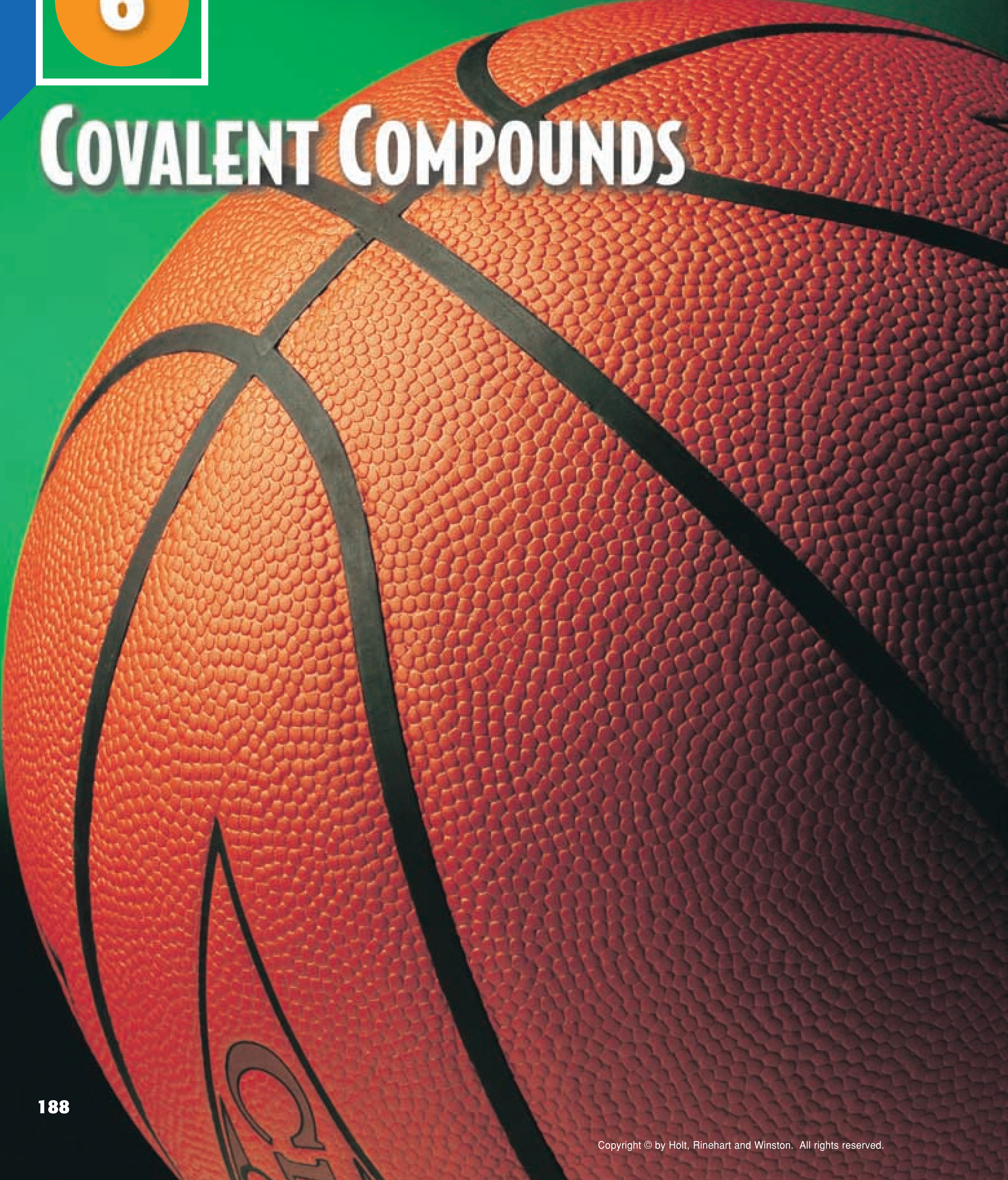
CHAPTER

6

COVALENT COMPOUNDS



Natural rubber comes from tropical trees. It is soft and sticky, so it has little practical use. However, while experimenting with rubber in 1839, Charles Goodyear dropped a mixture of sulfur and natural rubber on a hot stove by mistake. The heated rubber became tough and elastic because of the formation of covalent bonds. The resulting compound was *vulcanized rubber*, which is strong enough to make up a basketball that can take a lot of hard bounces.

START-UP ACTIVITY

SAFETY PRECAUTIONS

Ionic Versus Covalent

PROCEDURE

1. Clean and dry **three test tubes.** Place a small amount of **paraffin wax** into the first test tube. Place an equal amount of **table salt** into the second test tube. Place an equal amount of **sugar** into the third test tube.
2. Fill a **plastic-foam cup** halfway with **hot water.** Place the test tubes into the water. After 3 min, remove the test tubes from the water. Observe the contents of the test tubes, and record your observations.
3. Place a small amount of each substance on a **watch glass.** Crush each substance with a **spatula.** Record your observations.
4. Add **10 mL deionized water** to each test tube. Use a **stirring rod** to stir each test tube. Using a **conductivity device** (watch your teacher perform the conductivity tests), record the conductivity of each mixture.

ANALYSIS

1. Summarize the properties you observed for each compound.
2. Ionic bonding is present in many compounds that are brittle, have a high melting point, and conduct electric current when dissolved in water. Covalent bonding is present in many compounds that are not brittle, have a low melting point, and do not conduct electric current when mixed with water. Identify the type of bonding present in paraffin wax, table salt, and sugar.

Pre-Reading Questions

1. **What determines whether two atoms will form a bond?**
2. **How can a hydrogen atom, which has one valence electron, bond with a chlorine atom, which has seven valence electrons?**
3. **What happens in terms of energy after a hydrogen atom bonds with a chlorine atom?**

CONTENTS 6

SECTION

1 Covalent Bonds

Key Terms

- covalent bond
- molecular orbital
- bond length
- bond energy
- nonpolar covalent bond
- polar covalent bond
- dipole

Objectives

1. **Explain** the role and location of electrons in a covalent bond.
2. **Describe** the change in energy and stability that takes place as a covalent bond forms.
3. **Distinguish** between nonpolar and polar covalent bonds based on electronegativity differences.
4. **Compare** the physical properties of substances that have different bond types, and relate bond types to electronegativity differences.

Sharing Electrons

The diver shown in **Figure 1** is using a hot flame to cut metal under water. The flame is made by a chemical reaction in which hydrogen and oxygen gases combine. When these gases react, atoms and electrons rearrange to form a new, more stable compound: water.

You learned that electrons are rearranged when an ionic bond forms. When this happens, electrons transfer from one atom to another to form charged ions. The reaction of hydrogen and oxygen to form water causes another kind of change involving electrons. In this case, the neutral atoms share electrons.

Figure 1
This diver is using an oxyhydrogen torch. The energy released by the torch comes from a chemical reaction in which hydrogen and oxygen react to form water.

$$O_2 \quad + \quad 2H_2 \longrightarrow 2H_2O$$

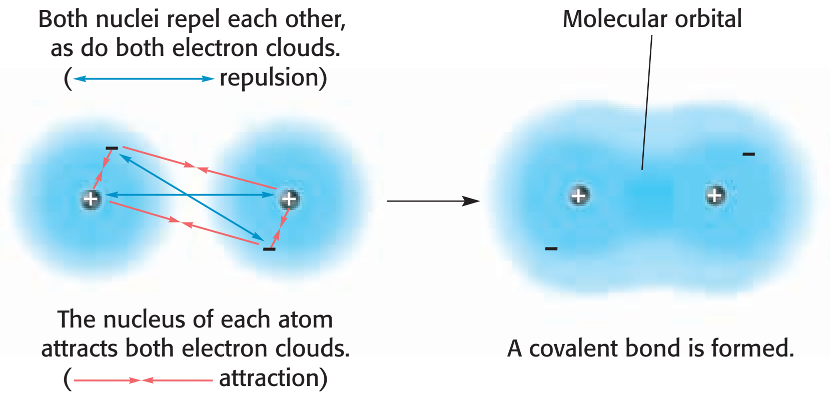


Figure 2
The positive nucleus of one hydrogen atom attracts the electron of the other atom. At the same time, the two atoms' positive nuclei repel each other. The two electron clouds also repel each other.

Forming Molecular Orbitals

The simplest example of sharing electrons is found in diatomic molecules, such as hydrogen, H_2. **Figure 2** shows the attractive and repulsive forces that exist when two hydrogen atoms are near one another. When these forces are balanced, the two hydrogen atoms form a bond. Because both atoms are of the same element, the attractive force of each atom is the same. Thus, neither atom will remove the electron from the other atom. Instead of transferring electrons to each other, the two hydrogen atoms share the electrons.

The result is a H_2 molecule that is more stable than either hydrogen atom is by itself. The H_2 molecule is stable because each H atom has a shared pair of electrons. This shared pair gives both atoms a stability similar to that of a helium configuration. Helium is stable because its atoms have filled orbitals.

The sharing of a pair of electrons is the bond that holds the two hydrogen atoms together. When two atoms share electrons, they form a **covalent bond.** The shared electrons move in the space surrounding the nuclei of the two hydrogen atoms. The space that these shared electrons move within is called a **molecular orbital.** As shown in **Figure 2,** a molecular orbital is made when two atomic orbitals overlap. Sugar and water, shown in **Figure 3,** have molecules with covalent bonds.

covalent bond
a bond formed when atoms share one or more pairs of electrons

molecular orbital
the region of high probability that is occupied by an individual electron as it travels with a wavelike motion in the three-dimensional space around one of two or more associated nuclei

Figure 3
The sugar, $C_{12}H_{22}O_{11}$, and water, H_2O, in the tea are examples of covalent, or molecular, compounds.

Energy and Stability

Most individual atoms have relatively low stability. (Noble gases are the exception.) They become more stable when they are part of a compound. Unbonded atoms also have high potential energy, as shown by the energy that is released when atoms form a compound.

After two hydrogen atoms form a covalent bond, each of them can have an electron configuration like that of helium, which has relatively low potential energy and high stability. Thus, bonding causes a decrease in energy for the atoms. This energy is released to the atoms' surroundings.

Energy Is Released When Atoms Form a Covalent Bond

Figure 4 shows the potential energy changes that take place as two hydrogen atoms come near one another. In part (a) of the figure, the distance between the two atoms is large enough that there are no forces between them. At this distance, the potential energy of the atoms is arbitrarily set at zero.

In part (b) of the figure, the potential energy decreases as the attractive electric force pulls the two atoms closer together. As the potential energy goes down, the system gives off energy. In other words, energy is released as the attractive force pulls the atoms closer. Eventually, the atoms get close enough that the attractive forces between the electrons of one atom and the nucleus of the other atom are balanced by the repulsive force caused by the two positively charged nuclei as they are forced closer together. The two hydrogen atoms are now covalently bonded. In part (c) of the figure, the two atoms have bonded, and they are at their lowest potential energy. If they get any closer, repulsive forces will take over between the nuclei.

