



# CHAPTER 5

## CHEMICAL COMPOUNDS

Look around you. Do you think you see anything composed of just one element... any objects consisting only of carbon, or of gold, or of hydrogen? The correct answer is almost certainly no. If you are lucky enough to have a diamond ring, you have a piece of carbon that is almost pure (although a gemologist would tell you that diamonds contain slight impurities that give each stone its unique character). If you have a “gold” ring, you have a *mixture* of gold with other metals, added to give the ring greater strength.

Even though a few elements, such as carbon and gold, are sometimes found in elemental form in nature, most of the substances we see around us consist of two or more elements that have combined chemically to form more complex substances called compounds. For example, in nature, the element hydrogen is combined with other elements, such as oxygen and carbon, in compounds such as the water and sugar used to make a soft drink. (Perhaps you are sipping one while you read.) In this chapter, you will learn to (1) define the terms *mixture* and *compound* more precisely, (2) distinguish between elements, compounds, and mixtures, (3) describe how elements combine to form compounds, (4) describe the characteristics of certain kinds of chemical compounds, and (5) describe how atoms in compounds are attached and how these atoms are arranged in space. The chapter will also expand your ability to visualize the basic structures of matter.



The flecks of gold in this pan are the only pure elements visible in this scene.

- 5.1 Classification of Matter
- 5.2 Compounds and Chemical Bonds
- 5.3 Ionic Compounds
- 5.4 Molecular Compounds
- 5.5 Drawing Lewis Structures
- 5.6 Resonance
- 5.7 Molecular Geometry from Lewis Structures

### Review Skills

The presentation of information in this chapter assumes that you can already perform the tasks listed below. You can test your readiness to proceed by answering the Review Questions at the end of the chapter. This might also be a good time to read the Chapter Objectives, which precede the Review Questions.

- Describe the particle nature of solids, liquids, and gases. (Section 3.1)
- Convert between the names and symbols for the common elements. (Table 3.1)
- Given a periodic table, write the number of the group to which each element belongs. (Figure 3.3)
- Given a periodic table, identify the alkali metals, alkaline earth metals, halogens, and noble gases. (Section 3.3)
- Using a periodic table, classify elements as metals, nonmetals, or metalloids. (Section 3.3)
- Describe the nuclear model of the atom. (Section 3.4)
- Define the terms ion, cation, and anion. (Section 3.4)
- Define the terms covalent bond, molecule, and diatomic. (Section 3.5)
- Describe the covalent bond in a hydrogen molecule,  $H_2$ . (Section 3.5)
- Write or identify the definition of atomic orbital. (Section 4.2)
- Write electron configurations and orbital diagrams for the nonmetallic elements. (Section 4.3)

## 5.1 Classification of Matter

Before getting started on your chemistry homework, you go into the kitchen to make some pasta for your six-year-old nephew. You run water into a pan, adding a few shakes of salt, and while you're waiting for it to boil, you pour a cup of coffee. When the water begins to boil, you pour in the pasta. Then you add some sugar to your coffee.

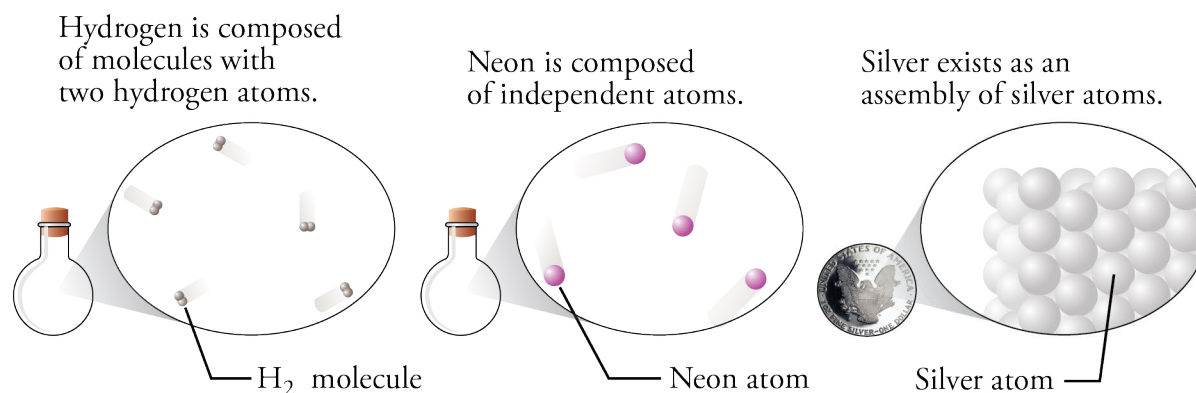
Pure water, the sucrose in white sugar, and the sodium chloride in table salt are all examples of chemical compounds. A **compound** is a substance that contains two or more elements, with the atoms of those elements always combining in the same whole-number ratio (Figure 5.1). There are relatively few chemical elements, but there are millions of chemical compounds. Compounds in our food fuel our bodies, and the compounds in gasoline fuel our cars. They can alter our moods and cure our diseases.

### OBJECTIVE 2

Water is composed of molecules that contain two atoms of hydrogen and one atom of oxygen. We describe the composition of water with the chemical formula  $\text{H}_2\text{O}$ . White sugar is a highly purified form of sucrose, whose chemical formula is  $\text{C}_{12}\text{H}_{22}\text{O}_{11}$ . Its molecules are composed of 12 carbon atoms, 22 hydrogen atoms, and 11 oxygen atoms. Sodium and chlorine atoms combine in a one-to-one ratio to form sodium chloride,  $\text{NaCl}$ , which is the primary ingredient in table salt.

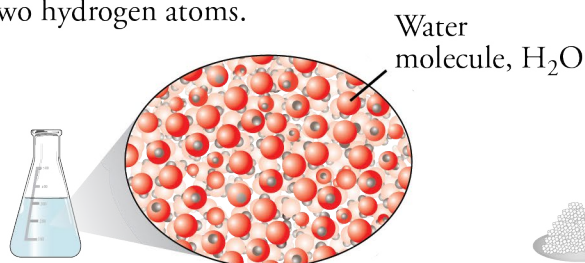
**Figure 5.1**  
Elements Versus Compounds

### ELEMENTS

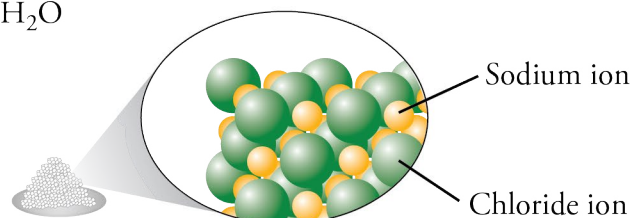


### COMPOUNDS

Water is composed of molecules that contain one oxygen atom and two hydrogen atoms.



Sodium chloride exists as an assembly of sodium and chloride ions, always in a one-to-one ratio.



Note that a **chemical formula** is a concise written description of the components of a chemical compound. It identifies the elements in the compound by their symbols and describes the relative number of atoms of each element with subscripts. If an element symbol in a formula is not accompanied by a subscript, the relative number of atoms of that element is assumed to be one.

OBJECTIVE 2

Pure water, sodium chloride, and sucrose always have the composition described in their chemical formulas. In other words, their composition is constant. Elements, too, have a constant composition described by a chemical formula. (We have seen that the formula for hydrogen is  $H_2$ .) When a substance has a constant composition—when it can be described by a chemical formula—it must by definition be either an element or a compound, and it is considered a **pure substance**. For example, the symbol Na refers to pure sodium. The formula  $Na_2CO_3$  refers to pure sodium carbonate and tells us that this compound is always composed of sodium, carbon, and oxygen in a constant atom ratio of 2:1:3.

OBJECTIVE 3

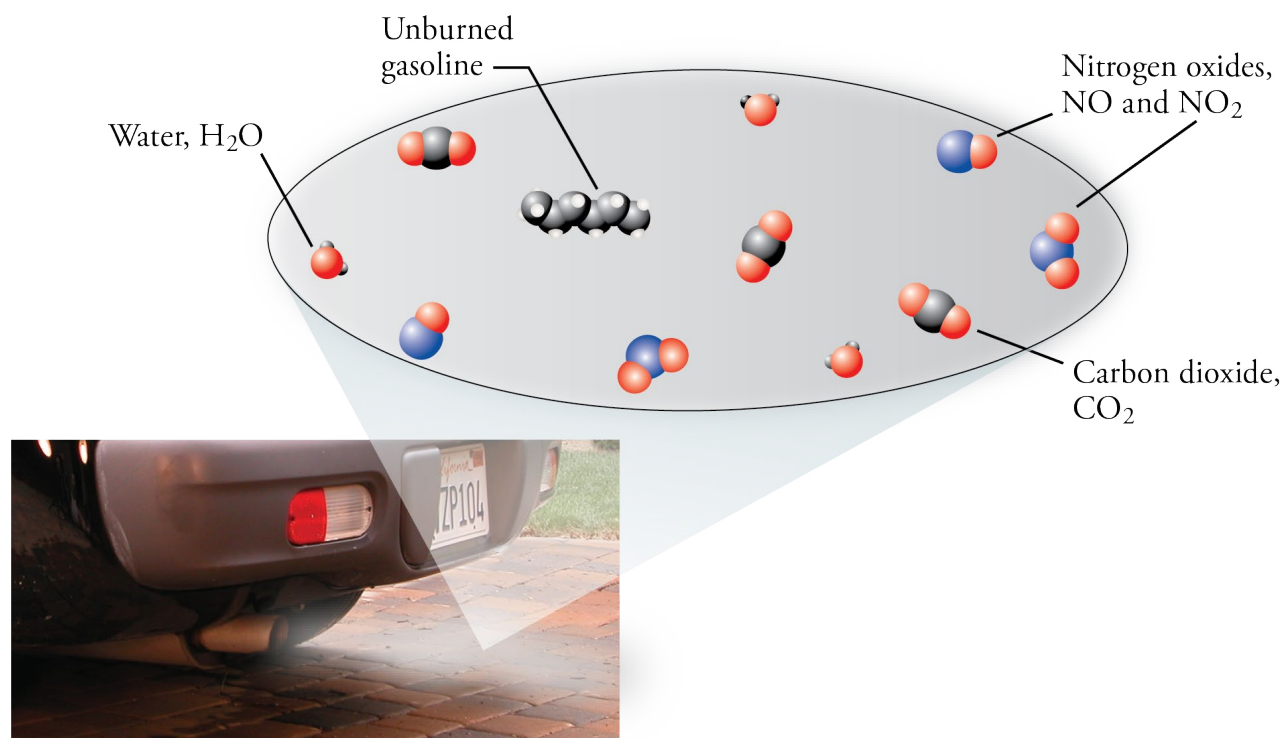
OBJECTIVE 4

**Mixtures** are samples of matter that contain two or more pure substances and have variable composition (Figure 5.2). For example, when salt,  $NaCl$ , and water,  $H_2O$ , are combined, we know the resulting combination is a mixture because we can vary the percentage of these two pure substances. You can add one, two, or ten teaspoons of salt to a pan of water, and the result will still be salt water.

OBJECTIVE 3

**Figure 5.2**  
Automobile Exhaust—a Mixture  
The components and amounts vary.

OBJECTIVE 3



The following sample study sheet and Figure 5.3 show the questions you can ask to discover whether a sample of matter is an element, a compound, or a mixture.

## Sample Study Sheet 5.1

### Classification of Matter

OBJECTIVE 3

OBJECTIVE 4

**TIP-OFF** You are asked to classify a sample of matter as a pure substance or a mixture; or you are asked to classify a pure substance as an element or a compound.

**GENERAL STEPS** The following general procedure is summarized in Figure 5.3.

- To classify a sample of matter as a pure substance or a mixture, ask one or both of the following questions:

*Does it have a constant composition?* If it does, it is a pure substance. If it has variable composition, it is a mixture.

*Can the sample as a whole be described with a chemical formula?* If it can, it is a pure substance. If it cannot, it is a mixture.

- To classify a pure substance as an element or a compound, ask the following question:

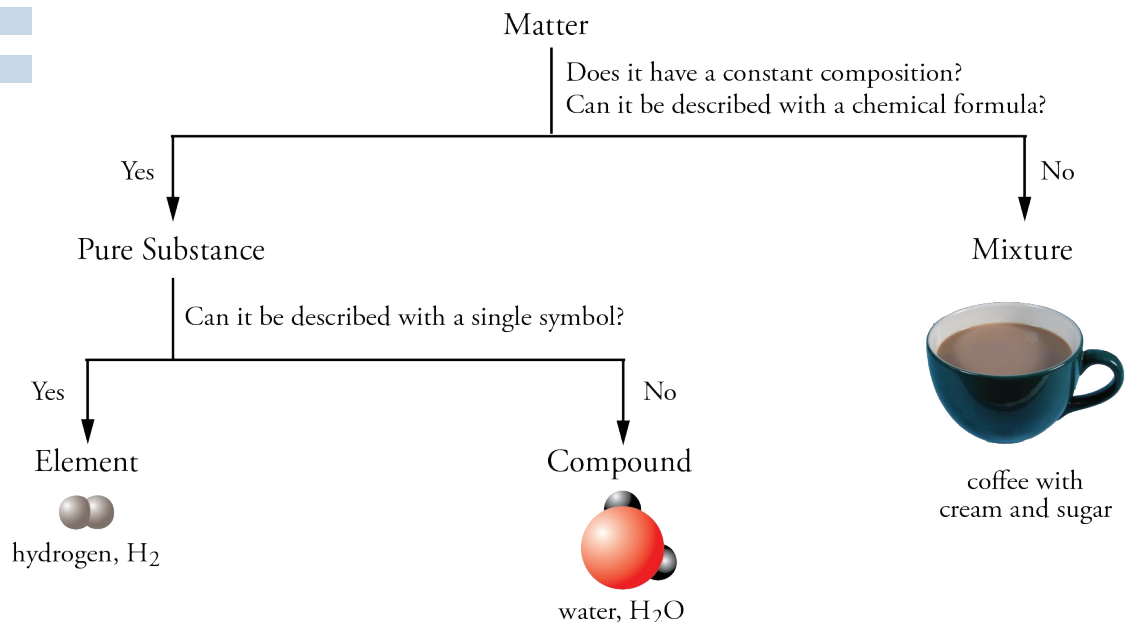
*Can it be described with a single symbol?* If it can, it is an element. If its chemical formula contains two or more different element symbols, it is a compound.

**EXAMPLE** See Example 5.1.

**Figure 5.3**  
Classification of Matter

OBJECTIVE 3

OBJECTIVE 4



**EXAMPLE 5.1 - Classification of Matter**

Many of us have a bottle in our medicine cabinet containing a mild disinfectant consisting of hydrogen peroxide and water. The liquid is about 3% hydrogen peroxide,  $\text{H}_2\text{O}_2$ , and about 97% water. Classify each of the following as a pure substance or a mixture. If it is a pure substance, is it an element or a compound?

- the liquid disinfectant
- the hydrogen peroxide,  $\text{H}_2\text{O}_2$ , used to make the disinfectant
- the hydrogen used to make hydrogen peroxide

*Solution*

- We know that the liquid disinfectant is a **mixture** for two reasons. It is composed of two pure substances ( $\text{H}_2\text{O}_2$  and  $\text{H}_2\text{O}$ ), and it has variable composition.
- Because hydrogen peroxide can be described with a formula,  $\text{H}_2\text{O}_2$ , it must be a **pure substance**. Because the formula contains symbols for two elements, it represents a **compound**.
- Hydrogen can be described with a single symbol, H or  $\text{H}_2$ , so it is a **pure substance** and an **element**.

OBJECTIVE 3

OBJECTIVE 4

**EXERCISE 5.1 - Classification of Matter**

The label on a container of double-acting baking powder tells us that it contains cornstarch, bicarbonate of soda (also called sodium hydrogen carbonate,  $\text{NaHCO}_3$ ), sodium aluminum sulfate, and acid phosphate of calcium (which chemists call calcium dihydrogen phosphate,  $\text{Ca}(\text{H}_2\text{PO}_4)_2$ ). Classify each of the following as a pure substance or a mixture. If it is a pure substance, is it an element or a compound?

- calcium
- calcium dihydrogen phosphate
- double-acting baking powder

OBJECTIVE 3

OBJECTIVE 4

**5.2 Compounds and Chemical Bonds**

The percentage of  $\text{H}_2\text{O}_2$  in the mixture of hydrogen peroxide and water that is used as a disinfectant can vary, but the percentage of hydrogen in the compound water is always the same. Why? One of the key reasons that the components of a given compound are always the same, and present in the same proportions, is that the atoms in a compound are joined together by special kinds of attractions called **chemical bonds**. Because of the nature of these attractions, the atoms combine in specific ratios that give compounds their constant composition. This section will introduce you to the different types of chemical bonds and provide you with the skills necessary to predict the types of chemical bonds between atoms of different elements.

## Equal and Unequal Sharing of Electrons

Let's first consider the compound hydrogen chloride, HCl. When HCl is dissolved in water, the resulting mixture is called hydrochloric acid. Not only is this mixture a very common laboratory agent, but it is also used in food processing and to treat the water in swimming pools.

In Section 3.5, we learned about the bond between hydrogen atoms in H<sub>2</sub> molecules. We saw that the two electrons in the H<sub>2</sub> molecule are shared equally between the atoms and can be viewed as an electron-charge cloud surrounding the hydrogen nuclei. This sharing creates a covalent bond that holds the atoms together. There is also a covalent bond between the hydrogen atom and the chlorine atom in each molecule of HCl. It is very similar to the covalent bond in hydrogen molecules, with one important exception.

H<sub>2</sub>

### OBJECTIVE 5

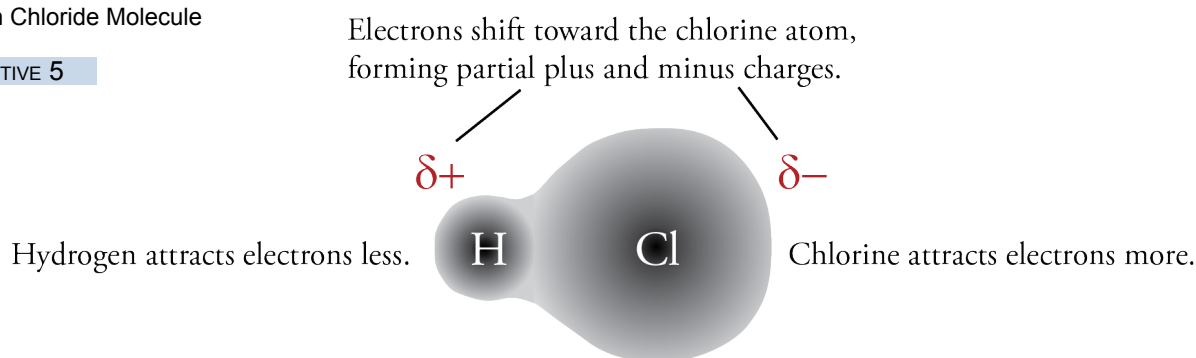
The difference between the H–Cl bond and the H–H bond is that the hydrogen and chlorine atoms in HCl do not share the electrons in the bond equally. In the hydrogen-chlorine bond, the two electrons are attracted more strongly to the chlorine atom than to the hydrogen atom. The negatively charged electrons in the bond shift toward the chlorine atom, giving it a partial negative charge,  $\delta^-$ , and giving the hydrogen atom a partial positive charge,  $\delta^+$  (Figure 5.4). The lower case Greek delta,  $\delta$ , is a symbol that represents *partial* or *fractional*.

### OBJECTIVE 7

When the electrons of a covalent bond are shared unequally, the bond is called a **polar covalent bond**. Due to the unequal sharing of the electrons in the bond, a polar covalent bond has one atom with a partial positive charge,  $\delta^+$ , and one atom with a partial negative charge,  $\delta^-$ .

**Figure 5.4**  
Hydrogen Chloride Molecule

### OBJECTIVE 5



If the electron-attracting ability of one atom in a bond is much greater than the others, there is a large shift in the electron cloud, and the partial charges are large. If the electron-attracting ability of one atom in a covalent bond is only slightly greater than the others, there is not much of a shift in the electron cloud, and the partial charges are small. When the difference in electron-attracting abilities is negligible (or zero), the atoms in the bond will have no significant partial charges. We call this type of bond a **nonpolar covalent bond**. The covalent bond between hydrogen atoms in H<sub>2</sub> is an example of a nonpolar covalent bond.

### OBJECTIVE 7

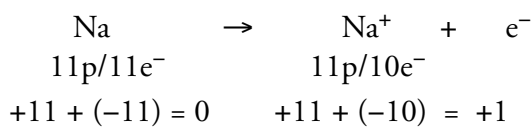
## Transfer of Electrons

Sometimes one atom in a bond attracts electrons so much more strongly than the other that one or more electrons are fully transferred from one atom to another. This commonly happens when metallic atoms combine with nonmetallic atoms. A nonmetallic atom usually attracts electrons so much more strongly than a metallic atom that one or more electrons shift from the metallic atom to the nonmetallic atom. For example, when the element sodium combines with the element chlorine to form sodium chloride, NaCl, the chlorine atoms attract electrons so much more strongly than the sodium atoms that one electron is transferred from each sodium atom to a chlorine atom.

OBJECTIVE 6

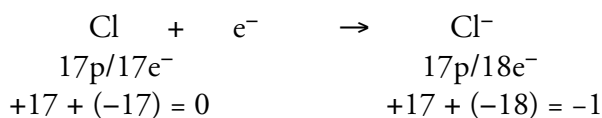
When an electron is transferred completely from one uncharged atom to another, the atom that loses the electron is left with one more proton than electron and acquires a +1 charge overall. It therefore becomes a cation. For example, when an uncharged sodium atom with 11 protons and 11 electrons loses an electron, it is left with 11 protons (a charge of +11) and 10 electrons (a charge of -10), yielding an overall +1 charge.

