8.4 Polar Bonds and Molecules

Connecting to Your World Snow covers approximately 23 percent of Earth's surface. Each individual snowflake is formed from as many as 100 snow crystals. The size and shape of each crystal depends mainly on the air temperature and amount of water vapor in the air at the time the snow forms. In this section, you will see that the polar bonds in water molecules influence the distinctive geometry of snowflakes.

Bond Polarity

Covalent bonds involve electron sharing between atoms. However, covalent bonds differ in terms of how the bonded atoms share the electrons. The character of the bonds in a given molecule depends on the kind and number of atoms joined together. These features, in turn, determine the molecular properties.

The bonding pairs of electrons in covalent bonds are pulled, as in the tug-of-war in Figure 8.22, between the nuclei of the atoms sharing the electrons. When the atoms in the bond pull equally (as occurs when identical atoms are bonded), the bonding electrons are shared equally, and the bond is a **nonpolar covalent bond.** Molecules of hydrogen (H_2), oxygen (O_2), and nitrogen (N_2) have nonpolar covalent bonds. Diatomic halogen molecules, such as Cl_2, are also nonpolar.

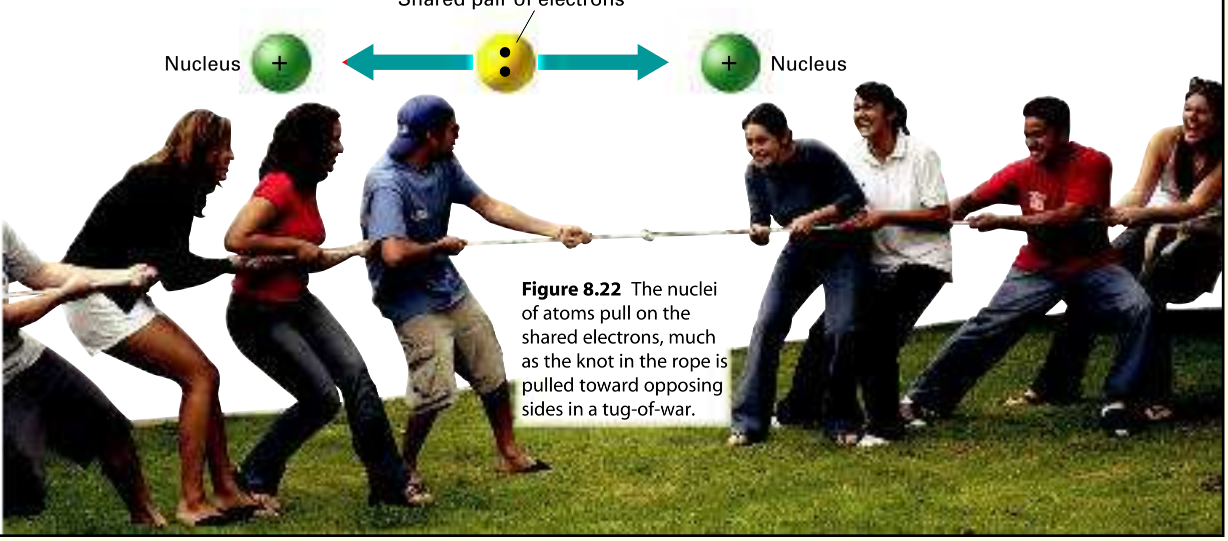


Figure 8.22 The nuclei of atoms pull on the shared electrons, much as the knot in the rope is pulled toward opposing sides in a tug-of-war.

Guide for Reading

Key Concepts

- How do electronegativity values determine the charge distribution in a polar bond?
- What happens to polar molecules between a pair of oppositely charged metal plates?
- How do intermolecular attractions compare with ionic and covalent bonds?
- Why do network solids have high melting points?

Vocabulary

nonpolar covalent bond
polar covalent bond
polar bond
polar molecule
dipole
van der Waals forces
dipole interactions
dispersion forces
hydrogen bonds
network solids

Reading Strategy

Relating Text and Visuals As you read, look closely at Figures 8.22 and 8.23. Explain how these illustrations help you understand how molecules attract each other.

8.4

1 FOCUS

Objectives

8.4.1 **Describe** how electronegativity values determine the distribution of charge in a polar molecule.

8.4.2 **Describe** what happens to polar molecules when they are placed between oppositely charged metal plates.

8.4.3 **Evaluate** the strength of intermolecular attractions compared with the strength of ionic and covalent bonds.

8.4.4 **Identify** the reason why network solids have high melting points.

