw **Lewis dot symbols**. These symbols represent the number of valence electrons in an atom as dots drawn around the atomic symbol. Here are the Lewis symbols for each of the row 2 elements:



In Chapter 4 we introduced the *octet rule*, which states that *an atom is stabilized by having its valence energy level filled*. For elements in row 2 and below, eight electrons are required to fill the valence level. The octet rule explains why the noble gases are so stable, and it also allows us to predict how main-group elements form chemical bonds. These elements fulfill the octet rule by gaining or losing electrons to form *ions*, or by sharing electrons between two atoms. We will explore these behaviors in the sections that follow.

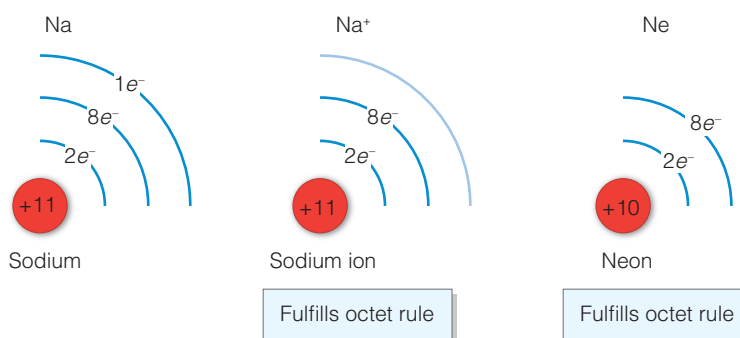
Main-group elements can fulfill the octet rule by gaining, losing, or sharing electrons. ■

5.2 Ions

Cations: Ions with a Positive Charge

Main-group metals fulfill the octet rule by losing electrons to form positively charged ions, called **cations** (pronounced *cat-eye-uns*). For example, consider sodium metal: Sodium has an electron configuration of $1s^2 2s^2 2p^6 3s^1$. Because $1s^2 2s^2 2p^6$ is the same configuration as neon, we often write it as $[\text{Ne}]3s^1$. To fill its valence level (level 3), sodium would have to gain seven electrons—an unlikely occurrence. However, by losing just one electron, sodium becomes electronically identical to neon—a very stable arrangement that fulfills the octet rule. As a result,

Na^+ , Mg^{2+} , and Ne are all *isoelectronic*—meaning they have the same electron configuration.



Alkali metals form +1 ions.

Alkaline earth metals form +2 ions. ■

Figure 5.3 Sodium has one valence electron. By losing its outermost electron, sodium becomes electronically identical to the noble gas neon and fulfills the octet rule.

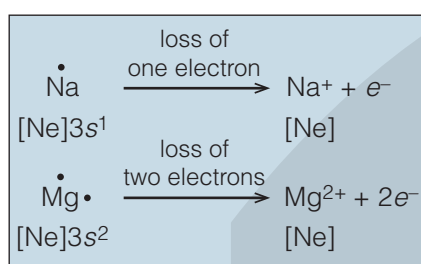


Figure 5.4 Sodium and magnesium both lose their valence electrons to become isoelectronic with neon.

sodium easily loses one electron to form Na^+ , a common ion (**Figure 5.3**). All of the alkali metals (metals in group 1A of the periodic table) lose one electron to form +1 ions.

Magnesium has an electron configuration of $[\text{Ne}]3s^2$. Just as sodium lost one electron to reach the $[\text{Ne}]$ electron configuration, magnesium can reach this electron configuration by losing two electrons (**Figure 5.4**). Because it loses two electrons, it has a charge of +2, which is written as Mg^{2+} . Each of the alkaline earth metals (group 2A) loses two electrons to form +2 ions (**Figure 5.5**).

The *transition metals* (elements in the *d* block of the periodic table) also tend to lose electrons to form positively charged ions. But unlike main-group metals, transition metal ions do not follow a simple pattern. Transition metals typically form ions having a charge between +1 and +4, and some transition metals form multiple charged ions. Metals in the lower part of the *p* block also behave this way (**Figure 5.6**).

Transition metal ions may have multiple charges. ■

Naming Cations

In general, metal cations are given the same name as the neutral metal. For example, the cation produced from sodium metal is simply called the *sodium ion*.

As mentioned earlier, some metals can have more than one charge. For example, iron commonly forms both +2 and +3 ions. Historically, these two ions were named as *ferrous* and *ferric* ions, respectively. Similarly, copper commonly forms

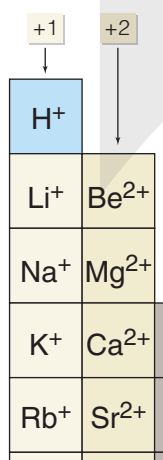


Figure 5.5 The group 1A elements (hydrogen and the alkali metals) form +1 ions. Group 2A elements (the alkaline earth metals) form +2 ions.

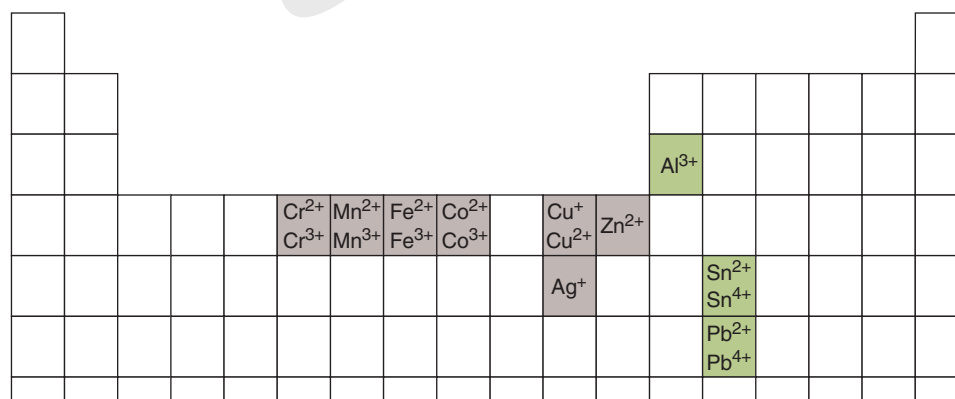


Figure 5.6 Many common transition and *p*-block metals form ions with multiple charges.

TABLE 5.1 Naming Ions with More Than One Charge

Atom	Ion	Older Name	Modern Name
Iron	Fe^{2+}	Ferrous	Iron(II)
	Fe^{3+}	Ferric	Iron(III)
Copper	Cu^+	Cuprous	Copper(I)
	Cu^{2+}	Cupric	Copper(II)

both +1 (*cuprous*) and +2 (*cupric*) ions. Although you will encounter these names occasionally, the modern style of naming these ions puts the charge in Roman numerals within parentheses immediately after the atom name. For example, the ferrous ion (Fe^{2+}) is named as iron(II), which is read as “iron-two”; the ferric ion (Fe^{3+}) is named as iron(III), read as “iron-three” (Table 5.1).

If an atom can form more than one cation, use Roman numerals after the atom name to specify the charge. ■

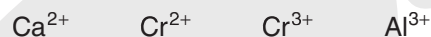
Example 5.1 Naming Cations

Name each of the following ions: Ag^+ , Pb^{2+} , and Pb^{4+} .

From Figure 5.6, we see that silver (Ag) forms only one ion. Therefore, we refer to Ag^+ as a *silver* ion. However, lead (Pb) forms two different ions. To distinguish them, we refer to Pb^{2+} as a lead(II) ion and Pb^{4+} as a lead(IV) ion.

TRY IT

1. Provide names for each of these cations:



Check it

Watch explanation

Anions: Ions with a Negative Charge

The nonmetals lie on the right-hand side of the periodic table. Unlike metals, the valence shells of most nonmetals are nearly full. To fulfill the octet rule, most nonmetals gain electrons to form negatively charged ions, called **anions** (pronounced *an-eye-uns*).

For example, fluorine has an electron configuration of $1s^2 2s^2 2p^5$. By gaining one electron, fluorine can achieve an electron configuration of $1s^2 2s^2 2p^6$, the same electron configuration as neon. This configuration fulfills the octet rule and provides tremendous stability. As a result, fluorine tends to aggressively “grab” an electron, forming a very stable ion with a charge of -1 (Figure 5.7). The other halogens (chlorine, bromine, iodine) also form -1 ions.

Nonmetals gain electrons to form anions. ■

Halogens form -1 ions. ■

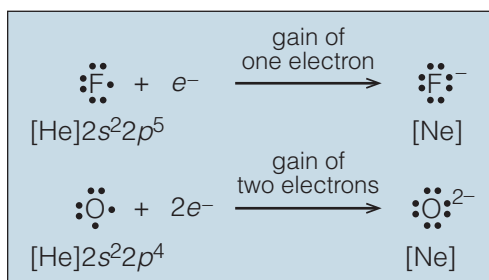


Figure 5.7 Fluorine and oxygen both gain electrons to become isoelectronic with neon.

Chalcogens form -2 ions. ■

	5A	6A	7A	8A
	-3	-2	-1	
	N ³⁻	O ²⁻	F ⁻	
	P ³⁻	S ²⁻	Cl ⁻	
			Br ⁻	
			I ⁻	

Figure 5.8 The halogens form -1 ions. Oxygen and sulfur form -2 ions. Nitrogen and phosphorus form -3 ions. The noble gases (shaded violet) have complete valence shells and do not form ions.

Oxygen, sulfur, and the atoms below them on the periodic table comprise a family called the *chalcogens* (group 6A). Each of these elements is two electrons short of a noble gas configuration. For example, oxygen has an electron configuration of $1s^2 2s^2 2p^4$. It needs two electrons to fill its outer valence level, and so it tends to gain two electrons, resulting in a charge of -2 . Sulfur also forms a stable ion with a charge of -2 .

