CHAPTER 9

Covalent Bonding

What You'll Learn

- You will analyze the nature of a covalent bond.
- You will name covalently bonded groups of atoms.
- You will determine the shapes of molecules.
- You will describe characteristics of covalent molecules.
- You will compare and contrast polar and nonpolar molecules.

Why It's Important

Most compounds, including those in living organisms, are covalently bonded.

Chemistry Online

Visit the Chemistry Web site at **chemistrymc.com** to find links about covalent bonding.

Herbicides and fertilizers used on crops are covalent compounds.

DISCOVERY LAB

Materials

Beral-type pipette
vinegar
vegetable oil

Oil and Vinegar Dressing

When preparing a meal, you combine different types of food. But when you mix different substances, do they always "mix"? How about making oil and vinegar dressing for tonight's salad?

Safety Precautions

Procedure

1. Fill the bulb of a Beral-type pipet about 1/3 full of vinegar and 1/3 full of vegetable oil. Shake the pipet and its contents. Record your observations.
2. Allow the contents to sit for about five minutes. Record your observations.

Analysis

Do oil and vinegar mix? What explanation can you give for your observations in this experiment? Why do the instructions on many types of salad dressings read "shake well before using"?

Section 9.1 The Covalent Bond

Objectives

- **Apply** the octet rule to atoms that bond covalently.
- **Describe** the formation of single, double, and triple covalent bonds.
- **Compare** and **contrast** sigma and pi bonds.
- **Relate** the strength of covalent bonds to bond length and bond dissociation energy.

Vocabulary

covalent bond
molecule
Lewis structure
sigma bond
pi bond
endothermic
exothermic

Worldwide, scientists are studying ways to increase food supplies, reduce pollution, and prevent disease. Understanding the chemistry of compounds that make up fertilizers, pollutants, and materials that carry genetic information is essential in developing new technologies in these areas. An understanding of the chemistry of compounds requires an understanding of their bonding.

Why do atoms bond?

You learned in Chapter 6 that all noble gases have particularly stable electron arrangements. This stable arrangement consists of a full outer energy level. A full outer energy level consists of two valence electrons for helium and eight valence electrons for all other noble gases. Because of this stability, noble gases, in general, don't react with other elements to form compounds.

You also learned in Chapter 8 that when metals and nonmetals react to form binary ionic compounds, electrons are transferred, and the resulting ions have noble-gas electron configurations. But sometimes two atoms that both need to gain valence electrons to become stable have a similar attraction for electrons. Sharing of electrons is another way that these atoms can acquire the electron configuration of noble gases. Recall from Chapter 6 that the octet rule states that atoms lose, gain, or share electrons to achieve a stable configuration of eight valence electrons, or an octet. Although exceptions to the octet rule exist, the rule provides a useful framework for understanding chemical bonds.

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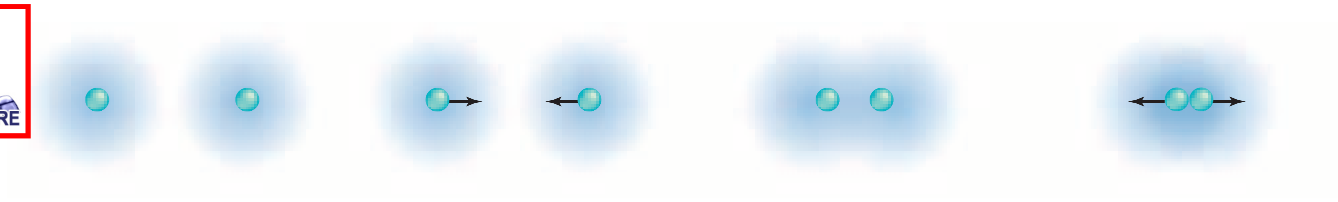


a The atoms are too far away from each other to have noticeable attraction or repulsion.

b Each nucleus attracts the other atom's electron cloud, but the electron clouds repel each other.

c The distance is right for the attraction of one atom's protons for the other atom's electrons to make the bond stable.

d If the atoms are forced closer together, the nuclei and electrons repel.

Figure 9-1

The overall force between two atoms is the result of electron-electron repulsion, nucleus-nucleus repulsion, and nucleus-electron attraction. The arrows in this diagram show the net force acting on two fluorine atoms as they move toward each other.

What is a covalent bond?

You know that certain atoms, such as magnesium and chlorine, transfer electrons from one atom to another, forming an ionic bond. However, the number of ionic compounds is quite small compared with the total number of known compounds. What type of bonding is found in all these other compounds that are not ionically bonded?

The atoms in these other compounds share electrons. The chemical bond that results from the sharing of valence electrons is a **covalent bond.** In a covalent bond, the shared electrons are considered to be part of the complete outer energy level of both atoms involved. Covalent bonding generally occurs when elements are relatively close to each other on the periodic table. The majority of covalent bonds form between nonmetallic elements.

A **molecule** is formed when two or more atoms bond covalently. The carbohydrates and simple sugars you eat; the proteins, fats, and DNA found in your body; and the wool, cotton, and synthetic fibers in the clothes you wear all consist of molecules formed from covalently bonded atoms.

Formation of a covalent bond Hydrogen (H_2), nitrogen (N_2), oxygen (O_2), fluorine (F_2), chlorine (Cl_2), bromine (Br_2), and iodine (I_2) occur in nature as diatomic molecules, not as single atoms because the molecules formed are more stable than the indiviual atoms. How do two atoms that do not give up electrons bond with each other?

Consider fluorine (F_2), which has an electron configuration of $1s^22s^22p^5$. Each fluorine atom has seven valence electrons and must have one additional electron to form an octet. As two fluorine atoms approach each other, as shown in **Figure 9-1**, two forces become important. A repulsive force occurs between the like-charged electrons and between the like-charged protons of the two atoms. An attractive force also occurs between the protons of one fluorine atom and the electrons of the other atom. As the fluorine atoms move closer, the attraction of both nuclei for the other atom's electrons increases until the maximum attraction is achieved. At the point of maximum attraction, the attractive forces balance the repulsive forces.

If the two nuclei move even closer, the repulsion between the like-charged nuclei and electron clouds will increase, resulting in repulsive forces that exceed attractive forces. Thus the most stable arrangement of atoms exists at the point of maximum attraction. At that point, the two atoms bond covalently and a molecule forms. Fluorine exists as a diatomic molecule because the sharing of one pair of electrons will give both fluorine atoms stable noble gas configurations. Each fluorine atom in the fluorine molecule has one bonding pair of electrons and three *lone pairs*, which are unshared pairs of electrons.

Single Covalent Bonds

Consider the formation of a hydrogen molecule, which is shown in **Figure 9-2**. Each covalently bonded atom equally attracts one pair of shared electrons. Thus, two electrons shared by two hydrogen nuclei belong to each atom simultaneously. Both hydrogen atoms have the noble gas configuration of helium ($1s^2$). The hydrogen molecule is more stable than individual hydrogen atoms.

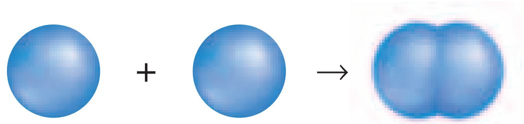


Figure 9-2

By sharing a pair of electrons, these hydrogen atoms have a full outer electron energy level and are stable. Refer to **Table C-1** in Appendix C for a key to atom color conventions.

When a single pair of electrons is shared, such as in a hydrogen molecule, a single covalent bond forms. The shared electron pair, often referred to as the bonding pair, is represented by either a pair of dots or a line in the Lewis structure for the molecule. **Lewis structures** use electron-dot diagrams to show how electrons are arranged in molecules. For example, a hydrogen molecule is represented as H : H or H—H. Hydrogen gas also is represented by the molecular formula H_2, which reflects the number of atoms in each molecule.

As you have seen, the halogens—group 7A elements—such as fluorine, have seven valence electrons. To attain an octet, one more electron is necessary. Therefore, group 7A elements will form a single covalent bond. You have seen how an atom from group 7A will form a covalent bond with another identical atom. Fluorine exists as F_2. Similarly, chlorine exists as Cl_2, bromine as Br_2, and iodine as I_2 because the molecule formed is more stable than the individual atoms. In addition, such bonds are often formed between the halogen and another element, such as carbon.

Group 6A elements share two electrons to form two covalent bonds. Oxygen is a group 6A element with an electron configuration of $1s^22s^22p^4$. Water is composed of two hydrogen atoms and one oxygen atom. Each hydrogen atom attains the noble gas configuration of helium as it shares one electron with oxygen. Oxygen, in turn, attains the noble gas configuration of neon by sharing one electron with each hydrogen atom. **Figure 9-3a** shows the Lewis structure for a molecule of water. Notice that two single covalent bonds are formed and two lone pairs remain on the oxygen atom.

Likewise, group 5A elements form three covalent bonds with atoms of nonmetals. Nitrogen is a group 5A element with the electron configuration of $1s^22s^22p^3$. Ammonia (NH_3) contains three single covalent bonds and one lone pair of electrons on the nitrogen atom. **Figure 9-3b** shows the Lewis structure

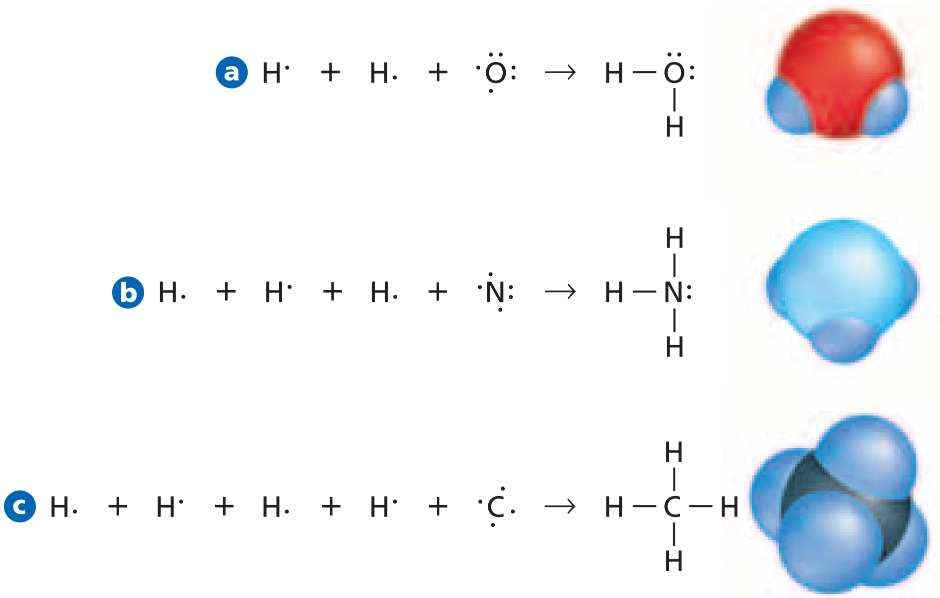
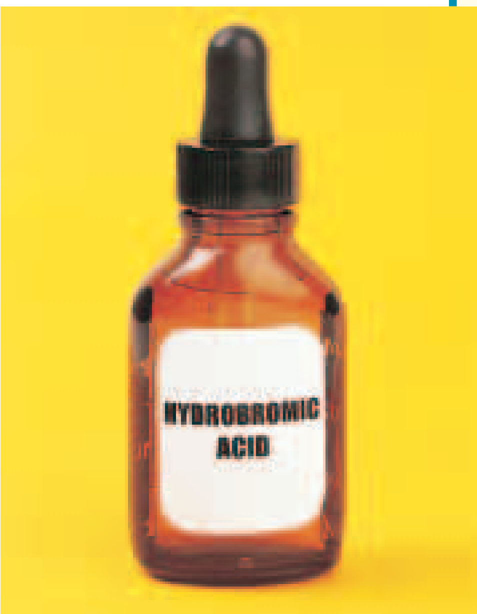


Figure 9-3

These chemical equations show how atoms share electrons to become stable. As can be seen by the Lewis structures (left side) for the molecules, after a reaction, all atoms in the molecules are stable according to the octet rule.

for an ammonia molecule. Nitrogen also forms similar compounds with group 7A elements, such as nitrogen trifluoride (NF_3), nitrogen trichloride (NCl_3), and nitrogen tribromide (NBr_3). Each of these group 7A atoms shares a pair of electrons with the nitrogen atom.