Guide for Reading

Build Vocabulary L2

Word Parts Polar, nonpolar, and dipole all have the same Latin stem *polus* meaning pole. A pole is part of a system (large or small) that has opposite electric or magnetic positions. Thus, Earth has magnetic north and south poles. Batteries have positive and negative poles or terminals.

Reading Strategy L2

Predict Having reviewed the root of the words polar, nonpolar, and dipole, ask students to predict what is discussed under the three red headings in this section. Have them check their predictions after reading.

2 INSTRUCT

Connecting to Your World

Have students read the opening paragraph. Ask, **What kind of bonds are in the water molecule?** *(polar)* **What is the shape of the snowflake?** *(pentagon)*

Section Resources

Print

- **Guided Reading and Study Workbook,** Section 8.4
- **Core Teaching Resources,** Section 8.4 Review
- **Transparency,** T93

Technology

- **Interactive Textbook with ChemASAP,** Animation 10, Simulation 8, Assessment 8.4
- **Go Online,** Section 8.4

Bond Polarity

Discuss L2

Ask, **What does the term *electronegativity* mean?** *(It describes the attraction an atom has for electrons when the atom is in a compound.)* Remind students that electronegativities are calculated values from 0.0 to 4.0 based on the properties of the elements. Ask, **Which element is the most electronegative? Which is the least?** *(Fluorine has an electronegativity value of 4.0; cesium, the element that has the lowest attraction for electrons, has a value of 0.7.)*

Discuss L2

When students use Table 8.3 to solve Conceptual Problem 8.3, they may be surprised to discover that bonds previously labeled ionic are now categorized as polar covalent. The reality is that most bonds are neither totally ionic nor totally covalent. Most fall on a continuum from ionic (complete transfer of electrons) to covalent (equal sharing of electrons). In between are polar-covalent bonds in which electrons are displaced more or less toward one atom or the other. Make sure that students understand the two uses of the Greek letter delta. A lowercase delta (δ) is used to denote partial charges; an uppercase delta (Δ) is used to denote a change in a variable.

CLASS Activity

A Magnetic Analogy L1

Purpose Students experience the attraction of opposite poles of a magnet as an experience similar to the attraction of opposite electric charges.

Materials magnets, iron filings, cardboard

Procedure Place some iron filings on a cardboard. Hold a magnet under the cardboard and move it to create patterns. Allow students to handle magnets and feel the pull when north and south poles are close. Have them write a sentence relating the experience to polar molecules.

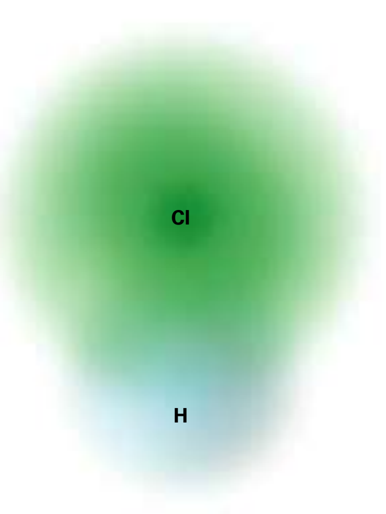


Figure 8.23 This electron-cloud picture of hydrogen chloride shows that the chlorine atom attracts the electron cloud more than the hydrogen atom does. **Inferring** ***Which atom is more electronegative, a chlorine atom or a hydrogen atom?***

A **polar covalent bond,** known also as a **polar bond,** is a covalent bond between atoms in which the electrons are shared unequally. **The more electronegative atom attracts electrons more strongly and gains a slightly negative charge. The less electronegative atom has a slightly positive charge.** Refer back to Table 6.2 in Chapter 6 to see the electronegativities of some common elements. The higher the electronegativity value, the greater the ability of an atom to attract electrons to itself.

Consider the hydrogen chloride molecule (HCl) shown in Figure 8.23. Hydrogen has an electronegativity of 2.1, and chlorine has an electronegativity of 3.0. These values are significantly different, so the covalent bond in hydrogen chloride is polar. The chlorine atom acquires a slightly negative charge. The hydrogen atom acquires a slightly positive charge. The lowercase Greek letter delta (δ) denotes that atoms involved in the covalent bond acquire only partial charges, less than 1+ or 1−.

$$\overset{\delta+}{H}-\overset{\delta-}{Cl}$$

The minus sign in this notation shows that chlorine has acquired a slightly negative charge. The plus sign shows that hydrogen has acquired a slightly positive charge. These partial charges are shown as clouds of electron density. The polar nature of the bond may also be represented by an arrow pointing to the more electronegative atom, as shown here.

