How does a snowflake get its shape?

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.

Bond Polarity

How do electronegativity values determine the charge distribution in a polar bond?

Covalent bonds differ in terms of how the bonded atoms share the electrons.

- The character of the 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 between the nuclei of the atoms sharing the electrons.

- 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.

The bonding pairs of electrons in covalent bonds are pulled 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 each bond formed is a **nonpolar covalent bond**.

A **polar covalent bond**, known also as a **polar bond**, is a covalent bond between atoms in which the electrons are shared unequally.

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 more strongly and gains a slightly negative charge. The less electronegative atom has a slightly positive charge.

The higher the electronegativity value, the greater the ability of an atom to attract electrons to itself.

Describing Polar Covalent Bonds

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.

Describing Polar Covalent Bonds

Hydrogen has an electronegativity of 2.1, and chlorine has an electronegativity of 3.0.

- The chlorine atom, with its higher electronegativity, acquires a slightly negative charge.
- The hydrogen atom acquires a slightly positive charge.

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

Describing Polar Covalent Bonds

The lowercase Greek letter delta (δ) denotes that atoms in the covalent bond acquire only partial charges, less than 1+ or 1–.

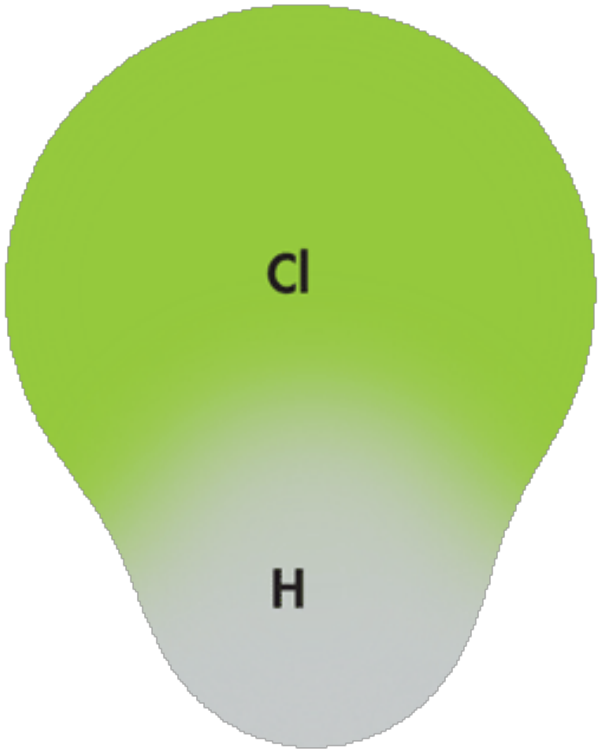
$$\overset{\delta+}{H}—\overset{\delta-}{Cl}$$

- The minus sign shows that chlorine has a slightly negative charge.
- The plus sign shows that hydrogen has acquired a slightly positive charge.

Describing Polar Covalent Bonds

These partial charges are shown as clouds of electron density.

- This electron-cloud picture of hydrogen chloride shows that the chlorine atom attracts the electron cloud more than the hydrogen atom does.