Potential Energy Determines Bond Length

Part (c) of **Figure 4** shows that when the two bonded hydrogen atoms are at their lowest potential energy, the distance between them is 75 pm. This distance is considered the length of the covalent bond between two hydrogen atoms. The distance between two bonded atoms at their minimum potential energy is known as the **bond length.**

bond length

the distance between two bonded atoms at their minimum potential energy; the average distance between the nuclei of two bonded atoms

Potential Energy Curve for H_2

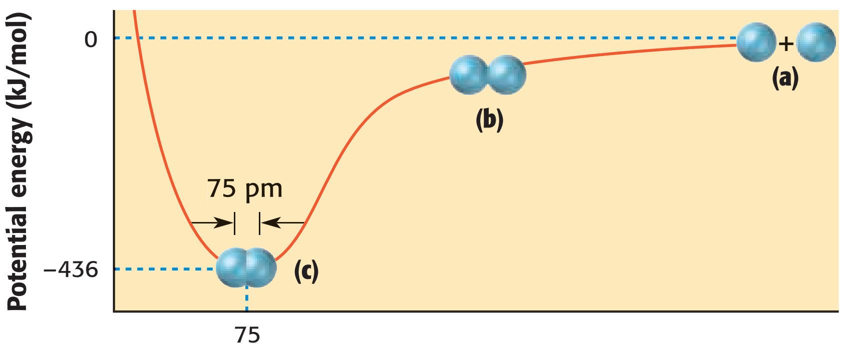


Figure 4
Two atoms form a covalent bond at a distance where attractive and repulsive forces balance. At this point, the potential energy is at a minimum.

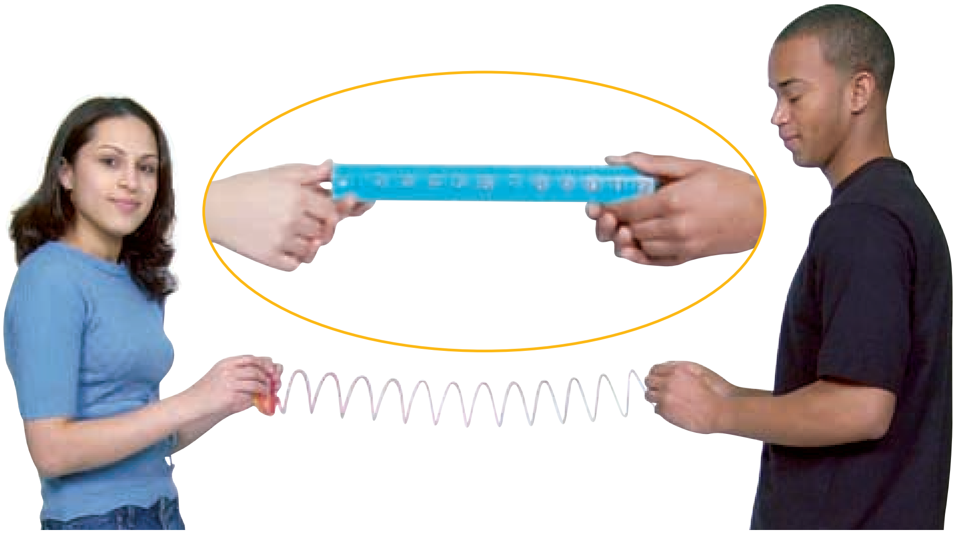


Figure 5
A covalent bond is more like a flexible spring than a rigid ruler, because the atoms can vibrate back and forth.

Bonded Atoms Vibrate, and Bonds Vary in Strength

Models often incorrectly show covalent bonds as rigid "sticks." If these bonds were in fact rigid, then the nuclei of the bonded atoms would be at a fixed distance from one another. Because the ruler held by the students in the top part of **Figure 5** is rigid, the students are at a fixed distance from one another.

However, a covalent bond is more flexible, like two students holding a spring. The two nuclei vibrate back and forth. As they do, the distance between them constantly changes. The bond length is in fact the average distance between the two nuclei.

At a bond length of 75 pm, the potential energy of H_2 is –436 kJ/mol. This means that 436 kJ of energy is released when 1 mol of bonds form. It also means that 436 kJ of energy must be supplied to break the bonds and separate the hydrogen atoms in 1 mol of H_2 molecules. The energy required to break a bond between two atoms is the **bond energy.** **Table 1** lists the energies and lengths of some common bonds in order of decreasing bond energy. Note that the bonds that have the highest bond energies (the "strongest" bonds) usually involve the elements H or F. Also note that stronger bonds generally have shorter bond lengths.

bond energy
the energy required to break the bonds in 1 mol of a chemical compound

Table 1 Bond Energies and Bond Lengths for Single Bonds

	Bond energy (kJ/mol)	Bond length (pm)		Bond energy (kJ/mol)	Bond length (pm)
H—F	570	92	**H—I**	299	161
C—F	552	138	**C—Br**	280	194
O—O	498	121	**Cl—Cl**	243	199
H—H	436	75	**C—I**	209	214
H—Cl	432	127	**Br—Br**	193	229
C—Cl	397	177	**F—F**	159	142
H—Br	366	141	**I—I**	151	266

Electronegativity and Covalent Bonding

The example in which two hydrogen atoms bond is simple because both atoms are the same. Also, each one has a single proton and a single electron, so the attractions are easy to identify. However, many covalent bonds form between two different atoms. These atoms often have different attractions for shared electrons. In such cases, electronegativity values are a useful tool to predict what kind of bond will form.

Topic Link

Refer to the "Periodic Table" chapter for more about electronegativity.

Atoms Share Electrons Equally or Unequally

Figure 6 lists the electronegativity values for several elements. In a molecule such as H_2, the values of the two atoms in the bond are equal. Because each one attracts the bonding electrons with the same force, they share the electrons equally. A **nonpolar covalent bond** is a covalent bond in which the bonding electrons in the molecular orbital are shared equally.

nonpolar covalent bond

a covalent bond in which the bonding electrons are equally attracted to both bonded atoms

What happens when the electronegativity values are not the same? If the values differ significantly, the two atoms form a different type of covalent bond. Think about a carbon atom bonding with an oxygen atom. The O atom has a higher electronegativity and attracts the bonding electrons more than the C atom does. As a result, the two atoms share the bonding electrons, but unequally. This type of bond is a **polar covalent bond.** In a polar covalent bond, the shared electrons, which are in a molecular orbital, are more likely to be found nearer to the atom whose electronegativity is higher.

polar covalent bond

a covalent bond in which a shared pair of electrons is held more closely by one of the atoms

If the difference in electronegativity values of the two atoms is great enough, the atom with the higher value may remove an electron from the other atom. An ionic bond will form. For example, the electronegativity difference between magnesium and oxygen is great enough for an O atom to remove two electrons from a Mg atom. **Figure 7** shows a model of how to classify bonds based on electronegativity differences. Keep in mind that the boundaries between bond types are arbitrary. This model is just one way that you can classify bonds. You can also classify bonds by looking at the characteristics of the substance.

Figure 6
Fluorine has an electronegativity of 4.0, the highest value of any element.

Electronegativities

H 2.2																
Li 1.0	Be 1.6											B 2.0	C 2.6	N 3.0	O 3.4	F 4.0
Na 0.9	Mg 1.3											Al 1.6	Si 1.9	P 2.2	S 2.6	Cl 3.2
K 0.8	Ca 1.0	Sc 1.4	Ti 1.5	V 1.6	Cr 1.7	Mn 1.6	Fe 1.9	Co 1.9	Ni 1.9	Cu 2.0	Zn 1.7	Ga 1.8	Ge 2.0	As 2.2	Se 2.5	Br 3.0
Rb 0.8	Sr 1.0	Y 1.2	Zr 1.3	Nb 1.6	Mo 2.2	Tc 1.9	Ru 2.2	Rh 2.3	Pd 2.2	Ag 1.9	Cd 1.7	In 1.8	Sn 1.9	Sb 2.0	Te 2.1	I 2.7
Cs 0.8	Ba 0.9	La 1.1	Hf 1.3	Ta 1.5	W 2.4	Re 1.9	Os 2.2	Ir 2.2	Pt 2.3	Au 2.5	Hg 2.0	Tl 1.8	Pb 2.1	Bi 2.0	Po 2.0	At 2.2
Fr 0.7	Ra 0.9	Ac 1.1														

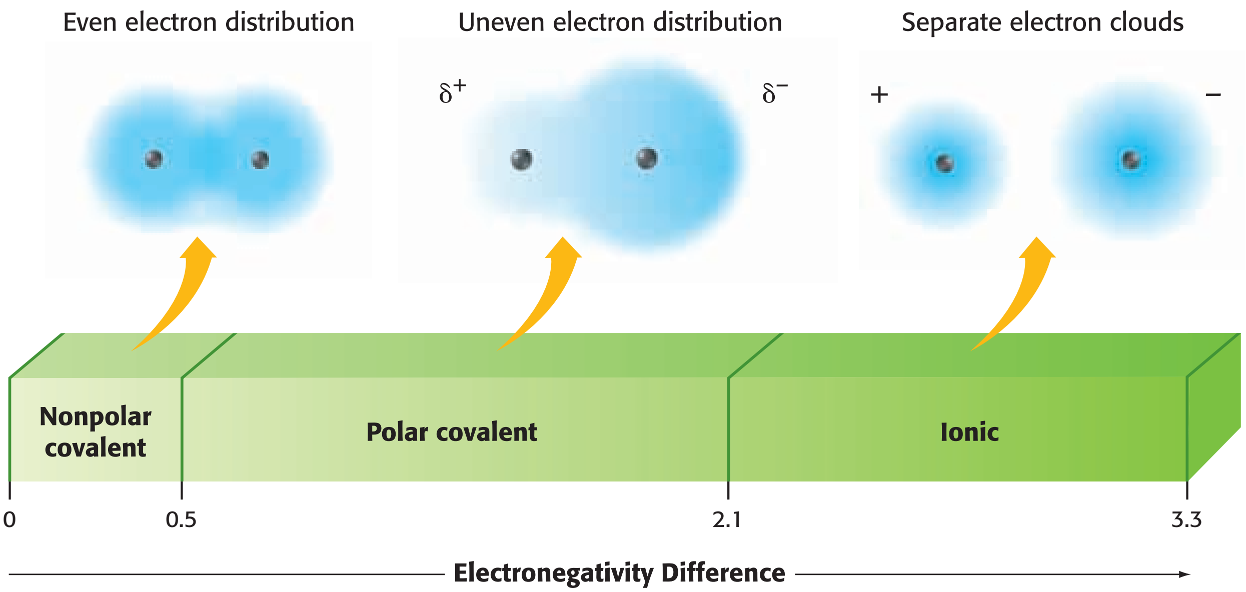


Figure 7
Electronegativity differences can be used to predict the properties of a bond. Note that there are no distinct boundaries between the bond types—the distinction is arbitrary.

