

Covalent Bonding

BIG Idea Covalent bonds form when atoms share electrons.

8.1 The Covalent Bond

MAIN Idea Atoms gain stability when they share electrons and form covalent bonds.

8.2 Naming Molecules

MAIN Idea Specific rules are used when naming binary molecular compounds, binary acids, and oxyacids.

8.3 Molecular Structures

MAIN Idea Structural formulas show the relative positions of atoms within a molecule.

8.4 Molecular Shapes

MAIN Idea The VSEPR model is used to determine molecular shape.

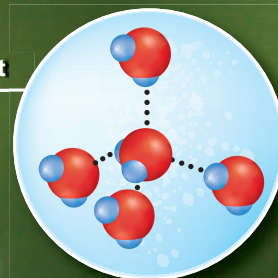
8.5 Electronegativity and Polarity

MAIN Idea A chemical bond's character is related to each atom's attraction for the electrons in the bond.

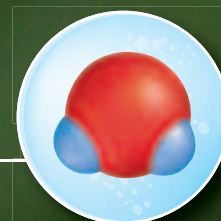
ChemFacts

- The spherical shape of a water drop is due to surface tension, a phenomenon caused by forces between molecules.
- Surface tension makes water act somewhat like an elastic film. Insects called water striders are able to walk on the filmlike surface of water.
- The chemical and physical properties of water make it a unique liquid.

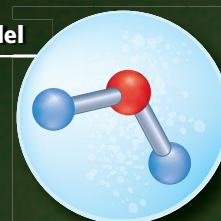
Spherical water droplet



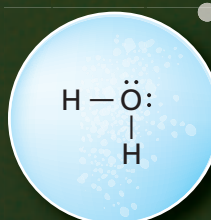
Space-filling model



Ball-and-stick model



Lewis structure



Start-Up Activities

LAUNCH Lab

What type of compound is used to make a Super Ball?

Super Balls are often made of a silicon compound called organosilicon oxide ($\text{Si}(\text{OCH}_2\text{CH}_3)_2\text{O}$).



Procedure

1. Read and complete the lab safety form.
2. Spread several **paper towels** across your desk or lab work area. Put on **lab gloves**. Place a **paper cup** on the paper towels.
3. Using a **graduated cylinder**, measure 20.0 mL of **sodium silicate solution**, and pour it into the cup. Add one drop of **food coloring** and 10.0 mL of **ethanol** to the cup. Stir the mixture clockwise with a **wooden splint** for 3 s.
WARNING: Keep ethanol away from flame and spark sources, as its vapors can be explosive.
4. Working over paper towels, pour the mixture onto one of your glove-covered palms. Gently squeeze out excess liquid as the mixture solidifies.
5. Roll the solid between glove-covered hands and form a ball. Drop it on the floor and observe what happens.
6. Store the ball in an airtight container. You will need to reshape the ball before using it again.

Analysis

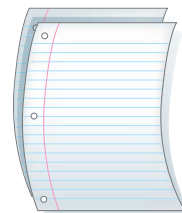
1. **Describe** the properties of the ball that you observed.
2. **Compare** the properties you observed with those of an ionic compound.

Inquiry How many electrons do silicon and oxygen atoms need to form octets? If both atoms must gain electrons, how can they form a bond with each other?

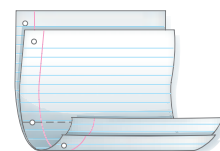
FOLDABLES™ Study Organizer

Bond Character Make the following Foldable to help you organize your study of the three major types of bonding.

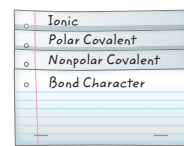
- ▶ **STEP 1** Collect two sheets of paper, and layer them about 2 cm apart vertically.



- ▶ **STEP 2** Fold up the bottom edges of the sheets to form three equal tabs. Crease the fold to hold the tabs in place.



- ▶ **STEP 3** Staple along the fold. Label the tabs as follows: *Bond Character*, *Nonpolar Covalent*, *Polar Covalent*, and *Ionic*.



FOLDABLES Use this Foldable with Section 8.1. As you read this section, summarize what you learn about bond character and how it affects the properties of compounds.

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Section 8.1

Objectives

- ▶ **Apply** the octet rule to atoms that form covalent bonds.
- ▶ **Describe** the formation of single, double, and triple covalent bonds.
- ▶ **Contrast** sigma and pi bonds.
- ▶ **Relate** the strength of a covalent bond to its bond length and bond dissociation energy.

Review Vocabulary

chemical bond: the force that holds two atoms together

New Vocabulary

covalent bond
molecule
Lewis structure
sigma bond
pi bond
endothermic reaction
exothermic reaction

The Covalent Bond

MAIN Idea Atoms gain stability when they share electrons and form covalent bonds.

Real-World Reading Link Have you ever run in a three-legged race? Each person in the race shares one of their legs with a teammate to form a single three-legged team. In some ways, a three-legged race mirrors how atoms share electrons and join together as a unit.

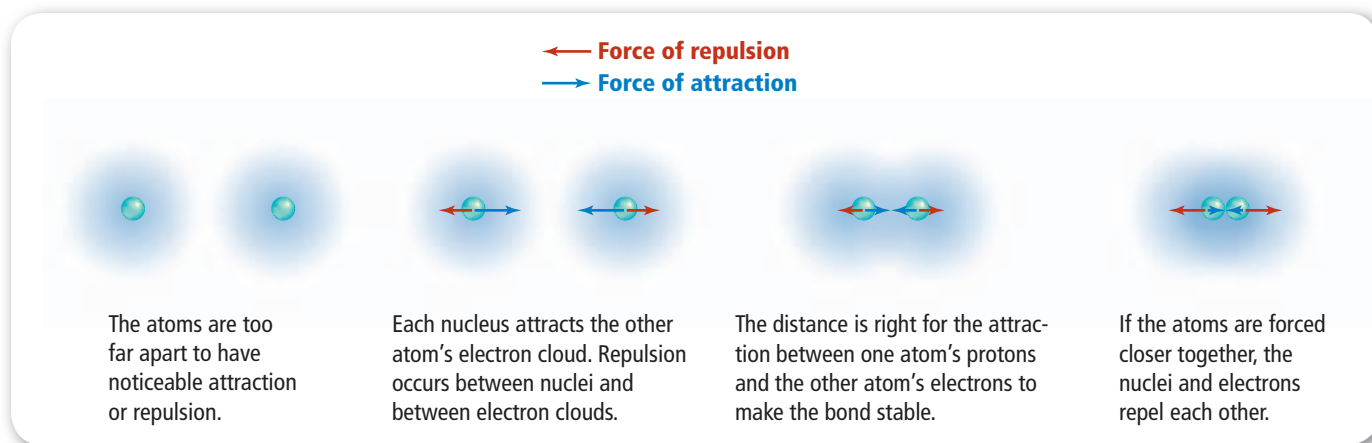
Why do atoms bond?

Understanding the bonding in compounds is essential to developing new chemicals and technologies. To understand why new compounds form, recall what you know about elements that do not tend to form new compounds—the noble gases. You read in Chapter 6 that all noble gases have stable electron arrangements. This stable arrangement consists of a full outer energy level and has lower potential energy than other electron arrangements. Because of their stable configurations, noble gases seldom form compounds.

Gaining stability The stability of an atom, ion, or compound is related to its energy; that is, lower energy states are more stable. In Chapter 7, you read that metals and nonmetals gain stability by transferring (gaining or losing) electrons to form ions. The resulting ions have stable noble-gas electron configurations. From the octet rule in Chapter 6, you know that atoms with a complete octet, a configuration of eight valence electrons, are stable. In this chapter, you will learn that the sharing of valence electrons is another way atoms can acquire the stable electron configuration of noble gases. The water droplets shown in **Figure 8.1** consist of water molecules formed when hydrogen and oxygen atoms share electrons.

■ **Figure 8.1** Each water droplet is made up of water molecules. Each water molecule is made up of two hydrogen atoms and one oxygen atom that have bonded by sharing electrons. The shapes of the drops are due to intermolecular forces acting on the water molecules.





What is a covalent bond?

You just read that atoms can share electrons to form stable electron configurations. How does this occur? Are there different ways in which electrons can be shared? How are the properties of these compounds different from those formed by ions? Read on to answer these questions.

Shared electrons Atoms in nonionic compounds share electrons. The chemical bond that results from sharing valence electrons is a **covalent bond**. A **molecule** is formed when two or more atoms bond covalently. In a covalent bond, the shared electrons are considered to be part of the outer energy levels of both atoms involved. Covalent bonding generally can occur between elements that are near each other on the periodic table. The majority of covalent bonds form between atoms of nonmetallic elements.

Covalent bond formation Diatomic molecules, such as hydrogen (H_2), nitrogen (N_2), oxygen (O_2), fluorine (F_2), chlorine (Cl_2), bromine (Br_2), and iodine (I_2), form when two atoms of each element share electrons. They exist this way because the two-atom molecules are more stable than the individual atoms.

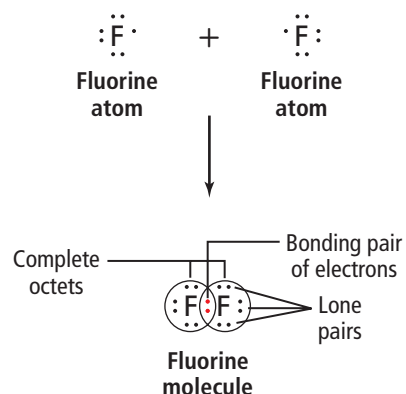
Consider fluorine, which has an electron configuration of $1s^2 2s^2 2p^5$. Each fluorine atom has seven valence electrons and needs another electron to form an octet. As two fluorine atoms approach each other, several forces act, as shown in **Figure 8.2**. Two repulsive forces act on the atoms, one from each atom's like-charged electrons and one from each atom's like-charged protons. A force of attraction also acts, as one atom's protons attract the other atom's electrons. As the fluorine atoms move closer, the attraction of the protons in each nucleus for the other atom's electrons increases until a point of maximum net attraction is achieved. At that point, the two atoms bond covalently and a molecule forms. If the two nuclei move closer, the repulsion forces increase and exceed the attractive forces.

The most stable arrangement of atoms in a covalent bond exists at some optimal distance between nuclei. At this point, the net attraction is greater than the net repulsion. Fluorine exists as a diatomic molecule because the sharing of one pair of electrons gives each fluorine atom a stable noble-gas configuration. As shown in **Figure 8.3**, each fluorine atom in the fluorine molecule has one pair of electrons that are covalently bonded (shared) and three pairs of electrons that are unbonded (not shared). Unbonded pairs are also known as lone pairs.

■ **Figure 8.2** The arrows in this diagram show the net forces of attraction and repulsion acting on two fluorine atoms as they move toward each other. The overall force between two atoms is the result of electron-electron repulsion, nucleus-nucleus repulsion, and nucleus-electron attraction. At the position of maximum net attraction, a covalent bond forms.

Relate How is the stability of the bond related to the forces acting on the atoms?

■ **Figure 8.3** Two fluorine atoms share a pair of electrons to form a covalent bond. Note that the shared electron pair gives each atom a complete octet.



Compare Melting Points

How can you determine the relationship between bond type and melting point? The properties of a compound depend on whether the bonds in the compound are ionic or covalent.

Procedure

1. Read and complete the lab safety form.
2. Create a data table for the experiment.
3. Using a **permanent marker**, draw three lines on the inside bottom of a **disposable, 9-inch aluminum pie pan** to create three, equal wedges. Label the wedges, *A*, *B*, and *C*.
4. Set the pie pan on a **hot plate**.
WARNING: Hot plate and metal pie pan will burn skin—handle with care.
5. Obtain samples of the following from your teacher and deposit them onto the labeled wedges as follows: **sugar crystals** ($C_{12}H_{22}O_{11}$), *A*; **salt crystals** ($NaCl$) *B*; **paraffin** ($C_{23}H_{48}$), *C*.
6. Predict the order in which the compounds will melt.

7. Turn the temperature knob on the hot plate to the highest setting. You will heat the compounds for 5 min. Assign someone to time the heating of the compounds.
8. Observe the compounds during the 5-min period. Record which compounds melt and the order in which they melt.
9. After 5 min, turn off the hot plate and remove the pie pan using a hot mitt or tongs.
10. Allow the pie pan to cool, and then place it in the proper waste container.

Analysis

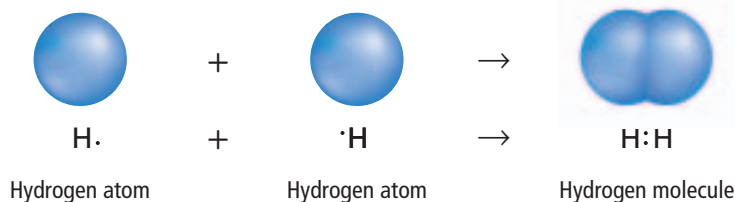
1. **State** Which solid melted first? Which solid did not melt?
2. **Apply** Based on your observations and data, describe the melting point of each solid as low, medium, high, or very high.
3. **Infer** Which compounds are bonded with ionic bonds? Which are bonded with covalent bonds?
4. **Summarize** how the type of bonding affects the melting points of compounds.

Single Covalent Bonds

When only one pair of electrons is shared, such as in a hydrogen molecule, it is a single covalent bond. The shared electron pair is often referred to as the bonding pair. For a hydrogen molecule, shown in **Figure 8.4**, each covalently bonded atom equally attracts the pair of shared electrons. Thus, the two shared electrons belong to each atom simultaneously, which gives each hydrogen atom the noble-gas configuration of helium ($1s^2$) and lower energy. The hydrogen molecule is more stable than either hydrogen atom is by itself.

Recall from chapter 5 that electron-dot diagrams can be used to show valence electrons of atoms. In a **Lewis structure**, they can represent the arrangement of electrons a molecule. A line or a pair of vertical dots between the symbols of elements represents a single covalent bond in a Lewis structure. For example, a hydrogen molecule is written as $H-H$ or $H:H$.

■ **Figure 8.4** When two hydrogen atoms share a pair of electrons, each hydrogen atom is stable because it has a full outer-energy level.



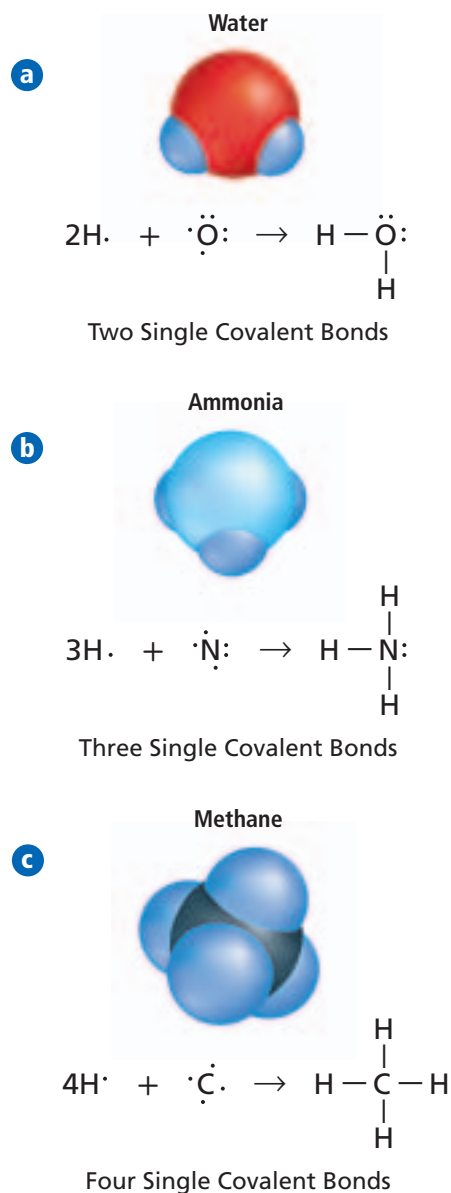
Group 17 and single bonds The halogens—the group 17 elements—such as fluorine have seven valence electrons. To form an octet, one more electron is needed. Therefore, atoms of group 17 elements form single covalent bonds with atoms of other nonmetals, such as carbon. You have already read that the atoms of some group 17 elements form covalent bonds with identical atoms. For example, fluorine exists as F_2 and chlorine exists as Cl_2 .