Group 4A elements will form four covalent bonds. A methane molecule (CH_4) is formed when one carbon atom bonds with four hydrogen atoms. Carbon, a group 4A element, has an electron configuration of $1s^2 2s^2 2p^2$. With four valence electrons, it must obtain four more electrons for a noble gas configuration. Therefore, when carbon bonds with other atoms, it forms four bonds. Because hydrogen, a group 1A element, has one valence electron, four hydrogen atoms are necessary to provide the four electrons needed by the carbon atom. The Lewis structure for methane is shown in **Figure 9-3c**. Carbon also forms single covalent bonds with other nonmetals, including group 7A elements.



When hydrogen bromide is dissolved in water, hydrobromic acid forms. This acid must be kept in a dark bottle because it decomposes when exposed to light.

EXAMPLE PROBLEM 9-1

Lewis Structure for a Molecule

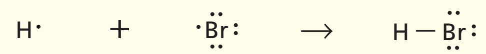
Hydrogen bromide (HBr) is used to manufacture several other bromides and has been medically used as a sedative. Draw the Lewis structure for this molecule.

1. Analyze the Problem

You are given that hydrogen and bromine form the molecule hydrogen bromide. Hydrogen, a group 1A element, has only one valence electron. Therefore, when hydrogen bonds with any nonmetal, it must share one pair of electrons. Bromine, a group 7A element, also needs one electron to complete its octet. Hydrogen and bromine bond with each other by one single covalent bond.

2. Solve for the Unknown

To draw the Lewis structure, first draw the electron-dot structure for each of the two atoms. Then show the sharing of the pairs of electrons by a single line.



Hydrogen atom | Bromine atom | Hydrogen bromide molecule

3. Evaluate the Answer

Each atom in the molecule has achieved a noble gas configuration and thus is stable.

For more practice with drawing Lewis structures, go to **Supplemental Practice Problems** in Appendix A.

PRACTICE PROBLEMS

Draw the Lewis structure for each of these molecules.

1. PH_3

2. H_2S

3. HCl

4. CCl_4

5. SiH_4

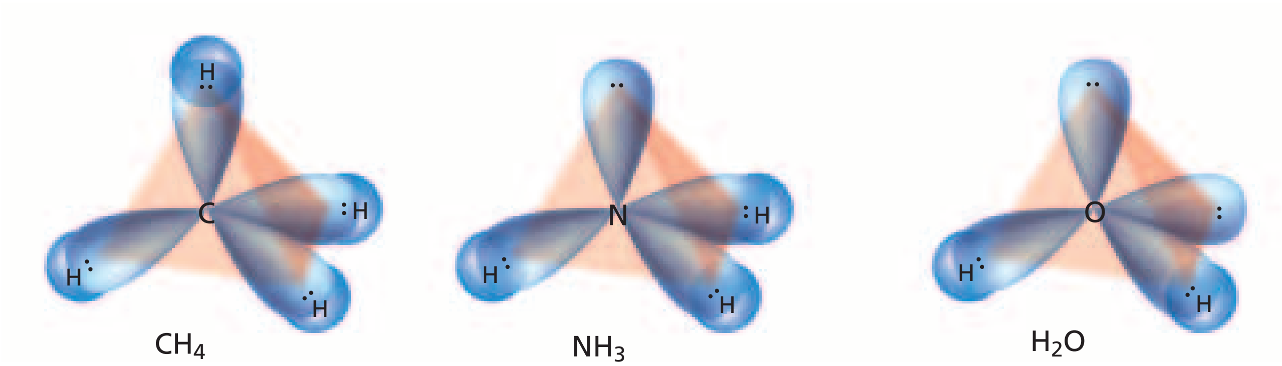


Figure 9-4

The sigma bonds form in these molecules as the atomic orbitals of hydrogen atoms overlap end to end with the orbitals of the central atoms.

The sigma bond Single covalent bonds also are called **sigma bonds,** symbolized by the Greek letter sigma (σ). A sigma bond occurs when the electron pair is shared in an area centered between the two atoms. When two atoms share electrons, the valence atomic orbital of one atom overlaps or merges with the valence atomic orbital of the other atom. A sigma bond results if the 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. **Figure 9-4** indicates the sigma bonds found in methane (CH_4), ammonia (NH_3), and water (H_2O). Sigma bonds can form from the overlap of an s orbital with another s orbital, an s orbital with a p orbital, or a p orbital with another p orbital. Does hydrogen use s or p orbitals to form bonds in **Figure 9-4**?

Multiple Covalent Bonds

In many molecules, atoms attain a noble-gas configuration by sharing more than one pair of electrons between two atoms, forming a multiple covalent bond. Atoms of the elements carbon, nitrogen, oxygen, and sulfur most often form multiple bonds. How do you know when two atoms will form a multiple bond? The number of valence electrons of an element is associated with the number of shared electron pairs needed to complete the octet and gives a clue as to how many covalent bonds can form.

Double and triple covalent bonds are examples of multiple bonds. A double covalent bond occurs when two pairs of electrons are shared. The atoms in an oxygen molecule (O_2) share two electron pairs, forming a double bond. Each oxygen atom has six valence electrons and must obtain two additional electrons for a noble-gas configuration. If each oxygen atom shares two electrons, a total of two pairs of electrons is shared between the two atoms. A double covalent bond results. See **Figure 9-5a**.

A triple covalent bond is formed when three pairs of electrons are shared between two atoms. Nitrogen (N_2) shares three electron pairs, producing a triple bond. One nitrogen atom needs three additional electrons to attain a noble-gas configuration. **Figure 9-5b** shows the triple bond formed between two nitrogen atoms.

a $:\ddot{O}\cdot + \cdot\ddot{O}: \rightarrow \ddot{O}{=}\ddot{O}$

b $:\dot{N}\cdot + \cdot\dot{N}: \rightarrow :N{\equiv}N:$

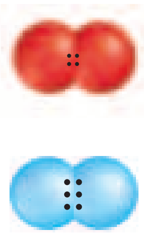
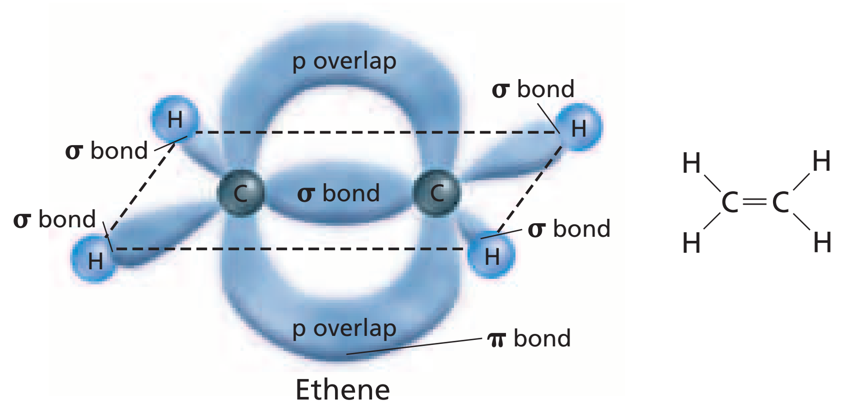


Figure 9-5

Multiple bonds form when more than one pair of electrons are shared. **a** The two pairs of electrons shared by oxygen atoms form a double bond. **b** Nitrogen shares three pairs of electrons, forming a triple bond.

Figure 9-6

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



The pi bond A **pi bond,** denoted by the Greek symbol pi (π), is formed when parallel orbitals overlap to share electrons, as shown in **Figure 9-6**. 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. A multiple bond consists of one sigma bond and at least one pi bond. A double covalent bond has one sigma bond and one pi bond. A triple covalent bond consists of one sigma bond and two pi bonds. See **Figure 9-6**. A pi bond always accompanies a sigma bond when forming double and triple bonds.

Strength of Covalent Bonds

Remember 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, the covalent bond can be broken. Some covalent bonds are broken more easily than others because they differ in strength. Several factors control the strength of covalent bonds.

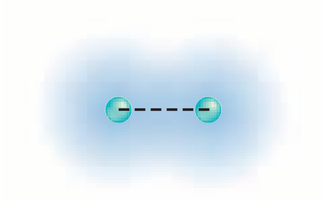


Figure 9-7

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

The strength of a covalent bond depends on how much distance separates bonded nuclei. The distance between the two bonding nuclei at the position of maximum attraction is called *bond length*, which is determined by the size of the atoms and how many electron pairs are shared. See **Figure 9-7.** The bond length of the single bond in an F_2 molecule is 1.43×10^{-10} m. The bond length of the double bond in O_2 is 1.21×10^{-10} m. The bond length of the triple bond in N_2 is 1.10×10^{-10} m. Although the sizes of the atoms are not the same, not much difference exists in the size of these molecules. How does the number of pairs of electrons shared in F_2, O_2, and N_2 relate to the bond lengths in each of these molecules? As the number of shared electron pairs increases, bond length decreases. A triple bond, sharing three electron pairs, has a shorter bond length than a single bond where only two electrons are shared. The shorter the bond length, the stronger the bond. Thus, single bonds, such as those in F_2, are weaker than double bonds, such as those in O_2. Double bonds are weaker than triple bonds, such as those in N_2.

An energy change accompanies the forming or breaking of a bond between atoms in a molecule. Energy is released when a bond forms. Energy must be added to break the bonds in a molecule. The amount of energy required to break a specific covalent bond is called *bond dissociation energy*. Breaking bonds always requires the addition of energy. Thus, bond dissociation energy is always a positive value. The bond dissociation energy of F_2 is 159 kJ/mol, of O_2 is 498 kJ/mol, and of N_2 is 945 kJ/mol. The sum of the bond dissociation energy values for all bonds in a compound is used to determine the amount of chemical potential energy available in a molecule of that compound.

Bond dissociation energy indicates the strength of a chemical bond because a direct relationship exists between bond energy and bond length. As two atoms are bonded closer together, greater amounts of bond energy are needed to separate them. Think back to the relative bond lengths of F_2, O_2, and N_2. Based on bond length, which of these three molecules would you predict to have the greatest bond energy? The least? Do the bond dissociation energies supplied in the previous paragraph confirm your predictions?

Figure 9-8

Energy is used to break C–C bonds in coal and O–O bonds of oxygen in air. Energy is released as heat and light due to the formation of CO_2. Coal burning is an exothermic reaction.

In chemical reactions, bonds in reactant molecules are broken and new bonds are formed as product molecules form. See **Figure 9-8.** The total energy change of the chemical reaction is determined from the energy of the bonds broken and formed. **Endothermic** reactions occur 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 product molecules. **Exothermic** reactions occur when more energy is released forming new bonds than is required to break bonds in the initial reactants. You will learn more about energy and chemical processes in Chapter 16.

Try at Home LAB

See page 956 in Appendix E for **Breaking Covalent Bonds**

Section 9.1 Assessment

6. What is a covalent bond? How does it differ from an ionic bond?

7. What is a single covalent bond? Why does it form?

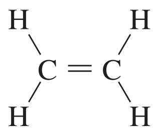
8. Why do multiple bonds form?

