SECTION 2.1

The Formation of Ionic and Covalent Bonds

Key Terms

octet rule
ionic bond
ionic compound
covalent bond
molecular compound
single bond
double bond
triple bond
bonding pair
lone pair
Lewis structure
polyatomic ion
polar covalent bond
electronegativity difference

Figure 2.1 The helium that was used to inflate these balloons is a noble gas. Noble gases are some of the very few elements that are found in nature in their elemental form.

Ninety-two naturally occurring elements combine to form the millions of different compounds that are found in nature. Very few of these elements, however, are found in their elemental form in nature. Some of the elements that are found in their elemental form are the noble gases, as illustrated in **Figure 2.1**. What property of atoms causes them to combine with atoms of other elements? Why are some combinations of elements much more common than others? Answers to these questions are based on the types of bonds that form between atoms of elements. Over the next few pages, you will examine some naturally occurring compounds and look for patterns in these compounds to find clues about the nature of chemical bonds.

Clues in Naturally Occurring Compounds

Scientists often study patterns in nature to better understand scientific concepts. Chemists learn a great deal about the nature of chemical bonds by observing trends in naturally forming compounds. For example, ores are metal compounds that are mined, as shown in **Figure 2.2**, to extract the pure metals. Ores are solid and consist of a metal combined with a non-metal, such as oxygen, sulfur, or a halogen, or with polyatomic ions such as carbonate ions, CO_3^{2-}. Very few metals are found in their elemental form in nature. The few metals that are found in their elemental form, such as gold and silver, are called precious metals. For a compound, such as an ore, to be in solid form, some type of strong attractive force must be holding the individual particles together.

Figure 2.2 Ores, consisting of metals combined with non-metals, are sometimes obtained from open pit mines, such as the copper mine shown here.

Clues in the Atmosphere

You can gain more insight into the nature of chemical bonds by examining the atmosphere. It contains the oxygen that you inhale and the carbon dioxide that you exhale. The atmosphere also contains water vapour that condenses to form clouds, which can then become snow or rain. The major component of the atmosphere is nitrogen. As well, there are traces of argon, methane, ozone, and hydrogen. Of these gases, only argon, a noble gas, is found as individual atoms, not bonded to any other atoms. The non-metal elements, oxygen, nitrogen, and hydrogen, are found in the atmosphere as diatomic molecules. This means that they are made up of two identical atoms bonded together. Carbon dioxide, water, and methane are examples of atoms of non-metal elements bonded together.

The following patterns can be discerned from these observations:

- Metals usually form bonds with non-metals. The compounds they form are solid.
- Non-metals can bond with one another to form gases, liquids, or solids.
- The only elements that are *never* found in a combined form in nature are the noble gases.

Stability of Atoms and the Octet Rule

Because atoms of the noble gases are always found as monatomic gases, and because atoms of all other elements are usually found chemically bonded to other atoms, you can infer that there is something very unique about the chemistry of noble gases, which prevents the atoms from forming bonds. Recall, from Chapter 1, that the noble gases are the only elements whose atoms have a filled valence shell, as shown in **Figure 2.3**. This leads to the conclusion that atoms that have filled valence shells do not tend to form chemical bonds with other atoms. Such atoms are referred to as stable.

Figure 2.3 Atoms of each of the noble gases except helium have eight electrons in their outer shell, giving them filled valence shells. Because helium is in Period 1, only its first shell, which holds a maximum of two electrons, is occupied. Thus, for helium, two electrons constitute a filled valence shell.

The observation that a filled valence shell makes atoms stable led early chemists to propose that when bonds form between atoms, they do so in a way that gives each atom a filled valence shell. Because, for most main-group elements, a filled valence shell contains eight electrons, this configuration is often called an *octet*. These observations led to the **octet rule** for bond formation, which is stated below.

octet rule a "rule of thumb" that allows you to predict the way in which bonds will form between atoms

The Octet Rule
When bonds form between atoms, the atoms gain, lose, or share electrons in such a way that they create a filled outer shell containing eight electrons.

