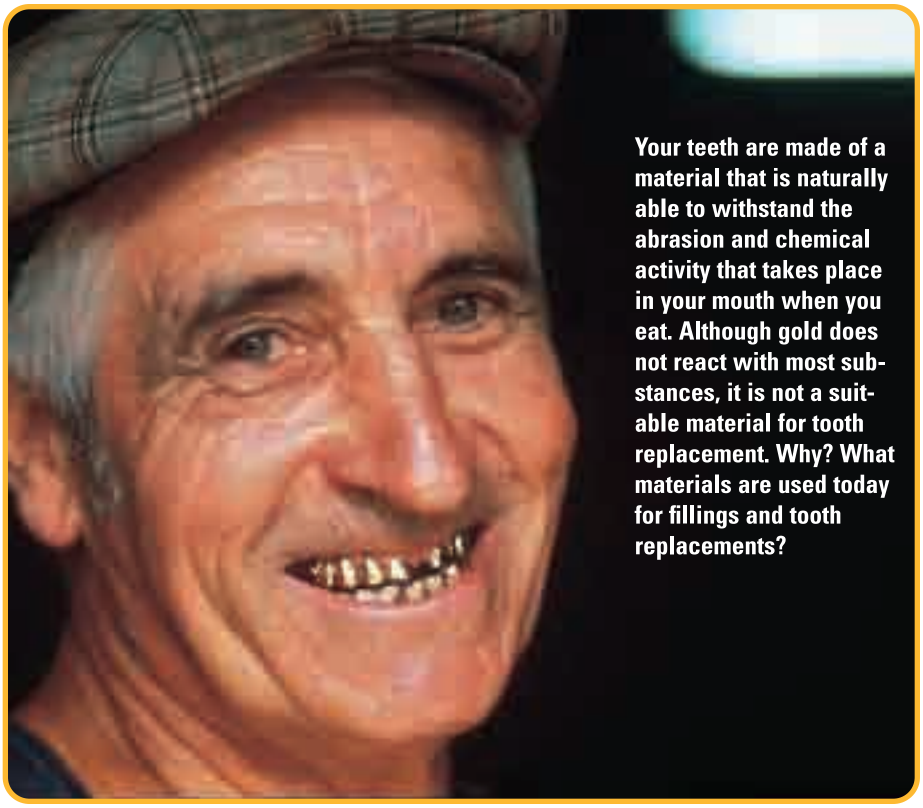
CHAPTER 3

Chemical Compounds and Bonding

The year was 1896. A chance discovery sent a message echoing from Yukon's Far North to the southern reaches of the United States: "Gold!" People migrated in great numbers to the Yukon Territory, hoping to make their fortunes. Within two years, these migrants transformed a small fishing village into bustling Dawson City—one of Canada's largest cities at the time. They also launched the country's first metal-mining industry.

Gold, like all metals, is shiny, malleable, ductile, and a good conductor of electricity and heat. Unlike most metals and other elements, however, gold is found in nature in its pure form, as an element. Most elements are chemically combined in the form of compounds. Why is this so? Why do atoms of some elements join together as compounds, while others do not? In this chapter, you will use the periodic trends you examined in Chapter 2 to help you answer these questions. You will learn about the bonds that hold elements together in compounds. At the same time, you will learn how to write chemical formulas and how to name compounds.



Your teeth are made of a material that is naturally able to withstand the abrasion and chemical activity that takes place in your mouth when you eat. Although gold does not react with most substances, it is not a suitable material for tooth replacement. Why? What materials are used today for fillings and tooth replacements?

Chapter Preview

Concepts and Skills You Will Need

Before you begin this chapter, review the following concepts and skills:

- drawing Lewis structures to represent valence electrons in the outer energy levels of atoms (Chapter 2, section 2.1)
- identifying and explaining periodic trends (Chapter 2, section 2.2)
- identifying elements by name and by symbol (Chapter 2, section 2.2)

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.

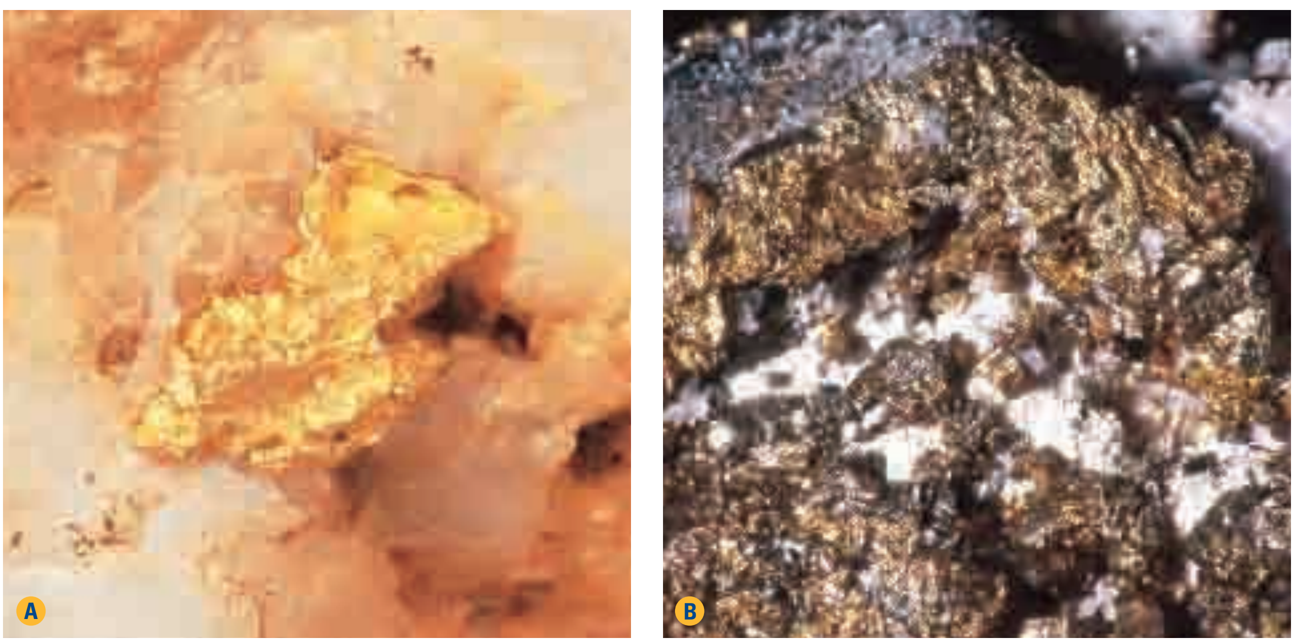


Figure 3.1 **Prospectors used the physical properties of gold and pyrite to distinguish between them. Can you tell which of these photos shows gold and which shows pyrite?**

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**.

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.

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?

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?

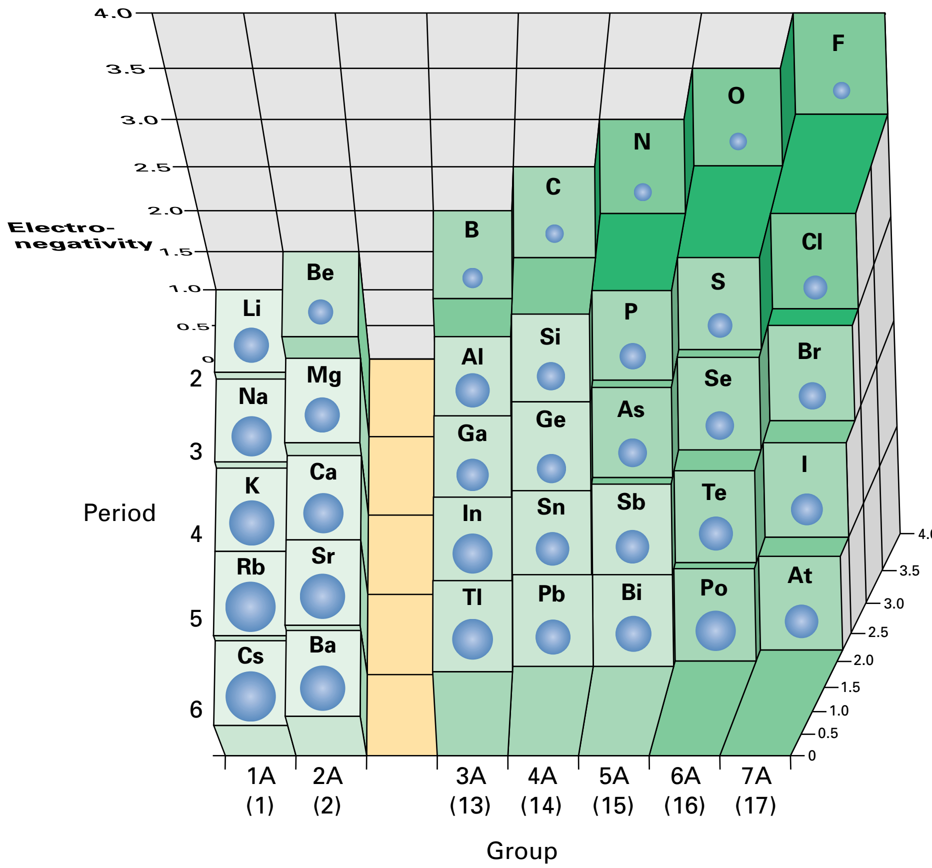


Figure 3.7 **Periodic trends for electronegativity (bars) and atomic size (spheres) are inversely related.**

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.

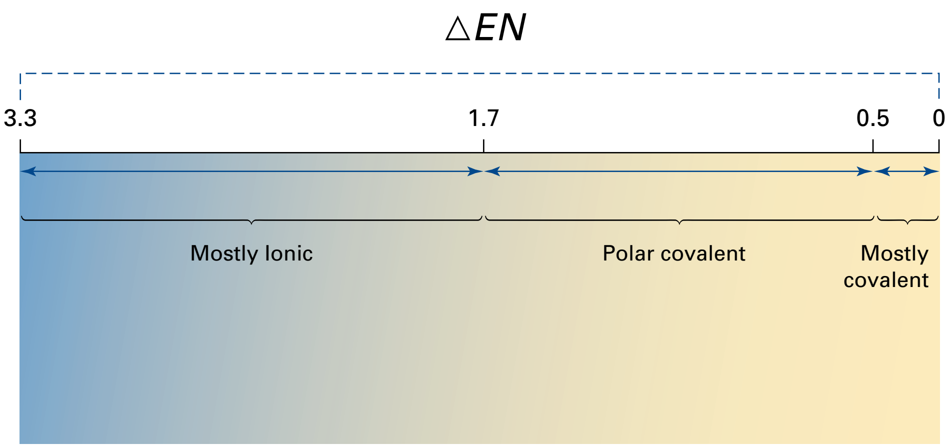


Figure 3.8 **Chemical bonds range in character from mostly ionic to mostly covalent.**

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.

Ionic and Covalent Bonding: The Octet Rule

3.2

In section 3.1, you reviewed your understanding of the physical properties of covalent and ionic compounds. You learned how to distinguish between an ionic bond and a covalent bond based on the difference between the electronegativities of the atoms. By considering what happens to electrons when atoms form bonds, you will be able to explain some of the characteristic properties of ionic and covalent compounds.

Section Preview/ Specific Expectations

In this section, you will

- **demonstrate** an understanding of the formation of ionic and covalent bonds, and **explain** the properties of the products
- **explain** how different elements combine to form covalent and ionic bonds, using the octet rule
- **represent** the formation of ionic and covalent bonds using diagrams
- **communicate** your understanding of the following terms: *octet rule, isoelectronic, pure covalent bond, diatomic elements, double bond, triple bond, molecular compounds, intramolecular forces, intermolecular forces, metallic bond, alloy*

The Octet Rule

Why do atoms form bonds? When atoms are bonded together, they are often more stable. We know that noble gases are the most stable elements in the periodic table. What evidence do we have? The noble gases are extremely unreactive. They do not tend to form compounds. What do the noble gases have in common? They have a filled outer electron energy level. When an atom loses, gains, or shares electrons through bonding to achieve a filled outer electron energy level, the resulting compound is often very stable.