Polar Molecules Have Positive and Negative Ends

Hydrogen fluoride, HF, in solution is used to etch glass, such as the vase shown in **Figure 8.** The difference between the electronegativity values of hydrogen and fluorine shows that H and F atoms form a polar covalent bond. The word *polar* suggests that this bond has ends that are in some way opposite one another, like the two poles of a planet, a magnet, or a battery. In fact, the ends of the HF molecule have opposite partial charges.

The electronegativity of fluorine (4.0) is much higher than that of hydrogen (2.2). Therefore, the shared electrons are more likely to be found nearer to the fluorine atom. For this reason, the fluorine atom in the HF molecule has a partial negative charge. In contrast, the shared electrons are less likely to be found nearer to the hydrogen atom. As a result, the hydrogen atom in the HF molecule has a partial positive charge. A molecule in which one end has a partial positive charge and the other end has a partial negative charge is called a **dipole.** The HF molecule is a dipole.

To emphasize the dipole nature of the HF molecule, the formula can be written as $H^{\delta+}F^{\delta-}$. The symbol δ is a lowercase Greek *delta,* which is used in science and math to mean *partial.* With polar molecules, such as HF, the symbol δ^{+} is used to show a partial positive charge on one end of the molecule. Likewise, the symbol δ^{-} is used to show a partial negative charge on the other end.

Although δ^{+} means a positive charge, and δ^{-} means a negative charge, these symbols do not mean that the bond between hydrogen and fluorine is ionic. An electron is not transferred completely from hydrogen to fluorine, as in an ionic bond. Instead, the atoms share a pair of electrons, which makes the bond covalent. However, the shared pair of electrons is more likely to be found nearer to the fluorine atom. This unequal distribution of charge makes the bond polar covalent.

dipole
a molecule or part of a molecule that contains both positively and negatively charged regions

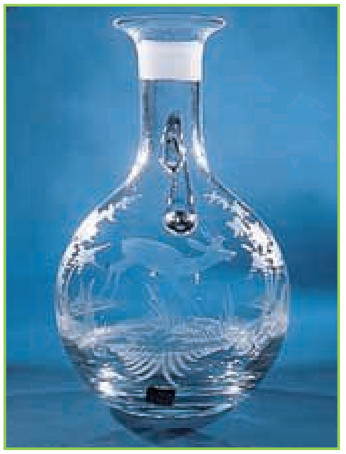


Figure 8
Hydrogen fluoride, HF, is an acid that is used to etch beautiful patterns in glass.

Table 2 Electronegativity Difference for Hydrogen Halides

Molecule	Electronegativity difference	Bond energy
H—F	1.8	570 kJ/mol
H—Cl	1.0	432 kJ/mol
H—Br	0.8	366 kJ/mol
H—I	0.5	298 kJ/mol

Polarity Is Related to Bond Strength

When examining the electronegativity differences between elements, you may notice a connection between electronegativity difference, the polarity of a bond, and the strength of that bond. The greater the difference between the electronegativity values of two elements joined by a bond, the greater the polarity of the bond. In addition, greater electronegativity differences tend to be associated with stronger bonds. Of the compounds listed in **Table 2,** H—F has the greatest electronegativity difference and thus the greatest polarity. Notice that H—F also requires the largest input of energy to break the bond and therefore has the strongest bond.

Electronegativity and Bond Types

You have learned that when sodium and chlorine react, an electron is removed from Na and transferred to Cl to form Na^+ and Cl^- ions. These ions form an ionic bond. However, when hydrogen and oxygen gas react, their atoms form a polar covalent bond by sharing electrons. How do you know which type of bond the atoms will form? Differences in electronegativity values provide one model that can tell you.

Bonds Can Be Classified by Bond Character

Figure 7 shows the relationship between electronegativity differences and the type of bond that forms between two elements. Notice the general rule that can be used to predict the type of bond that forms. If the difference in electronegativity is between 0 and 0.5, the bond is probably nonpolar covalent. If the difference in electronegativity is between 0.5 and 2.1, the bond is considered polar covalent. If the difference is larger than 2.1, then the bond is usually ionic. Remember that this method of classifying bonds is just one model. Another general rule states that covalent bonds tend to form between nonmetals, while a nonmetal and a metal will form an ionic bond.

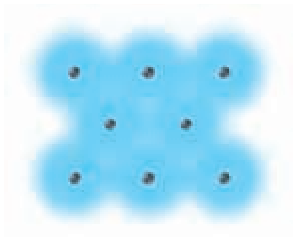
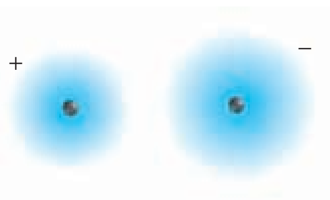
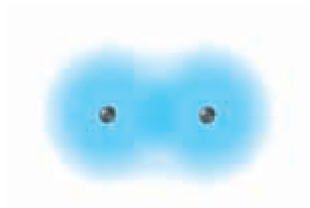
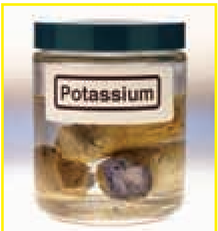
You can see how electronegativity differences provide information about bond character. Think about the bonds that form between the ions sodium and fluoride and between the ions calcium and oxide. The electronegativity difference between Na and F is 3.1. Therefore, they form an ionic bond. The electronegativity difference between Ca and O is 2.4. They also form an ionic bond. However, the larger electronegativity difference between Na and F means that the bond between them has a higher percentage of ionic character.

Next think about the bonds that form between carbon and chlorine and between aluminum and chlorine. The electronegativity difference between C and Cl is 0.6. These two elements form a polar covalent bond. The electronegativity difference between Al and Cl is 1.6. These two elements also form a polar covalent bond. However, the larger difference between Al and Cl means that the bond between these two elements is more polar, with greater partial charges, than the bond between C and Cl is.

Properties of Substances Depend on Bond Type

The type of bond that forms determines the physical and chemical properties of the substance. For example, metals, such as potassium, are very good electric conductors in the solid state. This property is the result of metallic bonding. Metallic bonds are the result of the attraction between the electrons in the outermost energy level of each metal atom and all of the other atoms in the solid metal. The metal atoms are held in the solid because all of the valence electrons are attracted to all of the atoms in the solid. These valence electrons can move easily from one atom to another. They are free to roam around in the solid and can conduct an electric current.

Table 3 Properties of Substances with Metallic, Ionic, and Covalent Bonds

Bond type	**Metallic**	**Ionic**	**Covalent**
			
Example substance	potassium	potassium chloride	chlorine
Melting point (°C)	63	770	–101
Boiling point (°C)	760	1500 (sublimes)	–34.6
Properties	• soft, silvery, solid • conductor as a solid 	• crystalline, white solid • conductor when dissolved in water 	• greenish yellow gas • not a good conductor 

In ionic substances, the overall attraction between all the cations and anions is very strong. Ionic compounds, such as potassium chloride, KCl, are made up of many K^+ and Cl^- ions. Each ion is held into place by many oppositely charged neighbors, so the forces—the ionic bonds—that hold them together are very strong and hard to break.

In molecular substances, such as Cl_2, the molecules are held together by sharing electrons. The shared electrons are attracted to the two bonding atoms, and they have little attraction for the atoms of other nearby molecules. Therefore, the attractive forces between separate Cl_2 molecules are very small compared to the attractive forces between the ions in KCl.

The difference in the strength of attraction between the basic units of ionic and molecular substances gives rise to different properties in the two types of substances. For example, the stronger the force between the ions or molecules of a substance in a liquid state, the more energy is required for the substance to change into a gas. **Table 3** shows that the strong forces in ionic substances, such as KCl, account for the high melting and boiling points they have compared to molecular substances, such as Cl_2. You will learn more about this relationship in a later chapter. The table also compares the conductivity of each substance.

1 Section Review

UNDERSTANDING KEY IDEAS

1. Describe the attractive forces and repulsive forces that exist between two atoms as the atoms move closer together.
2. Compare a bond between two atoms to a spring between two students.
3. In what two ways can two atoms share electrons when forming a covalent bond?
4. What happens in terms of energy and stability when a covalent bond forms?
5. How are the partial charges shown in a polar covalent molecule?
6. What information can be obtained by knowing the electronegativity differences between two elements?
7. Why do molecular compounds have low melting points and low boiling points relative to ionic substances?

CRITICAL THINKING

8. Why does the distance between two nuclei in a covalent bond vary?
9. How does a molecular orbital differ from an atomic orbital?
10. How does the strength of a covalent bond relate to bond length?
11. Compare the degree of polarity in HF, HCl, HBr, and HI.
12. Given that it has the highest electronegativity, can a fluorine atom ever form a nonpolar covalent bond? Explain your answer.
13. What does a small electronegativity difference reveal about the strength of a covalent bond?
14. Based on electronegativity values, which bond has the highest degree of ionic character: H—S, Si—Cl, or Cs—Br?

SECTION 2

Drawing and Naming Molecules

Key Terms

- valence electron
- Lewis structure
- unshared pair
- single bond
- double bond
- triple bond
- resonance structure

Objectives

1. **Draw** Lewis structures to show the arrangement of valence electrons among atoms in molecules and polyatomic ions.
2. **Explain** the differences between single, double, and triple covalent bonds.
3. **Draw** resonance structures for simple molecules and polyatomic ions, and recognize when they are required.
4. **Name** binary inorganic covalent compounds by using prefixes, roots, and suffixes.

Lewis Electron-Dot Structures

Both ionic and covalent bonds involve **valence electrons,** the electrons in the outermost energy level of an atom. In 1920, G. N. Lewis, the American chemist shown in **Figure 9,** came up with a system to represent the valence electrons of an atom. This system—known as electron-dot diagrams or **Lewis structures** —uses dots to represent valence electrons. Lewis's system is a valuable model for covalent bonding. However, these diagrams do not show the actual locations of the valence electrons. They are models that help you to keep track of valence electrons.

valence electron

an electron that is found in the outermost shell of an atom and that determines the atom's chemical properties

Lewis structure

a structural formula in which electrons are represented by dots; dot pairs or dashes between two atomic symbols represent pairs in covalent bonds

Lewis Structures Model Covalently Bonded Molecules

A Lewis structure shows only the valence electrons in an atom or molecule. The nuclei and the electrons of the inner energy levels (if any) of an atom are represented by the symbol of the element. With only one valence electron, a hydrogen atom has the electron configuration $1s^1$. When drawing hydrogen's Lewis structure, you represent the nucleus by the element's symbol, H. The lone valence electron is represented by a dot.

$$\mathrm{H}\cdot$$

When two hydrogen atoms form a nonpolar covalent bond, they share two electrons. These two electrons are represented by a pair of dots between the symbols.

$$\mathrm{H}:\mathrm{H}$$

This Lewis structure represents a stable hydrogen molecule in which both atoms share the same pair of electrons.

