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

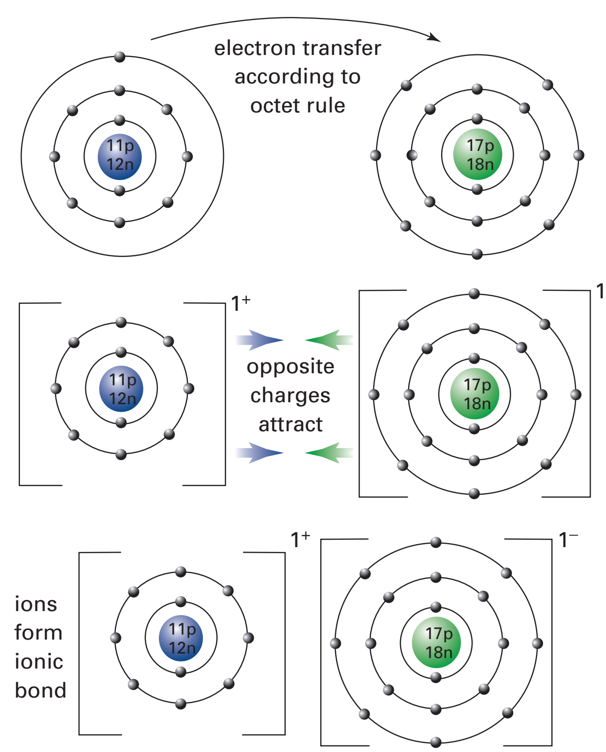

Figure 3.9 Sodium's electron is transferred to chlorine. The atoms become oppositely charged ions with stable octets. Because they are oppositely charged, they are strongly attracted to one another.

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.

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-}.

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.

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.

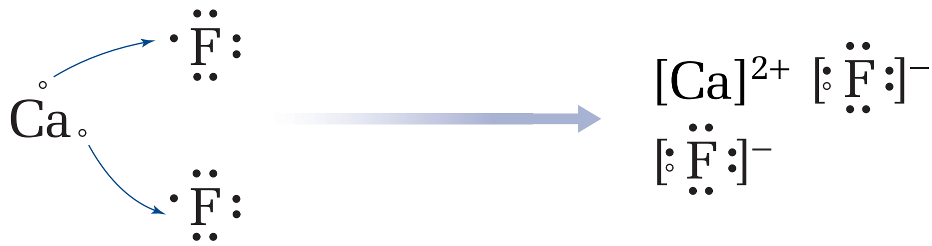

Figure 3.12 **These Lewis structures show the formation of bonds between one atom of calcium and two atoms of fluorine.**

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.

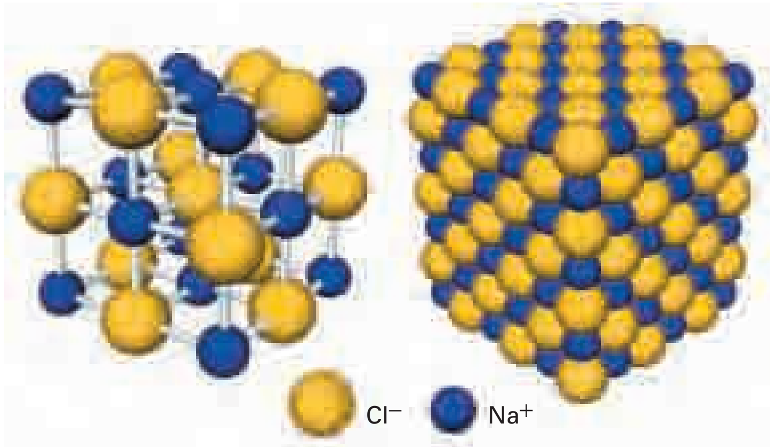

Figure 3.13 In solid sodium chloride, NaCl, sodium and chlorine are arranged in a rigid lattice pattern.

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?