As you read in Chapter 1, atoms of the transition elements and inner transition elements can have complex electron configurations. They can have more than eight electrons in their valence shells and, therefore, they do not follow the octet rule. Because main-group elements are much more common on Earth, however, a very large number of compounds that you study will follow the octet rule. Thus, the octet rule provides an important basis on which to predict how bonds will form.

ionic bond the attractive electrostatic force between a negative ion and a positive ion

ionic compound a chemical compound composed of ions that are held together by ionic bonds

The Formation of Ionic Bonds

An **ionic bond** is the attractive electrostatic force between oppositely charged ions. Thus, before an ionic bond can form, atoms must be ionized. According to the octet rule, atoms gain or lose electrons to attain a filled valence shell. In Chapter 1, Section 1.3, you learned that an atom of an element with fewer than four electrons in its valence shell, especially an alkali metal atom, can lose electrons relatively easily. You also learned that an atom with more than four electrons in its valence shell can gain electrons and form a stable ion. Thus, in general, a metal loses all of its valence electrons and becomes an ion with an octet of electrons in its outer shell. A non-metal gains enough electrons to fill its valence shell. These oppositely charged ions exert attractive electrostatic forces on each other, resulting in the formation of an ionic bond. A compound that is held together by ionic bonds is called an **ionic compound**.

Because ionic compounds must have an overall charge of zero, the number of electrons that are lost by the metal atoms must be equal to the number of electrons gained by the non-metal atoms. Two such examples are shown in **Figure 2.4**, in the form of Lewis diagrams. Notice that the electrons of the metals are depicted as open circles and the electrons of the non-metals are depicted as dots, so you can follow them throughout the process.

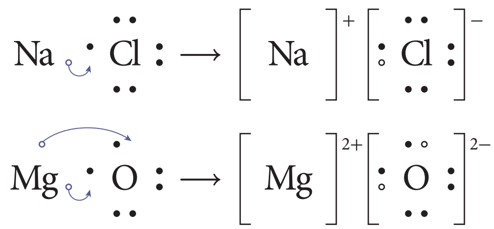

Figure 2.4 When metal atoms, such as sodium and magnesium, lose electrons, they have no valence electrons remaining. Therefore, there are no dots around the symbols for the metal ions.

In each example in **Figure 2.4**, the number of electrons gained by the non-metal atom is exactly the same as the number of electrons lost by the metal atom. It is also possible for the number of electrons gained by a non-metal atom to be different from the number of electrons lost by a metal atom. However, the total number of electrons gained by non-metal atoms must be the same as the total number of electrons lost by metal atoms. Examples of three such situations are shown in **Figure 2.5**.

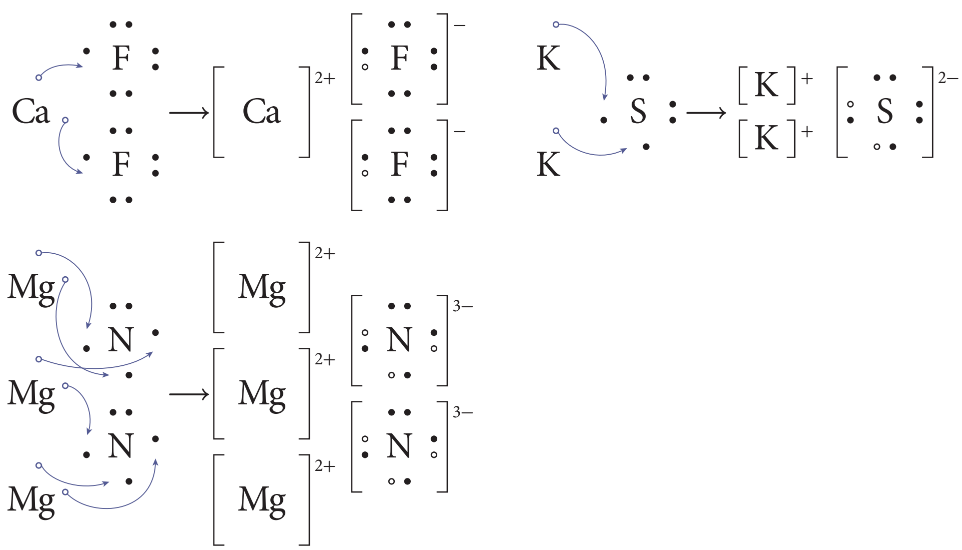

Figure 2.5 In each example, you can see that the total number of positive charges (electrons lost) on the metal ions is equal to the total number of negative charges (electrons gained) on the non-metal ions.

