Polar Covalent Bonds and Polar Molecules

3.3

Section Preview/ Specific Expectations

In this section, you will

- **construct** molecular models
- **predict** the polarity of a given bond, using electronegativity values
- **predict** the overall polarity of molecules, using electronegativity values and molecular models
- **communicate** your understanding of the following terms: *polar covalent bond, lone pairs, bonding pairs, polar molecule, dipolar molecules, non-polar molecule*

In section 3.2, you learned what kind of bond forms when the electronegativity difference between two atoms is very small or very large. You now understand how electrons are shared or transferred in bonds. Thus, you can explain the properties of ionic compounds, and some of the properties of covalent compounds.

How can you explain the wide variety of properties that covalent compounds have? Covalent compounds may be solids, liquids, or gases at different temperatures. Some covalent compounds dissolve in water, and some do not. In fact, water itself is a covalent compound! Examine Figures 3.22 and 3.23. Why are the bonds in water different from the bonds in dinitrogen monoxide? Both of these compounds are made up of two elements, and each molecule contains three atoms. The differences in the properties of these compounds are explained in part by the ΔEN of their bonds.

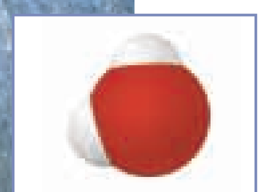


Figure 3.22 Water may be liquid, solid, or gas in nature. Why does the water that is sprayed up by this skier form a sheet?

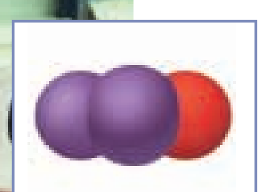


Figure 3.23 Dinitrogen monoxide, also known as laughing gas, boils at about –89°C. Laughing gas is used as an anaesthetic for dental work.

Polar Covalent Bonds: The "In-Between" Bonds

When two bonding atoms have an electronegativity difference that is greater than 0.5 but less than 1.7, they are considered to be a particular type of covalent bond called a **polar covalent bond**. In a polar covalent bond, the atoms have significantly different electronegativities. The electronegativity difference is not great enough, however, for the less electronegative atom to transfer its valence electrons to the other, more electronegative atom. The difference *is* great enough for the bonding electron pair to spend more time near the more electronegative atom than the less electronegative atom.

For example, the bond between oxygen and hydrogen in water has an electronegativity difference of 1.24. Because this value falls between 0.5 and 1.7, the bond is a polar covalent bond. The oxygen attracts the electrons more strongly than the hydrogen. Therefore, the oxygen has a slightly negative charge and the hydrogen has a slightly positive charge. Since the hydrogen does not completely transfer its electron to the oxygen, their respective charges are not +1 and −1, but rather δ^+ and δ^-. The symbol δ^+ (delta plus) stands for a partial positive charge. The symbol δ^- (delta minus) stands for a partial negative charge. Figure 3.24 illustrates the partial negative and positive charges across an oxygen-hydrogen bond. Figure 3.25 shows the polar covalent bond between hydrogen and chlorine.

PROBEWARE

If you have access to probeware, do the Chemistry 11 lab, Properties of Bonds, now.

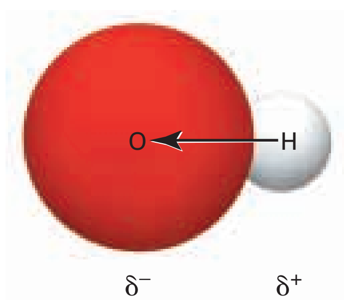


Figure 3.24 The O end of an O—H bond has a partial negative charge. The H end has a partial positive charge.

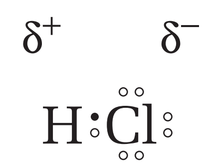


Figure 3.25 The Cl end of a H—Cl bond has a partial negative charge. The H end has a partial positive charge.

Try the following problems to practise identifying the partial charges across polar covalent bonds.