$$\overset{+\!\longrightarrow}{H-Cl}$$

The O—H bonds in the water molecule are also polar. The highly electronegative oxygen partially pulls the bonding electrons away from hydrogen. The oxygen acquires a slightly negative charge. The hydrogen is left with a slightly positive charge.

$$\overset{\delta-\,\delta-}{O}\text{ bonded to } {}_{\delta+}H \text{ and } H_{\delta+} \quad \textit{or} \quad H \nearrow O \nwarrow H$$

As shown in Table 8.3, the electronegativity difference between two atoms tells you what kind of bond is likely to form. Remember when you use the table that there is no sharp boundary between ionic and covalent bonds. As the electronegativity difference between two atoms increases, the polarity of the bond increases. If the electronegativity difference is greater than 2.0, it is very likely that electrons will be pulled away completely by one of the atoms. In that case, an ionic bond will form.

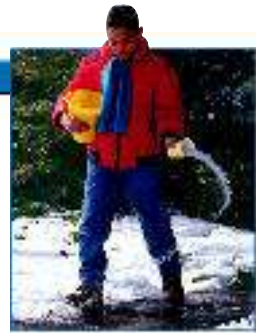
Table 8.3

Electronegativity Differences and Bond Types

Electronegativity difference range	Most probable type of bond	Example
0.0–0.4	Nonpolar covalent	H—H (0.0)
0.4–1.0	Moderately polar covalent	$\overset{\delta+}{H}-\overset{\delta-}{Cl}$ (0.9)
1.0–2.0	Very polar covalent	$\overset{\delta+}{H}-\overset{\delta-}{F}$ (1.9)
≥ 2.0	Ionic	Na^+Cl^- (2.1)

238 Chapter 8

CONCEPTUAL PROBLEM 8.3



Identifying Bond Type

Which type of bond (nonpolar covalent, moderately-polar covalent, or ionic) will form between each of the following pairs of atoms?

a. N and H **b.** F and F **c.** Ca and Cl **d.** Al and Cl

1 Analyze ***Identify the relevant concepts.***

In each case, the pairs of atoms involved in the bonding pair are given. The types of bonds depend on the electronegativity differences between the bonding elements. Use Table 6.2 to find the electronegativity difference, then use Table 8.3 to determine the bond type.

2 Solve ***Apply concepts to this situation.***

From Tables 6.2 and 8.3, the electronegativities, their differences, and the corresponding bond types are as follows.

a. N (3.0), H (2.1); 0.9 moderately polar covalent
b. F (4.0), F (4.0); 0.0; nonpolar covalent
c. Ca (1.0), Cl (3.0); 2.0; ionic
d. Al (1.5), Cl (3.0); 1.5; very polar covalent

Practice Problems

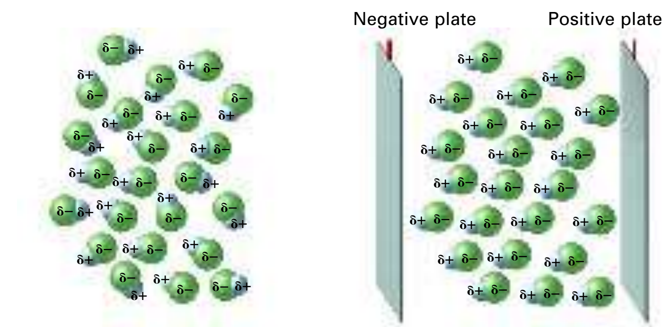
30. Identify the bonds between atoms of each pair of elements as nonpolar covalent, moderately polar covalent, very covalent, or ionic.
a. H and Br **b.** K and Cl **c.** C and O
d. Cl and F **e.** Li and O **f.** Br and Br

31. Place the following covalent bonds in order from least to most polar.
a. H—Cl **b.** H—Br
c. H—S **d.** H—C

Polar Molecules

The presence of a polar bond in a molecule often makes the entire molecule polar. In a **polar molecule,** one end of the molecule is slightly negative and the other end is slightly positive. In the hydrogen chloride molecule, for example, the partial charges on the hydrogen and chlorine atoms are electrically charged regions or poles. A molecule that has two poles is called a dipolar molecule, or **dipole.** The hydrogen chloride molecule is a dipole. Look at Figure 8.24. **When polar molecules are placed between oppositely charged plates, they tend to become oriented with respect to the positive and negative plates.**

Interactive Textbook
Animation 10 Learn to distinguish between polar and nonpolar molecules.
with **ChemASAP**



Electric field absent.
Polar molecules orient randomly.