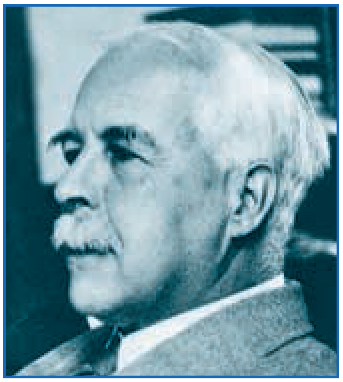


Figure 9
G. N. Lewis (1875–1946) not only came up with important theories of bonding but also gave a new definition to acids and bases.

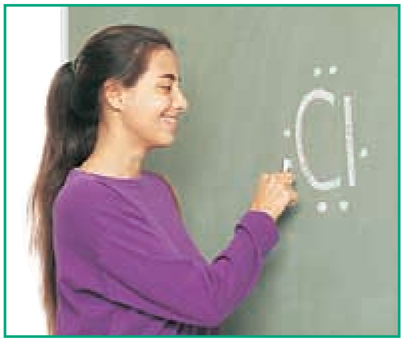


Figure 10
An electron configuration shows all of the electrons of an atom, while the Lewis structure, above, shows only the valence electrons.

Table 4 Lewis Structures of the Second-Period Elements

Element	Electron configuration	Number of valence electrons	Lewis structure (for bonding)
Li	$1s^2\mathbf{2s^1}$	1	Li·
Be	$1s^2\mathbf{2s^2}$	2	·Be·
B	$1s^2\mathbf{2s^22p^1}$	3	·B· (3 dots)
C	$1s^2\mathbf{2s^22p^2}$	4	·C· (4 dots)
N	$1s^2\mathbf{2s^22p^3}$	5	:N· (5 dots)
O	$1s^2\mathbf{2s^22p^4}$	6	:O· (6 dots)
F	$1s^2\mathbf{2s^22p^5}$	7	:F· (7 dots)
Ne	$1s^2\mathbf{2s^22p^6}$	8	:Ne: (8 dots)

Lewis Structures Show Valence Electrons

The Lewis structure of a chlorine atom shows only the atom's seven valence electrons. Its Lewis structure is written with three pairs of electrons and one unpaired electron around the element's symbol, as shown below and in **Figure 10.**

:C̈l·

Table 4 shows the Lewis structures of the elements in the second period of the periodic table as they would appear in a bond. Notice that as you go from element to element across the period, you add a dot to each side of the element's symbol. You do not begin to pair dots until all four sides of the element's symbol have a dot.

An element with an octet of valence electrons, such as that found in the noble gas Ne, has a stable configuration. When two chlorine atoms form a covalent bond, each atom contributes one electron to a shared pair. With this shared pair, both atoms can have a stable octet. This tendency of bonded atoms to have octets of valence electrons is called the *octet rule.*

:C̈l:C̈l:

unshared pair

a nonbonding pair of electrons in the valence shell of an atom; also called *lone pair*

single bond

a covalent bond in which two atoms share one pair of electrons

Each chlorine atom in Cl_2 has three pairs of electrons that are not part of the bond. These pairs are called **unshared pairs** or *lone pairs.* The pair of dots that represents the shared pair of electrons can also be shown by a long dash. Both notations represent a **single bond.**

:C̈l:C̈l: or :C̈l—C̈l:

For Lewis structures of bonded atoms, you may want to keep in mind that when the dots for the valence electrons are placed around the symbol, each side must contain an unpaired electron before any side can contain a pair of electrons. For example, see the Lewis structure for a carbon atom below.

$$\cdot\dot{\underset{\cdot}{\mathrm{C}}}\cdot$$

The electrons can pair in any order. However, any unpaired electrons are usually filled in to show how they will form a covalent bond. For example, think about the bonding between hydrogen and chlorine atoms.

$$\mathrm{H}\cdot \ + \ \cdot\ddot{\underset{\cdot\cdot}{\mathrm{Cl}}}\!: \ \longrightarrow \ \mathrm{H}{-}\ddot{\underset{\cdot\cdot}{\mathrm{Cl}}}\!:$$

SKILLS Toolkit 1

Drawing Lewis Structures with Many Atoms

1. Gather information.
- Draw a Lewis structure for each atom in the compound. When placing valence electrons around an atom, place one electron on each side before pairing any electrons.
- Determine the total number of valence electrons in the compound.

2. Arrange the atoms.
- Arrange the Lewis structure to show how the atoms bond in the molecule.
- Halogen and hydrogen atoms *often* bind to only one other atom and are *usually* at an end of the molecule.
- Carbon is often placed in the center of the molecule.
- You will find that, with the exception of carbon, the atom with the lowest electronegativity is often the central atom.

3. Distribute the dots.
- Distribute the electron dots so that each atom, except for hydrogen, beryllium, and boron, satisfies the octet rule.

4. Draw the bonds.
- Change each pair of dots that represents a shared pair of electrons to a long dash.

5. Verify the structure.
- Count the number of electrons surrounding each atom. Except for hydrogen, beryllium, and boron, all atoms must satisfy the octet rule. Check that the number of valence electrons is still the same number you determined in step 1.

SAMPLE PROBLEM A

Drawing Lewis Structures with Single Bonds

Draw a Lewis structure for CH_3I.

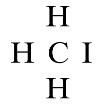
1 Gather information.

Draw each atom's Lewis structure, and count the total number of valence electrons.

number of dots: 14

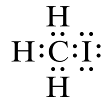
2 Arrange the atoms.

Arrange the Lewis structure so that carbon is the central atom.



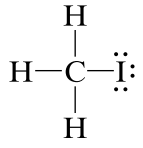
3 Distribute the dots.

Distribute one bonding pair of electrons between each of the bonded atoms. Then, distribute the remaining electrons, in pairs, around the remaining atoms to form an octet for each atom.



4 Draw the bonds.

Change each pair of dots that represents a shared pair of electrons to a long dash.



5 Verify the structure.

Carbon and iodine have 8 electrons, and hydrogen has 2 electrons. The total number of valence electrons is still 14.

PRACTICE HINT

You may have to try several Lewis structures until you get one in which all of the atoms, except hydrogen, beryllium, and boron, obey the octet rule.

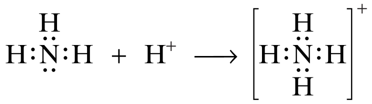
PRACTICE

PROBLEM SOLVING SKILL

1. Draw the Lewis structures for H_2S, CH_2Cl_2, NH_3, and C_2H_6.
2. Draw the Lewis structure for methanol, CH_3OH. First draw the CH_3 part, and then add O and H.

Lewis Structures for Polyatomic Ions

Lewis structures are also helpful in describing polyatomic ions, such as the ammonium ion, NH_4^+. An ammonium ion, shown in **Figure 11,** forms when ammonia, NH_3, is combined with a substance that easily gives up a hydrogen ion, H^+. To draw the Lewis structure of NH_4^+, first draw the structure of NH_3. With five valence electrons, a nitrogen atom can make a stable octet by forming three covalent bonds, one with each hydrogen atom. Then add H^+, which is simply the nucleus of a hydrogen atom, or a proton, and has no electrons to share. The H^+ can form a covalent bond with NH_3 by bonding with the unshared pair on the nitrogen atom.



The Lewis structure is enclosed in brackets to show that the positive charge is distributed over the entire ammonium ion.

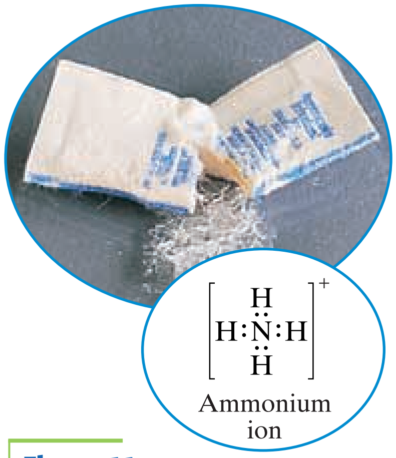


Figure 11
Smelling salts often have an unstable ionic compound made of two polyatomic ions: ammonium and carbonate.

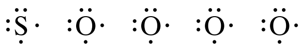
SAMPLE PROBLEM B

Drawing Lewis Structures for Polyatomic Ions

Draw a Lewis structure for the sulfate ion, SO_4^{2-}.

1 Gather information.

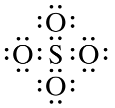
When counting the total number of valence electrons, add two additional electrons to account for the 2– charge on the ion.



number of dots: 30 + 2 = 32

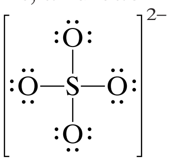
2 Arrange the atoms. Distribute the dots.

Sulfur has the lowest electronegativity, so it is the central atom. Distribute the 32 dots so that there are 8 dots around each atom.



3 Draw the bonds. Verify the structure.

- Change each bonding pair to a long dash. Place brackets around the ion and a 2– charge outside the bracket to show that the charge is spread out over the entire ion.
- There are 32 valence electrons, and each O and S has an octet.



PRACTICE HINT

- If the polyatomic ion has a negative charge, add the appropriate number of valence electrons. (For example, the net charge of 2– on SO_4^{2-} means that there are two more electrons than in the neutral atoms.)
- If the polyatomic ion has a positive charge, subtract the appropriate number of valence electrons. (For example, the net charge of 1+ on H_3O^+ means that there is one fewer electron than in the neutral atoms.)

PRACTICE

1. Draw the Lewis structure for ClO_3^-.
2. Draw the Lewis structure for the hydronium ion, H_3O^+.

PROBLEM SOLVING SKILL

Multiple Bonds

Atoms can share more than one pair of electrons in a covalent bond. Think about a nonpolar covalent bond formed between two oxygen atoms in an O_2 molecule. Each O has six valence electrons, as shown below.

If these oxygen atoms together shared only one pair of electrons, each atom would have only seven electrons. The octet rule would not be met.

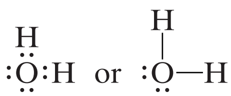
Bonds with More than One Pair of Electrons

To make an octet, each oxygen atom needs two more electrons to be added to its original six. To add two electrons, each oxygen atom must share two electrons with the other atom so that the two atoms share four electrons. The covalent bond formed by the sharing of two pairs of electrons is a **double bond,** shown in the Lewis structures below.

double bond

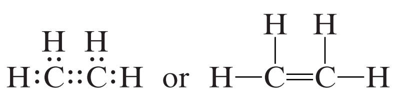
a covalent bond in which two atoms share two pairs of electrons

Atoms will form a single or a multiple bond depending on what is needed to make an octet. While two O atoms form a double bond in O_2, an O atom forms a single bond with each of two H atoms in a water molecule.



Another example of a molecule that has a double bond is ethene, C_2H_4, shown in **Figure 12.** Each H atom forms a single bond with a C atom. Each C atom below has two electrons that are not yet part of a bond.