Group 16 and single bonds An atom of a group 16 element can share two electrons and can form two covalent bonds. Oxygen is a group 16 element with an electron configuration of $1s^2 2s^2 2p^4$. Water is composed of two hydrogen atoms and one oxygen atom. Each hydrogen atom has the noble-gas configuration of helium when it shares one electron with oxygen. Oxygen, in turn, has the noble-gas configuration of neon when it shares one electron with each hydrogen atom. **Figure 8.5a** shows the Lewis structure for a molecule of water. Notice that the oxygen atom has two single covalent bonds and two unshared pairs of electrons.

Group 15 and single bonds Group 15 elements form three covalent bonds with atoms of nonmetals. Nitrogen is a group 15 element with the electron configuration of $1s^2 2s^2 2p^3$. Ammonia (NH_3) has three single covalent bonds. Three nitrogen electrons bond with the three hydrogen atoms leaving one pair of unshared electrons on the nitrogen atom. **Figure 8.5b** shows the Lewis structure for an ammonia molecule. Nitrogen also forms similar compounds with atoms of group 17 elements, such as nitrogen trifluoride (NF_3), nitrogen trichloride (NCl_3), and nitrogen tribromide (NBr_3). Each atom of these group 17 elements and the nitrogen atom share an electron pair.

Group 14 and single bonds Atoms of group 14 elements form four covalent bonds. A methane molecule (CH_4) forms when one carbon atom bonds with four hydrogen atoms. Carbon, a group 14 element, has an electron configuration of $1s^2 2s^2 2p^2$. With four valence electrons, carbon needs four more electrons for a noble gas configuration. Therefore, when carbon bonds with other atoms, it forms four bonds. Because a hydrogen atom, a group 1 element, has one valence electron, it takes four hydrogen atoms to provide the four electrons needed by a carbon atom. The Lewis structure for methane is shown in **Figure 8.5c**. Carbon also forms single covalent bonds with other nonmetal atoms, including those in group 17.

 **Reading Check** Describe how a Lewis structure shows a covalent bond.



■ **Figure 8.5** These chemical equations show how atoms share electrons and become stable. As shown by the Lewis structure for each molecule, all atoms in each molecule achieve a full outer energy level.

Describe For the central atom in each molecule, describe how the octet rule is met.



■ **Figure 8.6** The frosted-looking portions of this glass were chemically etched using hydrogen fluoride (HF), a weak acid. Hydrogen fluoride reacts with silica, the major component of glass, and forms gaseous silicon tetrafluoride (SiF_4) and water.

EXAMPLE Problem 8.1

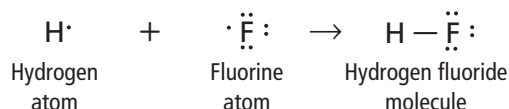
Lewis Structure of a Molecule The pattern on the glass shown in **Figure 8.6** was made by chemically etching its surface with hydrogen fluoride (HF). Draw the Lewis structure for a molecule of hydrogen fluoride.

1 Analyze the Problem

You are given the information that hydrogen and fluorine form the molecule hydrogen fluoride. An atom of hydrogen, a group 1 element, has only one valence electron. It can bond with any nonmetal atom when they share one pair of electrons. An atom of fluorine, a group 17 element, needs one electron to complete its octet. Therefore, a single covalent bond forms when atoms of hydrogen and fluorine bond.

2 Solve for the Unknown

To draw a Lewis structure, first draw the electron-dot diagram for each of the atoms. Then, rewrite the chemical symbols and draw a line between them to show the shared pair of electrons. Finally, add dots to show the unshared electron pairs.



3 Evaluate the Answer

Each atom in the new molecule now has a noble-gas configuration and is stable.

PRACTICE Problems

Extra Practice Page 979 and glencoe.com

Draw the Lewis structure for each molecule.

1. PH_3
2. H_2S
3. HCl
4. CCl_4
5. SiH_4
6. **Challenge** Draw a generic Lewis structure for a molecule formed between atoms of Group 1 and Group 16 elements.

VOCABULARY

ACADEMIC VOCABULARY

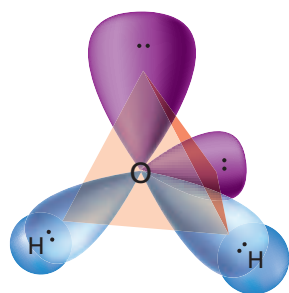
Overlap

to occupy the same area in part
The two driveways overlap at the street forming a common entrance.

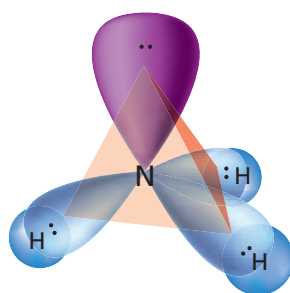
The sigma bond Single covalent bonds are also called **sigma bonds**, represented by the Greek letter sigma (σ). A sigma bond occurs when the pair of shared electrons is in an area centered between the two atoms. When two atoms share electrons, their valence atomic orbitals overlap end to end, concentrating the electrons in a bonding orbital between the two atoms. A bonding orbital is a localized region where bonding electrons will most likely be found. Sigma bonds can form when an s orbital overlaps with another s orbital or a p orbital, or two p orbitals overlap. Water (H_2O), ammonia (NH_3), and methane (CH_4) have sigma bonds, as shown in **Figure 8.7**.



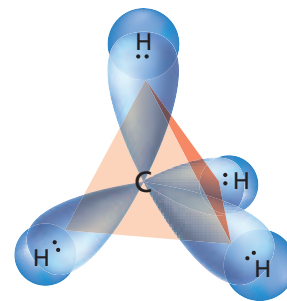
Reading Check List the orbitals that can form sigma bonds in a covalent compound.



Water (H₂O)



Ammonia (NH₃)



Methane (CH₄)

Multiple Covalent Bonds

In some molecules, atoms have noble-gas configurations when they share more than one pair of electrons with one or more atoms. Sharing multiple pairs of electrons forms multiple covalent bonds. A double covalent bond and a triple covalent bond are examples of multiple bonds. Carbon, nitrogen, oxygen, and sulfur atoms often form multiple bonds with other nonmetals. How do you know if two atoms will form a multiple bond? In general, the number of valence electrons needed to form an octet equals the number of covalent bonds that can form.

Double bonds A double covalent bond forms when two pairs of electrons are shared between two atoms. For example, atoms of the element oxygen only exist as diatomic molecules. Each oxygen atom has six valence electrons and must obtain two additional electrons for a noble-gas configuration, as shown in **Figure 8.8a**. A double covalent bond forms when each oxygen atom shares two electrons; a total of two pairs of electrons are shared between the two atoms.

Triple bonds A triple covalent bond forms when three pairs of electrons are shared between two atoms. Diatomic nitrogen (N₂) molecules contain a triple covalent bond. Each nitrogen atom shares three electron pairs, forming a triple bond with the other nitrogen atom as shown in **Figure 8.8b**.

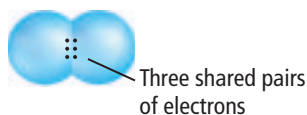
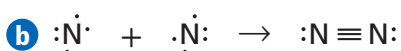
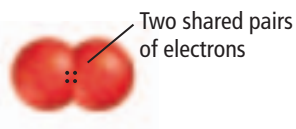
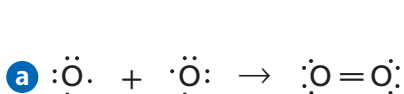
The pi bond A multiple covalent bond consists of one sigma bond and at least one pi bond. A **pi bond**, represented by the Greek letter pi (π), forms when parallel orbitals overlap and share electrons. The shared electron pair of a pi bond occupies the space above and below the line that represents where the two atoms are joined together.

■ **Figure 8.7** Sigma bonds formed in each of these molecules when the atomic orbital of each hydrogen atom overlapped end to end with the orbital of the central atom.

Interpret Identify the types of orbitals that overlap to form the sigma bonds in methane.

FOLDABLES

Incorporate information from this section into your Foldable.

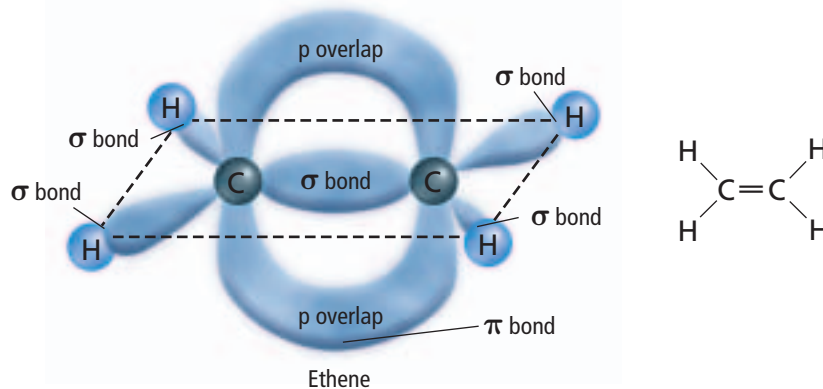


■ **Figure 8.8** Multiple covalent bonds form when two atoms share more than one pair of electrons. **a.** Two oxygen atoms form a double bond. **b.** A triple bond forms between two nitrogen atoms.

Chemistry  **online**

Personal Tutor For an online tutorial on multiple covalent bonds, visit glencoe.com.

■ **Figure 8.9** Notice how the multiple bond between the two carbon atoms in ethene (C_2H_4) consists of a sigma bond and a pi bond. The carbon atoms are close enough that the side-by-side p orbitals overlap and forms the pi bond. This results in a doughnut-shaped cloud around the sigma bond.



Concepts in Motion

Interactive Figure To see an animation of sigma and pi bonding, visit glencoe.com.

It is important to note that molecules having multiple covalent bonds contain both sigma and pi bonds. A double covalent bond, as shown in **Figure 8.9**, consists of one pi bond and one sigma bond. A triple covalent bond consists of two pi bonds and one sigma bond.

The Strength of Covalent Bonds

Recall that a covalent bond involves attractive and repulsive forces. In a molecule, nuclei and electrons attract each other, but nuclei repel other nuclei, and electrons repel other electrons. When this balance of forces is upset, a covalent bond can be broken. Because covalent bonds differ in strength, some bonds break more easily than others. Several factors influence the strength of covalent bonds.

Bond length The strength of a covalent bond depends on the distance between the bonded nuclei. The distance between the two bonded nuclei at the position of maximum attraction is called bond length, as shown in **Figure 8.10**. It is determined by the sizes of the two bonding atoms and how many electron pairs they share. Bond lengths for molecules of fluorine (F_2), oxygen (O_2), and nitrogen (N_2) are listed in **Table 8.1**. Notice that as the number of shared electron pairs increases, the bond length decreases.

Bond length and bond strength are also related: the shorter the bond length, the stronger the bond. Therefore, a single bond, such as that in F_2 , is weaker than a double bond, such as that in O_2 . Likewise, the double bond in O_2 is weaker than the triple bond in N_2 .



Reading Check Relate covalent bond type to bond length.

■ **Figure 8.10** Bond length is the distance from the center of one nucleus to the center of the other nucleus of two bonded atoms.

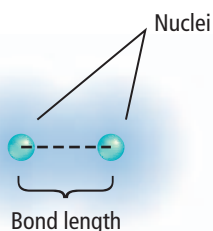


Table 8.1 Covalent Bond Type and Bond Length		
Molecule	Bond Type	Bond Length
F_2	single covalent	1.43×10^{-10} m
O_2	double covalent	1.21×10^{-10} m
N_2	triple covalent	1.10×10^{-10} m

Table 8.2 Bond-Dissociation Energy	
Molecule	Bond-Dissociation Energy
F ₂	159 kJ/mol
O ₂	498 kJ/mol
N ₂	945 kJ/mol

Bonds and energy An energy change occurs when a bond between atoms in a molecule forms or breaks. Energy is released when a bond forms, but energy must be added to break a bond. The amount of energy required to break a specific covalent bond is called bond-dissociation energy and is always a positive value. The bond-dissociation energies for the covalent bonds in molecules of fluorine, oxygen, and nitrogen are listed in **Table 8.2**.

Bond-dissociation energy also indicates the strength of a chemical bond because of the inverse relationship between bond energy and bond length. As indicated in **Table 8.1** and **Table 8.2**, the smaller bond length, the greater the bond-dissociation energy. The sum of the bond-dissociation energy values for all of the bonds in a molecule is the amount of chemical potential energy in a molecule of that compound.

The total energy change of a chemical reaction is determined from the energy of the bonds broken and formed. An **endothermic reaction** occurs when a greater amount of energy is required to break the existing bonds in the reactants than is released when the new bonds form in the products. An **exothermic reaction** occurs when more energy is released during product bond formation than is required to break bonds in the reactants. See **Figure 8.11**.



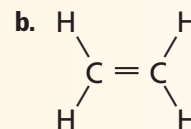
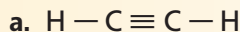
■ **Figure 8.11** Breaking the C–C bonds in charcoal and the O–O bonds in the oxygen in air requires an input of energy. Energy is released as heat and light when bonds form producing CO₂. Thus, the burning of charcoal is an exothermic reaction.

Section 8.1 Assessment

Section Summary

- ▶ Covalent bonds form when atoms share one or more pairs of electrons.
- ▶ Sharing one pair, two pairs, and three pairs of electrons forms single, double, and triple covalent bonds, respectively.
- ▶ Orbitals overlap directly in sigma bonds. Parallel orbitals overlap in pi bonds. A single covalent bond is a sigma bond but multiple covalent bonds are made of both sigma and pi bonds.
- ▶ Bond length is measured nucleus-to-nucleus. Bond-dissociation energy is needed to break a covalent bond.

7. **MAIN Idea** Identify the type of atom that generally forms covalent bonds.
8. **Describe** how the octet rule applies to covalent bonds.
9. **Illustrate** the formation of single, double, and triple covalent bonds using Lewis structures.
10. **Compare and contrast** ionic bonds and covalent bonds.
11. **Contrast** sigma bonds and pi bonds.
12. **Apply** Create a graph using the bond-dissociation energy data in **Table 8.2** and the bond-length data in **Table 8.1**. Describe the relationship between bond length and bond-dissociation energy.
13. **Predict** the relative bond-dissociation energies needed to break the bonds in the structures below.



Section 8.2

Objectives

- **Translate** molecular formulas into binary molecular compound names.
- **Name** acidic solutions.

Review Vocabulary

oxyanion: a polyatomic ion in which an element (usually a nonmetal) is bonded to one or more oxygen atoms

New Vocabulary

oxyacid

Naming Molecules

MAIN Idea Specific rules are used when naming binary molecular compounds, binary acids, and oxyacids.

