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 crystal forms. In this section, you will see that the polar bonds in water molecules influence the distinctive geometry of snowflakes.

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.

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.

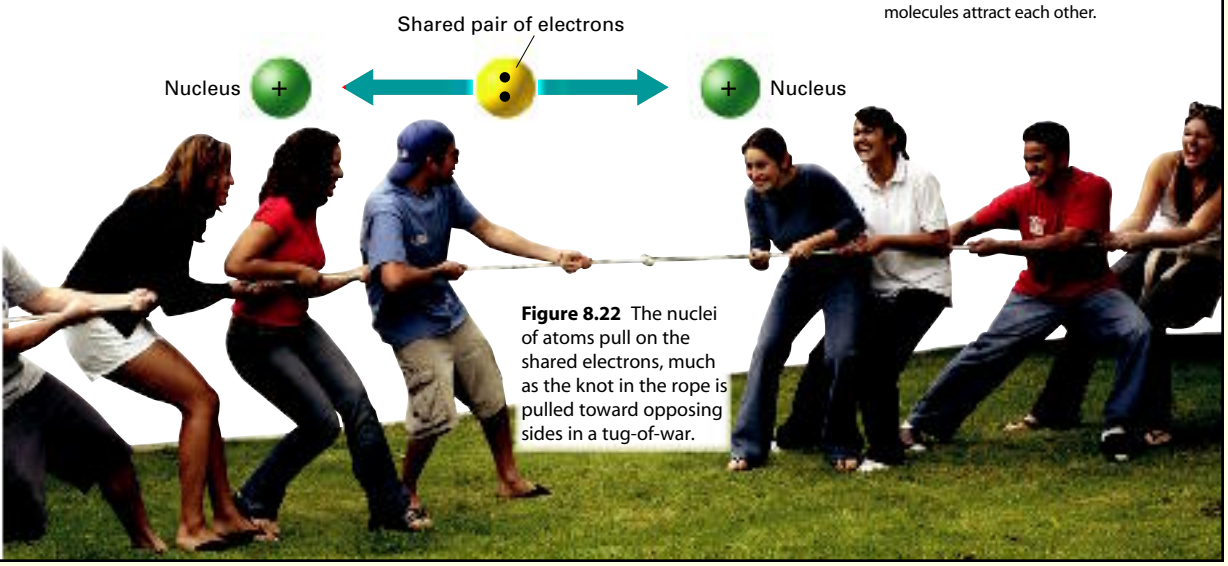


Figure 8.22 The nuclei of atoms pull on the shared electrons, much as the knot in the rope is pulled toward opposing sides in a tug-of-war.

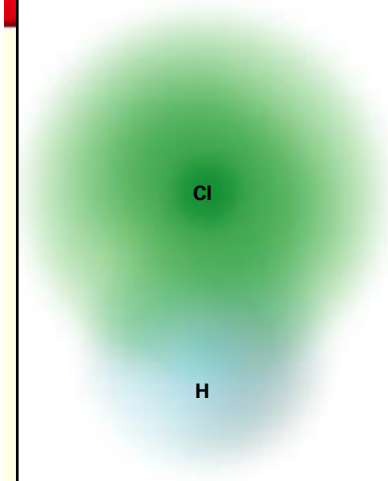


Figure 8.23 This electron-cloud picture of hydrogen chloride shows that the chlorine atom attracts the electron cloud more than the hydrogen atom does. **Inferring** ***Which atom is more electronegative, a chlorine atom or a hydrogen atom?***

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.

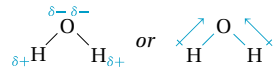
Consider the hydrogen chloride molecule (HCl) shown in Figure 8.23. Hydrogen has an electronegativity of 2.1, and chlorine has an electronegativity of 3.0. These values are significantly different, so the covalent bond in hydrogen chloride is polar. The chlorine atom acquires a slightly negative charge. 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−.

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

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.

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

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.



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	$\overset{\delta+}{\mathrm{H}}-\overset{\delta-}{\mathrm{Cl}}$ (0.9)
1.0–2.0	Very polar covalent	$\overset{\delta+}{\mathrm{H}}-\overset{\delta-}{\mathrm{F}}$ (1.9)
≥ 2.0	Ionic	Na^+Cl^- (2.1)

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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.
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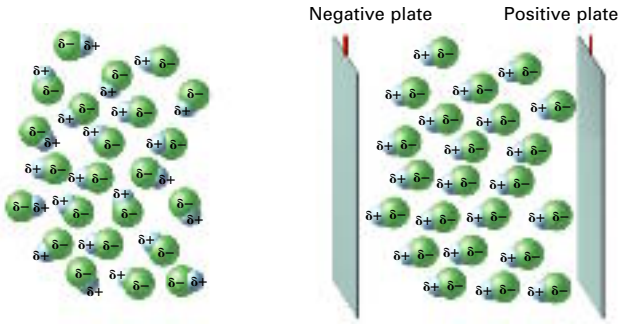

Electric field absent.
Polar molecules orient randomly.

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?***

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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.

$$\overleftarrow{O}=\overset{+}{C}=\overrightarrow{O}$$

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.

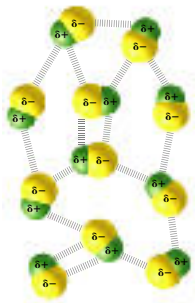
Van der Waals Forces

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.