What about group 5A elements, such as nitrogen and phosphorus? Consistent with the pattern just described, these atoms gain three electrons to fill their valence level and so form ions with a charge of -3 (**Figure 5.8**).

Naming Anions

When an atom gains electrons, we name the resulting anion by changing the end of the atom name to *-ide*. For example, chlorine atoms form *chloride* ions, oxygen atoms form *oxide* ions, and sulfur atoms form *sulfide* ions. A list of common anions is given in **Table 5.2**.

TABLE 5.2 Common Anions

Atom	Anion Symbol	Anion Name
Nitrogen	N ³⁻	Nitride
Phosphorus	P ³⁻	Phosphide
Oxygen	O ²⁻	Oxide
Sulfur	S ²⁻	Sulfide
Fluorine	F ⁻	Fluoride
Chlorine	Cl ⁻	Chloride
Bromine	Br ⁻	Bromide



Sports drinks contain ions that are commonly lost during exercise. They include sodium, potassium, chloride, and phosphate.

Example 5.2 Naming Ions and Predicting Charges

Predict the ions that would be formed from an atom of calcium and from an atom of sulfur. Name each ion.

Calcium belongs to the alkaline earth metal family. It has an electron configuration of $[\text{Ar}]4s^2$. Calcium loses its two valence electrons, resulting in a charge of $+2$. Cations are given the same name as the parent atom, so we refer to Ca^{2+} as the *calcium ion*.

Sulfur is a nonmetal with an electron configuration of $[\text{Ne}]3s^2 3p^4$. To fill its valence shell, sulfur gains two electrons, giving the ion a charge of -2 . We refer to S^{2-} as the *sulfide ion*.



Check it

Watch explanation

TRY IT

2. Use the periodic table to predict whether each atom would gain or lose electrons, and write the charge on the ion formed:

Cl Br O Be K

Polyatomic Ions

Polyatomic ions are groups of atoms that have an overall charge. Many of these ions, such as acetate and phosphate, are essential to life and common in many different materials and applications. Formulas and names for the most common

TABLE 5.3 Common Polyatomic Ions

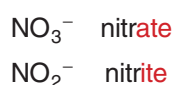
Formula	Name	Formula	Name
NH_4^+	Ammonium		
NO_3^-	Nitrate	SO_4^{2-}	Sulfate
CO_3^{2-}	Carbonate	SO_3^{2-}	Sulfite
HCO_3^-	Bicarbonate (also called hydrogen carbonate)	HSO_4^-	Bisulfate (also called hydrogen sulfate)
NO_2^-	Nitrite	ClO_4^-	Perchlorate
PO_4^{3-}	Phosphate	ClO_3^-	Chlorate
HPO_4^{2-}	Hydrogen phosphate	ClO_2^-	Chlorite
$\text{C}_2\text{H}_3\text{O}_2^-$	Acetate	ClO^-	Hypochlorite
OH^-	Hydroxide	CrO_4^{2-}	Chromate
CN^-	Cyanide	$\text{Cr}_2\text{O}_7^{2-}$	Dichromate
O_2^{2-}	Peroxide	MnO_4^-	Permanganate

polyatomic ions are given in **Table 5.3**. Notice that this table contains only one common polyatomic cation (ammonium). All others are anions.

Naming Polyatomic Ions

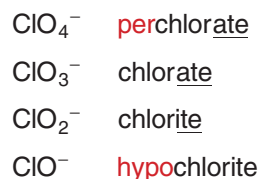
Although Table 5.3 contains many ion names, there are patterns that will help you keep these names organized. Notice that most of the polyatomic ions contain oxygen—these are called **oxyanions**. We name oxyanions by adding the suffix *-ate* to the root of the element. For example, the oxyanion from carbon (CO_3^{2-}) is *carbonate*, and the oxyanion formed from phosphorus (PO_4^{3-}) is *phosphate*.

Some elements form more than one oxyanion. In these cases we use the suffix *-ate* to indicate the ion with more oxygen atoms present, and the suffix *-ite* to indicate the ion with fewer oxygen atoms present. For example, there are two common nitrogen oxyanions:



“-ate is great, and -ite is lite”
 More oxygen atoms: *-ate*
 Fewer oxygen atoms: *-ite*

Chlorine forms four oxyanions. In this case, we use the prefix *per-* (meaning “more than”) to indicate the largest number of oxygen atoms, and the prefix *hypo-* (meaning “below”) to indicate the least number of oxygen atoms:



A Summary of the Common Ions

As you continue studying chemistry, you will find it essential to know the structure, formula, and charge of common monatomic and polyatomic ions. **Figure 5.9** summarizes the most common ions. You should be very familiar with these ions, because you will use them regularly throughout this course.

Monatomic atoms															
H ⁺															
Li ⁺	Be ²⁺												N ³⁻	O ²⁻	F ⁻
Na ⁺	Mg ²⁺											Al ³⁺	P ³⁻	S ²⁻	Cl ⁻
K ⁺	Ca ²⁺				Cr ²⁺	Mn ²⁺	Fe ²⁺	Co ²⁺		Cu ⁺	Zn ²⁺				Br ⁻
					Cr ³⁺	Mn ³⁺	Fe ³⁺	Co ³⁺		Cu ²⁺					
Rb ⁺	Sr ²⁺									Ag ⁺					
													Sn ²⁺		I ⁻
													Sn ⁴⁺		
													Pb ²⁺		
													Pb ⁴⁺		

Polyatomic atoms			
NH ₄ ⁺ Ammonium			
NO ₃ ⁻	Nitrate	SO ₄ ²⁻	Sulfate
CO ₃ ²⁻	Carbonate	SO ₃ ²⁻	Sulfite
HCO ₃ ⁻	Bicarbonate (Hydrogen carbonate)	HSO ₄ ⁻	Bisulfate (Hydrogen sulfate)
NO ₂ ⁻	Nitrite	ClO ₄ ⁻	Perchlorate
PO ₄ ³⁻	Phosphate	ClO ₃ ⁻	Chlorate
HPO ₄ ²⁻	Hydrogen phosphate	ClO ₂ ⁻	Chlorite
C ₂ H ₃ O ₂ ⁻	Acetate	ClO ⁻	Hypochlorite
OH ⁻	Hydroxide	CrO ₄ ²⁻	Chromate
CN ⁻	Cyanide	Cr ₂ O ₇ ²⁻	Dichromate
O ₂ ²⁻	Peroxide	MnO ₄ ⁻	Permanganate

**Practice****Common Ions**

How well do you know the common ions? Try this interactive game to practice and test your knowledge.

Figure 5.9 It is important to know the names, formulas, and charges for these common ions.

Example 5.3 Gathering Information from Ion Names

The four ions named below are less common and are not listed in Table 5.3. Which of these are polyatomic? Identify each one as a cation or an anion.

- a. bromate b. bromite c. palladium(II) d. selenide

From the suffixes *-ate* and *-ite*, we know that both bromate and bromite are oxyanions of the element bromine. Further, we know that bromate contains more oxygen atoms than bromite. The actual formula for bromate is BrO₃⁻, and the formula for bromite is BrO₂⁻.

Palladium is a transition metal, so it forms a cation. The (II) indicates that this ion is Pd²⁺.

Finally, the ending *-ide* indicates that selenide is a monatomic anion formed from the element selenium. Selenium lies just below sulfur on the periodic table, so we predict the charge of this ion to be -2.

**Check it****Watch explanation****TRY IT**

3. Write the symbol and the charge for each ion listed. Refer to the periodic table as needed.

calcium nitrate scandium(III) telluride

4. Name each of these ions:

Cs⁺ Fe²⁺ SO₄²⁻ As³⁻

5.3 Ionic Bonds and Compounds

Ionic Bonds and Ionic Lattices

Opposite charges attract each other. When positive and negative ions come near each other, they stick tightly together. The force of attraction between oppositely charged ions is called an **ionic bond**. A compound composed of oppositely charged ions is an **ionic compound**. Because metals form cations and nonmetals form anions, the compounds formed between metals and nonmetals are ionic compounds.

Ionic compounds contain many cations and anions, joined through ionic bonds. To understand how these ions fit together, let's consider the structure of a compound composed of sodium cations (Na^+) and chloride anions (Cl^-). A single Na^+ ion and a single Cl^- ion adhere to each other in an ionic bond. But what happens if additional ions are present? They pack together in a structure of alternating positive and negative charges that stretch out in three dimensions (**Figure 5.10**). This array of positive and negative ions is called an **ionic lattice**.

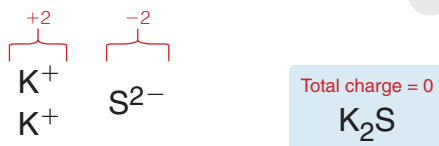
To represent the composition of compounds like this one, we use a **chemical formula** that indicates the type and amount of each element present. We represent ionic compounds using a specific type of chemical formula, called an **empirical formula**. An empirical formula gives the smallest whole-number ratio of atoms in a compound. Subscripts written after each atom indicate the number of that atom present. If no subscript is written, we understand the number of atoms to be one. In this instance, the formula is written simply as NaCl .

The empirical formula gives the smallest number of ions necessary to form a compound. This number of ions is called the **formula unit**. For example, the empirical formula for sodium chloride is NaCl ; a formula unit of sodium chloride contains one sodium ion and one chloride ion.

When writing empirical formulas for ionic compounds, we write the symbol or formula for the cation, followed by the anion. In the next section, we'll look at several more examples of empirical formulas and formula units.