9. What is the difference between a sigma bond and a pi bond?

10. How is bond length related to bond dissociation energy?

11. **Thinking Critically** From the following structures, predict the relative bond energies needed to break all of the bonds present.

a. H—C ≡ C—H **b.**



12. **Making Predictions** Draw the electron-dot diagrams for the elements sulfur, carbon, bromine, oxygen, and hydrogen. Using Lewis structures, predict the number of covalent bonds formed when

a. one atom of sulfur bonds with two atoms of hydrogen.

b. one atom of carbon bonds with two atoms of sulfur.

c. two atoms of bromine bond with one atom of sulfur.

d. one atom of carbon bonds with four atoms of bromine

e. one atom of sulfur bonds with two atoms of oxygen.

Section 9.2

Naming Molecules

Español

Objectives

- **Identify** the names of binary molecular compounds from their formulas.
- **Name** acidic solutions.

Vocabulary

oxyacid

You know that many atoms covalently bond to form molecules that behave as a single unit. These units can be represented by chemical formulas and names that are used to identify them. When naming molecules, the system of rules is similar to the one you used to name ionic compounds.

Naming Binary Molecular Compounds

The anesthetic dinitrogen oxide (N_2O), commonly known as nitrous oxide, is a covalently bonded compound. Because it contains only two different elements, it is a binary molecular compound. Binary molecular compounds are composed of two different nonmetals and do not contain metals or ions. Although many of these compounds have common names, they also have scientific names that reveal their composition. Use the following simple rules to name binary molecular compounds.

1. *The first element in the formula is always named first, using the entire element name.*
2. *The second element in the formula is named using the root of the element and adding the suffix -ide.*
3. *Prefixes are used to indicate the number of atoms of each type that are present in the compound.* The most common prefixes are shown in **Table 9-1**.

Table 9-1

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-

One exception to using these prefixes is that the first element in the formula never uses the prefix *mono-*. Also, to avoid awkward pronunciation, drop the final letter in the prefix when the element name begins with a vowel. For example, CO is carbon monoxide, not monocarbon monooxide.

EXAMPLE PROBLEM 9-2

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. This formula reveals what elements are present and how many atoms of each element exist in a molecule. 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

Name the first element present in the compound.

phosphorus

The second element is oxygen. The root of this name is *ox-*, so the second part of the name is *oxide*.

phosphorus oxide

From the formula P_2O_5, two phosphorus atoms and five oxygen atoms make up a molecule of the compound. From **Table 9-1**, *di-* is the prefix for two, and *penta-* is the prefix for five.

diphosphorus pentoxide.

The *a-* of *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 chemical formula for the compound, P_2O_5.

Diphosphorus pentoxide releases energy and produces fumes when it dissolves in water to form phosphoric acid.

PRACTICE PROBLEMS

Name the following binary covalent compounds.

13. CCl_4

14. As_2O_3

15. CO

16. SO_2

17. NF_3

Practice!

For more practice with naming binary covalent compounds, go to **Supplemental Practice Problems** in Appendix A.

Common names of some molecular compounds How frequently have you drunk an icy, cold glass of dihydrogen monoxide? Quite frequently, but you probably didn't call it that. You called it by its more common name, which is water. Remember from Chapter 8 that many ionic compounds have common names in addition to their scientific ones. Baking soda is sodium hydrogen carbonate and common table salt is sodium chloride. Many covalent compounds also have both common and scientific names.

Many binary molecular compounds were discovered and given common names long before the modern naming system was developed. **Table 9-2** lists some of these molecules, their common names, and the binary molecular compound names.

Table 9-2

Formulas and Names of Some Covalent Compounds		
Formula	**Common name**	**Molecular compound name**
H_2O	water	dihydrogen monoxide
NH_3	ammonia	nitrogen trihydride
N_2H_4	hydrazine	dinitrogen tetrahydride
N_2O	nitrous oxide (laughing gas)	dinitrogen monoxide
NO	nitric oxide	nitrogen monoxide

Careers Using Chemistry

Organic Chemist

Would you like to create molecular models, using state-of-the-art computers? That is one role of organic chemists, who study molecules based on carbon.

Organic chemists and chemists in related fields use computers to analyze the structure and motion of molecules. Sometimes they build a visual model of one molecule. Other times they examine entire molecular systems. This information is essential in finding new drugs and in mapping human genes. Molecular modeling is one of the fastest-growing fields in science.

Naming Acids

Water solutions of some molecules are acidic and are named as acids. Acids are important compounds with specific properties that will be discussed at length in Chapter 19. If the 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. When naming a binary acid, use the prefix *hydro-* to name the hydrogen part of the compound. The rest of the name consists of a form of the root of the second element plus the suffix *-ic*, followed by the word *acid*. For example, HBr in a water solution is called hydrobromic 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.

Naming oxyacids Another set of rules is used to name an acid that contains an oxyanion. An oxyanion is a polyatomic ion that contains oxygen. Any acid that contains hydrogen and an oxyanion is referred to as an **oxyacid.**

Because the name of an oxyacid depends on the oxyanion present in the acid, you must first identify the anion present. The name of an oxyacid consists of a form of the root of the anion, a suffix, and the word *acid*. If the anion suffix is *-ate*, it is replaced with the suffix *-ic*. When the anion ends in *-ite*, the suffix is replaced with *-ous*. Consider the oxyacid HNO_3. Its oxyanion is nitrate (NO_3^-). Following this rule, HNO_3 is named nitric acid. The anion of HNO_2 is the nitrite ion (NO_2^-). HNO_2 is nitrous acid. Notice that the hydrogen in an oxyacid is not part of the name.

It's important to remember that these hydrogen-containing compounds are named as acids only when they are in water solution. For example, at room temperature and pressure HCl is hydrogen chloride, a gas. When HCl is dissolved in water, it is hydrochloric acid.

Practice!

For more practice with naming acids, go to **Supplemental Practice Problems** in Appendix A.

PRACTICE PROBLEMS

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

18. HI

19. $HClO_3$

20. $HClO_2$

21. H_2SO_4

22. H_2S

Writing Formulas from Names

The name of a molecular compound reveals its composition and is important in communicating the nature of the compound. **Figure 9-9** can help you determine the name of a molecular covalent compound.

The name of any binary molecule allows you to write the correct formula with ease. Subscripts are determined from the prefixes used in the name because the name indicates the exact number of each atom present in the molecule. The formula for an acid can be derived from the name as well.

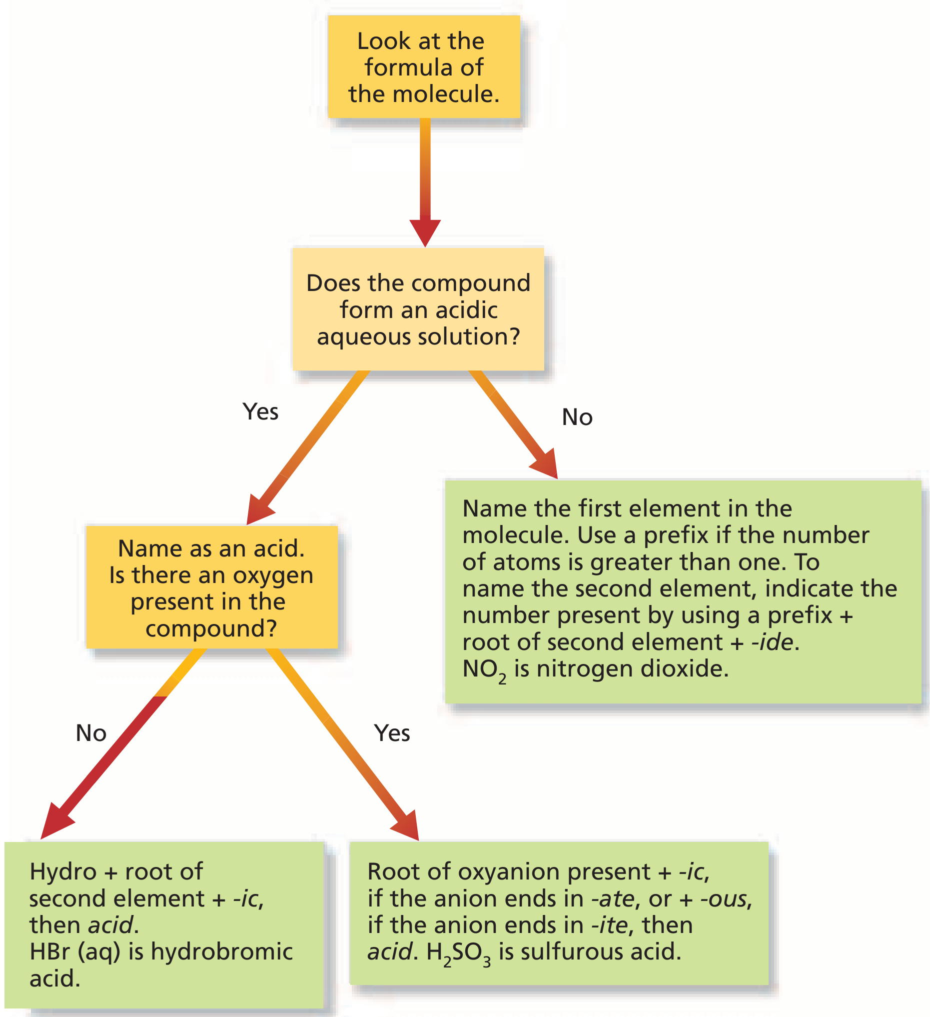


Figure 9-9

This flow chart summarizes how to name molecular compounds when their formulas are known.

Section 9.2 Assessment

23. What is a binary molecular compound?

24. Using the system of rules for naming binary molecular compounds, describe how you would name the molecule N_2O_4.

25. Compare and contrast naming binary acids and naming other binary covalent molecules.

26. What is the difference between a binary acid and an oxyacid?

27. Write the molecular formula for each of the following compounds.
- **a.** disulfur trioxide
- **b.** iodic acid
- **c.** dinitrogen monoxide
- **d.** hydrofluoric acid
- **e.** phosphorus pentachloride

28. **Thinking Critically** Write the molecular formula for each listed compound.
- **a.** dinitrogen trioxide and nitrogen monoxide
- **b.** hydrochloric acid and chloric acid
- **c.** sulfuric acid and sulfurous acid

29. **Making and Using Tables** Complete the following table.

Formulas and Names of Covalent Compounds

Formula	Name
PCl_5	
	hydrobromic acid
H_3PO_4	
	oxygen difluoride
SO_2	

Section 9.3

Molecular Structures

Objectives

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

Vocabulary

structural formula
resonance
coordinate covalent bond

You can now identify atoms that bond covalently and name the molecular compounds formed through covalent bonding. In order to predict the arrangement of atoms in each molecule, a model, or representation is used. Several different models can be used, as shown in **Figure 9-10**. 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 C-1** in Appendix C. These colors are used for identification of the atoms if the chemical symbol of the element is not present.

Structural Formulas

One of the most useful molecular models is the **structural formula,** which uses letter symbols and bonds to show relative positions of atoms. The structural formula can be predicted 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.

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. The following steps should be used to determine Lewis structures.

1. Predict the location of certain atoms.
   - **a.** 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.
   - **b.** The atom with the least attraction for shared electrons in the molecule is the central atom. This element usually is the one closer to the left on the periodic table. The central atom is located in the center of the molecule, and all other atoms become terminal atoms.
2. Find the total number of electrons available for bonding. This total is the number of valence electrons in the atoms in the molecule.
3. Determine the number of bonding pairs by dividing the number of electrons available for bonding by two.
4. Place one bonding pair (single bond) between the central atom and each of the terminal atoms.