With only six electrons, each C atom needs two more electrons to have an octet. The only way to complete the octets is to form a **double bond.**



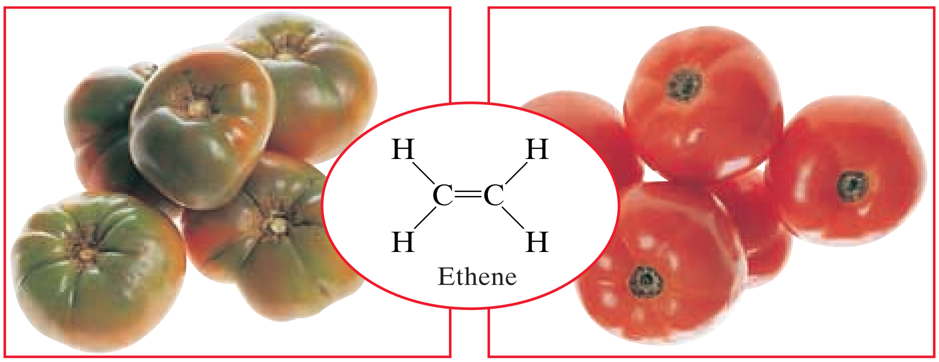


Figure 12
Most plants have a hormone called *ethene,* C_2H_4. Tomatoes release ethene, also called *ethylene,* as they ripen.

Carbon, oxygen, and nitrogen atoms often form double bonds by sharing two pairs of electrons. Carbon and nitrogen atoms may even share three pairs of electrons to form a **triple bond.** Think about the molecule N_2. With five valence electrons, each N atom needs three more electrons for a stable octet. Each N atom contributes three electrons to form three bonding pairs. The two N atoms form a triple bond by sharing these three pairs of electrons, or a total of six electrons. Because the two N atoms share the electrons equally, the triple bond is a nonpolar covalent bond.

$$:N:::N: \quad \text{or} \quad :N\equiv N:$$

triple bond

a covalent bond in which two atoms share three pairs of electrons

SAMPLE PROBLEM C

Drawing Lewis Structures with Multiple Bonds

Draw a Lewis structure for formaldehyde, CH_2O.

1 Gather information.

Draw each atom's Lewis structure, and count the total dots.

$\cdot\dot{\underset{\cdot}{C}}\cdot \quad H\cdot \quad H\cdot \quad \cdot\ddot{O}:$ total dots: 12

2 Arrange the atoms. Distribute the dots.

- Arrange the atoms so that carbon is the central atom.
- Distribute one pair of dots between each of the atoms. Then, starting with the outside atoms, distribute the rest of the dots, in pairs, around the atoms. You will run out of electrons before all of the atoms have an octet (left structure). C does not have an octet, so there must be a multiple bond. To obtain an octet for C, move one of the unshared pairs from the O atom to between the O and the C (right structure).

incorrect:
```
 :Ö:
H:C:H
```
correct:
```
 :O:
  ::
H:C:H
```

3 Draw the bonds. Verify the structure.

- Change each pair of dots that represents a shared pair of electrons to a long dash. Two pairs of dots represent a double bond.
- C and O atoms both have eight electrons, and each H atom has two electrons. The total number of valence electrons is still 12.

```
   :O:
    ‖
H—C—H
```

PRACTICE HINT

- Begin with a single pair of dots between each pair of bonded atoms. If no arrangement of single bonds provides a Lewis structure whose atoms satisfy the octet rule, the molecule might have multiple bonds.
- N and C can form single bonds or combinations of single and double or triple bonds.

PRACTICE

1. Draw the Lewis structures for carbon dioxide, CO_2, and carbon monoxide, CO.
2. Draw the Lewis structures for ethyne, C_2H_2, and hydrogen cyanide, HCN.

PROBLEM SOLVING SKILL

Figure 13
You can draw resonance structures for sulfur dioxide, SO_2, a chemical that can add to air pollution.

resonance structure

in chemistry, any one of two or more possible configurations of the same compound that have identical geometry but different arrangements of electrons

Topic Link

Refer to the "Ions and Ionic Compounds" chapter for more about naming ionic compounds.

Resonance Structures

Some molecules, such as ozone, O_3, cannot be represented by a single Lewis structure. Ozone has two Lewis structures, as shown below.

$$\ddot{O}=\ddot{O}-\ddot{O}: \leftrightarrow :\ddot{O}-\ddot{O}=\ddot{O}$$

Each O atom follows the octet rule, but the two structures use different arrangements of the single and double bonds. So which structure is correct? Neither structure is correct by itself. When a molecule has two or more possible Lewis structures, the two structures are called **resonance structures.** You place a double-headed arrow between the structures to show that the actual molecule is an average of the two possible states.

Another molecule that has resonance structures is sulfur dioxide, SO_2, shown in **Figure 13.** Sulfur dioxide released into the atmosphere is partly responsible for acid precipitation. The actual structure of SO_2 is an average, or a *resonance hybrid,* of the two structures. Although you draw the structures as if the bonds change places again and again, the bonds do not in fact move back and forth. The actual bonding is a mixture of the two extremes represented by each of the Lewis structures.

Naming Covalent Compounds

Covalent compounds made of two elements are named by using a method similar to the one used to name ionic compounds. Think about how the covalent compound SO_2 is named. The first element named is usually the first one written in the formula, in this case *sulfur.* Sulfur is the less-electronegative element. The second element named has the ending *-ide,* in this case *oxide.*

However, unlike the names for ionic compounds, the names for covalent compounds must often distinguish between two different molecules made of the same elements. For example, SO_2 and SO_3 cannot both be called *sulfur oxide.* These two compounds are given different names based on the number of each type of atom in the compound.

Prefixes Indicate How Many Atoms Are in a Molecule

The system of prefixes shown in **Table 5** is used to show the number of atoms of each element in the molecule. SO_2 and SO_3 are distinguished from one another by the use of prefixes in their names. With only two oxygen atoms, SO_2 is named *sulfur dioxide.* With three oxygen atoms, SO_3 is named *sulfur trioxide.* The following example shows how to use the system of prefixes to name P_2S_5.

P_2S_5

Prefix needed if there is more than one atom of the less-electronegative element + Name of less-electronegative element

diphosphorus

Prefix that shows the number of atoms of the more-electronegative element + Root name of more electronegative element + *ide*

pentasulfide

Table 5 Prefixes for Naming Covalent Compounds

Prefix	Number of atoms	Example	Name
mono-	1	CO	carbon monoxide
di-	2	SiO_2	silicon dioxide
tri-	3	SO_3	sulfur trioxide
tetra-	4	SCl_4	sulfur tetrachloride
penta-	5	$SbCl_5$	antimony pentachloride

Refer to Appendix A for a more complete list of prefixes.

Prefixes are added to the first element in the name only if the molecule contains more than one atom of that element. So, N_2O is named *dinitrogen oxide,* S_2F_{10} is named *disulfur decafluoride,* and P_4O_6 is named *tetraphosphorus hexoxide.* If the molecule contains only one atom of the first element given in the formula, the prefix *mono-* is left off. Both SO_2 and SO_3 have only one S atom each. Therefore, the names of both start with the word *sulfur.* Note that the vowels *a* and *o* are dropped from a prefix that is added to a word begining with a vowel. For example, CO is carbon monoxide, not carbon monooxide. Similarly, N_2O_4 is named *dinitrogen tetroxide,* not *dinitrogen tetraoxide.*

2 Section Review

UNDERSTANDING KEY IDEAS

1. Which electrons do a Lewis structure show?

2. In a polyatomic ion, where is the charge located?

3. How many electrons are shared by two atoms that form a triple bond?

4. What do resonance structures represent?

5. How do the names for SO_2 and SO_3 differ?

PRACTICE PROBLEMS

6. Draw a Lewis structure for an atom that has the electron configuration $1s^22s^22p^63s^23p^3$.

7. Draw Lewis structures for each compound:

a. BrF **c.** Cl_2O

b. $N(CH_3)_3$ **d.** ClO_2^-

8. Draw three resonance structures for SO_3.

9. Name the following compounds.

a. SnI_4 **c.** PCl_3

b. N_2O_3 **d.** CSe_2

10. Write the formula for each compound:

a. phosphorus pentabromide

b. diphosphorus trioxide

c. arsenic tribromide

d. carbon tetrachloride

CRITICAL THINKING

11. Compare and contrast the Lewis structures for krypton and radon.

12. Do you always follow the octet rule when drawing a Lewis structure? Explain.

13. What is incorrect about the name *monosulfur dioxide* for the compound SO_3?

SECTION 3

Molecular Shapes

KEY TERM

- **VSEPR theory**

OBJECTIVES

1. **Predict** the shape of a molecule using VSEPR theory.
2. **Associate** the polarity of molecules with the shapes of molecules, and relate the polarity and shape of molecules to the properties of a substance.

Determining Molecular Shapes

Lewis structures are two-dimensional and do not show the three-dimensional shape of a molecule. However, the three-dimensional shape of a molecule is important in determining the molecule's physical and chemical properties. Sugar, or sucrose, is an example. Sucrose has a shape that fits certain nerve receptors on the tongue. Once stimulated, the nerves send signals to the brain, and the brain interprets these signals as sweetness. Inside body cells, sucrose is processed for energy.

People who want to avoid sucrose in their diet often use a sugar substitute, such as sucralose, shown in **Figure 14.** These substitutes have shapes similar to that of sucrose, so they can stimulate the nerve receptors in the same way that sucrose does. However, sucralose has a different chemical makeup than sucrose does and cannot be processed by the body.

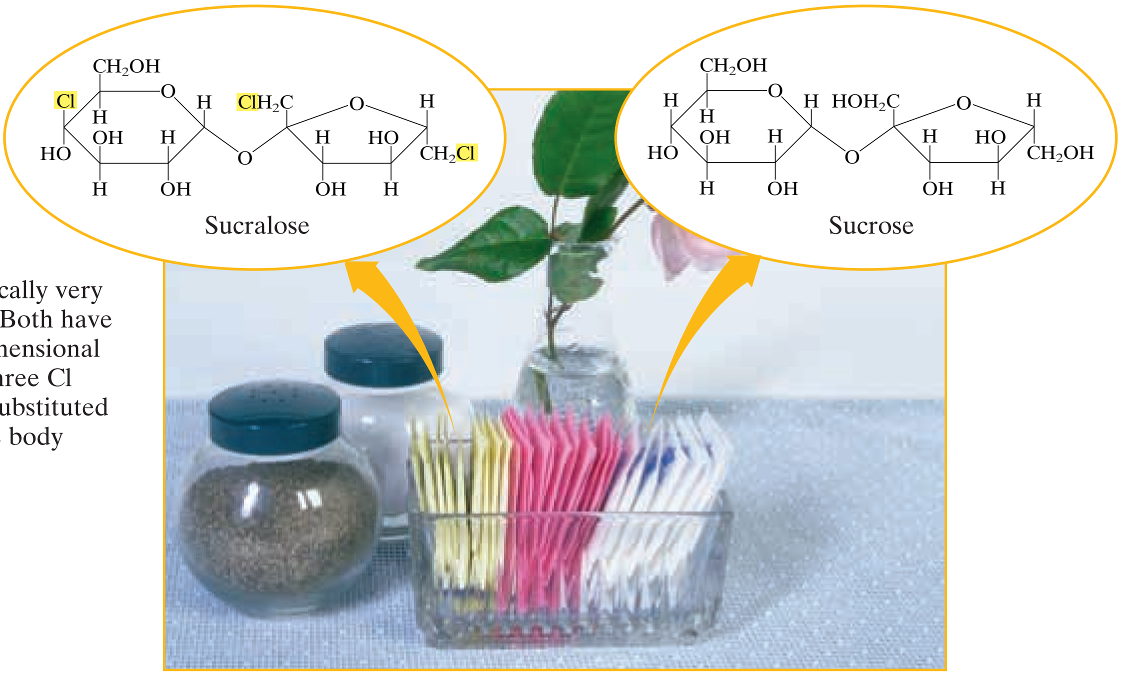
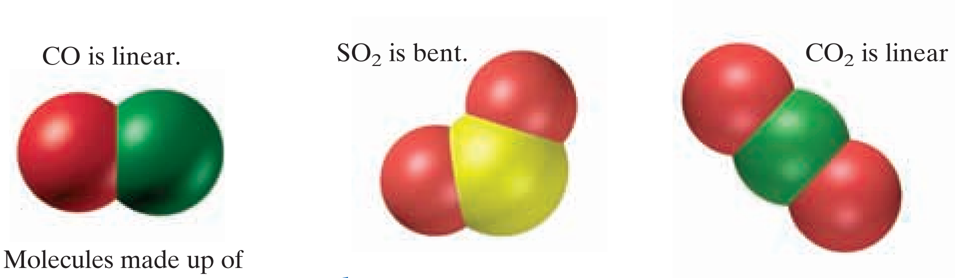