Real-World Reading Link You probably know that your mother's mother is your grandmother, and that your grandmother's sister is your great-aunt. But what do you call your grandmother's brother's daughter? Naming molecules requires a set of rules, just as naming family relationships requires rules.

Naming Binary Molecular Compounds

Many molecular compounds have common names, but they also have scientific names that reveal their composition. To write the formulas and names of molecules, you will use processes similar to those described in Chapter 7 for ionic compounds.

Start with a binary molecular compound. Note that a binary molecular compound is composed only of two nonmetal atoms—not metal atoms or ions. An example is dinitrogen monoxide (N_2O), a gaseous anesthetic that is more commonly known as nitrous oxide or laughing gas. The naming of nitrous oxide is explained in the following rules.

1. The first element in the formula is always named first, using the entire element name. **N is the symbol for nitrogen.**
2. The second element in the formula is named using its root and adding the suffix *-ide*. **O is the symbol for oxygen so the second word is oxide.**
3. Prefixes are used to indicate the number of atoms of each element that are present in the compound. **Table 8.3** lists the most common prefixes used. **There are two atoms of nitrogen and one atom of oxygen, so the first word is dinitrogen and second word is monoxide.**

There are exceptions to using the prefixes shown in **Table 8.3**. The first element in the compound name never uses the *mono-* prefix. For example, CO is carbon monoxide, not monocarbon monoxide. Also, if using a prefix results in two consecutive vowels, one of the vowels is usually dropped to avoid an awkward pronunciation. For example, notice that the oxygen atom in CO is called monoxide, not monooxide.

Concepts in Motion

Interactive Table Explore naming covalent compounds at glencoe.com.

Table 8.3		Prefixes in Covalent Compounds	
Number of Atoms	Prefix	Number of Atoms	Prefix
1	mono-	6	hexa-
2	di-	7	hepta-
3	tri-	8	octa-
4	tetra-	9	nona-
5	penta-	10	deca-

EXAMPLE Problem 8.2

Naming Binary Molecular Compounds Name the compound P_2O_5 , which is used as a drying and dehydrating agent.

1 Analyze the Problem

You are given the formula for a compound. The formula contains the elements and the number of atoms of each element in one molecule of the compound. Because only two different elements are present and both are nonmetals, the compound can be named using the rules for naming binary molecular compounds.

2 Solve for the Unknown

First, name the elements involved in the compound.

phosphorus

The first element, represented by P, is phosphorus.

oxide

The second element, represented by O, is oxygen.
Add the suffix *-ide* to the root of oxygen, *ox-*.

phosphorus oxide

Combine the names.

Now modify the names to indicate the number of atoms present in a molecule.

diphosphorus pentoxide

From the formula P_2O_5 , you know that two phosphorus atoms and five oxygen atoms make up a molecule of the compound. From Table 8.3, you know that *di-* is the prefix for two and *penta-* is the prefix for five. The *a* in *penta-* is not used because *oxide* begins with a vowel.

3 Evaluate the Answer

The name diphosphorus pentoxide shows that a molecule of the compound contains two phosphorus atoms and five oxygen atoms, which agrees with the compound's chemical formula, P_2O_5 .

PRACTICE Problems

Extra Practice Page 979 and glencoe.com

Name each of the binary covalent compounds listed below.

14. CO_2

15. SO_2


16. NF_3

17. CCl_4

18. **Challenge** What is the formula for diarsenic trioxide?

Common names for some molecular compounds Have you ever enjoyed an icy, cold glass of dihydrogen monoxide on a hot day? You probably have but you most likely called it by its common name, water. Recall from Chapter 7 that many ionic compounds have common names in addition to their scientific ones. For example, baking soda is sodium hydrogen carbonate and common table salt is sodium chloride.

Many binary molecular compounds, such as nitrous oxide and water, were discovered and given common names long before the present-day naming system was developed. Other binary covalent compounds that are generally known by their common names rather than their scientific names are ammonia (NH_3), hydrazine (N_2H_4), and nitric oxide (NO).

 **Reading Check Apply** What are the scientific names for ammonia, hydrazine, and nitric oxide?

Naming Acids

Water solutions of some molecules are acidic and are named as acids. Acids are important compounds with specific properties and will be discussed at length in Chapter 18. If a compound produces hydrogen ions (H^+) in solution, it is an acid. For example, HCl produces H^+ in solution and is an acid. Two common types of acids exist—binary acids and oxyacids.

Naming binary acids A binary acid contains hydrogen and one other element. The naming of the common binary acid known as hydrochloric acid is explained in the following rules.

1. The first word has the prefix *hydro-* to name the hydrogen part of the compound. The rest of the first word consists of a form of the root of the second element plus the suffix *-ic*. **HCl (hydrogen and chlorine) becomes *hydrochloric*.**
2. The second word is always *acid*. **Thus, HCl in a water solution is called *hydrochloric acid*.**

Although the term *binary* indicates exactly two elements, a few acids that contain more than two elements are named according to the rules for naming binary acids. If no oxygen is present in the formula for the acidic compound, the acid is named in the same way as a binary acid, except that the root of the second part of the name is the root of the polyatomic ion that the acid contains. For example, HCN , which is composed of hydrogen and the cyanide ion, is called *hydrocyanic acid* in solution.

Naming oxyacids An acid that contains both a hydrogen atom and an oxyanion is referred to as an **oxyacid**. Recall from Chapter 7 that an oxyanion is a polyatomic ion containing one or more oxygen atoms. The following rules explain the naming of nitric acid (HNO_3), an oxyacid.

1. First, identify the oxyanion present. The first word of an oxyacid's name consists of the root of the oxyanion and the prefix *per-* or *hypo-* if it is part of the name, and a suffix. If the oxyanion's name ends with the suffix *-ate*, replace it with the suffix *-ic*. If the name of the oxyanion ends with the suffix *-ite*, replace it with the suffix *-ous*. **NO_3 , the nitrate ion, becomes *nitric*.**
2. The second word of the name is always *acid*. **HNO_3 (hydrogen and the nitrate ion) becomes *nitric acid*.**

Table 8.4 shows how the names of several oxyacids follow these rules. Notice that the hydrogen in an oxyacid is not part of the name.

Compound	Oxyanion	Acid Suffix	Acid Name
HClO_3	chlorate	-ic	chloric acid
HClO_2	chlorite	-ous	chlorous acid
HNO_3	nitrate	-ic	nitric acid
HNO_2	nitrite	-ous	nitrous acid

Table 8.5
Formulas and Names of Some Covalent Compounds

Formula	Common Name	Molecular Compound Name
H ₂ O	water	dihydrogen monoxide
NH ₃	ammonia	nitrogen trihydride
N ₂ H ₄	hydrazine	dinitrogen tetrahydride
HCl	muriatic acid	hydrochloric acid
C ₉ H ₈ O ₄	aspirin	2-(acetyloxy)benzoic acid

You have learned that naming covalent compounds follows different sets of rules depending on the composition of the compound. **Table 8.5** summarizes the formulas and names of several covalent compounds. Note that an acid, whether a binary acid or an oxyacid, can have a common name in addition to its compound name.

PRACTICE Problems

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Name the following acids. Assume each compound is dissolved in water.

19. HI 20. HClO₃ 21. HClO₂ 22. H₂SO₄ 23. H₂S
 24. **Challenge** What is the formula for periodic acid?

Writing Formulas from Names

The name of a molecular compound reveals its composition and is important in communicating the nature of the compound. Given the name of any binary molecule, you should be able to write the correct chemical formula. The prefixes used in a name indicate the exact number of each atom present in the molecule and determine the subscripts used in the formula. If you are having trouble writing formulas from the names for binary compounds, you might want to review the naming rules listed on pages at the beginning of this section.

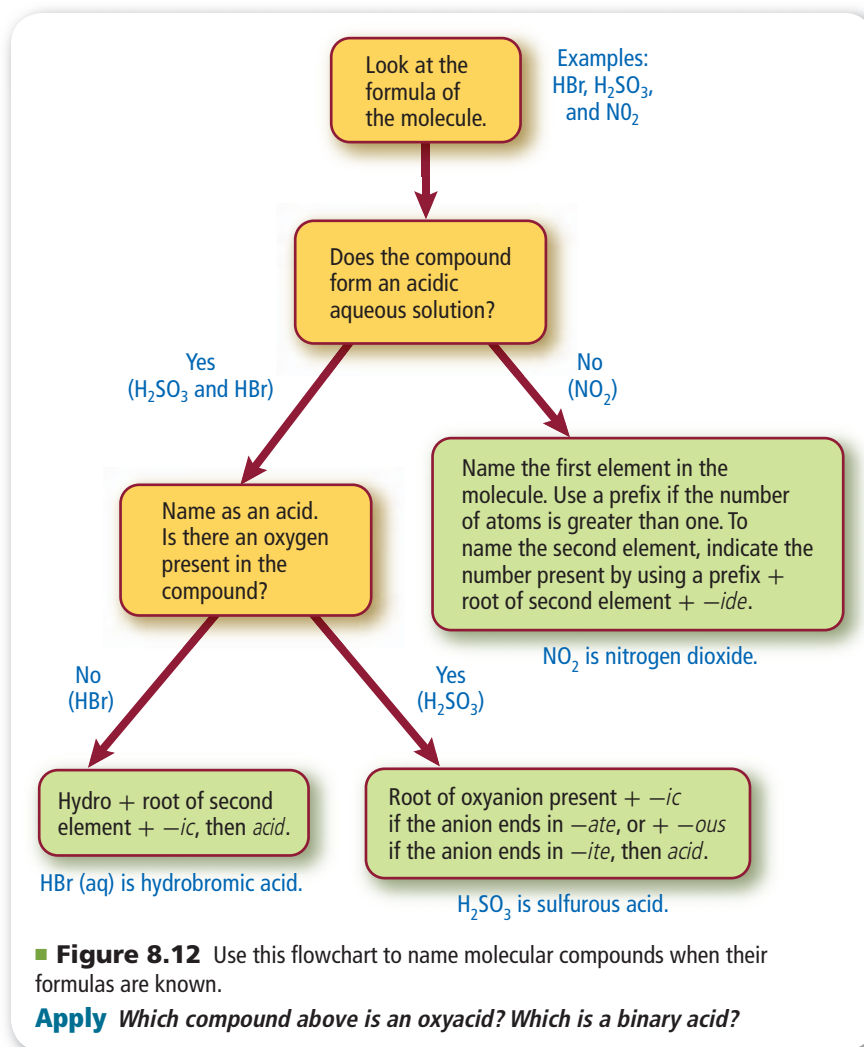
The formula for an acid can also be derived from the name. It is helpful to remember that all binary acids contain hydrogen and one other element. For oxyacids—acids containing oxyanions—you will need to know the names of the common oxyanions. If you need to review oxyanion names, see **Table 7.9** in the previous chapter.

PRACTICE Problems

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Give the formula for each compound.

25. silver chloride
 26. dihydrogen oxide
 27. chlorine trifluoride
 28. diphosphorus trioxide
 29. strontium acetate
 30. **Challenge** What is the formula for carbonic acid?



The flowchart in **Figure 8.12** can help you determine the name of a molecular covalent compound. To use the chart, start at the top and work downward by reading the text contained in the colored boxes and applying it to the formula of the compound you wish to name.

Section 8.2 Assessment

Section Summary

- ▶ Names of covalent molecular compounds include prefixes for the number of each atom present. The final letter of the prefix is dropped if the element name begins with a vowel.
- ▶ Molecules that produce H⁺ in solution are acids. Binary acids contain hydrogen and one other element. Oxyacids contain hydrogen and an oxyanion.

31. **MAIN** <Idea> **Summarize** the rules for naming binary molecular compounds.
32. **Define** a binary molecular compound.
33. **Describe** the difference between a binary acid and an oxyacid.
34. **Apply** Using the system of rules for naming binary molecular compounds, describe how you would name the molecule N₂O₄.
35. **Apply** Write the molecular formula for each of these compounds: *iodic acid*, *disulfur trioxide*, *dinitrogen monoxide*, and *hydrofluoric acid*.
36. **State** the molecular formula for each compound listed below.

a. dinitrogen trioxide	d. chloric acid
b. nitrogen monoxide	e. sulfuric acid
c. hydrochloric acid	f. sulfurous acid

Section 8.3

Objectives

- ▶ **List** the basic steps used to draw Lewis structures.
- ▶ **Explain** why resonance occurs, and identify resonance structures.
- ▶ **Identify** three exceptions to the octet rule, and name molecules in which these exceptions occur.

Review Vocabulary

ionic bond: the electrostatic force that holds oppositely charged particles together in an ionic compound

New Vocabulary

structural formula
resonance
coordinate covalent bond

Molecular Structures

MAIN Idea Structural formulas show the relative positions of atoms within a molecule.

Real-World Reading Link As a child, you might have played with plastic building blocks that connected only in certain ways. If so, you probably noticed that the shape of the object you built depended on the limited ways the blocks interconnected. Building molecules out of atoms works in a similar way.

Structural Formulas

In Chapter 7, you learned about the structure of ionic compounds—substances formed from ionic bonds. The covalent molecules you have read about in this chapter have structures that are different from those of ionic compounds. In studying the molecular structures of covalent compounds, models are used as representations of the molecule.


The molecular formula, which shows the element symbols and numerical subscripts, tells you the type and number of each atom in a molecule. As shown in **Figure 8.13**, there are several different models that can be used to represent a molecule. Note that in the ball-and-stick and space-filling molecular models, atoms of each specific element are represented by spheres of a representative color, as shown in **Table R-1** on page 968. These colors are used for identifying the atoms if the chemical symbol of the element is not present.

One of the most useful molecular models is the **structural formula**, which uses letter symbols and bonds to show relative positions of atoms. You can predict the structural formula for many molecules by drawing the Lewis structure. You have already seen some simple examples of Lewis structures, but more involved structures are needed to help you determine the shapes of molecules.

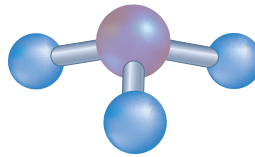
- **Figure 8.13** All of these models can be used to show the relative locations of atoms and electrons in the phosphorus trihydride (phosphine) molecule.
Compare and contrast the types of information contained in each model.

PH_3
Molecular formula

$\begin{array}{c} \text{H} - \ddot{\text{P}} - \text{H} \\ | \\ \text{H} \end{array}$
Lewis structure


Space-filling
molecular model

$\begin{array}{c} \text{H} - \text{P} - \text{H} \\ | \\ \text{H} \end{array}$
Structural formula


Ball-and-stick
molecular model

Lewis structures Although it is fairly easy to draw Lewis structures for most compounds formed by nonmetals, it is a good idea to follow a regular procedure. Whenever you need to draw a Lewis structure, follow the steps outlined in this Problem-Solving Strategy.

Problem-Solving Strategy

Drawing Lewis Structures

1. Predict the location of certain atoms.

The atom that has the least attraction for shared electrons will be the central atom in the molecule. This element is usually the one closer to the left side of the periodic table. The central atom is located in the center of the molecule; all other atoms become terminal atoms.

Hydrogen is always a terminal, or end, atom. Because it can share only one pair of electrons, hydrogen can be connected to only one other atom.

2. Determine the number of electrons available for bonding.

This number is equal to the total number of valence electrons in the atoms that make up the molecule.