According to the **octet rule**, atoms bond in order to achieve an electron configuration that is the same as the electron configuration of a noble gas. When two atoms or ions have the same electron configuration, they are said to be **isoelectronic** with one another. For example, Cl^- is isoelectronic with Ar because both have 18 electrons and a filled outer energy level. This rule is called the octet rule because all the noble gases (except helium) have eight electrons in their filled outer energy level. (Recall that helium's outer electron energy level contains only two electrons.)

Ionic Bonding

In Section 3.1 you learned that the electronegativity difference for the bond between sodium and chlorine is 2.1. Thus, the bond is an ionic bond. Sodium has a very low electronegativity, and chlorine has a very high electronegativity. Therefore, when sodium and chlorine interact, sodium transfers its valence electron to chlorine. As shown in Figure 3.9, sodium becomes Na^+ and chlorine becomes Cl^-.

How does the formation of an ionic bond between sodium and chlorine reflect the octet rule? Neutral sodium has one valence electron. When it loses this electron to chlorine, the resulting Na^+ cation has an electron energy level that contains eight electrons. It is isoelectronic with the noble gas neon. On the other hand, chlorine has an outer electron energy level that contains seven electrons. When chlorine gains sodium's

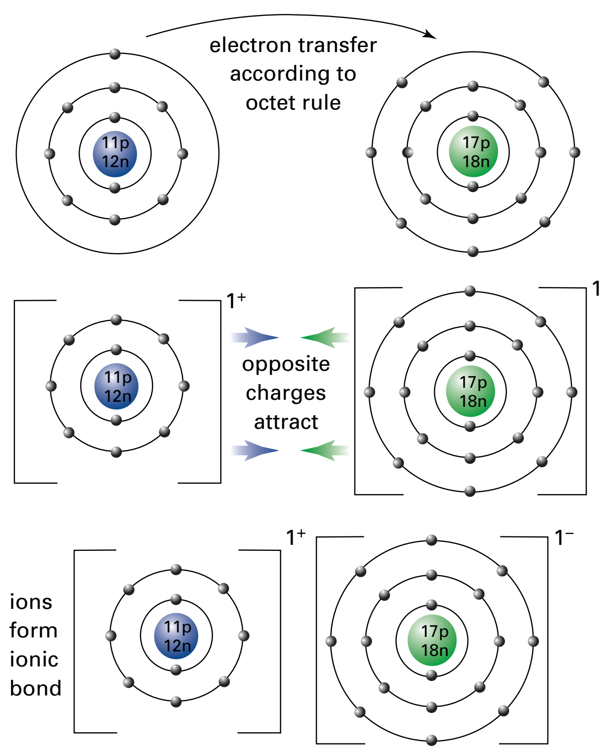


Figure 3.9 Sodium's electron is transferred to chlorine. The atoms become oppositely charged ions with stable octets. Because they are oppositely charged, they are strongly attracted to one another.

electron, it becomes an anion that is isoelectronic with the noble gas argon. As you can see in Figure 3.10, you can represent the formation of an ionic bond using Lewis structures.

Thus, in an ionic bond, electrons are transferred from one atom to another so that they form oppositely charged ions. The strong force of attraction between the oppositely charged ions is what holds them together.

Figure 3.10 These Lewis structures show the formation of a bond between a sodium atom and a chlorine atom.

PROBLEM TIP

When you draw Lewis structures to show the formation of a bond, you can use different colours or symbols to represent the electrons from different atoms. For example, use an "x" for an electron from sodium, and an "o" for an electron from chlorine. Or, use open and closed cirlces as is shown here. This will make it easier to see how the electrons have been transferred.

Transferring Multiple Electrons

In sodium chloride, NaCl, one electron is transferred from sodium to chlorine. In order to satisfy the octet rule, two or three electrons may be transferred from one atom to another. For example, consider what happens when magnesium and oxygen combine.

The electronegativity difference for magnesium oxide is $3.4 - 1.3 = 2.1$. Therefore, magnesium oxide is an ionic compound. Magnesium contains two electrons in its outer shell. Oxygen contains six electrons in its outer shell. In order to become isoelectronic with a noble gas, magnesium needs to lose two electrons and oxygen needs to gain two electrons. Hence, magnesium transfers its two valence electrons to oxygen, as shown in Figure 3.11. Magnesium becomes Mg^{2+}, and oxygen becomes O^{2-}.

Figure 3.11 These Lewis structures show the formation of a bond between a magnesium atom and an oxygen atom.

Try the following problems to practise representing the formation of ionic bonds between two atoms.

Practice Problems

2. For each bond below, determine ΔEN. Is the bond ionic or covalent?

(a) Ca—O
(b) K—Cl
(c) K—F
(d) Li—F
(e) Li—Br
(f) Ba—O

3. Draw Lewis structures to represent the formation of each bond in question 2.

Careers in Chemistry

Metallurgist

Alison Dickson

Alison Dickson is a metallurgist at Polaris, the world's northernmost mine. Polaris is located on Little Cornwallis Island in Nunavut. It is a lead and zinc mine, operated by Cominco Ltd., the world's largest producer of zinc concentrate.

After ore is mined at Polaris, metallurgists must separate the valuable lead- and zinc-bearing compounds from the waste or "slag." First the ore is crushed and ground with water to produce flour-like particles. Next a process called *flotation* is used to separate the minerals from the slag. In flotation, chemicals are added to the metal-containing compounds. The chemicals react with the lead and zinc to make them very insoluble in water, or *hydrophobic*. Air is then bubbled through the mineral and water mixture. The hydrophobic particles attach to the bubbles and float to the surface. They form a stable froth, or concentrate, which is collected. The concentrate is filtered and dried, and then stored for shipment.

Dickson says that she decided on metallurgy as a career because she wanted to do something that was "hands-on." After completing her secondary education in Malaysia, where she grew up, Dickson moved to Canada. She studied mining and mineral process engineering at the University of British Columbia.

Dickson says that she also wanted to do something adventurous. She wanted to travel and live in different cultures. As a summer student, Dickson worked at a Chilean copper mine. Her current job with Cominco involves frequent travel to various mines. "Every day provides a new challenge," Dickson says. When she is at Polaris, Dickson enjoys polar bear sightings on the tundra.

Making Career Connections

Are you interested in a career in mining and metallurgy? Here are two ways that you can get information:

1. Explore the web site of The Canadian Institute of Mining, Metallurgy and Petroleum. Go to **www.school.mcgrawhill.ca/resources/**, to Science Resources, then to Chemistry 11 to know where to go next. It has a special section for students who are interested in mining and metallurgy careers. This section lists education in the field, scholarships and bursaries, and student societies for mining and metallurgy.
2. To discover the variety of jobs that are available for metallurgists, search for careers at Infomine. Go to **www.school.mcgrawhill.ca/resources/**, to Science Resources, then to Chemistry 11 to know where to go next. Many of the postings are for jobs overseas.

Ionic Bonding That Involves More Than Two Ions

Sometimes ionic compounds contain more than one atom of each element. For example, consider the compound that is formed from calcium and fluorine. Because the electronegativity difference between calcium and fluorine is 3.0, you know that a bond between calcium and fluorine is ionic. Calcium has two electrons in its outer energy level, so it needs to lose two electrons according to the octet rule. Fluorine has seven electrons in its outer energy level, so it needs to gain one electron, again according to the octet rule. How do the electrons of these elements interact so that each element achieves a filled outer energy level?

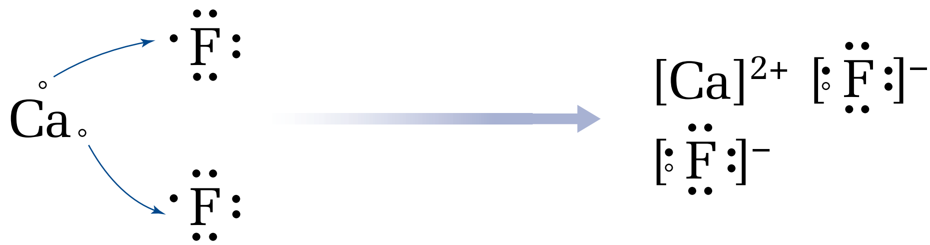


Figure 3.12 **These Lewis structures show the formation of bonds between one atom of calcium and two atoms of fluorine.**

Examine Figure 3.11. In an ionic bond, calcium tends to lose two electrons and fluorine tends to gain one electron. Therefore, one calcium atom bonds with two fluorine atoms. Calcium loses one of each of its valence electrons to each fluorine atom. Calcium becomes Ca^{2+}, and fluorine becomes F^{-}. They form the compound calcium fluoride, CaF_2.

In the following Practice Problems, you will predict the kind of ionic compound that will form from two elements.

Electronic Learning Partner

Your Chemistry 11 Electronic Learning Partner has an interactive simulation on forming ionic compounds.

Practice Problems

4. For each pair of elements, determine ΔEN.

- **(a)** magnesium and chlorine
- **(b)** calcium and chlorine
- **(c)** lithium and oxygen
- **(d)** sodium and oxygen
- **(e)** potassium and sulfur
- **(f)** calcium and bromine

5. Draw Lewis structures to show how each pair of elements in question 4 forms bonds to achieve a stable octet.

Explaining the Conductivity of Ionic Compounds

Now that you understand the nature of the bonds in ionic compounds, can you explain some of their properties? Consider electrical conductivity. Ionic compounds do not conduct electricity in their solid state. They are very good conductors in their liquid state, however, or when they are dissolved in water. To explain these properties, ask yourself two questions:

1. What is required for electrical conductivity?
2. What is the structure of ionic compounds in the liquid, solid, and dissolved states?

An electrical current can flow only if charged particles are available to move and carry the current. Consider sodium chloride as an example. Is there a mobile charge in solid sodium chloride? No, there is not. In the solid state, sodium and chlorine ions are bonded to each other by strong ionic bonds. Like all solid-state ionic compounds, the ions are arranged in a rigid lattice formation, as shown in Figure 3.13.

In the solid state, the ions cannot move very much. Thus, there is no mobile charge. Solid sodium chloride does not conduct electricity.

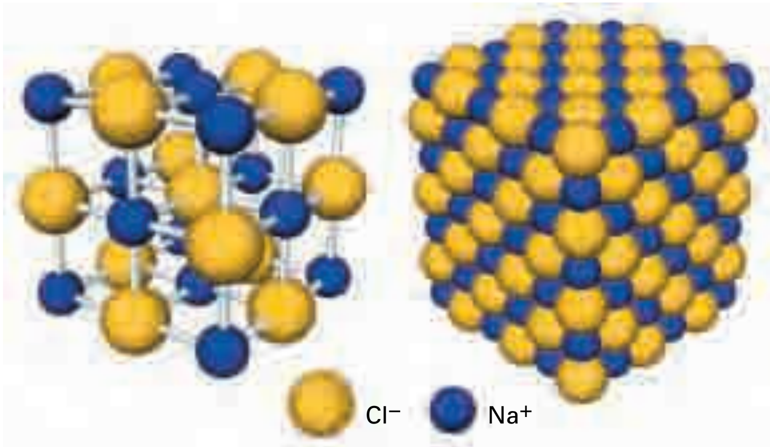


Figure 3.13 **In solid sodium chloride, NaCl, sodium and chlorine are arranged in a rigid lattice pattern.**

In molten sodium chloride, the rigid lattice structure is broken. The ions that make up the compound are free to move, and they easily conduct electricity. Similarly, when sodium chloride is dissolved, the sodium and chlorine ions are free to move. The solution is a good conductor of electricity, as shown in Figure 3.14. You will learn more about ionic compounds in solution in Chapter 9.

mind STRETCH

Go back to Table 3.1. What other properties of ionic compounds can you now explain with your new understanding of ionic bonding?