Ionic Compounds Containing Transition Metals

All of the examples in **Figures 2.4** and **2.5** include only main-group elements. You can determine the number of valence electrons of an atom of a main-group element by its group number. Occasionally, however, you will be working with transition metals. In Chapter 1, you read that the electron configuration of transition metals is quite complex. Therefore, it is not possible to predict the number of electrons that a transition metal atom can lose from its group number. In fact, the number of electrons that a transition metal can lose can vary. For example, an iron atom can lose either two electrons or three electrons. You can find the number of electrons that atoms of a transition element can lose by checking the periodic table. **Figure 2.6** shows you how to find the possible charges on the resulting ions after the metal atoms have become ionized. Notice that the possible charges on the ions are highlighted. The figure shows a few common transition metals that can form more than one possible ion. As stated above, iron atoms can lose two or three electrons. Thus iron atoms can form ions with charges of 2+ or 3+.

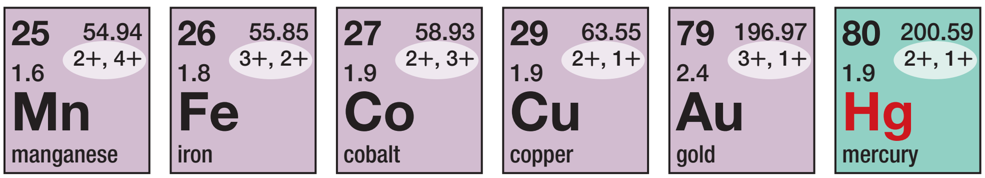

Figure 2.6 These cells are taken directly from the periodic table on page 24. The common ion charges are highlighted.

When you are working with transition metals, you will be given the charge or enough information to determine the charge on the ions. For example, you might be told that two iron atoms have combined with three oxygen atoms and asked to draw a Lewis diagram of the compound. If you do not know that the oxygen ion has a charge of 2−, you can find it in the periodic table. Since there are three oxygen ions, the total negative charge in the compound will be 6−. Thus, the total positive charge on the two iron ions must be 6+. Therefore, there must be a charge of 3+ on each iron ion. The Lewis diagram for this compound is shown in **Figure 2.7 (A)**.

You might also be asked to draw the Lewis diagram of a compound that contains one iron ion and two chloride ions. Since a chloride ion has a charge of 1−, the single iron ion must have a charge of 2+. The Lewis diagram of this compound is shown in **Figure 2.7 (B)**. It is important to remember that the electron configuration for iron atoms is complex. Iron atoms do not actually have two valance electrons or three valence electrons as shown in **Figure 2.7**. The iron atoms are drawn as though they have either two or three valence electrons only because these are the numbers of electrons that they can lose when they become ionized.

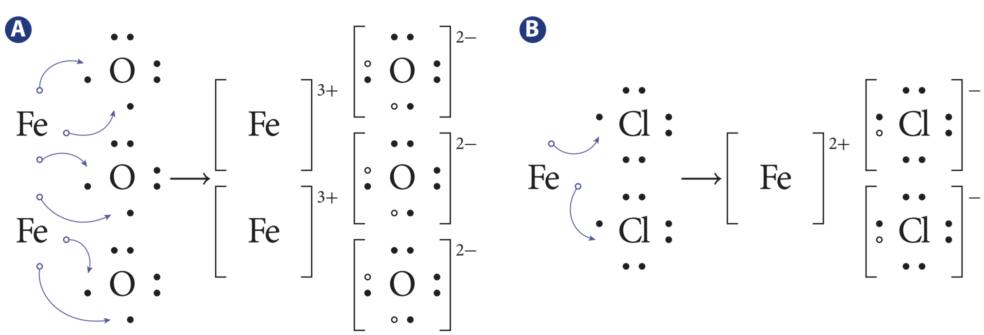

Figure 2.7 **(A)** Each of the two iron atoms loses three electrons to oxygen atoms. Thus, the resulting ions have a charge of 3+. **(B)** An iron atom loses one electron to one chlorine atom and a second electron to another chlorine atom. The resulting iron ion has a charge of 2+.

The Formation of Covalent Bonds

covalent bond the attraction between atoms that results from the sharing of electrons

molecular compound a chemical compound that is held together by covalent bonds

The octet rule states that atoms can also acquire a filled outer shell by sharing electrons. When the nuclei of two atoms are both attracted to one or more pairs of shared electrons, the attraction is called a **covalent bond**. A compound that is held together by covalent bonds is called a **molecular compound**. Molecular compounds consist of non-metal elements only. Examples of molecular compounds are water and carbon dioxide. Molecular compounds can be solid, liquid, or gas at room temperature.