OBJECTIVE 6



In contrast, an uncharged atom that gains an electron will have one more electron than proton, so it forms an anion with a -1 charge. When a chlorine atom gains an electron from a sodium atom, the chlorine atom changes from an uncharged atom with 17 protons and 17 electrons to an anion with 17 protons and 18 electrons and an overall -1 charge.

OBJECTIVE 6



Atoms can transfer one, two, or three electrons. Thus cations can have a +1, +2, or +3 charge, and anions can have a -1, -2, or -3 charge.

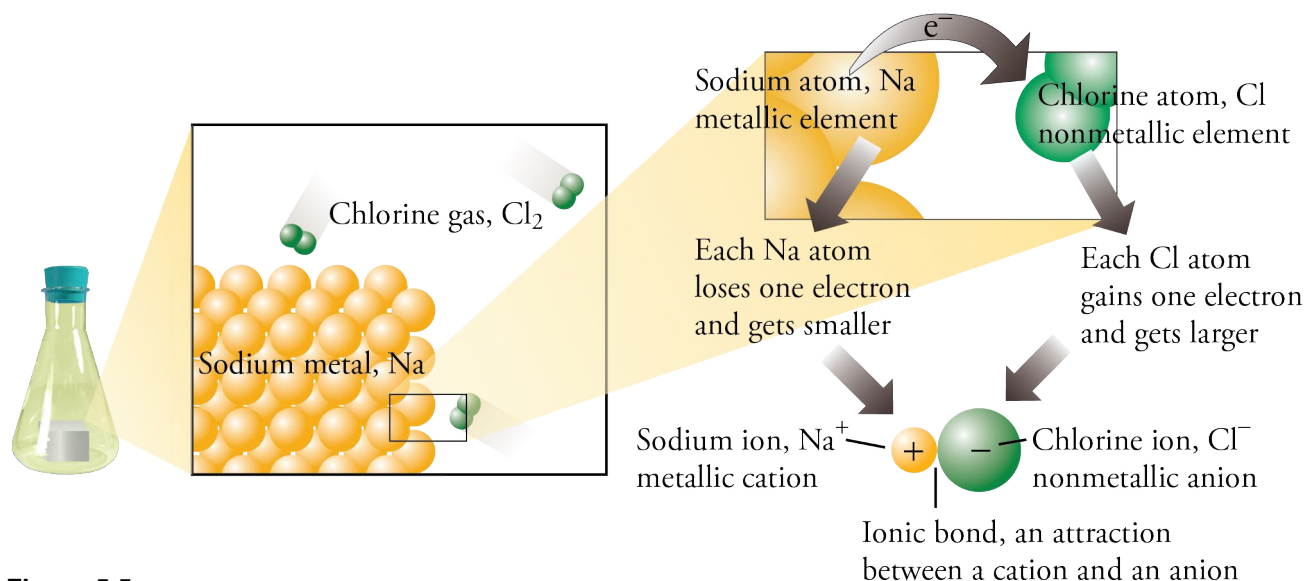
Because particles with opposite charges attract each other, there is an attraction between cations and anions. This attraction is called an **ionic bond**. For example, when an electron is transferred from a sodium atom to a chlorine atom, the attraction between the +1 sodium cation and the -1 chlorine anion is an ionic bond (Figure 5.5 on the next page).

You will see as you read more of this book that substances that have ionic bonds are very different from those that have all covalent bonds. For example, compounds that have ionic bonds, such as the sodium chloride in table salt, are solids at room temperature and pressure, but compounds with all covalent bonds, such as hydrogen chloride and water, can be gases and liquids as well as solids.



**The Salt-Encrusted Shore of The Dead Sea**

Salt (sodium chloride) is an ionic compound. Water is molecular.



**Figure 5.5**  
Ionic Bond Formation

OBJECTIVE 6

## Summary of Covalent and Ionic Bond Formation

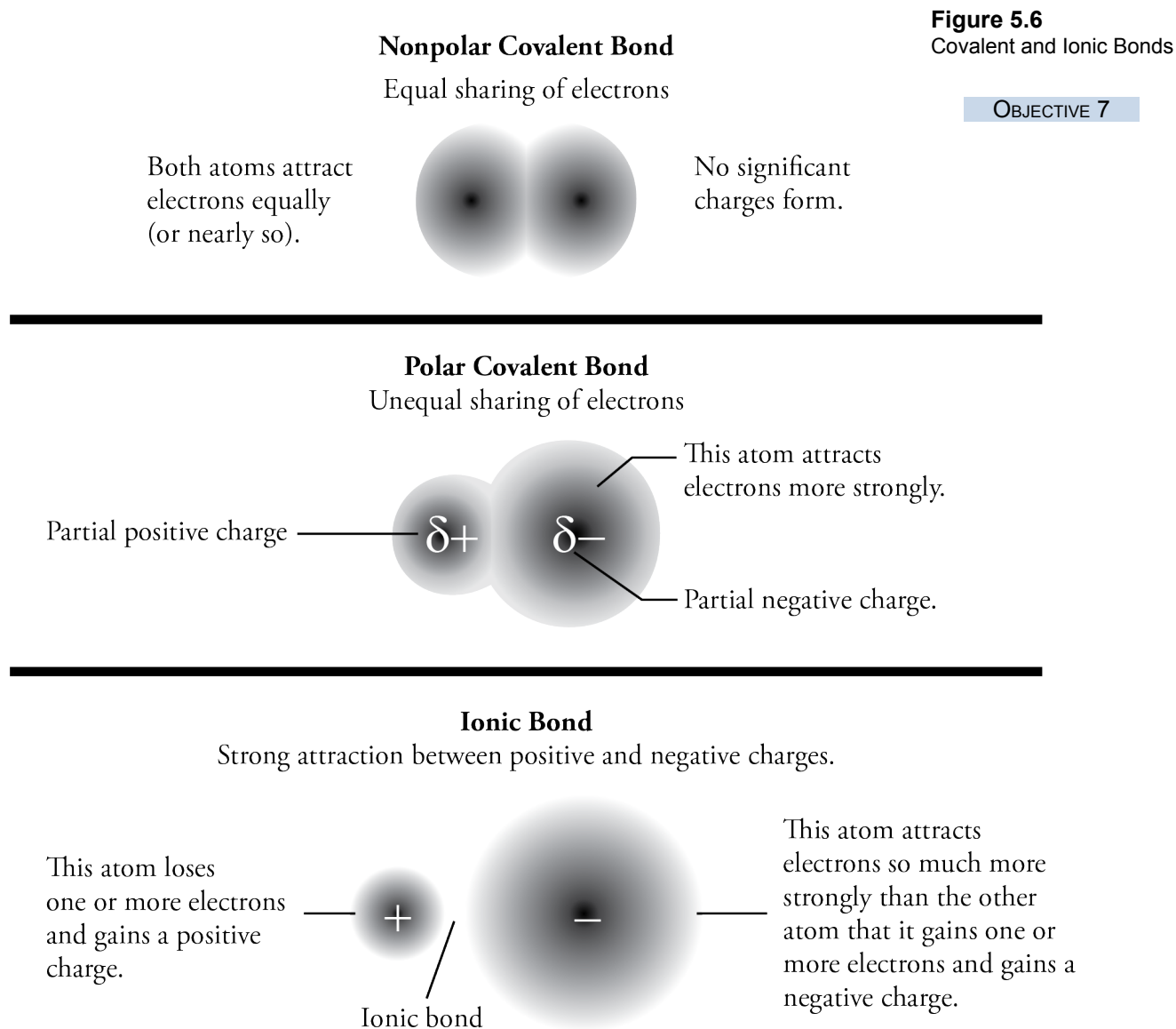
OBJECTIVE 7

- When atoms of different elements form chemical bonds, the electrons in the bonds can shift from one bonding atom to another.
- The atom that attracts electrons more strongly will acquire a negative charge, and the other atom will acquire a positive charge.
- The more the atoms differ in their electron-attracting ability, the more the electron cloud shifts from one atom toward another.
- If there is a large enough difference in electron-attracting ability, one, two, or three electrons can be viewed as shifting completely from one atom to another. The atoms become positive and negative ions, and the attraction between them is called an ionic bond.
- If the electron transfer is significant but not enough to form ions, the atoms acquire partial positive and partial negative charges. The bond in this situation is called a polar covalent bond.
- If there is no shift of electrons or if the shift is negligible, no significant charges will form, and the bond will be a nonpolar covalent bond.

It might help, when thinking about these different kinds of bonds, to compare them to a game of tug-of-war between two people. The people are like the atoms with a chemical bond between them, and the rope is like the electrons in the bond. If the two people tugging have the same (or about the same) strength, the rope will not move (or not move much). This leads to a situation that is like the nonpolar covalent bond. If one person is stronger than the other person, the rope will shift toward that person, the way the electrons in a polar covalent bond shift toward the atom that attracts them more. If one person can pull a lot harder than the other person can, the stronger person pulls the rope right out of the hands of the weaker one. This is similar to the formation of ions and ionic bonds, when a nonmetallic atom pulls one or more electrons away from a metallic atom.



Figure 5.6 summarizes the general differences between nonpolar covalent bonds, polar covalent bonds, and ionic bonds.



## Predicting Bond Type

The simplest way to predict whether a bond will be ionic or covalent is to apply the following rules.

**OBJECTIVE 8**

- When a nonmetallic atom bonds to another nonmetallic atom, the bond is covalent.
- When a metallic atom bonds to a nonmetallic atom, the bond is *usually* ionic.

Some bonds between a metallic atom and a nonmetallic atom are better described as covalent. For now, however, we will keep our guidelines simple. All nonmetal-nonmetal combinations lead to covalent bonds, and except when you are told otherwise, you can assume that all bonds between metallic atoms and nonmetallic atoms are ionic bonds.

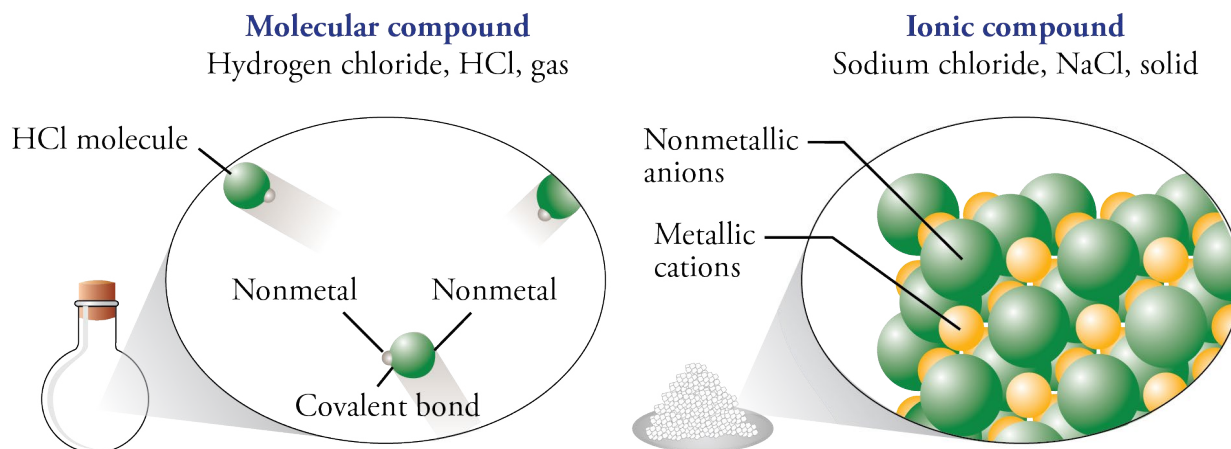
## Classifying Compounds

### OBJECTIVE 9

Compounds can be classified as molecular or ionic. **Molecular compounds** are composed of molecules, which are uncharged collections of atoms held together by all covalent bonds. **Ionic compounds** contain cations and anions held together by ionic bonds (Figure 5.7). You will see some exceptions later in this text, but for now, if a formula for a compound indicates that all the elements in it are nonmetals, you can assume that all of the bonds are covalent bonds, which form molecules, and that the compound is a molecular compound. We will assume that metal-nonmetal combinations lead to ionic bonds and ionic compounds.

**Figure 5.7**  
Classifying Compounds

### OBJECTIVE 9



### EXAMPLE 5.2 - Classifying Compounds

#### OBJECTIVE 9

Classify each of the following as either a molecular compound or an ionic compound.

- calcium chloride,  $\text{CaCl}_2$  (used for de-icing roads)
- ethanethiol,  $\text{C}_2\text{H}_5\text{SH}$  (a foul-smelling substance used to odorize natural gas)

#### *Solution*

- Calcium, Ca, is a metal, and chlorine, Cl, is a nonmetal. We expect the bonds between them to be ionic, so calcium chloride is an **ionic compound**.
- Carbon, hydrogen, and sulfur are all nonmetallic elements, so we expect the bonds between them to be covalent bonds. The formula,  $\text{C}_2\text{H}_5\text{SH}$ , tells us that ethanethiol is composed of molecules that each contain two carbon atoms, six hydrogen atoms, and one sulfur atom. Ethanethiol is a **molecular compound**.

### EXERCISE 5.2 - Classifying Compounds

#### OBJECTIVE 9

Classify each of the following substances as either a molecular compound or an ionic compound.

- formaldehyde,  $\text{CH}_2\text{O}$  (used in embalming fluids)
- magnesium chloride,  $\text{MgCl}_2$  (used in fireproofing wood and in paper manufacturing)

## 5.3 Ionic Compounds

As you saw in the last section, ionic compounds are substances composed of ions attracted to each other by ionic bonds. Let's consider how they play a part in a "typical" family's Fourth of July.

Before the family leaves the house to go to the holiday picnic, the kids are sent off to brush their teeth and change into clean clothes. Their toothpaste contains sodium fluoride, a common cavity-fighting ionic compound. The white shirts in their red, white, and blue outfits were bleached with the ionic compound sodium hypochlorite, the stains on their red pants were removed by potassium oxalate, and dyes were fixed to their blue hats by aluminum nitrate.

While the kids are getting ready, dad and mom get the picnic dinner together. The hot dogs they are packing are cured with the ionic compound sodium nitrite, and the buns contain calcium acetate as a mold inhibitor and calcium iodate as a dough conditioner. The soft drinks have potassium hydrogen carbonate to help trap the bubbles, the mineral water contains magnesium sulfate, and the glass for the bottles was made with a variety of ionic compounds. Because it will be dark before they get home, Mom packs a flashlight as well. Its rechargeable batteries contain the ionic compounds cadmium hydroxide and nickel hydroxide.

When our family gets to the park, they find themselves a place on the lawn, which was fertilized with a mixture of ionic compounds, including iron(II) sulfate. They eat their dinner and play in the park until it's time for the fireworks. The safety matches used to light the rockets contain barium chromate, and ionic compounds in the fireworks provide the colors: red from strontium chlorate, white from magnesium nitrate, and blue from copper(II) chloride.



Ionic compounds play an important role in creating the excitement of a Fourth of July celebration.

### Cations and Anions

Remember the sodium fluoride in the kids' toothpaste? It could be made from the reaction of sodium metal with the nonmetallic atoms in fluorine gas. Metallic atoms hold some of their electrons relatively loosely, and as a result, they tend to lose electrons and form cations. In contrast, nonmetallic atoms attract electrons more strongly than metallic atoms, and so nonmetals tend to gain electrons and form anions. Thus, when a metallic element and a nonmetallic element combine, the nonmetallic atoms often pull one or more electrons far enough away from the metallic atoms to form ions. The positive cations and the negative anions then attract each other to form ionic bonds. In the formation of sodium fluoride from sodium metal and fluorine gas, each sodium atom donates one electron to a fluorine atom to form a  $\text{Na}^+$  cation and an  $\text{F}^-$  anion. The  $\text{F}^-$  anions in toothpaste bind to the surface of your teeth, making them better able to resist tooth decay. This section provides you with more information about other cations and anions, including how to predict their charges.

#### OBJECTIVE 10



Sodium fluoride is added to toothpaste to help fight tooth decay.

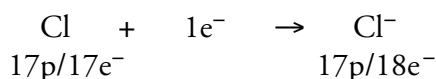
## Predicting Ionic Charges

It is useful to be able to predict the charges the atoms of each element are most likely to attain when they form ions. Because the periodic table can be used to predict ionic charges, it is a good idea to have one in front of you when you study this section.

We discovered in Chapter 3 that the atoms of the noble gases found in nature are uncombined with other atoms. The fact that the noble gas atoms do not gain, lose, or share their electrons suggests there must be something especially stable about having 2 (helium, He), 10 (neon, Ne), 18 (argon, Ar), 36 (krypton, Kr), 54 (xenon, Xe), or 86 (radon, Rn) electrons. This stability is reflected in the fact that nonmetallic atoms form anions in order to get the same number of electrons as the nearest noble gas.

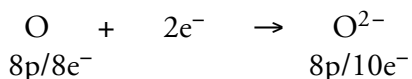
### OBJECTIVE 11

All of the halogens (group 17) have one less electron than the nearest noble gas. When halogen atoms combine with metallic atoms, they tend to gain one electron each and form  $-1$  ions (Figure 5.8). For example, uncharged chlorine atoms have 17 protons and 17 electrons. If a chlorine atom gains one electron, it will have 18 electrons like an uncharged argon atom. With a  $-18$  charge from the electrons and a  $+17$  charge from the protons, the resulting chlorine ion has a  $-1$  charge. The symbol for this anion is  $\text{Cl}^-$ . Note that the negative charge is indicated with a “ $-$ ” not “ $-1$ ” or “ $1-$ ”.



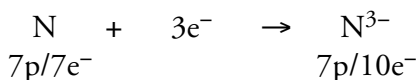
### OBJECTIVE 11

The nonmetallic atoms in group 16 (oxygen, O, sulfur, S, and selenium, Se) have two fewer electrons than the nearest noble gas. When atoms of these elements combine with metallic atoms, they tend to gain two electrons and form  $-2$  ions (Figure 5.8). For example, oxygen, in group 16, has atoms with eight protons and eight electrons. Each oxygen atom can gain two electrons to achieve ten, the same number as its nearest noble gas, neon. The symbol for this anion is  $\text{O}^{2-}$ . Notice that the charge is indicated with “ $2-$ ” not “ $-2$ ”.



### OBJECTIVE 11

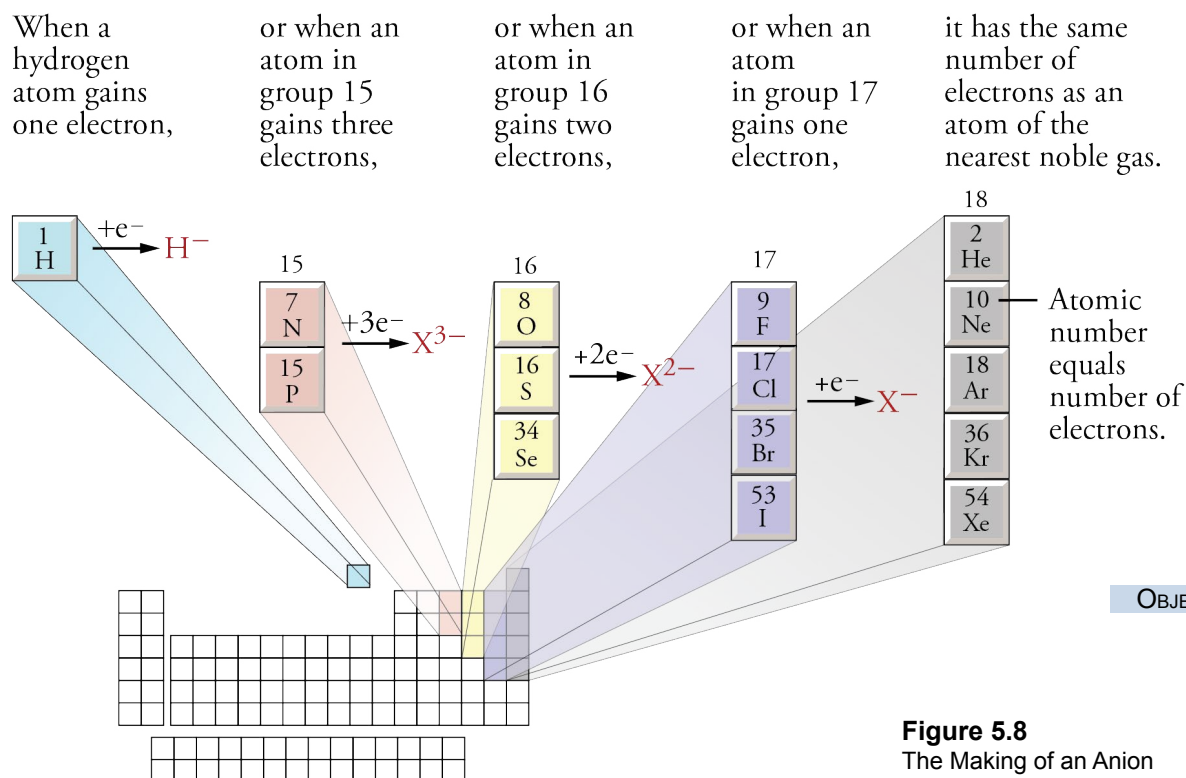
Nitrogen, N, and phosphorus, P, have three fewer electrons than the nearest noble gas. When atoms of these elements combine with metallic atoms, they tend to gain three electrons and form  $-3$  ions (Figure 5.8). For example, nitrogen atoms have seven protons and seven electrons. Each nitrogen atom can gain three electrons to achieve ten, like neon, forming a  $\text{N}^{3-}$  anion.