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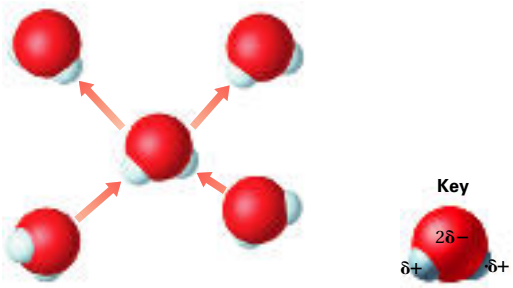


Figure 8.26 The strong hydrogen bonding between water molecules accounts for many properties of water, such as the fact that water is a liquid rather than a gas at room temperature.

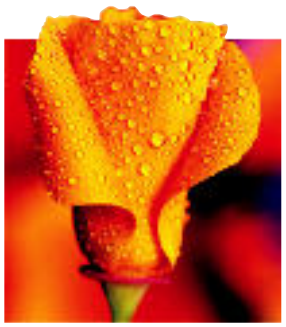
Hydrogen Bonds The dipole interactions in water produce an attraction between water molecules. Each O—H bond in the water molecule is highly polar, and the oxygen acquires a slightly negative charge because of its greater electronegativity. The hydrogens in water molecules acquire a slightly positive 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.



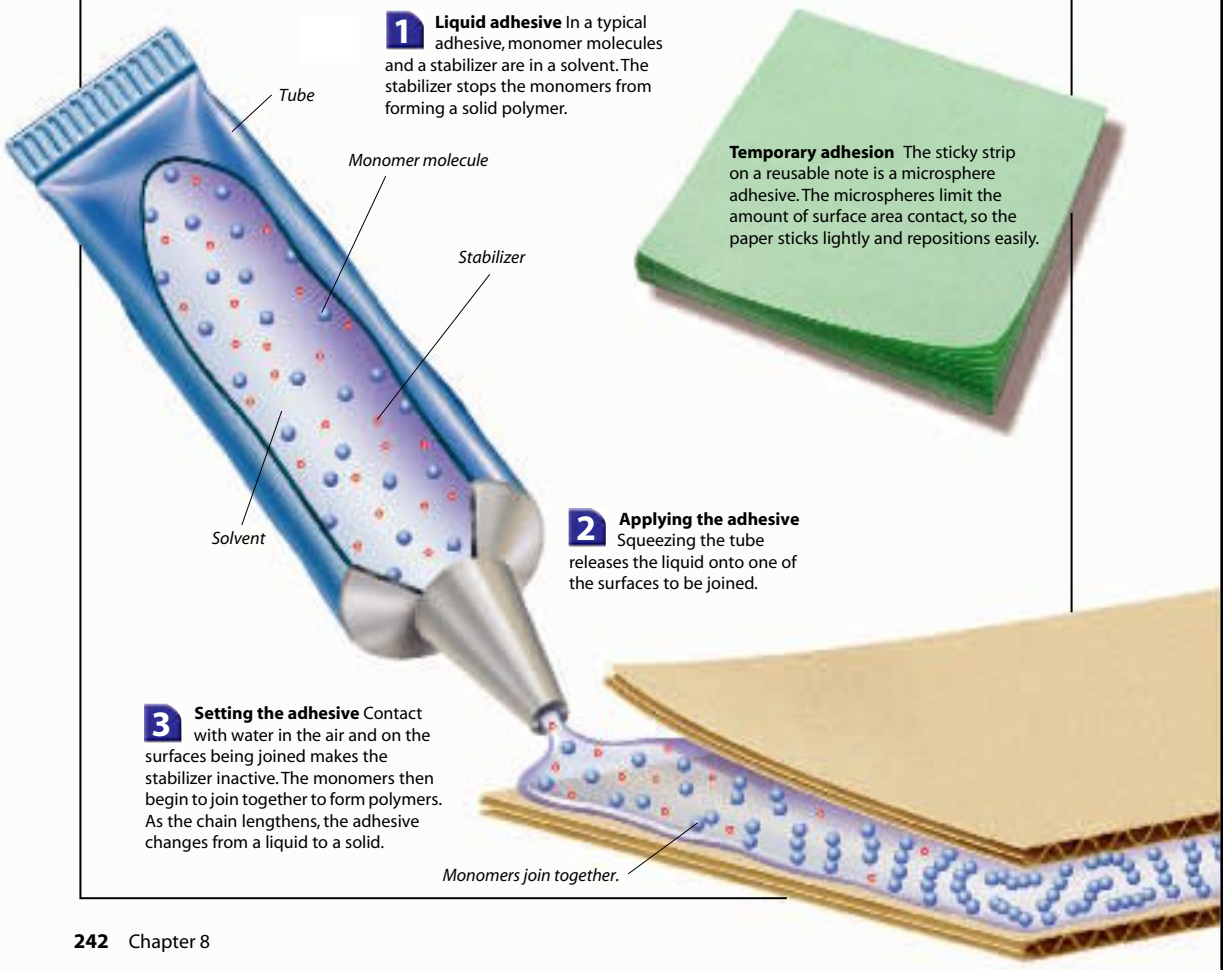
Technology & Society

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.



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.

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.

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.

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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

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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.
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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 (δ+) and slightly negative (δ−) 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.

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