Only unpaired electrons are likely to participate in chemical bonds. **Figure 2.8** shows how covalent bonds form when **(A)** hydrogen atoms share their only electrons and when **(B)** chlorine atoms share their only unpaired electrons.

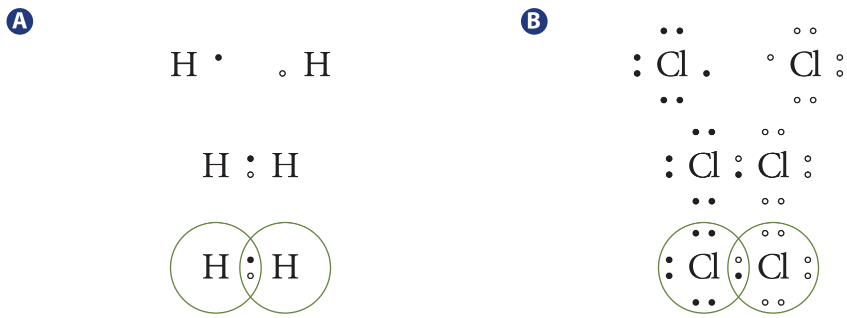

Figure 2.8 In both **(A)** and **(B)**, electrons of one atom are shown as open circles and electrons of the other atom are shown as dots, to help you follow the electrons. In the third row, circles surrounding each atom in the molecule show the filled shell of electrons for each atom.

Multiple Bonds

In some molecules, there are not enough valence electrons for two atoms to share one pair of electrons and form filled valence shells. For example, in carbon dioxide, the carbon atom has four valence electrons and each of the two oxygen atoms has six valence electrons, as shown in **Figure 2.9 (A)**. If the carbon atom shared one pair of electrons with each oxygen atom, each of the oxygen atoms would have only seven electrons and the carbon atom would have only six electrons in the valence shell, as shown in **Figure 2.9 (B)**. This configuration would not provide all atoms with filled outer shells. Instead, to complete an octet for each atom, the unpaired electrons on all atoms are rearranged, as shown in **Figure 2.9 (C)**, to become shared. The atoms now share four electrons as shown in **Figure 2.9 (D)**. It is important to remember that Lewis diagrams are just that—diagrams. It would be more correct to show electron clouds overlapping and forming new electron clouds with different shapes. However, it is more difficult to visualize the number of electrons in valence shells when using the electron cloud model.

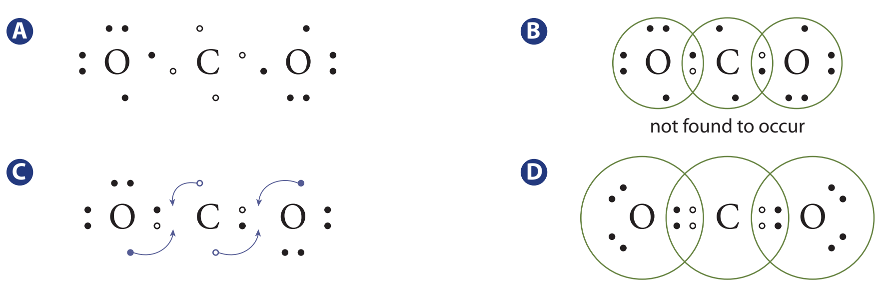

Figure 2.9 **(A)** Each oxygen atom has six valence electrons and the carbon atom has four. **(B)** If each oxygen atom shared two electrons with the carbon atom, neither would have a filled outer shell. **(C)** Instead, the remaining unpaired electrons in both oxygen atoms and the carbon atom are rearranged so that they can also be shared by the atoms. **(D)** When each oxygen atom shares four electrons (two pair) with the carbon atom, all of the atoms acquire an octet of electrons.

***Explain** how you could predict the number of bonds that an atom could form with other atoms.*

Types of Covalent Bonds and Electron Pairs

While one pair of shared electrons constitutes a **single bond**, two pairs of shared electrons make up a **double bond**. Compounds can also have triple bonds, which consist of three pairs of shared electrons. Nitrogen, the gas that makes up most of the atmosphere, is an example of a molecule that has a **triple bond**, as shown in **Figure 2.10**.