Predicting Formulas for Ionic Compounds

Some ionic compounds contain cations and anions with different charges. For example, consider the solid composed of potassium and sulfide ions. The potassium cation has a charge of $+1$ while the sulfide anion has a charge of -2 . To form a neutral solid, *the positive charges must equal the negative charges*. To balance the charges, the solid must contain two potassium ions for every one sulfide ion:



We therefore write the empirical formula for this compound as K_2S . Put another way, a formula unit of potassium sulfide contains two potassium ions and one sulfide ion.

We can also predict the formulas for ionic solids containing polyatomic ions. For example, consider the ionic compound produced from calcium (Ca^{2+}) and nitrate (NO_3^-) ions: For the positive and negative charges to balance, there must be two nitrate ions for every one calcium ion (**Figure 5.11**). We could write this formula as CaN_2O_6 . But it is better to write the formula as $\text{Ca}(\text{NO}_3)_2$ because this shows that two nitrates are attached, rather than some other arrangement of nitrogen and oxygen. If we have more than one polyatomic ion in the formula, we write that ion inside parentheses to show that the entire unit is repeating.

Metal cations and nonmetal anions form ionic bonds. ■

An empirical formula gives the smallest whole-number ratio of atoms in a compound. ■

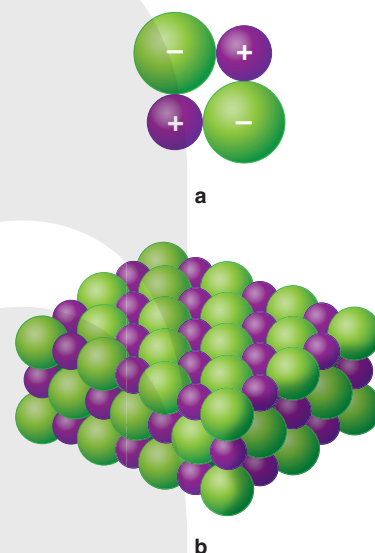


Figure 5.10 (a) Ions pack together in a framework of alternating positive and negative charges. (b) This packing results in a three-dimensional framework called an ionic lattice.

In an ionic compound, the total charge must equal zero. ■

**Practice****Balancing Charges**

To write an ionic compound formula correctly, you must balance the charges on the ions. Try this interactive to practice this skill.

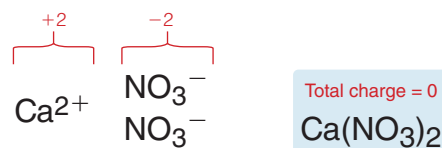


Figure 5.11 To balance the charges, this compound requires two nitrate ions for every one calcium ion.

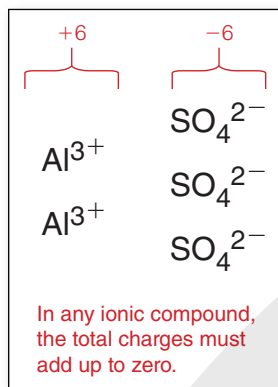


Figure 5.12 Three sulfate ions are required to balance the charge on two aluminum ions.

Example 5.4 Writing Formulas for Ionic Compounds

A compound is composed of two ions, aluminum and sulfate. What is the formula for this compound?

We know the aluminum ion has a charge of +3, and the sulfate ion has a charge of -2 (see Figure 5.9). For the charges of these ions to balance, we must have two aluminum ions for every three sulfate ions, as shown in **Figure 5.12**. Therefore, we write the formula for this compound as $\text{Al}_2(\text{SO}_4)_3$. As before, we put the repeating polyatomic ion in parentheses.

Example 5.5 Writing Formulas for Ionic Compounds

What is the formula for a compound composed of iron(III) and bromide ions?

Recall that iron is a transition metal, and it can have more than one possible charge. The name *iron(III)* indicates that this ion is Fe^{3+} . The bromide ion is Br^- . For the charges to balance, there must be three bromide ions for every one iron(III) ion. Therefore, we write this formula as FeBr_3 .

**Check it****Watch explanation****TRY IT**

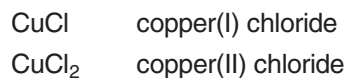
5. Predict the empirical formulas for compounds formed from these ions:

- | | |
|---------------------------|----------------------------|
| a. magnesium and chloride | b. potassium and phosphate |
| c. lead(II) and oxide | d. ammonium and carbonate |

Naming Ionic Compounds

To name an ionic compound, we give the cation name followed by the anion name. For example, NaCl is *sodium chloride*, and MgCl_2 is *magnesium chloride*. Because we know that a magnesium ion always has a +2 charge and a chloride ion has a -1 charge, there is no need to indicate the ratio of cations to anions. Given the name of the cation and anion present, we can determine the empirical formula.

For transition metals with more than one possible charge, it is important to include the charge of the ion in parentheses with the name. For example, copper and chloride ions form two different compounds, CuCl and CuCl_2 . In CuCl , the copper ion must have a charge of +1 to balance the charge from the chloride ion. In CuCl_2 , the copper ion must have a charge of +2. Therefore, we name these compounds as follows:



Compounds containing polyatomic ions are named in the same way as those containing monatomic ions. For example, the ionic compound MgSO_4 consists of a monatomic cation (magnesium) and a polyatomic anion (sulfate). Therefore, the name of this compound is magnesium sulfate.

Table 5.4 summarizes the names, formulas, and uses of several common ionic compounds. It is important to be able to convert between the name and empirical formulas for ionic compounds. This process is further illustrated in the examples that follow.

TABLE 5.4 Common Ionic Compounds

Compound	Formula	Application
Sodium chloride	NaCl	Table salt
Sodium fluoride	NaF	Fluoride treatment
Sodium bicarbonate	NaHCO ₃	Baking soda
Calcium oxide	CaO	Cement mix
Lithium carbonate	Li ₂ CO ₃	Treatment of bipolar disorder
Ammonium nitrate	NH ₄ NO ₃	Fertilizer

Example 5.6 Naming Ionic Compounds

Name the following compound: $\text{Fe}(\text{NO}_2)_2$.

The keys to solving this problem are to identify the ions present and to know their charges. The anion in this formula is nitrite, which has a charge of -1 . Because two NO_2^- ions are present, the charge on the iron (Fe) cation must be $+2$. Fe^{2+} is named as iron(II), and so the total compound is iron(II) nitrite.

Example 5.7 Writing the Formula for an Ionic Compound

Write the empirical formula for ammonium sulfide.

We know that ammonium is NH_4^+ and that sulfide is S^{2-} . For the charges to balance, there must be two ammonium ions for each sulfide ion. To show this, we put the ammonium formula in parentheses with a two on the outside. Listing the cation first, we write the formula for this compound as $(\text{NH}_4)_2\text{S}$.

TRY IT

6. Name each of these compounds:

- a. RbCl b. CuBr_2 c. ZnCO_3 d. K_2SO_4

7. Write the empirical formula for each compound named:

- a. zinc sulfide b. iron(III) oxide c. ammonium phosphate

8. Titanium is a transition metal that can have multiple ionic charges. The titanium compound TiO_2 is commonly used as an additive in paints. In this compound, what is the charge on the cation? What is the name of this compound?



Check it
Watch explanation

5.4 Covalent Bonding

Nonmetal–Nonmetal Bonds

In the previous section, we saw how ionic bonds form between metal cations and nonmetal anions. When two nonmetal atoms come together, a different type of bond occurs, called a **covalent bond**. In a covalent bond, two electrons are shared between two atoms.

Covalent bonds form between nonmetal atoms. ■

For example, consider the bond that forms between two hydrogen atoms. Each hydrogen atom has one proton and one electron. To form a covalent bond, the two electrons “pair up” in the space between the two nuclei (**Figure 5.13**).

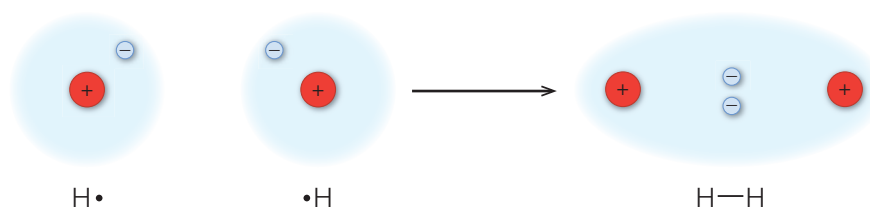


Figure 5.13 Two atoms form a covalent bond by sharing a pair of electrons.

A dash between two chemical symbols indicates a covalent bond. ■

The force of attraction between the nuclei and the two electrons holds the atoms together. We represent these shared electrons by drawing a dash between the symbols of the two atoms:



Molecules are held together by covalent bonds. ■

Remember that the first energy level holds only two electrons. By forming a covalent bond, each hydrogen atom completes its valence level.

When two hydrogen atoms combine, they form a *molecule*. In earlier chapters, we defined molecules as groups of atoms that bind together and behave as a unit. The bonds that hold molecules together are covalent bonds. In its elemental form, hydrogen is a gas composed entirely of these two-atom molecules (**Figure 5.14**).

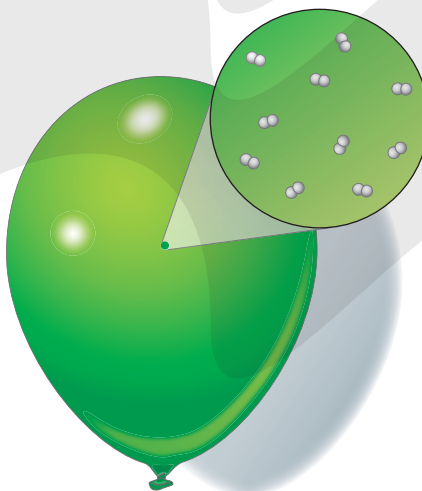
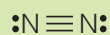
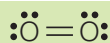
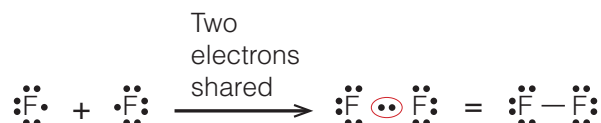


Figure 5.14 This balloon contains elemental hydrogen gas. The gas is composed of molecules containing two atoms each.