Practice Problems

11. Predict whether each bond will be covalent, polar covalent, or ionic.

(a) C—F **(c)** Cl—Cl **(e)** Si—H **(g)** Fe—O

(b) O—N **(d)** Cu—O **(f)** Na—F **(h)** Mn—O

12. For each polar covalent bond in problem 11, indicate the locations of the partial charges.

13. Arrange the bonds in each set in order of increasing polarity. (A completely polarized bond is an ionic bond.)

(a) H—Cl, O—O, N—O, Na—Cl

(b) C—Cl, Mg—Cl, P—O, N—N

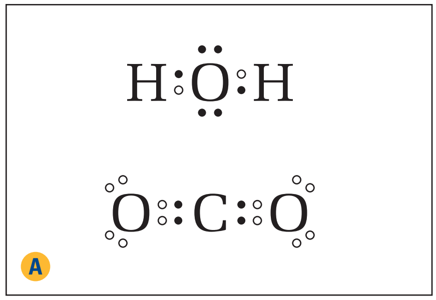
Electronic Learning Partner

Your Chemistry 11 Electronic Learning Partner has a film clip that shows the formation of bonds in water, hydrogen gas, and sodium chloride.

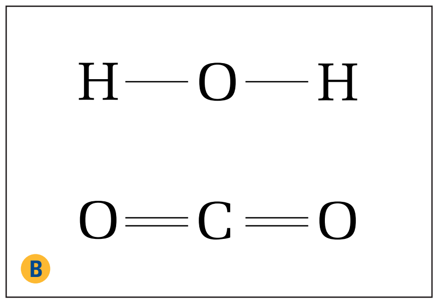
Comparing Molecular Models

Throughout this chapter, you have seen several different types of diagrams representing molecules. These diagrams, or models, are useful for highlighting various aspects of molecules and bonding. Examine Figure 3.26 to see the various strengths of the different models.

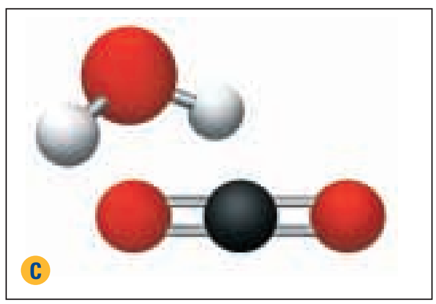
A A *Lewis structure* shows you exactly how many electrons are involved in each bond in a compound. Some Lewis structures show bonding pairs as lines between atoms.



B A *structural diagram* shows single bonds as single lines and multiple bonds as multiple lines. It does not show non-bonding pairs. It is less cluttered than a Lewis structure. It clearly shows whether the bonds involved are single, double, or triple bonds.



C A *ball-and-stick model* shows atoms as spheres and bonds as sticks. It accurately shows how the bonds within a molecule are oriented in three-dimensional space. The distances between the atoms are exaggerated, however. In this model, you can see the differences in the shapes of carbon dioxide and water.



D A *space-filling model* shows atoms as spheres. It is the most accurate representation of the shape of a real molecule.

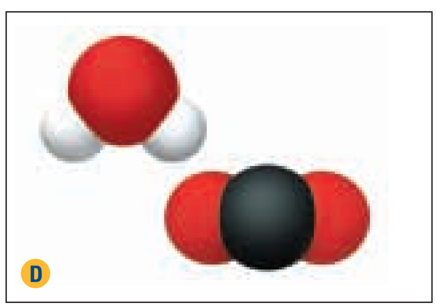


Figure 3.26 **You can compare a molecule of water with a molecule of carbon dioxide using a variety of different models.**

Consider a molecule of water and a molecule of carbon dioxide. Both water and carbon dioxide contain two atoms of the same element bonded to a third atom of another element. According to Figure 3.26, however, water and carbon dioxide molecules are different shapes. Why does carbon dioxide have a linear shape while water is bent?