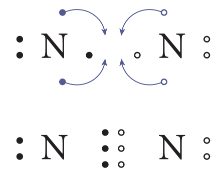

Figure 2.10 Two nitrogen atoms must share three pairs of electrons to complete an octet of electrons around each nitrogen atom.

Although there are rarely any unpaired electrons in molecular compounds, some electron pairs are shared while others are not. A pair of shared electrons is called a **bonding pair**. A pair of electrons that is not involved in a covalent bond is called a **lone pair**. These types of electron pairs are labelled in the water molecule in **Figure 2.11**. When a Lewis diagram is used to portray a complete molecular compound, as done in this figure, the diagram is called a **Lewis structure**.

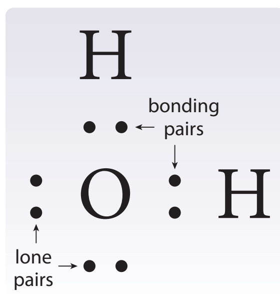

Figure 2.11 The oxygen atom in water has two bonding pairs and two lone pairs. The hydrogen atoms each have one bonding pair.

Drawing Lewis Structures

Although there are no specific steps that you can always follow to draw a Lewis structure, there are some guidelines that will help you. First, draw a Lewis diagram for each atom in the structure. Start with the atom that has the most unpaired electrons. Determine the number of bonds that each atom can form with other atoms. That number is the same as the number of unpaired electrons. Finally, try to fit the atoms together in a way that will create a filled outer shell for each atom. The example used in **Figure 2.12 (A)** has one carbon atom, two oxygen atoms, and two hydrogen atoms. Begin with the carbon atom. If both oxygen atoms are bonded to the carbon atom with double bonds, there will be no way to add the hydrogen atoms. If both hydrogen atoms are bonded to the carbon atom, there will be bonds for only one oxygen atom. The final result is shown in **Figure 2.12 (B)**.

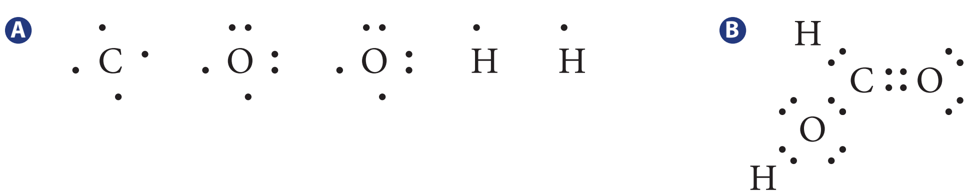

Figure 2.12 If you take the atoms in **(A)** and test different ways of connecting them, you will find that the Lewis structure in **(B)** creates filled outer shells for all of the atoms. This compound is commonly called formic acid, HCOOH.

single bond a covalent bond that results from atoms sharing one pair of electrons

double bond a covalent bond that results from atoms sharing two pairs of electrons

triple bond a covalent bond that results from atoms sharing three pairs of electrons

bonding pair a pair of electrons that is shared by two atoms, thus forming a covalent bond

lone pair a pair of electrons that is not part of a covalent bond

Lewis structure a Lewis diagram that portrays a complete molecular compound

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Polyatomic Ions and Bond Formation

polyatomic ion a molecular compound that has an excess or a deficit of electrons, and thus has a charge

When you first look at the structure in **Figure 2.13 (A)**, (ignoring the colour), it appears to be a typical Lewis structure. However, when you count the electrons, you will find a new feature in this compound. Count the number of electrons that are the same colour as the symbol, and you will find that they represent the number of valence electrons that an atom of that element has. The carbon atom has four black electrons, and each oxygen atom has six red electrons. The colour-coded electrons account for all of the valence electrons that are available. If there were no additional electrons, the atoms would not all have filled valence shells. To fill the shells, two electrons were added, as shown in green in **Figure 2.13 (A)**. These two electrons give the compound a negative charge of 2−. Nevertheless, it is a valid Lewis structure. Some molecular compounds, like non-metal atoms, can gain electrons to complete octets on all of their atoms. Such compounds are called **polyatomic ions** because they consist of two or more atoms. The correct diagram for polyatomic ions includes brackets and a number and sign, as shown in **Figure 2.13 (B)**.