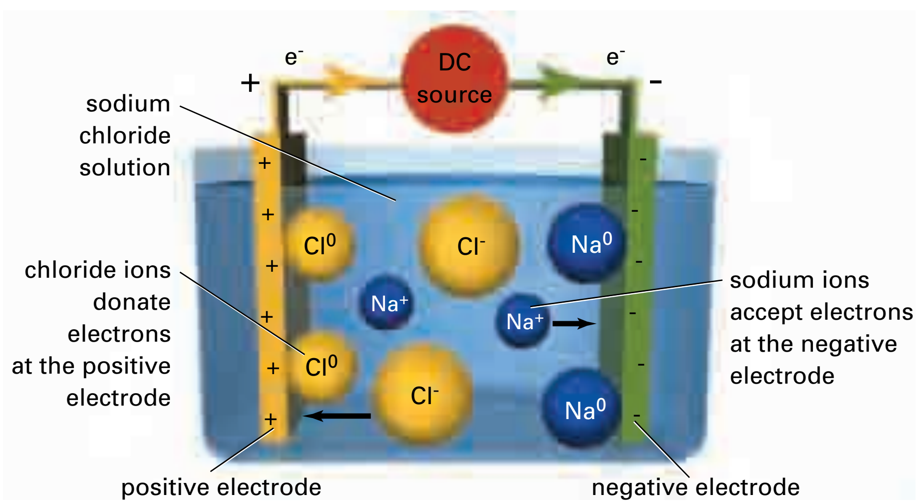


Figure 3.14 **Aqueous sodium chloride is a good conductor of electricity.**

You are probably familiar with the ionic crystals in caves. Stalagmites and stalactites are crystal columns that form when water, containing dissolved lime, drips very slowly from the ceiling of a cave onto the floor below. How do these ionic crystals grow?

When a clear solution of an ionic compound is poured over a seed crystal of the same compound, the ions align themselves according to the geometric arrangement in the seed crystal. You will observe this for yourself in Investigation 3-A.

SKILL FOCUS

Performing and recording

Analyzing and interpreting

Investigation 3-A

Crystalline Columns

In this investigation, you will prepare a supersaturated solution of sodium acetate. (A supersaturated solution contains more dissolved solute at a specific temperature than is normally possible.) Then you will use the solution to prepare your own ionic crystal.

Question

How can you build a crystal column on your laboratory bench?

Prediction

If the solution drips slowly enough, a tall crystal column will form.

Safety Precautions

Materials

water
sodium acetate trihydrate crystals
balance
10 mL graduated cylinder
100 mL Erlenmeyer flask
hot plate
squirt bottle
burette
burette stand
forceps

Procedure

1. Place 50 g of sodium acetate trihydrate in a clean 100 mL Erlenmeyer flask.
2. Add 5 mL of water. Heat the solution slowly.
3. Swirl the flask until the solid completely dissolves. If any crystals remain inside the flask or on the neck, wash them down with a small amount of water.
4. Remove the flask from the heat. Pour the solution into a clean, dry burette.
5. Raise the burette as high as it will go. Place it on the lab bench where you intend to grow your crystal column.
6. Pour some sodium acetate trihydrate crystals onto the lab bench. Using clean and dry forceps, choose a relatively large crystal (the seed crystal). Place it directly underneath the burette spout.
7. Turn the buret stopcock slightly so that the solution drips out slowly. Adjust the position of the seed crystal so that the drops fall on it. (You can drip the solution right onto the bench, or onto a glass plate if you prefer.)
8. Observe the crystal for 10 min. Record your observations about the crystal column or your apparatus. Continue to make observations every 10 min.

Analysis

1. Describe your observations.
2. Why does the column form upward?
3. What was the purpose of the seed crystal?
4. What improvements would you suggest for better results in the future?

Conclusions

5. What kind of change is taking place when a crystal forms? Is the change chemical or physical? Explain.

Application

6. Repeat steps 1 to 3 in the Procedure. Once you remove the solution from the heat, seal the flask with a clean, dry rubber stopper. Allow the flask to cool to room temperature. Next, remove the stopper and carefully add only one crystal of sodium acetate trihydrate to the flask. Record your observations, and explain what is happening.

Covalent Bonding

You have learned what happens when the electronegativity difference between two atoms is greater than 1.7. The atom with the lower electronegativity transfers its valence electron(s) to the atom with the higher electronegativity. The resulting ions have opposite charges. They are held together by a strong ionic bond.

What happens when the electronegativity difference is very small? What happens when the electronegativity difference is zero? As an example, consider chlorine. Chlorine is a yellowish, noxious gas. What is it like at the atomic level? Each chlorine atom has seven electrons in its outer energy level. In order for chlorine to achieve the electron configuration of a noble gas according to the octet rule, it needs to gain one electron. When two chlorine atoms bond together, their electronegativity difference is zero. The electrons are equally attracted to each atom.

Therefore, instead of transferring electrons, the two atoms each share one electron with each other. In other words, each atom contributes one electron to a covalent bond. *A covalent bond consists of a pair of shared electrons*. Thus, each chlorine atom achieves a filled outer electron energy level, satisfying the octet rule. Examine Figure 3.15 to see how to represent a covalent bond with a Lewis structure.

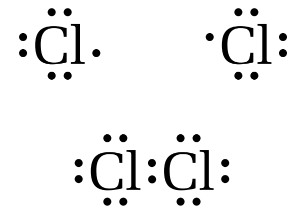


Figure 3.15 These Lewis structures show the formation of a bond between two atoms of chlorine.

When two atoms of the same element form a bond, they share their electrons equally in a **pure covalent bond**. Elements with atoms that bond to each other in this way are known as **diatomic elements**.

CHEM FACT

Some examples of diatomic elements are chlorine, Cl_2, bromine, Br_2, iodine, I_2, nitrogen, N_2, and hydrogen, H_2.

When atoms such as carbon and hydrogen bond to each other, their electronegativities are so close that they share their electrons almost equally. Carbon and hydrogen have an electronegativity difference of only $2.6 - 2.2 = 0.4$. In Figure 3.16, you can see how one atom of carbon forms a covalent bond with four atoms of hydrogen. The compound methane, CH_4, is formed.

Each hydrogen atom shares one of its electrons with the carbon. The carbon shares one of its four valence electrons with each hydrogen. Thus, each hydrogen atom achieves a filled outer energy level, and so does carbon. (Recall that elements in the first period need only two electrons to fill their outer energy level.) *When analyzing Lewis structures that show covalent bonds, count the shared electrons as if they belong to each of the bonding atoms*. In the following Practice Problems, you will represent covalent bonding using Lewis structures.

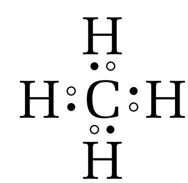


Figure 3.16 This Lewis structure shows a molecule of methane, CH_4.

Practice Problems

6. Show the formation of a covalent bond between two atoms of each diatomic element.

(a) iodine
(b) bromine
(c) hydrogen
(d) fluorine

7. Use Lewis structures to show the simplest way in which each pair of elements forms a covalent compound, according to the octet rule.

(a) hydrogen and oxygen
(b) chlorine and oxygen
(c) carbon and hydrogen
(d) iodine and hydrogen
(e) nitrogen and hydrogen
(f) hydrogen and rubidium

Electronic Learning Partner

Your Chemistry 11 Electronic Learning Partner has several animations that show ionic and covalent bonding.

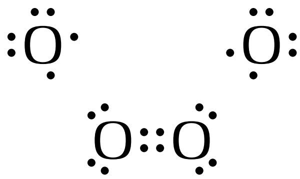


Figure 3.17 These Lewis structures show the formation of a double bond between two atoms of oxygen.

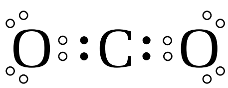


Figure 3.18 This Lewis structure shows the double bond in a molecule of carbon dioxide, CO_2.

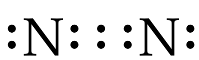


Figure 3.19 This Lewis structure shows the triple bond in a molecule of nitrogen, N_2.

PROBLEM TIP

When drawing Lewis structures to show covalent bonding, you can use lines between atoms to show the bonding pairs of electrons. One line (–) signifies a single bond. Two lines (=) signify a double bond. Three lines (≡) signify a triple bond. Non-bonding pairs are shown as dots in the usual way.

Multiple Covalent Bonds

Atoms sometimes transfer more than one electron in ionic bonding. Similarly, in covalent bonding, atoms sometimes need to share two or three pairs of electrons, according to the octet rule. For example, consider the familiar diatomic element oxygen. Each oxygen atom has six electrons in its outer energy level. Therefore, each atom requires two additional electrons to achieve a stable octet. When two oxygen atoms form a bond, they share two pairs of electrons, as shown in Figure 3.17. This kind of covalent bond is called a **double bond**.

Double bonds can form between different elements, as well. For example, consider what happens when carbon bonds to oxygen in carbon dioxide. To achieve a stable octet, carbon requires four electrons, and oxygen requires two electrons. Hence, two atoms of oxygen bond to one atom of carbon. Each oxygen forms a double bond with the carbon, as shown in Figure 3.18.

When atoms share three pairs of electrons, they form a **triple bond**. Diatomic nitrogen contains a triple bond, as you can see in Figure 3.19. Try the following problems to practise representing covalent bonding using Lewis structures. Watch for multiple bonding!

Practice Problems

8. One carbon atom is bonded to two sulfur atoms. Use a Lewis structure to represent the bonds.
9. A molecule contains one hydrogen atom bonded to a carbon atom, which is bonded to a nitrogen atom. Use a Lewis structure to represent the bonds.
10. Two carbon atoms and two hydrogen atoms bond together, forming a molecule. Each atom achieves a full outer electron level. Use a Lewis structure to represent the bonds.

Explaining the Low Conductivity of Covalent Compounds

Covalent compounds have a wider variety of properties than ionic compounds. Some dissolve in water, and some do not. Some conduct electricity when molten or dissolved in water, and some do not. If you consider only covalent compounds that contain bonds with an electronegativity difference that is less than 0.5, you will notice greater consistency. For example, consider the compounds carbon disulfide, CS_2, dichlorine monoxide, Cl_2O, and carbon tetrachloride, CCl_4. What are some of the properties of these compounds? They all have low boiling points. None of them conducts electricity in the solid, liquid, or gaseous state.

How do we explain the low conductivity of these pure covalent compounds? The atoms in each compound are held together by strong covalent bonds. Whether the compound is in the liquid, solid, or gaseous state, these bonds do not break. Thus, covalent compounds (unlike ionic compounds) do not break up into ions when they melt or boil. Instead, their atoms remain bonded together as molecules. For this reason, covalent compounds are also called **molecular compounds**. The molecules that make up a pure covalent compound cannot carry a current, even if the compound is in its liquid state or in solution.

Evidence for Intermolecular Forces

You have learned that pure covalent compounds are not held together by ionic bonds in lattice structures. They do form liquids and solids at low temperatures, however. Something must hold the molecules together when a covalent compound is in its liquid or solid state. The forces that bond the *atoms* to each other within a molecule are called **intramolecular forces**. Covalent bonds are intramolecular forces. In comparison, the forces that bond *molecules* to each other are called **intermolecular forces**.

You can see the difference between intermolecular forces and intramolecular forces in Figure 3.20. Because pure covalent compounds have low melting and boiling points, you know that the intermolecular forces must be very weak compared with the intramolecular forces. It does not take very much energy to break the bonds that hold the molecules to each other.

There are several different types of intermolecular forces. You will learn more about them in section 3.3, as well as in Chapters 8 and 11.