Electronic Learning Partner

Your Chemistry 11 Electronic Learning Partner has a video clip that explains how to draw Lewis structures.

Web LINK

There are ways of representing molecules in addition to the ones shown in Figure 3.26. Seach for some examples on the Internet. Go to **www.school.mcgrawhill.ca/resources/** for some ideas on where to start.

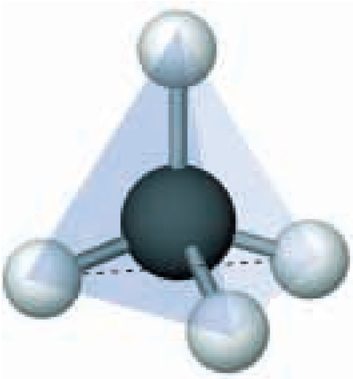


Figure 3.27 A tetrahedron has four equal sides.

To understand why molecules have different shapes, consider how electron arrangement affects shape. The Lewis structure for water, for example, shows that the oxygen is surrounded by four electron pairs. As shown in figure 3.26, two of the pairs are involved in bonding with the hydrogen atoms and two of the pairs are not. Electron pairs that are not involved in bonding are called **lone pairs**. Electron pairs that are involved in bonding are called **bonding pairs**.

Electron pairs are arranged around molecules so that they are a maximum distance from each other. This makes sense, because electrons are negatively charged and they repel each other. The shape that allows four electron pairs to be a maximum distance from each other around an atom is a tetrahedron. Figure 3.27 shows a tetrahedron.

The Shape of a Water Molecule

In a water molecule, there are four electron pairs around the oxygen atom. Two of these pairs bond with the hydrogen. The electron pairs are arranged in a shape that is nearly tetrahedral. When you draw the molecule, however, you draw only the oxygen atom and the two hydrogen atoms. This is where the bent shape comes from, as you can see in Figure 3.28.

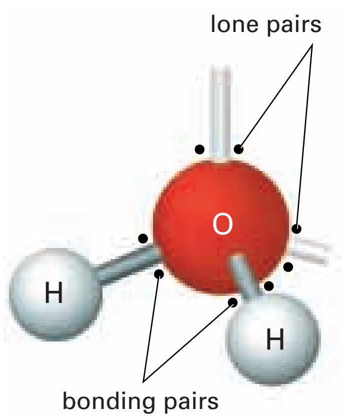


Figure 3.28 Two non-bending pairs account for water's bent shape.

The Shape of a Carbon Dioxide Molecule

Now consider carbon dioxide, CO_2. Why does a carbon dioxide molecule have a linear shape? Examine the Lewis structure for carbon dioxide. The central carbon atom is surrounded by eight electrons (four pairs), like the oxygen atom in a water molecule. In a carbon dioxide molecule, though, all the electrons are involved in bonding. There are no lone pairs. Because the bonding electrons spend most of their time between the carbon and oxygen atoms, they are arranged in a straight line. This allows them to be as far away from each other as possible, as you can see in Figure 3.29.

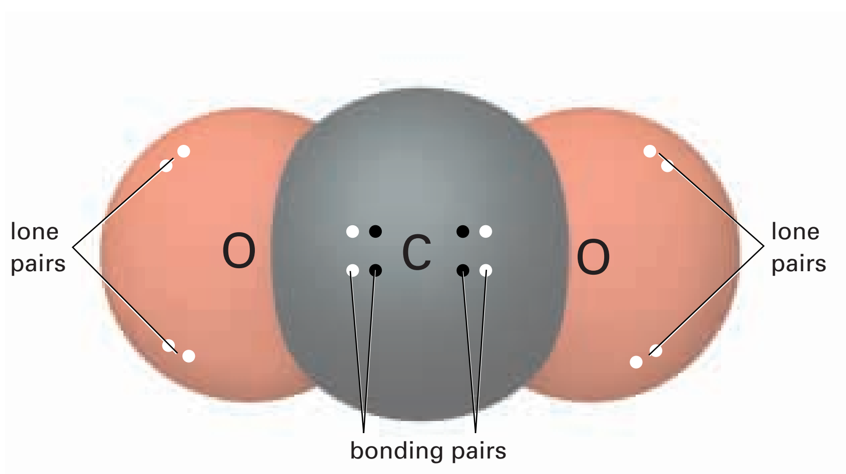


Figure 3.29 Carbon dioxide is linear in shape.