3. Determine the number of bonding pairs.

To do this, divide the number of electrons available for bonding by two.

4. Place the bonding pairs.

Place one bonding pair (single bond) between the central atom and each of the terminal atoms.

5. Determine the number of bonding pairs remaining.

To do this, subtract the number of pairs used in Step 4 from the total number of bonding pairs determined in Step 3. These remaining pairs include lone pairs as well as pairs used in double and triple bonds. Place lone pairs around each terminal atom (except H atoms) bonded to the central atom to satisfy the octet rule. Any remaining pairs will be assigned to the central atom.

6. Determine whether the central atom satisfies the octet rule.

Is the central atom surrounded by four electron pairs? If not, it does not satisfy the octet rule. To satisfy the octet rule, convert one or two of the lone pairs on the terminal atoms into a double bond or a triple bond between the terminal atom and the central atom. These pairs are still associated with the terminal atom as well as with the central atom. Remember that carbon, nitrogen, oxygen, and sulfur often form double and triple bonds.

Apply the Strategy

Study Example Problems 8.3 through 8.5 to see how the steps in the Problem-Solving Strategy are applied.

EXAMPLE Problem 8.3

Lewis Structure for a Covalent Compound with Single Bonds Ammonia is a raw material used in the manufacture of many materials, including fertilizers, cleaning products, and explosives. Draw the Lewis structure for ammonia (NH₃).

1 Analyze the Problem

Ammonia molecules consist of one nitrogen atom and three hydrogen atoms. Because hydrogen must be a terminal atom, nitrogen is the central atom.

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Dimensional Analysis
page 956

2 Solve for the Unknown

Find the total number of valence electrons available for bonding.

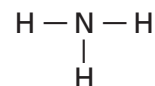
$$1 \text{ N atom} \times \frac{5 \text{ valence electrons}}{1 \text{ N atom}} + 3 \text{ H atoms} \times \frac{1 \text{ valence electron}}{1 \text{ H atom}} \\ = 8 \text{ valence electrons}$$

There are 8 valence electrons available for bonding.

$$\frac{8 \text{ electrons}}{2 \text{ electrons/pair}} = 4 \text{ pairs}$$

Determine the total number of bonding pairs. To do this, divide the number of available electrons by two.

Four pairs of electrons are available for bonding.



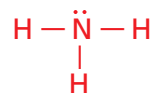
Place a bonding pair (a single bond) between the central nitrogen atom and each terminal hydrogen atom.

Determine the number of bonding pairs remaining.

$$4 \text{ pairs total} - 3 \text{ pairs used} \\ = 1 \text{ pair available}$$

Subtract the number of pairs used in these bonds from the total number of pairs of electrons available.

The remaining pair—a lone pair—must be added to either the terminal atoms or the central atom. Because hydrogen atoms can have only one bond, they have no lone pairs.



Place the remaining lone pair on the central nitrogen atom.

3 Evaluate the Answer

Each hydrogen atom shares one pair of electrons, as required, and the central nitrogen atom shares three pairs of electrons and has one lone pair, providing a stable octet.

PRACTICE Problems

Extra Practice Page 980 and glencoe.com

37. Draw the Lewis structure for BH₃.
38. **Challenge** A nitrogen trifluoride molecule contains numerous lone pairs. Draw its Lewis structure.

EXAMPLE Problem 8.4

Lewis Structure for a Covalent Compound with Multiple Bonds Carbon dioxide is a product of all cellular respiration. Draw the Lewis structure for carbon dioxide (CO_2).

1 Analyze the Problem

The carbon dioxide molecule consists of one carbon atom and two oxygen atoms. Because carbon has less attraction for shared electrons, carbon is the central atom, and the two oxygen atoms are terminal.

2 Solve for the Unknown

Find the total number of valence electrons available for bonding.

$$1 \text{ C atom} \times \frac{4 \text{ valence electrons}}{1 \text{ C atom}} + 2 \text{ O atoms} \times \frac{6 \text{ valence electrons}}{1 \text{ O atom}} \\ = 16 \text{ valence electrons}$$

There are 16 valence electrons available for bonding.

$$\frac{16 \text{ electrons}}{2 \text{ electrons/pair}} = 8 \text{ pairs}$$

Determine the total number of bonding pairs by dividing the number of available electrons by two.

Eight pairs of electrons are available for bonding.

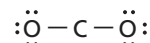


Place a bonding pair (a single bond) between the central carbon atom and each terminal oxygen atom.

Determine the number of bonding pairs remaining. Subtract the number of pairs used in these bonds from the total number of pairs of electrons available.

$$8 \text{ pairs total} - 2 \text{ pairs used} \\ = 6 \text{ pairs available}$$

Subtract the number of pairs used in these bonds from the total number of pairs of electrons available.



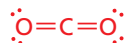
Add three lone pairs to each terminal oxygen atom.

Determine the number of bonding pairs remaining.

$$6 \text{ pairs available} - 6 \text{ pairs used} \\ = 0 \text{ pairs available}$$

Subtract the lone pairs from the pairs available.

Examine the incomplete structure above (showing the placement of the lone pairs). Note that the carbon atom does not have an octet and that there are no more electron pairs available. To give the carbon atom an octet, the molecule must form double bonds.



Use a lone pair from each O atom to form a double bond with the C atom.

3 Evaluate the Answer

Both carbon and oxygen now have an octet, which satisfies the octet rule.



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PRACTICE Problems

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39. Draw the Lewis structure for ethylene, C_2H_4 .

40. Challenge A molecule of carbon disulfide contains both lone pairs and multiple-covalent bonds. Draw its Lewis structure.

Lewis structures for polyatomic ions Although the unit acts as an ion, the atoms within a polyatomic ion are covalently bonded. The procedure for drawing Lewis structures for polyatomic ions is similar to drawing them for covalent compounds. The main difference is in finding the total number of electrons available for bonding. Compared to the number of valence electrons present in the atoms that make up the ion, more electrons are present if the ion is negatively charged and fewer are present if the ion is positive. To find the total number of electrons available for bonding, first find the number available in the atoms present in the ion. Then, subtract the ion charge if the ion is positive, and add the ion charge if the ion is negative.

EXAMPLE Problem 8.5

Lewis Structure for a Polyatomic Ion Draw the correct Lewis structure for the polyatomic ion phosphate (PO_4^{3-}).

1 Analyze the Problem

You are given that the phosphate ion consists of one phosphorus atom and four oxygen atoms and has a charge of $3-$. Because phosphorus has less attraction for shared electrons than oxygen, phosphorus is the central atom and the four oxygen atoms are terminal atoms.

2 Solve for the Unknown

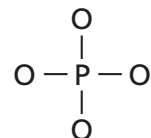
Find the total number of valence electrons available for bonding.

$$1 \text{ P atom} \times \frac{5 \text{ valence electrons}}{\text{P atom}} + 4 \text{ O atoms} \times \frac{6 \text{ valence electrons}}{\text{O atom}}$$

$$+ 3 \text{ electrons from the negative charge} = 32 \text{ valence electrons}$$

$$\frac{32 \text{ electrons}}{2 \text{ electrons/pair}} = 16 \text{ pair}$$

Determine the total number of bonding pairs.



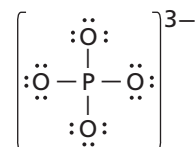
Draw single bonds from each terminal oxygen atom to the central phosphorus atom.

$$16 \text{ pairs total} - 4 \text{ pairs used} \\ = 12 \text{ pairs available}$$

Subtract the number of pairs used from the total number of pairs of electrons available.

Add three lone pairs to each terminal oxygen atom.

$$12 \text{ pairs available} - 12 \text{ lone pairs used} = 0$$



Subtracting the lone pairs used from the pairs available verifies that there are no electron pairs available for the phosphorus atom. The Lewis structure for the phosphate ion is shown.

3 Evaluate the Answer

All of the atoms have an octet, and the group has a net charge of $3-$.

PRACTICE Problems

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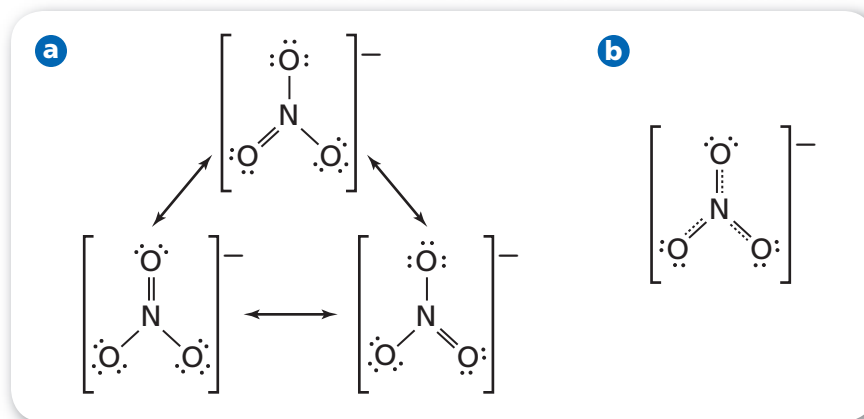
41. Draw the Lewis structure for the NH_4^+ ion.
42. **Challenge** The ClO_4^- ion contains numerous lone pairs. Draw its Lewis structure.

Real-World Chemistry Phosphorus and Nitrogen



Algal blooms Phosphorus and nitrogen are nutrients required for algae growth. Both can enter lakes and streams from discharges of sewage and industrial waste, and in fertilizer runoff. If these substances build up in a body of water, a rapid growth of algae, known as an algal bloom, can occur, forming a thick layer of green slime over the water's surface. When the algae use up the supply of nutrients, they die and decompose. This process reduces the amount of dissolved oxygen in the water that is available to other aquatic organisms.

■ **Figure 8.14** The nitrate ion (NO_3^-) exhibits resonance. **a.** These resonance structures differ only in the location of the double bond. The locations of the nitrogen and oxygen atoms stay the same. **b.** The actual nitrate ion is like an average of the three resonance structures in **a.** The dotted lines indicate possible locations of the double bond.



VOCABULARY

SCIENCE USAGE V. COMMON USAGE

Resonance

Science usage: a phenomenon related to the stability of a molecule; a large vibration in a mechanical system caused by a small periodic stimulus
The new molecule had several resonance structures.

Common usage: a quality of richness or variety
The sound of the orchestra had resonance.

Resonance Structures

Using the same sequence of atoms, it is possible to have more than one correct Lewis structure when a molecule or polyatomic ion has both a double bond and a single bond. Consider the polyatomic ion nitrate (NO_3^-), shown in **Figure 8.14a**. Three equivalent structures can be used to represent the nitrate ion.

Resonance is a condition that occurs when more than one valid Lewis structure can be written for a molecule or ion. The two or more correct Lewis structures that represent a single molecule or ion are referred to as resonance structures. Resonance structures differ only in the position of the electron pairs, never the atom positions. The location of the lone pairs and bonding pairs differs in resonance structures. The molecule O_3 and the polyatomic ions NO_3^- , NO_2^- , SO_3^{2-} , and CO_3^{2-} commonly form resonance structures.

It is important to note that each molecule or ion that undergoes resonance behaves as if it has only one structure. Refer to **Figure 8.14b**. Experimentally measured bond lengths show that the bonds are identical to each other. They are shorter than single bonds but longer than double bonds. The actual bond length is an average of the bonds in the resonance structures.

PRACTICE Problems

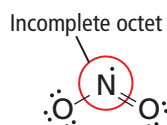
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Draw the Lewis resonance structures for the following molecules.



46. **Challenge** Draw the Lewis resonance structure for the ion SO_3^{2-} .

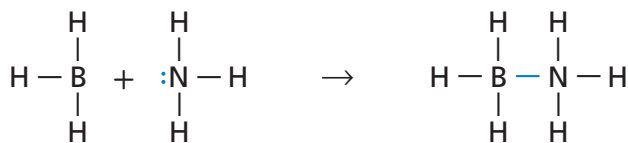
■ **Figure 8.15** The central nitrogen atom in this NO_2 molecule does not satisfy the octet rule; the nitrogen atom has only seven electrons in its outer energy level.



Exceptions to the Octet Rule

Generally, atoms attain an octet when they bond with other atoms. Some molecules and ions, however, do not obey the octet rule. There are several reasons for these exceptions.

Odd number of valence electrons First, a small group of molecules might have an odd number of valence electrons and be unable to form an octet around each atom. For example, NO_2 has five valence electrons from nitrogen and 12 from oxygen, totaling 17, which cannot form an exact number of electron pairs. See **Figure 8.15**. ClO_2 and NO are other examples of molecules with odd numbers of valence electrons.



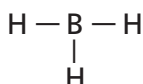
The boron atom has no electrons to share, whereas the nitrogen atom has two electrons to share.

The nitrogen atom shares both electrons to form the coordinate covalent bond.

■ **Figure 8.16** In this reaction between boron trihydride (BH_3) and ammonia (NH_3), the nitrogen atom donates both electrons that are shared by boron and ammonia, forming a coordinate covalent bond.

Interpret Does the coordinate covalent bond in the product molecule satisfy the octet rule?

Suboctets and coordinate covalent bonds Another exception to the octet rule is due to a few compounds that form suboctets—stable configurations with fewer than eight electrons present around an atom. This group is relatively rare, and BH_3 is an example. Boron, a group 3 nonmetal, forms three covalent bonds with other nonmetallic atoms.



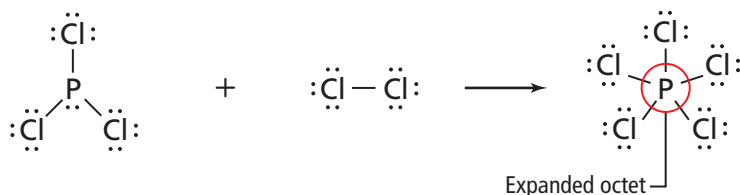
The boron atom shares only six electrons, too few to form an octet. Such compounds tend to be reactive and can share an entire pair of electrons donated by another atom.

A **coordinate covalent bond** forms when one atom donates both of the electrons to be shared with an atom or ion that needs two electrons to form a stable electron arrangement with lower potential energy. Refer to **Figure 8.16**. Atoms or ions with lone pairs often form coordinate covalent bonds with atoms or ions that need two more electrons.

Expanded octets The third group of compounds that does not follow the octet rule has central atoms that contain more than eight valence electrons. This electron arrangement is referred to as an expanded octet. An expanded octet can be explained by considering the d orbital that occurs in the energy levels of elements in period three or higher. An example of an expanded octet, shown in **Figure 8.17**, is the bond formation in the molecule PCl_5 . Five bonds are formed with ten electrons shared in one s orbital, three p orbitals, and one d orbital. Another example is the molecule SF_6 , which has six bonds sharing 12 electrons in an s orbital, three p orbitals, and two d orbitals. When you draw the Lewis structure for these compounds, extra lone pairs are added to the central atom or more than four bonding atoms are present in the molecule.

✓ **Reading Check Summarize** three reasons why some molecules do not conform to the octet rule.

■ **Figure 8.17** Prior to the reaction of PCl_3 and Cl_2 , every reactant atom follows the octet rule. After the reaction, the product, PCl_5 , has an expanded octet containing ten electrons.



EXAMPLE Problem 8.6

Lewis Structure: Exception to the Octet Rule Xenon is a noble gas that will form a few compounds with nonmetals that strongly attract electrons. Draw the correct Lewis structure for xenon tetrafluoride (XeF_4).