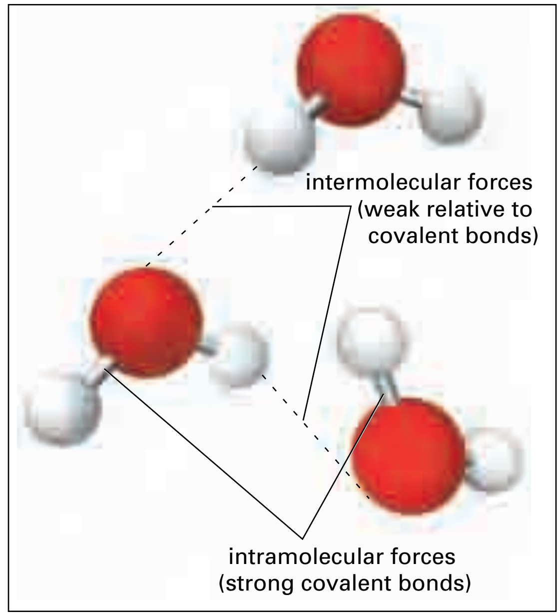


Figure 3.20 **Strong intramolecular forces (covalent bonds) hold the atoms in molecules together. Relatively weak intermolecular forces act between molecules.**

Metallic Bonding

In this chapter, you have seen that non-metals tend to form ionic bonds with metals. Non-metals tend to form covalent bonds with other non-metals and with themselves. How do metals bond to each other?

We know that elements that tend to form ionic bonds have very different electronegativities. Metals bonding to themselves or to other metals do not have electronegativity differences that are greater than 1.7. Therefore, metals probably do not form ionic bonds with each other.

Evidence bears this out. A pure metal, such as sodium, is soft enough to be cut with a butter knife. Other pure metals, such as copper or gold, can be drawn into wires or hammered into sheets. Ionic compounds, by contrast, are hard and brittle.

Do metals form covalent bonds with each other? No. They do not have enough valence electrons to achieve stable octets by sharing electrons. Although metals do not form covalent bonds, however, they do share their electrons.

In metallic bonding, atoms release their electrons to a shared pool of electrons. You can think of a metal as a non-rigid arrangement of metal ions in a sea of free electrons, as shown in Figure 3.21. The force that holds metal atoms together is called a **metallic bond**. Unlike ionic or covalent bonding, metallic bonding does not have a particular orientation in space. Because the electrons are free to move, the metal ions are not rigidly held in a lattice formation. Therefore, when a hammer pounds metal, the atoms can slide past one another. This explains why metals can be easily hammered into sheets.

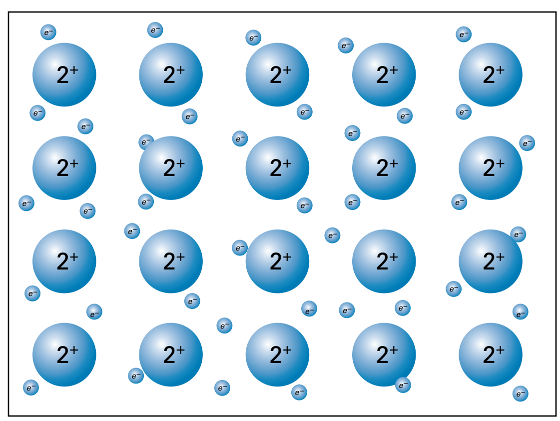


Figure 3.21 **In magnesium metal, the two valence electrons from each atom are free to move in an "electron sea." The valence electrons are shared by all the metal ions.**

Pure metals contain metallic bonds, as do alloys. An **alloy** is a homogeneous mixture of two or more metals. Different alloys can have different amounts of elements. Each alloy, however, has a uniform composition throughout. One example of an alloy is bronze. Bronze contains copper, tin, and lead, joined together with metallic bonds. You will learn more about alloys in Chapter 4 and Chapter 8.

Section Wrap-up

In this section, you learned how to distinguish between an ionic bond and a pure covalent bond. You learned how to represent these bonds using Lewis structures. You were also introduced to metallic bonding.

In section 3.3, you will learn about "in between" covalent bonds with ΔEN greater than 0.5 but less than 1.7. You will learn how the nature of these bonds influences the properties of the compounds that contain them. As well, you will examine molecules in greater depth. You will explore ways to visualize them in three dimensions, which will help you further understand the properties of covalent compounds.

Section Review

1 K/U Use Lewis structures to show how each pair of elements forms an ionic bond.

(a) magnesium and fluorine

(b) potassium and bromine

(c) rubidium and chlorine

(d) calcium and oxygen

2 K/U Use Lewis structures to show how the following elements form covalent bonds.

(a) one silicon atom and two oxygen atoms

(b) one carbon atom, one hydrogen atom, and three chlorine atoms

(c) two nitrogen atoms

(d) two carbon atoms bonded to each other—three hydrogen atoms bonded to one of the carbon atoms, and one hydrogen atom and one oxygen atom bonded to the other carbon atom

3 K/U Use what you know about electronegativity differences to decide what kind of bond would form between each pair of elements.

(a) palladium and oxygen

(b) carbon and bromine

(c) silver and sulfur

(d) sodium and iodine

(e) beryllium and fluorine

(f) phosphorus and calcium

4 C "In general, the farther away two elements are from each other in the periodic table, the more likely they are to participate in ionic bonding." Do you agree with this statement? Explain why or why not.

5 C Covalent bonding and metallic bonding both involve electron sharing. Explain how covalent bonding is different from metallic bonding.

6 MC Ionic compounds are extremely hard. They hold their shape extremely well.

(a) Based on what you know about ionic bonding within an ionic crystal, explain these properties.

(b) Give two reasons to explain why, in spite of these properties, it is not practical to make tools out of ionic compounds.

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.

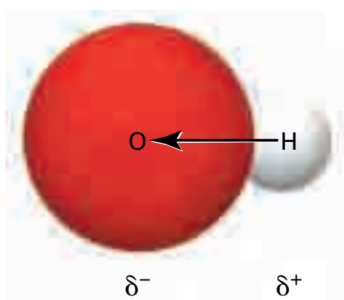


Figure 3.24 The O end of an O—H bond has a partial negative charge. The H end has a partial positive charge.

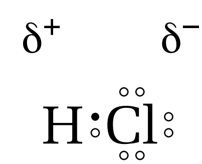


Figure 3.25 The Cl end of a H—Cl bond has a partial negative charge. The H end has a partial positive charge.

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 **(b)** O—N **(c)** Cl—Cl **(d)** Cu—O **(e)** Si—H **(f)** Na—F **(g)** Fe—O **(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.

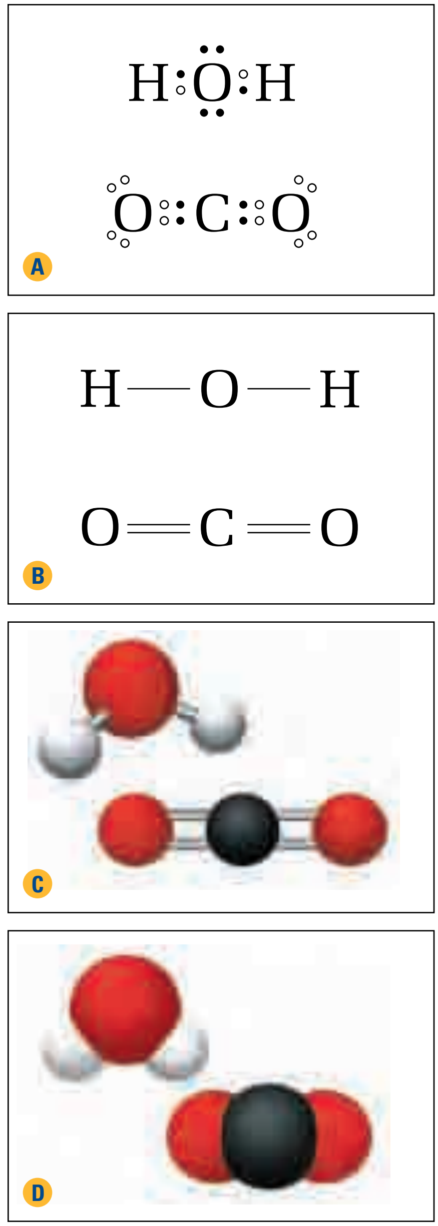


Figure 3.26 **You can compare a molecule of water with a molecule of carbon dioxide using a variety of different models.**

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.

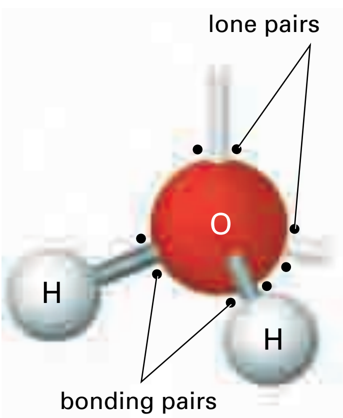


Figure 3.28 Two non-bending pairs account for water's bent shape.

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.

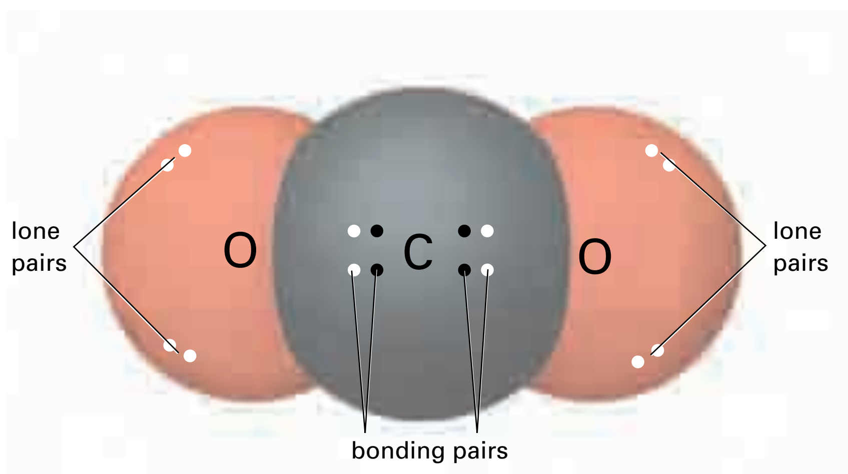


Figure 3.29 Carbon dioxide is linear in shape.

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.

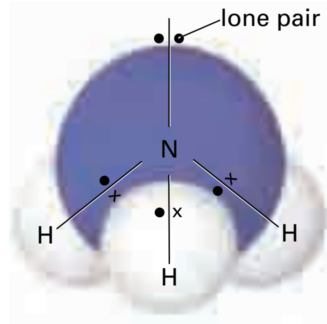


Figure 3.30 **An ammonia molecule is shaped like a pyramid.**

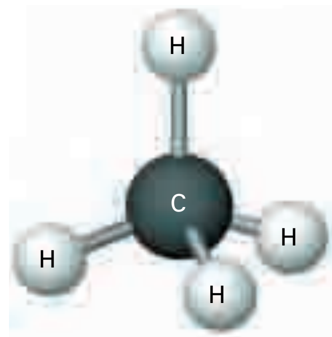


Figure 3.31 **A methane molecule is shaped as though its hydrogen atoms were on the corners of a tetrahedron.**

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.

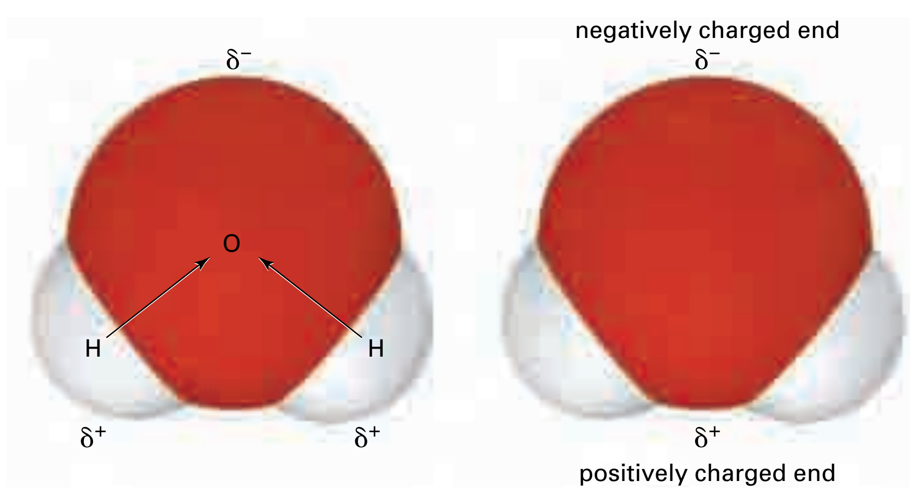


Figure 3.33 **Water is a polar molecule because of its shape and the polarity of its bonds.**

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.