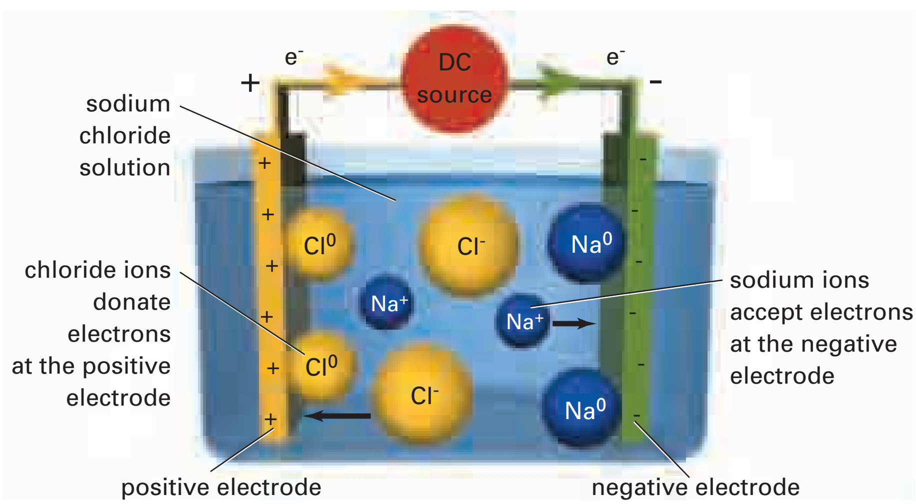

Figure 3.14 Aqueous sodium chloride is a good conductor of electricity.

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.

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.

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**.

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.

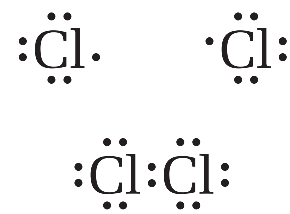

Figure 3.15 These Lewis structures show the formation of a bond between two atoms of chlorine.

CHEM FACT

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

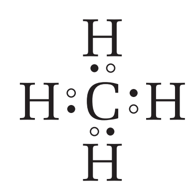

Figure 3.16 This Lewis structure shows a molecule of methane, CH_4.

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.

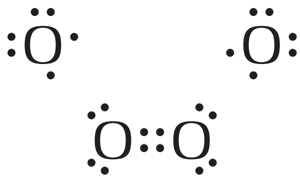

Figure 3.17 These Lewis structures show the formation of a double bond between two atoms of oxygen.

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!

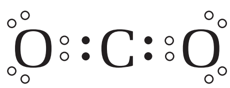

Figure 3.18 This Lewis structure shows the double bond in a molecule of carbon dioxide, CO_2.

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.

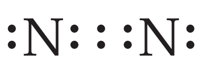

Figure 3.19 This Lewis structure shows the triple bond in a molecule of nitrogen, N_2.

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.

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.

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.

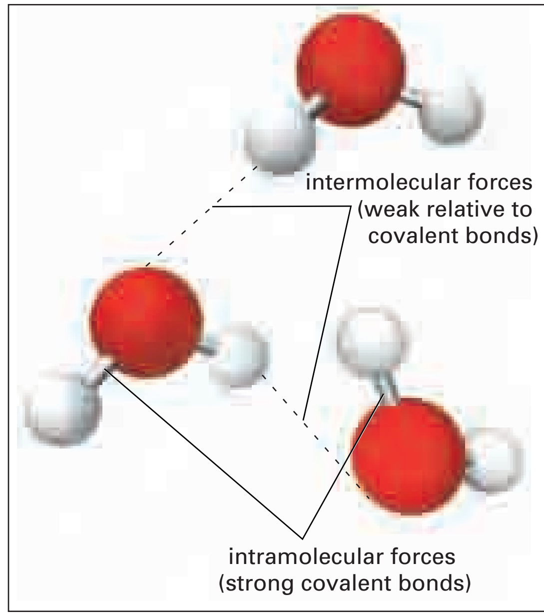

Figure 3.20 **Strong intramolecular forces (covalent bonds) hold the atoms in molecules together. Relatively weak intermolecular forces act between molecules.**

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.

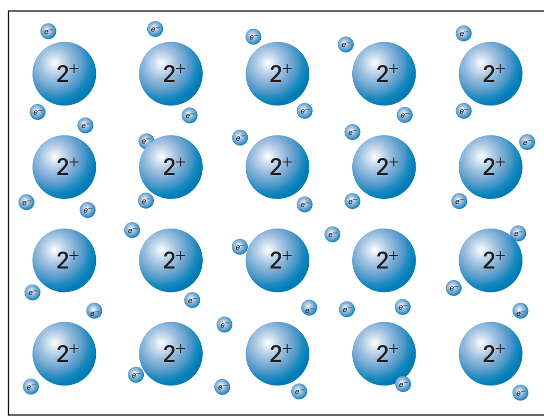

Figure 3.21 **In magnesium metal, the two valence electrons from each atom are free to move in an "electron sea." The valence electrons are shared by all the metal ions.**

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.