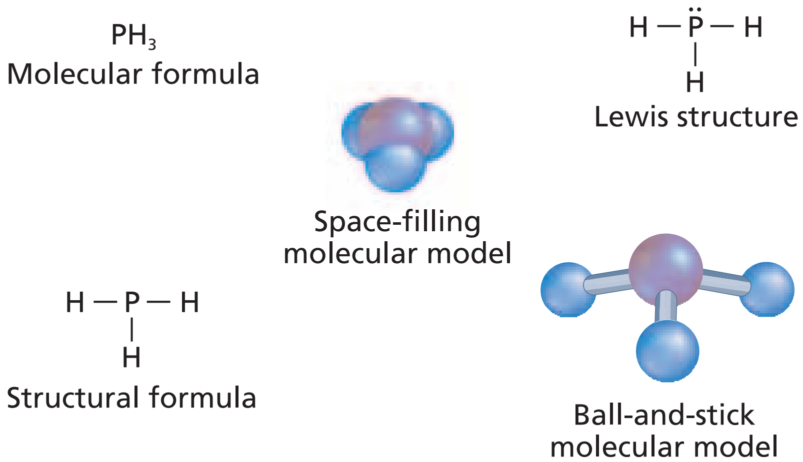


Figure 9-10
All of these models can be used to show the relative locations of atoms and electrons in the phosphorus trihydride (phosphine) molecule.

5. Subtract the number of pairs you used in step 4 from the number of bonding pairs you determined in step 3. The remaining electron pairs include lone pairs as well as pairs used in double and triple bonds. Place lone pairs around each terminal atom bonded to the central atom to satisfy the octet rule. Any remaining pairs are assigned to the central atom.

6. If the central atom is not surrounded by four electron pairs, it does not have an octet. You must convert one or two of the lone pairs on the terminal atoms to 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, in general, carbon, nitrogen, oxygen, and sulfur can form double or triple bonds with the same element or with another element.

EXAMPLE PROBLEM 9-3

Lewis Structure: Covalent Compound with Single Bonds

Ammonia is a raw material for the manufacture of many materials, including fertilizers, cleaning products, and explosives. Draw the Lewis structure for ammonia (NH_3).

Ammonia is a common ingredient in many cleaning products.

1. Analyze the Problem

You are given that the ammonia molecule consists of one nitrogen atom and three hydrogen atoms. Because hydrogen must be a terminal atom, nitrogen is the central atom.

2. Solve for the Unknown

Find the total number of valence electrons.

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

$= 8$ valence electrons

Determine the total number of bonding pairs.

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

Draw single bonds from each H to the N.

$$\begin{array}{c} \text{H}-\text{N}-\text{H} \\ | \\ \text{H} \end{array}$$

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

4 pairs total $-$ 3 pairs used $=$ 1 pair available

One lone pair remains to 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 atom, N.

$$\begin{array}{c} \text{H}-\ddot{\text{N}}-\text{H} \\ | \\ \text{H} \end{array}$$

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.

EXAMPLE PROBLEM 9-4

Lewis Structure: 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

You are given that 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.

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

$= 16$ valence electrons

Determine the total number of bonding pairs.

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

Draw single bonds from each O to the central C atom.

O—C—O

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

8 pairs total − 2 pairs used = 6 pairs available

Add three lone pairs to each terminal oxygen. Subtract the lone pairs used from the pairs available.

6 pairs available − 6 lone pairs used = 0

No electron pairs remain available for the carbon atom.

:Ö—C—Ö:

Carbon does not have an octet, so use a lone pair from each oxygen atom to form a double bond with the carbon atom.

Ö=C=Ö

3. Evaluate the Answer

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

As a waste product, carbon dioxide is one of the gases exhaled.

The polyatomic ions you learned about in Chapter 8 are related to covalent compounds. Although the unit acts as an ion, the atoms within the ion itself are covalently bonded. The structures of these ions can also be represented by Lewis structures.

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 9-5

Lewis Structure: 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. Phosphorus has less attraction for shared electrons, so it is the central atom, and the four oxygen atoms are terminal.

2. Solve for the Unknown

Find the total number of valence electrons.

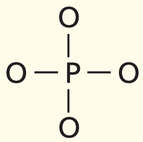
$$1\ \cancel{\text{P atom}} \times \frac{5\text{ valence electrons}}{\cancel{\text{P atom}}} + 4\ \cancel{\text{O atoms}} \times \frac{6\text{ valence electrons}}{\cancel{\text{O atom}}}$$

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

Determine the total number of bonding pairs.

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

Draw single bonds from each O to the central P.



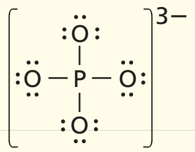
Subtract the number of pairs used in these bonds from the total number of pairs of electrons available.

16 pairs total − 4 pairs used = 12 pairs available

Add three lone pairs to each terminal oxygen. Subtract the lone pairs used from the pairs available.

12 pairs available − 12 lone pairs used = 0

No electron pairs remain available for the phosphorus atom, resulting in the following Lewis structure for the phosphate ion.



3. Evaluate the Answer

Phosphorus and all four oxygen atoms have an octet, and the group of atoms has a net charge of −3.

One phosphate compound, $MgH_4(PO_4)_2$, is used to fireproof wood. This compound contains the polyatomic ion phosphate.

PRACTICE PROBLEMS

Draw a Lewis structure for each of the following:

30. NF_3

31. CS_2

32. BH_3

33. ClO_4^-

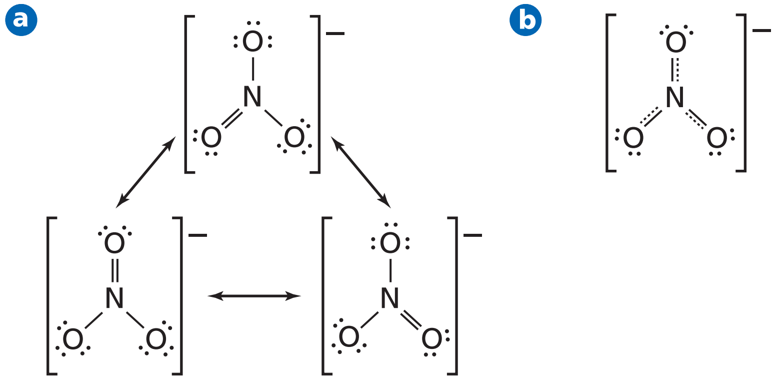
34. NH_4^+

Practice!

For more practice with drawing Lewis structures, go to Supplemental Practice Problems in Appendix A.

Figure 9-11

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 line indicates possible locations of the double bond.



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 9-11a**. 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 often 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 actual molecule or ion that undergoes resonance behaves as if it has only one structure. See **Figure 9-11b.** 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.

For more practice with drawing Lewis resonance structures, go to **Supplemental Practice Problems** in Appendix A.

PRACTICE PROBLEMS

Draw the Lewis resonance structures for the following.

35. SO_3

36. SO_2

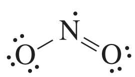
37. O_3

38. NO_2^-

Exceptions to the Octet Rule

The Lewis structure is focused on the attainment of an octet by all atoms when they bond with other elements. Some molecules and ions, however, do not obey the octet rule. Three reasons exist for these exceptions.

First, a small group of molecules has an odd number of valence electrons and cannot 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. ClO_2 and NO are other examples of molecules with odd numbers of valence electrons.



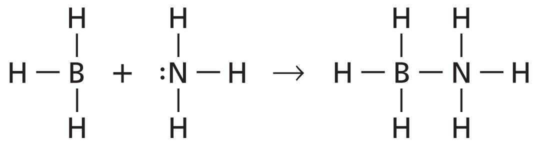
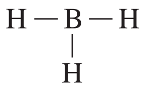


Figure 9-12

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.

Second, some compounds form with fewer than eight electrons present around an atom. This group is relatively rare, and BH_3 is an example. Boron, a group 3A nonmetal, forms three covalent bonds with other nonmetallic atoms.



A total of six electrons is shared by the boron atom, which is two less than the number needed for an octet. Such compounds tend to be reactive and can share an entire pair of electrons donated by another atom. When one atom donates a pair of electrons to be shared with an atom or ion that needs two electrons to become stable, a **coordinate covalent bond** forms, as in **Figure 9-12**. Atoms or ions with lone pairs often form coordinate covalent bonds with atoms or ions that need two more electrons.

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 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.

EXAMPLE PROBLEM 9-6

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

Find the total number of valence electrons.

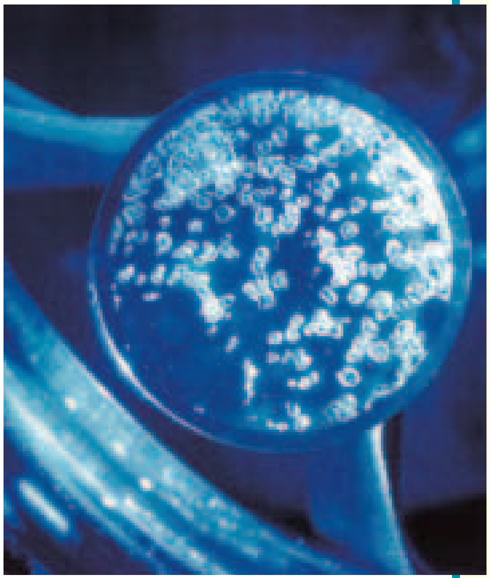
$$1\ \cancel{\text{Xe atom}} \times \frac{8\text{ valence electrons}}{\cancel{\text{Xe atom}}} + 4\ \cancel{\text{F atoms}} \times \frac{7\text{ valence electrons}}{\cancel{\text{F atom}}}$$

$$= 36\text{ valence electrons}$$

Determine the number of bonding pairs.

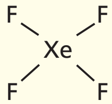
$$\frac{36\ \cancel{\text{electrons}}}{2\ \cancel{\text{electrons}}\text{/pair}} = 18\text{ pairs}$$

Continued on next page



Xenon compounds, such as the XeF_4 shown here, are highly toxic because they are so reactive.

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



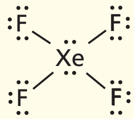
Determine the number of remaining pairs.

18 pairs available − 4 pairs used = 14 pairs available

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

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

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.

PRACTICE PROBLEMS

Practice!

For more practice with drawing Lewis structures, go to **Supplemental Practice Problems** in Appendix A.

Draw the correct Lewis structures for the following molecules, which contain expanded octets.

39. SF_6

40. PCl_5

41. ClF_3

You now know how to draw Lewis structures for molecules and polyatomic ions. You can use them to determine the number of bonding pairs between atoms and the number of lone pairs present. Next, you will learn to describe molecular structure and predict the angles in a molecule, both of which determine the three-dimensional molecular shape.

Section 9.3 Assessment

42. What is the role of the central atom when drawing the Lewis structure for a molecule?

43. What is resonance?

44. List three exceptions to the octet rule.

45. What is a coordinate covalent bond?

46. What is an expanded octet?

47. **Thinking Critically** Draw the resonance structures for the N_2O molecule.

48. **Formulating Models** Draw the Lewis structures for the following molecules and ions.

a. CN^- **c.** HCO_3^-

b. SiF_4 **d.** AsF_6^-

Section 9.4 Molecular Shape

Español

The shape of a molecule determines many of its physical and chemical properties. Molecular shape, in turn, is determined by the overlap of orbitals that share electrons. Theories have been developed to explain the overlap of bonding orbitals and are used to predict the shape of the molecule.

VSEPR Model

Many chemical reactions, especially those in living things, depend on the ability of two compounds to contact each other. The shape of the molecule determines whether or not molecules can get close enough to react.

Once a Lewis structure is drawn, you can determine the molecular geometry, or shape, of the molecule. The model used to determine the molecular shape is referred to as the **V**alence **S**hell **E**lectron **P**air **R**epulsion model, or **VSEPR model.** This model is based on an arrangement that minimizes the repulsion of shared and unshared pairs of electrons around the central atom.

The repulsions among electron pairs in a molecule result in atoms existing at fixed angles to each other. The angle formed by any two terminal atoms and the central atom is a *bond angle*. Bond angles predicted by VSEPR are supported by experimental evidence. Shared electron pairs repel one another. Lone pairs of electrons also are important in determining the shape of the molecule. Because lone pairs are not shared between two nuclei, they occupy a slightly larger orbital than shared electrons. Shared bonding orbitals are pushed together slightly by lone pairs.

