3.1 Classifying Chemical Compounds

Section Preview/ Specific Expectations

In this section, you will

- **describe** how electron arrangement and forces in atoms can explain the periodic trend associated with electronegativity
- **perform** a Thought Lab to classify compounds as ionic or covalent according to their properties
- **predict** the ionic character of a given bond using electronegativity values
- **communicate** your understanding of the following terms: *chemical bonds, ionic bond, covalent bond, electronegativity*

As you learned in the chapter opener, most elements do not exist in nature in their pure form, as elements. Gold, silver, and platinum are three metals that can be found in Earth's crust as elements. They are called "precious metals" because this occurrence is so rare. Most other metals, and most other elements, are found in nature only as compounds.

As the prospectors in the Yukon gold rush were searching for the element gold, they were surrounded by compounds. The streams they panned for gold ran with water, H_2O, a compound that is essential to the survival of nearly every organism on this planet. To sustain their energy, the prospectors ate food that contained, among other things, starch. Starch is a complex compound that consists of carbon, hydrogen, and oxygen. To flavour their food, they added sodium chloride, NaCl, which is commonly called table salt. Sometimes a compound called pyrite, also known as "fool's gold," tricked a prospector. Pyrite (iron disulfide, FeS_2) looks almost exactly like gold, as you can see in Figure 3.1. Pyrite, however, will corrode, and it is not composed of rare elements. Thus, it was not valuable to a prospector.

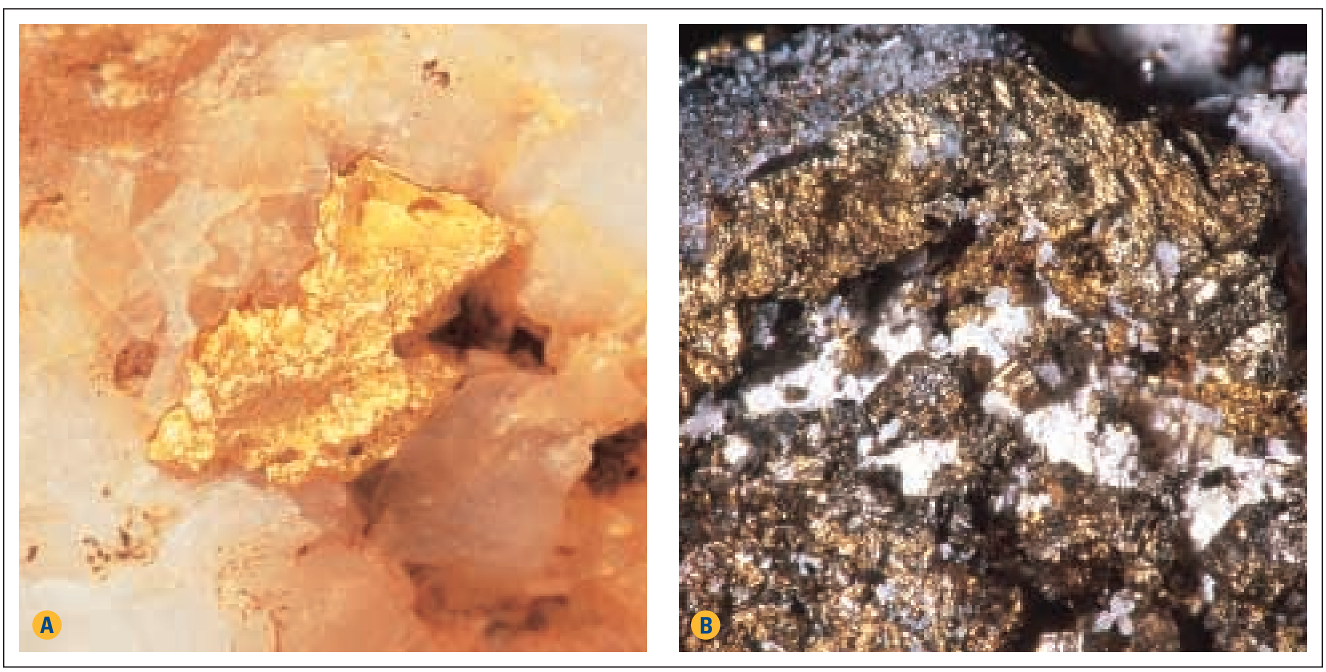

Figure 3.1 **Prospectors used the physical properties of gold and pyrite to distinguish between them. Can you tell which of these photos shows gold and which shows pyrite?**

There are only about 90 naturally occurring elements. In comparison, there are thousands upon thousands of different compounds in nature, and more are constantly being discovered. Elements combine in many different ways to form the astonishing variety of natural and synthetic compounds that you see and use every day.