The shapes of the water molecule and the carbon dioxide molecule, as shown in the diagram you have seen, make sense based on what we know about electron pairs. These shapes have also been supported by experiment. You will learn more about experimental evidence for the structure of carbon dioxide and water later in this chapter.

Drawing the Lewis structure of a molecule can help you determine the molecule's shape. In Figure 3.30, you can see the shape of the ammonia, NH_3, molecule. The ammonia molecule has three bonding electron pairs and one lone pair on its central atom, all arranged in a nearly tetrahedral shape. Because there is one lone pair, the molecule's shape is pyramidal. The molecule methane, CH_4, is shown in Figure 3.31. This molecule has four bonding pairs on its central atom and no lone pairs. It is shaped like a perfectly symmetrical tetrahedron.

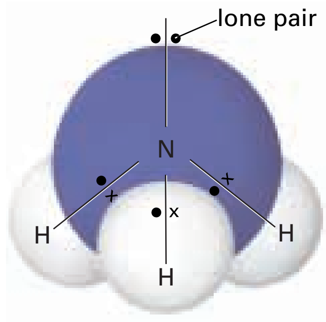


Figure 3.30 **An ammonia molecule is shaped like a pyramid.**

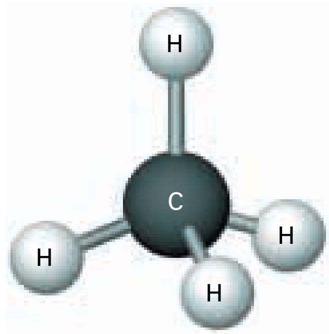


Figure 3.31 **A methane molecule is shaped as though its hydrogen atoms were on the corners of a tetrahedron.**

Canadians in Chemistry

Dr. Geoffrey Ozin

His work has flown on a Space Shuttle, and it has been hailed as art. It may well be part of the next computing revolution.

What does Dr. Geoffrey Ozin do? As little as possible, for he believes in letting the atoms do most of the work. This approach has made him one of the more celebrated chemists in Canada. Time and again, he has brought together organic and inorganic molecules, polymers, and metals in order to create materials with just the right structure for a specific purpose.

Self-assembly is the key. Atoms and molecules are driven into pre-designed shapes by intermolecular forces and geometrical constraints. At the University of Toronto, Dr. Ozin teaches his students the new science of intentional design, instead of the old trial-and-error methods.

Born in London, England, in 1943, Geoffrey Ozin earned a doctorate in chemistry at Oxford University. He joined the University of Toronto in 1969. Ozin's father was a tailor. In a way, Ozin is continuing the family tradition. Ozin, however, uses ionic and covalent bonds, atoms and molecules, acids, gases, and solutions to fashion his creations.

In 1996, Dr. Ozin demonstrated the self-assembly of crystals with a porous structure in space, under the conditions (such as microgravity) found aboard a Space Shuttle. Since then, he has shown how the self-assembly of many materials can be controlled to produce their structure.

Dr. Ozin's latest achievement involves structure. Ozin was part of an international research team that created regular microscopic cavities inside a piece of silicon. This material can transmit light photons in precisely regulated ways. In the future, this material might be used to build incredibly fast computers that function by means of photons instead of electrons!

Polar Bonds and Molecular Shapes

Water molecules are attracted to one another. Because we are surrounded by water, we are surrounded by evidence of this attraction. Re-examine the water skier in Figure 3.22. If water molecules did not attract one another, do you think the spray from the ski would form a "sheet" as shown? Try filling a glass with water. As you near the rim, add water very slowly. If you are careful, you can fill the glass so that the water bulges over the rim. After a rainfall, you have probably seen beads of water on the surface of vehicles. In Figure 3.32, you can see further evidence of the attraction of water molecules to one another.