To make sense of the VSEPR model, consider balloons of similar size tied together, as shown in **Figure 9-13**. Each balloon represents an orbital and, therefore, the repulsive force that keeps other electrons from entering this space. When each set of balloons is connected at a central point that represents the central atom, the balloons naturally form a shape that minimizes interactions between the balloons. You will build additional examples of VSEPR models in the **miniLAB** on page 261.

Examine **Table 9-3**, which indicates some common shapes of molecules. First, consider molecules that contain no lone pairs of electrons. When only two pairs of electrons are shared off the central atom ($BeCl_2$), these bonding electrons will seek the maximum separation, which is a bond angle of 180°, and the molecular shape is linear. Three bonding electron pairs also will separate as much as possible ($AlCl_3$), forming a trigonal planar shape with 120° bond angles. When the central atom has four pairs of bonding electrons (CH_4), the shape is tetrahedral with bond angles of 109.5°.

It is important to remember that a lone pair also occupies a position in space. Recall phosphine (PH_3) shown in **Figure 9-10,** which has three single covalent bonds and one lone pair. You might expect the structure to be tetrahedral because of the four bonding positions around the central atom. However, the lone pair takes up a greater amount of space than the shared pairs. The geometry of PH_3 is trigonal pyramidal, and the bond angles are 107.3°.

Now consider a water molecule (H_2O), which has two single covalent bonds and two lone pairs according to its Lewis structure. Although a water molecule has four electron pairs off the central atom, it is not tetrahedral because the two lone pairs occupy more space than do the paired electrons. The water molecule has a bent shape with a bond angle of 104.5°.

Objectives

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

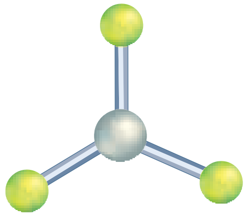
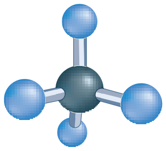
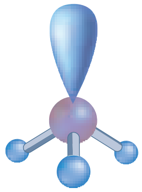
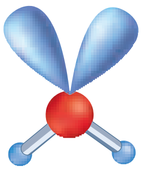
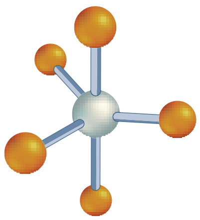
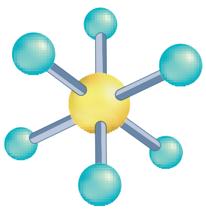
Vocabulary

VSEPR model
hybridization

Figure 9-13

Electrons in a molecule are located as far apart as they can be, just as these balloons are arranged.

Table 9-3

Molecular Shapes						
Example	**Total pairs**	**Shared pairs**	**Lone pairs**	***Molecular shape**	**Bond Angle**	**Hybrid Orbitals**
$BeCl_2$	2	2	0	 **Linear**	180°	sp
$AlCl_3$	3	3	0	 **Trigonal planar**	120°	sp^2
CH_4	4	4	0	 **Tetrahedral**	109.5°	sp^3
PH_3	4	3	1	 **Trigonal pyramidal**	107.3°	sp^3
H_2O	4	2	2	 **Bent**	104.5°	sp^3
$NbBr_5$	5	5	0	 **Trigonal bipyramidal**	(vertical to horizontal) 90° /120° (horizontal to horizontal)	sp^3d
SF_6	6	6	0	 **Octahedral**	90°	sp^3d^2

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

CONTENTS

Hybridization

A hybrid results from combining two of the same type of object, and it has characteristics of both. Atomic orbitals undergo hybridization during bonding. Let's consider the bonding involved in the methane molecule (CH_4). The carbon atom has four valence electrons with the electron configuration of $[He]2s^22p^2$. You might expect the two unpaired p electrons to bond with other atoms and the 2s electrons to remain a lone pair. However, carbon atoms undergo **hybridization,** a process in which atomic orbitals are mixed to form new, identical *hybrid orbitals.* Each hybrid orbital contains one electron that it can share with another atom, as shown in **Figure 9-14.** Carbon is the most common element that undergoes hybridization. Because the four hybrid orbitals form from one s and three p orbitals, this hydrid orbital is called an sp^3 orbital.

Look again at **Table 9-3** on the previous page. Notice that the number of atomic orbitals mixed to form the hybrid orbital equals the total number of pairs of electrons. 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. To have this shape, one s and two p orbitals on the central atom Al must mix to 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 9-3.** Both compounds contain three atoms. Why is H_2O sp^3? There are two lone pairs on the central atom (oxygen) in H_2O. Therefore, there must be four hybrid orbitals—two for bonding and two for the lone pairs.

Recall from Section 9.1 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.

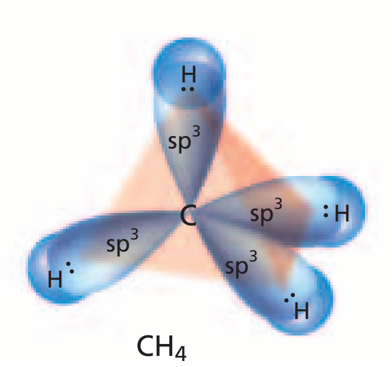


Figure 9-14

One s and three p orbitals hybridize to form four sp^3 orbitals. According to VSEPR, a tetrahedral shape minimizes repulsion between the orbitals.

miniLAB

Building VSEPR Models

Formulating Models The VSEPR model states that pairs of valence electrons on a central atom repel each other and are arranged so that the repulsions are as small as possible. In this **miniLAB**, you will use marshmallows and gumdrops to build models of substances, showing examples of the VSEPR model.

Materials regular-sized marshmallows (3); mini-sized marshmallows (9); small gumdrops (3); toothpicks, cut in half

Procedure

1. Draw Lewis structures for methane (CH_4), ammonia (NH_3), and water (H_2O). Notice the location of each shared and unshared pair of electrons.
2. Using your Lewis structures, build a VSEPR model for each molecule. Use a mini-marshmallow to represent both the hydrogen atom and the region of space containing the pair of electrons shared by hydrogen and the central atom. Use a regular-sized marshmallow to represent the space occupied by an unshared pair of electrons and a small gumdrop to represent a central atom. Use small pieces of toothpicks to attach the marshmallows and gumdrops to each other. Sketch each of your models.

Analysis

1. How did drawing a Lewis structure help you to determine the geometry of each of your substances?
2. Why was a mini-marshmallow used to show a shared pair of electrons and a regular marshmallow an unshared pair?
3. How can the VSEPR model help to predict the bond angles for these substances?

EXAMPLE PROBLEM 9-7

Finding the Shape of a Molecule

Phosphorus trihydride is produced when organic materials rot, and it smells like rotten fish. What is the shape of a molecule of phosphorus trihydride? Determine the bond angle, and identify the type of hybrid.

Phosphine, or phosphorus trihydride, is formed when phosphorus-containing organic material, such as this potato, rots.

1. Analyze the Problem

You are given a molecule of phosphorus trihydride that contains one phosphorus atom bonded to three hydrogen atoms. The hydrogen atoms are terminal atoms, and phosphorus is the central atom.

2. Solve for the Unknown

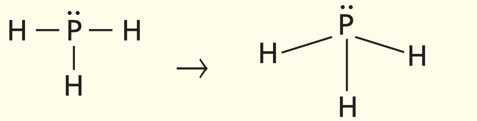
Find the total number of valence electrons and the number of electron pairs.

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

$$= 8\text{ valence electrons}$$

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

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



Lewis structure → Molecular shape

The molecular geometry is trigonal pyramidal with a bond angle of 107°. With four bonding positions, the molecule is an sp^3 hybrid.

4. Evaluate the Answer

Each atom has a stable electron configuration, and all electron pairs are accounted for.

Practice! For more practice with molecular geometry and hybridization, go to **Supplemental Practice Problems** in Appendix A.

PRACTICE PROBLEMS

Determine the molecular geometry, bond angle, and type of hybridization for the following.

49. BF_3

50. NH_4^+

51. OCl_2

52. BeF_2

53. CF_4

Section 9.4 Assessment

54. What is the VSEPR model?

55. What are the bond angles in a molecule with a tetrahedral shape?

56. What is hybridization?

57. What are the hybrid orbitals in a molecule with a tetrahedral shape?

58. **Thinking Critically** Compare the molecules PF_3 and PF_5. What is the molecular shape of each of the two molecules? What type of hybrid orbital is in each molecule? Why is the shape different?

59. **Making and Using Tables** Make a table that contains the Lewis structure, molecular shape, bond angle, and type of hybrid for the following molecules: CS_2, CH_2O, H_2Se, CCl_2F_2, and NCl_3.

Section 9.5 Electronegativity and Polarity

You now know that the type of bond that forms when two elements react depends on which elements are involved. What makes one type of bond form when carbon burns and another type form when iron corrodes? The answer lies in how much attraction each type of atom has for electrons.

Objectives

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

Vocabulary

polar covalent

Electronegativity Difference and Bond Character

Electron affinity is a measure of the tendency of an atom to accept an electron. Excluding noble gases, electron affinity increases as the atomic number increases within a given period and decreases with an increase in atomic number within a group. The scale of electronegativities allows a chemist to evaluate the electron affinity of specific atoms when they are incorporated into a compound. Recall from Chapter 6 that electronegativity indicates the relative ability of an atom to attract electrons in a chemical bond.

Look at the electronegativity values shown in **Figure 9-15.** Can you observe any trends? Note that fluorine has the highest electronegativity value (3.98), while francium has the lowest (0.7). The same trends appear with electronegativities that can be observed with electron affinities. Because noble gases do not generally form compounds, individual electronegativity values for helium, neon, and argon are not given. However, larger noble gases like xenon do occasionally bond with highly electronegative atoms such as fluorine.

Figure 9-15

Electronegativity values are not measured quantities. They are values assigned by Linus Pauling comparing the abilities of atoms to attract shared electrons with the ability of fluorine to do so.

Electronegativities

Metal
Metalloid
Nonmetal

1 H 2.20																
3 Li 0.98	4 Be 1.57											5 B 2.04	6 C 2.55	7 N 3.04	8 O 3.44	9 F 3.98
11 Na 0.93	12 Mg 1.31											13 Al 1.61	14 Si 1.90	15 P 2.19	16 S 2.58	17 Cl 3.16
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														

Elements not included on this table have no measured electronegativity or form relatively few bonds.

Lanthanide series	58 Ce 1.12	59 Pr 1.13	60 Nd 1.14	61 Pm —	62 Sm 1.17	63 Eu —	64 Gd 1.20	65 Tb —	66 Dy 1.22	67 Ho 1.23	68 Er 1.24	69 Tm 1.25	70 Yb —	71 Lu 1.0
Actinide series	90 Th 1.3	91 Pa 1.5	92 U 1.7	93 Np 1.3	94 Pu 1.3									

In the 1940s, while experimenting with a device called a magnetron that generates microwaves, a scientist noticed that a candy bar in his pocket was melting. This accidental discovery led to the creation of the microwave oven, which cooks foods based on the polarity of the molecules involved. Early microwave ovens did not sell well because they were about the size and weight of a refrigerator and cost several thousand dollars. Eventually, smaller, lower-cost versions appeared for the home kitchen.

The character and type of a chemical bond can be predicted using the electronegativity difference of the elements that are bonded. For identical atoms, which have an electronegativity difference of zero, the electrons in the bond are equally shared between the two atoms and the bond is considered nonpolar covalent, which is a pure covalent bond. Chemical bonds between atoms of different elements are never completely ionic or covalent, and the character of the bond depends on how strongly the bonded atoms attract electrons. A covalent bond formed between atoms of different elements does not have equal sharing of the electron pair because there is a difference in electronegativity. Unequal sharing results in a **polar covalent** bond. Large differences in electronegativity indicate that an electron was transferred from one atom to another, resulting in bonding that is primarily ionic.