Because there are so many compounds, chemists have developed a classification system to organize them according to their properties, such as melting point, boiling point, hardness, conductivity, and solubility. In the following Express Lab, you will use the property of magnetism to show that an element has formed a compound.

ExpressLab A Metal and a Compound

Humans have invented ways to extract iron from its compounds in order to take advantage of its properties. Does iron remain in its uncombined elemental form once it has been extracted? No, it doesn't. Instead, it forms rust, or iron(III) oxide, Fe_2O_3. How do we know that rust and iron are different substances? One way to check is to test a physical property, such as magnetism. In this activity, you will use magnetism to compare the properties of iron and rust.

Safety Precautions

Procedure

1. Obtain a new iron nail and a rusted iron nail from your teacher.
2. Obtain a thin, white piece of cardboard and a magnet. Wrap your magnet in plastic to keep it clean.
3. Test the iron nail with the magnet. Record your observations.
4. Gently rub the rusted nail with the other nail over the cardboard. Some rust powder will collect on the cardboard.
5. Hold up the cardboard horizontally. Move the magnet back and forth under the cardboard. Record your observations.

Analysis

1. How did the magnet affect the new iron nail? Based on your observations, is iron magnetic?
2. What did you observe when you moved the magnet under the rust powder?
3. What evidence do you have to show that iron and rust are different substances?
4. Consider what you know about iron and rust from your everyday experiences. Is it more likely that rust will form from iron, or iron from rust?

Properties of Ionic and Covalent Compounds

Based on their physical properties, compounds can be classified into two groups: ionic compounds and covalent compounds. Some of the properties of ionic and covalent compounds are summarized in Table 3.1.

Table 3.1 Comparing Ionic and Covalent Compounds

Property	Ionic compound	Covalent compound
state at room temperature	crystalline solid	liquid, gas, solid
melting point	high	low
electrical conductivity as a liquid	yes	no
solubility in water	most have high solubility	most have low solubility
conducts electricity when dissolved in water	yes	not usually

In the following Thought Lab, you will use the properties of various compounds to classify them as covalent or ionic.

ThoughtLab: Ionic or Covalent?

Imagine that you are a chemist. A colleague has just carried out a series of tests on the following compounds:

ethanol
carbon tetrachloride
glucose
table salt (sodium chloride)
water
potassium permanganate

You take the results home to organize and analyze them. Unfortunately your colleague labelled the tests by sample number and forgot to write down which compound corresponded to each sample number. You realize, however, that you can use the properties of the compounds to identify them. Then you can use the compounds' properties to decide whether they are ionic or covalent.

Procedure

1. Copy the following table into your notebook.

Sample	Compound name	Dissolves in water?	Conductivity as a liquid or when dissolved in water	Melting point	Appearance	Covalent or ionic?
1		yes	high	801°C	clear, white crystalline solid	
2		yes	low	0.0°C	clear, colourless liquid	
3		yes	high	240°C	purple, crystalline solid	
4		yes	low	146°C	white powder	
5		no	low	–23°C	clear, colourless liquid	
6		yes	low	–114°C	clear, colourless liquid	

2. Based on what you know about the properties of compounds, decide which compound corresponds to each set of properties. Write your decisions in your table. Once you have identified the samples, share your results as a class and come to a consensus. **Hint:** Carbon tetrachloride is not soluble in water.
3. Examine the properties associated with each compound. Decide whether each compound is ionic or covalent. If you are unsure, leave the space blank. Discuss your results as a class, and come to a consensus.

Analysis

1. Write down the reasoning you used to identify each compound, based on the properties given.
2. Write down the reasoning you used to decide whether each compound was ionic or covalent.
3. Were you unsure how to classify any of the compounds? Which ones, and why?
4. Think about the properties in the table you filled in, as well as your answers to questions 1 to 3. Which property is most useful for deciding whether a compound is ionic or covalent?
5. Suppose that you could further subdivide the covalent compounds into two groups, based on their properties. Which compounds would you group together? Explain your answer.

Applications

6. Use a chemistry reference book or the Internet to find an MSDS for ethanol, carbon tetrachloride, and potassium permanganate.
   (a) Write down the health hazards associated with each compound.
   (b) What precautions would a chemist who was performing tests on ethanol and carbon tetrachloride need to take?