1 Analyze the Problem

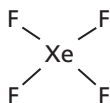
You are given that a molecule of xenon tetrafluoride consists of one xenon atom and four fluorine atoms. Xenon has less attraction for electrons, so it is the central atom.

2 Solve for the Unknown

First, find the total number of valence electrons.

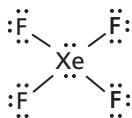
$$1 \text{ Xe atom} \times \frac{8 \text{ valence electrons}}{1 \text{ Xe atom}} + 4 \text{ F atoms} \times \frac{7 \text{ valence electrons}}{1 \text{ F atom}} = 36 \text{ valence electrons}$$

$$\frac{36 \text{ electrons}}{2 \text{ electrons/pair}} = 18 \text{ pairs}$$



18 pairs available – 4 pairs used = 14 pairs available

$$14 \text{ pairs} - 4 \text{ F atoms} \times \frac{3 \text{ pairs}}{1 \text{ F atom}} = 2 \text{ pairs unused}$$



Determine the total number of bonding pairs.

Use four bonding pairs to bond the four F atoms to the central Xe atom.

Determine the number of remaining pairs.

Add three pairs to each F atom to obtain an octet. Determine how many pairs remain.

Place the two remaining pairs on the central Xe atom.

3 Evaluate the Answer

This structure gives xenon 12 total electrons—an expanded octet—for a total of six bond positions. Xenon compounds, such as the XeF_4 shown here, are toxic because they are highly reactive.

PRACTICE Problems

Extra Practice Page 980 and glencoe.com

Draw the expanded octet Lewis structure for each molecule.



49. **Challenge** Draw the Lewis structure for the molecule formed when six fluorine atoms and one sulfur atom bond covalently.

Section 8.3 Assessment

Section Summary

- Different models can be used to represent molecules.
- Resonance occurs when more than one valid Lewis structure exists for the same molecule.
- Exceptions to the octet rule occur in some molecules.

- 50. **MAIN Idea** Describe the information contained in a structural formula.
- 51. **State** the steps used to draw Lewis structures.
- 52. **Summarize** exceptions to the octet rule by correctly pairing these molecules and phrases: odd number of valence electrons, PCl_5 , ClO_2 , BH_3 , expanded octet, less than an octet.
- 53. **Evaluate** A classmate states that a binary compound having only sigma bonds displays resonance. Could the classmate's statement be true?
- 54. **Draw** the resonance structures for the dinitrogen oxide (N_2O) molecule.
- 55. **Draw** the Lewis structures for CN^- , SiF_4 , HCO_3^- , and AsF_6^- .

Section 8.4

Objectives

- Summarize the VSEPR bonding theory.
- Predict the shape of, and the bond angles in, a molecule.
- Define hybridization.

Review Vocabulary

atomic orbital: the region around an atom's nucleus that defines an electron's probable location

New Vocabulary

VSEPR model
hybridization

Molecular Shapes

MAIN Idea The VSEPR model is used to determine molecular shape.

Real-World Reading Link Have you ever rubbed two balloons in your hair to create a static electric charge on them? If you brought the balloons together, their like charges would cause them to repel each other. Molecular shapes are also affected by the forces of electric repulsion.

VSEPR Model

The shape of a molecule determines many of its physical and chemical properties. Often, shapes of reactant molecules determine whether or not they can get close enough to react. Electron densities created by the overlap of the orbitals of shared electrons determine molecular shape. Theories have been developed to explain the overlap of bonding orbitals and can be used to predict the shape of the molecule.

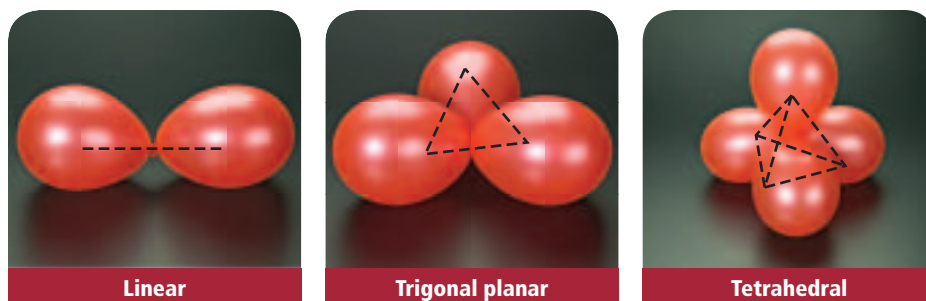
The molecular geometry, or shape, of a molecule can be determined once a Lewis structure is drawn. The model used to determine the molecular shape is referred to as the **Valence Shell Electron Pair Repulsion model**, or **VSEPR model**. This model is based on an arrangement that minimizes the repulsion of shared and unshared electron pairs around the central atom.

Bond angle To understand the VSEPR model better, imagine balloons that are inflated to similar sizes and tied together, as shown in **Figure 8.18**. Each balloon represents an electron-dense region. The repulsive force of this electron-dense region keeps other electrons from entering this space. When a set of balloons is connected at a central point, which represents a central atom, the balloons naturally form a shape that minimizes interactions between the balloons.

The electron pairs in a molecule repel one another in a similar way. These forces cause the atoms in a molecule to be positioned at fixed angles relative to one another. The angle formed by two terminal atoms and the central atom is a bond angle. Bond angles predicted by VSEPR are supported by experimental evidence.

Unshared pairs of electrons are also important in determining the shape of the molecule. These electrons occupy a slightly larger orbital than shared electrons. Therefore, shared bonding orbitals are pushed together by unshared pairs.

■ **Figure 8.18** Electron pairs in a molecule are located as far apart as they can be, just as these balloons are arranged. Two pairs form a linear shape. Three pairs form a trigonal planar shape. Four pairs form a tetrahedral shape.



VOCABULARY

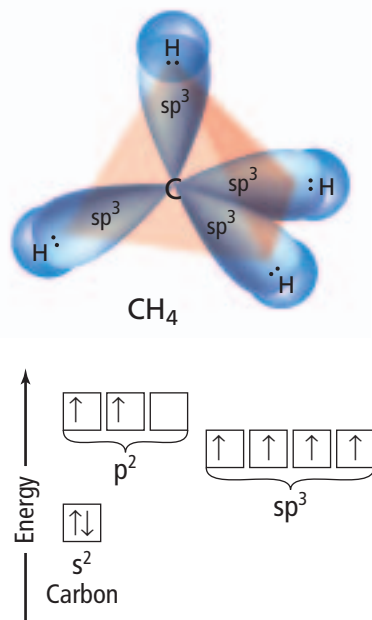
WORD ORIGIN

Trigonal planar

comes from the Latin words *trigonum*, which means *triangular*, and *plan-*, which means *flat*

■ **Figure 8.19** A carbon atom's 2s and 2p electrons occupy the hybrid sp^3 orbitals. Notice that the hybrid orbitals have an intermediate amount of potential energy when compared with the energy of the original s and p orbitals. According to VSEPR theory, a tetrahedral shape minimizes repulsion between the hybrid orbitals in a CH_4 molecule.

Identify How many faces does the tetrahedral shape formed by the sp^3 orbitals have?



Concepts in Motion

Interactive Figure To see an animation of molecular shapes, visit glencoe.com.

Connection to Biology

The shape of food molecules is important to our sense of taste. The surface of your tongue is covered with taste buds, each of which contains from 50 to 100 taste receptor cells. Taste receptor cells can detect five distinct tastes—sweet, bitter, salty, sour, and umami (the taste of MSG, monosodium glutamate)—but each receptor cell responds best to only one taste.

The shapes of food molecules are determined by their chemical structures. When a molecule enters a taste bud, it must have the correct shape for the nerve in each receptor cell to respond and send a message to the brain. The brain then interprets the message as a certain taste. When such molecules bind to sweet receptors, they are sensed as sweet. The greater the number of food molecules that fit a sweet receptor cell, the sweeter the food tastes. Sugars and artificial sweeteners are not the only sweet molecules. Some proteins found in fruits are also sweet molecules. Some common molecular shapes are illustrated in **Table 8.5**.

Hybridization

A hybrid occurs when two things are combined and the result has characteristics of both. For example, a hybrid automobile uses both gas and electricity as energy sources. During chemical bonding, different atomic orbitals undergo hybridization. To understand this, consider the bonding involved in the methane molecule (CH_4). The carbon atom has four valence electrons with the electron configuration $[He]2s^2 2p^2$. You might expect the two unpaired p electrons to bond with other atoms and the 2s electrons to remain an unshared pair. However, carbon atoms undergo **hybridization**, a process in which atomic orbitals mix and form new, identical hybrid orbitals.

The hybrid orbitals in a carbon atom are shown in **Figure 8.19**. Note that each hybrid orbital contains one electron that it can share with another atom. The hybrid orbital is called an sp^3 orbital because the four hybrid orbitals form from one s orbital and three p orbitals. Carbon is the most common element that undergoes hybridization.

The number of atomic orbitals that mix and form the hybrid orbital equals the total number of pairs of electrons, as shown in **Table 8.5**. In addition, the number of hybrid orbitals formed equals the number of atomic orbitals mixed. For example, $AlCl_3$ has a total of three pairs of electrons and VSEPR predicts a trigonal planar molecular shape. This shape results when one s and two p orbitals on the central atom, Al, mix and form three identical sp^2 hybrid orbitals.

Lone pairs also occupy hybrid orbitals. Compare the hybrid orbitals of $BeCl_2$ and H_2O in **Table 8.6**. Both compounds contain three atoms. Why does an H_2O molecule contain sp^3 orbitals? There are two lone pairs on the central oxygen atom in H_2O . Therefore, there must be four hybrid orbitals—two for bonding and two for the lone pairs.

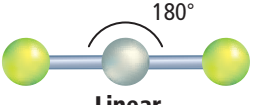
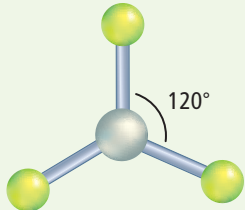
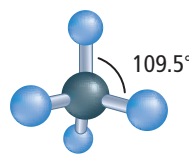
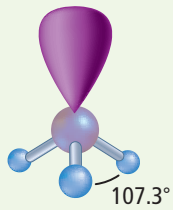
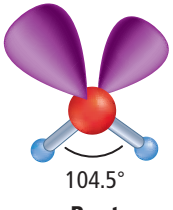
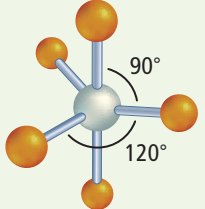
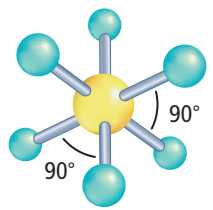
Recall from Section 8.1 that multiple covalent bonds consist of one sigma bond and one or more pi bonds. Only the two electrons in the sigma bond occupy hybrid orbitals such as sp and sp^2 . The remaining unhybridized p orbitals overlap to form pi bonds. It is important to note that single, double, and triple covalent bonds contain only one hybrid orbital. Thus, CO_2 , with two double bonds, forms sp hybrid orbitals.



Reading Check State the number of electrons that are available for bonding in a hybrid sp^3 orbital.

Table 8.6 Molecular Shapes

Interactive Table Explore molecular shapes at glencoe.com.

Molecule	Total Pairs	Shared Pairs	Lone Pairs	Hybrid Orbitals	Molecular Shape*
BeCl ₂	2	2	0	sp	 <p>Linear</p>
AlCl ₃	3	3	0	sp ²	 <p>Trigonal planar</p>
CH ₄	4	4	0	sp ³	 <p>Tetrahedral</p>
PH ₃	4	3	1	sp ³	 <p>Trigonal pyramidal</p>
H ₂ O	4	2	2	sp ³	 <p>Bent</p>
NbBr ₅	5	5	0	sp ³ d	 <p>Trigonal bipyramidal</p>
SF ₆	6	6	0	sp ³ d ²	 <p>Octahedral</p>

The BeCl₂ molecule contains only two pairs of electrons shared with the central Be atom. These bonding electrons have the maximum separation, a bond angle of 180°, and the molecular shape is linear.

The three bonding electron pairs in AlCl₃ have maximum separation in a trigonal planar shape with 120° bond angles.

When the central atom in a molecule has four pairs of bonding electrons, as CH₄ does, the shape is tetrahedral. The bond angles are 109.5°.

PH₃ has three single covalent bonds and one lone pair. The lone pair takes up a greater amount of space than the shared pairs. There is stronger repulsion between the lone pair and the bonding pairs than between two bonding pairs. The resulting geometry is trigonal pyramidal, with 107.3° bond angles.

Water has two covalent bonds and two lone pairs. Repulsion between the lone pairs causes the angle to be 104.5°, less than both tetrahedral and trigonal pyramidal. As a result, water molecules have a bent shape.

The NbBr₅ molecule has five pairs of bonding electrons. The trigonal bipyramidal shape minimizes the repulsion of these shared electron pairs.

As with NbBr₅, SF₆ has no unshared electron pairs on the central atom. However, six shared pairs arranged about the central atom result in an octahedral shape.

*Balls represent atoms, sticks represent bonds, and lobes represent lone pairs of electrons.

EXAMPLE Problem 8.7

Find the Shape of a Molecule Phosphorus trihydride, a colorless gas, is produced when organic materials, such as fish flesh, rot. What is the shape of a phosphorus trihydride molecule? Identify the bond angle size and hybrid orbitals.

1 Analyze the Problem

You are given the information that a phosphorus trihydride molecule has three, terminal hydrogen atoms bonded to a central phosphorus atom.

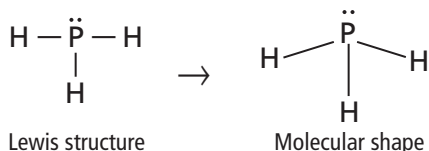
2 Solve for the Unknown

Find the total number of valence electrons and the number of electron pairs.

$$1 \text{ P atom} \times \frac{5 \text{ valence electrons}}{1 \text{ P atom}} + 3 \text{ H atoms} \times \frac{1 \text{ valence electron}}{1 \text{ H atom}} = 8 \text{ valence electrons}$$

$$\frac{8 \text{ electrons}}{2 \text{ electrons/pair}} = 4 \text{ pairs}$$

Determine the total number of bonding pairs.



Lewis structure

Molecular shape

Draw the Lewis structure, using one pair of electrons to bond each H atom to the central P atom and assigning the lone pair to the P atom.

The molecular shape is trigonal pyramidal with a 107° bond angle and sp^3 hybrid orbitals.

3 Evaluate the Answer

All electron pairs are used and each atom has a stable electron configuration.

PRACTICE Problems

Extra Practice Page 980 and glencoe.com

Determine the molecular shape, bond angle, and hybrid orbitals for each molecule.



60. **Challenge** For a NH_4^+ ion, identify its molecular shape, bond angle, and hybrid orbitals.

Section 8.4 Assessment

Section Summary

- ▶ VSEPR model theory states that electron pairs repel each other and determine both the shape of and bond angles in a molecule.
- ▶ Hybridization explains the observed shapes of molecules by the presence of equivalent hybrid orbitals.