Atoms sometimes share two or even three pairs of electrons in covalent bonds. We represent double covalent bonds using two dashes between the atoms, and triple covalent bonds using three dashes. We will discuss covalent bonding in more detail in Chapter 9.

As a second example, let's look at the bonding that occurs between two fluorine atoms. Each fluorine atom contains seven valence electrons and therefore needs only one more electron to complete its valence level. Two fluorine atoms can form a single covalent bond. By forming this bond, the atoms fill their valence shell with eight electrons and satisfy the octet rule.

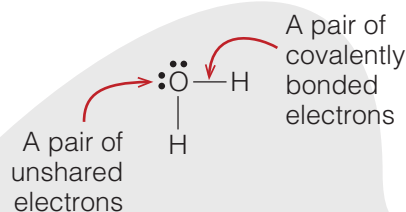


As in the hydrogen example, we use a dash to represent two shared electrons. We call this type of drawing a **Lewis structure**. Lewis structures depict the arrangement of valence electrons within a molecule or polyatomic ion. We will explore Lewis structures further in Chapter 9, when we take a more detailed look at the bonding and properties of molecules.

Hydrogen and fluorine are two of seven elements that exist as *diatomic* (“two-atom”) molecules in their elemental forms. The others are nitrogen, oxygen, and the rest of the halogens (**Figure 5.15**).

Covalent Compounds

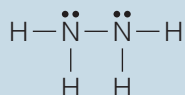
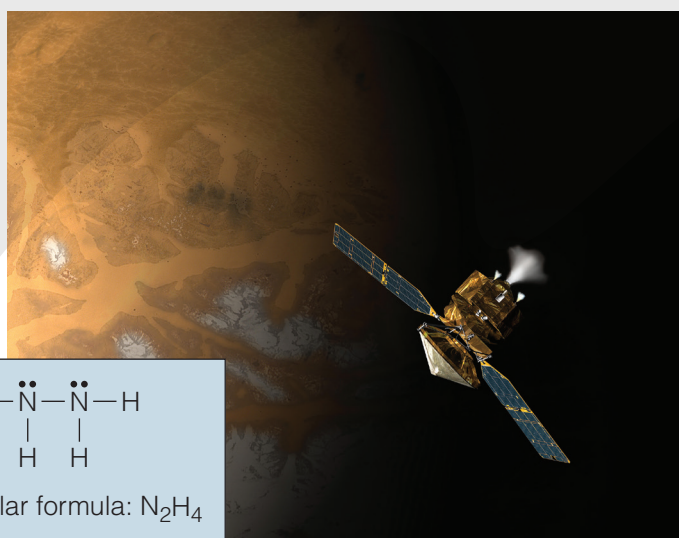
Covalent compounds form when different elements combine through covalent bonds, forming discrete molecules. Water is an example of a covalent compound. In a water molecule, an oxygen atom covalently bonds to two hydrogen atoms:



The valence level of each hydrogen atom is filled with two electrons. What about the oxygen atom? It has four unshared electrons and two covalent bonds. Between the unshared and the shared electrons, the oxygen atom has eight electrons in its valence level and fulfills the octet rule.

To describe covalent compounds, we often use **molecular formulas**. This type of chemical formula gives the actual number of atoms in the molecule rather than the simplest whole-number ratio.

For example, consider hydrazine, a fuel used for rocket thrusters (**Figure 5.16**). A hydrazine molecule contains two nitrogen atoms and four hydrogen atoms. The empirical formula for this compound is the smallest whole-number ratio, or NH_2 . However, chemists usually prefer to write this compound using the molecular formula: N_2H_4 .



Molecular formula: N_2H_4
Empirical formula: NH_2

Covalent bonds often lead to complex structures. Consider the molecule octane, a component of gasoline (**Figure 5.17**): One molecule of octane contains 25 different covalent bonds. Larger compounds may contain hundreds or thousands of covalent bonds. Because of this complex bonding, elements can often combine in many different ratios.

The Magnificent Seven

Elements that form Diatomic Molecules

Hydrogen: H_2

Nitrogen: N_2

Oxygen: O_2

Fluorine: F_2

Chlorine: Cl_2

Bromine: Br_2

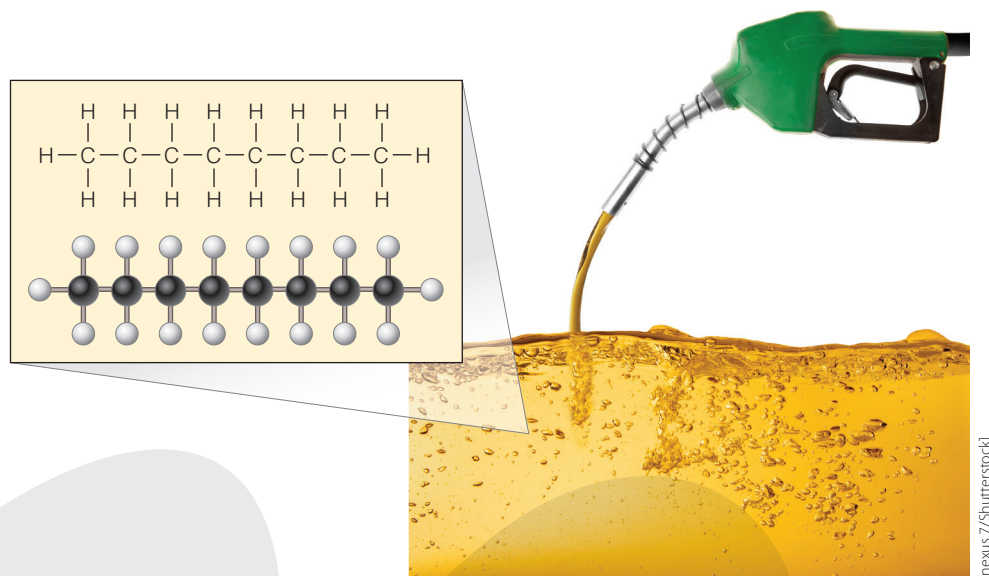
Iodine: I_2

Figure 5.15 Seven elements exist as diatomic molecules.

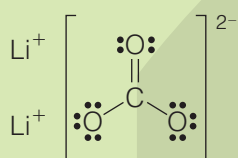
Covalent compounds fulfill the octet rule by sharing electrons. ■

Figure 5.16 We usually describe covalent molecules, such as the rocket fuel hydrazine, by their molecular formula rather than their empirical formula.

Figure 5.17 A molecule of octane is composed of hydrogen and carbon atoms, held together by covalent bonds.



[iStockphoto.com]



Lithium carbonate is an ionic compound used to treat bipolar disorder. In polyatomic ions like carbonate (CO_3^{2-}), the atoms are held together with covalent bonds.

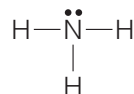
Each unique bonding arrangement produces a different compound. For example, **Table 5.5** lists several of the compounds that form between phosphorus and oxygen.

TABLE 5.5 Covalent Compounds Containing Phosphorus and Oxygen

Compound Name	Formula
Phosphorus monoxide	PO
Diphosphorus trioxide	P ₂ O ₃
Diphosphorus tetroxide	P ₂ O ₄
Tetraphosphorus decoxide	P ₄ O ₁₀

Example 5.8 Interpreting Lewis Structures

The Lewis structure for a molecule of ammonia (NH_3) is shown below. In this structure, how many electrons does the nitrogen atom share through covalent bonds? How many of the valence nitrogen electrons are not shared? Does this nitrogen atom have a complete octet?



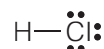
Each dash represents two shared electrons. We see from the structure that nitrogen forms three covalent bonds to hydrogen. Because each dash represents two electrons, we can say nitrogen has six shared electrons. The two dots above the nitrogen represent nonbonded (unshared) electrons. Combining the six shared and two unshared electrons, the nitrogen atom has eight electrons in its valence shell—a complete octet.



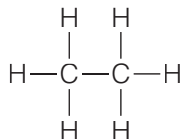
Check it
Watch explanation

TRY IT

9. The Lewis structure for the compound HCl is shown below. How many bonded and nonbonded electrons are in the valence of the chlorine atom? Does this atom fulfill the octet rule?



10. Consider the Lewis structure shown below. What is the molecular formula for this compound? What is the empirical formula?



Naming Covalent Compounds

Covalent compounds containing only two elements are called *binary covalent compounds*. These compounds are named in a manner that is similar to ionic compounds. The element that is lower and farther to the left on the periodic table is named first, and the full element name is used. The element that is nearer to the upper right on the periodic table is named as though it were an anion, by changing the end of the atom name to *-ide*.

However, there is one complicating factor: Because covalent compounds can form in many different ratios, covalent compounds use a series of prefixes (Table 5.6) to indicate the number of atoms present. A prefix is assigned to both the first and second part of the name. If the molecule contains only one atom of the first element, the prefix *mono-* is not used.

For example, phosphorus and chlorine commonly form two compounds that have the formulas PCl_3 and PCl_5 . How do we name these compounds? Phosphorus is to the left of chlorine on the periodic table (Figure 5.18), so we name phosphorus first and then name chlorine as the anion (*chloride*). Using the prefixes in Table 5.6, we refer to PCl_3 as *phosphorus trichloride*, and PCl_5 as *phosphorus pentachloride*.