Electric field on.
Polar molecules line up.

Figure 8.24 When polar molecules, such as HCl, are placed in an electric field, the slightly negative ends of the molecules become oriented toward the positively charged plate and the slightly positive ends of the molecules become oriented toward the negatively charged plate. **Predicting** ***What would happen if, instead, carbon dioxide molecules were placed between the plates? Why?***

CONCEPTUAL PROBLEM 8.3

Answers

30. a. moderately polar covalent **b.** ionic **c.** moderately to very polar covalent **d.** moderately to very polar covalent **e.** ionic **f.** nonpolar covalent

31. c and d (tie), b, a

Practice Problems Plus L2

Which type of bond (nonpolar, covalent, moderately-covalent, or ionic) will form between each of the following pairs of ions?

a. C and H *(nonpolar covalent)*
b. Be and F *(ionic)*
c. Si and Cl *(very polar covalent)*

Polar Molecules

Discuss L2

Draw diagrams of molecules on the board or overhead projector. Use electronegativity values to determine which element has the partial positive charge and which has the partial negative charge. Explain how the attraction of electrons can be described with a vector. The shape of the molecule determines the direction of the vector. If the sum of the vectors acting on the central atom in one direction is cancelled by the sum of the vectors acting in the other direction, the molecule has no dipole. If they don't cancel, the molecule is polar.

Answers to...
Figure 8.23 chlorine
Figure 8.24 Carbon dioxide molecules would not change orientation because the CO_2 molecule is nonpolar.

TEACHER Demo

Observing Evidence of Polarity L2

Purpose Students observe what happens to polar molecules when they are placed near positively charged materials

Materials 2 burets, 2 beakers, ring stand, buret clamps, turpentine, water, rubber rod, glass rod, fur cloth

Safety Wear safety goggles and gloves.

Procedure Attach two burets to a ring stand. Fill one with water and the other with turpentine. Place a beaker under each. **CAUTION!** *Turpentine is an irritant.* Rub a rubber rod with a fur cloth to charge it negatively. Bring the rod near the open end of the water-containing buret as you slowly open the stopcock. Have students observe the behavior of the water. Repeat the procedure using turpentine. Then repeat using a charged glass rod for both liquids.

Expected Outcomes The polar water molecules are attracted to both negatively and positively charged materials, but the nonpolar turpentine molecules are not attracted to either.

Attractions Between Molecules

Discuss L2

The sharing of electrons binds atoms into molecules. If no other forces were present, all covalently bonded molecules would be gases at any temperature. No attractive forces would tend to pull them together to form a liquid or solid. However, forces do exist between covalently bonded molecules. These include van der Waals forces and hydrogen bonding.

Go Online SCI LINKS
For: Links on Intermolecular Forces
Visit: www.SciLinks.org
Web Code: cdn-1084

The effect of polar bonds on the polarity of an entire molecule depends on the shape of the molecule and the orientation of the polar bonds. A carbon dioxide molecule, for example, has two polar bonds and is linear.

$$\overset{\longleftarrow+}{\mathrm{O}{=}\mathrm{C}}\overset{+\longrightarrow}{{=}\mathrm{O}}$$

Note that the carbon and oxygens lie along the same axis. Therefore, the bond polarities cancel because they are in opposite directions. Carbon dioxide is thus a nonpolar molecule, despite the presence of two polar bonds.

The water molecule also has two polar bonds. However, the water molecule is bent rather than linear. Therefore, the bond polarities do not cancel and a water molecule is polar.

Attractions Between Molecules

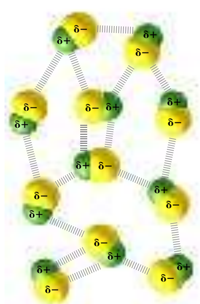
Molecules can attract each other by a variety of forces. **Intermolecular attractions are weaker than either ionic or covalent bonds.** Nevertheless, you should not underestimate the importance of these forces. Among other things, these attractions are responsible for determining whether a molecular compound is a gas, a liquid, or a solid at a given temperature.

Van der Waals Forces The two weakest attractions between molecules are collectively called **van der Waals forces,** named after the Dutch chemist Johannes van der Waals (1837–1923). Van der Waals forces consist of dipole interactions and dispersion forces.

Dipole interactions occur when polar molecules are attracted to one another. The electrical attraction involved occurs between the oppositely charged regions of polar molecules, as shown in Figure 8.25. The slightly negative region of a polar molecule is weakly attracted to the slightly positive region of another polar molecule. Dipole interactions are similar to but much weaker than ionic bonds.