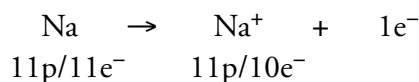
### OBJECTIVE 11

Hydrogen has one less electron than helium, so when it combines with metallic atoms, it forms a  $-1$  ion,  $\text{H}^-$  (Figure 5.8). Anions like  $\text{H}^-$ ,  $\text{Cl}^-$ ,  $\text{O}^{2-}$ , and  $\text{N}^{3-}$ , which contain single atoms with a negative charge, are called **monatomic anions**.

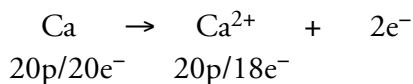
Some metallic atoms lose enough electrons to create a cation that has the same number of electrons as the nearest smaller noble gas. For example, the alkali metals in group 1 all have one more electron than the nearest noble gas. When they react with nonmetallic atoms, they lose one electron and form  $+1$  ions (Figure 5.9 on page 184). For example, sodium has atoms with 11 protons and 11 electrons. If an atom of sodium loses one electron, it will have ten electrons like uncharged neon. With a  $-10$



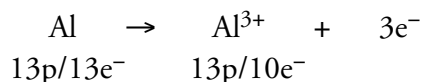
charge from the electrons and a +11 charge from the protons, the sodium ions have a +1 overall charge. The symbol for this cation is  $\text{Na}^+$ . Note that the charge is indicated with a “+” instead of “+1” or “1+”.



The alkaline earth metals in group 2 all have two more electrons than the nearest noble gas. When they react with nonmetallic atoms, they tend to lose two electrons and form +2 ions (Figure 5.9 on the next page). For example, calcium has atoms with 20 protons and 20 electrons. Each calcium atom can lose two electrons to achieve 18, the same number as its nearest noble gas, argon. The symbol for this cation is  $\text{Ca}^{2+}$ . Note that the charge is indicated with “2+” not “+2”.



Aluminum atoms and the atoms of the group 3 metals have three more electrons than the nearest noble gas. When they react with nonmetallic atoms, they tend to lose three electrons and form +3 ions (Figure 5.9 on the next page). For example, uncharged aluminum atoms have 13 protons and 13 electrons. Each aluminum atom can lose three electrons to achieve ten, like neon, forming an  $\text{Al}^{3+}$  cation.



Cations like  $\text{Na}^+$ ,  $\text{Ca}^{2+}$ , and  $\text{Al}^{3+}$ , which are single atoms with a positive charge, are called **monatomic cations**.

OBJECTIVE 11

OBJECTIVE 11

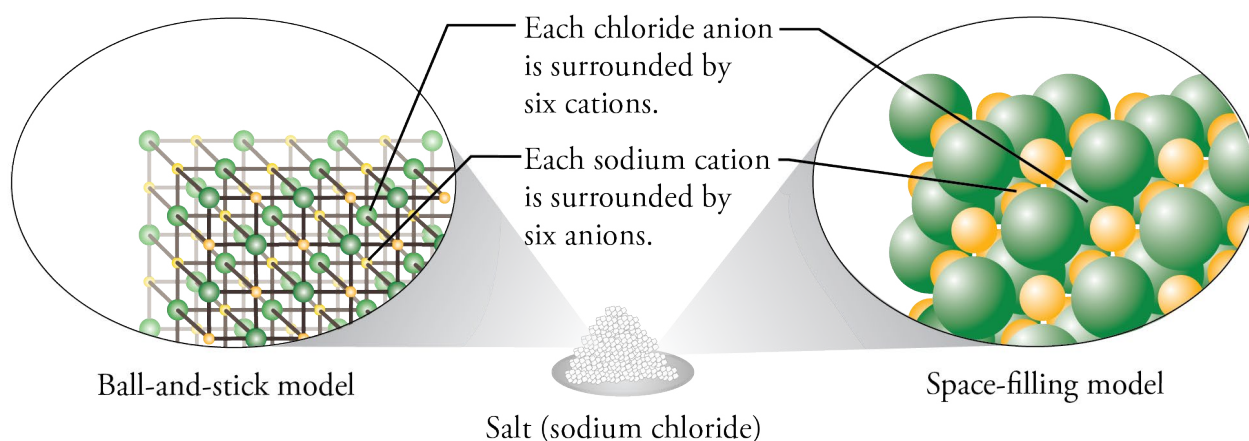


Several of the monatomic cations play important roles in our bodies. For example, we need calcium ions in our diet for making bones and teeth. Iron(II) ions are found in hemoglobin molecules in red blood cells that carry oxygen from our lungs to the tissues of our bodies. Potassium, sodium, and chloride ions play a crucial role in the transfer of information between nerve cells. Enzymes (chemicals in the body that increase the speed of chemical reactions) often contain metallic cations, such as manganese(II) ions, iron(III) ions, copper(II) ions, and zinc ions. For example,  $\text{Zn}^{2+}$  ions are in the center of the enzyme alcohol dehydrogenase, which is the enzyme in our livers that accelerates the breakdown of the ethanol consumed in alcoholic beverages.

## Structure of Ionic Compounds

Figure 5.11 shows the solid structure of the ionic compound sodium chloride,  $\text{NaCl}$ . We have already seen that the particles that form the structure of ionic compounds are cations and anions, and the attractions that hold them together are ionic bonds. When atoms gain electrons and form anions, they get larger. When atoms lose electrons and form cations, they get significantly smaller. Thus the chloride ions are larger than the sodium ions. The ions take the arrangement that provides the greatest cation-anion attraction while minimizing the anion-anion and cation-cation repulsions. Each sodium ion is surrounded by six chloride ions, and each chloride ion is surrounded by six sodium ions.

Any ionic compound that has the same arrangement of cations and anions as  $\text{NaCl}$  is said to have the sodium chloride crystal structure. The ionic compounds in this category include  $\text{AgF}$ ,  $\text{AgCl}$ ,  $\text{AgBr}$ , and the oxides and sulfides of the alkaline earth metals, such as  $\text{MgO}$ ,  $\text{CaS}$ , etc. The sodium chloride crystal structure is just one of many different possible arrangements of ions in solid ionic compounds.



OBJECTIVE 12

**Figure 5.11**  
Sodium Chloride Structure

## Polyatomic Ions

When an electric current is run through a purified saltwater solution (brine), hydrogen gas, chlorine gas, and an ionic compound called sodium hydroxide,  $\text{NaOH}$ , form. The sodium hydroxide, commonly called caustic soda or lye, is a very important compound that is used in paper production, vegetable oil refining, and to make many different compounds, such as soap and rayon. Like sodium chloride,  $\text{NaCl}$ , sodium

hydroxide, NaOH, contains a cation and an anion, but unlike the monatomic  $\text{Cl}^-$  anion in NaCl, the hydroxide ion,  $\text{OH}^-$ , in NaOH is a **polyatomic ion**, a charged *collection* of atoms held together by covalent bonds.

## OBJECTIVE 13

The ammonium ion can take the place of a monatomic cation in an ionic crystal structure. For example, the crystal structure of ammonium chloride,  $\text{NH}_4\text{Cl}$ , which is found in fertilizers, is very similar to the crystal structure of cesium chloride,  $\text{CsCl}$ , which is used in brewing, mineral waters, and to make fluorescent screens. In each structure, the chloride ions form a cubic arrangement with chloride ions at the corners of each cube. In cesium chloride, the cesium ions sit in the center of each cube, surrounded by eight chloride ions. Ammonium chloride has the same general structure as cesium chloride, with ammonium ions playing the same role in the  $\text{NH}_4\text{Cl}$  structure as cesium ions play in  $\text{CsCl}$ . The key idea is that because of its overall positive charge, the polyatomic ammonium ion acts like the monatomic cesium ion,  $\text{Cs}^+$  (Figure 5.12).

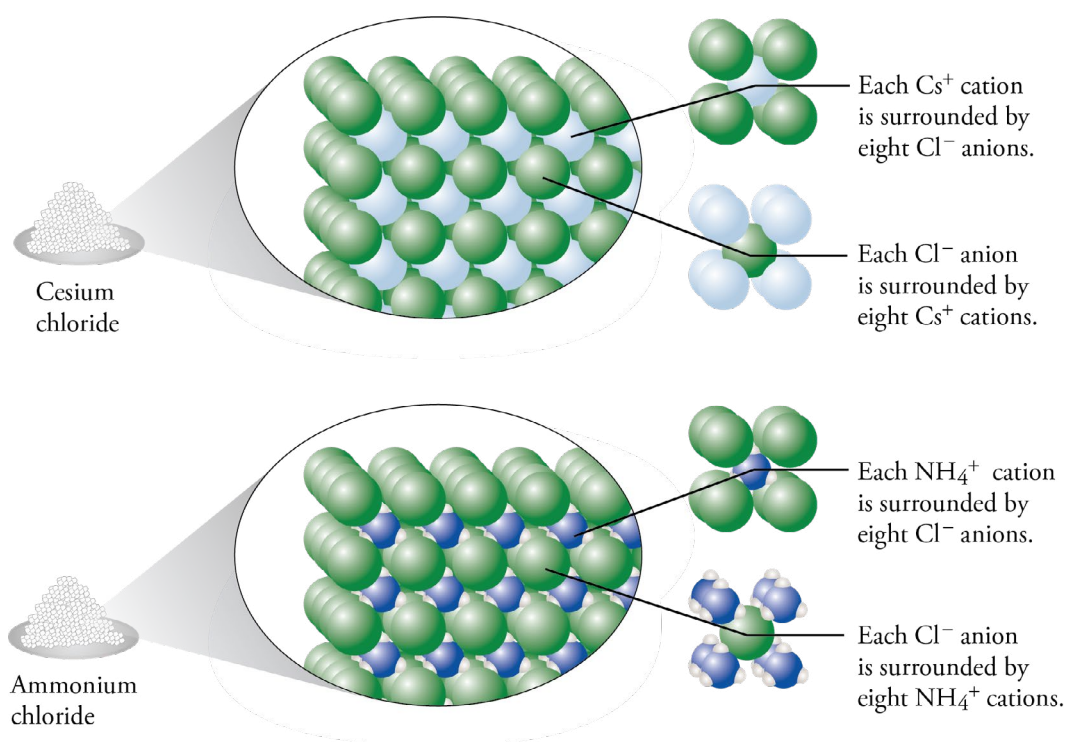


Figure 5.12

Cesium Chloride Crystal Structure

OBJECTIVE 13

There are many polyatomic anions that can take the place of monatomic anions. For example, zinc hydroxide, used as an absorbent in surgical dressings, has a similar structure to zinc chloride, which is used in embalming and taxidermist's fluids. The hydroxide ion,  $\text{OH}^-$ , plays the same role in the structure of  $\text{Zn}(\text{OH})_2$  as the chloride ion,  $\text{Cl}^-$ , plays in  $\text{ZnCl}_2$ . (Note that to show that there are two hydroxide ions for each zinc ion, the OH is in parentheses, with a subscript of 2.)



## 5.4 Molecular Compounds

It's Monday morning, and you'd like a cup of coffee, but when you try cranking up the stove to reheat yesterday's brew, nothing happens. Apparently, the city gas line has sprung a leak and been shut down for repairs. The coffee cravings are strong, so you rummage in the garage until you find that can of Sterno left over from your last camping trip. You're saved. Both Sterno and natural gas contain compounds that burn and release heat, but the compounds in each of these substances are different. Natural gas is mostly methane,  $\text{CH}_4$ , while Sterno contains several substances, including methanol,  $\text{CH}_3\text{OH}$ .

The oxygen atom in methanol molecules makes methanol's properties very different than methane's. Methane is a colorless, odorless, and tasteless gas. Methanol, or wood alcohol, is a liquid with a distinct odor, and is poisonous in very small quantities.

Chemists have discovered that part of the reason the small difference in structure leads to large differences in properties lies in the nature of covalent bonds and the arrangement of those bonds in space. The remaining sections of this chapter provide a model for explaining how covalent bonds form, teach you how to describe the resulting molecules with Lewis structures, and show how Lewis structures can be used to predict the three-dimensional geometric arrangement of atoms in molecules.

We can use what we know from Chapter 4 about the electron configurations of atoms to begin to understand why atoms form bonds as they do. To describe the formation of covalent bonds in molecules, we use a model called the valence-bond model, but before the assumptions of this model are described, let's revisit some of the important issues relating to the use of models for describing the physical world.

### The Strengths and Weaknesses of Models

When developing a model of physical reality, scientists take what they think is true and simplify it enough to make it useful. Such is the case with their description of the nature of molecules. Scientific understanding of molecular structure has advanced tremendously in the last few years, but the most sophisticated descriptions are too complex and mathematical to be understood by anyone but the most highly trained chemists and physicists. To be useful to the rest of us, the descriptions have been translated into simplified versions of what scientists consider to be true.

Such models have advantages and disadvantages. They help us to visualize, explain, and predict chemical changes, but we need to remind ourselves now and then that they are only models and that as models, they have their limitations. For example, because a model is a simplified version of what we think is true, the processes it depicts are sometimes described using the phrase *as if*. When you read, "It is *as if* an electron were promoted from one orbital to another," the phrase is a reminder that we do not necessarily think this is what really happens. We merely find it *useful* to talk about the process *as if* this is the way it happens.



Methane or methanol? We use both to heat food. How do their molecules—and properties—differ?

One characteristic of models is that they change with time. Because our models are simplifications of what we think is real, we are not surprised when they sometimes fail to explain experimental observations. When this happens, the model is altered to fit the new observations.

The valence-bond model for covalent bonds, described below, has its limitations, but it is still extremely useful. For example, you will see in Chapter 12 that it helps us understand the attractions between molecules and predict relative melting points and boiling points of substances. The model is also extremely useful in describing the mechanisms of chemical changes. Therefore, even though it strays a bit from what scientists think is the most accurate description of real molecules, the valence-bond model is the most popular model for explaining covalent bonding.

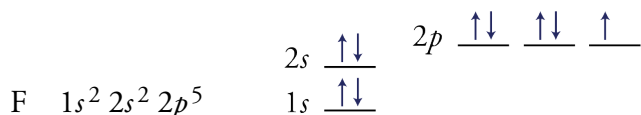
### The Valence-Bond Model

The *valence-bond model*, which is commonly used to describe the formation of covalent bonds, is based on the following assumptions:

- Only the highest-energy electrons participate in bonding.
- Covalent bonds usually form to pair unpaired electrons.

#### OBJECTIVE 15

Fluorine is our first example. Using the process described in Section 4.3, we can draw its electron configuration and orbital diagram.



The first assumption of our model states that only the highest energy electrons of fluorine atoms participate in bonding. There are two reasons why this is a reasonable assumption. First, we know from the unreactive nature of helium atoms that the  $1s^2$  electron configuration is very stable, so we assume that these electrons in a fluorine atom are less important than others are for the creation of bonds. The second reason is that the electrons in  $2s$  and  $2p$  orbitals have larger electron clouds and are therefore more available for interaction with other atoms. The  $2s^2$  and  $2p^5$  electrons are the fluorine atom's valence electrons. **Valence electrons** are the highest-energy  $s$  and  $p$  electrons in an atom.

In a similar way, we can show that for other elements, some electrons are more influential in the formation of chemical bonds than others are. Of chlorine's 17 electrons, for example, only seven are important in predicting how chlorine will bond. Of sulfur's 16 electrons, only six are important; of phosphorus's 15 electrons, only five are important. The number of valence electrons for an atom of each nonmetallic element is equal to the element's "A-group" number on the periodic table (Figure 5.13). For example, the nonmetallic elements in group 7A (F, Cl, Br, and I) have seven valence electrons, those in group 6A (O, S, and Se) have six valence electrons, those in group 5A (N and P) have five, and carbon (C) in group 4A has four.

#### OBJECTIVE 16

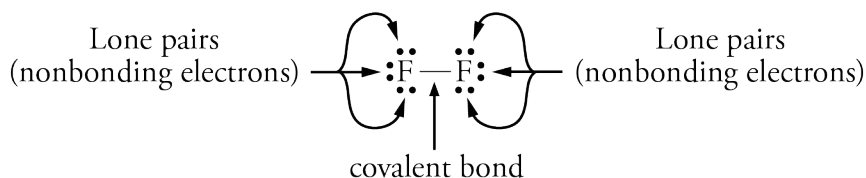


When atoms other than the noble gas atoms form bonds, they often have eight electrons around them in total. When the unpaired electron of one fluorine atom pairs with an unpaired electron of another fluorine atom to form a covalent bond, each fluorine atom has an octet of eight valence electrons around it: two from the two-electron covalent bond and six from its three lone pairs.



Usually, the covalent bonds in the electron-dot symbols for molecules are indicated with lines. The nonbonding pairs of electrons (lone pairs) are described with dots. Each atom of fluorine in a  $F_2$  molecule has 1 covalent bond and 3 lone pairs.

## OBJECTIVE 17(A)



## OBJECTIVE 18

This way of depicting a molecule or polyatomic ion—using the elements' symbols to represent atoms, using lines to show two-electron covalent bonds, and using dots to represent lone pairs—is called a **Lewis structure**. Although the bonds in Lewis structures can be described either with lines or with dots, in this text they will be described with lines.

## OBJECTIVE 19

Electron-dot symbols are derived by placing valence electrons (represented by dots) around the element's symbol. Starting on any of the four sides of the symbol, we place one dot at a time until there are up to four unpaired electrons around it. If there are more than four valence electrons for an atom, the remaining electrons are added one by one to the unpaired electrons to form up to four pairs.

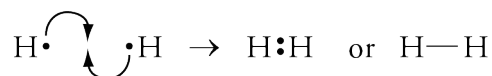


## OBJECTIVE 17(B)

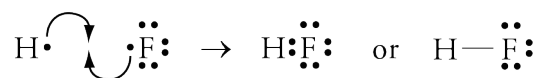
Each hydrogen atom in its ground state has one valence electron in a  $1s$  orbital. Its electron-dot symbol and orbital diagram are therefore



Because atoms become more stable when they pair their unpaired electrons, hydrogen atoms combine to form hydrogen molecules,  $H_2$ , which allow each atom to share two electrons.

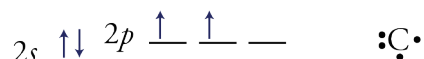


Hydrogen atoms can also combine with fluorine atoms to form HF molecules.



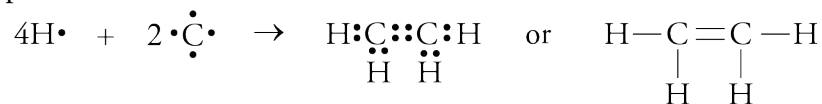
## OBJECTIVE 17(C)

Carbon is in group 4A on the periodic table, so we predict that it has four valence electrons. Looking at the orbital diagram for these electrons, we might expect carbon to form two covalent bonds (to pair its two unpaired electrons) and have one lone pair.





Carbon atoms also frequently form **double bonds**, in which they share four electrons with another atom, often another carbon atom. Double bonds are represented by double lines in a Lewis structure. Ethene (commonly called ethylene),  $C_2H_4$ , is an example.



Note that each carbon atom in  $C_2H_4$  has four bonds total, two single bonds to hydrogen atoms and two bonds to the other carbon atom.

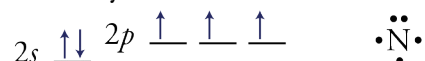
The bond between the carbon atoms in ethyne (commonly called acetylene),  $C_2H_2$ , is a **triple bond**, which can be viewed as the sharing of six electrons between two atoms.



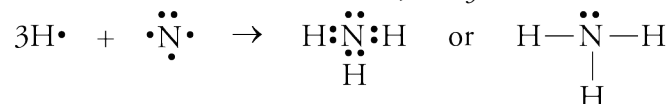
Note that each carbon atom in  $C_2H_2$  has four bonds total, one single bond to a hydrogen atom and three bonds to the other carbon atom.

#### OBJECTIVE 17(D)

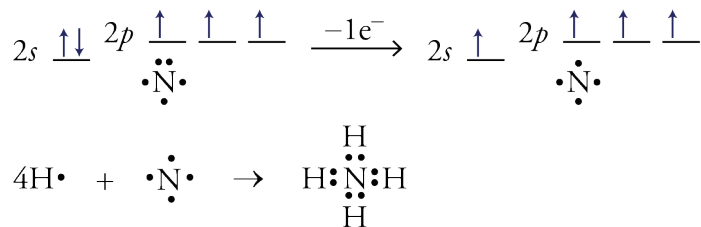
Nitrogen is in group 5A, so it has five valence electrons. Its orbital diagram and electron-dot symbol are



Each nitrogen atom has three unpaired electrons, and as the model predicts, it usually forms three covalent bonds. For example, a nitrogen atom can bond to three hydrogen atoms to form an ammonia molecule,  $NH_3$ :

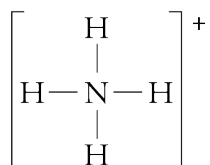


Another common bonding pattern for nitrogen atoms is four bonds with no lone pairs. The nitrogen atom in an ammonium polyatomic ion,  $NH_4^+$ , is an example. This pattern and the positive charge on the ion can be explained by the loss of one electron from the nitrogen atom. It is *as if* an uncharged nitrogen atom loses one electron from the  $2s$  orbital, leaving it with four unpaired electrons and the ability to make four bonds.

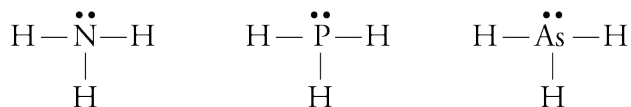


#### OBJECTIVE 17(E)

Because nitrogen must lose an electron to form this bonding pattern, the overall structure of the ammonium ion has a +1 charge. The Lewis structures of polyatomic ions are usually enclosed in brackets, with the overall charge written outside the brackets on the upper right.