Table Salt: An Ionic Compound

Sodium chloride, NaCl, is a familiar compound. You know it as table salt. The sodium in sodium chloride plays a vital role in body functions. We need to ingest about 500 mg of sodium a day. Too much sodium chloride, however, may contribute to high blood pressure. In the winter, sodium chloride is put on roads and sidewalks to melt the ice, as shown in Figure 3.2. Although this use of sodium chloride increases the safety of pedestrians and drivers, there are several drawbacks. For example, the saltwater discolours and damages footwear, and it corrodes the metal bodies of cars and trucks. Also, as shown in Figure 3.3, deer and moose that are attracted to the salt on the roads can be struck by vehicles.

Figure 3.2 **Sodium chloride is used to melt ice because salt water has a lower melting point than pure water.**

Figure 3.3 **This moose was attracted to the sodium chloride that was put on the road to melt snow and ice. Humans, like most organisms, need sodium to maintain normal body functions.**

Sodium chloride is a typical ionic compound. Like most ionic compounds, it is a crystalline solid at room temperature. It melts at a very high temperature, at 801°C. As well, it dissolves easily in water. A solution of sodium chloride in water is a good conductor of electricity. Liquid sodium chloride is also a good electrical conductor.

Carbon Dioxide: A Covalent Compound

The cells of most organisms produce carbon dioxide, CO_2, during cellular respiration: the process that releases energy from food. Plants, like the ones shown in Figure 3.4, synthesize their own food from carbon dioxide and water using the Sun's energy.

Carbon dioxide has most of the properties of a typical covalent compound. It has a low melting point (–79°C). At certain pressures and temperatures, carbon dioxide is a liquid. Liquid carbon dioxide is a weak conductor of electricity.

Figure 3.4 **Plants use carbon dioxide and water to produce their own food, using the Sun's energy.**

Figure 3.5 The bubbles fizzing out of the soft drink contain carbon dioxide.

Carbon dioxide is somewhat soluble in water, especially at high pressures. This is why soft drinks are bottled under pressure. When you open a bottle of pop, some of the carbon dioxide comes out of solution. Often, this happens too quickly, as you can see in Figure 3.5. A solution of carbon dioxide in water is a weak conductor of electricity.

What Is Bonding?

Why are carbon dioxide and sodium chloride so different? Why can we divide compounds into two categories that display distinct physical properties? The answers come from an understanding of **chemical bonds**: the forces that attract atoms to each other in compounds. *Bonding involves the interaction between the valence electrons of atoms.* Usually the formation of a bond between two atoms creates a compound that is more stable than either of the two atoms on their own.

The different properties of ionic and covalent compounds result from the manner in which chemical bonds form between atoms in these compounds. Atoms can either exchange or share electrons.

When two atoms exchange electrons, one atom loses its valence electron(s) and the other atom gains the electron(s). This kind of bonding usually occurs between a metal and a non-metal. Recall, from Chapter 2, that metals have low ionization energies and non-metals have high electron affinities. That is, metals tend to lose electrons and non-metals tend to gain them. When atoms exchange electrons, they form an **ionic bond**.

Atoms can also share electrons. This kind of bond forms between two non-metals. It can also form between a metal and a non-metal when the metal has a fairly high ionization energy. *When atoms share electrons, they form a* **covalent bond**.

mind STRETCH

Do you think that water is a covalent compound or an ionic compound? List water's physical properties. Can you tell whether water is a covalent compound or an ionic compound based only on its physical properties? Why or why not?

How can you determine whether the bonds that hold a compound together are ionic or covalent? Examining the physical properties of the compound is one method. This method is not always satisfactory, however. Often a compound has some ionic characteristics and some covalent characteristics. You saw this in the previous Thought Lab.

For example, hydrogen chloride, also known as hydrochloric acid, has a low melting point and a low boiling point. (It is a gas at room temperature.) These properties might lead you to believe that hydrogen chloride is a covalent compound. Hydrogen chloride, however, is extremely soluble in water, and the water solution conducts electricity. These properties are characteristic of an ionic compound. Is there a clear, theoretical way to decide whether the bond between hydrogen and chlorine is ionic or covalent? The answer lies in a periodic trend.

Electronegativity: Attracting Electrons

When two atoms form a bond, each atom attracts the other atom's electrons in addition to its own. The **electronegativity** of an atom is a measure of an atom's ability to attract electrons in a chemical bond. *EN* is used to symbolize electronegativity. There is a specific electronegativity associated with each element.