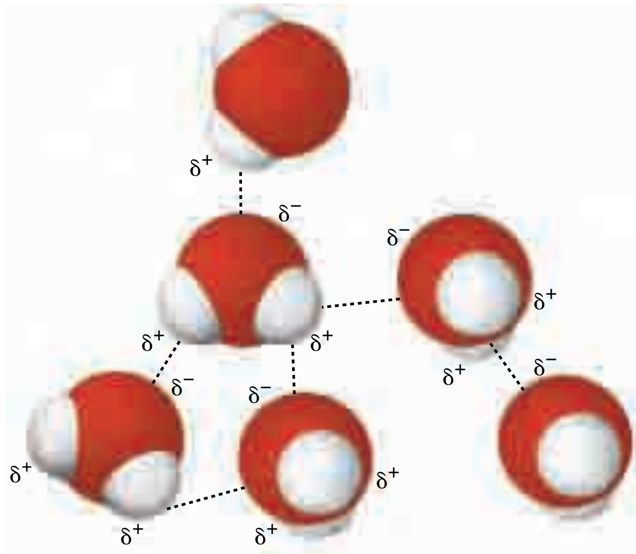


Figure 3.34 The negative ends of water molecules attracts the positive ends. Some of the resulting intermolecular forces are shown here.

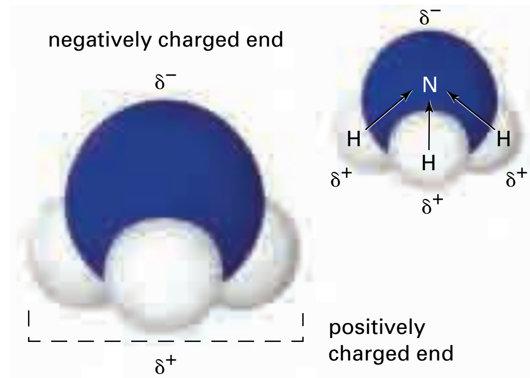


Figure 3.35 Because an ammonia molecule contains polar bonds and is asymmetrical, it is a polar molecule.

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.

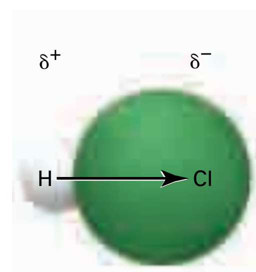


Figure 3.36 Hydrogen chloride contains one polar bond. Therefore, the molecule is polar.

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.

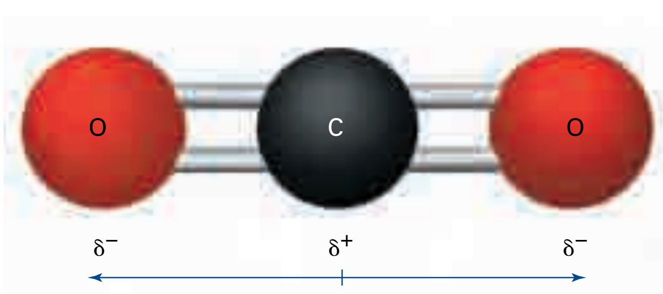


Figure 3.37 Carbon dioxide is a non-polar molecule because it is symmetrical.

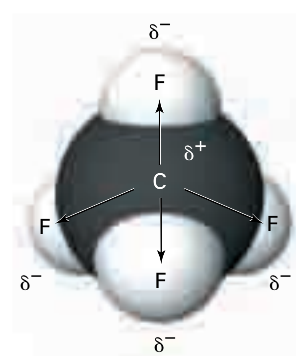


Figure 3.38 Carbon tetrafluoride, CF_4, contains four polar bonds. Because of its symmetry, however, it is a non-polar molecule.

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

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.

Writing Chemical Formulas and Naming Chemical Compounds

3.4

Section Preview/ Specific Expectations

In this section, you will

- **write** the formulas of binary and tertiary compounds, including compounds that contain elements with multiple valences
- **communicate** formulas using IUPAC and traditional systems
- **recognize** the formulas of compounds in various contexts
- **communicate** your understanding of the following terms: *chemical formula, valence, polyatomic ions, zero sum rule, chemical nomenclature, binary compound, Stock system, tertiary compounds*

You have used Lewis structures to demonstrate how ionic and covalent bonds form between atoms. When given two elements, you determined how many atoms of each element bond together to form a compound, according to the octet rule. For example, you used the periodic table and your understanding of the octet rule to determine how calcium and bromine bond to form an ionic compound. Using a Lewis structure, you determined that calcium and bromine form a compound that contains two bromine atoms for every calcium atom, as shown in Figure 3.39.

Chemical Formulas

Lewis structures are helpful for keeping track of electron transfers in bonding and for making sure that the octet rule is obeyed. As well, Lewis structures can be used to help determine the ratio of the atoms in a compound. To communicate this ratio, chemists use a special kind of shorthand called a **chemical formula**. A chemical formula provides two important pieces of information:

1. the elements that make up the compound
2. the number of atoms of each element that are present in a compound

The order in which the elements are written also communicates important information. The less electronegative element or ion is usually listed first in the formula, and the more electronegative element or ion comes second. For example, the ionic compound that is formed from calcium and bromine is written $CaBr_2$. Calcium, a metal with low electronegativity, is written first. The subscript 2 after the bromine indicates that there are two bromine atoms for every calcium atom.

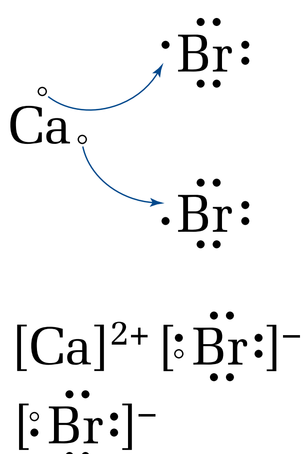


Figure 3.39 These Lewis structures show the formation of calcium bromide.

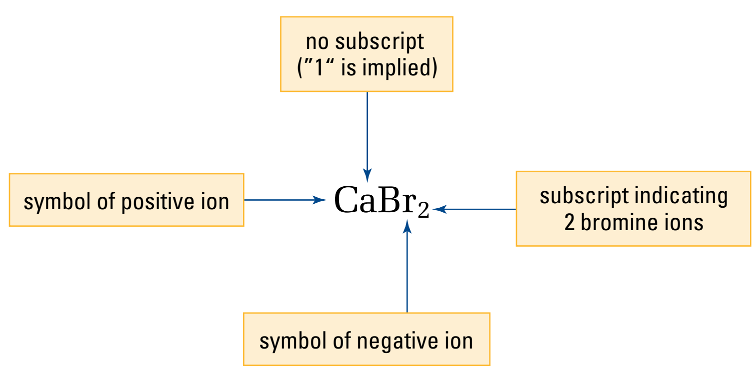


Figure 3.40 $CaBr_2$ is the chemical formula of the compound formed by calcium and bromine. When a subscript is omitted, only one atom is present per formula unit.

What a Chemical Formula Represents

For covalent compounds, the chemical formula represents how many of each type of atom are in each molecule. For example, the formula NH_3 signifies that a molecule of ammonia contains one nitrogen atom and three hydrogen atoms. The formula C_2H_6 tells you that a molecule of propane contains two atoms of carbon and six atoms of hydrogen.

For ionic compounds, the formula represents *a ratio rather than a discrete particle*. For example, the formula for magnesium oxide, MgO, signifies that magnesium and oxygen exist in a one-to-one atomic ratio. Recall that MgO exists in a lattice structure held together by ionic bonds, as shown in Figure 3.41. The formula MgO represents the ratio in which ions are present in the compound.

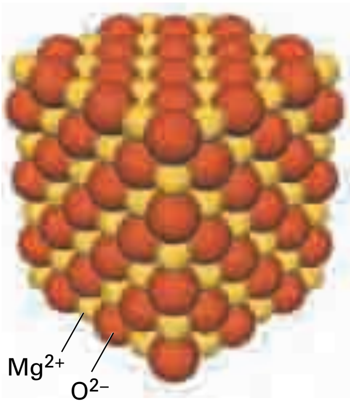


Figure 3.41 In a crystal of magnesium oxide, MgO, magnesium and oxygen atoms exist in a 1:1 ratio.

CHECKPOINT

What does the formula of calcium bromide represent?

Using Valence Numbers to Describe Bonding Capacity

You have seen how Lewis structures can help you draw models of ionic, covalent, and polar covalent compounds. When you draw a Lewis structure, you can count how many electrons are needed by each atom to achieve a stable octet. Thus, you can find out the ratio in which the atoms combine. Once you know the ratio of the atoms, you can write the chemical formula of the compound. Drawing Lewis structures can become overwhelming, however, when you are dealing with large molecules. Is there a faster and easier method for writing chemical formulas?

Every element has a certain capacity to combine with other atoms. An atom of a Group 1 (IA) element, for example, has the capacity to lose one electron from its valence level in order to bond with another atom. A number is assigned to each element to describe the element's bonding capacity. This number is called the **valence**. Thus, Group 1 (IA) elements, such as sodium and lithium, have a valence of +1. The 1 indicates that these elements tend to have one electron involved in bonding. This makes sense, because Group 1 elements have only one electron in their outer electron energy level. The + indicates that these elements tend to give up their electrons, becoming positively charged ions. They may transfer their electrons, or they may attract the electron relatively weakly in a polar covalent bond.

On the other hand, Group 17 (VIIA) elements (the halogens) have a valence of –1. Again, the 1 indicates that these elements tend to have one electron involved in bonding. However, they need to *gain* an electron to achieve a stable octet. In general, halogens become more negatively charged when they participate in bonding.

As a general rule, if two atoms form an ionic bond, the valence tells you the charges on the ions that are formed. If a covalent bond is formed, the valence tells you how many electrons the atoms contribute to the covalent bond.

You can use the periodic table to predict valence numbers. For example, Group 2 (IIA) elements have two electrons in their outer energy level. To achieve a stable octet, they need to lose these two electrons. Therefore, the valence for all Group 2 elements is +2.

Practice Problems

14. Use the periodic table to predict the most common valences of the atoms in Groups 16 (VIA) and 17 (VIIA).

15. If you had to assign a valence to the noble gases, what would it be? Explain your answer.

The smaller atoms of elements in the first two periods usually have only one common valence, which is easily determined from the periodic table. Many larger elements, however, have more than one valence because the electron distribution in these elements is much more complex. Therefore, you will have to memorize the valences of the elements that are commonly used in this course. Some useful valences are listed in Table 3.3, with the most common valences listed first.

Table 3.3 Common Valences of Selected Elements

1	2	3	4	5	6	7	8	9	10	11	12	13	14	15	16	17
H(+1)																
Li(+1)	Be(+2)													N(−3)	O(−2)	F(−1)
Na(+1)	Mg(+2)													P(−3)	S(−2)	Cl(−1)
K(+1)	Ca(+2)				Cr(+3) Cr(+2) Cr(+6)		Fe(+3) Fe(+2)	Co(+2) Co(+3)	Ni(+2) Ni(+3)	Cu(+2) Cu(+1)	Zn(+2)	Ga(+3)				Br(−1)
Rb(+1)	Sr(+2)									Ag(+1)	Cd(+2)		Sn(+4) Sn(+2)			I(−1)
Cs(+1)	Ba(+2)									Au(+3) Au(+1)	Hg(+2) Hg(+1)		Pb(+2) Pb(+4)			

Polyatomic Ions

Some compounds contain ions that are made from more than one atom. These ions are called **polyatomic ions**. (The prefix *poly* means "many.") Calcium carbonate, $CaCO_3$, which is found in chalk, contains one calcium cation and one polyatomic anion called carbonate, CO_3^{2-}.