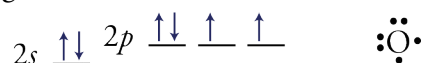


Phosphorus is in group 5A, so its atoms have five valence electrons, in this case, in the  $3s^23p^3$  configuration. Arsenic, also in group 5A, has atoms with a  $4s^24p^3$  configuration for their valence electrons. Because these valence configurations are similar to nitrogen's,  $2s^22p^3$ , the model correctly predicts that phosphorus and arsenic atoms will form bonds like nitrogen. For example, they each form three bonds to hydrogen atoms and have one lone pair in  $\text{NH}_3$ ,  $\text{PH}_3$ , and  $\text{AsH}_3$ .

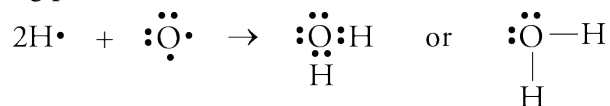


On the other hand, phosphorus and arsenic atoms exhibit bonding patterns that are not possible for nitrogen atoms. For example, molecules such as  $\text{PCl}_5$  and  $\text{AsF}_5$  have five bonds and no lone pairs. If you take other chemistry courses, you are likely to see compounds with bonding patterns like this, but because they are somewhat uncommon, you will not see them again in this text.

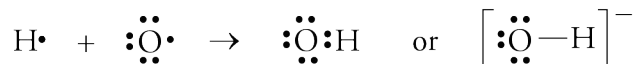
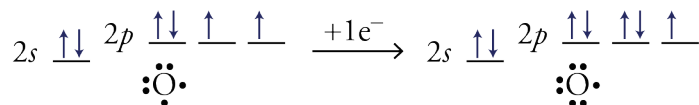
The most common bonding pattern for oxygen atoms is two covalent bonds and two lone pairs. Our model explains this in terms of the valence electrons'  $2s^22p^4$  electron configuration.



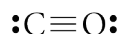
The two unpaired electrons are able to participate in two covalent bonds, and the two pairs of electrons remain two lone pairs. The oxygen atom in a water molecule has this bonding pattern.



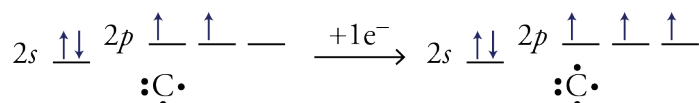
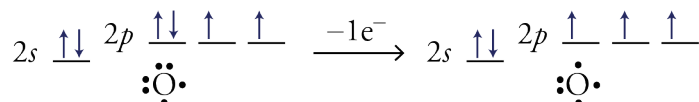
In another common bonding pattern, oxygen atoms gain one electron and form one covalent bond with three lone pairs. The oxygen atom in the hydroxide ion has this bonding pattern.



In rare circumstances, carbon and oxygen atoms can form triple bonds, leaving each atom with one lone pair. The carbon monoxide molecule,  $\text{CO}$ , is an example:



According to the valence-bond model, it is *as if* an electron is transferred from the oxygen atom to the carbon atom as the bonding in  $\text{CO}$  occurs. This gives each atom three unpaired electrons to form the triple bond, with one lone pair each left over.

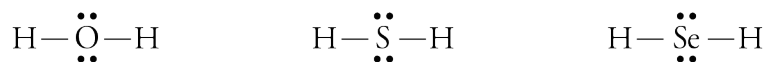


OBJECTIVE 17(F)

OBJECTIVE 17(G)

OBJECTIVE 17(H)

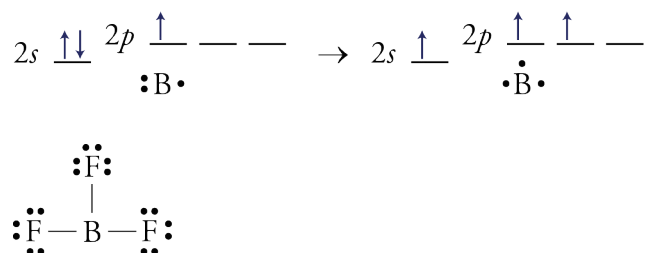
Like oxygen, sulfur and selenium are in group 6A on the periodic table, so they too have six valence electrons, with  $3s^23p^4$  and  $4s^24p^4$  electron configurations, respectively. We therefore expect sulfur atoms and selenium atoms to have bonding patterns similar to oxygen's. For example, they all commonly form two bonds and have two lone pairs, as in molecules such as  $H_2O$ ,  $H_2S$ , and  $H_2Se$ .



Sulfur and selenium atoms have additional bonding patterns that are not possible for oxygen atoms. For example, they can form six bonds in molecules such as  $SF_6$  and  $SeF_6$ . You will not see these somewhat uncommon bonding patterns again in this text.

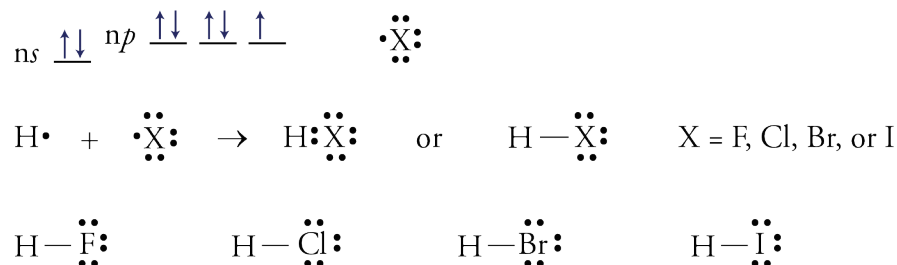
## OBJECTIVE 17(i)

When the three equivalent B–F bonds form in boron trifluoride,  $BF_3$ , it is *as if* one of the boron atom's valence electrons were promoted from its  $2s$  orbital to an empty  $2p$  orbital. This leaves three unpaired electrons to form three covalent bonds.



## OBJECTIVE 17(j)

The elements in group 7A all have the  $ns^2np^5$  configuration for their valence electrons ( $3s^23p^5$  for chlorine,  $4s^24p^5$  for bromine, etc.). Thus they all commonly form one covalent bond and have three lone pairs. For example, their atoms all form one bond to a hydrogen atom to form HF, HCl, HBr, and HI.



Chlorine, bromine, and iodine atoms have additional, less common bonding patterns that you might see in other chemistry courses.

Table 5.1 summarizes the bonding patterns described in this section. The patterns listed there are not the only possible ones that these elements can have, but any patterns not listed are rare. In Section 5.5, you will be asked to draw Lewis structures from formulas, and knowledge of the common bonding patterns will help you to propose structures and evaluate their stability.



**Table 5.1** Covalent Bonding Patterns

OBJECTIVE 20

| Element        | Frequency of pattern | Number of bonds | Number of lone pairs | Example                      |
|----------------|----------------------|-----------------|----------------------|------------------------------|
| H              | always               | 1               | 0                    | H–                           |
| B              | most common          | 3               | 0                    | –B–<br>                      |
| C              | most common          | 4               | 0                    | –C– or –C= or =C= or –C≡<br> |
|                | rare                 | 3               | 1                    | ≡C:                          |
| N, P, & As     | most common          | 3               | 1                    | –N̈– or –N̈= or ≡N̈:         |
|                | common               | 4               | 0                    | –N̈–<br>                     |
| O, S, & Se     | most common          | 2               | 2                    | –Ö– or Ö<br>               |
|                | common               | 1               | 3                    | –Ö:                         |
|                | rare                 | 3               | 1                    | ≡Ö:                         |
| F, Cl, Br, & I | most common          | 1               | 3                    | –Ẍ:                         |

## 5.5 Drawing Lewis Structures

After studying chemistry all morning, you go out to mow the lawn. While adding gasoline to the lawnmower's tank, you spill a bit, so you go off to get some soap and water to clean it up. By the time you get back, the gasoline has all evaporated, which starts you wondering...why does gasoline evaporate so much faster than water? We are not yet ready to explain this, but part of the answer is found by comparing the substances' molecular structures and shapes. Lewis structures provide this information. You will see in Chapter 12 that the ability to draw Lewis structures for the chemical formulas of water,  $\text{H}_2\text{O}$ , and hexane,  $\text{C}_6\text{H}_{14}$ , (one of the major components of gasoline) will help you to explain their relative rates of evaporation. The ability to draw Lewis structures will be important for many other purposes as well, including explaining why the soap would have helped clean up the spill if the gasoline had not evaporated so quickly.

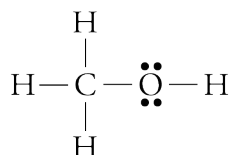
### Simple Procedure

There are two general techniques for drawing Lewis structures. Because atoms of the nonmetallic elements usually have their most common bonding pattern, the simplest technique for drawing Lewis structures is to attempt to give each atom its most common bonding pattern as listed on Table 5.1.

OBJECTIVE 21

OBJECTIVE 22

To illustrate how Lewis structures can be drawn using the information on Table 5.1, let's figure out the Lewis structure of methanol,  $\text{CH}_3\text{OH}$ , which is often called methyl alcohol or wood alcohol. Methanol is a poisonous liquid that is used as a solvent. When drawing its Lewis structure, we assume that the carbon atom will have four bonds (represented by four lines), the oxygen atom will have two bonds and two lone pairs, and each hydrogen atom will have one bond. The Lewis structure below meets these criteria.

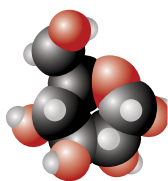


OBJECTIVE 23

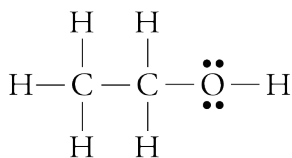
Methanol,  $\text{CH}_3\text{OH}$ 

OBJECTIVE 24

(methyl alcohol)

Web  
Molecules

Methanol is an alcohol, which is a category of organic compounds, not just the intoxicating compound in certain drinks. **Alcohols** are organic compounds that possess one or more  $-\text{OH}$  groups attached to a hydrocarbon group (a group that contains only carbon and hydrogen). Ethanol,  $\text{C}_2\text{H}_5\text{OH}$ , is the alcohol in alcoholic beverages (see Special Topic 5.1: *Molecular Shapes, Intoxicating Liquids, and the Brain*), while the alcohol in rubbing alcohol is usually 2-propanol (Figure 5.15). These alcohols are also called ethyl alcohol,  $\text{C}_2\text{H}_5\text{OH}$ , and isopropyl alcohol,  $\text{C}_3\text{H}_7\text{OH}$ .

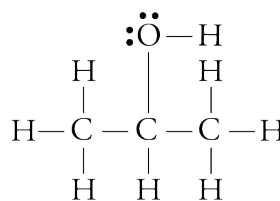


OBJECTIVE 23

Ethanol,  $\text{C}_2\text{H}_5\text{OH}$ 

OBJECTIVE 24

(ethyl alcohol)

2-propanol,  $\text{C}_3\text{H}_7\text{OH}$ 

(isopropyl alcohol)

Check to see that each of these compounds follows our guidelines for drawing Lewis structures.

**Figure 5.15**  
Products Containing Alcohols



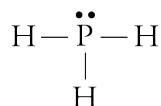
### EXAMPLE 5.3 - Drawing Lewis Structures from Formulas

Draw a Lewis structure for each of the following formulas:

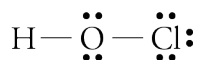
- phosphine,  $\text{PH}_3$  (used to make semiconductors)
- hypochlorous acid,  $\text{HOCl}$  (used to bleach textiles)
- CFC-11,  $\text{CCl}_3\text{F}$  (used as a refrigerant)
- $\text{C}_2\text{H}_2$ , acetylene (burned in oxyacetylene torches)

*Solution*

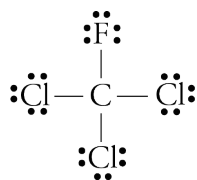
- Phosphorus atoms usually have three covalent bonds and one lone pair, and hydrogen atoms have one covalent bond and no lone pairs. The following Lewis structure for  $\text{PH}_3$  gives each of these atoms its most common bonding pattern.



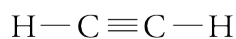
- Hydrogen atoms have one covalent bond and no lone pairs, oxygen atoms usually have two covalent bonds and two lone pairs, and chlorine atoms usually have one covalent bond and three lone pairs.



- Carbon atoms usually have four covalent bonds and no lone pairs. Fluorine and chlorine atoms usually have one covalent bond and three lone pairs. The fluorine atom can be put in any of the four positions around the carbon atom.



- Carbon atoms form four bonds with no lone pairs, and hydrogen atoms form one bond with no lone pairs. To achieve these bonding patterns, there must be a triple bond between the carbon atoms.



OBJECTIVE 21

### EXERCISE 5.3 - Drawing Lewis Structures from Formulas

Draw a Lewis structure for each of the following formulas:

- nitrogen triiodide,  $\text{NI}_3$  (explodes at the slightest touch)
- hexachloroethane,  $\text{C}_2\text{Cl}_6$  (used to make explosives)
- hydrogen peroxide,  $\text{H}_2\text{O}_2$  (a common antiseptic)
- ethylene (or ethene),  $\text{C}_2\text{H}_4$  (used to make polyethylene)

OBJECTIVE 21

## General Procedure

The shortcut described above works well for many simple uncharged molecules, but it does not work reliably for molecules that are more complex or for polyatomic ions. To draw Lewis structures for these, you can use the stepwise procedure described in the following sample study sheet.

### Sample Study Sheet 5.2

#### Drawing Lewis Structures from Formulas

##### OBJECTIVE 22

**TIP-OFF** In this chapter, you may be given a chemical formula for a molecule or polyatomic ion and asked to draw a Lewis structure, but there are other, more subtle tip-offs that you will see in later chapters.

**GENERAL STEPS** SEE FIGURE 5.16 FOR A SUMMARY OF THESE STEPS.

**STEP 1** Determine the total number of valence electrons for the molecule or polyatomic ion. (Remember that the number of valence electrons for a representative element is equal to its group number, using the A-group convention for numbering groups. For example, chlorine, Cl, is in group 7A, so it has seven valence electrons. Hydrogen has one valence electron.)

- For uncharged molecules, the total number of valence electrons is the sum of the valence electrons of each atom.
- For polyatomic cations, the total number of valence electrons is the sum of the valence electrons for each atom minus the charge.
- For polyatomic anions, the total number of valence electrons is the sum of the valence electrons for each atom plus the charge.

**STEP 2** Draw a reasonable skeletal structure, using single bonds to join all the atoms. One or more of the following guidelines might help with this step. (They are clarified in the examples that follow.)

- Try to arrange the atoms to yield the most typical number of bonds for each atom. Table 5.1 lists the most common bonding patterns for the nonmetallic elements.
- Apply the following guidelines in deciding what element belongs in the center of your structure.
  - Hydrogen and fluorine atoms are never in the center.
  - Oxygen atoms are rarely in the center.
  - The element with the fewest atoms in the formula is often in the center.
  - The atom that is capable of making the most bonds is often in the center.
- Oxygen atoms rarely bond to other oxygen atoms.
- The molecular formula often reflects the molecular structure. (See Example 5.7.)
- Carbon atoms commonly bond to other carbon atoms.

**STEP 3** Subtract two electrons from the total for each of the single bonds (lines) described in Step 2 above. This tells us the number of electrons that still need to be distributed.

**STEP 4** Try to distribute the remaining electrons as lone pairs to obtain a total of eight electrons around each atom except hydrogen and boron. The atoms in reasonable Lewis structures are often surrounded by an octet of electrons. The following are some helpful observations pertaining to octets.

- In a reasonable Lewis structure, carbon, nitrogen, oxygen, and fluorine always have eight electrons around them.
- Hydrogen will always have a total of two electrons from its one bond.
- Boron can have fewer than eight electrons but never more than eight.
- The nonmetallic elements in periods beyond the second period (P, S, Cl, Se, Br, and I) usually have eight electrons around them, but they can have more. (In this text, they will always have eight electrons around them, but if you go on to take other chemistry courses, it will be useful to know that they can have more.)
- The bonding properties of the metalloids arsenic, As, and tellurium, Te, are similar to those of phosphorus, P, and sulfur, S, so they usually have eight electrons around them but can have more.

**STEP 5** Do one of the following.

- If in Step 4 you were able to obtain an octet of electrons around each atom other than hydrogen and boron, and if you used all of the remaining valence electrons, go to step 6.
- If you have electrons remaining after each of the atoms other than hydrogen and boron have their octet, you can put more than eight electrons around elements in periods beyond the second period. (You will not need to use this procedure for any of the structures in this text, but if you take more advanced chemistry courses, it will be useful.)
- If you do not have enough electrons to obtain octets of electrons around each atom (other than hydrogen and boron), convert one lone pair into a multiple bond for each two electrons that you are short. (See Example 5.5.)

If you would need two more electrons to get octets, convert one lone pair in your structure to a second bond between two atoms.

If you would need four more electrons to get octets, convert two lone pairs into bonds. This could mean creating a second bond in two different places or creating a triple bond in one place.

If you would need six more electrons to get octets, convert three lone pairs into bonds.

Etc.

**STEP 6** Check your structure to see if all of the atoms have their most common bonding pattern (Table 5.1).

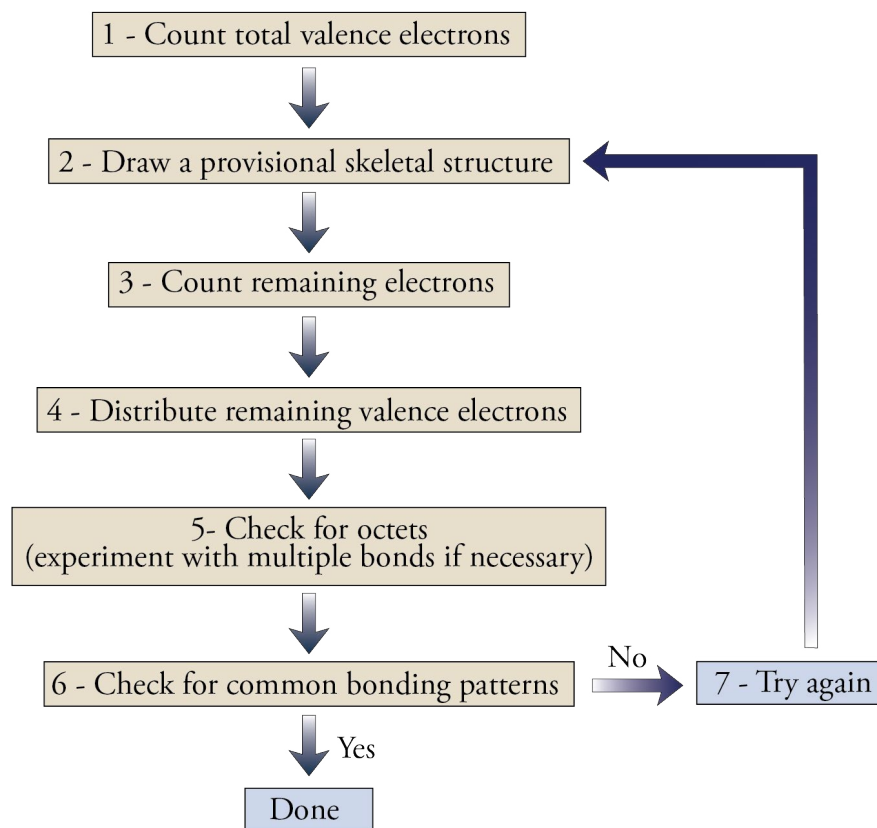
- If each atom has its most common bonding pattern, your structure is a reasonable structure. Skip Step 7.
- If one or more atoms are without their most common bonding pattern, continue to Step 7.

**STEP 7** If necessary, try to rearrange your structure to give each atom its most common bonding pattern. One way to do this is to return to Step 2 and try another skeleton. (This step is unnecessary if all of the atoms in your structure have their most common bonding pattern.)

**EXAMPLE** See Examples 5.4 to 5.7.

**Figure 5.16**  
Lewis Structure Procedure

## OBJECTIVE 22



## EXAMPLE 5.4 - Drawing Lewis Structures

## OBJECTIVE 22

Draw a reasonable Lewis structure for methyl bromide,  $\text{CH}_3\text{Br}$ , which is an ozone depleting gas used as a fumigant.

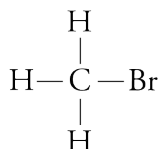
*Solution*

Let's start with the stepwise procedure for drawing Lewis structures.

**STEP 1** To determine the number of valence electrons for  $\text{CH}_3\text{Br}$ , we note that carbon is in group 4A, so its atoms have four valence electrons; hydrogen has one valence electron; bromine is in group 7A, so its atoms have seven valence electrons.

$$\text{CH}_3\text{Br} \quad \text{Number of valence } e^- = \overset{\text{C}}{1(4)} + \overset{\text{H}}{3(1)} + \overset{\text{Br}}{1(7)} = 14$$

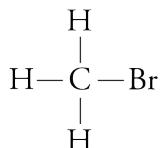
**STEP 2** Before setting up the skeleton for  $\text{CH}_3\text{Br}$ , we check Table 5.1, which reminds us that carbon atoms usually have four bonds, hydrogen atoms always have one bond, and bromine atoms most commonly have one bond. Thus the following skeleton is most reasonable.



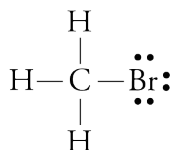
**STEP 3** We started with 14 total valence electrons for  $\text{CH}_3\text{Br}$ , and we have used eight of them for the four bonds in the skeleton we drew in Step 2.

$$\begin{aligned} \text{Number } e^- \text{ remaining} &= \text{total valence } e^- - \text{number of bonds} \left( \frac{2 e^-}{1 \text{ bond}} \right) \\ &= 14 - 4(2) = 6 \end{aligned}$$

**STEP 4** After Step 3, we have the following skeleton for  $\text{CH}_3\text{Br}$  and six valence electrons still to distribute.