As you can see in Figure 3.6, electronegativity is a periodic property, just as atomic size, ionization energy, and electron affinity are. Atomic size, ionization energy, and electron affinity, however, are properties of single atoms. In contrast, electronegativity is a property of atoms that are involved in chemical bonding.

Electronegativities

1 H 2.20																	2 He -
3 Li 0.98	4 Be 1.57											5 B 2.04	6 C 2.55	7 N 3.04	8 O 3.44	9 F 3.98	10 Ne -
11 Na 0.93	12 Mg 1.31											13 Al 1.61	14 Si 1.90	15 P 2.19	16 S 2.58	17 Cl 3.16	18 Ar -
19 K 0.82	20 Ca 1.00	21 Sc 1.36	22 Ti 1.54	23 V 1.63	24 Cr 1.66	25 Mn 1.55	26 Fe 1.83	27 Co 1.88	28 Ni 1.91	29 Cu 1.90	30 Zn 1.65	31 Ga 1.81	32 Ge 2.01	33 As 2.18	34 Se 2.55	35 Br 2.96	36 Kr -
37 Rb 0.82	38 Sr 0.95	39 Y 1.22	40 Zr 1.33	41 Nb 1.6	42 Mo 2.16	43 Tc 2.10	44 Ru 2.2	45 Rh 2.28	46 Pd 2.20	47 Ag 1.93	48 Cd 1.69	49 In 1.78	50 Sn 1.96	51 Sb 2.05	52 Te 2.1	53 I 2.66	54 Xe -
55 Cs 0.79	56 Ba 0.89		72 Hf 1.3	73 Ta 1.5	74 W 1.7	75 Re 1.9	76 Os 2.2	77 Ir 2.2	78 Pt 2.2	79 Au 2.4	80 Hg 1.9	81 Tl 1.8	82 Pb 1.8	83 Bi 1.9	84 Po 2.0	85 At 2.2	86 Rn -
87 Fr 0.7	88 Ra 0.9		104 Rf -	105 Db -	106 Sg -	107 Bh -	108 Hs -	109 Mt -	110 Uun -	111 Uuu -	112 Uub -	113 -	114 Uuq -	115 -	116 Uuh -	117 -	118 Uuo -

57 La 1.10	58 Ce 1.12	59 Pr 1.13	60 Nd 1.14	61 Pm -	62 Sm 1.17	63 Eu -	64 Gd 1.20	65 Tb -	66 Dy 1.22	67 Ho 1.23	68 Er 1.24	69 Tm 1.25	70 Yb -	71 Lu 1.0
89 Ac 1.1	90 Th 1.3	91 Pa 1.5	92 U 1.7	93 Np 1.3	94 Pu 1.3	95 Am -	96 Cm -	97 BK -	98 Cf -	99 Es -	100 Fm -	101 Md -	102 No -	103 Lr -

Figure 3.6 **Electronegativity is a periodic trend. It increases up a group and across a period.**

The trend for electronegativity is the reverse of the trend for atomic size. Examine Figure 3.7, on the next page, to see what this means. In general, as atomic size decreases from left to right across a period, electronegativity increases. Why? The number of protons (which attract electrons) in the nucleus increases. At the same time, the number of filled, inner electron energy levels (which shield the protons from valence electrons) remains the same. Thus the electrons are pulled more tightly to the nucleus, resulting in a smaller atomic size. The atom attracts a bonding pair of electrons more strongly, because the bonding pair can move closer to the nucleus.

In the second period, for example, lithium has the largest atomic size and the lowest electronegativity. As atomic size decreases across the second period, the electronegativity increases. Fluorine has the smallest atomic size in the third period (except for neon) and the highest electronegativity. Because noble gases do not usually participate in bonding, their electronegativities are not given.

Similarly, as atomic size increases down a group, electronegativity decreases. As you move down a group, valence electrons are less strongly attracted to the nucleus because the number of filled electron energy levels between the nucleus and the valence electrons increases. In a compound, increasing energy levels between valence electrons and the nucleus mean that the nucleus attracts bonding pairs less strongly.

For example, in Group 2 (IIA), beryllium has the smallest atomic radius and the largest electronegativity. As atomic size increases down the group, electronegativity decreases.

Figure 3.7 shows the relationship between atomic size and electronegativity for the main-group elements in periods 2 to 6.

CHECKPOINT

Which element is the most electronegative? Not including the noble gases, which element is the least electronegative?