Describing Polar Covalent Bonds

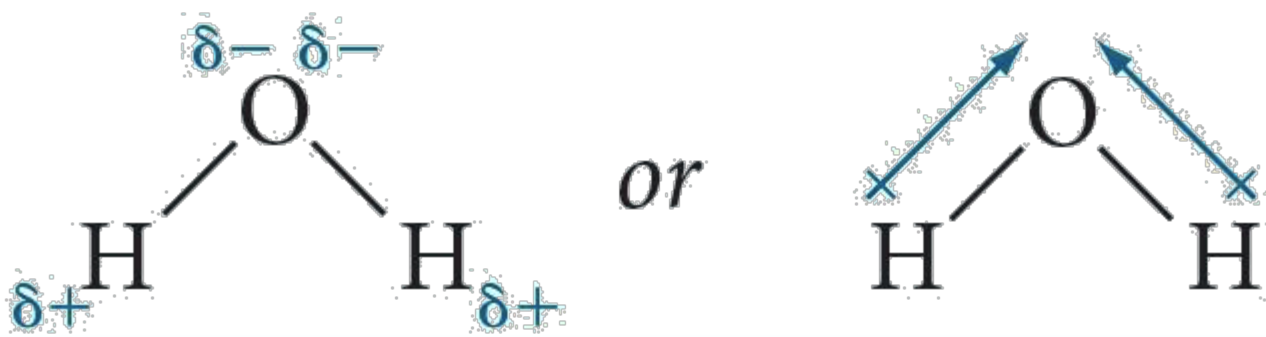
The polar nature of the bond may also be represented by an arrow pointing to the more electronegative atom.

H—Cl

Describing Polar Covalent Bonds

The O—H bonds in a water molecule are also polar.

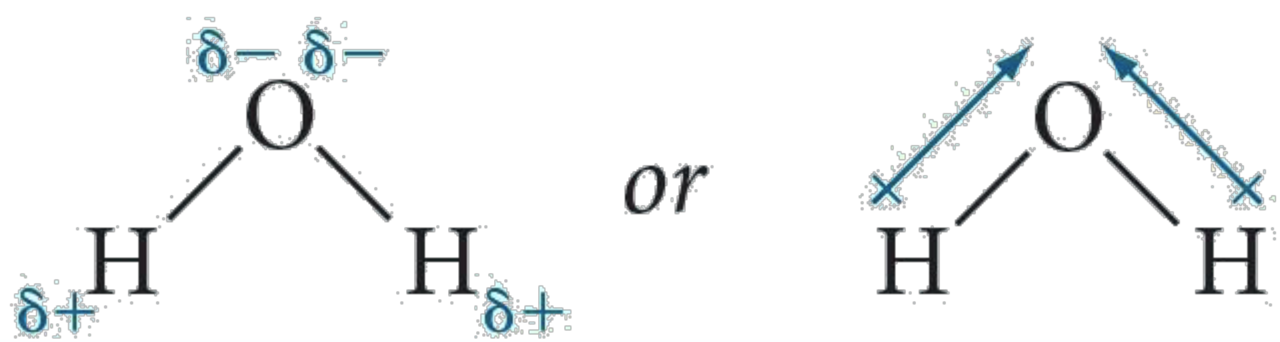
- The highly electronegative oxygen partially pulls the bonding electrons away from hydrogen.



Describing Polar Covalent Bonds

The O—H bonds in a water molecule are also polar.

- The oxygen acquires a slightly negative charge.
- The hydrogen is left with a slightly positive charge.



Describing Polar Covalent Bonds

The electronegativity difference between two atoms tells you what kind of bond is likely to form.

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)

Describing Polar Covalent Bonds

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 difference is more than 2.0, the electrons will likely be pulled away completely by one of the atoms.
  - In that case, an ionic bond will form.

Identifying Bond Type

Which type of bond (nonpolar covalent, moderately polar covalent, very 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.

2 Solve Apply concepts to this problem.

Identify the electronegativities of each atom using Table 6.2.

a. N(3.0), H(2.1)

b. F(4.0), F(4.0)

c. Ca(1.0), Cl(3.0)

d. Al(1.5), Cl(3.0)

2 Solve Apply concepts to this problem.

Calculate the electronegativity difference between the two atoms.

a. N(3.0), H(2.1); 0.9

b. F(4.0), F(4.0); 0.0

c. Ca(1.0), Cl(3.0); 2.0

d. Al(1.5), Cl(3.0); 1.5

The electronegativity difference between two atoms is expressed as the absolute value. So, you will never express the difference as a negative number.

2 Solve Apply concepts to this problem.

Based on the electronegativity difference, determine the bond type using Table 8.4.