PCl_3
Phosphorus trichloride

		P		Cl	

The atom to the left is named first...

...the atom to the right is named as the anion.

Figure 5.18 When naming a covalent compound, the atom that lies farthest left on the periodic table comes first.

Example 5.9 Naming Covalent Compounds

Nitrogen and oxygen form two covalent compounds, NO_2 and N_2O_4 . Name each of these compounds.

Because nitrogen is to the left of oxygen on the periodic table, we name nitrogen first (*nitrogen*) and then oxygen as the anion (*oxide*). The first compound, NO_2 , is called *nitrogen dioxide*. The second compound, N_2O_4 , is called *dinitrogen tetroxide*. Notice that we use the prefix on the first name only if more than one atom is present.

TRY IT

11. Write the names of these covalent compounds:



When naming covalent compounds, use prefixes to indicate how many atoms are present. ■

TABLE 5.6 Prefixes for Naming Covalent Compounds

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

Pent- or Penta-

If the root name of the atom begins with a vowel, we remove the *-a* from the end of the prefix to make it easier to pronounce. For example, PCl_5 is phosphorus **pentachloride**, but P_2O_5 is diphosphorus **pentoxide**.



Check it
Watch explanation

In the first example, MgBr_2 , magnesium is a metal and bromine is a nonmetal — so this is an ionic compound. We therefore name the compound simply by naming the cation first and then the anion. This compound is magnesium bromide.

In the second example, FeCl_3 , iron is a metal and chlorine is a nonmetal. Again, this is an ionic compound. Remember that iron forms cations with more than one charge, so we must specify the charge in parentheses. Because the cation is bound to three chloride ions, this ion is Fe^{3+} , or iron(III). This compound is iron(III) chloride.

In the third example, SF_6 , both sulfur and fluoride are nonmetals. Therefore this is a covalent compound, and we must use prefixes to indicate the number of each atom present. Because sulfur is to the left of chlorine on the periodic table, it is named first. This compound is sulfur hexafluoride.

TRY IT

12. Identify each of these compounds as ionic or covalent, and write its name:



Check it

Watch explanation

5.6 Aqueous Solutions: How Ionic and Covalent Compounds Differ

One of the most important differences between ionic and covalent compounds is how they behave when combined with water. To understand this critical difference, let's begin with some fundamental ideas: When a substance such as salt or sugar mixes with water, it disperses through the liquid, forming a homogeneous mixture called a **solution** (Figure 5.20). (If the liquid is water, we call it an *aqueous solution*.) When this happens, we say that the solid has *dissolved*. Compounds that dissolve in water are said to be **soluble** in water; those that do not are *insoluble*.



Figure 5.20 Ocean water contains many dissolved compounds. It is an aqueous solution.

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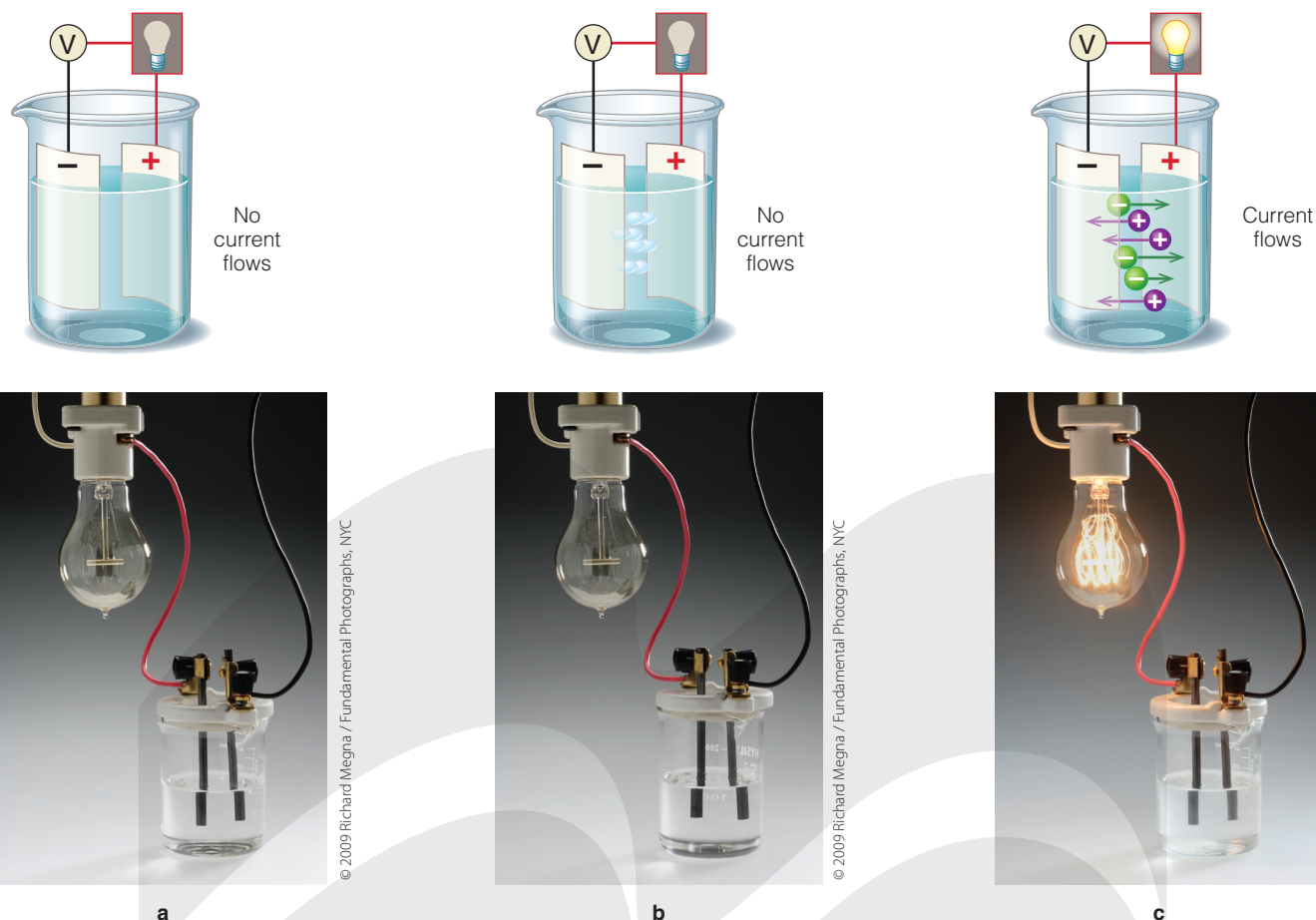


Figure 5.21 (a) Pure water is a poor conductor of electricity. (b) Nonionic compounds, such as sugar, may dissolve in water, but they do not increase the solution's ability to conduct electricity. (c) Ionic compounds, like salt, dissociate into ions; the resulting solution conducts electricity.

 **Explore**
Figure 5.21

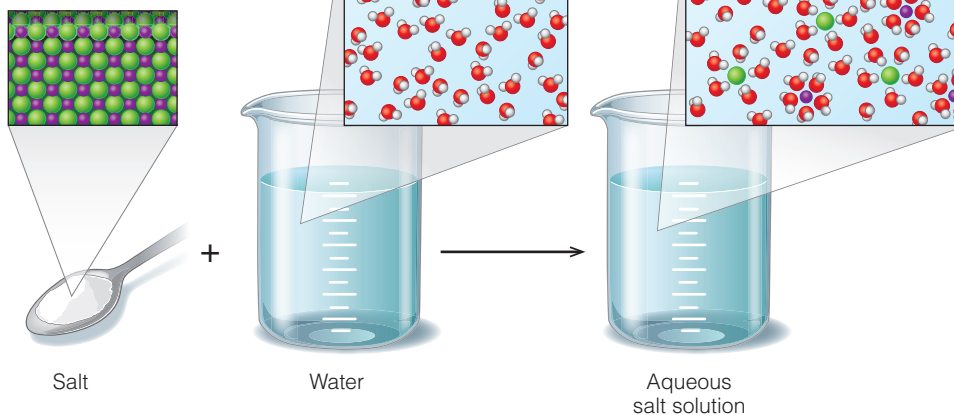
Devices that test for water purity often test how well the water conducts electricity. If a water sample conducts electricity well, we know that ionic compounds are present.

Pure water is a poor conductor of electricity (**Figure 5.21**). However, if ionic compounds are dissolved in water, the resulting solutions conduct electricity much more efficiently. Because of this property, we refer to aqueous ionic solutions as **electrolyte solutions**, and we call the ionic compounds *electrolytes*.

When ionic compounds dissolve in water, the positive and negative ions are pulled away from each other and surrounded by water ions (**Figure 5.22**). This

Figure 5.22 When an ionic solid like salt dissolves in water, the water molecules pull the ions away from the solid and into solution.

 **Explore**
Figure 5.22



process of pulling apart the ions in an ionic solid is called **dissociation**. The dissolved ions help carry electric current through the aqueous solutions.

As a general rule, covalent compounds do not form ions in water. Because of this, aqueous solutions containing only covalent compounds are not electrolytic (Figure 5.21b).

5.7 Acids—An Introduction

Most covalent compounds do not form ions when dissolved in water, but this rule has one important exception: **Acids** are covalent compounds that produce H^+ ions in aqueous solution. Most acids contain a covalent bond between hydrogen and a species that can form a stable anion. When dissolved in water, this bond breaks to produce a hydrogen cation and a corresponding anion.

For example, HCl and HNO_3 are both acidic molecules. When dissolved in water, these compounds *ionize* (form two ions):

- HCl ionizes to form H^+ and Cl^- in aqueous solution.
- HNO_3 ionizes to form H^+ and NO_3^- in aqueous solution.