Why do water molecules "stick together"? To answer this question, you need to consider both the nature of the bonds within a water molecule and its shape.

Figure 3.32 The shape of water droplets is evidence that water molecules are attracted to one another. This property of water can be explained by the polarity of its O—H bonds.

The Polar Water Molecule

First consider the shape of a water molecule. You have discovered that a water molecule has a bent shape. Each oxygen-hydrogen bond is polar. The hydrogen atom has a partial positive charge and the oxygen atom has a partial negative charge. You know that the bonds are polar, but what about the molecule as a whole? Because the molecule is bent, there is a partial negative charge on the oxygen end and a partial positive charge on the hydrogen end, as shown in Figure 3.33.

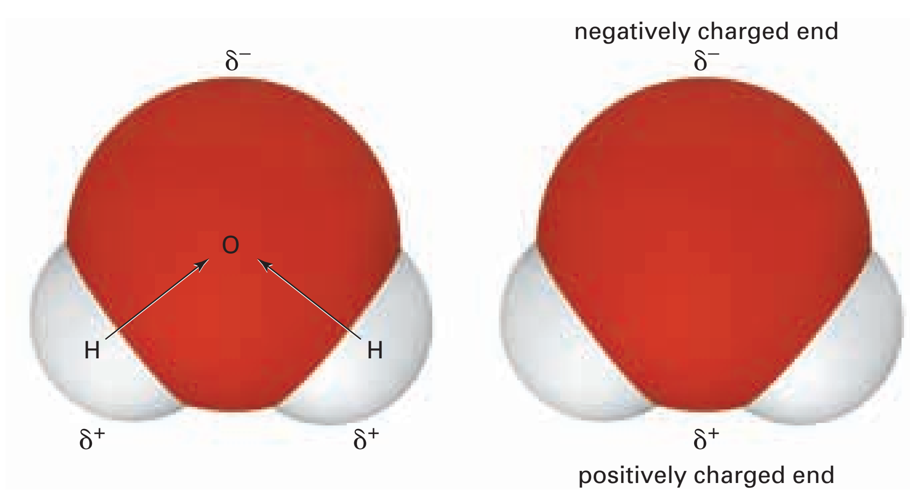


Figure 3.33 Water is a polar molecule because of its shape and the polarity of its bonds.

Because the water molecule *as a whole* has a partial negative charge on one end and a partial positive charge on the other end, it is called a **polar molecule**. Because water is polar, its negative and positive ends attract each other. This explains why liquid water "sticks" to itself. Figure 3.34 shows how water molecules attract each other in the liquid state.

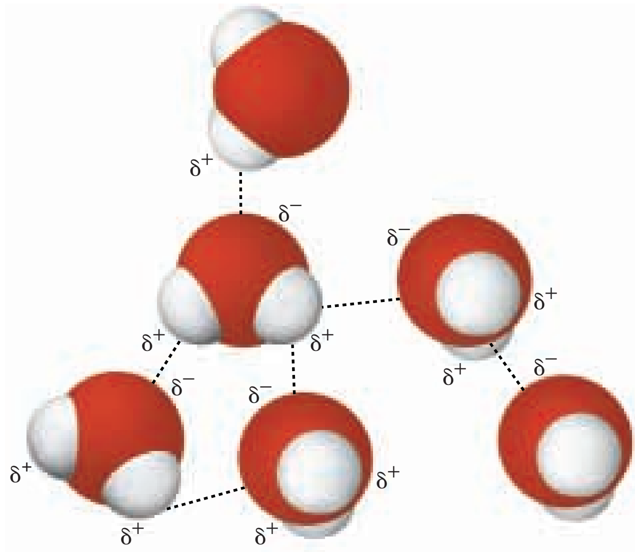


Figure 3.34 **The negative ends of water molecules attracts the positive ends. Some of the resulting intermolecular forces are shown here.**