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

Describing Polar Covalent 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.

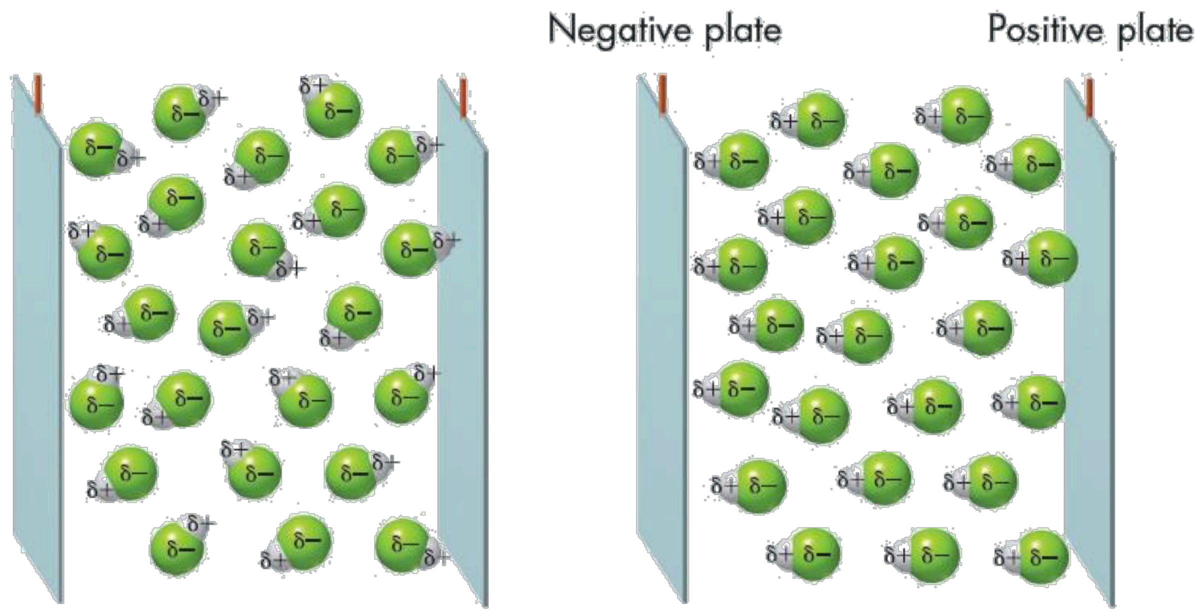
Describing Polar Covalent Molecules

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.

Describing Polar Covalent Molecules

When polar molecules are placed between oppositely charged plates, they tend to become oriented with respect to the positive and negative plates.



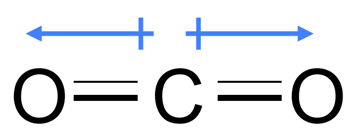
Electric field is absent. Polar molecules orient randomly.

Electric field is on. Polar molecules line up.

Describing Polar Covalent Molecules

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 has two polar bonds and is linear.



Describing Polar Covalent Molecules

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.

What is the difference between an ionic bond and a very polar covalent bond?

What is the difference between an ionic bond and a very polar covalent bond?

Two atoms will form an ionic bond rather than a very polar covalent bond if the two atoms have a slightly higher difference in electronegativity—a difference of more than 2.0. There is no sharp boundary between ionic and covalent bonds.

Attractions Between Molecules

How do the strengths of intermolecular attractions compare with ionic and covalent bonds?

Molecules can be attracted to each other by a variety of different forces.

Intermolecular attractions are weaker than either ionic or covalent bonds.

- 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.

- Van der Waals forces consist of dipole interactions and dispersion forces.