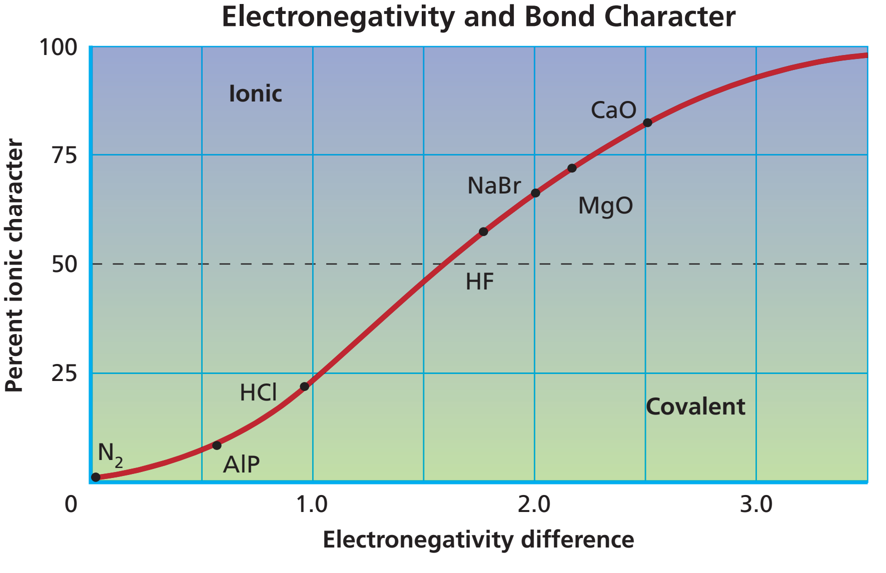
Bonding often is not clearly ionic or covalent. As the difference in electronegativity increases, the bond becomes more ionic in character. An electronegativity difference of 1.70 is considered 50 percent covalent and 50 percent ionic. Generally, ionic bonds form when the electronegativity difference is greater than 1.70. However, this cutoff is sometimes inconsistent with experimental observations of two nonmetals bonding together. **Figure 9-16** 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?

Polar Covalent Bonds

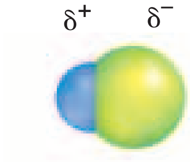
Why are some bonds polar covalent? Sharing is not always equal. Consider a tug-of-war when the two sides are not of equal strength. Although both sides share the rope, one side pulls more of the rope toward its side. Polar covalent bonds form because not all atoms that share electrons attract them equally. When a polar bond forms, the shared pair of electrons is pulled toward one of the atoms. The electrons spend more time around that atom than they do around the other atom. Partial charges occur at the ends of the bond. Using the symbols δ^-, partially negative, and δ^+, partially positive, next to a model of the molecule indicates the polarity of a polar covalent bond. See **Figure 9-17**. The more electronegative atom is located at the partially negative end, while the less electronegative atom is found at the partially positive end. The resulting polar bond often is referred to as a a dipole (two poles).

Figure 9-16

This graph shows how the percent ionic character of a bond depends on the difference in electronegativity of the atoms that form it. Above 50% ionic character, bonds are mostly ionic. What is the percent ionic character of a pure covalent bond?



Electronegativity Cl = 3.16
Electronegativity H = 2.20
Difference = 0.96



δ^+ H—Cl δ^-

Figure 9-17

In a molecule containing hydrogen and chlorine, chlorine has the higher electronegativity. Therefore, the shared pair of electrons is with the chlorine atom more often than it is with the hydrogen atom. The symbols indicating partial charge of each end of the molecule reflect this unequal sharing of electrons.

Molecular polarity Molecules are either nonpolar or polar, depending on the location and nature of the covalent bonds they contain. One way to distinguish polar from nonpolar molecules is that nonpolar molecules are not attracted by an electric field. Polar molecules tend to align with an electric field because polar molecules have a greater electron density on one side of the molecule. A polar molecule has a partial negative charge on one side, while the other side of the molecule has a partial positive charge. The molecule is a dipole because of the two partial charges.

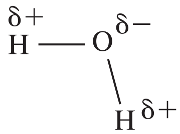
Polar molecule or not? Let's compare water (H_2O) and carbon tetrachloride (CCl_4) molecules to see why some molecules are polar and some are not. Both molecules contain polar covalent bonds. Using **Figure 9-15** on page 263, you can see that the electronegativity difference between one hydrogen atom and one oxygen atom is 1.24. The electronegativity difference between one chlorine atom and one carbon atom is 0.61. Although these electronegativity differences vary quite a bit, both the H—O and the C—Cl bonds are considered to be polar covalent.

δ^+ δ^-
H—O

δ^+ δ^-
C—Cl

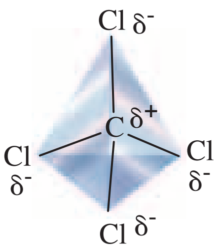
Both molecules contain more than one polar covalent bond. But water molecules are polar and carbon tetrachloride molecules are not. Examine the geometry of the molecules to see the reason for this difference.

The shape of H_2O determined by VSEPR is bent because there are two lone pairs of electrons on the central oxygen atom. Because the polar H—O bonds are not symmetric in a water molecule, the molecule has a definite positive end and a definite negative end. Thus, it is polar.



H_2O

CCl_4



The shape of CCl_4 is tetrahedral. This molecule is symmetric. The electrical charge measured at any distance from its center is identical to the charge measured at the same distance on the opposite side. The average center of the negative charge is located on the carbon atom. The positive center also is located on the carbon atom. Because the partial charges are balanced, CCl_4 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 (NH_3) shown in **Figure 9-18** polar or not? It contains a central nitrogen atom and three terminal hydrogen atoms. It has a trigonal pyramidal shape because of the lone pair of electrons present on the nitrogen atom. Using **Figure 9-15,** you can find that the electronegativity difference of hydrogen and nitrogen is 0.84. So, each N—H bond is polar covalent.

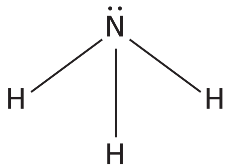


Figure 9-18

Ammonia, represented by the Lewis structure, is used in the manufacture of certain synthetic fibers.

The charge distribution is unequal because the molecule is not symmetric. Thus, the molecule is polar. The **CHEMLAB** and the **How it Works** feature at the end of this chapter are based on the polarity of molecules.

Solubility of polar molecules The ability of a substance to dissolve in another substance is known as the physical property solubility. The bond type and the shape of the molecules present determine solubility. Polar molecules and ionic compounds usually are soluble in polar substances, but nonpolar molecules dissolve only in nonpolar substances. See **Figure 9-19.** Solubility will be covered in detail in Chapter 15.

Figure 9-19

Oil, most petroleum products, and other symmetric covalent molecules are nonpolar, whereas water and other asymmetric molecules are usually polar. When polar and nonpolar substances are mixed, they separate into two layers, as seen when oil floats on water.

Properties of Covalent Compounds

You've probably noticed how similar table salt, an ionic solid, and table sugar, a covalent solid, are in appearance. But if you heat salt on the stove, it won't melt, even if the temperature is high. Sugar, on the other hand, melts at a relatively low temperature. Does type of bonding affect properties?

Differences in properties are a result of differences in attractive forces. In a covalent compound, the covalent bond between atoms in molecules is quite strong, but the attraction between individual molecules is relatively weak. The weak forces of attraction between individual molecules are known as *intermolecular forces*, or *van der Waals forces*. Intermolecular forces, which are discussed at length in Chapter 13, 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. For nonpolar substances, the attraction between the molecules is weak and is called a dispersion force, or induced dipole. The effect of dispersion forces is investigated in the **problem-solving LAB** on the next page. The force between polar molecules is stronger and is called a dipole-dipole force. This force is the attraction of one end of the dipole to the oppositely charged end of another dipole. The more polar the molecule, the stronger the dipole-dipole force. The third force, a hydrogen bond, is an especially strong intermolecular force that is formed between the hydrogen end of one dipole and a fluorine, oxygen, or nitrogen atom on another dipole.

Many physical properties of covalent molecular solids are due to intermolecular forces. The melting and boiling points of molecular substances are relatively low compared with those of ionic substances. That's why salt doesn't melt when you heat it but sugar does. 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. Hardness is also due to the intermolecular forces between individual molecules, so covalent molecules form relatively soft solids. Paraffin is a common example of a covalent solid.

In the solid state, molecules line up in a pattern forming a crystal lattice similar to that of an ionic solid, but with less attraction between particles. The structure of the crystal lattice depends on the shape of the molecule and the type of intermolecular force. Most information about molecules, including properties, molecular shape, bond length, and bond angle, has been determined by studying molecular solids.

Chemistry Online

Topic: Molecular Shapes
To learn more about molecular shapes and behavior, visit the Chemistry Web site at **chemistrymc.com**

Activity: Research the structure of an amino acid, lipid, or other biological molecule. Make a model or poster to explain how the polarity of each bond and the overall molecule affect the shape, function, and reactivity of the molecule.

PRACTICE PROBLEMS

Decide whether each of the following molecules is polar or nonpolar.

60. SCl_2 **61.** H_2S **62.** CF_4 **63.** CS_2

Practice!
For more practice with determining polarity, go to **Supplemental Practice Problems** in Appendix A.

problem-solving LAB

How do dispersion forces determine the boiling point of a substance?

Making and Using Graphs The strength of the dispersion force between nonpolar molecules determines the temperature at which these substances boil. Measuring the boiling point allows one to get a fairly good estimate of the relative strength of this force among different molecules.

Analysis

The table lists the molecular masses and boiling points for eight compounds composed of carbon and hydrogen atoms. Construct a graph showing the relationship between these quantities. Infer from the graph how dispersion forces change with increasing molecular mass. Then, draw the molecular structure for each molecule.

Thinking Critically

1. How do the relationship demonstrated by the graph and the geometry of the different hydrocarbons help to explain the effect the dispersion force has on molecules of different size?
2. How would you use your graph to predict the properties of similar but larger compounds, such as decane ($C_{10}H_{22}$)?

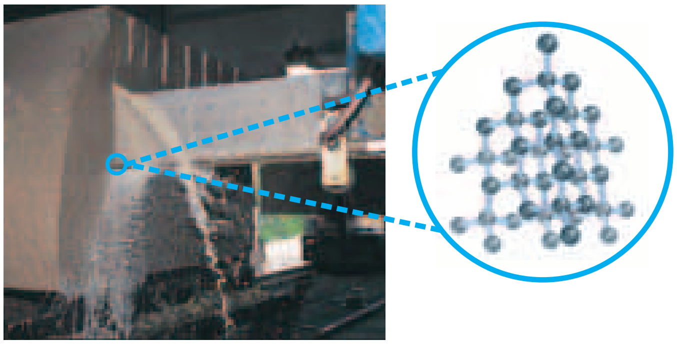
Name	Formula	M (amu)	T_{bp} (°C)
Methane	CH_4	16	−161.48
Ethane	C_2H_6	30	−88.6
Propane	C_3H_8	44	−42.1
Butane	C_4H_{10}	58	−0.5
Pentane	C_5H_{12}	72	36.06
Hexane	C_6H_{14}	86	68.73
Heptane	C_7H_{16}	100	98.5
Octane	C_8H_{18}	114	125.7

Covalent Network Solids

A number of solids are composed only of atoms interconnected by a network of covalent bonds. These solids are often called covalent network solids. Quartz is a network solid, as is diamond. See **Figure 9-20.** In contrast to molecular solids, network solids are typically brittle, nonconductors of heat or electricity, and extremely hard. In a diamond, four other carbon atoms surround each carbon atom. This tetrahedral arrangement forms a strongly bonded crystal system that is extremely hard and has a very high melting point.