Dispersion forces, the weakest of all molecular interactions, are caused by the motion of electrons. They occur even between non-polar molecules. When the moving electrons happen to be momentarily more on the side of a molecule closest to a neighboring molecule, their electric force influences the neighboring molecule's electrons to be momentarily more on the opposite side. This causes an attraction between the two molecules similar to, but much weaker than, the force between permanently polar molecules. The strength of dispersion forces generally increases as the number of electrons in a molecule increases. The halogen diatomic molecules, for example, attract each other mainly by means of dispersion forces. Fluorine and chlorine have relatively few electrons and are gases at ordinary room temperature and pressure because of their especially weak dispersion forces. The larger number of electrons in bromine generates larger dispersion forces. Bromine molecules therefore attract each other sufficiently to make bromine a liquid at ordinary room temperature and pressure. Iodine, with a still larger number of electrons, is a solid at ordinary room temperature and pressure.

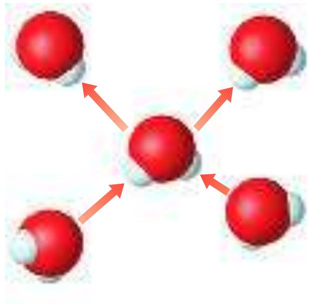
Figure 8.25 Polar molecules are attracted to one another by dipole interactions, a type of van der Waals force.



Facts and Figures

Molecular Modeling

The widespread availability of fast computers has greatly benefited chemists interested in the shapes of molecules. Molecular modeling computer programs are becoming an important tool in the pharmaceutical industry for designing drugs. These programs create models of molecules composed of thousands of atoms. Using a special viewing apparatus, chemists can see three-dimensional representations of complex molecules. Models can be rotated, shrunk, and cut in pieces. Through computer technology called virtual reality, a chemist can virtually get inside a molecular structure to examine it from every possible angle.



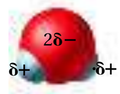


Figure 8.26 The strong hydrogen bonding between water molecules accounts for many properties of water, such as the fact that water is a liquid rather than a gas at room temperature.

Hydrogen Bonds The dipole interactions in water produce an attraction between water molecules. Each O—H bond in the water molecule is highly polar, and the oxygen acquires a slightly negative charge because of its greater electronegativity. The hydrogens in water molecules acquire a slightly positive charge. The positive region of one water molecule attracts the negative region of another water molecule, as illustrated in Figure 8.26. This attraction between the hydrogen of one water molecule and the oxygen of another water molecule is strong compared to other dipole interactions. This relatively strong attraction, which is also found in hydrogen-containing molecules other than water, is called a hydrogen bond. Figure 8.26 illustrates hydrogen bonding in water.

Hydrogen bonds are attractive forces in which a hydrogen covalently bonded to a very electronegative atom is also weakly bonded to an unshared electron pair of another electronegative atom. This other atom may be in the same molecule or in a nearby molecule. Hydrogen bonding always involves hydrogen. It is the only chemically reactive element with valence electrons that are not shielded from the nucleus by other electrons.

Remember that for a hydrogen bond to form, a covalent bond must already exist between a hydrogen atom and a highly electronegative atom, such as oxygen, nitrogen, or fluorine. The combination of this strongly polar bond and the lack of shielding effect in a hydrogen atom is responsible for the relative strength of hydrogen bonds. A hydrogen bond has about 5% of the strength of an average covalent bond. Hydrogen bonds are the strongest of the intermolecular forces. They are extremely important in determining the properties of water and biological molecules such as proteins. Figure 8.27 shows how the relatively strong attractive forces between water molecules cause the water to form small drops on a waxy surface.

Checkpoint ***What are hydrogen bonds?***

Figure 8.27 The strong attractions between water molecules cause the water to pull together into small drops rather than spread over the surface of the flower.

Discuss

L2

Construct a concept map using the following terms: intermolecular attractions, van der Waals forces, dispersion forces, dipole interactions, and hydrogen bonding. Use the concept map to introduce and discuss each force in detail. Use the halogen family to demonstrate dispersion forces, carbon dioxide to discuss dipole interactions, and water to discuss hydrogen bonding. Ask, **Why does He have a lower boiling point than Rn?** *(Rn has more electrons, so its intermolecular forces are stronger than those of He.)* Point out that hydrogen bonds only occur between molecules in which hydrogen is bonded to a strongly electronegative atom such as nitrogen, oxygen, or fluorine. Because living systems contain molecules rich in oxygen and nitrogen, hydrogen bonding plays an important role in the chemistry of living systems. Hydrogen bonding explains many of the physical properties of water.