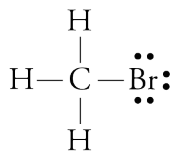


Hydrogen atoms never have lone pairs, and the carbon atom has an octet of electrons around it from its four bonds. In contrast, the bromine atom needs six more electrons to obtain an octet, so we put the remaining six electrons around the bromine atom as three lone pairs.



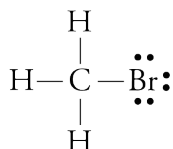
**STEP 5** The structure drawn in Step 4 for  $\text{CH}_3\text{Br}$  has an octet of electrons around the carbon and bromine atoms. Hydrogen has its one bond. We have also used all of the valence electrons. Thus we move on to Step 6.

**STEP 6** All of the atoms in the structure drawn for  $\text{CH}_3\text{Br}$  have their most common bonding pattern, so we have a reasonable Lewis structure.



**STEP 7** Because the atoms in our structure for  $\text{CH}_3\text{Br}$  have their most common bonding pattern, we skip this step.

**Shortcut** The shortcut to drawing Lewis structures described at the beginning of this section can often be used for uncharged molecules such as  $\text{CH}_3\text{Br}$ . Carbon atoms usually have four bonds and no lone pairs, hydrogen atoms always have one bond, and bromine atoms most commonly have one bond and three lone pairs. The only way to give these atoms their most common bonding patterns is with the following Lewis structure, which is the same Lewis structure we arrived at with the stepwise procedure.



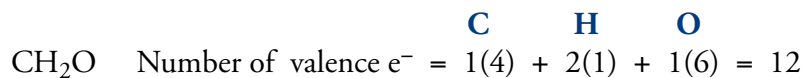
## EXAMPLE 5.5 - Drawing Lewis Structures

## OBJECTIVE 22

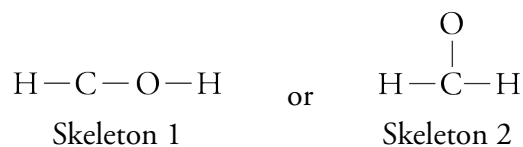
Formaldehyde,  $\text{CH}_2\text{O}$ , has many uses, including the preservation of biological specimens. Draw a reasonable Lewis structure for formaldehyde.

*Solution*

**STEP 1** Carbon is in group 4A, so its atoms have four valence electrons. Hydrogen has one valence electron. Oxygen is in group 6A, so its atoms have six valence electrons.

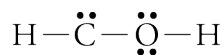


**STEP 2** There are two possible skeletons for this structure. We will work with skeleton 1 first.

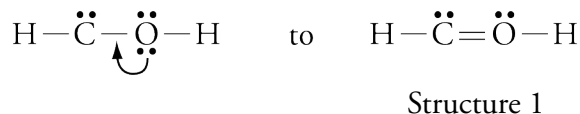


**STEP 3** (Skeleton 1) Number of  $e^-$  remaining =  $12 - 3(2) = 6$

**STEP 4** (Skeleton 1) The most common bonding pattern for oxygen atoms is two bonds and two lone pairs. It is possible for carbon atoms to have one lone pair, although that is rare. Thus we might use our six remaining electrons as two lone pairs on the oxygen atom and one lone pair on the carbon atom.



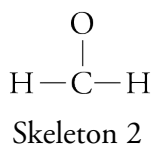
**STEP 5** (Skeleton 1) The structure above leaves the carbon atom without its octet. Because we are short two electrons for octets, we convert one lone pair to another bond. Converting the lone pair on the carbon to a multiple bond would still leave the carbon with only six electrons around it and would put ten electrons around the oxygen atom. Carbon and oxygen atoms always have eight electrons around them in a reasonable Lewis structure. Thus we instead try converting one of the lone pairs on the oxygen atom to another C–O bond.



**STEP 6** (Skeleton 1) In structure 1, the carbon atom and the oxygen atom have rare bonding patterns. This suggests that there might be a better way to arrange the atoms. We proceed to Step 7.

**STEP 7** To attempt to give each atom its most common bonding pattern, we return to Step 2 and try another skeleton.

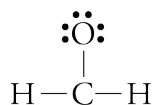
**STEP 2** (Skeleton 2) The only alternative to skeleton 1 for  $\text{CH}_2\text{O}$  is



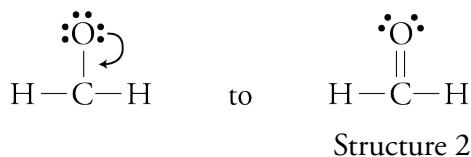
**STEP 3** (Skeleton 2) Number of  $e^-$  remaining =  $12 - 3(2) = 6$



**STEP 4** (Skeleton 2) Oxygen atoms commonly have one bond and three lone pairs, so we might use our remaining electrons as three lone pairs on the oxygen atom.



**STEP 5** (Skeleton 2) In Step 4, we used all of the remaining electrons but left the carbon atom with only six electrons around it. Because we are short two electrons needed to obtain octets of electrons around each atom, we convert one lone pair into another bond.



**STEP 6** (Skeleton 2) All of the atoms have their most common bonding pattern, so we have a reasonable Lewis structure.

**STEP 7** (Skeleton 2) Because each atom in Structure 2 has its most common bonding pattern, we skip this step.

Because Structure 2 is the only arrangement that gives each atom its most common bonding pattern, we could have skipped the stepwise procedure and used the shortcut for this molecule.

### EXAMPLE 5.6 - Drawing Lewis Structures

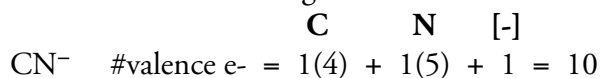
The cyanide polyatomic ion,  $\text{CN}^-$  is similar in structure to carbon monoxide,  $\text{CO}$ . Although they work by different mechanisms, they are both poisons that can disrupt the use of oxygen,  $\text{O}_2$ , in organisms. Draw a reasonable Lewis structure for the cyanide ion.

OBJECTIVE 22

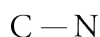
#### Solution

The shortcut does not work for polyatomic ions, so we use the stepwise procedure for  $\text{CN}^-$ .

**STEP 1** Carbon is in group 4A, so its atoms have four valence electrons. Nitrogen is in group 5A, so its atoms have five valence electrons. Remember to add one electron for the  $-1$  charge.

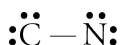


**STEP 2** There is only one way to arrange two atoms.

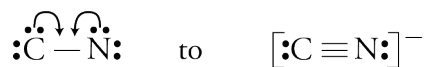


**STEP 3** Number  $e^-$  remaining =  $10 - 1(2) = 8$

**STEP 4** Both the carbon atom and the nitrogen atom need six more electrons. We do not have enough electrons to provide octets by the formation of lone pairs. You will see in the next step that we will make up for this lack of electrons with multiple bonds. For now, we might put two lone pairs on each atom.

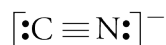


**STEP 5** Because we are short four electrons (or two pairs) in order to obtain octets, we convert two lone pairs into bonds. If we convert one lone pair from each atom into another bond, we get octets of electrons around each atom.



**STEP 6** The nitrogen atom now has its most common bonding pattern, three bonds and one lone pair. The carbon atom has a rare bonding pattern, so we proceed to Step 7.

**STEP 7** There is no other way to arrange the structure and still have an octet of electrons around each atom, so the Lewis structure for  $\text{CN}^-$  is



Remember to put the Lewis structures for polyatomic ions in brackets and show the charge on the outside upper right.

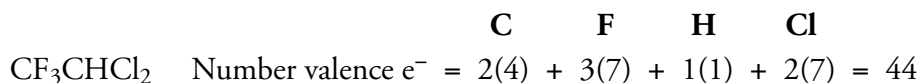
### EXAMPLE 5.7 - Drawing Lewis Structures

#### OBJECTIVE 22

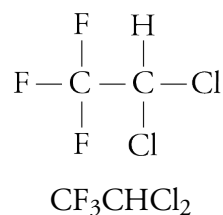
Draw a reasonable Lewis structure for  $\text{CF}_3\text{CHCl}_2$ , the molecular formula for HCFC-123, which is one of the hydrochlorofluorocarbons used as a replacement for more damaging chlorofluorocarbons.

#### Solution

**STEP 1** Carbon is in group 4A, so its atoms have four valence electrons. Chlorine and fluorine are in group 7A, so their atoms have seven valence electrons. Hydrogen has one valence electron.

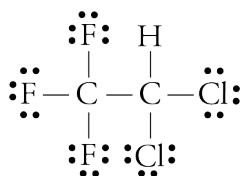


**STEP 2** The best way to start our skeleton is to remember that carbon atoms often bond to other carbon atoms. Thus we link the two carbon atoms together in the center of the skeleton. We expect hydrogen, fluorine, and chlorine atoms to form one bond, so we attach all of them to the carbon atoms. There are many ways that they could be arranged, but the way the formula has been written (especially the separation of the two carbons) gives us clues:  $\text{CF}_3\text{CHCl}_2$ , tells us that the three fluorine atoms are on one carbon atom, and the hydrogen and chlorine atoms are on the second carbon.



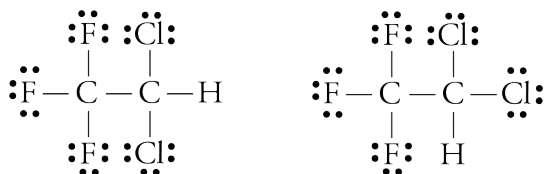
**STEP 3** Number of  $e^-$  remaining =  $44 - 7(2) = 30$

**STEP 4** We expect halogen atoms to have three lone pairs, so we can use the 30 remaining electrons to give us three lone pairs for each fluorine and chlorine atom.



Structure 1

We could also represent  $\text{CF}_3\text{CHCl}_2$ , with the following Lewis structures:



Structure 2

Structure 3

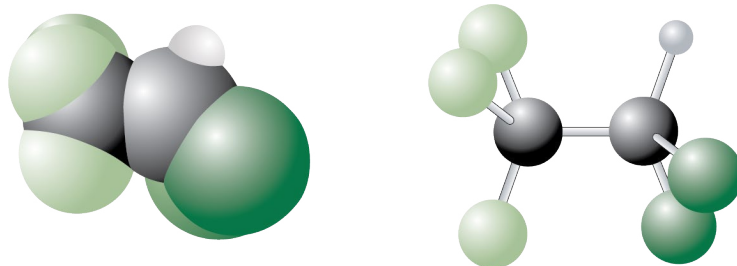
Structures 2 and 3 might look like they represent different molecules, but they actually represent the same molecule as Structure 1. To confirm that this is true, picture yourself sitting on the hydrogen atom in either the space filling or the ball-and-stick model shown in Figure 5.17 for  $\text{CF}_3\text{CHCl}_2$ . Atoms connected to each other by single bonds, such as the two carbon atoms in this molecule, are constantly rotating with respect to each other. Thus your hydrogen atom is sometimes turned toward the top of the molecule (somewhat like in Structure 1), sometimes toward the bottom of the structure (somewhat like Structure 3), and sometimes in one of the many possible positions in between. (Section 5.7 describes how you can predict molecular shapes. If you do not see at this point why Structures 1, 2, and 3 all represent the same molecule, you might return to this example after reading that section.)

**STEP 5** We have used all of the valence electrons, and we have obtained octets of electrons around each atom other than hydrogen. Thus we move to Step 6.

**STEP 6** All of the atoms in our structure have their most common bonding pattern, so we have a reasonable Lewis structure.

**STEP 7** Because each atom in our structure has its most common bonding pattern, we skip this step.

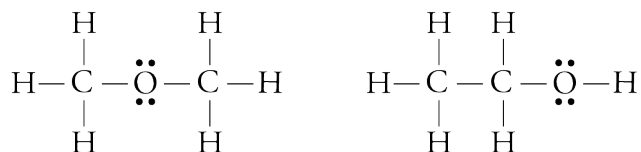
**Figure 5.17**  
Models of  $\text{CF}_3\text{CHCl}_2$



## More Than One Possible Structure

### OBJECTIVE 22

It is possible to generate more than one reasonable Lewis structure for some formulas. When this happens, remember that the more common the bonding pattern is (as summarized in Table 5.1), the more stable the structure. Even after you apply this criterion, you may still be left with two or more Lewis structures that are equally reasonable. For example, the following Lewis structures for  $C_2H_6O$  both have the most common bonding pattern for all of their atoms.



In fact, both these structures describe actual compounds. The first structure is dimethyl ether, and the second is ethanol. Substances that have the same molecular formula but different structural formulas are called **isomers**. We can write the formulas for these two isomers so as to distinguish between them:  $CH_3OCH_3$  represents dimethyl ether, and  $CH_3CH_2OH$  represents ethanol.

### EXAMPLE 5.8 - Drawing Lewis Structures

### OBJECTIVE 22

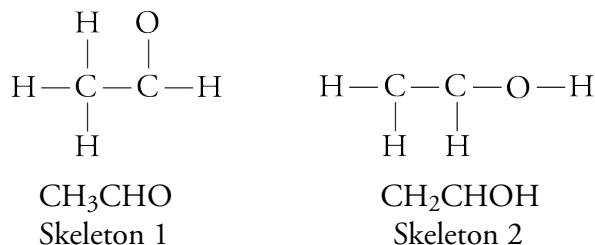
Acetaldehyde can be converted into the sedative chloral hydrate (the “Mickey Finn” or knockout drops often mentioned in detective stories). In the first step of the reaction that forms chloral hydrate, acetaldehyde,  $CH_3CHO$ , changes to its isomer,  $CH_2CHOH$ . Draw a reasonable Lewis structures for each of these isomers.

#### Solution

**STEP 1** Both molecules have the molecular formula  $C_2H_4O$ .

$$\begin{array}{ccc}
 \text{C} & \text{H} & \text{O} \\
 \text{Number valence } e^- = & 2(4) + & 4(1) + & 1(6) = 18
 \end{array}$$

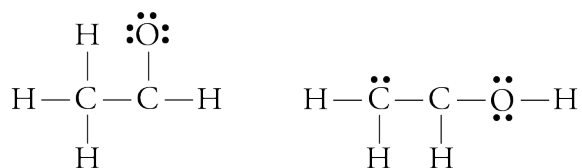
**STEP 2** Because we expect carbon atoms to bond to other carbon atoms, we can link the two carbon atoms together to start each skeleton, which is then completed to try to match each formula given.



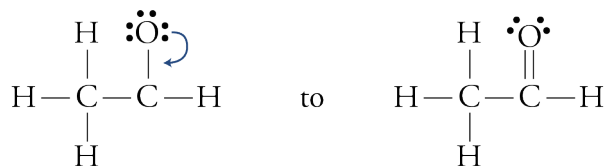
**STEP 3** Number of  $e^-$  remaining =  $18 - 6(2) = 6$

**STEP 4** For Skeleton 1, we can add three lone pairs to the oxygen atom to give it its octet. For Skeleton 2, we can add two lone pairs to the oxygen atom to get its most common bonding pattern of two bonds and two lone pairs. We can place

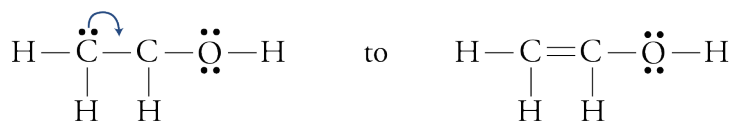
the remaining two electrons on one of the carbon atoms as a lone pair.



**STEP 5** In each case, we have one carbon atom with only six electrons around it, so we convert one lone pair into another bond.



Structure 1



Structure 2

**STEP 6** All of the atoms in both structures have their most common bonding pattern, so we have two reasonable Lewis structures representing isomers. Structure 1 is acetaldehyde (or ethanal) and structure 2 is ethenol.

**STEP 7** Because each atom in our structures has its most common bonding pattern, we skip this step.

### EXERCISE 5.4 - Lewis Structures

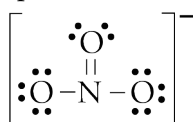
Draw a reasonable Lewis structure for each of the following formulas.

- |                            |                           |
|----------------------------|---------------------------|
| a. $\text{CCl}_4$          | f. $\text{N}_2\text{H}_4$ |
| b. $\text{Cl}_2\text{O}$   | g. $\text{H}_2\text{O}_2$ |
| c. $\text{COF}_2$          | h. $\text{NH}_2\text{OH}$ |
| d. $\text{C}_2\text{Cl}_6$ | i. $\text{NCl}_3$         |
| e. $\text{BCl}_3$          |                           |

OBJECTIVE 22

## 5.6 Resonance

As the valence-bond model was being developed, chemists came to recognize that it did not always describe all molecules and polyatomic ions adequately. For example, if we follow the procedures in Section 5.5 to draw a Lewis structure for the nitrate ion,  $\text{NO}_3^-$ , we would obtain a structure that chemists have discovered is not an accurate description of the ion's bonds:

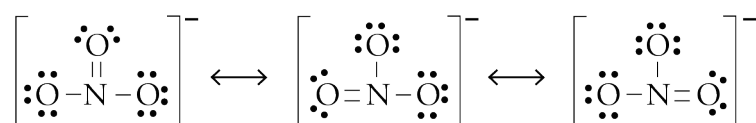


This Lewis structure shows two different types of bonds, single and double. Because more energy is required to break a double bond than to break a single bond, we say that a double bond is stronger than a single bond. Double bonds also have a shorter bond length (the distance between the nuclei of the two atoms in the bond) than single bonds do. Thus, if the above Lewis structure for nitrate were correct, the nitrate polyatomic ion would have one bond that is shorter and stronger than the other two.

This is not the case. Laboratory analyses show all three of the bonds in the nitrate ion to be the same strength and the same length. Interestingly, analysis of the bonds suggests they are longer than double bonds and shorter than single bonds. They are also stronger than single bonds but not as strong as double bonds. In order to explain how these characteristics are possible for the nitrate ion and for molecules and polyatomic ions like it, the valence-bond model had to be expanded.

## OBJECTIVE 25

The model now allows us to view certain molecules and polyatomic ions *as if* they were able to resonate—to switch back and forth—between two or more different structures. For example, the nitrate ion can be viewed *as if* it resonates between the three different structures below. Each of these structures is called a **resonance structure**. The hypothetical switching from one resonance structure to another is called **resonance**. In chemical notation, the convention is to separate the resonance structures with double headed arrows.



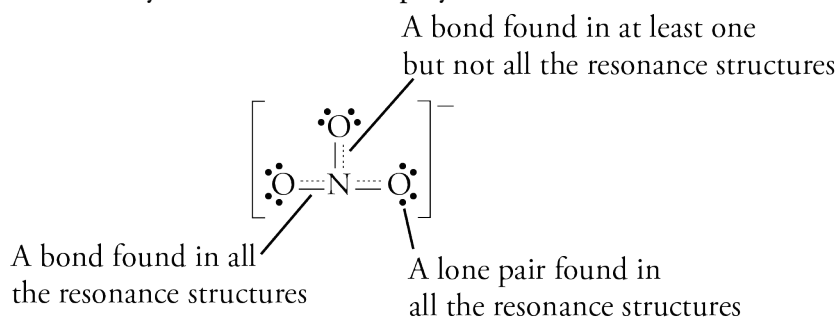
It is important to stress that the nitrate ion is not really changing from one resonance structure to another, but chemists find it useful, as an intermediate stage in the process of developing a better description of the nitrate ion, to think of it as if it were doing so. In actuality, the ion behaves as if it were a blend of the three resonance structures.

## OBJECTIVE 25

We can draw a Lewis-like structure that provides a better description of the actual character of the nitrate ion by blending the resonance structures into a single **resonance hybrid**:

- STEP 1** Draw the skeletal structure, using solid lines for the bonds that are found in all of the resonance structures.
- STEP 2** Where there is sometimes a bond and sometimes not, draw a dotted line.
- STEP 3** Draw only those lone pairs that are found on every one of the resonance structures. (Leave off the lone pairs that are on one or more resonance structure but not on all of them.)

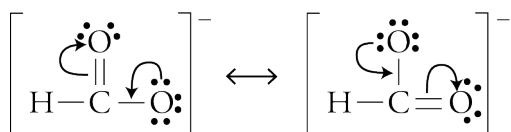
The resonance hybrid for the nitrate polyatomic ion is



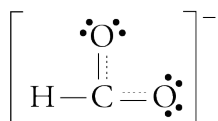
Resonance is possible whenever a Lewis structure has a multiple bond and an adjacent atom with at least one lone pair. The arrows in the following generalized structures show how you can think of the electrons shifting as one resonance structure changes to another.



For example, the two resonance structures for the formate ion,  $\text{HCO}_2^-$ , are

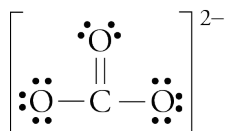


To generate the second resonance structure from the first, we imagine one lone pair dropping down to form another bond, and pushing an adjacent bond off to form a lone pair. The arrows show this *hypothetical* shift of electrons, which leads to the resonance hybrid below.



### EXERCISE 5.5 - Resonance

Draw all of the reasonable resonance structures and the resonance hybrid for the carbonate ion,  $\text{CO}_3^{2-}$ . A reasonable Lewis structure for the carbonate ion is



You can find a more comprehensive description of resonance at the textbook's Web site.

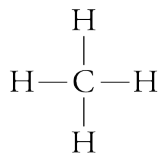
OBJECTIVE 25

OBJECTIVE 25

## 5.7 Molecular Geometry from Lewis Structures

The shapes of molecules play a major role in determining their function. For example, the shape of ethanol molecules allows them to attach to specific sites on nerve cell membranes and slow the transfer of information from one neuron to another. The shapes of the molecules in our food determine whether they taste sweet or bitter. The fat substitute Olestra is indigestible because it does not fit into the enzyme that digests natural fat. The purpose of this section is to show you how to use Lewis structures to predict three-dimensional shapes of simple molecules.