- 61. **MAIN Idea** Summarize the VSEPR bonding theory.
- 62. **Define** the term *bond angle*.
- 63. **Describe** how the presence of a lone electron pair affects the spacing of shared bonding orbitals.
- 64. **Compare** the size of an orbital that has a shared electron pair with one that has a lone pair.
- 65. **Identify** the type of hybrid orbitals present and bond angles for a molecule with a tetrahedral shape.
- 66. **Compare** the molecular shapes and hybrid orbitals of PF_3 and PF_5 molecules. Explain why their shapes differ.
- 67. **List** in a table, the Lewis structure, molecular shape, bond angle, and hybrid orbitals for molecules of CS_2 , CH_2O , H_2OSe , CCl_2F_2 , and NCl_3 .

Section 8.5

Objectives

- Describe how electronegativity is used to determine bond type.
- Compare and contrast polar and nonpolar covalent bonds and polar and nonpolar molecules.
- Generalize about the characteristics of covalently bonded compounds.

Review Vocabulary

electronegativity: the relative ability of an atom to attract electrons in a chemical bond

New Vocabulary

polar covalent bond

Figure 8.20 Electronegativity values are derived by comparing an atom's attraction for shared electrons to that of a fluorine's atom attraction for shared electrons. Note that the electronegativity values for the lanthanide and actinide series, which are not shown, range from 1.12 to 1.7.

Electronegativity and Polarity

MAIN Idea A chemical bond's character is related to each atom's attraction for the electrons in the bond.

Real-World Reading Link The stronger you are, the more easily you can do pull-ups. Just as people have different abilities for doing pull-ups, atoms in chemical bonds have different abilities to attract (pull) electrons.

Electron Affinity, Electronegativity, and Bond Character

The type of bond formed during a reaction is related to each atom's attraction for electrons. Electron affinity is a measure of the tendency of an atom to accept an electron. Excluding noble gases, electron affinity increases with increasing atomic number within a period and decreases with increasing atomic number within a group. The scale of electronegativities—shown in **Figure 8.20**—allows chemists to evaluate the electron affinity of specific atoms in a compound. Recall from Chapter 6 that electronegativity indicates the relative ability of an atom to attract electrons in a chemical bond. Note that electronegativity values were assigned, whereas electron affinity values were measured.

Electronegativity The version of the periodic table of the elements shown in **Figure 8.20** lists electronegativity values. Note that fluorine has the greatest electronegativity value (3.98), while francium has the least (0.7). Because noble gases do not generally form compounds, individual electronegativity values for helium, neon, and argon are not listed. However, larger noble gases, such as xenon, sometimes bond with highly electronegative atoms, such as fluorine.

Electronegativity Values for Selected Elements

1 H 2.20											5 B 2.04	6 C 2.55	7 N 3.04	8 O 3.44	9 F 3.98			
3 Li 0.98	4 Be 1.57											13 Al 1.61	14 Si 1.90	15 P 2.19	16 S 2.58	17 Cl 3.16		
11 Na 0.93	12 Mg 1.31	19 K 0.82	20 Ca 1.00	21 Sc 1.36	22 Ti 1.54	23 V 1.63	24 Cr 1.66	25 Mn 1.55	26 Fe 1.83	27 Co 1.88	28 Ni 1.91	29 Cu 1.90	30 Zn 1.65	31 Ga 1.81	32 Ge 2.01	33 As 2.18	34 Se 2.55	35 Br 2.96
37 Rb 0.82	38 Sr 0.95	39 Y 1.22	40 Zr 1.33	41 Nb 1.6	42 Mo 2.16	43 Tc 2.10	44 Ru 2.2	45 Rh 2.28	46 Pd 2.20	47 Ag 1.93	48 Cd 1.69	49 In 1.78	50 Sn 1.96	51 Sb 2.05	52 Te 2.1	53 I 2.66		
55 Cs 0.79	56 Ba 0.89	57 La 1.10	72 Hf 1.3	73 Ta 1.5	74 W 1.7	75 Re 1.9	76 Os 2.2	77 Ir 2.2	78 Pt 2.2	79 Au 2.4	80 Hg 1.9	81 Tl 1.8	82 Pb 1.8	83 Bi 1.9	84 Po 2.0	85 At 2.2		
87 Fr 0.7	88 Ra 0.9	89 Ac 1.1																

	Metal
	Metalloid
	Nonmetal

Table 8.7		EN Difference and Bond Character
Electronegativity Difference	Bond Character	
> 1.7	mostly ionic	
0.4 – 1.7	polar covalent	
< 0.4	mostly covalent	
0	nonpolar covalent	

Bond character A chemical bond between atoms of different elements is never completely ionic or covalent. The character of a bond depends on how strongly each of the bonded atoms attracts electrons. As shown in **Table 8.7**, the character and type of a chemical bond can be predicted using the electronegativity difference of the elements that bond. Electrons in bonds between identical atoms have an electronegativity difference of zero—meaning that the electrons are equally shared between the two atoms. This type of bond is considered nonpolar covalent, or a pure covalent bond. On the other hand, because different elements have different electronegativities, the electron pairs in a covalent bond between different atoms are not shared equally. Unequal sharing results in a **polar covalent bond**. When there is a large difference in the electronegativity between bonded atoms, an electron is transferred from one atom to the other, which results in bonding that is primarily ionic.

Bonding is not often clearly ionic or covalent. An electronegativity difference of 1.70 is considered 50 percent covalent and 50 percent ionic. As the difference in electronegativity increases, the bond becomes more ionic in character. Generally, ionic bonds form when the electronegativity difference is greater than 1.70. However, this cutoff is sometimes inconsistent with experimental observations of two nonmetals bonding together. **Figure 8.21** summarizes the range of chemical bonding between two atoms. What percent ionic character is a bond between two atoms that have an electronegativity difference of 2.00? Where would LiBr be plotted on the graph?



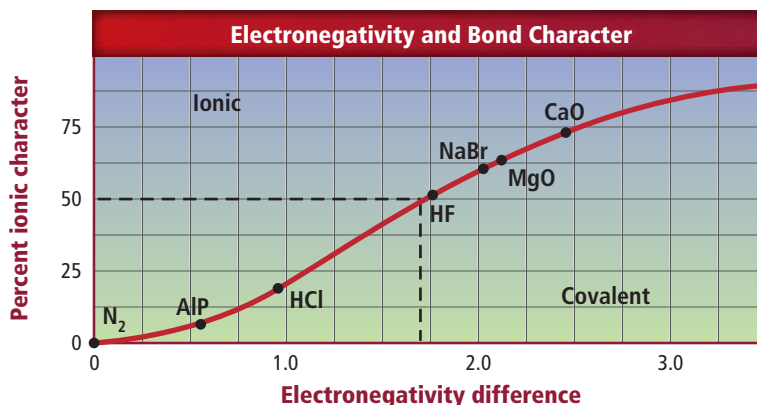
Reading Check Analyze What is the percent ionic character of a pure covalent bond?

■ **Figure 8.21** This graph shows that the difference in electronegativity between bonding atoms determines the percent ionic character of the bond. Above 50% ionic character, bonds are mostly ionic.

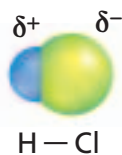


Graph Check

Determine the percent ionic character of calcium oxide.



Electronegativity	Cl = 3.16
Electronegativity	H = 2.20
Difference	= 0.96



Concepts in Motion

Interactive Figure To see an animation of bond types, visit glencoe.com.

■ **Figure 8.22** Chlorine's electronegativity is higher than that of hydrogen. Therefore, in a molecule containing hydrogen and chlorine, the shared pair of electrons is with the chlorine atom more often than it is with the hydrogen atom. Symbols are used to indicate the partial charge at each end of the molecule from this unequal sharing of electrons.

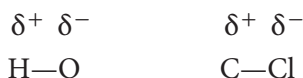
Polar Covalent Bonds

As you just learned, polar covalent bonds form because not all atoms that share electrons attract them equally. A polar covalent bond is similar to a tug-of-war in which the two teams are not of equal strength. Although both sides share the rope, the stronger team pulls more of the rope toward its side. When a polar bond forms, the shared electron pair or pairs are pulled toward one of the atoms. Thus, the electrons spend more time around that atom than the other atom. This results in partial charges at the ends of the bond.


The Greek letter delta (δ) is used to represent a partial charge. In a polar covalent bond, δ^- represent a partial negative charge and δ^+ represents a partial positive charge. As shown in **Figure 8.22**, δ^- and δ^+ can be added to a molecular model to indicate the polarity of the covalent bond. The more-electronegative atom is at the partially negative end, while the less-electronegative atom is at the partially positive end. The resulting polar bond often is referred to as a dipole (two poles).

Molecular polarity Covalently bonded molecules are either polar or nonpolar; which type depends on the location and nature of the covalent bonds in the molecule. A distinguishing feature of nonpolar molecules is that they are not attracted by an electric field. Polar molecules, however, are attracted by an electric field. Because polar molecules are dipoles with partially charged ends, they have an uneven electron density. This results in the tendency of polar molecules to align with an electric field.

Polarity and molecular shape You can learn why some molecules are polar and some are not by comparing water (H_2O) and carbon tetrachloride (CCl_4) molecules. Both molecules have polar covalent bonds. According to the data in **Figure 8.20**, the electronegativity difference between a hydrogen atom and an oxygen atom is 1.24. The electronegativity difference between a chlorine atom and a carbon atom is 0.61. Although these electronegativity differences vary, a H—O bond and a C—Cl bond are considered to be polar covalent.



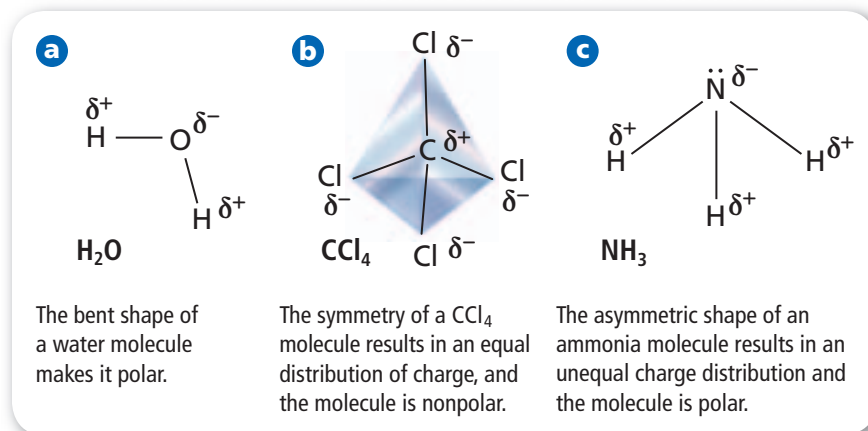
According to their molecular formulas, both molecules have more than one polar covalent bond. However, only the water molecule is polar.

 **Reading Check Apply** Why does a statically charged balloon cause a slow stream of water from a faucet to bend when placed next to it?

CAREERS IN CHEMISTRY

Flavor Chemist A flavor chemist, or flavorist, must know how chemicals react and change in different conditions. A degree in chemistry is an asset, but is not required. Most flavorists work for companies that supply flavors to the food and beverage industries. A certified flavorist trains for five years in a flavor laboratory, passes an oral examination, and then works under supervision for another two years. For more information on chemistry careers, visit glencoe.com.

■ **Figure 8.23** A molecule's shape determines its polarity.



The shape of a H₂O molecule, as determined by VSEPR, is bent because the central oxygen atom has lone pairs of electrons, as shown in **Figure 8.23a**. Because the polar H—O bonds are asymmetric in a water molecule, the molecule has a definite positive end and a definite negative end. Thus, it is polar.

A CCl₄ molecule is tetrahedral, and therefore, symmetrical, as shown in **Figure 8.23b**. The electric charge measured at any distance from its center is identical to the charge measured at the same distance to the opposite side. The average center of the negative charge is located on the chlorine atom. The positive center is also located on the carbon atom. Because the partial charges are balanced, CCl₄ is a nonpolar molecule. Note that symmetric molecules are usually nonpolar, and molecules that are asymmetric are polar as long as the bond type is polar.

Is the molecule of ammonia (NH₃), shown in **Figure 8.23c**, polar? It has a central nitrogen atom and three terminal hydrogen atoms. Its shape is a trigonal pyramidal because of the lone pair of electrons present on the nitrogen atom. Using **Figure 8.20**, you can find that the electronegativity difference of hydrogen and nitrogen is 0.84 making each N—H bond polar covalent. The charge distribution is unequal because the molecule is asymmetric. Thus, the molecule is polar.

Solubility of polar molecules The physical property known as solubility is the ability of a substance to dissolve in another substance. The bond type and the shape of the molecules present determine solubility. Polar molecules and ionic compounds are usually soluble in polar substances, but nonpolar molecules dissolve only in nonpolar substances, as shown in **Figure 8.24**. Solubility is discussed in detail in Chapter 14.

■ **Figure 8.24** Symmetric covalent molecules, such as oil and most petroleum products, are nonpolar. Asymmetric molecules, such as water, are usually polar. As shown in this photo, polar and nonpolar substances usually do not mix.

Infer Will water alone clean oil from a fabric?



Properties of Covalent Compounds

Table salt, an ionic solid, and table sugar, a covalent solid, are similar in appearance. However, these compounds behave differently when heated. Salt does not melt, but sugar melts at a relatively low temperature. Does the type of bonding in a compound affect its properties?

Intermolecular forces Differences in properties are a result of differences in attractive forces. In a covalent compound, the covalent bonds between atoms in molecules are strong, but the attraction forces between molecules are relatively weak. These weak attraction forces are known as intermolecular forces, or van der Waals forces, which are discussed in Chapter 12. Intermolecular forces vary in strength but are weaker than the bonds that join atoms in a molecule or ions in an ionic compound.

There are different types of intermolecular forces. Between nonpolar molecules, the force is weak and is called a dispersion force, or induced dipole. The force between oppositely charged ends of two polar molecules is called a dipole-dipole force. The more polar the molecule, the stronger the dipole-dipole force. The third force, a hydrogen bond, is especially strong. It forms between the hydrogen end of one dipole and a fluorine, oxygen, or nitrogen atom on another dipole.

DATA ANALYSIS LAB

Based on Real Data*

Interpret Data

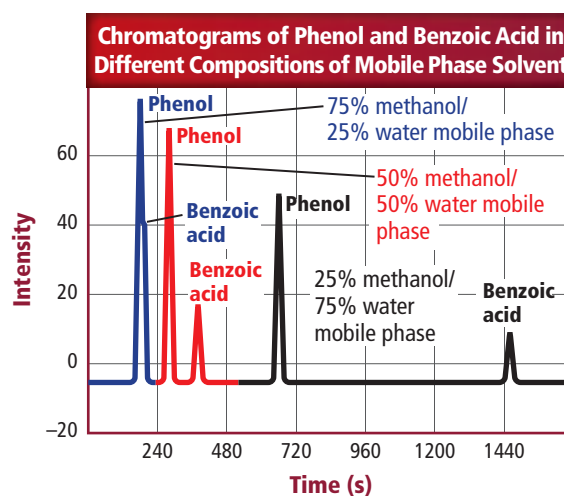
How does the polarity of the mobile phase affect chromatograms? Chromatography is a technique in which a moving phase transports and separates the components of a mixture. A chromatograph is created by recording the intensity of each component carried in the moving phase versus time. The peak intensities on the chromatograph indicate the amount of each component present in the mixture.

High-performance liquid chromatography, or HPLC, is used by analytical chemists to separate mixtures of solutes. During HPLC, components that are strongly attracted to the extracting solvent are retained longer by the moving phase and tend to appear early on a chromatograph. Several scientists performed HPLC using a methanol-water mixture as the extracting solvent to separate a phenol-benzoic acid mixture. Their results are shown in the graph to the right.