Polyatomic ions are in fact *charged molecules*. For example, the carbonate ion consists of one carbon atom covalently bonded to three oxygen atoms. The entire molecule has a charge of –2. Therefore, its valence is –2, as well. Polyatomic ions remain unchanged in simple chemical reactions because of the strong bonds that hold the component atoms together. They behave as a single unit and should be treated as a single ion.

Table 3.4 on the next page gives the valences, formulas, and names of many common polyatomic ions.

Language LINK

It is important to learn the names and valences of the five most common polyatomic ions: nitrate, carbonate, chlorate, sulfate, and phosphate. These ions form many of the chemicals in nature and in common use. While the task seems overwhelming, it may help to learn the "big five" using a mnemonic, or memory aid. You can use the following mnemonic to remember their names, valences, and number of oxygen atoms:

NICK the CAMEL had a CLAM for SUPPER in PHOENIX.

The first letter identifies the polyatomic ion. The number of vowels represents the valence. The number of consonants represents the number of oxygen atoms. For example, NICK (nitrate) has three consonants and one vowel. Therefore, nitrate contains three oxygen atoms and has a valence of –1. (All of these valences are negative.)

Try to come up with your own mnemonic.

The most common polyatomic cation is the ammonium ion, $[NH_4^+]$. The five atoms in NH_4^+ form a particle with a +1 charge. Because the atoms are bonded together strongly, the polyatomic ion is not altered in most chemical reactions. For example, when ammonium chloride is dissolved in water, the only ions in the solution are ammonium ions and chloride ions.

Table 3.4 Names and Valences of Some Common Ions

Valence = –1			
Ion	**Name**	**Ion**	**Name**
CN^-	cyanide	$H_2PO_3^-$	dihydrogen phosphite
CH_3COO^-	acetate	$H_2PO_4^-$	dihydrogen phosphate
ClO^-	hypochlorite	MnO_4^-	permanganate
ClO_2^-	chlorite	NO_2^-	nitrite
ClO_3^-	chlorate	NO_3^-	nitrate
ClO_4^-	perchlorate	OCN^-	cyanate
HCO_3^-	hydrogen carbonate	HS^-	hydrogen sulfide
HSO_3^-	hydrogen sulfite	OH^-	hydroxide
HSO_4^-	hydrogen sulfate	SCN^-	thiocyanate
Valence = –2			
Ion	**Name**	**Ion**	**Name**
CO_3^{2-}	carbonate	O_2^{2-}	peroxide
$C_2O_4^{2-}$	oxalate	SiO_3^{2-}	silicate
CrO_4^{2-}	chromate	SO_3^{2-}	sulfite
$Cr_2O_7^{2-}$	dichromate	SO_4^{2-}	sulfate
HPO_3^{2-}	hydrogen phosphite	$S_2O_3^{2-}$	thiosulfate
HPO_4^{2-}	hydrogen phosphate		
Valence = –3			
Ion	**Name**	**Ion**	**Name**
AsO_3^{3-}	arsenite	PO_3^{3-}	phosphite
AsO_4^{3-}	arsenate	PO_4^{3-}	phosphate

Writing Chemical Formulas Using Valences

You can use valences to write chemical formulas. This method is faster than using Lewis structures to determine chemical formulas. As well, you can use this method for both ionic and covalent compounds. In order to write a chemical formula using valences, you need to know which elements (or polyatomic ions) are in the compound, and their valences. You also need to know how to use the **zero sum rule**: *For neutral chemical formulas containing ions, the sum of positive valences plus negative valences of the atoms in a compound must equal zero.*

In the compound potassium fluoride, KF, each potassium ion has a charge of +1. Each fluoride ion has a charge of –1. Because there is one of each ion in the formula, the sum of the valences is zero.

What is the formula of a compound that consists of magnesium and chlorine? You know that the valence of magnesium, Mg, is +2. The valence of chlorine, Cl, is −1. The formula MgCl is not balanced, however, because it does not yet obey the zero sum rule. How can you balance this formula? You might be able to see, at a glance, that two chlorine atoms are needed for every magnesium atom. If it is not obvious how to balance a formula, you can follow these steps:

1. Write the unbalanced formula. Remember that the metal is first and the non-metal is second.

   Mg Cl

2. Place the valence of each element on top of the appropriate symbol.

   $$\overset{+2}{\text{Mg}}\quad\overset{-1}{\text{Cl}}$$

3. Using arrows, bring the numbers (without the signs) down to the subscript positions *by crossing over.*

   $$\overset{+2}{\text{Mg}}\quad\overset{-1}{\text{Cl}}$$

4. Check the subscripts. Any subscript of "1" can be removed.

   $MgCl_2$

You can check your formula by drawing a Lewis structure, as shown in Figure 3.42.

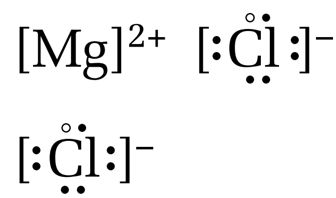


Figure 3.42 This Lewis structure represents magnesium chloride, $MgCl_2$. Each atom has achieved a stable octet.

Practice Problems

16. Write a balanced formula for a compound that contains sulfur and each of the following elements. Use a valence of −2 for sulfur.

(a) sodium
(b) calcium
(c) barium
(d) aluminum
(e) rubidium
(f) hydrogen

17. Write a balanced formula for a compound that contains calcium and each of the following elements.

(a) oxygen
(b) sulfur
(c) chlorine
(d) bromine
(e) phosphorus
(f) fluorine

PROBLEM TIP

After the crossing over step, you may need to reduce the subscripts to their lowest terms. For example, Mg_2O_2 becomes MgO. Be_2O_2 becomes BeO. Remember, formulas for ionic compounds represent ratios of ions.

How do you write and balance formulas that contain polyatomic ions? The same steps can be used, as long as you keep the atoms that belong to a polyatomic ion together. The easiest way to do this is to place brackets around the polyatomic ion at the beginning.

For example, suppose that you want to write a balanced formula for a compound that contains potassium and the phosphate ion. Use the following steps as a guide.

1. Write the unbalanced formula. Place brackets around any polyatomic ions that are present.

$$\begin{array}{ll} K & PO_4 \\ K & (PO_4) \end{array}$$

2. Write the valence of each ion above it. (Refer to Table 3.4.)

$$\begin{array}{ll} +1 & -3 \\ K & (PO_4) \end{array}$$

3. Cross over, and write the subscripts.

$$K_3(PO_4)_1$$

4. Tidy up the formula. Remember that you omit the subscript if there is only one particle in the ionic compound or molecule. Here the brackets are no longer needed, so they can be removed.

$$K_3PO_4$$

Pay close attention to the brackets when you are writing formulas that contain polyatomic ions. For example, how would you write the formula for a compound that contains ammonium and phosphate ions?

PROBLEM TIP

In this formula, the brackets must remain around the ammonium ion to distinguish the subscripts. The subscript 4 refers to how many hydrogen atoms are in each ammonium ion. The subscript 3 refers to how many ammonium ions are needed to form an ionic compound with the phosphate ion.

1. Write the unbalanced formula. Place brackets around any polyatomic ions that are present.

$$\begin{array}{ll} NH_4 & PO_4 \\ (NH_4) & (PO_4) \end{array}$$

2. Write the valence above each ion.

$$\begin{array}{ll} +1 & -3 \\ (NH_4) & (PO_4) \end{array}$$

3. Cross over, and write in the subscripts.

$$(NH_4)_3(PO_4)_1$$

4. Tidy up the formula. Remove the brackets only when the polyatomic ion has a subscript of "1".

$$(NH_4)_3PO_4$$

Try the following problems to practise writing formulas for compounds containing polyatomic ions.

Practice Problems

18. Use the information in Table 3.4 to write a chemical formula for a compound that contains sodium and each of the following polyatomic ions.

(a) nitrate **(c)** sulfite **(e)** thiosulfate

(b) phosphate **(d)** acetate **(f)** carbonate

19. Repeat question 19 using magnesium instead of sodium.

Naming Chemical Compounds

When writing a chemical formula, you learned that you write the metal element first. Similarly, the metal comes first when naming a chemical compound. For example, sodium chloride is formed from the metal sodium and the non-metal chlorine. Think of other names you have seen in this chapter, such as beryllium chloride, calcium oxide, and aluminum oxide. In each case, the metal is first and the non-metal is second. In other words, the cation is first and the anion is second. This is just one of the rules in **chemical nomenclature**: the system that is used in chemistry for naming compounds.

A chemical formula identifies a specific chemical compound because it reveals the composition of the compound. Similarly, the name of a compound distinguishes the compound from all other compounds.

Figure 3.43 **Many common chemicals have trivial, or common, names.**

In the early days of chemistry, there were no rules for naming compounds. Often, compounds received the names of people or places. Some of the original names are still used today. They are called *trivial* or *common* names because they tell little or nothing about the chemistry of the compounds. For example, potassium nitrate, KNO_3, is commonly known as saltpetre. The Greek word for rock is petra, and saltpetre is a salt found crusted on rocks. The chemical name of a compound of ammonium and chloride ions, is ammonium chloride NH_4Cl. Long before NH_4Cl received this name, however, people commonly referred to it as sal ammoniac. They mined this ionic compound near the ancient Egyptian temple of Ammon in Libya. The name sal ammoniac literally means "salt of Ammon." Figure 3.43 shows other examples of common (trivial) names for familiar compounds.

Early chemists routinely gave trivial names to substances before understanding their chemical structure and behaviour. This situation changed during the mid- to late-1800s. By this time, chemistry was firmly established as a science. Chemists observed and discovered new patterns of chemical relationships (such as periodicity). As well, chemists discovered new chemical compounds with tremendous frequency. The rapidly increasing number of chemical compounds required a more organized method of nomenclature.

The International Union of Pure and Applied Chemistry (IUPAC) was formed in 1919 by a group of chemists. The main aim of IUPAC was to establish international standards for masses, measurement, names, and symbols used in the discipline of chemistry. To further that aim, IUPAC developed, and continues to develop, a consistent and thorough system of nomenclature for compounds.

Table 3.5 contains the IUPAC names of selected common compounds as well as their common names.

Table 3.5 Common Chemical Compounds

IUPAC name	Chemical formula	Common name	Use or propety
aluminum oxide	Al_2O_3	alumina	abrasive
calcium carbonate	$CaCO_3$	limestone, marble	building, sculpting
calcium oxide	CaO	lime	neutralizing acidified lakes
hydrochloric acid	HCl	muriatic acid	cleaning metal
magnesium hydroxide	$Mg(OH)_2$	milk of magnesia	antacid
dinitrogen monoxide	N_2O	laughing gas	used in dentistry as an anaesthetic
silicon dioxide	SiO_2	quartz sand	manufacturing glass
sodium carbonate	Na_2CO_3	washing soda	general cleaner
sodium chloride	$NaCl$	table salt	enhancing flavour
sodium hydrogen carbonate	$NaHCO_3$	baking soda	making baked goods rise
sodium hydroxide	$NaOH$	lye	neutralizing acids
sodium thiosulfate	NaS_2O_3	hypo	fixer in photography

Naming Binary Compounds Containing a Metal and a Non-metal

A **binary compound** is an inorganic compound that contains two elements. Binary compounds may contain a metal and a non-metal or two non-metals. Binary compounds are often ionic compounds. To name a binary ionic compound, name the cation first and the anion second. For example, the compound that contains sodium and chlorine is called sodium chloride.

In the subsections that follow, you will examine the rules for naming metals and non-metals in binary compounds.

Naming Metals in Chemical Compounds: The Stock System

The less electronegative element in a binary compound is always named first. Often this element is a metal. You use the same name as the element. For example, *sodium* chloride, $NaCl$, *calcium* oxide, CaO, and *zinc* sulfide, ZnS, contain the metals sodium, calcium, and zinc.