Figure 14
Sucralose is chemically very similar to sucrose. Both have the same three-dimensional shape. However, three Cl atoms have been substituted in sucralose, so the body cannot process it.



a Molecules made up of only two atoms, such as CO, have a linear shape.

b Although SO_2 and CO_2 have the same numbers of atoms, they have different shapes because the numbers of electron groups surrounding the central atoms differ.

Figure 15
Molecules with three or fewer atoms have shapes that are in a flat plane.

A Lewis Structure Can Help Predict Molecular Shape

The shape of a molecule made of only two atoms, such as H_2 or CO, is easy to determine. As shown in **Figure 15,** only a linear shape is possible when there are two atoms. Determining the shapes of molecules made of more than two atoms is more complicated. Compare carbon dioxide, CO_2, and sulfur dioxide, SO_2. Both molecules are made of three atoms. Although the molecules have similar formulas, their shapes are different. Notice that CO_2 is linear, while SO_2 is bent.

Obviously, the formulas CO_2 and SO_2 do not provide any information about the shapes of these molecules. However, there is a model that can be used to predict the shape of a molecule. This model is based on the **valence shell electron pair repulsion (VSEPR) theory.** Using this model, you can predict the shape of a molecule by examining the Lewis structure of the molecule.

valence shell electron pair repulsion (VSEPR) theory
a theory that predicts some molecular shapes based on the idea that pairs of valence electrons surrounding an atom repel each other

Electron Pairs Can Determine Molecular Shape

According to the VSEPR theory, the shape of a molecule is determined by the valence electrons surrounding the central atom. For example, examine the Lewis structure for CO_2.

$$\ddot{\underset{..}{O}}=C=\ddot{\underset{..}{O}}$$

Notice the two double bonds around the central carbon atom. Because of their negative charge, electrons repel each other. Therefore, the two shared pairs that form each double bond repel each other and remain as far apart as possible. These two sets of two shared pairs are farthest apart when they are on opposite sides of the carbon atom. Thus, the shape of a CO_2 molecule is linear. You'll read about SO_2's bent shape later.

Now think about what happens when the central atom is surrounded by three shared pairs. Look at the Lewis structure for BF_3, which has boron, an example of an atom that does not always obey the octet rule.

$$:\ddot{\underset{..}{F}}-\underset{\underset{:\underset{..}{F}:}{|}}{B}-\ddot{\underset{..}{F}}:$$

Notice the three single bonds around the central boron atom. Like three spokes of a wheel, these shared pairs of electrons extend from the central boron atom. The three F atoms, each of which has three unshared pairs, will repel each other and will be at a maximum distance apart. This molecular shape is known as *trigonal planar,* as shown in **Figure 16.**

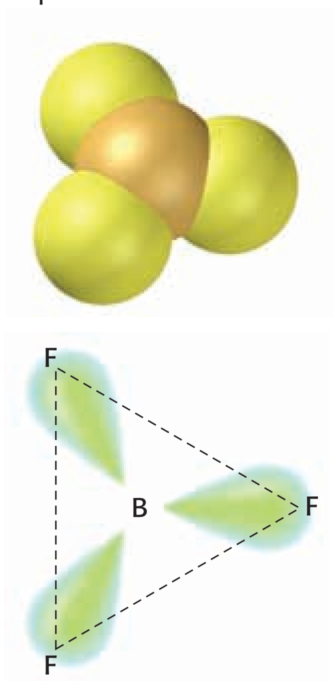


Figure 16
Trigonal planar molecules, such as BF_3, are flat structures in which three atoms are evenly spaced around the central atom.

internet connect

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Topic: Molecular Shapes
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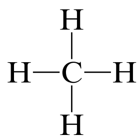
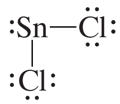
Next, think about what happens when the central atom is surrounded by four shared pairs of electrons. Examine the Lewis structure for methane, CH_4, shown below.



Notice that four single bonds surround the central carbon atom. On a flat plane the bonds are not as far apart as they can be. Instead, the four shared pairs are farthest apart when each pair of electrons is positioned at the corners of a tetrahedron, as shown in **Figure 17.** Only the electron clouds around the central atom are shown.

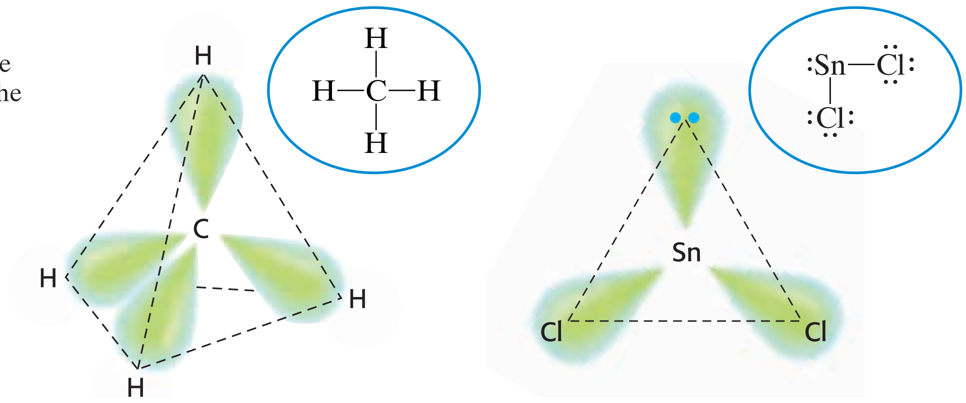
In CO_2, BF_3, and CH_4, all of the valence electrons of the central atom form shared pairs. What happens to the shape of a molecule if the central atom has an unshared pair? Tin(II) chloride, $SnCl_2$, gives an example. Examine the Lewis structure for $SnCl_2$, shown below.



Notice that the central tin atom has two shared pairs and one unshared pair of electrons. In VSEPR theory, unshared pairs occupy space around a central atom, just as shared pairs do. The two shared pairs and one unshared pair of the tin atom cause the shape of the $SnCl_2$ molecule to be bent, as shown in **Figure 17.**

The unshared pairs of electrons *influence* the shape of a molecule but are not visible in the space-fill model. For example, the shared and unshared pairs of electrons in $SnCl_2$ form a trigonal planar geometry, but the molecule has a bent shape. The bent shape of SO_2, shown in **Figure 15** on the previous page, is also due to unshared pairs. However, in the case of SO_2, there are two unshared pairs.

Figure 17
The electron clouds around the central atom help determine the shape of a molecule.



a A molecule whose central atom is surrounded by four shared pairs of electrons, such as CH_4, has a tetrahedral shape.

b A molecule whose central atom is surrounded by two shared pairs and one unshared pair, such as $SnCl_2$, has a bent shape.

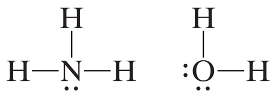
SAMPLE PROBLEM D

Predicting Molecular Shapes

Determine the shapes of NH_3 and H_2O.

1 Gather information.

Draw the Lewis structures for NH_3 and H_2O.



2 Count the shared and unshared pairs.

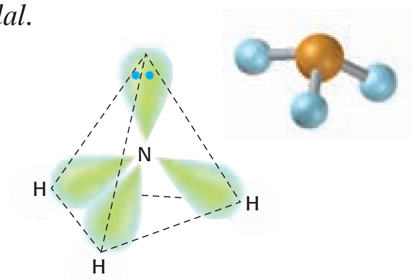
Count the number of shared and unshared pairs of electrons around each central atom.

NH_3 has three shared pairs and one unshared pair.

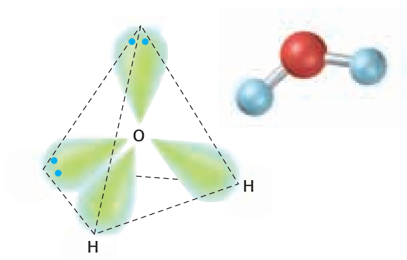
H_2O has two shared pairs and two unshared pairs.

3 Apply VSEPR theory.

- Use VSEPR theory to find the shape that allows the shared and unshared pairs of electrons to be spaced as far apart as possible.
- The ammonia molecule will have the shape of a pyramid. This geometry is called *trigonal pyramidal*.



- The water molecule will have a bent shape.



4 Verify the Structure.

For both molecules, be sure that all atoms, except hydrogen, obey the octet rule.

PRACTICE HINT

- Keep in mind that the geometry is difficult to show on the printed page because the atoms are arranged in three dimensions.
- If the sum of the shared and unshared pairs of electrons in each molecule is four, the electron pairs have *tetrahedral* geometry. However, the shape of the molecule is based on the number of shared pairs of electrons present. That is, the shape is based only on the position of the atoms and not on the position of the unshared pairs of electrons.

PRACTICE

Predict the shapes of the following molecules and polyatomic ions.

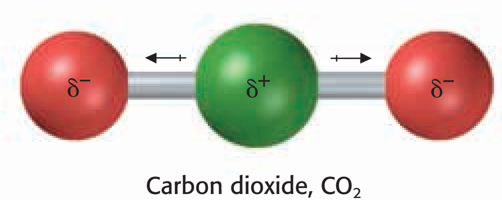
1 **a.** NH_2Cl **c.** NO_3^-

b. NOCl **d.** NH_4^+

PROBLEM SOLVING SKILL

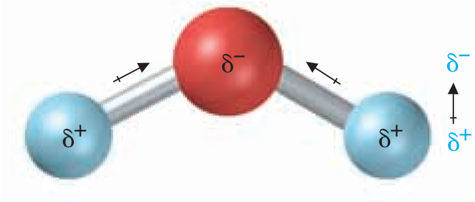
Figure 18
Molecules of both water and carbon dioxide have polar bonds. The symbol ⟶ shows a dipole.

a Because CO_2 is linear, the molecule is nonpolar.