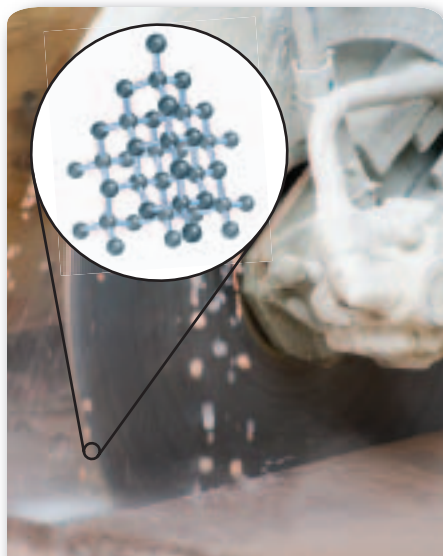
Think Critically

- 1. Explain** the different retention times shown on the chromatograms.
- 2. Infer** from the graph the component, phenol or benzoic acid, that is in excess. Explain your answer.
- 3. Infer** which component of the mixture has more polar molecules.
- 4. Determine** the most effective composition of the mobile phase (of those tested) for separating phenol from benzoic acid. Explain.

Data and Observations



*Data obtained from: Joseph, Seema M. and Palasota, John A. 2001. The combined effects of pH and percent methanol on the HPLC separation of benzoic acid and phenol. *Journal of Chemical Education* 78:1381.



■ **Figure 8.25** Network solids are often used in cutting tools because of their extreme hardness. Here, a diamond-tipped saw blade cuts through stone.

Forces and properties The properties of covalent molecular compounds are related to the relatively weak intermolecular forces holding the molecules together. These weak forces result in the relatively low melting and boiling points of molecular substances compared with those of ionic substances. That is why, when heated moderately, sugar melts but salt does not. Weak intermolecular forces also explain why many molecular substances exist as gases or vaporize readily at room temperature. Oxygen (O_2), carbon dioxide (CO_2), and hydrogen sulfide (H_2S) are examples of covalent gases. Because the hardness of a substance depends on the intermolecular forces between individual molecules, many covalent molecules are relatively soft solids. Paraffin, found in candles and other products, is a common example of a covalent solid.

In the solid phase, molecules align to form a crystal lattice. This molecular lattice is similar to that of an ionic solid, but with less attraction between particles. The structure of the lattice is affected by molecular shape and the type of intermolecular force. Most molecular information has been determined by studying molecular solids.

Covalent Network Solids

There are some solids, often called covalent network solids, that are composed only of atoms interconnected by a network of covalent bonds. Quartz and diamond are two common examples of network solids. In contrast to molecular solids, network solids are typically brittle, non-conductors of heat or electricity, and extremely hard. Analyzing the structure of a diamond explains some of its properties. In a diamond, each carbon atom is bonded to four other carbon atoms. This tetrahedral arrangement, which is shown in **Figure 8.25**, forms a strongly bonded crystal system that is extremely hard and has a very high melting point.

Section 8.5 Assessment

Section Summary

- ▶ The electronegativity difference determines the character of a bond between atoms.
- ▶ Polar bonds occur when electrons are not shared equally forming a dipole.
- ▶ The spatial arrangement of polar bonds in a molecule determines the overall polarity of a molecule.
- ▶ Molecules attract each other by weak intermolecular forces. In a covalent network solid, each atom is covalently bonded to many other atoms.

- 68. **MAIN Idea** **Summarize** how electronegativity difference is related to bond character.
- 69. **Describe** a polar covalent bond.
- 70. **Describe** a polar molecule.
- 71. **List** three properties of a covalent compound in the solid phase.
- 72. **Categorize** bond types using electronegativity difference.
- 73. **Generalize** Describe the general characteristics of covalent network solids.
- 74. **Predict** the type of bond that will form between the following pair of atoms:
 - a. H and S
 - b. C and H
 - c. Na and S.
- 75. **Identify** each molecule as polar or nonpolar: SCl_2 , CS_2 , and CF_4 .
- 76. **Determine** whether a compound made of hydrogen and sulfur atoms is polar or nonpolar.
- 77. **Draw** the Lewis structures for the molecules SF_4 and SF_6 . Analyze each structure to determine whether the molecule is polar or nonpolar.

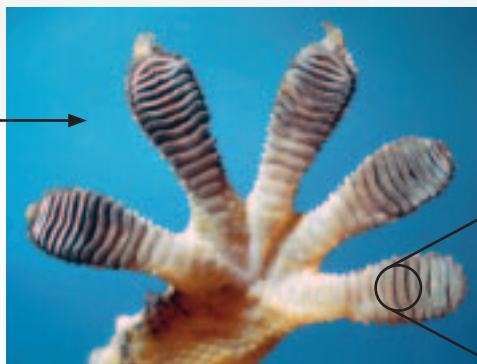
HOW IT WORKS

Sticky Feet: How Geckos Grip

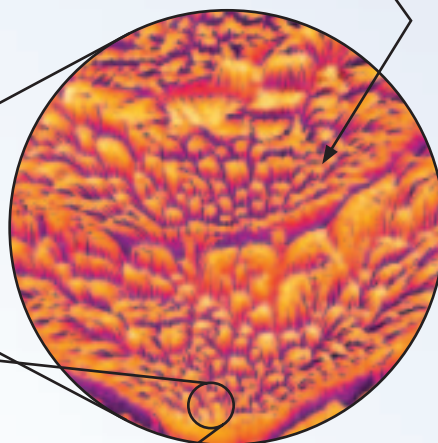
For a gecko, hanging from a wall or a ceiling is no great feat. The key to a gecko's amazing grip is found on each of its toes. Researchers have determined that a gecko's grip depends on the sticking power of atoms themselves.



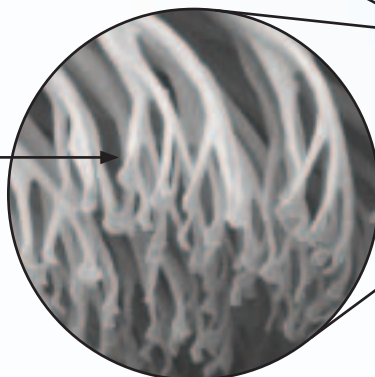
1 Gecko toe The bottom of a gecko's toe is covered with millions of tiny hairs, called setae, arranged in rows.



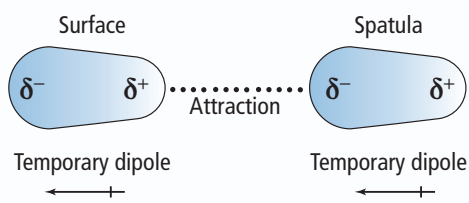
2 Spatulae Setae are complex structures. The end of each seta has microscopic branches called spatulae.



3 Surface area Each seta has a relatively enormous surface area because of its vast number of spatulae.



4 Sticking Van der Waals forces form between a surface and a gecko's spatulae. When multiplied by the spatulae's vast surface areas, the sum of the weak van der Waals forces is more than enough to balance the pull of gravity and hold a gecko in place.



5 Letting go A gecko simply curls its toes when it wants to move. This reduces the amount of surface contact and the van der Waals forces, and a gecko loses its grip.



WRITING in Chemistry

Invent Using their knowledge of how geckos stick to surfaces, scientists are developing applications for geckolike materials. Some possible applications include mini-robots that climb walls and tape that sticks even under water. What uses for a new sticky geckolike material can you think of? For more on gecko-tech, visit glencoe.com.

CHEMLAB

MODEL MOLECULAR SHAPES

Background: Covalent bonding occurs when atoms share valence electrons. In the Valence Shell Electron Pair Repulsion (VSEPR) theory, the way in which valence electrons of bonding atoms are positioned is the basis for predicting a molecule's shape. This method of visualizing shape is also based on the molecule's Lewis structure.

Question: How do the Lewis structure and the positions of valence electrons affect the shape of the covalent compound?

Materials

molecular model kit

Safety Precautions 

Procedure

1. Read and complete the lab safety form.
2. Create a table to record your data.
3. Note and record the color used to represent each of the following atoms in the molecular model kit: hydrogen (H), oxygen (O), phosphorus (P), carbon (C), fluorine (F), sulfur (S), and nitrogen (N).
4. Draw the Lewis structures of the H_2 , O_2 , and N_2 molecules.
5. Obtain two hydrogen atoms and one connector from the molecular model kit, and assemble a hydrogen (H_2) molecule. Observe that your model represents a single-bonded diatomic hydrogen molecule.
6. Obtain two oxygen atoms and two connectors from the molecular model kit, and assemble an oxygen (O_2) molecule. Observe that your model represents a double-bonded diatomic oxygen molecule.
7. Obtain two nitrogen atoms and three connectors from the molecular model kit, and assemble one nitrogen (N_2) molecule. Observe that your model represents a triple-bonded diatomic nitrogen molecule.
8. Recognize that diatomic molecules such as those formed in this lab are always linear. Diatomic molecules are made up of only two atoms and two points (atoms) can only be connected by a straight line.
9. Draw the Lewis structure of water (H_2O), and construct its molecule.



10. Classify the shape of the H_2O molecule using information in **Table 8.6**.
11. Repeat Steps 9 and 10 for the PH_3 , CF_4 , CO_2 , SO_3 , HCN , and CO molecules.

Analyze and Conclude

1. **Think Critically** Based on the molecular models you built and observed in this lab, rank single, double, and triple bonds in order of increasing flexibility and increasing strength.
2. **Observe and Infer** Explain why H_2O and CO_2 molecules have different shapes.
3. **Analyze and Conclude** One of the molecules from this lab undergoes resonance. Identify the molecule that has three resonance structures, draw the structures, and explain why resonance occurs.
4. **Recognize Cause and Effect** Use the electronegativity difference to determine the polarity of the molecules in Steps 9–11. Based on their calculated bond polarities and the models constructed in this lab, determine the molecular polarity of each structure.

INQUIRY EXTENSION

Model Use a molecular model kit to build the two resonance structures of ozone (O_3). Then, use Lewis structures to explain how you can convert between the two resonance structures by interchanging a lone pair for a covalent bond.

**BIG Idea** Covalent bonds form when atoms share electrons.**Section 8.1 The Covalent Bond****MAIN Idea** Atoms gain stability when they share electrons and form covalent bonds.**Vocabulary**

- covalent bond (p. 241)
- endothermic reaction (p. 247)
- exothermic reaction (p. 247)
- Lewis structure (p. 242)
- molecule (p. 241)
- pi bond (p. 245)
- sigma bond (p. 244)

Key Concepts

- Covalent bonds form when atoms share one or more pairs of electrons.
- Sharing one pair, two pairs, and three pairs of electrons forms single, double, and triple covalent bonds, respectively.
- Orbitals overlap directly in sigma bonds. Parallel orbitals overlap in pi bonds. A single covalent bond is a sigma bond but multiple covalent bonds are made of both sigma and pi bonds.
- Bond length is measured nucleus-to-nucleus. Bond dissociation energy is needed to break a covalent bond.

Section 8.2 Naming Molecules**MAIN Idea** Specific rules are used when naming binary molecular compounds, binary acids, and oxyacids.**Vocabulary**

- oxyacid (p. 250)

Key Concepts

- Names of covalent molecular compounds include prefixes for the number of each atom present. The final letter of the prefix is dropped if the element name begins with a vowel.
- Molecules that produce H^+ in solution are acids. Binary acids contain hydrogen and one other element. Oxyacids contain hydrogen and an oxyanion.

Section 8.3 Molecular Structures**MAIN Idea** Structural formulas show the relative positions of atoms within a molecule.**Vocabulary**

- coordinate covalent bond (p. 259)
- resonance (p. 258)
- structural formula (p. 253)

Key Concepts

- Different models can be used to represent molecules.
- Resonance occurs when more than one valid Lewis structure exists for the same molecule.
- Exceptions to the octet rule occur in some molecules.

Section 8.4 Molecular Shapes**MAIN Idea** The VSEPR model is used to determine molecular shape.**Vocabulary**

- hybridization (p. 262)
- VSEPR model (p. 261)

Key Concepts

- VSEPR model theory states that electron pairs repel each other and determine both the shape of and bond angles in a molecule.
- Hybridization explains the observed shapes of molecules by the presence of equivalent hybrid orbitals.

Section 8.5 Electronegativity and Polarity**MAIN Idea** A chemical bond's character is related to each atom's attraction for the electrons in the bond.**Vocabulary**

- polar covalent bond (p. 266)

Key Concepts

- The electronegativity difference determines the character of a bond between atoms.
- Polar bonds occur when electrons are not shared equally forming a dipole.
- The spatial arrangement of polar bonds in a molecule determines the overall polarity of a molecule.
- Molecules attract each other by weak intermolecular forces. In a covalent network solid, each atom is covalently bonded to many other atoms.

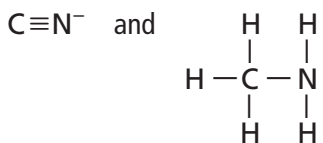
Section 8.1

Mastering Concepts

78. What is the octet rule, and how is it used in covalent bonding?
79. Describe the formation of a covalent bond.
80. Describe the bonding in molecules.
81. Describe the forces, both attractive and repulsive, that occur as two atoms move closer together.
82. How could you predict the presence of a sigma or pi bond in a molecule?

Mastering Problems

83. Give the number of valence electrons in N, As, Br, and Se. Predict the number of covalent bonds needed for each of these elements to satisfy the octet rule.
84. Locate the sigma and pi bonds in each of the molecules shown below.
 - a. $\begin{array}{c} \text{O} \\ || \\ \text{H}-\text{C}-\text{H} \end{array}$
 - b. $\text{H}-\text{C}\equiv\text{C}-\text{H}$
85. In the molecules CO, CO₂, and CH₂O, which C—O bond is the shortest? Which C—O bond is the strongest?
86. Consider the carbon-nitrogen bonds shown below:



Which bond is shorter? Which is stronger?

87. Rank each of the molecules below in order of the shortest to the longest sulfur-oxygen bond length.
 - a. SO₂
 - b. SO₃²⁻
 - c. SO₄²⁻

Section 8.2

Mastering Concepts

88. Explain how molecular compounds are named.
89. When is a molecular compound named as an acid?
90. Explain the difference between sulfur hexafluoride and disulfur tetrafluoride.
91. **Watches** The quartz crystals used in watches are made of silicon dioxide. Explain how you use the name to determine the formula for silicon dioxide.

Mastering Problems

92. Complete Table 8.8.

Table 8.8 Acid Names

Formula	Name
HClO ₂	
H ₃ PO ₄	
H ₂ Se	
HClO ₃	

93. Name each molecule.
 - a. NF₃
 - b. NO
 - c. SO₃
 - d. SiF₄
94. Name each molecule.
 - a. SeO₂
 - b. SeO₃
 - c. N₂F₄
 - d. S₄N₄
95. Write the formula for each molecule.
 - a. sulfur difluoride
 - b. silicon tetrachloride
 - c. carbon tetrafluoride
 - d. sulfurous acid
96. Write the formula for each molecule.
 - a. silicon dioxide
 - b. bromous acid
 - c. chlorine trifluoride
 - d. hydrobromic acid

Section 8.3

Mastering Concepts

97. What must you know in order to draw the Lewis structure for a molecule?
98. **Doping Agent** Material scientists are studying the properties of polymer plastics doped with AsF₅. Explain why the compound AsF₅ is an exception to the octet rule.
99. **Reducing Agent** Boron trihydride (BH₃) is used as reducing agent in organic chemistry. Explain why BH₃ often forms coordinate covalent bonds with other molecules.
100. Antimony and chlorine can form antimony trichloride or antimony pentachloride. Explain how these two elements can form two different compounds.