Many of the common metals are transition elements that have more than one possible valence. For example, tin is able to form the ions Sn^{2+} and Sn^{4+}, iron can form Fe^{2+} and Fe^{3+}, and copper can form Cu^{+} and Cu^{2+}. (The most common transition metals with more than one valence number are listed in Table 3.3.) The name of a compound must identify which ion is present in the compound. To do this, the element's name is used, followed by the valence in parentheses, written in Roman numerals. Therefore, Sn^{4+} is tin(IV), Fe^{3+} is iron(III), and Cu^{2+} is copper(II). This naming method is called the **Stock system** after Alfred Stock, a German chemist who first used it. Some examples of Stock system names are listed in Table 3.6.

Another Method for Naming Metals with Two Valences

In a method that predates the Stock system, two different endings are used to distinguish the valences of metals. The ending *-ic* is used to represent the *larger* valence number. The ending *-ous* is used to represent the *smaller* valence number. Thus, the ions Sn^{2+} and Sn^{4+} are named stann*ous* ion and stann*ic* ion. To use this system, you need to know the Latin name of an element. For example, the two ions of lead are the plumbous and plumbic ions. See Table 3.6 for more examples.

This naming method has several drawbacks. Many metals have more than two oxidation numbers. For example, chromium can form three different ions, and manganese can form five different ions. Another drawback is that the name does not tell you what the valence of the metal is. It only tells you that the valence is the smaller or larger of two.

Table 3.6 Two ways to Name Cations with Two Valences

element	ion	Stock system	Alternative system
copper	Cu^{+}	copper(I)	cuprous
	Cu^{2+}	copper(II)	cupric
mercury	Hg_2^{2+}	mercury(I)	mercurous
	Hg^{2+}	mercury(II)	mercuric
lead	Pb^{2+}	lead(II)	plumbous
	Pb^{4+}	lead(IV)	plumbic

mind STRETCH

Some types of ionic compounds can absorb water so that each formula unit is attached to a specific number of water molecules. They are called *hydrates*. $BaOH_2 \bullet 8H_2O$ is called barium hydroxide octahydrate. $CaSO_4 \bullet 2H_2O$ is called calcium sulfate dihydrate. Can you see the pattern? Try naming $MgSO_4 \bullet 7H_2O$. Use Table 3.8 to help you. You will learn more about hydrates in Chapter 6.

Naming Non-Metals in Chemical Compounds

To distinguish the non-metal from the metal in the name of a chemical compound, the non-metal (or more electronegative element) is always written second. Its ending changes to *-ide*. For example, hydrogen changes to *hydride*, carbon changes to *carbide*, sulfur changes to *sulfide*, and iodine changes to *iodide*.

Putting It All Together

To name a binary compound containing metal and a non-metal, write the name of the metal first and the name of the non-metal second. For example, a compound that contains potassium as the cation and bromine as the anion is called *potassium bromide*. Be sure to indicate the valence if necessary, using the Stock system. For example, a compound that contains Pb^{2+} and oxygen is called *lead(II) oxide*.

Practice Problems

20. Write the IUPAC name for each compound.

(a) Al_2O_3 **(d)** Cu_2S
(b) $CaBr_2$ **(e)** Mg_3N_2
(c) Na_3P **(f)** HgI_2

21. Write the formula of each compound.

(a) iron(II) sulfide **(d)** cobaltous chloride
(b) stannous oxide **(e)** manganese(II) iodide
(c) chromium(II) oxide **(f)** zinc oxide

Naming Compounds That Contain Hydrogen

If a binary compound contains hydrogen as the less electronegative element, "hydrogen" is used first in the name of the compound. For example, HCl is called *hydrogen chloride* and H_2S is called *hydrogen sulfide*. Sometimes hydrogen can be the anion, usually in a compound that contains a Group 1 metal. If hydrogen is the anion, its ending must be changed to *-ide*. For example, NaH is called *sodium hydride* and LiH is called *lithium hydride*. Hydrogen-containing compounds can also be formed with the Group 15 elements. These compounds are usually referred to by their common names as opposed to their IUPAC names. For example, NH_3 is called *ammonia*, PH_3 is called *phosphine*, AsH_3 is called *arsine*, and SbH_3 is called *stibine*.

Many compounds that contain hydrogen are also acids. For example, H_2SO_4, hydrogen sulfate, is also called sulfuric acid. You will learn about acid nomenclature in Chapter 10.

Practice Problems

22. Write the IUPAC name for each compound.

(a) H_2Se **(d)** LiH

(b) HCl **(e)** CaH_2

(c) HF **(f)** PH_3

Naming Compounds That Contain Polyatomic Ions

Many compounds contain one or more polyatomic ions. Often these compounds contain three elements, in which case they are called **tertiary compounds**. Although they are not binary compounds, they still contain one type of anion and one type of cation. The same naming rules that apply to binary compounds apply to these compounds as well. For example, NH_4Cl is called *ammonium chloride*. Na_2SO_4 is called *sodium sulfate*. $NiSO_4$ is called *nickel(II) sulfate*. NH_4NO_3 is called *ammonium nitrate*.

The non-metals in the periodic table are greatly outnumbered by the metals. There are many negatively charged polyatomic ions, however, to make up for this. In fact, polyatomic anions are commonly found in everyday chemicals. Refer back to Table 3.4 for the names of the most common polyatomic anions.

When you are learning the names of polyatomic ions, you will notice a pattern. For example, consider the polyatomic ions that contain chlorine and oxygen:

ClO^-	hypochlorite
ClO_2^-	chlorite
ClO_3^-	chlorate
ClO_4^-	perchlorate

Can you see the pattern? Each ion has the same valence, but different numbers of oxygen atoms. The base ion is the one with the "ate" ending chlorate. It contains three oxygen atoms. When the ending is changed to "ite," subtract an oxygen atom from the chlorate ion. The resulting chlorite ion contains two oxygen atoms. Add "hypo" to "chlorite," and subtract one more oxygen atom. The resulting hypochlorite ion has one oxygen atom. Adding "per" to "chlorate," means that you should add an oxygen to the chlorate ion. The perchlorate ion has four oxygen atoms.

The base "ate" ions do not always have three oxygen atoms like chlorate does. Consider the polyatomic ions that contain sulfur and oxygen. In this case, the base ion, sulfate, SO_4^{2-}, contains four oxygen atoms. The sulfite ion, SO_3^{2-}, therefore contains three oxygen atoms. The hyposulfite ion, SO_2^{2-}, contains two oxygen atoms. Once you know the meanings of the prefixes and suffixes, you need only memorize the formulas of the "ate" ions. You can work out the formulas for the related ions using their prefixes and suffixes. The meanings of the prefixes and suffixes are summarized in Table 3.7. In this table, the "*x*" stands for the number of oxygen atoms in the "ate" ion.

Table 3.7 Meaning of prefixes and suffixes

Prefix and suffix			Number of oxygen atoms
hypo		ite	$x - 2$ oxygen atoms
		ite	$x - 1$ oxygen atoms
		ate	x oxygen atoms
per		ate	$x + 1$ oxygen atoms

Language LINK

The prefix "thio" in the name of a polyatomic ion means that an oxygen atom in the root "ate" ion has been replaced by a sulfur atom. For example, the sulfate ion is SO_4^{2-}, while the thiosulfate ion is $S_2O_3^{2-}$. Notice that the valence does not change.

Practice Problems

23. Write the IUPAC name for each compound.

(a) $(NH_4)_2SO_3$ **(d)** $Ni(OH)_2$

(b) $Al(NO_2)_3$ **(e)** Ag_3PO_4

(c) Li_2CO_3 **(f)** $Cu(CH_3COO)_2$

Naming Binary Compounds That Contain Two Non-Metals

To indicate that a binary compound is made up of two non-metals, a prefix is usually added to both non-metals in the compound. This prefix indicates the *number of atoms of each element* in one molecule or formula unit of the compound. For example, P_2O_5 is named *diphosphorus pentoxide*. Alternatively, the Stock System may be used, and P_2O_5 can be named *phosphorus (V) oxide*. $AsBr_3$ is named phosphorus tribromide. The prefix *mono-* is often left out when there is only one atom of the first element in the name. A list of numerical prefixes is found in Table 3.8.

Table 3.8 Numerical Prefixes for Binary Compounds That Contain Two Non-Metals

Number	Prefix
1	mono-
2	di-
3	tri-
4	tetra-
5	penta-
6	hexa-
7	hepta-
8	octa-
9	nona-
10	deca-

Practice Problems

24. Write the IUPAC name for each compound.

(a) SF_6 **(c)** PCl_5

(b) N_2O_5 **(d)** CF_4

Section Wrap-up

In section 3.4, you learned how to name ionic and covalent compounds. You also learned how to write their formulas. In Chapter 4, you will learn how compounds and elements interact in nature, in the laboratory, and in everyday life. These interactions are responsible for the tremendous variety of substances and materials found on Earth.

Section Review

Unit Project Prep

In this section, you saw some common names for chemicals. Before you begin your Unit Project, create a list of names of chemicals found in common household cleaning products. If you know only the common name, find out the chemical name.

1. K/U Write an unambiguous name for each compound.
   - **(a)** K_2CrO_4
   - **(b)** NH_4NO_3
   - **(c)** Na_2SO_4
   - **(d)** $Sr_3(PO_4)_2$
   - **(e)** KNO_2
   - **(f)** $Ba(ClO)_2$
2. K/U Write the name of each binary compound.
   - **(a)** $MgCl_2$
   - **(b)** Na_2O
   - **(c)** $FeCl_3$
   - **(d)** CuO
   - **(e)** ZnS
   - **(f)** $AlBr_3$
3. K/U Write the formula of each compound.
   - **(a)** sodium hydrogen carbonate
   - **(b)** potassium dichromate
   - **(c)** sodium hypochlorite
   - **(d)** lithium hydroxide
   - **(e)** potassium permanganate
   - **(f)** ammonium chloride
   - **(g)** calcium phosphate
   - **(h)** sodium thiosulfate
4. K/U Write the formula and name of two possible compounds that could be formed from each pair of elements.
   - **(a)** vanadium and oxygen
   - **(b)** iron and sulfur (–2)
   - **(c)** nickel and oxygen
5. C The formula for hydrogen peroxide is H_2O_2. Explain why it is not correct to write the formula for this covalent compound as HO.

CHAPTER 3 Review

Reflecting on Chapter 3

Summarize this chapter in the format of your choice. Here are a few ideas to use as guidelines:

- Elements combine to form a wide variety of compounds.
- Ionic compounds and covalent compounds have characteristic properties. You can use these properties to classify various compounds.
- You can use the concepts of electron arrangement and forces in atoms to explain the periodic trend of electronegativity.
- The electronegativity difference between elements can be used to predict what kinds of compounds the elements will form.
- Lewis structures can represent the formation of ionic and covalent compounds according to the octet rule.
- You can explain the conductivity of covalent and ionic compounds using an understanding of covalent and ionic bonding.
- By predicting the shapes of molecules, you can predict their polarity.
- The polarity of molecules can be used to explain the range of boiling points and melting points among compounds that contain molecules with similar masses.
- There is a methodical way to unambiguously name compounds and write their chemical formulas.

Reviewing Key Terms

For each of the following terms, write a sentence that shows your understanding of its meaning.

chemical bonds	ionic bond
covalent bond	electronegativity
octet rule	isoelectronic
pure covalent bond	diatomic elements
double bond	triple bond
molecular compounds	intramolecular forces
intermolecular forces	metallic bond
alloy	polar covalent bond
lone pairs	bonding pairs
polar molecule	dipolar molecules
non-polar molecule	chemical formula
valence	polyatomic ions
zero sum rule	chemical nomenclature
binary compound	Stock system
tertiary compounds	

Knowledge/Understanding

1. Both electronegativity and electron affinity describe electron attraction. Explain how they are different.
2. Calculate ΔEN for each bond.
   (a) Zn—O
   (b) Mg—I
   (c) Co—Cl
   (d) N—O
3. Indicate whether each bond in question 2 is ionic or covalent.
4. Name three characteristics of covalent compounds and three characteristics of ionic compounds.
5. Give two examples of ionic compounds and two examples of covalent compounds.
6. How does a property of noble gases lead to the octet rule?
7. Draw a Lewis structure to represent each ionic compound.
   (a) potassium bromide
   (b) calcium fluoride
   (c) magnesium oxide
   (d) lithium oxide
8. Draw a Lewis structure to represent each covalent compound.
   (a) CO_2
   (b) NaH
   (c) NF_3
9. Describe, in detail, what happens when an ionic bond forms between calcium and chlorine. Use Lewis structures to illustrate your description.
10. Diatomic elements (such as oxygen, nitrogen, and chlorine) tend to exist at room temperature as gases. Explain why this is true using your understanding of bonding.
11. A solid covalent compound has both intermolecular and intramolecular forces. Do solid ionic compounds contain intermolecular forces? Explain your answer.
12. Distinguish between a non-polar covalent bond and a polar covalent bond.
13. Distinguish between an ionic bond and a polar covalent bond.