Carbon dioxide, CO_2
(no molecular dipole)

b Because H_2O has a bent shape, the molecule is polar.



Water, H_2O
(overall molecular dipole)

Molecular Shape Affects a Substance's Properties

A molecule's shape affects both the physical and chemical properties of the substance. Recall that both sucrose and sucralose have a shape that allows each molecule to fit into certain nerve endings on the tongue and stimulate a sweet taste. If bending sucrose or sucralose molecules into a different shape were possible, the substances might not taste sweet. Shape determines many other properties. One property that shape determines is the polarity of a molecule.

Shape Affects Polarity

The polarity of a molecule that has more than two atoms depends on the polarity of each bond and the way the bonds are arranged in space. For example, compare CO_2 and H_2O. Oxygen has a higher electronegativity than carbon does, so each oxygen atom in CO_2 attracts electrons more strongly. Therefore, the shared pairs of electrons are more likely to be found near each oxygen atom than near the carbon atom. Thus, the double bonds between carbon and oxygen are polar. As shown in **Figure 18,** each oxygen atom has a partial negative charge, while the carbon atom has a partial positive charge.

Notice also that CO_2 has a linear shape. This shape determines the overall polarity of the molecule. The polarities of the double bonds extend from the carbon atom in opposite directions. As a result, they cancel each other and CO_2 is nonpolar overall even though the individual covalent bonds are polar.

Now think about H_2O. Oxygen has a higher electronegativity than hydrogen does, so oxygen attracts the shared pairs more strongly than either hydrogen does. As a result, each covalent bond between hydrogen and oxygen is polar. The O atom has a partial negative charge, while each H atom has a partial positive charge. Notice also that H_2O has a bent shape. Because the bonds are at an angle to each other, their polarities do not cancel each other. As a result, H_2O is polar.

You can think of a molecule's overall polarity in the same way that you think about forces on a cart. If you and a friend pull on a wheeled cart in equal and opposite directions—you pull the cart westward and your friend pulls the cart eastward—the cart does not move. The pull forces cancel each other in the same way that the polarities on the CO_2 molecule cancel each other. What happens if the two of you pull with equal force but in nonopposite directions? If you pull the cart northward and your friend pulls it westward, the cart moves toward the northwest. Because the cart has a net force applied to it, it moves. The water molecule has a net partial positive charge on the H side and a net negative charge on the O side. As a result, the molecule has an overall charge and is therefore polar.

Polarity Affects Properties

Because CO_2 molecules are nonpolar, the attractive force between them is very weak. In contrast, the attractive force between polar H_2O molecules is much stronger. The H atoms (with partial positive charges) attract the O atoms (with partial negative charges) on other water molecules. The attractive force between polar water molecules contributes to the greater amount of energy required to separate these polar molecules. The polarity of water molecules also adds to their attraction to positively and negatively charged objects. Other properties realted to polarity and molecular shape will be discussed in a later chapter.

3 Section Review

UNDERSTANDING KEY IDEAS

1. In VSEPR theory, what information about a central atom do you need in order to predict the shape of a molecule?

2. What is the only shape that a molecule made up of two atoms can have?

3. Explain how Lewis structures help predict the shape of a molecule.

4. Explain how a molecule that has polar bonds can be nonpolar.

5. Give one reason why water molecules are attracted to each other.

PRACTICE PROBLEMS

6. Determine the shapes of Br_2 and HBr. Which molecule is more polar and why?

7. Use VSEPR theory to determine the shapes of each of the following.

a. SCl_2

b. PF_3

c. NCl_3

d. NH_4^+

8. Predict the shape of the CCl_4 molecule. Is the molecule polar or nonpolar? Explain your answer.

CRITICAL THINKING

9. Can a molecule made up of three atoms have a linear shape? Explain your answer.

10. Why is knowing something about the shape of a molecule important?

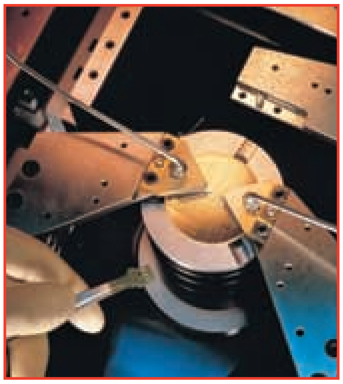
11. The electron pairs in a molecule of NH_3 form a tetrahedron. Why does the NH_3 molecule have a trigonal pyramidal shape rather than a tetrahedral shape?

14
Si
Silicon
28.0855
[Ne]$3s^2 3p^2$

SILICON

Element Spotlight

Where Is Si?
Earth's crust
27.72% by mass



Many integrated circuit chips can be made on the same silicon wafer. The wafer will be cut up into individual chips.

internet connect
www.scilinks.org
Topic: Silicon
SciLinks code: HW4116
SCILINKS. Maintained by the National Science Teachers Association

Silicon and Semiconductors

Silicon's most familiar use is in the production of microprocessor chips. Computer microprocessor chips are made from thin slices, or wafers, of a pure silicon crystal. The wafers are doped with elements such as boron, phosphorus, and arsenic to confer semiconducting properties on the silicon. A photographic process places patterns for several chips onto one wafer. Gaseous compounds of metals are allowed to diffuse into the open spots in the pattern, and then the pattern is removed. This process is repeated several times to build up complex microdevices on the surface of the wafer. When the wafer is finished and tested, it is cut into individual chips.

Industrial Uses

- Silicon and its compounds are used to add strength to alloys of aluminum, magnesium, copper, and other metals.
- When doped with elements of Group 13 or Group 15, silicon is a semiconductor. This property is important in the manufacture of computer chips and photovoltaic cells.
- Quartz (silicon dioxide) crystals are used for piezoelectric crystals for radio-frequency control oscillators and digital watches and clocks.

Real-World Connection Organic compounds containing silicon, carbon, chlorine, and hydrogen are used to make silicone polymers, which are used in water repellents, electrical insulation, hydraulic fluids, lubricants, and caulks.

A Brief History

1800 — **1900** — **2000**

1811: Joseph Louis Gay-Lussac and Louis Thenard prepare impure amorphous silicon from silicon tetrafluoride.

1824: Jöns Jacob Berzelius prepares pure amorphous silicon and is credited with the discovery of the element.

1854: Henri S. C. Deville prepares crystalline silicon.

1904: F. S. Kipping produces the first silicone compound.

1943: Commercial production of silicone rubber, oils, and greases begins in the United States.

1958: Jack Kilby and Robert Noyce produce the first integrated circuit on a silicon chip.

Questions

1. Research and identify five items that you encounter on a regular basis and that are constructed by using silicon.
2. Research piezoelectric materials, and identify how piezoelectric materials that contain silicon are used in science and industry.

Chapter Highlights 6

KEY TERMS / KEY IDEAS

SECTION ONE Covalent Bonds

Key terms: **covalent bond**, **molecular orbital**, **bond length**, **bond energy**, **nonpolar covalent bond**, **polar covalent bond**, **dipole**

- Covalent bonds form when atoms share pairs of electrons.
- Atoms have less potential energy and more stability after they form a covalent bond.
- The greater the electronegativity difference, the greater the polarity of the bond.
- The physical and chemical properties of a compound are related to the compound's bond type.

SECTION TWO Drawing and Naming Molecules

Key terms: **valence electron**, **Lewis structure**, **unshared pair**, **single bond**, **double bond**, **triple bond**, **resonance structure**

- In a Lewis structure, the element's symbol represents the atom's nucleus and inner-shell electrons, and dots represent the atom's valence electrons.
- Two atoms form single, double, and triple bonds depending on the number of electron pairs that the atoms share.
- Some molecules have more than one valid Lewis structure. These structures are called *resonance structures.*
- Molecular compounds are named using the elements' names, a system of prefixes, and *-ide* as the ending for the second element in the compound.

SECTION THREE Molecular Shapes

Key terms: **VSEPR theory**

- VSEPR theory states that electron pairs in the valence shell stay as far apart as possible.
- VSEPR theory can be used to predict the shape of a molecule.
- Molecular shapes predicted by VSEPR theory include linear, bent, trigonal planar, tetrahedral, and trigonal pyramidal.
- The shape of a molecule affects the molecule's physical and chemical properties.

KEY SKILLS

Drawing Lewis Structures with Single Bonds
Skills Toolkit 1 p. 201
Sample Problem A p. 202

Drawing Lewis Structures for Polyatomic Ions
Sample Problem B p. 203

Drawing Lewis Structures with Multiple Bonds
Sample Problem C p. 205

Predicting Molecular Shapes
Sample Problem D p. 211

6 CHAPTER REVIEW

USING KEY TERMS

1. How are bond length and potential energy related?
2. Describe the difference between a shared pair and an unshared pair of electrons.
3. How are the inner-shell electrons represented in a Lewis structure?
4. What term is used to describe the situation when two or more correct Lewis structures represent a molecule?
5. How is VSEPR theory useful?
6. Describe a molecular dipole.
7. Why is the electronegativity of an element important?
8. What type of bond results if two atoms share six electrons?
9. Contrast a polar covalent bond and a nonpolar covalent bond.
10. Describe how bond energy is related to the breaking of covalent bonds.

UNDERSTANDING KEY IDEAS

Covalent Bonds

11. How does a covalent bond differ from an ionic bond?
12. How are bond energy and bond strength related?
13. Why is a spring a better model than a stick for a covalent bond?
14. Describe the energy changes that take place when two atoms form a covalent bond.
15. Predict whether the bonds between the following pairs of elements are ionic, polar covalent, or nonpolar covalent.
    - **a.** Na—F
    - **b.** H—I
    - **c.** N—O
    - **d.** Al—O
    - **e.** S—O
    - **f.** H—H
16. Where are the bonding electrons between two atoms?
17. Arrange the following diatomic molecules in order of increasing bond polarity.
    - **a.** I—Cl
    - **b.** H—F
    - **c.** H—Br
18. What determines the electron distribution between two atoms in a bond?
19. Explain why the melting and boiling points of covalent compounds are usually lower than those of ionic compounds.

Drawing and Naming Molecules

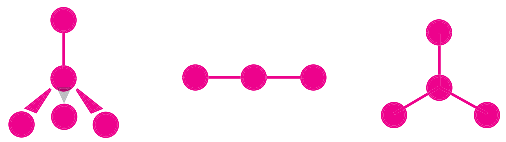
20. Draw the Lewis structures for boron, nitrogen, and phosphorus.
21. Describe a weakness of using Lewis structures to model covalent compounds.
22. What do the dots in a Lewis structure represent?
23. How does a Lewis structure show a bond between two atoms that share four electrons?
24. Why are resonance structures used to model certain molecules?