Go Online NSTA SCILINKS

Download a worksheet on **Intermolecular Forces** for students to complete, and find additional teacher support from NSTA SciLinks.

Facts and Figures

Molecular Modeling

The Dutch physicist Johannes van der Waals (1837–1923) surmised that real gas behavior could not be described exactly by the ideal gas law because the law assumes that gas molecules have no volume and that no attractive forces exist between gas molecules. Van der Waals adjusted the ideal gas equation by introducing two factors: one to take cognizance of the real volume of gas molecules; the other to account for intermolecular forces between molecules. For his work in developing an understanding of intermolecular attractions, he was awarded the Nobel Prize in Physics in 1910, and van der Waals forces, the weak attractive forces he had been studying, were named in his honor.

Answers to...

Checkpoint a hydrogen bond is the attraction between a hydrogen atom bonded to a very electronegative atom and an unshared electron pair of another electronegative atom.

Technology & Society

Make sure that students are familiar with polymerization. Ask, **What is a monomer?** *(a single small molecule)* **What is a polymer?** *(a large molecule made up of many repeated monomer units)* **In an adhesive, what causes the monomer to form a polymer?** *(moisture in the air or on the surfaces to be joined together)* **Why does the adhesive not polymerize in the tube?** *(A stabilizer is added to the monomer that prevents polymerization until the monomer comes in contact with water.)*

Students can use library and Internet resources to find out more about any of the following topics:

- the adhesives used in ancient times, as long ago as 3300 years ago,
- natural adhesives from plants and animals including simple flour and water paste, animal glue, casein glue, and blood albumin glue,
- thermoplastic adhesives that respond to heating and cooling,
- the structure of epoxy resins,
- remoistenable adhesives used on stamps and envelopes.

Students may also want to investigate the adsorption theory of adhesion. What is the mechanism by which substances stick to each other? In theory, all substances can be categorized as adhesives if they undergo intermolecular attractions or bonding with other contacting substances.

Answers to...

Interpreting Diagrams A stabilizer stops the monomers from forming a solid polymer.

Technology & Society

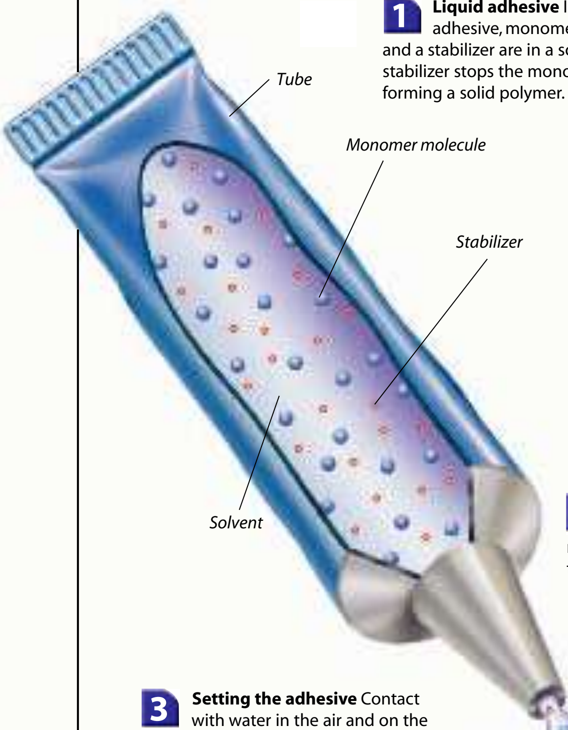
The Chemistry of Adhesives

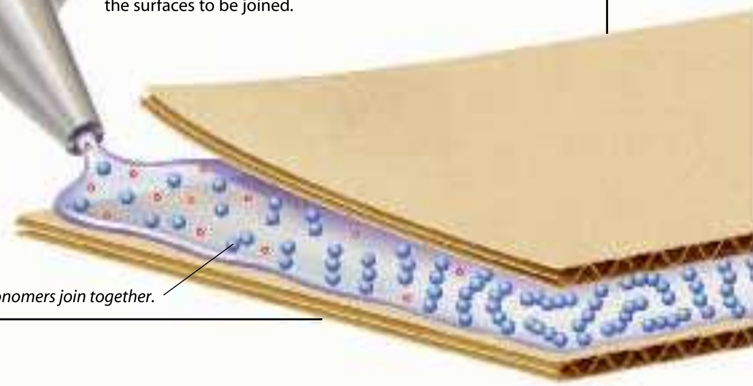
Most common adhesives are sticky substances that you can use to bind two surfaces together. The diagram shows how some adhesives work. The adhesive is a liquid until the surfaces are in position. Then the adhesive sets into a solid. Adhesion can work in three ways: Molecules of the polymer may fill crevices in the surfaces being connected, the molecules may also become attracted by intermolecular forces, or they may react by forming covalent bonds.