Van der Waals Forces

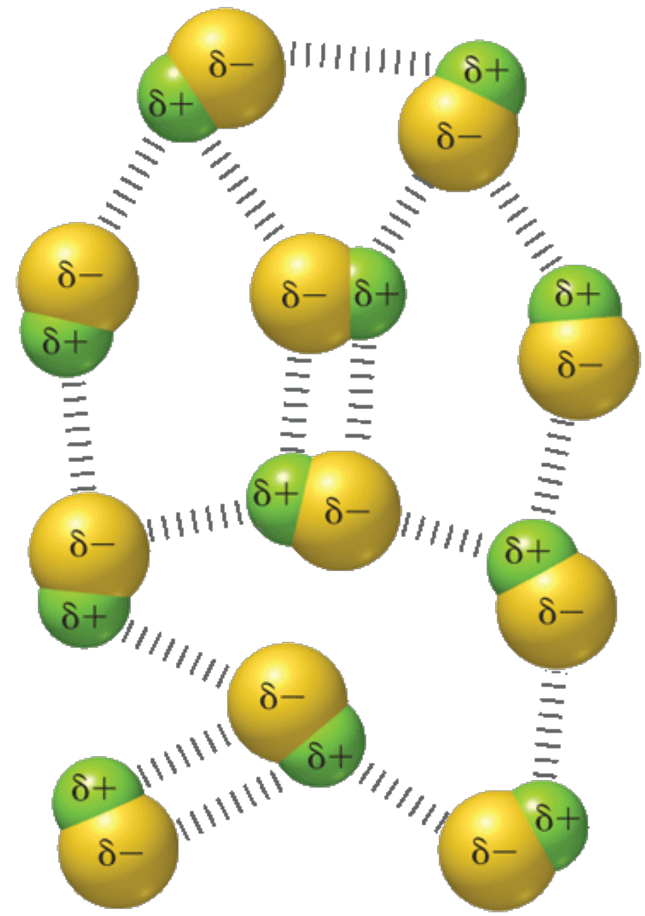
Dipole interactions occur when polar molecules are attracted to one another.

- The electrical attraction occurs between the oppositely charged regions of polar molecules.
- Dipole interactions are similar to, but much weaker than, ionic bonds.

Van der Waals Forces

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.



Van der Waals Forces

Dispersion forces, the weakest of all molecular interactions, are caused by the motion of electrons.

- They occur even between nonpolar molecules.

Van der Waals Forces

Dispersion forces, the weakest of all molecular interactions, are caused by the motion of electrons.

- 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.

Van der Waals Forces

Dispersion forces, the weakest of all molecular interactions, are caused by the motion of electrons.

- The strength of dispersion forces generally increases as the number of electrons in a molecule increases.

Van der Waals Forces

- Fluorine and chlorine have relatively few electrons and are gases at ordinary room temperature and pressure because of their especially weak dispersion forces.
- Bromine molecules therefore attract each other sufficiently to make bromine a liquid under ordinary room temperature and pressure.
- Iodine, with a still larger number of electrons, is a solid at ordinary room temperature and pressure.