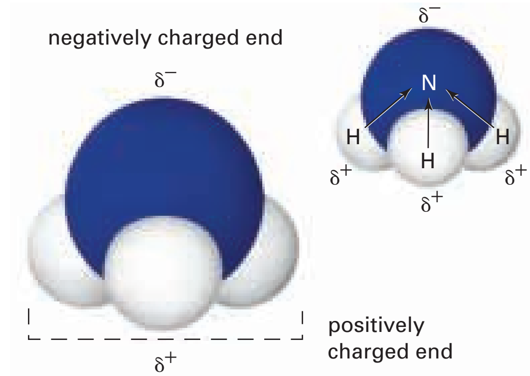


Figure 3.35 **Because an ammonia molecule contains polar bonds and is asymmetrical, it is a polar molecule.**

Two other examples of polar molecules are ammonia and hydrogen chloride, shown in Figures 3.35 and 3.36. Polar molecules are also called **dipolar molecules** because they have a negative pole and a positive pole.

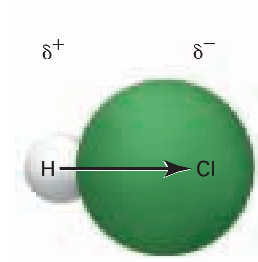


Figure 3.36 **Hydrogen chloride contains one polar bond. Therefore, the molecule is polar.**

The Non-Polar Carbon Dioxide Molecule

The bond between carbon and oxygen is polar. It has an electronegativity difference of 1.0. Does this mean that carbon dioxide, a molecule that contains two carbon-oxygen double bonds, is a polar molecule? No, it does not. The oxygen atoms have partial negative charges, and the carbon atom has a partial positive charge. The molecule, however, is straight and symmetrical. As you can see in Figure 3.37, the effects of the polar bonds cancel each other out. Therefore, while carbon dioxide contains polar bonds, it is a **non-polar molecule**. It has neither a positive pole nor a negative pole.

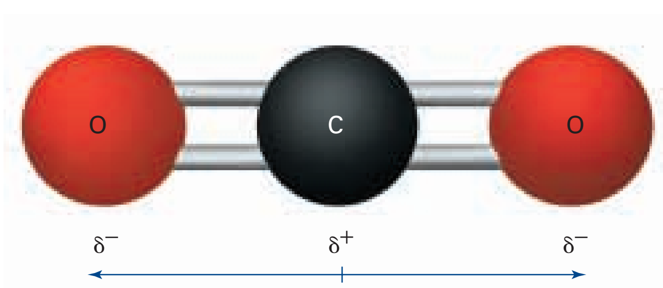


Figure 3.37 **Carbon dioxide is a non-polar molecule because it is symmetrical.**

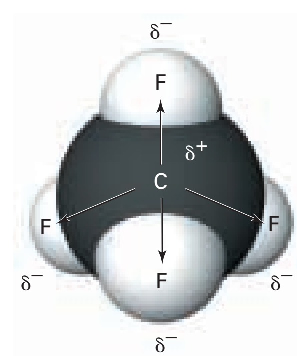


Figure 3.38 **Carbon tetrafluoride, CF_4, contains four polar bonds. Because of its symmetry, however, it is a non-polar molecule.**

Carbon tetrafluoride, CF_4, shown in Figure 3.38, is another example of a non-polar molecule that contains polar bonds.

SKILL FOCUS

Modelling concepts

Analyzing and interpreting

Communicating results

Investigation 3-B

Modelling Molecules

We cannot see molecules with our eyes or with a light microscope. We can predict their shapes, however, based on what we know about their electron configurations. In this investigation, you will practise working with a kit to build models of molecules.

Question

How can you build models of molecules to help you predict their shape and polarity?