Interpreting Diagrams ***Explain the purpose of a stabilizer in an adhesive.***

Permanent adhesion An epoxy resin attached this full-size automobile to the billboard. Epoxy resins are often stored in two parts that are mixed just before the epoxy is used. Strong binding forces in these adhesives make them heat- and water-resistant.

Temporary adhesion The sticky strip on a reusable note is a microsphere adhesive. The microspheres limit the amount of surface area contact, so the paper sticks lightly and repositions easily.





1 Liquid adhesive In a typical adhesive, monomer molecules and a stabilizer are in a solvent. The stabilizer stops the monomers from forming a solid polymer.

2 Applying the adhesive Squeezing the tube releases the liquid onto one of the surfaces to be joined.

3 Setting the adhesive Contact with water in the air and on the surfaces being joined makes the stabilizer inactive. The monomers then begin to join together to form polymers. As the chain lengthens, the adhesive changes from a liquid to a solid.

Intermolecular Attractions and Molecular Properties

At room temperature, some compounds are gases, some are liquids, and some are solids. The physical properties of a compound depend on the type of bonding it displays—in particular, on whether it is ionic or covalent. A great range of physical properties occurs among covalent compounds. This is mainly because of widely varying intermolecular attractions.

The melting and boiling points of most compounds composed of molecules are low compared with those of ionic compounds. In most solids formed from molecules, only the weak attractions between molecules need to be broken. However, a few solids that consist of molecules do not melt until the temperature reaches 1000°C or higher, or they decompose without melting at all. Most of these very stable substances are **network solids** (or network crystals), solids in which all of the atoms are covalently bonded to each other. **Melting a network solid would require breaking covalent bonds throughout the solid.**

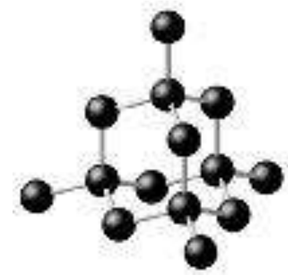
Diamond is an example of a network solid. As shown in Figure 8.28, each carbon atom in a diamond is covalently bonded to four other carbons, interconnecting carbon atoms throughout the diamond. Cutting a diamond requires breaking a multitude of these bonds. Diamond does not melt; rather, it vaporizes to a gas at 3500°C and above.

Silicon carbide, with the formula SiC and a melting point of about 2700°C, is also a network solid. Silicon carbide is so hard that it is used in grindstones and as an abrasive as illustrated in Figure 8.29. The molecular structures of silicon carbide and diamond are similar to each other. You can think of samples of diamond, silicon carbide, and other network solids as single molecules.

Checkpoint ***What substance is an example of a network solid?***

Figure 8.28 Diamond is a network-solid form of carbon. Diamond has a three-dimensional structure, with each carbon at the center of a tetrahedron.

Diamond



Go Online NSTA SciLinks
For: Links on Diamond
Visit: www.SciLinks.org
Web Code: cdn-1085

Figure 8.29 Silicon carbide, a network solid, is so hard that it is used in this grindstone to wear down the end of a hardened steel cutting tool to form a sharp edge.

Interactive Textbook
Simulation 8 Relate melting and boiling points to the strength of intermolecular forces.
with ChemASAP

Intermolecular Attractions and Molecular Properties

Go Online NSTA SciLinks
Download a worksheet on **Diamond** for students to complete, and find additional teacher support from NSTA SciLinks.

TEACHER Demo

Evidence of Hydrogen Bonding L2

Purpose Students see how hydrogen bonding affects the surface tension and viscosity of fluids.

Materials Four 300-mL stoppered flasks, each containing 100 mL of one of these liquids: glycerol, ethylene glycol, water, ethanol

Procedure Gently rotate one of the flasks so that its contents begin to swirl. Have students measure the time interval between your ceasing to rotate the bottle and the disappearance of the swirling vortex. Repeat the procedure for each fluid. Be sure to rotate each flask the same number of times.