Lewis structures are useful for showing how the atoms in a molecule are connected by covalent bonds, but they do not always give a clear description of how the atoms are arranged in space. For example, the Lewis structure for methane,  $\text{CH}_4$ , shows the four covalent bonds connecting the central carbon atom to the hydrogen atoms:

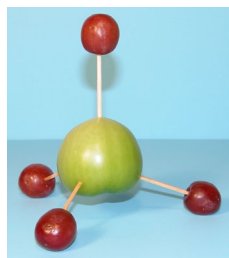
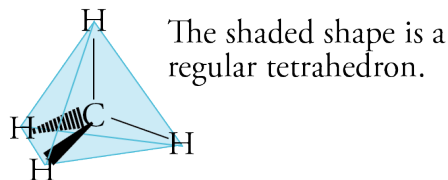


However, this Lewis structure seems to indicate that the five atoms are all located in the same plane and that the angles between the atoms are all  $90^\circ$  or  $180^\circ$ . This is not true. The actual shape of a molecule can be more accurately predicted by recognizing that the negatively charged electrons that form covalent bonds and lone pairs repel each other. Therefore, the most stable arrangement of the electron-groups is the molecular shape that keeps the groups as far away from each other as possible.

## OBJECTIVE 26

## OBJECTIVE 27

The best way to keep the negative charges for the four covalent bonds in a methane molecule as far apart as possible is to place them in a three-dimensional molecular shape called **tetrahedral**, with angles of  $109.5^\circ$  between the bonds. The angle formed by straight lines (representing bonds) connecting the nuclei of three adjacent atoms is called a **bond angle**.



**Figure 5.18**  
A Green Tomato-Grape  
Ball-and-Stick Model for  
the  $\text{CH}_4$  Molecule

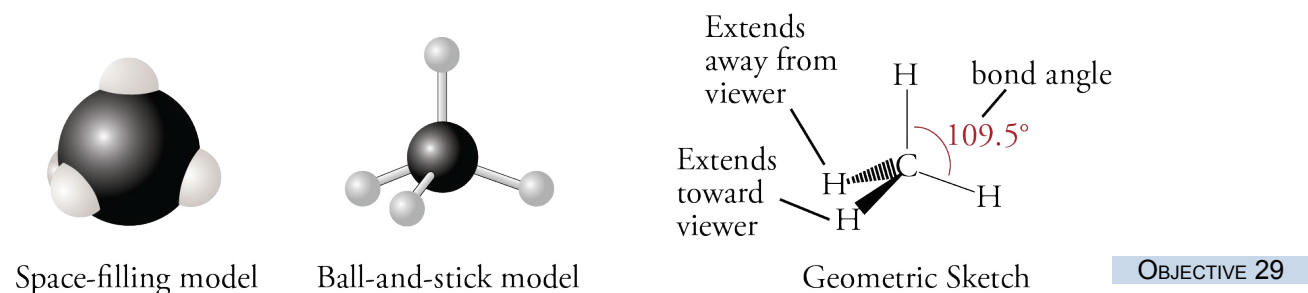
## OBJECTIVE 28

The tetrahedral arrangement is very common and important in chemistry, so be sure that you have a clear image in your mind of the three-dimensional shape of a molecule such as methane. You can make a ball-and-stick model of this molecule using a firm piece of fruit such as an apple or a green tomato for the carbon atom, four toothpicks to represent the axes through each of the four bonds, and four grapes to represent the hydrogen atoms. Try to arrange the toothpicks and grapes so that each grape is an equal distance from the other three. When you think you have successfully constructed a  $\text{CH}_4$  fruit molecule, turn your model over and over, so that a different grape is on top each time. If your model is really tetrahedral, it will always look the same, no matter which grape is on top (Figure 5.18).

Three ways to represent the methane molecule are shown in Figure 5.19. The first image, a *space-filling model*, provides the most accurate representation of the electron-charge clouds for the atoms in  $\text{CH}_4$ . A *ball-and-stick model*, the second image, emphasizes the molecule's correct molecular shape and shows the covalent bonds more clearly. The third image, a *geometric sketch*, shows a simple technique for describing three-dimensional tetrahedral structures with a two-dimensional drawing. Two hydrogen atoms are connected to the central carbon atom with solid lines. Picture these as being in the same plane as the carbon atom. A third hydrogen atom, connected to the central carbon with a solid wedge, comes out of the plane toward you. The

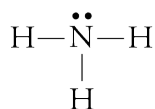


fourth hydrogen atom, connected to the carbon atom by a dashed wedge, is situated back behind the plane of the page.

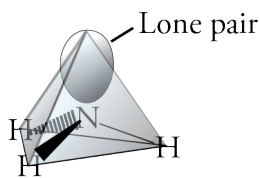


**Figure 5.19**  
Three Ways to Describe a Methane Molecule

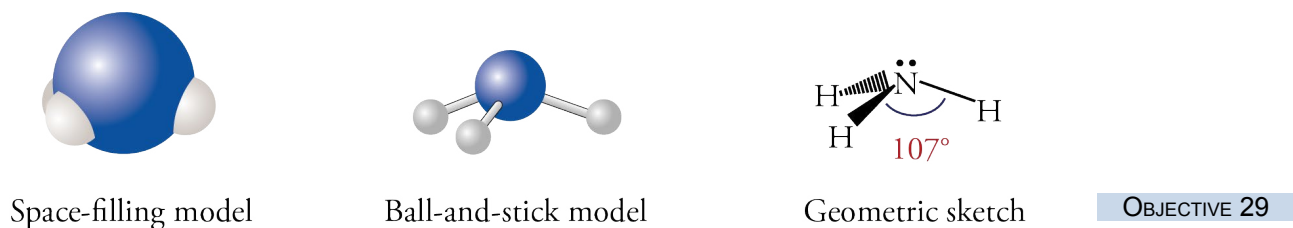
The nitrogen atom in an ammonia molecule,  $\text{NH}_3$ , forms three covalent bonds and in addition has a lone pair of electrons. A lone pair on a central atom must be considered in predicting a molecule's shape.



Like the carbon atom in a methane molecule, the nitrogen atom has four electron-groups around it, so the ammonia molecule has a shape that is very similar to the shape of a  $\text{CH}_4$  molecule.



The lone pair on the nitrogen atom repels neighboring electron-groups more strongly than the bond pairs do, so the lone pair in the ammonia molecule pushes the bond pairs closer together than the bond pairs for methane. The bond angle is about  $107^\circ$  instead of  $109.5^\circ$ . Figure 5.20 shows three ways to represent the ammonia molecule.

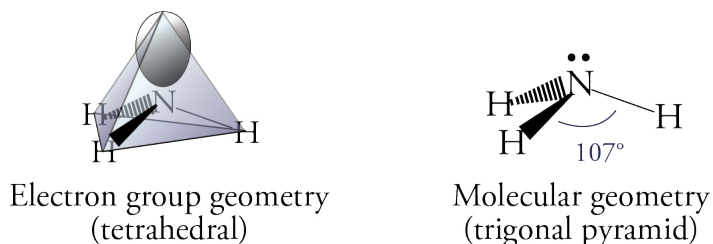


**Figure 5.20**  
Three Ways to Describe an Ammonia Molecule

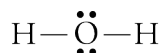
There are two ways to describe the geometry of the ammonia molecule. We can describe the arrangement of all the electron groups, including the lone pair, and call the shape tetrahedral, or we can describe the arrangement of the atoms only, without

considering the lone pair, and call ammonia's shape a **trigonal pyramid** (Figure 5.21). The three hydrogen atoms represent the corners of the pyramid's base, and the nitrogen atom forms the peak. In this book, we will call the geometry that describes all of the electron groups, including the lone pairs, the **electron group geometry**. The shape that describes the arrangement of the atoms only—treating the lone pairs as invisible—will be called the **molecular geometry**.

**Figure 5.21**  
Geometry of an Ammonia  
Molecule



The Lewis structure of water shows that the oxygen atom has four electron-groups around it: two covalent bonds and two lone pairs.



**OBJECTIVE 30**

We predict that the four groups would be distributed in a tetrahedral electron group arrangement to keep their negative charges as far apart as possible. Water's molecular geometry, which describes the arrangement of the atoms only, is called **bent**. The two lone pairs on the oxygen atom exert a stronger repulsion than the two bond groups and push the bond groups closer together than they would be if all of the electron groups around the oxygen atom were identical. The bond angle becomes  $105^\circ$  instead of the  $109.5^\circ$  predicted for four identical electron groups (Figure 5.22).

**Figure 5.22**  
Three ways to Describe  
a Water Molecule

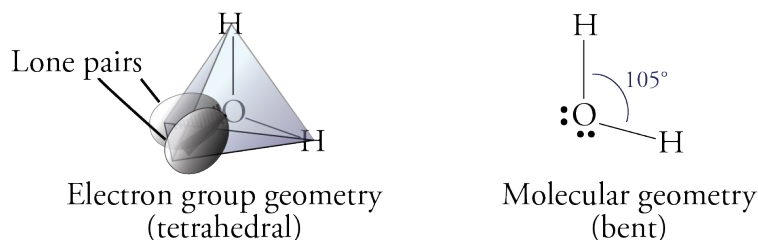
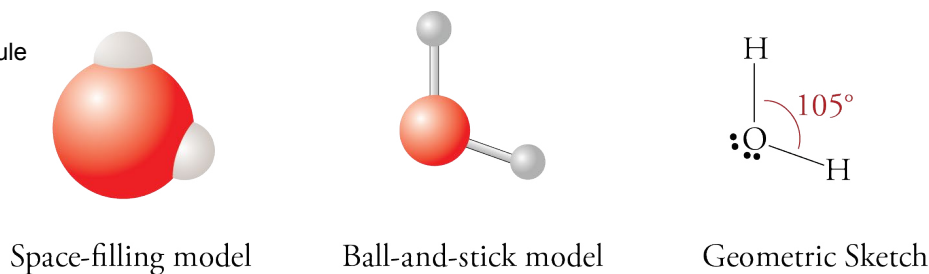


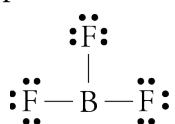
Figure 5.23 shows three ways to represent a water molecule.

**Figure 5.23**  
Geometry of a Water Molecule

**OBJECTIVE 29**

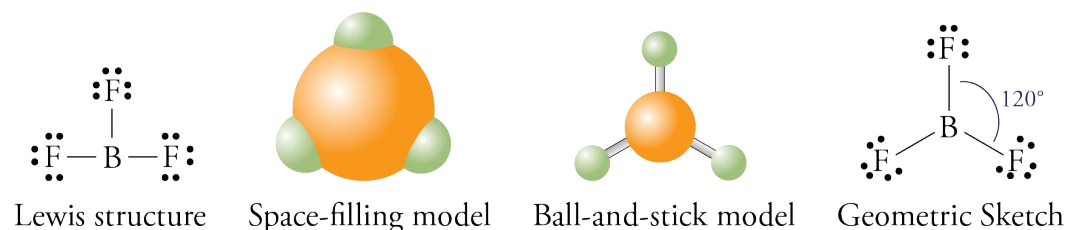


The boron atom in the center of a  $\text{BF}_3$  molecule is surrounded by three electron groups, which are each composed of a two-electron covalent bond to a fluorine atom.

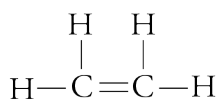


The geometric arrangement that keeps three electron groups as far apart as possible is called **trigonal planar** (often called **triangular planar**) and leads to angles of  $120^\circ$  between the groups. Therefore, the B–F bonds in  $\text{BF}_3$  molecules have a trigonal planar arrangement with  $120^\circ$  angles between any two fluorine atoms (Figure 5.24).

## OBJECTIVE 30

**Figure 5.24**Four Ways to Describe a Boron Trifluoride,  $\text{BF}_3$ , Molecule

Molecules quite often have more than one central atom (an atom with two or more atoms bonded to it). To describe the arrangement of atoms in such molecules, we need to consider the central atoms one at a time. Let's look at the ethene,  $\text{C}_2\text{H}_4$ , molecule:



There are three electron groups around each of ethene's carbon atoms (two single bonds to hydrogen atoms and the double bond to the other carbon). Because these groups repel each other, the most stable arrangement around each carbon atom—with the electron groups as far apart as possible—is a triangle with all three groups in the same plane and with angles of about  $120^\circ$  between the groups. That is, ethene has a trigonal planar arrangement of atoms around each carbon.

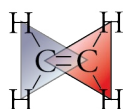
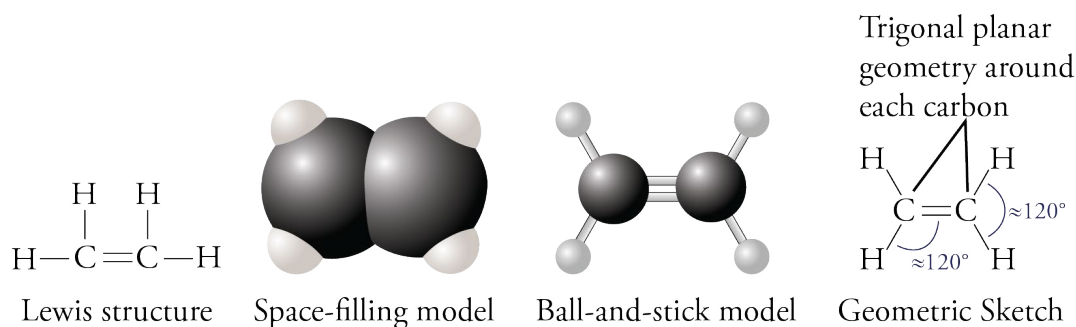


Figure 5.25 shows four ways to describe a  $\text{C}_2\text{H}_4$  molecule. Because the bond groups around each carbon are not identical (two are single bonds and one is a double bond), the angles are only approximately  $120^\circ$ .

**Figure 5.25**Four Ways to Describe an Ethene (Ethylene),  $\text{C}_2\text{H}_4$ , Molecule

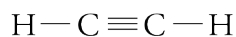
When atoms have only two electron groups around them, the groups are arranged in a line, with bond angles of  $180^\circ$ . This arrangement is called **linear**. For example, the carbon atom in  $\text{HCN}$  has two electron groups around it (one single bond and one

## OBJECTIVE 30

triple bond), and the best way to keep these groups apart is in a linear shape.



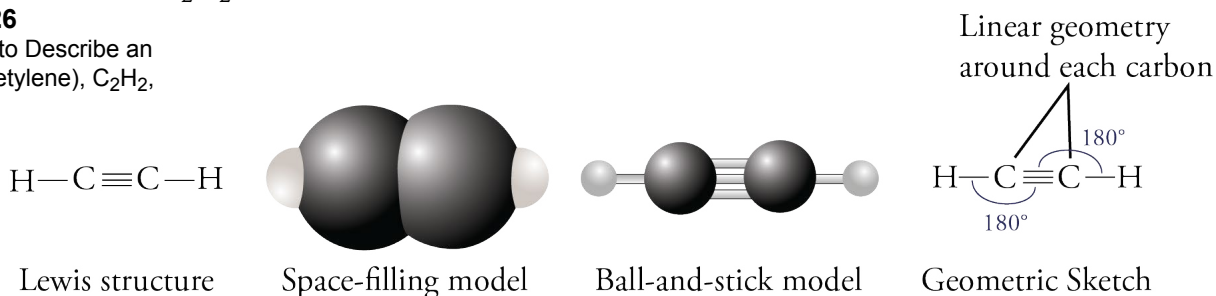
There are two central atoms in an acetylene,  $\text{C}_2\text{H}_2$ , molecule.



Because each of acetylene's central atoms has two electron groups around it, the arrangement of atoms around each is linear. Figure 5.26 shows four ways to describe a  $\text{C}_2\text{H}_2$  molecule.

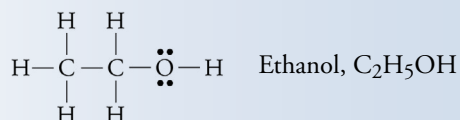
**Figure 5.26**

Four Ways to Describe an Ethyne (Acetylene),  $\text{C}_2\text{H}_2$ , Molecule

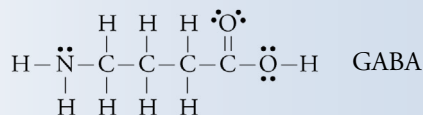


## SPECIAL TOPIC 5.1 Molecular Shapes, Intoxicating Liquids, and the Brain

Even if we do not all have firsthand experience with alcoholic beverages, we all know that their consumption slows brain activity. Ethanol, or ethyl alcohol, is the chemical in alcoholic beverages that causes this change.



Information is transferred through our nervous system when one nerve cell causes the next nerve cell to fire. This firing is regulated by the attachment of molecules with *specific shapes* to large protein molecules that form part of the nerve cell's membrane. When certain small molecules of the correct shape attach to the protein structures in the cell membranes, the cell is caused to fire. When certain other molecules attach, the firing of the cell is inhibited. For example, when a molecule called gamma-aminobutyric acid, or GABA, attaches to a protein molecule in certain nerve cells, it causes changes that inhibit the firing of the cell.



The GABA molecules are constantly attaching to the protein and then leaving again. When the GABA molecules are attached, the information transfer between nerve cells is inhibited. Anything that would make it easier for the GABA molecules to attach to the protein would lead to a slowing of the transfer of information between nerve cells.

Ethanol molecules can attach to the same protein as the GABA molecules, but they attach to a different site on the protein molecule. They change the shape of the protein in such a way that it becomes easier for the GABA molecules to find their position on the protein. Thus, when ethanol is present, the GABA molecules will become attached to the protein more often, inhibiting the firing of the cell. In this way, ethanol helps to slow the transfer of nerve information.

This slowing of the transfer of nerve information to the brain may not be a big problem to someone having a glass of wine with dinner at home, but when a person has a few drinks and drives a car, the consequences can be serious. If a deer runs in front of the car, we want the “put on the brake” signal sent from eyes to brain and then from brain to foot as quickly as possible.

The following sample study sheet summarizes a systematic procedure for predicting electron group geometries, drawing geometric sketches, and describing molecular geometries.

**TIP-OFF** In this chapter, you are given a Lewis structure for a molecule or polyatomic ion (or a formula from which you can draw a Lewis structure), and asked to (1) name the electron group geometry around one or more atoms in the structure, (2) draw a geometric sketch of the structure, and/or (3) name the molecular geometry around one or more of the atoms in the structure. You will find in later chapters that there are other tip-offs for these tasks.

**GENERAL STEPS** Follow these steps.

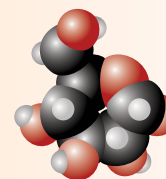
- STEP 1** To determine the name of the electron group geometry around each atom that is attached to two or more other atoms, count the number of electron groups around each “central” atom and apply the guidelines found on Table 5.2. An electron group can be a single bond, a lone pair, or a multiple bond. (Double and triple bonds count as one group.)
- STEP 2** Use one or more of the geometric sketches shown on Table 5.2 as models for the geometric sketch of your molecule. If the groups around the central atom are identical, the bond angle is exact. If the groups are not identical, the angles are approximate.
- STEP 3** To determine the name of the molecular geometry around each atom that has two or more atoms attached to it, count the number of bond groups and lone pairs, and then apply the guidelines found on Table 5.2. Single, double, and triple bonds all count as one bond group. Note that if all of the electron groups attached to the atom are bond groups (no lone pairs), the name of the molecular geometry is the same as the name of the electron group geometry.

**EXAMPLE** See Examples 5.9 and 5.10.

## Sample Study Sheet 5.3


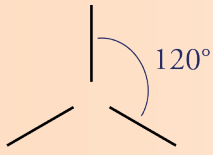
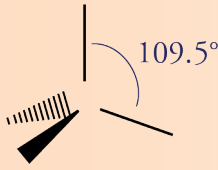
### Predicting Molecular Geometry

OBJECTIVE 30



Web  
Molecules

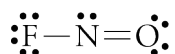
**Table 5.2** Electron Group and Molecular Geometry

| $e^-$ groups | $e^-$ group geometry | General geometric sketch  | Bond angles   | Bond groups | Lone pairs | Molecular geometry |
|--------------|----------------------|---|---------------|-------------|------------|--------------------|
| 2            | linear               |  | $180^\circ$   | 2           | 0          | linear             |
| 3            | trigonal planar      |  | $120^\circ$   | 3           | 0          | trigonal planar    |
|              |                      |   |               | 2           | 1          | bent               |
| 4            | tetrahedral          |  | $109.5^\circ$ | 4           | 0          | tetrahedral        |
|              |                      |   |               | 3           | 1          | trigonal pyramid   |
|              |                      |   |               | 2           | 2          | bent               |

## EXAMPLE 5.9 - Predicting Molecular Geometry

## OBJECTIVE 30

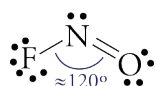
Nitrosyl fluoride, NOF, is used as an oxidizer in rocket fuels.



Using its Lewis structure, (a) write the name of the electron group geometry around the nitrogen atom, (b) draw a geometric sketch of the molecule, including bond angles, and (c) write the name of the molecular geometry around the nitrogen atom.