We will explore the behavior of acids in Chapters 6 and 12. For now, it is important that you be able to identify and name common acids. The most common acids are listed in **Table 5.7**. When writing the formulas for acids, we typically write the formula with H first, as though it were the cation, followed by the anion.

Naming Acids

Binary Acids

Binary acids consist of H^+ and a single nonmetal element. The most common of these acids are those formed from the halogens: HF , HCl , HBr , and HI . These acids are named by combining the prefix *hydro-*, the root name of the halogen, and the suffix *-ic acid*:

HF	hydrofluoric acid
HCl	hydrochloric acid
HBr	hydrobromic acid
HI	hydroiodic acid

Oxyacids

Oxyacids are compounds that dissociate to form H^+ and an oxyanion. There are two rules for naming acids that dissociate to form oxyanions:

1. If the anion ends in *-ate*, name the acid by changing the suffix to *-ic acid*.
For example:

NO_3^-	nitrate ion	HNO_3	nitric acid
CO_3^{2-}	carbonate ion	H_2CO_3	carbonic acid

Derivatives of the sulfur and phosphorus oxyanions deviate slightly from this rule:

SO_4^{2-}	sulfate ion	H_2SO_4	sulfuric acid
PO_4^{3-}	phosphate ion	H_3PO_4	phosphoric acid

2. If the anion ends in *-ite*, name the acid by changing the suffix to *-ous acid*.

NO_2^-	nitrite ion	HNO_2	nitrous acid
-----------------	-------------	----------------	--------------

Acids produce H^+ ions in water. ■



Johan Larson/Shutterstock

Acids are *corrosive*, meaning they destroy many substances, including metal surfaces. They can also cause severe burns to the skin and should be handled with care.

TABLE 5.7 Common Acids

Formula	Name
HF	Hydrofluoric acid
HCl	Hydrochloric acid
HBr	Hydrobromic acid
HI	Hydroiodic acid
H_2CO_3	Carbonic acid
HNO_3	Nitric acid
HNO_2	Nitrous acid
H_2SO_4	Sulfuric acid
H_3PO_4	Phosphoric acid
$\text{HC}_2\text{H}_3\text{O}_2$	Acetic acid

Example 5.11 Naming Acids of Oxyanions

Name each of these acids, using the guidelines described earlier:

- a. HClO_4 b. H_2CrO_4

In water, HClO_4 ionizes to form H^+ and ClO_4^- ions. Because ClO_4^- is the perchlorate ion (see Table 5.3), HClO_4 is named *perchloric acid*. Similarly, CrO_4^{2-} is the chromate ion, so H_2CrO_4 is *chromic acid*.



Check it

Watch explanation

TRY IT

13. Name these acids:

- a. HF b. HClO c. $\text{HC}_2\text{H}_3\text{O}_2$

14. Write a formula for the acidic, ionic, and covalent compounds shown here.

- a. chlorous acid b. zinc chlorate c. boron trichloride



Capstone Video

Capstone Question

Ascorbic acid, more commonly known as vitamin C, is an essential part of your diet (**Figure 5.23**). A related compound, calcium ascorbate, is a common food additive and vitamin supplement with the chemical formula $\text{Ca}(\text{C}_6\text{H}_7\text{O}_6)_2$. Based on this, what is the formula and charge of the ascorbate ion? Using this information, predict (a) the empirical formula for sodium ascorbate, (b) the empirical formula for aluminum ascorbate, and (c) both the molecular and empirical formulas for ascorbic acid.



Kevin Revell

Figure 5.23 Vitamin C is a common component of citrus. A related compound, calcium citrate, is a common vitamin supplement.

SUMMARY

Chemical bonding involves the gain, loss, or sharing of valence electrons. A key factor in chemical bonding is the octet rule, which states that atoms are stabilized by the presence of eight electrons in their valence shells. Atoms that fulfill the octet rule have completely filled *s* and *p* sublevels in their valence shell.

To fulfill the octet rule, many atoms gain or lose electrons, forming ions. Metals tend to lose electrons to form positive ions (cations) while nonmetals tend to gain electrons to form negative ions (anions). Polyatomic ions are groups of atoms that contain an overall charge.

Ionic compounds are a combination of positive ions (cations) and negative ions (anions). In any ionic compound, the total charge must be equal to zero. When naming an ionic compound, we give the name of the cation first, followed by the name of the anion. Ionic compounds bind together in lattices of alternating charges. We describe an ionic compound by its empirical formula, which is the lowest whole-number ratio of atoms in that compound.

In covalent bonds, electrons are shared between two nonmetal atoms. Covalent compounds form discrete units called molecules. We typically describe a covalent solid by its molecular formula, which gives the number of each type of atom present in the molecule. When naming covalent compounds, we use prefixes to indicate the number of each type of atom present.

Ionic and covalent compounds behave differently in water. When ionic compounds dissolve in water, they dissociate into their component cations and anions. Dissolved ions enhance water's ability to conduct electricity. Because of this trait, ionic compounds are sometimes referred to as electrolytes. In contrast, most covalent compounds remain intact when dissolved in water.

Acids are covalent compounds that ionize in water to produce H^+ ions and a corresponding anion. The names of acids derive from the names of the anions they produce in solution.

Continuing Cade's Work

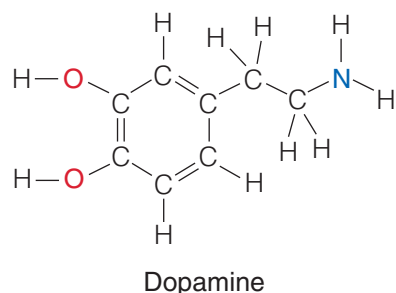


Figure 5.24 Dopamine is connected to mood, memory, and motor control.

A lot has changed since John Cade began using lithium carbonate to treat bipolar disorder. Today we have a much better (though far from complete) understanding of how ions and compounds affect our brain's function. For example, scientists now know that the covalent compound dopamine (**Figure 5.24**) plays a critical role in the working of the brain. Dopamine conveys signals between nerve cells, and it affects brain functions such as mood, memory, and motor control. Parkinson's disease (a degenerative disorder affecting muscle control) arises from a drop in dopamine levels. Other medical and cognitive issues, including drug addiction, perception of pain, appetite, and sexual gratification, all involve dopamine levels.

To perform its function, dopamine binds to cells in the central nervous system at special locations on the cell surface called *receptor sites*. When dopamine docks to a receptor site, it activates the site in much the same way that a key activates a lock. Like a key, the molecule's *size* and *shape* (along with other features) are critically important to its function. The shape of a molecule depends on the electronic structure of its atoms and on the covalent bonds that hold the atoms together. We'll explore the shape of molecules in much more detail in Chapter 9.

Medicinal chemists often search for molecules that can mimic the function of biological molecules like dopamine. They explore how slight changes in molecular structure (and therefore in molecule size and shape) affect the molecule's ability to bind to a receptor site. For molecules in the brain, these small differences in structure create profound differences in function.

For example, look at the three molecules in **Figure 5.25**. Do you notice their similarity to dopamine? The first molecule is adrenaline, a hormone that stimulates the nervous system. The second is ephedrine, a commercial decongestant and appetite suppressant. The third is methamphetamine, a devastatingly addictive, mood-altering drug. Like dopamine, each of these molecules affects brain function. But their small differences in size and shape affect how they bind to receptors, causing different responses in mood and behavior. ⚠️

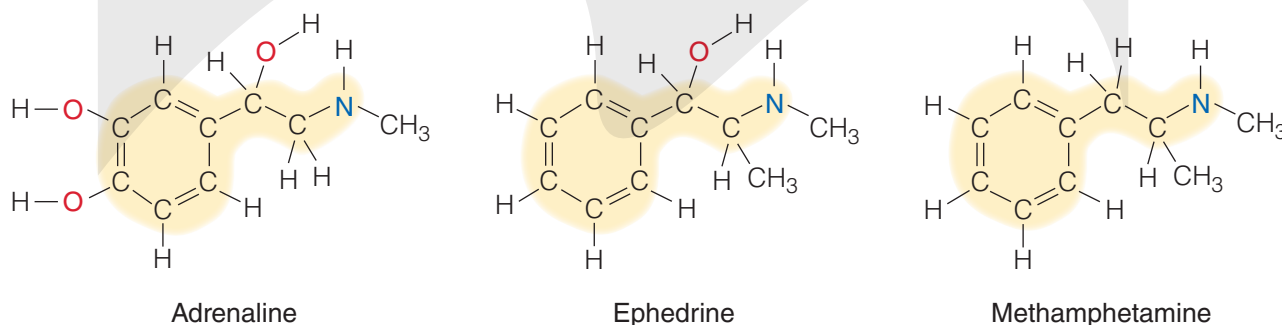


Figure 5.25 These compounds are similar to dopamine, and they also affect brain function. The yellow shading highlights their structural similarities.

Key Terms

5.1 Lewis Symbols and the Octet Rule

Lewis dot symbol A method of representing the valence structure of an atom or ion that involves using dots around the atomic symbol to indicate valence electrons.

5.2 Ions

cation A positively charged ion.

anion A negatively charged ion.

polyatomic ion A group of covalently bonded atoms with an overall charge.

oxyanion A negatively charged polyatomic ion that contains oxygen.

5.3 Ionic Bonds and Compounds

ionic bond A force of attraction between oppositely charged ions.

ionic compound A compound composed of oppositely charged ions.

ionic lattice A tightly packed array of alternating positive and negative charges; the characteristic arrangement of ions in an ionic solid.

chemical formula A representation of the type and amount of each element present in a compound.

empirical formula A chemical formula that gives the smallest whole-number ratio of atoms in a compound.

formula unit In ionic compounds, the smallest number of ions necessary to form a compound; the combination of atoms described by an empirical formula.