Materials

molecular model kit
pen
paper

Procedure

1. Obtain a model kit from your teacher.
2. Draw a Lewis structure for each molecule below.
   (a) hydrogen bonded to a hydrogen: H_2
   (b) chlorine bonded to a chlorine: Cl_2
   (c) oxygen bonded to two hydrogens: H_2O
   (d) carbon bonded to two oxygens: CO_2
   (e) nitrogen bonded to three hydrogens: NH_3
   (f) carbon bonded to four chlorines: CCl_4
   (g) boron bonded to three fluorines: BF_3
3. Build a three-dimensional model of each molecule using your model kit.
4. Sketch the molecular models you have built.
5. In your notebook, make a table like the one below. Give it a title, fill in your data, and exchange your table with a classmate.

Compound	Lewis structure for compound	Sketch of predicted shape of molecule

Analysis

1. Compare your models with the models that your classmates built. Discuss any differences.
2. How did your Lewis structures help you predict the shape of each molecule?

Conclusion

3. Summarize the strengths and limitations of creating molecular models using molecular model kits.

Applications

4. Calculate the electronegativity difference for each bond in the molecules you built. Show partial charges. Based on the electronegativity difference and the predicted shape of each molecule, decide whether the molecule is polar or non-polar.
5. Look back through Chapter 3, and locate some different simple molecules. Build models of these molecules. Predict whether they are polar or non-polar.

Properties of Polar and Non-Polar Molecules

Because water is made up of polar molecules with positive and negative ends that attract one another, water tends to "stick" to itself. This means that it has a high melting point and boiling point, relative to other covalent compounds. For example, carbon dioxide is made up of non-polar molecules. These molecules do not attract each other as much as polar molecules do, because they do not have positive and negative poles. Compounds that are made up of non-polar molecules generally have lower melting points and boiling points than compounds that are made up of polar molecules. In fact, compounds with non-polar molecules, like carbon dioxide, are often gases at room temperature.

CHEM FACT

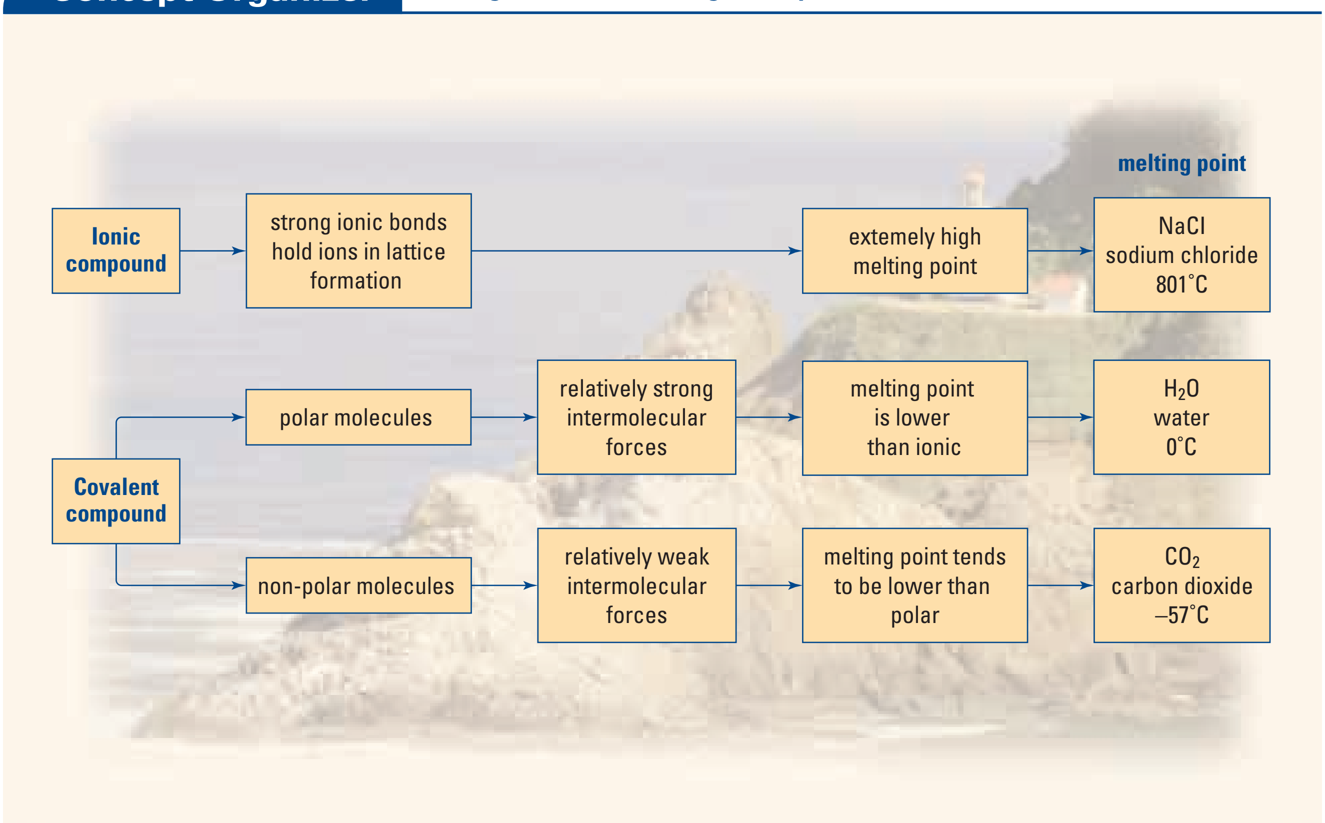
Differences in melting points and boiling points are due in part to the polarity of the molecules that make up the compounds. They are also due to the different masses of the individual molecules.