14. Without calculating ΔEN, arrange each set of bonds from most polar to least polar. Then calculate ΔEN for each bond to check your arrangement.
   (a) Mn—O, Mn—N, Mn—F
   (b) Be—F, Be—Cl, Be—Br
   (c) Ti—Cl, Fe—Cl, Cu—Cl, Ag—Cl, Hg—Cl
15. What kind of diagram would you use (Lewis structure, structural diagram, ball-and-stick model, or space-filling model) to illustrate each idea?
   (a) When sodium and chlorine form an ionic bond, sodium loses an electron and chlorine gains an electron.
   (b) Water, H_2O, is a molecule that has a bent shape.
   (c) Each oxygen atom in carbon dioxide, CO_2, has two bonding pairs and two non-bonding pairs.
   (d) Silicon tetrachloride, $SiCl_4$, is a tetrahedral molecule.
16. Write the valences for the elements in each compound. If the compound is ionic, indicate the charge that is associated with each valence.
   (a) AgCl (b) Mn_3P_2
   (c) PCl_5 (d) CH_4
   (e) TiO_2 (f) HgF_2
   (g) CaO (h) FeS
17. Write the formula of each compound.
   (a) tin(II) fluoride
   (b) barium sulfate
   (c) hydrogen cyanide
   (d) cesium bromide
   (e) ammonium hydrogen phosphate
   (f) sodium periodate
   (g) potassium bromate
   (h) sodium cyanate
18. Write the name of each compound.
   (a) HIO_2 (b) $KClO_4$
   (c) CsF (d) $NiCl_2$
   (e) $NaHSO_4$ (f) $Al_2(SO_3)_3$
   (g) $K_2Cr_2O_7$ (h) $Fe(IO_4)_3$
19. Name each compound in two different ways.
   (a) FeO (b) $SnCl_4$
   (c) $CuCl_2$ (d) $CrBr_3$
   (e) PbO_2 (f) HgO

Inquiry

20. Suppose that you have two colourless compounds. You know that one is an ionic compound and the other is a covalent compound. Design an experiment to determine which compound is which. Describe the tests you would perform and the results you would expect.
21. You have two liquids, A and B. You know that one liquid contains polar molecules, and the other liquid contains non-polar molecules. You do not know which is which, however. You pour each liquid so that it falls in a steady, narrow stream. As you pour, you hold a negatively charged ebonite rod to the stream. The stream of liquid A is deflected toward the rod. The rod does not affect the stream of liquid B. Which liquid is polar? Explain your answer.

Communication

22. Explain how you would predict the most common valences of the elements of the second period (Li, Be, B, C, N, O, F, and Ne) if you did not have access to the periodic table. Use Lewis structures to illustrate your explanation.
23. Create a concept map to summarize what you learned in this chapter about the nature of bonding and the ways in which bonding models help to explain physical and chemical properties.
24. Compare and contrast ionic bonding and metallic bonding. Include the following ideas:
   - Metals do not bond to other metals in definite ratios. Metals do bond to non-metals in definite ratios.
   - Solid ionic compounds do not conduct electricity, but solid metals do.
25. Explain why it was important for chemists worldwide to decide on a system for naming compounds.

Making Connections

26. Chemists do not always agree on names, not just for compounds but even for elements. As new elements are synthesized in laboratories, they must be named. Until 1997, there was a controversy over the names of elements 104 to

109 (called the *transfermium* elements). The periodic table at the back of this textbook gives the names that have now been accepted. Do some research to find out what other names were proposed for those elements. Find out what justification was given for the alternative names and the accepted names. Then write an essay in which you evaluate the choice that was made. Do you agree or disagree? Justify your opinion.

Answers to Practice Problems and Short Answers to Section Review Questions:

Practice Problems: **1.(a)** 1.24, covalent **(b)** 0.50, covalent **(c)** 1.85, ionic **(d)** 1.94, ionic **(e)** 1.78, ionic **(f)** 0.49, covalent **(g)** 1.73, ionic **(h)** 2.03, ionic **2.(a)** 2.44, ionic **(b)** 2.34, ionic **(c)** 3.16, ionic **(d)** 3.00, ionic **(e)** 1.98, ionic **(f)** 2.55, ionic **3.(a)** one calcium atom gives up two electrons to one oxygen atom **(b)** one potassium atom gives up one electron to one chlorine atom **(c)** one potassium atom gives up one electron to one fluorine atom **(d)** one lithium atom gives up one electron to one fluorine atom **(e)** one lithium atom gives up one electron to one bromine atom **(f)** one barium atom gives up two electrons to one oxygen atom **4.(a)** 1.85 **(b)** 2.16 **(c)** 2.46 **(d)** 2.51 **(e)** 1.76 **(f)** 1.96 **5.(a)** one magnesium atom gives up one electron to each of two chlorine atoms **(b)** one calcium atom gives up one electron to each of two chlorine atoms **(c)** two lithium atoms each give up one electron to one oxygen atom **(d)** two sodium atoms each give up one electron to one oxygen atom **(e)** two potassium atoms each give up one electron to one sulfur atom **(f)** one calcium atom gives up one electron to each of two bromine atoms **6.(a)** Each iodine atom has seven electrons. Two iodine atoms bonded together share one pair of electrons so each has access to eight electrons. **(b)** Each bromine atom has seven electrons. Two bromine atoms bonded together share one pair of electrons so that each has access to eight electrons. **(c)** Each hydrogen atom has one electron. Two hydrogen atoms bonded together share one pair of electrons so that each has access to two electrons. **(d)** Each fluorine atom has seven electrons. Two fluorine atoms bonded together share one pair of electrons so that each has access to eight electrons. **7.(a)** One hydrogen atom bonds to one oxygen atom, sharing one electron pair. **(b)** Two chlorine atoms bond to one oxygen atom. Each chlorine atom shares one pair of electrons with the oxygen atom. **(c)** One carbon atom bonds to four hydrogen atoms. Each hydrogen atom shares one pair of electrons with the carbon atom. **(d)** One iodine atom bonds to one hydrogen atom. They share an electron pair. **(e)** One nitrogen atom bonds to three hydrogen atoms. Each hydrogen atom shares a pair of electrons with the nitrogen atom. **(f)** One hydrogen atom bonds to one rubidium atom. They share a pair of electrons. **8.** The carbon atom shares two pairs of electrons with each sulfur atom, so it has two double bonds. **9.** The carbon atom shares one pair of electrons with hydrogen, and three pairs of electrons with the nitrogen atom. **10.** The two carbon atoms share three pairs of electrons in a triple bond. Each carbon atom shares one pair of electrons with one hydrogen atom. **11.(a)** polar covalent **(b)** covalent **(c)** covalent **(d)** polar covalent **(e)** covalent **(f)** ionic **(g)** polar covalent **(h)** ionic **12.(a)** C δ^+, F δ^- **(d)** Cu δ^+, O δ^- **(g)** Fe δ^+, O δ^- **13.(a)** O—O, N—O, H—Cl, Na—Cl **(b)** N—N, P—O, C—Cl, Mg—C **14.** Group 13, 3, Group 16, –2, Group 17, –1 **15.** 0, do not tend to gain or lose electrons **16.(a)** Na_2S **(b)** CaS **(c)** BaS **(d)** Al_2S_3 **(e)** Rb_2S **(f)** H_2S **17.(a)** CaO **(b)** CaS **(c)** $CaCl_2$ **(d)** $CaBr_2$ **(e)** Ca_3P_2 **(f)** CaF_2 **18.(a)** $NaNO_3$ **(b)** Na_3PO_4 **(c)** Na_2SO_3 **(d)** $NaCH_3COO$ **(e)** $Na_2S_2O_3$ **(f)** Na_2CO_3 **19.** $Mg(NO_3)_2$ **(b)** $Mg_3(PO_4)_2$ **(c)** $MgSO_3$ **(d)** $Mg(CH_3COO)_2$ **(e)** MgS_2O_3 **20.(a)** aluminum oxide **(b)** calcium bromide **(c)** sodium phosphide **(d)** copper(I) sulfide **(e)** magnesium nitride **(f)** mercury(II) iodide **21.(a)** FeS **(b)** SnO (c) CrO **(d)** $CoCl_2$ **(e)** MnI_2 **(f)** ZnO **22.(a)** hydrogen selenide **(b)** hydrogen chloride **(c)** hydrogen fluoride **(d)** lithium hydride **(e)** calcium hydride **(f)** phosphorus(III) hydride **23.(a)** ammonium sulfite **(b)** aluminum nitrite **(c)** lithium carbonate **(d)** nickel(II) hydroxide **(e)** silver phosphate **(f)** copper(II) acetate **24.(a)** sulfur hexafluoride **(b)** dinitrogen pentoxide **(c)** phosphorus pentachloride **(d)** carbon tetrafluoride

Section Review: **3.1:** **4.(a)** Li, La, Zn, Si, Br **(b)** Cs, Y, Ga, P, Cl **5.(a)** 0.40, covalent **(b)** 1.89, ionic **(c)** 0.96, covalent **(d)** 2.16, ionic **6.(a)** low **(b)** covalent **3.2:** **1.(a)** one magnesium atom gives up one electron to each of two fluorine atoms **(b)** one potassium electron gives up one electron to one bromine atom **(c)** one rubidium atom gives up one electron to one chlorine atom **(d)** one calcium atom gives up two electrons to one oxygen atom **2.(a)** The hydrogen atom and chlorine atoms all bond to the carbon atom. **(b)** Each hydrogen and chlorine atom shares one electron pair with the carbon atom. **(c)** The two nitrogen atoms share three pairs of electrons in a triple covalent bond. **3.(a)** covalent **(b)** covalent **(c)** covalent **(d)** ionic **3.3:** **1.(a)** 1.94, ionic **(b)** 0.35, covalent **(c)** 2.23, ionic **(d)** 1.54, polar covalent **(e)** 0.86, polar covalent **(f)** 0.61, polar covalent **2.(d)** Si δ^+, O δ^- **(e)** S δ^+, O δ^- **(f)** C δ^+, Cl δ^- **4.(a)** O—F, H—Br, H—Cl, K—Br **(b)** C—H, C—Br, C—O, C—F **7.(a)** 1.26, non-polar molecule **(b)** 0.97, polar molecule **(c)** 1.55, non-polar molecule **3.4:** **1.(a)** potassium chromate **(b)** ammonium nitrate **(c)** sodium sulfate **(d)** strontium phosphate **(e)** potassium nitrite **(f)** barium hypochlorite **2.(a)** magnesium chloride **(b)** sodium oxide **(c)** iron(III) chloride **(d)** copper(II) oxide **(e)** zinc sulfide **(f)** aluminum bromide **3.(a)** $NaHCO_3$ **(b)** $K_2Cr_2O_7$ **(c)** NaClO **(d)** LiOH **(e)** $KMnO_4$ **(f)** NH_4Cl **(g)** $Ca_3(PO_4)_2$ **(h)** $Na_2S_2O_3$ **4.(a)** any two of: vanadium(II) oxide, VO, vanadium(III) oxide, V_2O_3, vanadium(IV) oxide, VO_2, vanadium(V) oxide, V_2O_5 **(b)** iron(II) sulfide, FeS, iron(I) sulfide, Fe_2S **(c)** nickel(II) oxide, NiO, nickel(III) oxide, Ni_2O_3