Figure 9-20

In diamond, each carbon atom is bonded to four other carbon atoms. Network solids, such as diamond, are often used in cutting tools because of their hardness.



Section 9.5 Assessment

64. Define electronegativity.
65. How is electronegativity difference used in determining the type of bond that occurs between two atoms?
66. Describe a polar covalent bond.
67. What is a polar molecule?
68. List three properties of a covalent compound.
69. **Thinking Critically** Predict the type of bond that will form between the following atoms.
    a. H and S **b.** C and H **c.** Na and S
70. **Drawing Conclusions** Draw the Lewis structure for the SF_4 and SF_6 molecules and determine if each molecule is polar or nonpolar.

CHEMLAB 9

Chromatography

Paper chromatography is a common way to separate various components of a mixture. The components of the mixture separate because different substances are selectively absorbed by paper due to differences in polarity. In this field or laboratory investigation, you will separate the various pigments found in leaves. You also will calculate the ratio called R_f for each of them. The ratio R_f compares the distance traveled by a substance, D_s, to the distance traveled by the solvent, D_f. The ratio is written as $R_f = D_s / D_f$.

Problem

How can a mixture be separated based on the polarity of substances in the mixture?

Objectives

- **Separate** pigments found in leaves.
- **Determine** the R_f value for each of the pigments in the leaves.

Materials

chromatography paper (3 pieces)
2-L plastic soft drink bottle
pencils (2)
metric ruler
tape
scissors or metal snips
aluminum foil
acetone
fresh leaf samples from three different species of deciduous trees or outdoor plants.

Safety Precautions

- **Acetone is a flammable liquid. Do not use near flames or sparks.**
- **Do not allow acetone to contact skin.**
- **Perform procedure in an area with proper ventilation.**

Pre-Lab

1. Read the entire **CHEMLAB**.
2. Prepare all written materials that you will take into the laboratory. Be sure to include safety precautions, procedure notes, and a data table in which to record your observations.
3. What is polarity? How is polarity related to how chromatography works?
4. Predict what will happen when a mixture of leaf pigments is placed on a piece of paper and a solvent is allowed to move through the paper, moving the pigment with it.
5. Suppose that the pigments in two samples contain red pigment and that the red pigment in sample A is more soluble in acetone than the red pigment in sample B. Form a hypothesis regarding which red pigment has the higher R_f value. Explain your answer.

Paper Chromatography

Leaf sample	D_f (cm)	Colors	D_s (cm)	R_f
1				
2				
3				

Procedure

1. For each leaf sample, crush the leaves and soak them in a small amount of acetone to make a concentrated solution of the pigments in the leaves.
2. Cut the top off a 2-liter bottle. Cut small notches, as shown in the figure, so that a pencil can rest across the top of the bottle.
3. Cut three pieces of 3-cm wide chromatography paper to a length of about 18 cm. Label the top of each paper with a number. Assign a number to each pigment sample used. Draw pencil lines about 5 cm from the bottom of the end of each paper.
4. On the pencil line of paper 1, put a dot from the first sample. Make sure the dot is concentrated but not wide. Do the same for the other samples on their respective papers. Tape the papers to the pencil, as shown in the figure.
5. Put enough acetone in the 2-liter bottle so that when the papers are put in the bottle, the solvent touches only the bottom 1 cm of each paper, as shown in the figure. **CAUTION:** *Do not allow acetone to come in contact with skin. Use in area with proper ventilation.*
6. Carefully lower the chromatography papers into the acetone and put the pencil into the notches at the top of the bottle. Cover the top with aluminum foil. Allow the chromatograms to develop for about 35-40 minutes.

7. When the chromatograms are finished, remove them from the bottle. Mark the highest point reached by the solvent. Then, allow the papers to air dry.

Cleanup and Disposal

1. Dispose of the acetone as directed by your teacher.
2. Throw the chromatography paper in the trash can.

Analyze and Conclude

1. **Observing and Inferring** Record in the data table the colors that are found in each of the chromatograms. Space is allowed for three colors, but some samples may contain fewer or more than three colors.
2. **Measuring** For each strip, measure the distance the solvent traveled from the pencil line (D_f). For each color, measure from the top of the original marker dot to the farthest point the color traveled (D_s). Record these values in your data table.
3. **Interpreting Data** Calculate the R_f values for each of the pigments in each chromatogram and record them in the data table.
4. **Comparing and Contrasting** Describe the differences between the pigments in each of the samples.
5. **Applying Concepts** Will a polar solvent, such as water, cause a difference in how the pigments are separated? Explain your answer.
6. **Error Analysis** What could be done to improve the measurements you used to calculate R_f?

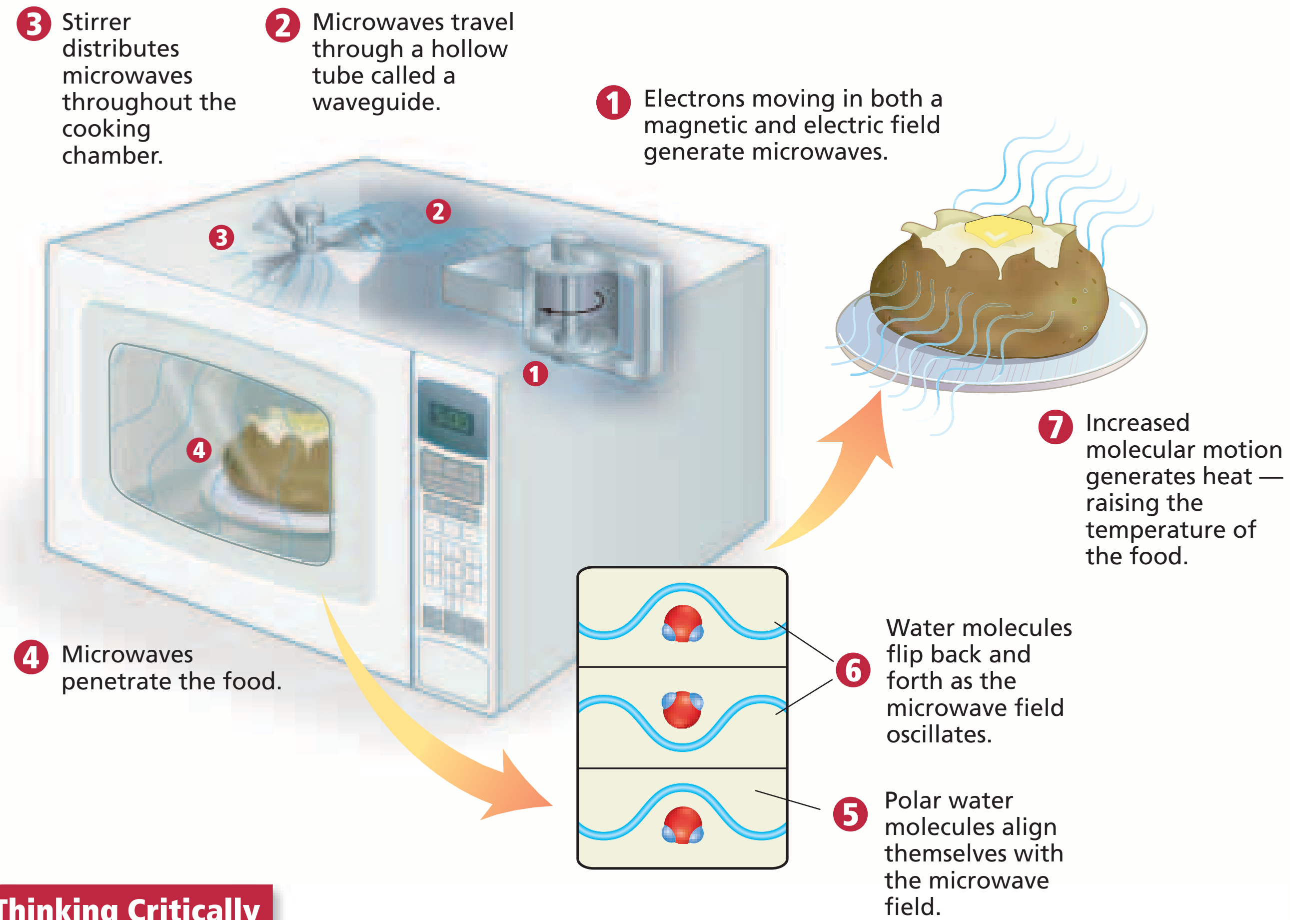
Real-World Chemistry

1. Use your results to explain what happens to leaves in autumn.
2. How might chromatography be used to analyze the composition of the dye in a marker?

How It Works

Microwave Oven

Today, about 90 percent of home kitchens in the United States have a microwave oven. The polar nature and small size of water molecules allow a microwave oven to cook food without a conventional source of heat. Many large molecules, such as those in sugars and fats, are nonpolar and are not easily heated by microwaves because the molecules cannot easily and rapidly realign themselves. Also, substances that are in the form of crystals usually do not heat well because the bonds that form the crystal structure prevent the molecules from easily realigning themselves.



Thinking Critically

1. **Predicting** How well would you expect a microwave oven to heat liquid carbon dioxide (CO_2)? CO_2 is a small molecule and its structure is O=C=O. Explain your reasoning.

2. **Predicting** Would microwaves have the same heating effect on ice as they have on liquid water? What about table salt (NaCl)? Justify your predictions.

Summary

9.1 The Covalent Bond

- A covalent bond is formed when atoms share one or more pairs of electrons.
- Molecules, formed when atoms share electrons, are more stable than their constituent atoms.
- Sharing a single pair of electrons results in a single covalent bond. Two atoms sharing more than one pair of electrons results in a multiple bond.
- A double covalent bond results when two pairs of electrons are shared between atoms. Sharing three pairs of electrons results in a triple covalent bond.
- When an electron pair is shared by the direct overlap of bonding orbitals, a sigma bond results. The overlap of parallel orbitals forms a pi bond. Single bonds are sigma bonds. Multiple bonds involve both sigma and pi bonds.
- Bond length depends on the sizes of the bonded atoms and the number of electron pairs they share. Bond dissociation energy is the energy needed to break a covalent bond. Bond length and bond dissociation energy are directly related.

9.2 Naming Molecules

- Names of covalent molecular compounds include prefixes that tell the number of each atom present.
- Molecules that produce hydrogen ions in solution are acids and are named accordingly.

9.3 Molecular Structures

- The Lewis structure is used to show the distribution of shared and lone pairs of electrons in a molecule.
- Resonance occurs when more than one valid Lewis structure exists for the same molecule.
- Exceptions to the octet rule occur when an odd number of valence electrons exists between the bonding atoms, not enough electrons are available for an octet, or more than eight electrons are shared.
- Coordinate covalent bonding occurs when one atom of the bonding pair supplies both shared electrons.

9.4 Molecular Shape

- The valence shell electron pair repulsion, or VSEPR, model can be used to predict the three-dimensional shape of a molecule. 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.
- Two orbitals form two sp hybrid orbitals, and the molecule is linear. Three orbitals, forming three sp^2 hybrid orbitals, form a molecule that is trigonal planar. Four orbitals, forming four sp^3 hybrid orbitals, form a molecule that is tetrahedral.

9.5 Electronegativity and Polarity

- Electronegativity is the tendency of an atom to attract electrons and is related to electron affinity. The electronegativity difference between two bonded atoms is used to determine the type of bond that most likely occurs.
- Polar bonds occur when electrons are not shared equally, resulting in an unequal distribution of charge and the formation of a dipole.
- The spatial arrangement of polar bonds in a molecule determines the overall polarity of a molecule.
- Weak intermolecular forces, also called van der Waals forces, hold molecules together in the liquid and solid phases. These weak attractive forces determine properties. Molecular solids tend to be soft and have low melting and boiling points.
- Covalent network solids result when each atom is covalently bonded to many other atoms in the solid. These solids are hard and have high melting points.