Figure 2.13 **(A)** If you count the number of electrons, you will get 24. Because this is two more electrons than the sum of the valence electrons in three oxygen atoms and one carbon atom, the compound has a charge of 2−. **(B)** To show that this compound is a polyatomic ion, it is bracketed and a 2− is placed outside the brackets.

Typically, electrons in Lewis structures are not colour coded so you cannot easily see whether there are any extra electrons. Nevertheless, you can quickly determine whether a Lewis structure represents a neutral molecular compound or a polyatomic ion by first counting the electrons and comparing that number with the total number of valence electrons that each atom would have. For example, the structure in **Figure 2.13 (A)** has 24 electrons. Add up the number of valence electrons by reasoning that each oxygen atom has six valence electrons and the carbon atom has four valence electrons, giving a total of 22 electrons. You immediately know that you must have two extra electrons, giving you a negatively charged polyatomic ion.

A negatively charged polyatomic ion can bond to a positively charged ion to form an ionic compound in the same way that a metal ion and a non-metal ion can bond to form an ionic compound. **Figure 2.14** shows two examples of ionic compounds that contain the carbonate ion.

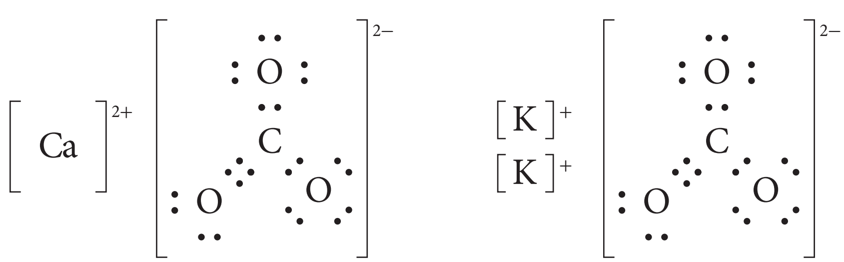


Figure 2.14 Polyatomic ions, like simple ions, must combine with oppositely charged ions that will give the final compound a neutral charge.

Positively Charged Polyatomic Ions

Positively charged polyatomic ions also exist, but the only one that is common is the ammonium ion. The ammonium ion forms when the molecular compound ammonia combines with a hydrogen ion, as shown in **Figure 2.15**. Ammonia has three bonded pairs and one lone pair. The hydrogen ion bonds with the lone pair to form an ammonium ion.

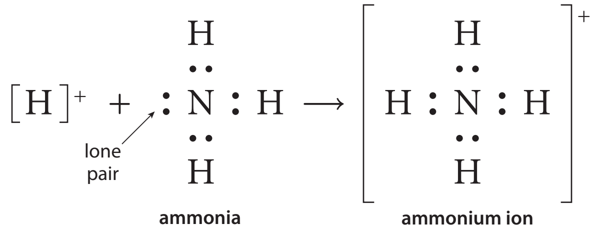


Figure 2.15 The ammonium ion is the only common positively charged polyatomic ion. Try not to confuse it with ammonia, which is a neutral molecular compound.

The ammonium ion forms ionic compounds by bonding with negatively charged ions. It can bond with a simple negatively charged ion, such as the chloride ion, as shown in **Figure 2.16 (A)**. It can also bond with a negatively charged polyatomic ion, such as the carbonate ion shown in **Figure 2.16 (B)**.