5.4 Covalent Bonding

covalent bond A bond in which two electrons are shared between atoms; covalent bonds typically form between nonmetals.

Lewis structure A depiction of the arrangement of valence electrons in a molecule or polyatomic ion, in which the Lewis symbols for atoms are shown connected by dashes representing covalent bonds.

covalent compounds Compounds formed by covalent bonds; these compounds form discrete groups of atoms called molecules.

molecular formula A formula that gives the actual number of atoms in the molecule.

5.6 Aqueous Solutions: How Ionic and Covalent Compounds Differ

solution A homogeneous mixture; for example, a solid mixed in a liquid.

soluble Having the ability to be dissolved in a liquid.

electrolyte solution An aqueous solution containing dissociated ions; this type of solution conducts electricity more effectively than pure water.

dissociation The process by which ions are pulled apart from a solid lattice when an ionic compound dissolves in water.

5.7 Acids—An Introduction

acid A covalent compound that produces H^+ ions in aqueous solution.

oxyacid A covalent compound that dissociates in aqueous solution to form H^+ and an oxyanion.

Additional Problems

5.1 Lewis Symbols and the Octet Rule

15. Using the periodic table, predict the number of valence electrons in each of these atoms:

Li C Si Kr Se

17. Write Lewis dot symbols to show the valence structures of each of these atoms:

Na N H As Sb

19. Write the electron configuration for the following atoms. Indicate which electrons are the valence electrons.

Mg N P I

21. Indicate whether each of these species fulfills the octet rule:

- a sodium atom
- a Na^+ ion
- a fluorine nucleus with 9 electrons
- a fluorine nucleus with 10 electrons

16. Using the periodic table, predict the number of valence electrons in each of these atoms:

Be Mg Ca Ge I

18. Write Lewis dot symbols to show the valence structures of each of these atoms:

Be F Ar Cs S

20. Write the electron configuration for the following atoms. Indicate which electrons are the valence electrons.

Be S Ge Br

22. Indicate whether each of these species fulfills the octet rule:

- a magnesium nucleus surrounded by 10 electrons
- a phosphorus atom
- an argon atom

5.2 Ions

23. What family of elements forms only +1 ions?

24. What family of elements forms only +2 ions?

- 25.** Potassium has an electronic structure of $[\text{Ar}]4s^1$. What is the electronic structure of the potassium ion (K^+)?
- 26.** Calcium has an electronic structure of $[\text{Ar}]4s^2$. What is the electronic structure of the calcium ion (Ca^{2+})?
- 27.** Write the electronic structure for each of these atoms and ions:
- a lithium atom
 - a lithium ion, Li^+
 - a sodium atom
 - a sodium ion, Na^+
- 28.** Write the electronic structure for each of these atoms and ions:
- a magnesium atom
 - a magnesium ion, Mg^{2+}
 - a beryllium atom
 - a beryllium ion, Be^{2+}
- 29.** Using the periodic table as a reference, predict the charge for each of these ions:
- a beryllium ion
 - a strontium ion
 - a sodium ion
 - a cesium ion
- 30.** Using the periodic table as a reference, predict the charge for each of these ions:
- a potassium ion
 - a barium ion
 - a calcium ion
 - a lithium ion
- 31.** What two charges are most common for a copper ion?
- 32.** What two charges are most common for an iron ion?
- 33.** Name each of the following cations:
- Na^+
 - Mg^{2+}
 - Cr^{2+}
 - Cr^{3+}
- 34.** Name each of the following cations:
- K^+
 - Ca^{2+}
 - Co^{2+}
 - Co^{3+}
- 35.** Name each of the following cations:
- Fe^{2+}
 - Fe^{3+}
 - Rb^+
 - Ba^{2+}
- 36.** Name each of the following cations:
- Sn^{2+}
 - Sn^{4+}
 - Ag^+
 - Be^{2+}
- 37.** Using the periodic table as a reference, write the symbol and charge for each cation:
- strontium
 - zinc
 - copper(II)
 - manganese(III)
- 38.** Using the periodic table as a reference, write the symbol and charge for each cation:
- aluminum
 - lead(II)
 - lead(IV)
 - magnesium
- 39.** What family of elements forms only -1 ions? What family of elements typically forms -2 ions?
- 40.** Unlike the other nonmetals, the noble gases do not form stable ions. Why is this so?
- 41.** The electronic structure of fluorine is $[\text{He}]2s^2 2p^5$. What is the electronic structure of the fluoride ion (F^-)?
- 42.** The electronic structure of oxygen is $[\text{He}]2s^2 2p^4$. What is the electronic structure of the oxide ion (O^{2-})?
- 43.** Write the electronic structure for each of these atoms and ions:
- a chlorine atom
 - a chloride ion, Cl^-
 - a bromine atom
 - a bromide ion, Br^-
- 44.** Write the electronic structure for each of these atoms and ions:
- a nitrogen atom
 - a nitride ion, N^{3-}
 - a sulfur atom
 - a sulfide ion, S^{2-}
- 45.** Indicate whether each atom would gain or lose electrons to fulfill the octet rule:
- Na
 - S
 - Mg
 - Br
- 46.** Indicate whether each atom would gain or lose electrons to fulfill the octet rule:
- Ba
 - O
 - K
 - F
- 47.** Determine whether the following would gain or lose electrons to fulfill the octet rule:
- a calcium atom
 - an atom in the halogen family
 - an atom with an electron configuration of $[\text{Ar}]4s^2 3d^{10} 4p^4$
 - an atom with an electron configuration of $[\text{Xe}]6s^2$
- 48.** Determine whether the following would gain or lose electrons to fulfill the octet rule:
- an alkaline earth metal
 - an oxygen atom
 - an atom with an electron configuration of $[\text{Ar}]4s^2 3d^{10} 4p^5$
 - an atom with an electron configuration of $[\text{Kr}]5s^1$

- 49.** Using the periodic table as a reference, write the symbol and charge for each of these ions:
- fluoride
 - iodide
 - oxide
 - selenide
-
- 51.** Name each of the following anions:
- F^-
 - S^{2-}
 - O^{2-}
 - I^-
-
- 53.** Using the periodic table as a reference, predict the charge of each of these ions:
- beryllium ion
 - oxide ion
 - chloride ion
-
- 55.** What charges would you expect on each of these ions?
- a halogen ion
 - an alkali metal ion
 - an ion formed from a neutral atom with electron configuration $[Ne]3s^23p^5$
-
- 57.** Write the name and the charge of the ion formed from each of these atoms:
- K
 - Rb
 - Cl
 - Br
-
- 59.** Write the symbol and charge for each of these ions:
- a fluoride ion
 - a strontium ion
 - a beryllium ion
 - a phosphide ion
-
- 61.** Identify each of these anions as monatomic or polyatomic:
- nitride
 - nitrate
 - sulfite
 - sulfide
-
- 63.** What two suffixes commonly indicate oxyanions?
-
- 65.** Write the formula and charge for each of these polyatomic ions:
- ammonium
 - carbonate
 - hydroxide
 - acetate
-
- 67.** Write the formula and charge for each of these polyatomic ions:
- chlorate
 - sulfite
 - hypochlorite
 - permanganate
-
- 69.** Write the symbol or formula and charge for each of these ions:
- tin(IV)
 - cupric ion
 - fluoride
 - sulfate
-
- 50.** Using the periodic table as a reference, write the symbol and charge for each of these ions:
- chloride
 - bromide
 - sulfide
 - phosphide
-
- 52.** Name each of the following anions:
- Cl^-
 - Br^-
 - P^{3-}
 - Te^{2-}
-
- 54.** Using the periodic table as a reference, predict the charge of each of these ions:
- bromide ion
 - sodium ion
 - barium ion
-
- 56.** What charges would you expect on each of these ions?
- an ion formed from a calcium atom
 - an alkaline earth metal ion
 - an ion formed from a neutral atom with electron configuration $[Ne]3s^1$
-
- 58.** Write the name and the charge of the ion formed from each of these atoms:
- Mg
 - Ca
 - O
 - S
-
- 60.** Write the symbol and charge for each of these ions:
- a ferrous ion
 - a copper(II) ion
 - a nitride ion
 - a rubidium ion
-
- 62.** Identify each of these anions as monatomic or polyatomic:
- bromate
 - bromite
 - bromide
 - perbromate
-
- 64.** Four common oxyanions are formed from bromine: BrO_4^- , BrO_3^- , BrO_2^- , and BrO^- . Name each of these ions.
-
- 66.** Write the formula and charge for each of these polyatomic ions:
- nitrate
 - nitrite
 - sulfate
 - bicarbonate
-
- 68.** Write the formula and charge for each of these polyatomic ions:
- cyanide
 - peroxide
 - dichromate
 - bisulfate
-
- 70.** Write the symbol or formula and charge for each of these ions:
- lead(II)
 - aluminum
 - bromide
 - chlorate

71. Write the symbol or formula and charge for each of these ions:

- zinc
- chromate
- sulfite
- phosphide

72. Write the symbol or formula and charge for each of these ions:

- iodide
- hydrogen phosphate
- iodate
- chromium(II)

5.3 Ionic Bonds and Compounds

73. Predict the empirical formulas for compounds formed from these ions:

- lithium and chloride
- calcium and bromide
- oxide and calcium
- iron(II) and phosphide

74. Predict the empirical formulas for compounds formed from these ions:

- sodium and fluoride
- chromium(III) and chloride
- silver and sulfide
- lithium and nitrite