Vocabulary

- coordinate covalent bond (p. 257)
- covalent bond (p. 242)
- endothermic (p. 247)
- exothermic (p. 247)
- hybridization (p. 261)
- Lewis structure (p. 243)
- molecule (p. 242)
- oxyacid (p. 250)
- pi bond (p. 246)
- polar covalent (p. 264)
- resonance (p. 256)
- sigma bond (p. 245)
- structural formula (p. 252)
- VSEPR model (p. 259)

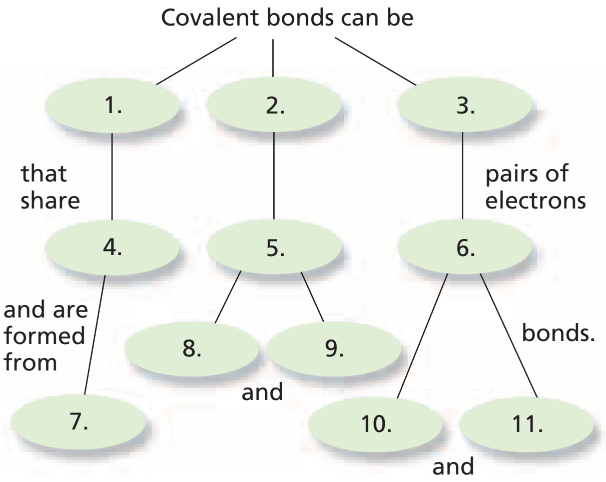
chemistrymc.com/vocabulary_puzzlemaker

CHAPTER 9 ASSESSMENT

Go to the Chemistry Web site at **chemistrymc.com** *for additional Chapter 9 Assessment.*

Concept Mapping

71. Complete the concept map using the following terms: double bonds, one, pi, sigma, single bonds, three, triple bonds, two. Each term can be used more than once.



Mastering Concepts

72. What is the octet rule, and how is it used in covalent bonding? (9.1)

73. Describe the formation of a covalent bond. (9.1)

74. Describe the bonding in molecules. (9.1)

75. Describe the forces, both attractive and repulsive, that occur as two atoms come closer together. (9.1)

76. How could you predict the presence of a sigma or pi bond in a molecule? (9.1)

77. Explain how molecular compounds are named. (9.2)

78. When is a molecular compound named as an acid? (9.2)

79. What must be known in order to draw the Lewis structure for a molecule? (9.3)

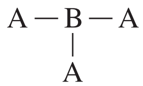
80. On what is the VSEPR model based? (9.4)

81. What is the molecular shape of each of the following molecules? Estimate the bond angle for each assuming no lone pair. (9.4)

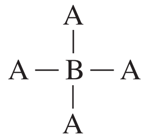
a. A—B

b. A—B—A

c.



d.



82. What is the maximum number of hybrid orbitals a carbon atom can form? (9.4)

83. Explain the theory of hybridization and determine the number of hybrid orbitals present in the molecule PCl_5. (9.4)

84. Describe the trends in electronegativity in the periodic table. (9.5)

85. Explain the difference between nonpolar molecules and polar molecules. (9.5)

86. Compare the location of bonding electrons in a polar covalent bond with those in a nonpolar covalent bond. Explain your answer. (9.5)

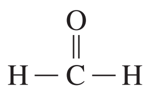
87. What is the difference between a covalent molecular solid and a covalent network solid? Do their physical properties differ? Explain your answer. (9.5)

Mastering Problems

Covalent Bonds (9.1)

88. 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.

89. Locate the sigma and pi bonds in the following molecule.



90. Locate the sigma and pi bonds in the following molecule.

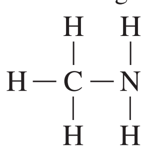
H—C ≡ C—H

Bond Length (9.1)

91. Consider the molecules CO, CO_2, and CH_2O. Which C—O bond is shorter? In which molecule is the C—O bond stronger?

92. Consider the carbon-nitrogen bonds in the following:

$C{\equiv}N^-$ and



Which bond is shorter? Which is stronger?

93. Rank each of the molecules below in order of the shortest to the longest sulfur-oxygen bond length.

a. SO_2 **b.** SO_3^{2-} **c.** SO_4^{2-}

Naming Covalent Compounds (9.2)

94. Name each of the following solutions as an acid.

a. $HClO_2$
b. H_3PO_4
c. H_2Se
d. $HClO_3$

95. Name each of the following molecules.

a. NF_3
b. NO
c. SO_3
d. SiF_4

96. Name each of the following molecules.

a. SeO_2
b. SeO_3
c. N_2F_4
d. S_4N_4

Writing Formulas (9.2)

97. Write the formula for each of the following.

a. sulfur difluoride
b. silicon tetrachloride
c. carbon tetrafluoride
d. sulfurous acid

98. Write the formula for each of the following.

a. silicon dioxide
b. bromous acid
c. chlorine trifluoride
d. hydrobromic acid

Lewis Structures (9.3)

99. Draw the Lewis structure for each of these molecules or ions.

a. H_2S
b. BF_4^-
c. SO_2
d. $SeCl_2$

100. Draw the Lewis structure for each of these molecules or ions.

a. SeF_2
b. ClO_2^-
c. PO_3^{3-}
d. $POCl_3$
e. GeF_4

101. Which of the following elements are capable of forming molecules in which an atom has an expanded octet? Explain your answer.

a. B
b. C
c. P
d. O
e. Se

102. Draw three resonance structures for the polyatomic ion CO_3^{2-}.

103. Draw two resonance structures for the polyatomic ion CHO_2^-.

104. Draw the Lewis structure for each of the following molecules that have central atoms that do not obey the octet rule.

a. PCl_5
b. BF_3
c. ClF_5
d. BeH_2

Molecular Shape (9.4)

105. Predict the molecular shape and bond angle, and identify the hybrid orbitals for each of the following. Drawing the Lewis structure may help you.

a. SCl_2
b. NH_2Cl
c. HOF
d. BF_3

106. For each of the following, predict the molecular shape.

a. COS **b.** CF_2Cl_2

107. Identify the expected hybrid on the central atom for each of the following. Drawing the Lewis structure may help you.

a. XeF_4
b. TeF_4
c. KrF_2
d. OF_2

Electronegativity and Polarity (9.5)

108. 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

109. For each of the bonds listed, tell which atom is more negatively charged.

a. C–H
b. C–N
c. C–S
d. C–O

110. Predict which of the following bonds is the most polar.

a. C–O
b. Si–O
c. C–Cl
d. C–Br

111. Rank the following bonds according to increasing polarity.

a. C–H
b. N–H
c. Si–H
d. O–H
e. Cl–H

112. Consider the following and determine if they are polar. Explain your answers.

a. H_3O^+
b. PCl_5
c. H_2S
d. CF_4

113. Why is the CF_4 molecule nonpolar even though it contains polar bonds?

114. Use Lewis structures to predict the molecular polarities for sulfur difluoride, sulfur tetrafluoride, and sulfur hexafluoride.

Mixed Review

Sharpen your problem solving skills by answering the following.

115. Consider the following molecules and determine which of the molecules are polar. Explain your answer.

a. CH_3Cl
b. ClF
c. NCl_3
d. BF_3
e. CS_2

116. Arrange the following bonds in order of least to greatest polar character.

a. C–O
b. Si–O
c. Ge–O
d. C–Cl
e. C–Br

117. Draw the Lewis structure for ClF_3 and identify the hybrid orbitals.

118. Use the Lewis structure for SF_4, to predict the molecular shape and identify the hybrid orbitals.

119. Write the formula for each of these molecules.

a. chlorine monoxide
b. arsenic acid
c. phosphorus pentachloride
d. hydrosulfuric acid

120. Name each of the following molecules.

a. PCl_3
b. Cl_2O_7
c. P_4O_6
d. NO

Thinking Critically

121. Concept Mapping Design a concept map that will link both the VSEPR model and the hybridization theory to molecular shape.

122. Making and Using Tables Complete the table using Chapters 8 and 9.

Table 9-4

Properties and Bonding			
Solid	**Bond description**	**Characteristic of solid**	**Example**
Ionic			
Covalent molecular			
Metallic			
Covalent network			

123. Drawing Conclusions Consider each of the following characteristics and determine whether the molecule is more likely to be polar or nonpolar.

a. a solid at room temperature
b. a gas at room temperature
c. attracted to an electric current

Writing in Chemistry

124. Research chromatography and write a paper discussing how it is used to separate mixtures.

125. Research laundry detergents. Write a paper to explain why they are used to clean oil and grease out of fabrics.

Cumulative Review

Refresh your understanding of previous chapters by answering the following.

126. The mass of the same liquid is given in the table below for the following volumes. Graph the volume on the *x*-axis and the mass on the *y*-axis. Calculate the slope of the graph. What information will the slope give you? (Chapter 2)

Table 9-5

Mass vs. Volume	
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

127. Which group 3A element is expected to exhibit properties that are significantly different from the remaining family members? (Chapter 7)

128. Write the correct chemical formula or name the following compounds. (Chapter 8)

a. NaI
b. calcium carbonate
c. $Fe(NO_3)_3$
d. $Sr(OH)_2$
e. potassium chlorate
f. copper(II) sulfate
g. $CoCl_2$
h. ammonium phosphate
i. silver acetate
j. $Mg(BrO_3)_2$

STANDARDIZED TEST PRACTICE
CHAPTER 9

Use the questions and the test-taking tip to prepare for your standardized test.

1. The common name of SiI_4 is tetraiodosilane. What is its molecular compound name?

a. silane tetraiodide
b. silane tetraiodine
c. silicon iodide
d. silicon tetraiodide

2. Which of the following compounds contains at least one pi bond?

a. CO_2 **c.** AsI_3
b. $CHCl_3$ **d.** BeF_2

3. What is the Lewis structure for silicon disulfide?

a. 

b. 

c. 

d. :S::Si::S:

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

a. 4 **c.** 6
b. 5 **d.** 7

5. Chloroform ($CHCl_3$) was one of the first anesthetics used in medicine. The chloroform molecule contains 26 valence electrons in total. How many of these valence electrons take part in covalent bonds?

a. 26 **c.** 8
b. 13 **d.** 4

6. Which is the strongest type of intermolecular bond?

a. ionic bond
b. dipole-dipole force
c. dispersion force
d. hydrogen bond

7. All of the following compounds have bent molecular shapes EXCEPT ______________.

a. BeH_2 **c.** H_2O
b. H_2S **d.** SeH_2

8. Which of the following compounds is NOT polar?

a. H_2S **c.** SiH_3Cl
b. CCl_4 **d.** AsH_3

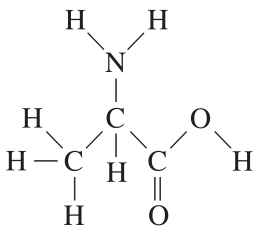
Interpreting Tables Use the table to answer the following questions.

Bond Dissociation Energies at 298 K			
Bond	**kJ/mol**	**Bond**	**kJ/mol**
Cl—Cl	242	N≡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		

9. Which of the following diatomic gases has the shortest bond between its two atoms?

a. HI **c.** Cl_2
b. O_2 **d.** N_2

10. Approximately how much energy will it take to break all of the bonds present in the molecule below?



a. 5011 kJ/mol **c.** 4318 kJ/mol
b. 3024 kJ/mol **d.** 4621 kJ/mol

TEST-TAKING TIP

Wear A Watch If you are taking a timed test, you should make sure that you pace yourself and do not spend too much time on any one question, but you shouldn't spend time staring at the clock. When each section of the test begins, set your watch for noon. This will make it very easy for you to figure out how many minutes have passed. After all, it is much easier to know that you started at 12:00 (according to your watch) and you'll be done at 12:30 than it is to figure out that you started at 10:42 and that time will run out at 11:12.