*Solution*

- Table 5.2 tells us that because there are three electron groups around the nitrogen in nitrosyl fluoride (two bond groups and one lone pair), the electron group geometry around the nitrogen atom is **trigonal planar**.
- The geometric sketch for NOF is

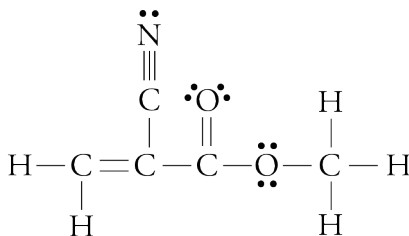


- The nitrogen atom in NOF has one single bond, one double bond, and one lone pair. According to Table 5.2, an atom with two bond groups and one lone pair has the molecular geometry called **bent**.

## EXAMPLE 5.10 - Predicting Molecular Geometry

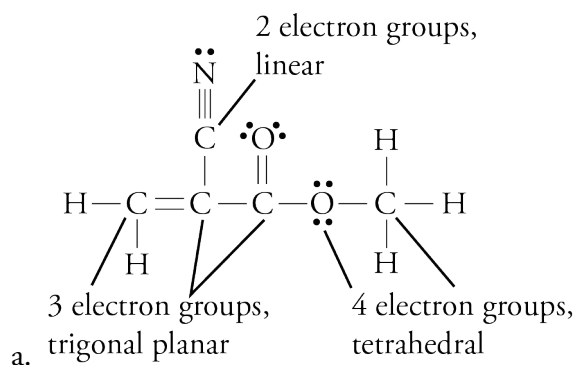
## OBJECTIVE 30

Methyl cyanoacrylate is the main ingredient in Super Glue.

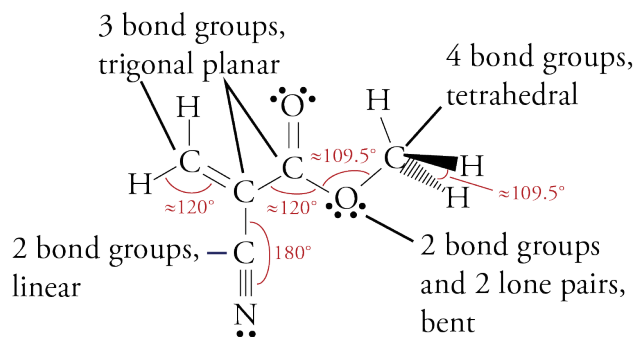


Using the Lewis structure above, (a) write the name of the electron group geometry around each atom that has two or more atoms attached to it, (b) draw a geometric sketch of the molecule, including bond angles, and (c) write the name of the molecular geometry around each atom that has two or more atoms attached to it. (This example was chosen because it contains several different types of electron groups and molecular shapes. The structures in the exercises and problems that accompany this chapter are much simpler.)

Solution



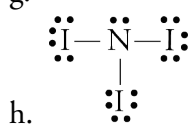
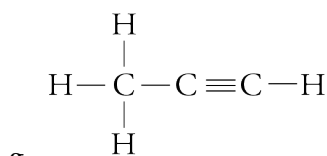
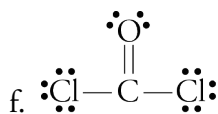
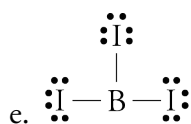
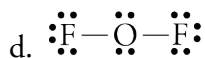
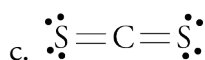
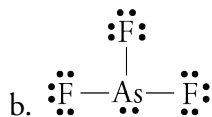
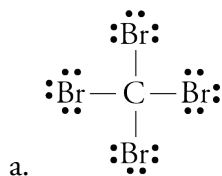
b and c. There are many ways to sketch the molecule and still show the correct geometry. The following is just one example. (Remember that your assigned exercises will be simpler than this one.)



### EXERCISE 5.6 - Molecular Geometry

For each of the Lewis structures given below, (a) write the name of the electron group geometry around each atom that has two or more atoms attached to it, (b) draw a geometric sketch of the molecule, including bond angles, and (c) write the name of the molecular geometry around each atom that has two or more atoms attached to it.

OBJECTIVE 30



Chapter  
Glossary

**Compound** A substance that contains two or more elements, the atoms of these elements always combining in the same whole-number ratio.

**Chemical formula** A concise written description of the components of a chemical compound. It identifies the elements in the compound by their symbols and indicates the relative number of atoms of each element with subscripts.

**Pure substance** A sample of matter that has constant composition. There are two types of pure substances, elements and compounds.

**Mixture** A sample of matter that contains two or more pure substances and has variable composition.

**Chemical bond** An attraction between atoms or ions in chemical compounds. Covalent bonds and ionic bonds are examples.

**Polar covalent bond** A covalent bond in which electrons are shared unequally, leading to a partial negative charge on the atom that attracts the electrons more and to a partial positive charge on the other atom.

**Nonpolar covalent bond** A covalent bond in which the difference in electron-attracting ability of two atoms in a bond is negligible (or zero), so the atoms in the bond have no significant charges.

**Ionic bond** The attraction between a cation and an anion.

**Molecular compound** A compound composed of molecules. In such compounds, all of the bonds between atoms are covalent bonds.

**Ionic compound** A compound that consists of ions held together by ionic bonds.

**Monatomic anions** Negatively charged particles, such as  $\text{Cl}^-$ ,  $\text{O}^{2-}$ , and  $\text{N}^{3-}$ , that contain single atoms with a negative charge.

**Monatomic cations** Positively charged particles, such as  $\text{Na}^+$ ,  $\text{Ca}^{2+}$ , and  $\text{Al}^{3+}$ , that contain single atoms with a positive charge.

**Polyatomic ion** A charged collection of atoms held together by covalent bonds.

**Valence electrons** The electrons that are most important in the formation of chemical bonds. The number of valence electrons for the atoms of an element is equal to the element's A-group number on the periodic table. The highest energy *s* and *p* electrons for an atom.

**Electron-dot symbol** A representation of an atom that consists of its elemental symbol surrounded by dots representing its valence electrons.

**Lone pair** Two electrons that are not involved in the covalent bonds between atoms but are important for explaining the arrangement of atoms in molecules. They are represented by pairs of dots in Lewis structures.

**Lewis structure** A representation of a molecule that consists of the elemental symbol for each atom in the molecule, lines to show covalent bonds, and pairs of dots to indicate lone pairs.

**Hydrocarbons** Compounds that contain only carbon and hydrogen.



**Organic chemistry** The branch of chemistry that involves the study of carbon-based compounds.

**Double bond** A link between atoms that results from the sharing of four electrons. It can be viewed as two 2-electron covalent bonds.

**Triple bond** A link between atoms that results from the sharing of six electrons. It can be viewed as three 2-electron covalent bonds.

**Alcohols** Compounds that contain a hydrocarbon group with one or more -OH groups attached.

**Isomers** Compounds that have the same molecular formula but different molecular structures.

**Resonance structures** Two or more Lewis structures for a single molecule or polyatomic ion that differ in the positions of lone pairs and multiple bonds but not in the positions of the atoms in the structure. It is as if the molecule or ion were able to shift from one of these structures to another by shifting pairs of electrons from one position to another.

**Resonance** The hypothetical switching from one resonance structure to another.

**Resonance hybrid** A structure that represents the average of the resonance structures for a molecule or polyatomic ion.

**Tetrahedral** The molecular shape that keeps the negative charge of four electron groups as far apart as possible. This shape has angles of  $109.5^\circ$  between the atoms.

**Bond angle** The angle formed by straight lines (representing bonds) connecting the nuclei of three adjacent atoms.

**Trigonal pyramid** The molecular geometry formed around an atom with three bonds and one lone pair.

**Electron group geometry** A description of the arrangement of all the electron groups around a central atom in a molecule or polyatomic ion, including the lone pairs.

**Molecular geometry** The description of the arrangement of all the atoms around a central atom in a molecule or polyatomic ion. This description does not consider lone pairs.

**Bent** The molecular geometry formed around an atom with two bond groups and two lone pairs or two bond groups and one lone pair.

**Trigonal planar (often called triangular planar)** The geometric arrangement that keeps three electron groups as far apart as possible. It leads to angles of  $120^\circ$  between the groups.

**Linear geometry** The geometric arrangement that keeps two electron groups as far apart as possible. It leads to angles of  $180^\circ$  between the groups.

You can test yourself on the glossary terms at the textbook's Web site.

## Chapter Objectives

**The goal of this chapter is to teach you to do the following.**

1. Define all of the terms in the Chapter Glossary.

### Section 5.1 Classification of Matter

2. Convert between a description of the number of atoms of each element found in a compound and its chemical formula.
3. Given a description of a form of matter, classify it as a pure substance or mixture.
4. Given a description of a pure substance, classify it as an element or compound.

### Section 5.2 Chemical Compounds and Chemical Bonds

5. Describe the polar covalent bond between two nonmetallic atoms, one of which attracts electrons more than the other one does. Your description should include a rough sketch of the electron-cloud that represents the electrons involved in the bond.
6. Describe the process that leads to the formation of ionic bonds between metallic and nonmetallic atoms.
7. Describe the difference between a nonpolar covalent bond, a polar covalent bond, and an ionic bond. Your description should include rough sketches of the electron-clouds that represent the electrons involved in the formation of each bond.
8. Given the names or formulas for two elements, identify the bond that would form between them as covalent or ionic.
9. Given a formula for a compound, classify it as either a molecular compound or an ionic compound.

### Section 5.3 Ionic Compounds

10. Explain why metals usually combine with nonmetals to form ionic bonds.
11. Write the ionic charges acquired by the following elements:
  - a. group 17 – halogens
  - b. oxygen, sulfur, and selenium
  - c. nitrogen and phosphorus
  - d. hydrogen
  - e. group 1 - alkali metals
  - f. group 2 - alkaline earth metals
  - g. group 3 elements
  - h. aluminum
  - i. iron, silver, copper, and zinc
12. Describe the crystal structure of sodium chloride, NaCl.
13. Describe the similarities and differences between the ionic structure of cesium chloride and ammonium chloride.

### Section 5.4 Molecular Compounds

14. Describe the advantages and disadvantages of using models to describe physical reality.
15. Describe the assumptions that lie at the core of the valence-bond model.
16. Write or identify the number of valence electrons in an atom of any representative element.

17. Use the valence-bond model to explain how (a) fluorine atoms form one covalent bond and have three lone pairs in molecules such as  $F_2$ , (b) hydrogen atoms can form one covalent bond and have no lone pairs in molecules such as HF, (c) carbon atoms can form four covalent bonds and have no lone pairs in molecules such as  $CH_4$ , (d) nitrogen, phosphorus, and arsenic atoms can form three covalent bonds and have one lone pair in molecules such as  $NH_3$ ,  $PH_3$ , and  $AsH_3$ , (e) nitrogen atoms can form four covalent bonds and have no lone pairs in structures such as  $NH_4^+$ , (f) oxygen, sulfur, and selenium atoms can form two covalent bonds and have two lone pairs in molecules such as  $H_2O$ ,  $H_2S$ , and  $H_2Se$  (g) oxygen atoms can form one covalent bond and have three lone pairs in structures such as  $OH^-$ , (h) carbon and oxygen atoms can form three covalent bonds and have one lone pair in molecules such as CO, (i) boron atoms can form three covalent bonds and have no lone pairs in molecules such as  $BF_3$ , (j) halogen atoms can form one covalent bond and have three lone pairs in molecules such as HF, HCl, HBr, and HI.
18. Give a general description of the information provided in a Lewis structure.
19. Write electron-dot symbols for the representative elements.
20. Identify the most common number of covalent bonds and lone pairs for the atoms of each of the following elements: hydrogen, the halogens (group 17), oxygen, sulfur, selenium, nitrogen, phosphorus, and carbon.

### Section 5.5 Drawing Lewis Structures

21. Given a chemical formula, draw a Lewis structure for it that has the most common number of covalent bonds and lone pairs for each atom.
22. Given a formula for a molecule or polyatomic ion, draw a reasonable Lewis structure for it.
23. Convert between the following systematic names and their chemical formulas: methanol, ethanol, and 2-propanol.
24. Given one of the following names for alcohols, write its chemical formula: methyl alcohol, ethyl alcohol, and isopropyl alcohol.

### Section 5.6 Resonance

25. Given a Lewis structure or enough information to draw one, predict whether it is best described in terms of resonance, and if it is, draw all of the reasonable resonance structures and the resonance hybrid for the structure.

### Section 5.7 Molecular Geometry from Lewis Structures

26. Describe the tetrahedral molecular shape.
27. Explain why the atoms in the  $CH_4$  molecule have a tetrahedral molecular shape.
28. Describe the information given by a space-filling model, a ball-and-stick model, and a geometric sketch.
29. Draw geometric sketches, including bond angles, for  $CH_4$ ,  $NH_3$ , and  $H_2O$ .
30. Given a Lewis structure or enough information to write one, do the following.
  - a. Write the name of the electron group geometry around each atom that has two or more atoms attached to it.
  - b. Draw a geometric sketch of the molecule, including bond angles (or approximate bond angles).
  - c. Write the name of the molecular geometry around each atom that has two or more atoms attached to it.

## Review Questions

- Using the A-group convention, what is the group number of the column on the periodic table that includes the element chlorine, Cl?

For problems 2-7, write in each blank the word or words that best complete each sentence.

- An atom or group of atoms that has lost or gained one or more electrons to create a charged particle is called a(n) \_\_\_\_\_.
- An atom or collection of atoms with an overall positive charge is a(n) \_\_\_\_\_.
- An atom or collection of atoms with an overall negative charge is a(n) \_\_\_\_\_.
- A(n) \_\_\_\_\_ bond is a link between atoms that results from the sharing of two electrons.
- A(n) \_\_\_\_\_ is an uncharged collection of atoms held together with covalent bonds.
- A molecule like  $H_2$ , which is composed of two atoms, is called \_\_\_\_\_.
- Describe the particle nature of solids, liquids, and gases. Your description should include the motion of the particles and the attractions between the particles.
- Describe the nuclear model of the atom.
- Describe the hydrogen molecule,  $H_2$ . Your description should include the nature of the link between the hydrogen atoms and a sketch that shows the two electrons in the molecule.
- Complete the following table.

| Element Name | Element symbol | Group number of periodic table | Metal, nonmetal, or metalloid? |
|--------------|----------------|--------------------------------|--------------------------------|
|              | Li             |                                |                                |
| carbon       |                |                                |                                |
|              | Cl             |                                |                                |
| oxygen       |                |                                |                                |
|              | Cu             |                                |                                |
| calcium      |                |                                |                                |
|              | Sc             |                                |                                |

- Write the name of the group on the periodic table to which each of the following belongs.
  - chlorine
  - xenon
  - sodium
  - magnesium
- Write the definition of orbital.
- Write a complete electron configuration and an orbital diagram for each of the following.
  - oxygen, O
  - phosphorus, P

Complete the following statements by writing one of these words or phrases in each blank.

|  |                       |
|--|-----------------------|
| 109.5°                                   | most                  |
| A-group                                  | most common           |
| anion-anion and cation-cation repulsions | negative              |
| anions                                   | negligible (or zero)  |
| arranged in space                        | never                 |
| as far apart                             | noble gas             |
| as if                                    | octet                 |
| cation-anion attraction                  | oxygen                |
| cations                                  | pair unpaired         |
| change                                   | partial negative      |
| compounds                                | partial positive      |
| constant composition                     | plus                  |
| covalent                                 | polyatomic ions       |
| double                                   | positive              |
| eight                                    | predict               |
| eight valence                            | rarely                |
| electron-charge clouds                   | repel                 |
| electron group                           | <i>s</i> and <i>p</i> |
| explain                                  | smaller               |
| fewest                                   | solids                |
| four electrons                           | structural            |
| function                                 | subscripts            |
| gain                                     | symbols               |
| gases                                    | triple                |
| highest-energy                           | true                  |
| larger                                   | two                   |
| liquids                                  | uncharged             |
| lone pairs                               | useful                |
| lose                                     | usually ionic         |
| minus                                    | variable              |
| molecular                                | visualize             |
| molecules                                | whole-number ratio    |
| more                                     |                       |

15. A compound is a substance that contains two or more elements, the atoms of those elements always combining in the same \_\_\_\_\_.
16. There are relatively few chemical elements, but there are millions of chemical \_\_\_\_\_.
17. A chemical formula is a concise written description of the components of a chemical compound. It identifies the elements in the compound by their \_\_\_\_\_ and indicates the relative number of atoms of each element with \_\_\_\_\_.
18. When a substance has a(n) \_\_\_\_\_, it must by definition be either an element or a compound, and it is considered a pure substance.
19. Mixtures are samples of matter that contain two or more pure substances and have \_\_\_\_\_ composition.

20. When the difference in electron-attracting ability between atoms in a chemical bond is \_\_\_\_\_, the atoms in the bond will have no significant partial charges. We call this type of bond a nonpolar covalent bond.
21. Because particles with opposite charges attract each other, there is an attraction between \_\_\_\_\_ and \_\_\_\_\_. This attraction is called an ionic bond.
22. Compounds that have ionic bonds, such as the sodium chloride in table salt, are \_\_\_\_\_ at room temperature and pressure, but compounds with all covalent bonds, such as hydrogen chloride and water, can be \_\_\_\_\_ and \_\_\_\_\_ as well as solids.
23. The atom in a chemical bond that attracts electrons more strongly acquires a(n) \_\_\_\_\_ charge, and the other atom acquires a(n) \_\_\_\_\_ charge. If the electron transfer is significant but not enough to form ions, the atoms acquire \_\_\_\_\_ and \_\_\_\_\_ charges. The bond in this situation is called a polar covalent bond.
24. When a nonmetallic atom bonds to another nonmetallic atom, the bond is \_\_\_\_\_.
25. When a metallic atom bonds to a nonmetallic atom, the bond is \_\_\_\_\_.
26. Molecular compounds are composed of \_\_\_\_\_, which are collections of atoms held together by all covalent bonds.
27. Metallic atoms hold some of their electrons relatively loosely, and as a result, they tend to \_\_\_\_\_ electrons and form cations. In contrast, nonmetallic atoms attract electrons more strongly than metallic atoms, so nonmetals tend to \_\_\_\_\_ electrons and form anions.
28. Nonmetallic atoms form anions to get the same number of electrons as the nearest \_\_\_\_\_.
29. When atoms gain electrons and form anions, they get \_\_\_\_\_. When atoms lose electrons and form cations, they get significantly \_\_\_\_\_.
30. The ions in ionic solids take the arrangement that provides the greatest \_\_\_\_\_ while minimizing the \_\_\_\_\_.
31. When developing a model of physical reality, scientists take what they think is \_\_\_\_\_ and simplify it enough to make it \_\_\_\_\_.
32. Models have advantages and disadvantages. They help us to \_\_\_\_\_, \_\_\_\_\_, and \_\_\_\_\_ chemical changes, but we need to remind ourselves now and then that they are only models and that as models, they have their limitations.
33. One characteristic of models is that they \_\_\_\_\_ with time.
34. The valence-bond model, which is commonly used to describe the formation of covalent bonds, is based on the assumptions that only the \_\_\_\_\_ electrons participate in bonding, and covalent bonds usually form to \_\_\_\_\_ electrons.
35. Valence electrons are the highest-energy \_\_\_\_\_ electrons in an atom.
36. When the columns in the periodic table are numbered by the \_\_\_\_\_ convention, the number of valence electrons in each atom of a representative element is equal to the element's group number in the periodic table.
37. Paired valence electrons are called \_\_\_\_\_.

38. The noble gases (group 8A) have a(n) \_\_\_\_\_ of electrons (except for helium, which has only two electrons total), and they are so stable that they rarely form chemical bonds with other atoms.
39. When atoms other than the noble-gas atoms form bonds, they often have \_\_\_\_\_ electrons around them in total.
40. Atoms can form double bonds, in which \_\_\_\_\_ are shared between atoms. Double bonds are represented by \_\_\_\_\_ lines in a Lewis structure.
41. When forming four bonds to hydrogen atoms in a methane molecule, each carbon atom behaves \_\_\_\_\_ one electron is promoted from the  $2s$  orbital to the  $2p$  orbital.
42. A(n) \_\_\_\_\_ bond can be viewed as the sharing of six electrons between two atoms.
43. The shortcut for drawing Lewis structures for which we try to give each atom its \_\_\_\_\_ bonding pattern works well for many simple uncharged molecules, but it does not work reliably for molecules that are more complex or for \_\_\_\_\_.
44. For \_\_\_\_\_ molecules, the total number of valence electrons is the sum of the valence electrons of each atom.
45. For polyatomic cations, the total number of valence electrons is the sum of the valence electrons for each atom \_\_\_\_\_ the charge.
46. For polyatomic anions, the total number of valence electrons is the sum of the valence electrons for each atom \_\_\_\_\_ the charge.
47. Hydrogen and fluorine atoms are \_\_\_\_\_ in the center of a Lewis structure.
48. Oxygen atoms are \_\_\_\_\_ in the center of a Lewis structure.
49. The element with the \_\_\_\_\_ atoms in the formula is often in the center of a Lewis structure.
50. The atom that is capable of making the \_\_\_\_\_ bonds is often in the center of a Lewis structure.
51. Oxygen atoms rarely bond to other \_\_\_\_\_ atoms.
52. In a reasonable Lewis structure, carbon, nitrogen, oxygen, and fluorine always have \_\_\_\_\_ electrons around them.
53. In a reasonable Lewis structure, hydrogen will always have a total of \_\_\_\_\_ electrons from its one bond.
54. Boron can have fewer than eight electrons around it in a reasonable Lewis structure but never \_\_\_\_\_ than eight.
55. Substances that have the same molecular formula but different \_\_\_\_\_ formulas are called isomers.
56. The shapes of molecules play a major role in determining their \_\_\_\_\_.
57. Lewis structures are useful for showing how the atoms in a molecule are connected by covalent bonds, but they do not always give a clear description of how the atoms are \_\_\_\_\_.