Mastering Problems

101. Draw three resonance structures for the polyatomic ion CO₃²⁻.
102. Draw the Lewis structures for these molecules, each of which has a central atom that does not obey the octet rule.
 - a. PCl₅
 - b. BF₃
 - c. ClF₅
 - d. BeH₂

- 103.** Draw two resonance structures for the polyatomic ion HCO_2^- .
- 104.** Draw the Lewis structure for a molecule of each of these compounds and ions.
- a. H_2S c. SO_2
 b. BF_4^- d. SeCl_2
- 105.** Which elements in the list below are capable of forming molecules in which one of its atoms has an expanded octet? Explain your answer.
- a. B d. O
 b. C e. Se
 c. P

Section 8.4

Mastering Concepts

- 106.** What is the basis of the VSEPR model?
- 107.** What is the maximum number of hybrid orbitals a carbon atom can form?
- 108.** What is the molecular shape of each molecule? Estimate the bond angle for each molecule, assuming that there is not a lone pair.
- a. A—B
- b. A—B—A
- c. $\begin{array}{c} \text{A}—\text{B}—\text{A} \\ | \\ \text{A} \end{array}$
- d. $\begin{array}{c} \text{A} \\ | \\ \text{A}—\text{B}—\text{A} \\ | \\ \text{A} \end{array}$
- 109. Parent Compound** PCl_5 is used as a parent compound to form many other compounds. Explain the theory of hybridization and determine the number of hybrid orbitals present in a molecule of PCl_5 .

Mastering Problems

- 110.** Complete Table 8.9 by identifying the expected hybrid on the central atom. You might find drawing the molecule's Lewis structure helpful.

Table 8.9 Structures

Formula	Hybrid Orbital	Lewis Structure
XeF_4		
TeF_4		
KrF_2		
OF_2		

- 111.** Predict the molecular shape of each molecule.
- a. COS b. CF_2Cl_2
- 112.** For each molecule listed below, predict its molecular shape and bond angle, and identify the hybrid orbitals. Drawing the Lewis structure might help you.
- a. SCl_2 c. HOF
 b. NH_2Cl d. BF_3

Section 8.5

Mastering Concepts

- 113.** Describe electronegativity trends in the periodic table.
- 114.** Explain the difference between nonpolar molecules and polar molecules.
- 115.** Compare the location of bonding electrons in a polar covalent bond with those in a nonpolar covalent bond. Explain your answer.
- 116.** What is the difference between a covalent molecular solid and a covalent network solid? Do their physical properties differ? Explain your answer.

Mastering Problems

- 117.** For each pair, indicate the more polar bond by circling the negative end of its dipole.
- a. C—S, C—O
 b. C—F, C—N
 c. P—H, P—Cl
- 118.** For each of the bonds listed, tell which atom is more negatively charged.
- a. C—H c. C—S
 b. C—N d. C—O
- 119.** Predict which bond is the most polar.
- a. C—O c. C—Cl
 b. Si—O d. C—Br
- 120.** Rank the bonds according to increasing polarity.
- a. C—H d. O—H
 b. N—H e. Cl—H
 c. Si—H
- 121. Refrigerant** The refrigerant known as freon-14 is an ozone-damaging compound with the formula CF_4 . Why is the CF_4 molecule nonpolar even though it contains polar bonds?
- 122.** Determine if these molecules and ion are polar. Explain your answers.
- a. H_3O^+ c. H_2S
 b. PCl_5 d. CF_4
- 123.** Use Lewis structures to predict the molecular polarities for sulfur difluoride, sulfur tetrafluoride, and sulfur hexafluoride.

Mixed Review

- 124.** Write the formula for each molecule.
- chlorine monoxide
 - arsenic acid
 - phosphorus pentachloride
 - hydrosulfuric acid
- 125.** Name each molecule.
- PCl_3
 - Cl_2O_7
 - P_4O_6
 - NO
- 126.** Draw the Lewis structure for each molecule or ion.
- SeF_2
 - ClO_2^-
 - PO_3^{3-}
 - POCl_3
 - GeF_4
- 127.** Determine which of the molecules are polar. Explain your answers.
- CH_3Cl
 - ClF
 - NCl_3
 - BF_3
 - CS_2
- 128.** Arrange the bonds in order of least to greatest polar character.
- C—O
 - Si—O
 - Ge—O
 - C—Cl
 - C—Br
- 129. Rocket Fuel** In the 1950s, the reaction of hydrazine with chlorine trifluoride (ClF_3) was used as a rocket fuel. Draw the Lewis structure for ClF_3 and identify the hybrid orbitals.
- 130.** Complete **Table 8.10**, which shows the number of electrons shared in a single covalent bond, a double covalent bond, and a triple covalent bond. Identify the group of atoms that will form each of these bonds.

Table 8.10 Shared Pairs

Bond Type	Number of Shared Electrons	Atoms that Form the Bond
Single covalent		
Double covalent		
Triple covalent		

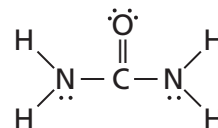
Think Critically

- 131. Organize** Design a concept map that explains how VSEPR model theory, hybridization theory, and molecular shape are related.
- 132. Compare and contrast** the two covalent compounds identified by the names arsenic(III) oxide and diarsenic trioxide.
- 133. Make and Use Tables** Complete **Table 8.11**, using what you learned in Chapters 7 and 8.

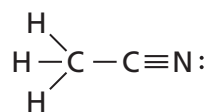
Table 8.11 Properties and Bonding

Solid	Bond Description	Characteristic of Solid	Example
Ionic			
Covalent molecular			
Metallic			
Covalent network			

- 134. Apply** Urea, whose structure is shown below, is a compound used in manufacturing plastics and fertilizers. Identify the sigma bond, pi bonds, and lone pairs present in a molecule of urea.



- 135. Analyze** For each of the characteristics listed below, identify the polarity of a molecule with that characteristic.
- solid at room temperature
 - gas at room temperature
 - attracted to an electric current
- 136. Apply** The structural formula for acetonitrile, CH_3CN , is shown below.



Examine the structure of the acetonitrile molecule. Determine the number of carbon atoms in the molecule, identify the hybrid present in each carbon atom, and explain your reasoning.

Challenge Problem

137. Examine the bond-dissociation energies for the various bonds listed in Table 8.12.

Bond	Bond-Dissociation Energy (kJ/mol)	Bond	Bond-Dissociation Energy (kJ/mol)
C—C	348	O—H	467
C=C	614	C—N	305
C≡C	839	O=O	498
N—N	163	C—H	416
N=N	418	C—O	358
N≡N	945	C=O	745

- Draw the correct Lewis structures for C_2H_2 and $HCOOH$.
- Determine the amount of energy needed to break apart each of these molecules.

Cumulative Review

138. Table 8.13 lists a liquid's mass and volume data. Create a line graph of this data with the volume on the x -axis and the mass on the y -axis. Calculate the slope of the line. What information does the slope give you? (Chapter 2)

Volume	Mass
4.1 mL	9.36 g
6.0 mL	14.04 g
8.0 mL	18.72 g
10.0 mL	23.40 g

139. Write the correct chemical formula for each compound. (Chapter 7)
- calcium carbonate
 - potassium chlorate
 - silver acetate
 - copper(II) sulfate
 - ammonium phosphate
140. Write the correct chemical name for each compound. (Chapter 7)
- NaI
 - $Fe(NO_3)_3$
 - $Sr(OH)_2$
 - $CoCl_2$
 - $Mg(BrO_3)_2$

Additional Assessment

WRITING in Chemistry

141. **Antifreeze** Research ethylene glycol, an antifreeze-coolant, to learn its chemical formula. Draw its Lewis structure and identify the sigma and pi bonds.
142. **Detergents** Choose a laundry detergent to research and write an essay about its chemical composition. Explain how it removes oil and grease from fabrics.

DBQ Document-Based Questions

Luminol Crime-scene investigators often use the covalent compound luminol to find blood evidence. The reaction between luminol, certain chemicals, and hemoglobin, a protein in blood, produces light. Figure 8.26 shows a ball-and-stick model of luminol.

Data obtained from: Fleming, Declan., 2002. The Chemiluminescence of Luminol, Exemplarchem, Royal Society of Chemistry.

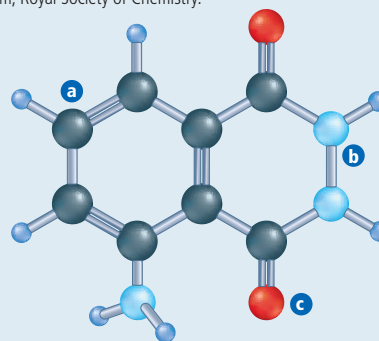
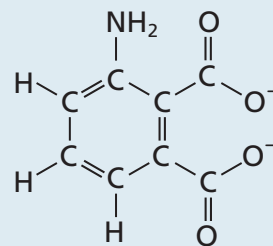


Figure 8.26

143. Determine the molecular formula for luminol and draw its Lewis structure.
144. Indicate the hybrid present on the atoms labeled A, B, and C in Figure 8.26.



APA ion

Figure 8.27

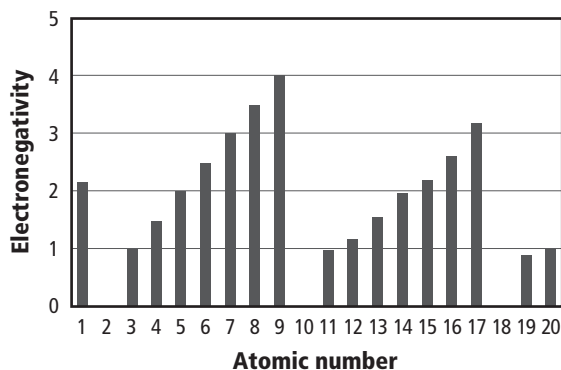
145. When luminol comes in contact with the iron ion in hemoglobin, it reacts to produce Na_2APA , water, nitrogen, and light energy. Given the structural formula of the APA ion in Figure 8.27, write the chemical formula for the polyatomic APA ion.

Cumulative Standardized Test Practice

Multiple Choice

- The common name of SiI_4 is tetraiodosilane. What is its molecular compound name?
 - silane tetraiodide
 - silane tetraiodine
 - silicon iodide
 - silicon tetraiodide
- Which compound contains at least one pi bond?
 - CO_2
 - CHCl_3
 - AsI_3
 - BeF_2

Use the graph below to answer Questions 3 and 4.



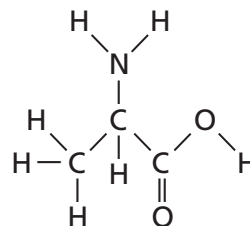
- What is the electronegativity of the element with atomic number 14?
 - 1.5
 - 1.9
 - 2.0
 - 2.2
- An ionic bond would form between which pairs of elements?
 - atomic number 3 and atomic number 4
 - atomic number 7 and atomic number 8
 - atomic number 4 and atomic number 18
 - atomic number 8 and atomic number 12
- Which is the Lewis structure for silicon disulfide?
 - :S::Si::S:
 - $\text{:}\ddot{\text{S}}\text{:}\text{:}\ddot{\text{S}}\text{:}\text{:}\ddot{\text{S}}\text{:}$
 - $\text{:}\ddot{\text{S}}\text{:}\text{:}\ddot{\text{S}}\text{:}$
 - $\text{:}\ddot{\text{S}}\text{:}\ddot{\text{S}}\text{:}\ddot{\text{S}}\text{:}$

- The central selenium atom in selenium hexafluoride forms an expanded octet. How many electron pairs surround the central Se atom?
 - 4
 - 5
 - 6
 - 7

Use the table below to answer Questions 7 and 8.

Bond Dissociation Energies at 298 K			
Bond	kJ/mol	Bond	kJ/mol
Cl-Cl	242	$\text{N}\equiv\text{N}$	945
C-C	345	O-H	467
C-H	416	C-O	358
C-N	305	C=O	745
H-I	299	O=O	498
H-N	391		

- Which diatomic gas has the shortest bond between its two atoms?
 - HI
 - O_2
 - Cl_2
 - N_2
- Approximately how much energy will it take to break all the bonds present in the molecule below?

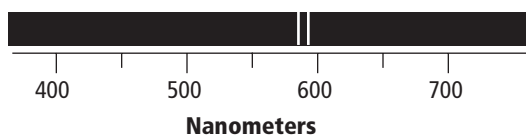


- 3024 kJ/mol
 - 4318 kJ/mol
 - 4621 kJ/mol
 - 5011 kJ/mol
- Which compound does NOT have a bent molecular shape?
 - BeH_2
 - H_2S
 - H_2O
 - SeH_2
 - Which compound is nonpolar?
 - H_2S
 - CCl_4
 - SiH_3Cl
 - AsH_3

Short Answer

11. Oxyacids contain hydrogen and an oxyanion. There are two different oxyacids that contain hydrogen, nitrogen, and oxygen. Identify these two oxyacids. How can they be distinguished on the basis of their names and formulas?

Use the atomic emission spectrum below to answer Questions 12 and 13.



12. Estimate the wavelength of the photons being emitted by this element.
13. Find the frequency of the photons being emitted by this element.

Extended Response

Use the table below to answer Question 14.

Percent Abundance of Silicon Isotopes		
Isotope	Mass	Percent Abundance
^{28}Si	27.98 amu	92.21 %
^{29}Si	28.98 amu	4.70 %
^{30}Si	29.97 amu	3.09 %

14. Your lab partner calculates the average atomic mass of these three silicon isotopes. His average atomic mass value is 28.98 amu. Explain why your lab partner is incorrect, and show how to calculate the correct average atomic mass.

SAT Subject Test: Chemistry

Use the list of separation techniques below to answer Questions 15 to 17.

- | | |
|--------------------|-------------------|
| A. filtration | D. chromatography |
| B. distillation | E. sublimation |
| C. crystallization | |
15. Which technique separates components of a mixture with different boiling points?
16. Which technique separates components of a mixture based on the size of its particles?
17. Which technique is based on the stronger attraction some components have for the stationary phase compared to the mobile phase?

Use the table below to answer Questions 18 to 19.

Electron-Dot Structures

Group	1	2	13	14	15	16	17	18
Diagram	Li·	·Be·	·B·	·C·	·N·	·O·	·F·	·Ne·

18. Based on the Lewis structures shown, which elements will combine in a 2:3 ratio?
- | | |
|---------------------------|------|
| A. lithium and carbon | D. 6 |
| B. beryllium and fluorine | E. 8 |
| C. beryllium and nitrogen | |
| D. boron and oxygen | |
| E. boron and carbon | |
19. How many electrons will beryllium have in its outer energy level after it forms an ion to become chemically stable?
- | | |
|------|------|
| A. 0 | D. 6 |
| B. 2 | E. 8 |
| C. 4 | |

NEED EXTRA HELP?

If You Missed Question . . .	1	2	3	4	5	6	7	8	9	10	11	12	13	14	15	16	17	18	19
Review Section . . .	8.2	8.1	8.5	8.5	8.3	8.3	8.1	8.1	8.4	8.5	8.2	5.2	5.2	4.3	3.3	3.3	3.3	7.2	7.2