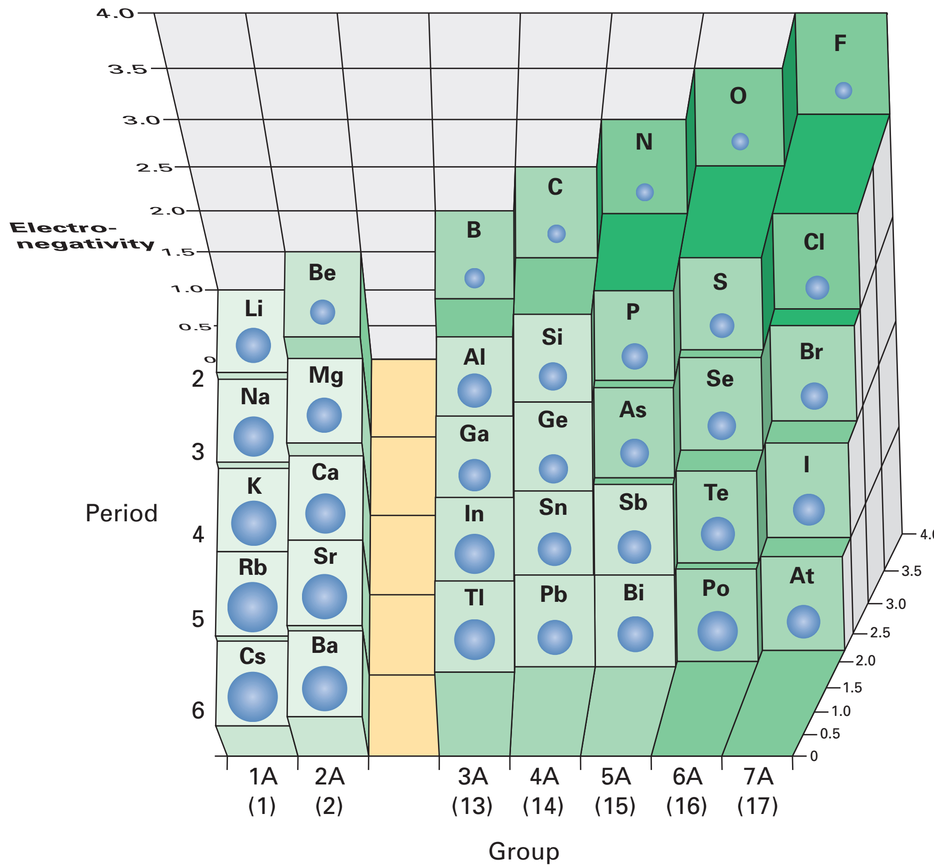

Figure 3.7 Periodic trends for electronegativity (bars) and atomic size (spheres) are inversely related.

Predicting Bond Type Using Electronegativity

You can use the differences between electronegativities to decide whether the bond between two atoms is ionic or covalent. The symbol ΔEN stands for the difference between two electronegativity values. When calculating the electronegativity difference, the smaller electronegativity is always subtracted from the larger electronegativity, so that the electronegativity difference is always positive.

How can the electronegativity difference help you predict the type of bond? By the end of this section, you will understand the aswer to this question. Consider three different substances: potassium fluoride, KF, oxygen, O_2, and hydrochloric acid, HCl. Potassium fluoride is an ionic compound made up of a metal and a non-metal that have very different electronegativities. Potassium's electronegativity is 0.82. Fluorine's electronegativity is 3.98. Therefore, ΔEN for the bond between potassium and fluorine is 3.16.

Now consider oxygen. This element exists as units of two atoms held together by covalent bonds. Each oxygen atom has an electronegativity of 3.44. The bond that holds the oxygen atoms together has an electronegativity difference of 0.00 because each atom in an oxygen molecule has an equal attraction for the bonding pair of electrons.

Finally, consider hydrogen chloride, or hydrochloric acid. Hydrogen has an electronegativity of 2.20, and chlorine has an electronegativity of 3.16. Therefore, the electronegativity difference for the chemical bond in hydrochloric acid, HCl, is 0.96. Hydrogen chloride is a gas at room temperature, but its water solution conducts electricity. Is hydrogen chloride a covalent compound or an ionic compound? Its ΔEN can help you decide, as you will see below.

The Range of Electronegativity Differences

When two atoms have electronegativities that are identical, as in oxygen, they share their bonding pair of electrons equally between them in a covalent bond. When two atoms have electronegativities that are very different, as in potassium fluoride, the atom with the lower electronegativity loses an electron to the atom with the higher electronegativity. In potassium fluoride, potassium gives up its valence electron to fluorine. Therefore, the bond is ionic.