75. Write the empirical formula for each of these compounds:

- aluminum chloride
- iron(II) sulfide
- calcium sulfate
- aluminum oxide

76. Write the empirical formula for each of these compounds:

- iron(III) nitrate
- copper(II) nitrate
- ammonium phosphate
- ammonium phosphide

77. Write the empirical formula for each of these compounds:

- chromium(III) acetate
- zinc chlorate
- silver nitrate
- lead(II) carbonate

78. Write the empirical formula for each of these compounds:

- tin(IV) chloride
- ammonium chlorite
- lithium bicarbonate
- cobalt(III) hydroxide

79. Write the empirical formula for each of these compounds:

- chromium(III) hypochlorite
- potassium permanganate
- sodium cyanide
- lead(II) perchlorate

80. Write the empirical formula for each of these compounds:

- tin(IV) chloride
- ammonium chlorate
- lithium bisulfate
- sodium hydrogen phosphate

81. In these compounds, determine the charge on the transition metal cation:

- SnCl_2
- SnCl_4
- $\text{Pb}(\text{NO}_3)_2$
- FeCO_3

82. In these compounds, determine the charge on the transition metal cation:

- Ag_3PO_4
- $\text{Cu}(\text{C}_2\text{H}_3\text{O}_2)_2$
- InBr_3
- $\text{Cr}_2(\text{SO}_4)_3$

83. Name each of these ionic compounds:

- NaBr
- K_2O
- FeBr_3
- CuS

84. Name each of these ionic compounds:

- KOH
- $\text{Cu}(\text{C}_2\text{H}_3\text{O}_2)_2$
- K_2CrO_4
- NH_4Cl

85. Name each of these ionic compounds:

- FeCO_3
- $\text{Al}(\text{NO}_2)_3$
- $\text{Ba}(\text{NO}_3)_2$
- $(\text{NH}_4)_2\text{SO}_4$

86. Name each of these ionic compounds:

- $\text{Na}_2\text{Cr}_2\text{O}_7$
- AgOH
- ZnCO_3
- $\text{Cr}_2(\text{SO}_4)_3$

5.4 Covalent Bonding

87. When two nonmetals bond together, why do they form a covalent bond rather than an ionic bond?

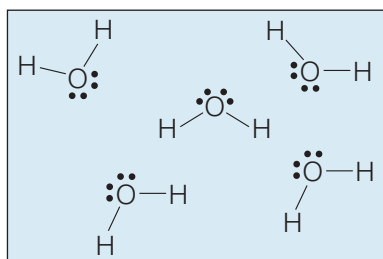
88. What seven elements exist as diatomic molecules in their elemental forms?

89. How many electrons are shared in a covalent bond? When drawing structures, how do we typically represent covalent bonds?

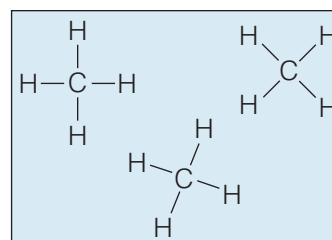
90. The Lewis structure of an H_2S molecule is shown. In this structure, how many electrons does the sulfur atom share through covalent bonds? How many of the valence electrons are not shared? Does this sulfur atom have a complete octet?



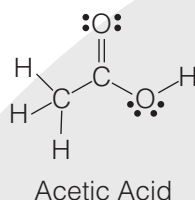
91. This figure shows a group of water molecules. How many water molecules are in this image? How many covalent bonds are present in each molecule?



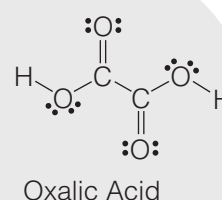
92. This figure shows a group of CH_4 molecules. How many CH_4 molecules are in this image? How many covalent bonds are present in each molecule?



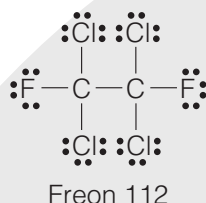
93. Acetic acid, shown here, is the main component of vinegar. Give the molecular formula and the empirical formula for this compound.



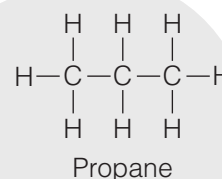
94. The structure of oxalic acid is shown here. Give the molecular formula and the empirical formula for this compound.



95. Compounds such as Freon[®] 112 (referred to as chlorofluorocarbons, or CFCs) were used as refrigerants for many years, but they were phased out because of their harmful effects on Earth's atmosphere. Write the molecular and empirical formulas for Freon 112.



96. Propane is a natural gas that is widely used as a heating fuel. Give the molecular and empirical formulas for propane.



97. Why is it necessary to use prefixes when naming binary covalent compounds?

98. When naming binary covalent compounds, when is a prefix not used?

99. Name each of these covalent compounds:

a. SCl_2 b. NF_3 c. N_2O_4 d. P_4O_{10}

100. Name each of these covalent compounds:

a. SO_3 b. CCl_4 c. N_2F_4 d. S_2Cl_2

101. Write molecular formulas for each of these covalent compounds:

a. arsenic tribromide
b. dinitrogen pentoxide
c. disulfur dioxide

102. Write molecular formulas for each of these covalent compounds:

a. disulfur dioxide
b. selenium tetrafluoride
c. tetraphosphorus trisulfide

5.5 Distinguishing Ionic and Covalent Compounds

103. When looking at a binary compound (one that has just two elements), how can we tell if it is ionic or covalent?

104. Why is it acceptable to write the formula of acetylene as C_2H_2 but not acceptable to write the formula of magnesium oxide as Mg_2O_2 ?

105. Determine whether these compounds contain ionic bonds or covalent bonds:

a. NaBr b. PCl_3 c. MnF_2

106. Determine whether these compounds contain ionic bonds or covalent bonds:

a. CO_2 b. N_2 c. KCl

107. Indicate whether these compounds would form an ionic lattice or discrete molecules:

- a. KCl b. CCl₄ c. P₄O₁₀ d. Na₂S

109. Indicate whether each of these compounds is ionic or covalent. Correctly name each compound.

- a. NaBr b. PBr₃ c. MgBr₂ d. SBr₂

111. Indicate whether each of these compounds is ionic or covalent. Correctly name each compound.

- a. SiCl₄ b. AlCl₃ c. BBr₃ d. Na₂SO₃

113. Write the correct chemical formula for each compound, using empirical formulas for ionic compounds and molecular formulas for covalent compounds.

- a. manganese(III) chloride b. phosphorus trichloride
c. sulfur dioxide d. titanium(IV) oxide

108. Indicate whether these compounds would form an ionic lattice or discrete molecules:

- a. CO₂ b. MgF₂ c. Ca(NO₃)₂ d. Na₃PO₄

110. Indicate whether each of these compounds is ionic or covalent. Correctly name each compound.

- a. SO₃ b. ZnO c. CO d. Fe₂O₃

112. Indicate whether each of these compounds is ionic or covalent. Correctly name each compound.

- a. MgSO₄ b. SO₃ c. NaHCO₃ d. CO₂

114. Write the correct chemical formula for each compound, using empirical formulas for ionic compounds and molecular formulas for covalent compounds.

- a. silver bromide b. selenium dibromide
c. sulfur trioxide d. copper(II) sulfite

5.6 Aqueous Solutions: How Ionic and Covalent Compounds Differ

115. What does the term *electrolyte* mean? What types of compounds are likely to be electrolytes?

116. By itself, water is a poor conductor of electricity. What must be present for water to conduct electricity efficiently?

117. When sodium sulfate is dissolved in water, it dissociates. What ions are present in an aqueous solution of sodium sulfate?

118. Ethylene glycol (C₂H₆O₂) is a covalent compound. Describe what happens to the C₂H₆O₂ molecules as this compound dissolves in water.

119. Which of these compounds are likely to dissociate in an aqueous solution? How can you tell?

- a. KCl b. CaBr₂ c. CO₂ d. C₂H₆O

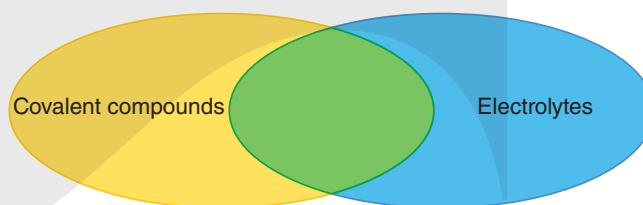
120. Which of these compounds are likely to dissociate in an aqueous solution? How can you tell?

- a. SO₃ b. Zn(ClO₄)₂ c. CS₂ d. MgF₂

5.7 Acids—An Introduction

121. What are acids? How do acids differ from most covalent compounds?

122. Place these compounds in the following Venn diagram: NaCl, CCl₄, and HCl.



123. Name the following acids:

- a. HCl b. HBr c. HI

124. H₂S is an acidic, foul-smelling gas. It is often called *hydrogen sulfide*. Name this compound using the rules for naming binary acids.

125. Name the following acids:

- a. HNO₃ b. HNO₂ c. HClO₄ d. HClO₂

126. Name the following acids:

- a. H₂SO₄ b. H₂SO₃ c. HClO₃ d. HClO

127. The formate ion is a biologically important ion with the formula CHO₂⁻. Based on this information, what is the name of the acid having the formula HCHO₂?

128. The selenate ion has the formula SeO₄²⁻. Based on this information, what is the name of the acid having the formula H₂SeO₄?

129. Classify each of these compounds as an ionic compound, a covalent compound, or an acid. Name each compound.

- a. NaNO₂ b. N₂O₄ c. HNO₂ d. KNO₂

130. Classify each of these compounds as an ionic compound, a covalent compound, or an acid. Name each compound.

- a. K₂SO₄ b. SO₂ c. H₂SO₄ d. NaHSO₄