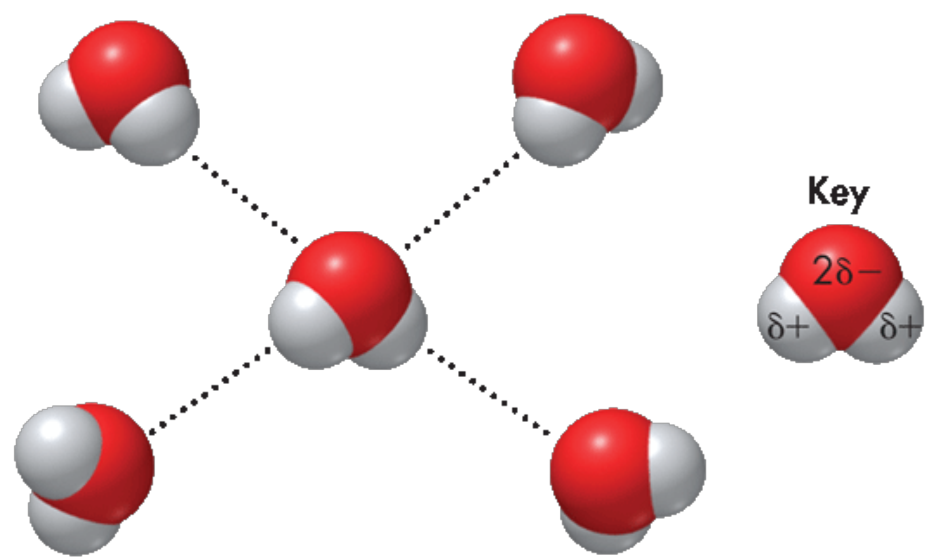
Hydrogen Bonds

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.

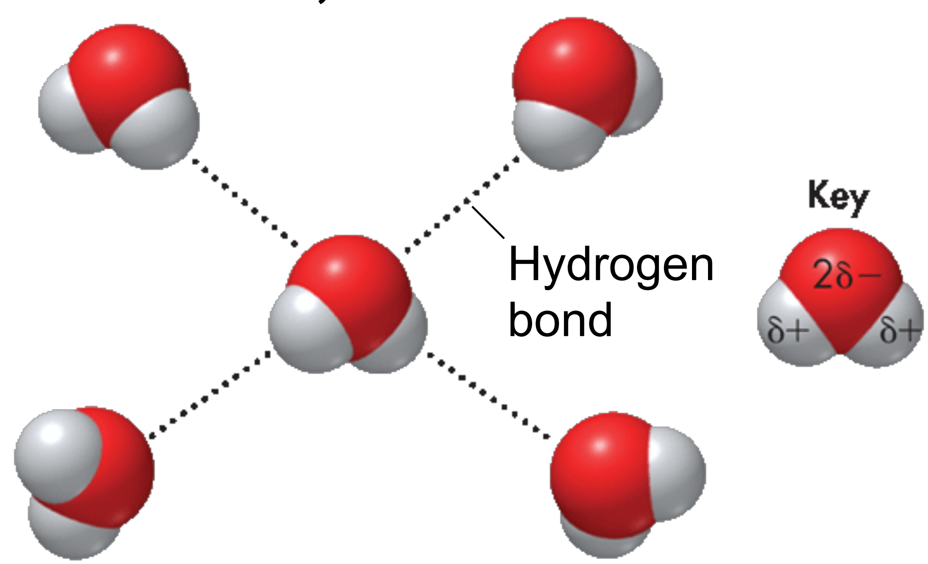
Hydrogen Bonds

The positive region of one water molecule attracts the negative region of another water molecule.



Hydrogen Bonds

This relatively strong attraction, which is also found in hydrogen-containing molecules other than water, is called a hydrogen bond.



Hydrogen Bonds

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.

- The other atom may be in the same molecule or in a nearby molecule.
- Hydrogen bonding always involves hydrogen.

Hydrogen Bonds

A hydrogen bond has about 5 percent of the strength of the 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.

How does a snowflake get its shape?

How does a snowflake get its shape?

A snowflake's shape is determined by the interactions of hydrogen bonds during its formation.

Why are hydrogen bonds important?

Why are hydrogen bonds important?

Hydrogen bonds are the strongest of the intermolecular forces and are extremely important in determining the properties of water and biological molecules such as proteins.

Intermolecular Attractions and Molecular Properties

Why are the properties of covalent compounds so diverse?

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.

The diversity of physical properties among covalent compounds 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 by molecules, only the weak attractions between molecules need to be broken.

A few solids that consist of molecules do not melt until the temperature reaches 1000°C or higher.

- 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.

- Each carbon atom in a diamond is covalently bonded to four other carbons, interconnecting carbon atoms throughout the diamond.
- Diamond does not melt; rather, it vaporizes to a gas at 3500°C and above.