25. Name the following covalent compounds.
  a. SF_4
  b. XeF_4
  c. PBr_5
  d. N_2O_5
  e. Si_3N_4

Molecular Shapes

26. Why do electron pairs around a central atom stay as far apart as possible?

27. Name the following molecular shapes.



28. Two molecules have different shapes but the same composition. Can you conclude that they have the same physical and chemical properties? Explain you answer.

29. a. What causes H_2O to have a bent shape rather than a linear shape?
  b. How does this bent shape relate to the polarity of the water molecule?

PRACTICE PROBLEMS

PROBLEM SOLVING SKILL

Sample Problem A Drawing Lewis Structures with Single Bonds

30. Draw Lewis structures for the following molecules. Remember that hydrogen can form only a single bond.
  a. NF_3
  b. CH_3OH
  c. ClF
  d. CCl_2F_2
  e. $HOCl$

Sample Problem B Drawing Lewis Structures for Polyatomic Ions

31. Draw Lewis structures for the following polyatomic ions.
  a. OH^-
  b. O_2^{2-}
  c. NO^{2-}
  d. NO^{2+}
  e. AsO_4^{3-}

Sample Problem C Drawing Lewis Structures with Multiple Bonds

32. Draw Lewis structures for the following molecules.
  a. O_2
  b. CS_2
  c. N_2O

Sample Problem D Predicting Molecular Shapes

33. Determine the shapes of the following compounds.
  a. CF_4
  b. Cl_2O

34. Draw the shapes of the following polyatomic ions.
  a. NH_4^+
  b. OCl^-
  c. CO_3^{2-}

MIXED REVIEW

35. a. Determine the shapes of SCl_2, PF_3, and NCl_3.
  b. Which of these molecules has the greatest polarity?

36. Draw three resonance structures for NO_3^-.

37. Draw Lewis structures for the following polyatomic ions.
  a. CO_3^{2-}
  b. O_2^{2-}
  c. PO_4^{3-}

38. Name the following compounds, draw their Lewis structures, and determine their shapes.
  a. $SiCl_4$
  b. BCl_3
  c. NBr_3

39. How does an ionic compound differ from a molecular compound?

40. Explain why a halogen is unlikely to form a double bond with another element.

41. According to VSEPR theory, what molecular shapes are associated with the following types of molecules?

a. AB
b. AB_2
c. AB_3
d. AB_4

42. What types of atoms tend to form the following types of bonding?

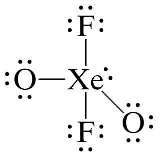
a. ionic
b. covalent
c. metallic

CRITICAL THINKING

43. What is the difference between a dipole and electronegativity difference?

44. Why does F generally form covalent bonds with great polarity?

45. Unlike other elements, noble gases are relatively inert. When noble gases do react, they do not follow the octet rule. Examine the following Lewis structure for the molecule XeO_2F_2.



a. Explain why the valence electrons of Xe do not follow the octet rule.
b. How many unshared pairs of electrons are in this molecule?
c. How many electrons make up all of the shared pairs in this molecule?

46. Ionic compounds tend to have higher boiling points than covalent substances do. Both ammonia, NH_3, and methane, CH_4, are covalent compunds, yet the boiling point of ammonia is 130°C higher than that of methane. What might account for this large difference?

47. Draw the Lewis structure of NH_4^+. Examine this structure to explain why this five-atom group exists only as a cation.

48. The length of a covalent bond varies depending on the type of bond formed. Triple bonds are generally shorter than double bonds, and double bonds are generally shorter than single bonds. Predict how the lengths of the C—C bond in the following molecules compare.

a. C_2H_6
b. C_2H_4
c. C_2H_2

ALTERNATIVE ASSESSMENT

49. Natural rubber consists of long chains of carbon and hydrogen atoms covalently bonded together. When Goodyear accidentally dropped a mixture of sulfur and rubber on a hot stove, the energy joined these chains together to make vulcanized rubber. Vulcan was the Roman god of fire. The carbon-hydrogen chains in vulcanized rubber are held together by two sulfur atoms that form covalent bonds between the chains. These covalent bonds are commonly called *disulfide bridges.* Explore other molecules that have such disulfide bridges. Present your findings to the class.

50. Devise a set of criteria that will allow you to classify the following substances as covalent, ionic, or metallic: $CaCO_3$, Cu, H_2O, NaBr, and C (graphite). Show your criteria to your teacher.

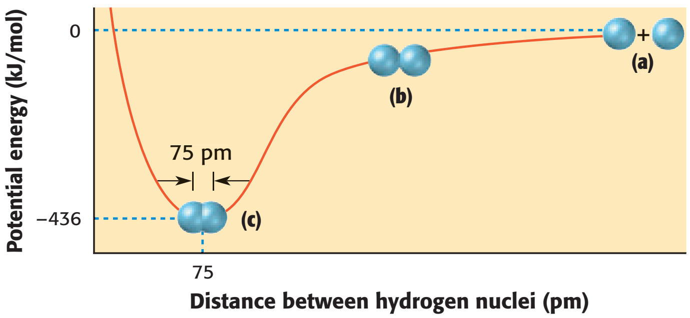
CONCEPT MAPPING

51. Use the following terms to create a concept map: *valence electrons, nonpolar, covalent compounds, polar, dipoles,* and *Lewis structures.*

FOCUS ON GRAPHING

Study the graph below, and answer the questions that follow.
For help in interpreting graphs, see Appendix B, "Study Skills for Chemistry."

Potential Energy Curve for H_2



52. What do the blue spheres represent on this graph?

53. What are the coordinates of the minimum (the lowest point) of the graph?

54. What relationship does the graph describe?

55. What is significant about the distance between the hydrogen nuclei at the lowest point on the graph?

56. When the distance between the hydrogen nuclei is greater than 75 pm is the slope positive or is it negative?

57. Miles measures the energy required to hold two magnets apart at varying distances. He notices that it takes less and less energy to hold the magnets apart as the distance between them increases. Compare the results of Miles's experiment with the data given in the graph.

TECHNOLOGY AND LEARNING

58. Graphing Calculator

Classifing Bonding Type

The graphing calculator can run a program that classifies bonding between atoms according to the difference between the atoms' electronegativities. Use this program to determine the electronegativity difference between bonded atoms and to classify bonding type.

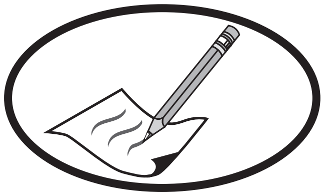
Go to Appendix C. If you are using a TI-83 Plus, you can download the program BONDTYPE and data sets and run the application as directed. If you are using another calculator, your teacher will provide you with the keystrokes and data sets to use. After you have graphed the data sets, answer the questions below.

a. Which element pair or pairs have a pure covalent bond?

b. What type of bond does the pair H and O have?

c. What type of bond does the pair Ca and O have?

6 STANDARDIZED TEST PREP



UNDERSTANDING CONCEPTS

***Directions (1–3):* For *each* question, write on a separate sheet of paper the letter of the correct answer.**

1. Which of these combinations is likely to have a polar covalent bond?
   A. two atoms of similar size
   B. two atoms of very different size
   C. two atoms with different electronegativities
   D. two atoms with the same number of electrons

2. According to VSEPR theory, which of these is caused by repulsion between electron pairs surrounding an atom?
   F. breaking of a chemical bond
   G. formation of a sea of electrons
   H. formation of a covalent chemical bond
   I. separation of electron pairs as much as possible

3. How many electrons are shared in a double covalent bond?
   A. 2　　**C.** 6
   B. 4　　**D.** 8

***Directions (4–6):* For *each* question, write a short response.**

4. How can the difference in number of valence electrons between nitrogen and carbon account for the fact that the boiling point of ammonia, NH_3, is 130°C higher than that of methane, CH_4.

5. Why don't scientists need VSEPR theory to predict the shape of HCl?

6. What are the attractive and repulsive forces involved in a covalent bond and how do their total strengths compare?

READING SKILLS

***Directions (7–9):* Read the passage below. Then answer the questions.**

Although water is a polar molecule, pure water does not carry an electric current. It is a good solvent for many ionic compounds, and solutions of ionic compounds in water do carry electric currents. The charged particles in solution move freely, carrying electric charges. Even a dilute solution of ions in water becomes a good conductor. Without ions in solution, there is very little electrical conductivity.

7. Why is a solution of sugar in water not a good electrical conductor?
   F. Sugar does not form ions in solution.
   G. The ionic bonds of sugar molecules are too strong to carry a current.
   H. Not enough sugar dissolves for the solution to become a conductor.
   I. A solution of sugar in water is not very conductive because it is mostly water, which is not very conductive.

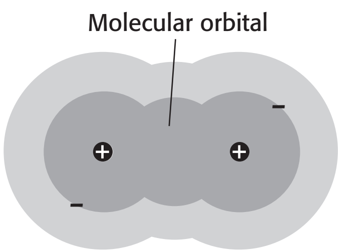
8. Why do molten ionic compounds generally conduct electric current well, while molten covalent compounds generally do not?
   A. Ionic compounds are more soluble in water.
   B. Ionic compounds have more electrons than compounds.
   C. When they melt, ionic compounds separate into charged particles.
   D. Most ionic compounds contain a metal atom which carries the electric current.

9. If water is not a good conductor of electric current, why is it dangerous to handle an electrical appliance when your hands are wet or when you are standing on wet ground?

INTERPRETING GRAPHICS

***Directions (10–13):* For *each* question below, record the correct answer on a separate sheet of paper.**

Use the diagram below to answer question 10.



10 The diagram above **best** represents which type of chemical bond?

F. ionic
G. metallic
H. nonpolar covalent
I. polar covalent

The table below shows the connection between electronegativity and bond strength (kilojoules per mole). Use it to answer questions 11 through 13.

Electronegativity Difference for Hydrogen Halides

Molecule	Electronegativity difference	Bond energy
H—F	1.8	570 kJ/mol
H—Cl	1.0	432 kJ/mol
H—Br	0.8	366 kJ/mol
H—I	0.5	298 kJ/mol

11 Which of these molecules has the smallest partial positive charge on the hydrogen end of the molecule?

A. HF
B. HCl
C. HBr
D. HI

12 How does the polarity of the bond between a halogen and hydrogen relate to the number of electrons of the halogen atom?

F. Polarity is not related to the number of electrons of the halogen atom.
G. Polarity decreases as the number of unpaired halogen electrons increases.
H. Polarity decreases as the total number of halogen atom electrons increases.
I. Polarity decreases as the number of valence electrons of the halogen atom increases.

13 Based on the information in this table, how does the electronegativity difference in a covalent bond relate to the strength of the bond?

Test TIP

Take time to read each question completely on a standardized test, including all of the answer choices. Consider each answer choice before determining which one is correct.

CHAPTER

7

The Mole and Chemical Composition