Expected Outcome From longest time to shortest: ethanol, water, ethylene glycol, and glycerol. Point out that hydrogen bonding increases a liquid's surface tension and viscosity. Thus, liquids with a relatively high degree of hydrogen bonding generally stop swirling more quickly. Have students order the fluids according to increasing hydrogen bonding. *(ethanol, water, ethylene glycol, glycerol)* Show the structural formulas of the compounds to demonstrate the origins of the hydrogen bonding.

Answers to...

Checkpoint diamond

3 ASSESS

Evaluate Understanding L2

Have students account for the generally low melting points of covalent compounds in terms of bonding. *(Intermolecular attractions between covalent compounds are generally not as strong as ionic bonds.)*

Reteach L1

Help students list the types of intermolecular attractions that operate between molecules. Give examples of compounds that would fit into each category. Ask students to name other examples.

Connecting Concepts

In dipole interactions, the slightly positive end of a polar molecule attracts the slightly negative end of another polar molecule. Dispersion forces arise from the movement of electrons in any molecule, polar or nonpolar. If the electrons in one molecule happen to be momentarily more on one side, the electrons in a neighboring molecule can be repulsed. Weak attractions result between the temporary dipoles.

Interactive Textbook

If your class subscribes to the Interactive Textbook, use it to review key concepts in Section 8.4.

with **ChemASAP**

Table 8.4 summarizes some of the characteristic differences between ionic and covalent (molecular) substances. Note that ionic compounds have higher melting points than molecular compounds. Ionic compounds also tend to be soluble in water.

Table 8.4

Characteristics of Ionic and Covalent Compounds

Characteristic	Ionic compound	Covalent compound
Representative unit	Formula unit	Molecule
Bond formation	Transfer of one or more electrons between atoms	Sharing of electron pairs between atoms
Type of elements	Metallic and nonmetallic	Nonmetallic
Physical state	Solid	Solid, liquid, or gas
Melting point	High (usually above 300°C)	Low (usually below 300°C)
Solubility in water	Usually high	High to low
Electrical conductivity of aqueous solution	Good conductor	Poor to nonconducting

8.4 Section Assessment

32. Key Concept How do electronegativity values determine the charge distribution in a polar covalent bond?

33. Key Concept What happens when polar molecules are between oppositely charged metal plates?

34. Key Concept Compare the strengths of intermolecular attractions to the strengths of ionic bonds and covalent bonds.

35. Key Concept Explain why network solids have high melting points.

36. Not every molecule with polar bonds is polar. Explain this statement. Use CCl_4 as an example.

37. Draw the electron dot structure for each molecule. Identify polar covalent bonds by assigning slightly positive ($\delta+$) and slightly negative ($\delta-$) symbols to the appropriate atoms.

a. HOOH **b.** BrCl **c.** HBr **d.** H_2O

38. How does a network solid differ from most other covalent compounds?

Connecting Concepts

Dipole Interactions and Dispersion Forces Explain how dipole interactions and dispersion forces are related. First, explain what produces the attractions between polar molecules. Then explain what produces dispersion forces between molecules. Identify what is similar and what is different in the two mechanisms of intermolecular attraction.

Interactive Textbook

Assessment 8.4 Test yourself on the important concepts of Section 8.4.

with **ChemASAP**

Section 8.4 Assessment

32. The more electronegative atom attracts electrons more strongly and gains a partial negative charge. The less electronegative atom has a partial positive charge.

33. Polar molecules tend to become oriented with respect to the positive and negative plates.

34. Intermolecular attractions are weaker than either ionic or covalent bonds.

35. Melting a network solid requires breaking covalent bonds throughout the solid.

36. The atoms in CCl_4 are oriented so that the bond polarities cancel.

37. a. $\overset{\delta+}{\mathrm{H}}:\overset{\delta-}{\ddot{\underset{..}{\mathrm{O}}}}:\overset{\delta-}{\ddot{\underset{..}{\mathrm{O}}}}:\overset{\delta+}{\mathrm{H}}$ **b.** $:\overset{\delta+}{\ddot{\underset{..}{\mathrm{Br}}}}:\overset{\delta-}{\ddot{\underset{..}{\mathrm{Cl}}}}:$ **c.** $\overset{\delta+}{\mathrm{H}}:\overset{\delta-}{\ddot{\underset{..}{\mathrm{Br}}}}:$

d. $\overset{\delta+}{\mathrm{H}}:\overset{\delta-}{\ddot{\underset{..}{\mathrm{O}}}}:\overset{\delta+}{\mathrm{H}}$

38. The atoms in a network solid are covalently bonded in a large array (or crystal), which can be thought of as a single molecule.