58. The actual shape of a molecule can be predicted by recognizing that the negatively charged electrons that form covalent bonds and lone pairs \_\_\_\_\_ each other. Therefore, the most stable arrangement of the electron groups is the molecular shape that keeps the groups as far away from each other as possible.
59. The geometric arrangement in which four electron groups are as far apart as possible is a tetrahedral arrangement, with bond angles of \_\_\_\_\_.
60. A space-filling model provides the most accurate representation of the \_\_\_\_\_ for the atoms in  $\text{CH}_4$ .
61. The most stable arrangement for electron groups is one in which they are \_\_\_\_\_ as possible.
62. In this book, we call the geometry that describes all of the electron groups, including the lone pairs, the \_\_\_\_\_ geometry. The shape that describes the arrangement of the atoms only—treating the lone pairs as invisible—is called the \_\_\_\_\_ geometry.

## Chapter Problems

OBJECTIVE 3

OBJECTIVE 4

OBJECTIVE 3

OBJECTIVE 4



OBJECTIVE 2

### Section 5.1 Classification of Matter

63. Classify each of the following as a pure substance or a mixture. If it is a pure substance, is it an element or a compound? Explain your answer.
- apple juice
  - potassium (A serving of one brand of apple juice provides 6% of the recommended daily allowance of potassium.)
  - ascorbic acid (vitamin C),  $\text{C}_6\text{H}_8\text{O}_6$ , in apple juice
64. Classify each of the following as a pure substance or a mixture. If it is a pure substance, is it an element or a compound? Explain your answer.
- fluorine (used to make fluorides, such as those used in toothpaste)
  - toothpaste
  - calcium fluoride,  $\text{CaF}_2$ , from the naturally occurring ore fluorite (It is used to make sodium monofluorophosphate, which is added to some toothpastes.)
65. Write the chemical formula for each of the following compounds. List the symbols for the elements in the order that the elements are mentioned in the description.
- a compound with molecules that consist of two nitrogen atoms and three oxygen atoms.
  - a compound with molecules that consist of one sulfur atom and four fluorine atoms.
  - a compound that contains one aluminum atom for every three chlorine atoms.
  - a compound that contains two lithium atoms and one carbon atom for every three oxygen atoms.



66. Write the chemical formula for each of the following compounds. List the symbols for the elements in the order that the elements are mentioned in the description.
- a compound with molecules that consist of two phosphorus atoms and five oxygen atoms.
  - a compound with molecules that consist of two hydrogen atoms and one sulfur atom.
  - a compound that contains three calcium atoms for every two nitrogen atoms.
  - a compound with molecules that consist of 12 carbon atoms, 22 hydrogen atoms, and 11 oxygen atoms.

OBJECTIVE 2

### Section 5.2 Chemical Compounds and Chemical Bonds

67. Hydrogen bromide, HBr, is used to make pharmaceuticals that require bromine in their structure. Each hydrogen bromide molecule has one hydrogen atom bonded to one bromine atom by a polar covalent bond. The bromine atom attracts electrons more than does the hydrogen atom. Draw a rough sketch of the electron-cloud that represents the electrons involved in the bond.
68. Iodine monochloride, ICl, is a compound used to make carbon-based (organic) compounds that contain iodine and chlorine. It consists of diatomic molecules with one iodine atom bonded to one chlorine atom by a polar covalent bond. The chlorine atom attracts electrons more than does the iodine atom. Draw a rough sketch of the electron cloud that represents the electrons involved in the bond.
69. Atoms of potassium and fluorine form ions and ionic bonds in a very similar way to atoms of sodium and chlorine. Each atom of one of these elements loses one electron, and each atom of the other element gains one electron. Describe the process that leads to the formation of the ionic bond between potassium and fluorine atoms in potassium fluoride. Your answer should include mention of the charges that form on the atoms.
70. Atoms of magnesium and oxygen form ions and ionic bonds in a similar way to atoms of sodium and chlorine. The difference is that instead of having each atom gain or lose one electron, each atom of one of these elements loses two electrons, and each atom of the other element gains two electrons. Describe the process that leads to the formation of the ionic bond between magnesium and oxygen atoms in magnesium oxide. Your answer should include mention of the charges that form on the atoms.
71. Explain how a nonpolar covalent bond, a polar covalent bond, and an ionic bond differ. Your description should include rough sketches of the electron-clouds that represent the electrons involved in the formation of each bond.
72. Write a chemical formula that represents both a molecule and a compound. Write a formula that represents a compound but not a molecule.
73. Would you expect the bonds between the following atoms to be ionic or covalent bonds?
- N-O
  - Al-Cl

OBJECTIVE 5

OBJECTIVE 5

OBJECTIVE 6

OBJECTIVE 6

OBJECTIVE 7

OBJECTIVE 8

## OBJECTIVE 8

74. Would you expect the bonds between the following atoms to be ionic or covalent bonds?

- a. Li-F      b. C-N

## OBJECTIVE 9



## OBJECTIVE 9

75. Classify each of the following as either a molecular compound or an ionic compound.

- a. acetone,  $\text{CH}_3\text{COCH}_3$  (a common paint solvent)  
b. sodium sulfide,  $\text{Na}_2\text{S}$  (used in sheep dips)

76. Classify each of the following as either a molecular compound or an ionic compound.

- a. cadmium fluoride,  $\text{CdF}_2$  (a starting material for lasers)  
b. sulfur dioxide,  $\text{SO}_2$  (a food additive that inhibits browning and bacterial growth)

### Section 5.3 Ionic Compounds

## OBJECTIVE 10

77. Explain why metals usually combine with nonmetals to form ionic bonds.

78. How many protons and electrons do each of the following ions have?

- a.  $\text{Be}^{2+}$   
b.  $\text{S}^{2-}$

79. How many protons and electrons do each of the following ions have?

- a.  $\text{N}^{3-}$   
b.  $\text{Ba}^{2+}$

## OBJECTIVE 11

80. When atoms of the following elements form ions, what charge or charges would the ions have?

- |             |             |
|-------------|-------------|
| a. hydrogen | e. iron     |
| b. Ca       | f. Zn       |
| c. fluorine | g. nitrogen |
| d. K        | h. Sc       |

## OBJECTIVE 11

81. When atoms of the following elements form ions, what charge or charges would the ions have?

- |           |              |
|-----------|--------------|
| a. oxygen | e. magnesium |
| b. Br     | f. Al        |
| c. silver | g. sodium    |
| d. Cu     | h. Li        |

## OBJECTIVE 12

82. Silver bromide,  $\text{AgBr}$ , is the compound on black and white film that causes the color change when the film is exposed to light. It has a structure similar to that of sodium chloride. What are the particles that form the basic structure of silver bromide? What type of attraction holds these particles together? Draw a rough sketch of the structure of solid silver bromide.

## OBJECTIVE 13

83. Describe the crystal structures of cesium chloride and ammonium chloride. How are they similar, and how are they different?

### Section 5.4 Molecular Compounds

## OBJECTIVE 14

84. Describe the advantages and disadvantages of using models to describe physical reality.

## OBJECTIVE 15

85. Describe the assumptions that lie at the core of the valence-bond model.

**86.** How many valence electrons do the atoms of each of the following elements have? Write the electron configuration for these electrons. (For example, fluorine has seven valence electrons, which can be described as  $2s^2 2p^5$ .)

- a. nitrogen, N      b. sulfur, S      c. iodine, I      d. argon, Ar

OBJECTIVE 16

**87.** How many valence electrons do the atoms of each of the following elements have? Write the electron configuration for these electrons. (For example, fluorine has seven valence electrons, which can be described as  $2s^2 2p^5$ .)

- a. oxygen, O      b. boron, B      c. neon, Ne      d. phosphorus, P      e. carbon, C

OBJECTIVE 16

**88.** Draw electron-dot symbols for each of the following elements.

- a. nitrogen, N      b. sulfur, S      c. iodine, I      d. argon, Ar

OBJECTIVE 19

**89.** Draw electron-dot symbols for each of the following elements.

- a. oxygen, O      b. boron, B      c. neon, Ne      d. phosphorus, P      e. carbon, C

OBJECTIVE 19

**90.** To which group on the periodic table would atoms with the following electron-dot symbols belong? List the group numbers using the 1-18 convention and using the A-group convention.

- a.  $\cdot\ddot{X}\cdot$       b.  $\cdot\ddot{X}\cdot$       c.  $\cdot\ddot{X}\cdot$

**91.** To which group on the periodic table would atoms with the following electron-dot symbols belong? List the group numbers using the 1-18 convention and using the A-group convention.

- a.  $\cdot\ddot{X}\cdot$       b.  $\cdot\ddot{X}\cdot$       c.  $\cdot\ddot{X}\cdot$

**92.** For each of the following elements, sketch all of the ways mentioned in Section 5.4 that their atoms could look in a Lewis structure. For example, fluorine has only one bonding pattern, and it looks like  $\text{—}\ddot{\text{F}}\cdot$ .

- a. nitrogen, N      b. boron, B      c. carbon, C

**93.** For each of the following elements, sketch all of the ways mentioned in Section 5.4 that their atoms could look in a Lewis structure. For example, fluorine has only one bonding pattern, and it looks like  $\text{—}\ddot{\text{F}}\cdot$ .

- a. hydrogen, H      b. oxygen, O      c. chlorine, Cl

**94.** Use the valence-bond model to explain the following.

OBJECTIVE 17

- Fluorine atoms have one bond and three lone pairs in  $\text{F}_2$ .
- Carbon atoms have four bonds and no lone pairs in  $\text{CH}_4$ .
- Nitrogen atoms have three bonds and one lone pair in  $\text{NH}_3$ .
- Sulfur atoms have two bonds and two lone pairs in  $\text{H}_2\text{S}$ .
- Oxygen atoms have one bond and three lone pairs in  $\text{OH}^-$ .

**95.** Use the valence-bond model to explain the following.

OBJECTIVE 17

- Phosphorus atoms have three bonds and one lone pair in  $\text{PH}_3$ .
- Nitrogen atoms have four bonds and no lone pairs in  $\text{NH}_4^+$ .
- Oxygen atoms have two bonds and two lone pairs in  $\text{H}_2\text{O}$ .
- Boron atoms have three bonds and no lone pair in  $\text{BF}_3$ .
- Chlorine atoms have one bond and three lone pairs in  $\text{HCl}$ .

96. Based on your knowledge of the most common bonding patterns for the nonmetallic elements, predict the formulas with the lowest subscripts for the compounds that would form from the following pairs of elements. (For example, hydrogen and oxygen can combine to form  $\text{H}_2\text{O}$  and  $\text{H}_2\text{O}_2$ , but  $\text{H}_2\text{O}$  has lower subscripts.)

- a. C and H                      b. S and H                      c. B and F

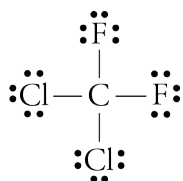
97. Based on your knowledge of the most common bonding patterns for the nonmetallic elements, predict the formulas with the lowest subscripts for the compounds that would form from the following pairs of elements. (For example, hydrogen and oxygen can combine to form  $\text{H}_2\text{O}$  and  $\text{H}_2\text{O}_2$ , but  $\text{H}_2\text{O}$  has lower subscripts.)

- a. P and I                      b. O and Br                      c. N and Cl

OBJECTIVE 18

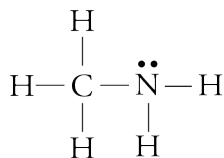


98. The following Lewis structure is for CFC-12, which is one of the ozone-depleting chemicals that has been used as an aerosol can propellant and as a refrigerant. Describe the information given in this Lewis structure.



OBJECTIVE 18

99. Describe the information given in the following Lewis structure for methylamine, a compound used to make insecticides and rocket propellants.



OBJECTIVE 20

100. Write the most common number of covalent bonds and lone pairs for atoms of each of the following nonmetallic elements.

- a. H                                      c. sulfur  
b. iodine                                d. N

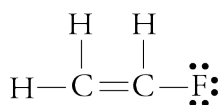
OBJECTIVE 20

101. Write the most common number of covalent bonds and lone pairs for atoms of each of the following nonmetallic elements.

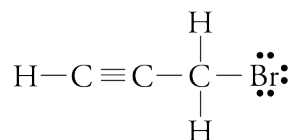
- a. C                                      c. oxygen  
b. phosphorus                      d. Br

### Section 5.5 Drawing Lewis Structures

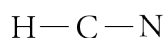
102. Copy the following Lewis structure and identify the single bonds, the double bond, and the lone pairs.



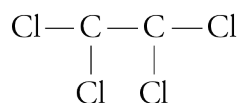
103. Copy the following Lewis structure and identify the single bonds, the triple bond, and the lone pairs.



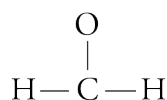
104. Draw a Lewis structure for each of the following formulas.
- oxygen difluoride,  $\text{OF}_2$  (an unstable, colorless gas)
  - bromoform,  $\text{CHBr}_3$  (used as a sedative)
  - phosphorus triiodide,  $\text{PI}_3$  (used to make organic compounds)
105. Draw a Lewis structure for each of the following formulas.
- nitrogen trifluoride,  $\text{NF}_3$  (used in high-energy fuels)
  - chloroethane,  $\text{C}_2\text{H}_5\text{Cl}$  (used to make the gasoline additive tetraethyl lead)
  - hypobromous acid,  $\text{HOBr}$  (used as a wastewater disinfectant)
106. Draw Lewis structures for the following compounds by adding any necessary lines and dots to the skeletons given.
- hydrogen cyanide,  $\text{HCN}$  (used to manufacture dyes and pesticides)



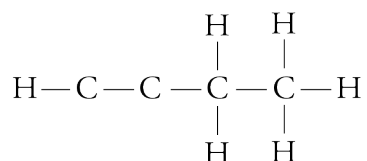
- dichloroethene,  $\text{C}_2\text{Cl}_4$  (used to make perfumes)



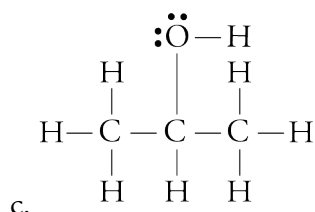
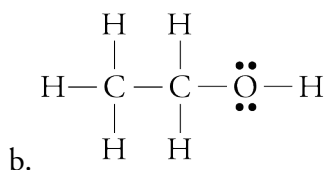
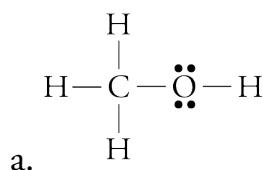
107. Draw Lewis structures for the following compounds by adding any necessary lines and dots to the skeletons given.
- formaldehyde,  $\text{H}_2\text{CO}$  (used in embalming fluids)



- 1-butyne,  $\text{C}_4\text{H}_6$  (a specialty fuel)



108. Write two different names for each of the following alcohols.



OBJECTIVE 21

OBJECTIVE 22

OBJECTIVE 21

OBJECTIVE 22

OBJECTIVE 21

OBJECTIVE 22



OBJECTIVE 21

OBJECTIVE 23

OBJECTIVE 24

- 109.** For each of the following molecular compounds, identify the atom that is most likely to be found in the center of its Lewis structure. Explain why.  
 a.  $\text{CBr}_4$       b.  $\text{SO}_2$       c.  $\text{H}_2\text{S}$       d.  $\text{NOF}$
- 110.** For each of the following molecular compounds, identify the atom that is most likely to be found in the center of its Lewis structure. Explain why.  
 a.  $\text{BI}_3$       b.  $\text{SO}_3$       c.  $\text{AsH}_3$       d.  $\text{HCN}$
- 111.** Calculate the total number of valence electrons for each of the following formulas.  
 a.  $\text{HNO}_3$       b.  $\text{CH}_2\text{CHF}$
- 112.** Calculate the total number of valence electrons for each of the following formulas.  
 a.  $\text{H}_3\text{PO}_4$       b.  $\text{HC}_2\text{H}_3\text{O}_2$
- 113.** Draw a reasonable Lewis structure for each of the following formulas.  
 a.  $\text{Cl}_4$       b.  $\text{O}_2\text{F}_2$       c.  $\text{HC}_2\text{F}$       d.  $\text{NH}_2\text{Cl}$   
 e.  $\text{PH}_3$       f.  $\text{S}_2\text{F}_2$       g.  $\text{HNO}_2$       h.  $\text{N}_2\text{F}_4$   
 i.  $\text{CH}_2\text{CHCH}_3$
- 114.** Draw a reasonable Lewis structure for each of the following formulas.  
 a.  $\text{H}_2\text{S}$       b.  $\text{CHBr}_3$       c.  $\text{NF}_3$       d.  $\text{Br}_2\text{O}$   
 e.  $\text{H}_2\text{CO}_3$       f.  $\text{H}_2\text{S}_2$       g.  $\text{HOCl}$       h.  $\text{BBr}_3$   
 i.  $\text{CH}_3\text{CH}_2\text{CHCH}_2$

OBJECTIVE 21

OBJECTIVE 22

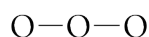
OBJECTIVE 21

OBJECTIVE 22

### Section 5.6 Resonance

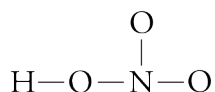
OBJECTIVE 25

- 115.** Draw a reasonable Lewis structure for the ozone molecule,  $\text{O}_3$ , using the skeleton below. The structure is best described in terms of resonance, so draw all of its reasonable resonance structures and the resonance hybrid that summarizes these structures.



OBJECTIVE 25

- 116.** Draw a reasonable Lewis structure for a nitric acid,  $\text{HNO}_3$ , using the skeleton below. The structure is best described in terms of resonance, so draw all of its reasonable resonance structures and the resonance hybrid that summarizes these structures.



### Section 5.7 Molecular Geometry from Lewis Structures

OBJECTIVE 27

- 117.** Explain why the atoms in the  $\text{CH}_4$  molecule are arranged with a tetrahedral molecular shape.

OBJECTIVE 28

OBJECTIVE 29

- 118.** Compare and contrast the information given in the Lewis structure, the space-filling model, the ball-and-stick model, and the geometric sketch of a methane molecule,  $\text{CH}_4$ .

OBJECTIVE 28

OBJECTIVE 29

- 119.** Compare and contrast the information given in the Lewis structure, the space-filling model, the ball-and-stick model, and the geometric sketch of an ammonia molecule,  $\text{NH}_3$ .

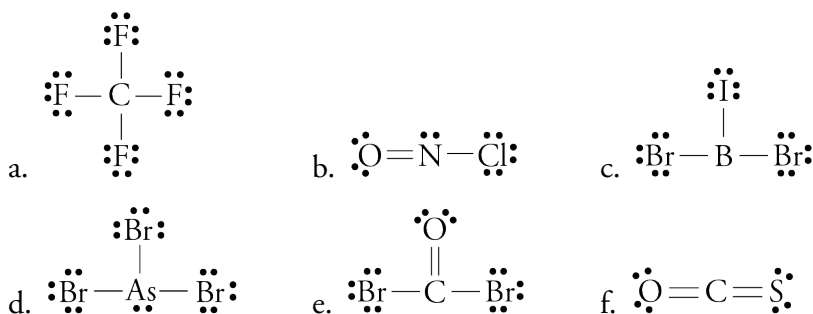
OBJECTIVE 28

OBJECTIVE 29

- 120.** Compare and contrast the information given in the Lewis structure, the space-filling model, the ball-and-stick model, and the geometric sketch of a water molecule,  $\text{H}_2\text{O}$ .

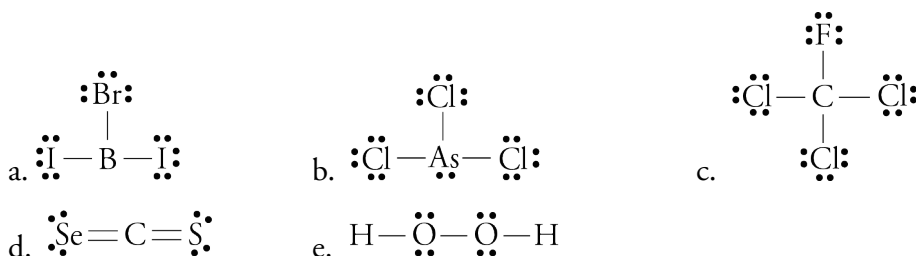
121. Although both  $\text{CO}_2$  molecules and  $\text{H}_2\text{O}$  molecules have three atoms,  $\text{CO}_2$  molecules are linear, and  $\text{H}_2\text{O}$  molecules are bent. Why?
122. Although both  $\text{BF}_3$  molecules and  $\text{NH}_3$  molecules have four atoms, the  $\text{BF}_3$  molecules are planar, and  $\text{NH}_3$  molecules are pyramidal. Why?
123. Using the symbol X for the central atom and Y for the outer atoms, draw the general geometric sketch for a three-atom molecule with linear geometry.
124. Using the symbol X for the central atom and Y for the outer atoms, draw the general geometric sketch for a molecule with trigonal planar geometry.
125. Using the symbol X for the central atom and Y for the outer atoms, draw the general geometric sketch for a molecule with tetrahedral geometry.
126. For each of the Lewis structures given below, (1) write the name of the electron group geometry around each atom that has two or more atoms attached to it, (2) draw a geometric sketch of the molecule, including bond angles (or approximate bond angles), and (3) write the name of the molecular geometry around each atom that has two or more atoms attached to it.

OBJECTIVE 30



127. For each of the Lewis structures given below, (1) write the name of the electron group geometry around each atom that has two or more atoms attached to it, (2) draw a geometric sketch of the molecule, including bond angles (or approximate bond angles), and (3) write the name of the molecular geometry around each atom that has two or more atoms attached to it.

OBJECTIVE 30



### Discussion Topic

128. It has been suggested that there is really only one type of chemical bond—that ionic and covalent bonds are not really fundamentally different. What arguments can be made for and against this position?