It is not always clear whether atoms share electrons or transfer them. Atoms with different electronegativities can share electrons unequally without exchanging them. How unequal does the sharing have to be before the bond is considered ionic?

Figure 3.8 shows the range of electronegativity differences. These values go from mostly covalent at 0.0 to mostly ionic at 3.3. Chemists consider bonds with an electronegativity difference that is greater than 1.7 to be ionic, and bonds with an electronegativity difference that is less than 1.7 to be covalent.

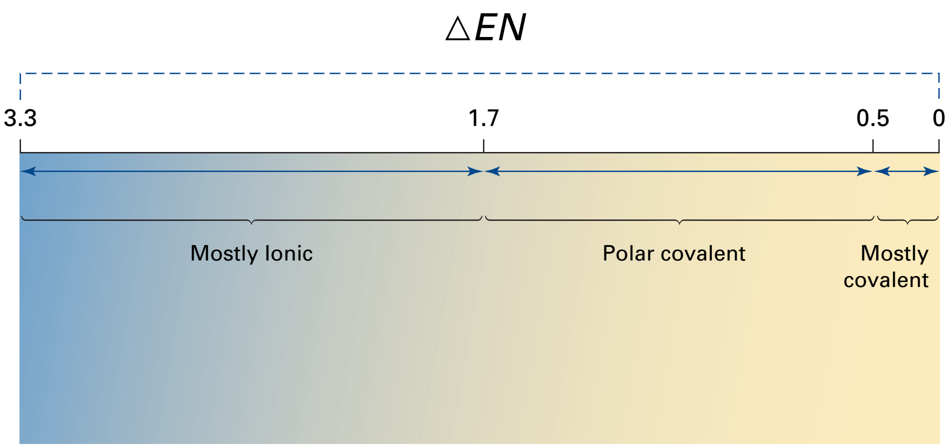

Figure 3.8 **Chemical bonds range in character from mostly ionic to mostly covalent.**

Table 3.2 shows how you can think of bonds as having a percent ionic character or percent covalent character, based on their electronegativity differences. When bonds have nearly 50% ionic or covalent character, they have characteristics of both types of bonding.

Table 3.2 Character of Bonds

Electronegativity difference	0.00	0.65	0.94	1.19	1.43	1.67	1.91	2.19	2.54	3.03
Percent ionic character	0%	10%	20%	30%	40%	50%	60%	70%	80%	90%
Percent covalent character	100%	90%	80%	70%	60%	50%	40%	30%	20%	10%

Based on Table 3.2, what kind of bond forms between hydrogen and chlorine? ΔEN for the bond in hydrogen chloride, HCl, is 0.96. This is lower than 1.7. Therefore, the bond in hydrogen chloride is a covalent bond.

Calculate ΔEN and predict bond character in the following Practice Problem.

Practice Problems

1. Determine ΔEN for each bond shown. Indicate whether each bond is ionic or covalent.

(a) O—H
(b) C—H
(c) Mg—Cl
(d) B—F
(e) Cr—O
(f) C—N
(g) Na—I
(h) Na—Br

Section Wrap-up

In this section, you learned that most elements do not exist in their pure form in nature. Rather, they exist as different compounds. You reviewed the characteristic properties of ionic and covalent compounds. You considered the periodic nature of electronegativity, and you learned how to use the electronegativity difference to predict the type of bond. You learned, for example, that ionic bonds form between two atoms with very different electronegativities.

In section 3.2, you will explore ionic and covalent bonding in terms of electron transfer and sharing. You will use your understanding of the nature of bonding to explain some properties of ionic and covalent compounds.

Section Review

1 K/U Name the typical properties of an ionic compound. Give two examples of ionic compounds.

2 K/U Name the typical properties of covalent compounds. Give two examples of covalent compounds.

3 C In your own words, describe and explain the periodic trend for electronegativity.

4 K/U Based only on their position in the periodic table, arrange the elements in each set in order of increasing attraction for electrons in a bond.

(a) Li, Br, Zn, La, Si
(b) P, Ga, Cl, Y, Cs

5 K/U Determine ΔEN for each bond. Indicate whether the bond is ionic or covalent.

(a) N—O
(b) Mn—O
(c) H—Cl
(d) Ca—Cl

6 I A chemist analyzes a white, solid compound and finds that it does not dissolve in water. When the compound is melted, it does not conduct electricity.

(a) What would you expect to be true about this compound's melting point?
(b) Are the atoms that make up this compound joined with covalent or ionic bonds? Explain.