Section Wrap-up

In section 3.2, you learned about the strong bonds that hold ions in clearly-defined lattice patterns. You learned that these bonds are responsible for the properties of ionic compounds. You also learned how to describe the properties of compounds that are made up of molecules with covalent bonds. In this section, you discovered that the properties of compounds with polar covalent bonds depend on their shape. The following Concept Organizer summarizes some of the properties of covalent compounds that are made up of polar and non-polar molecules.

In sections 3.2 and 3.3, you learned how to represent compounds using Lewis structures and molecular models. In the next section, you will learn how chemists name compounds and represent them using symbols.

Concept Organizer Melting Point and Bonding Concepts



Section Review

Unit Project Prep

Before beginning your Unit Project, think about properties of compounds that would be useful in common chemical products. What kinds of properties would an abrasire or a window-cleaning fluid need to have? What kinds of compounds exhibit these properties?

1 K/U Determine ΔEN for each bond. Is the bond ionic, covalent, or polar covalent?

(a) B—F
(b) C—H
(c) Na—Cl
(d) Si—O
(e) S—O
(f) C—Cl

2 K/U For each polar covalent bond in question 1, label the partial negative and partial positive charges on each end.

3 C Explain how a non-polar molecule can contain polar bonds.

4 K/U Arrange each set of bonds in order of increasing polarity, using only their position in the periodic table.

(a) H—Cl, H—Br, O—F, K—Br
(b) C—O, C—F, C—H, C—Br

5 K/U Check your arrangements in question 4 by determining the ΔEN for each bond. Explain any discrepancies between your two sets of predictions.

6 I A molecule of chloroform, $CHCl_3$, has the same shape as a molecule of methane, CH_4. However, methane's boiling point is −164°C and chloroform's boiling point is 62°C. Explain the difference between the two boiling points.

7 K/U Determine the shape of each molecule by drawing a Lewis structure and considering the distribution of electron pairs around the atoms. Determine ΔEN for the bonds. Use the ΔEN and the shape to predict whether the molecule is polar or non-polar.

(a) $SiCl_4$
(b) PCl_3

8 MC How would Earth and life on Earth be different if water were a non-polar molecule? Write a paragarph explaining your ideas.