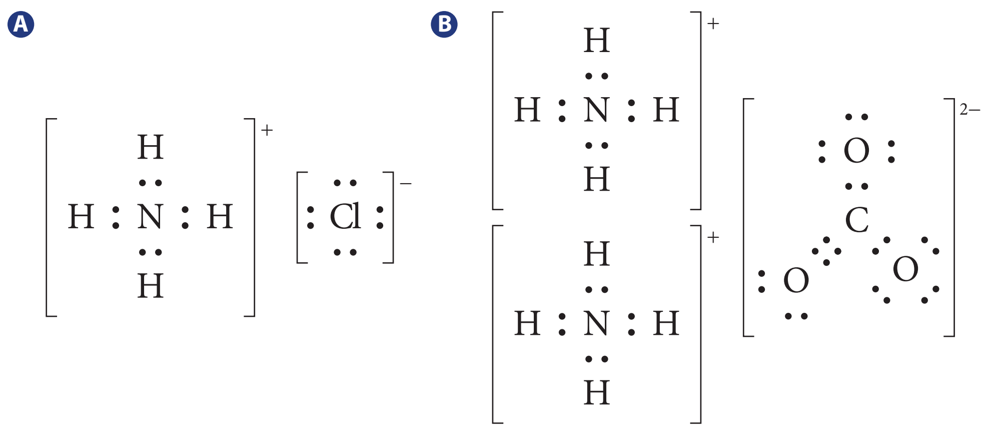


Figure 2.16 **(A)** The ammonium ion behaves like any other positively charged ion and forms ionic compounds, such as ammonium chloride, by bonding with negatively charged simple ions. **(B)** The ammonium ion can also form ionic compounds, such as ammonium carbonate, by bonding with negatively charged polyatomic ions.

Learning Check

1. State the octet rule, and give one example of how it can be applied.
2. When a calcium atom becomes ionized, it has a charge of 2+. When a bromine atom becomes ionized, it has a charge of 1–. Explain how ionic bonds can form between calcium and bromine to produce a compound that has a zero net charge.
3. Given a Lewis structure with four non-metal atoms, how would you determine whether it is a molecular compound with no charge or a polyatomic ion?
4. Draw a Lewis structure of two oxygen atoms that are covalently bonded together to form an oxygen molecule. Identify the bonding pairs and the lone pairs.
5. How do double bonds and triple bonds form? Why do they form?
6. Describe a situation in which two atoms that are covalently bonded together can be part of an ionic compound.

The Importance of Electronegativity in Bond Formation

Based on what you just learned about ionic bonds and covalent bonds, you might assume that they are two separate and distinct types of connections between atoms. However, like ionic bonds, covalent bonds also involve electrostatic attractions between positively charged nuclei and negatively charged electrons. To understand how electrostatic attraction influences the nature of bonds, recall the concept of electronegativity, which you learned about in Chapter 1.

Electronegativity is an indicator of the relative ability of an atom of a given element to attract shared electrons. Shared electrons constitute a covalent bond. Thus, the relative electronegativities of the elements of the two atoms that are bonded together should provide information about the nature of the bond. Although Lewis diagrams are drawn as though no electrons are shared between two nuclei in ionic compounds, the positively charged nucleus of each ion is attracting the negatively charged electrons of the other ions. Thus, the concept of electronegativity also applies to ionic compounds.

Electronegativity Difference and Bond Type

What do the relative electronegativities of elements tell you about the nature of bonds? If the electronegativity of one of the two atoms that are bonded together is greater than the electronegativity of the other atom, the electrons will be attracted more strongly to the first atom. In general, electrons spend more time around the atoms with the greater electronegativity.

Figure 2.17 illustrates a bond between a carbon atom and a chlorine atom. Of course, the carbon atom is bonded to other atoms as well as the chlorine atom. The electronegativity of the chlorine atom (3.2) is higher than the electronegativity of the carbon atom (2.6). The arrow indicates that the shared electrons are more strongly attracted to the chlorine atom, and thus spend more time there. The Greek letter delta, δ, is often used to represent "partial." Therefore, the symbols δ+ and δ− indicate that the carbon atom is partially positively charged and the chorine atom is partially negatively charged.

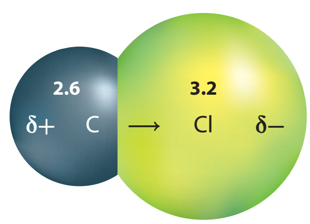


Figure 2.17 Because the shared electrons in this bond spend more time near the chlorine nucleus, the chlorine atom is slightly negatively charged. This leaves the carbon atom slightly positively charged.

Describe *How do you know that the electrons will spend more time near the chlorine atom than the carbon atom? Describe the data that tell you this.*

polar covalent bond a covalent bond around which there is an uneven distribution of electrons, making one end slightly positively charged and the other end slightly negatively charged

electronegativity difference the difference between the electronegativities of two atoms

Covalent bonds, in which the electron distribution is unequal, are called **polar covalent bonds**. These bonds are often referred to simply as polar bonds. Because these bonds have a positive "pole" and a negative "pole," they are sometimes also called *bond dipoles*. Depending on the difference in the electronegativities of the bonded atoms, some covalent bonds are only slightly polar while others are extremely polar. Chemists have devised a system for classifying the extent of the polarity of the bonds by calculating the **electronegativity difference** (ΔEN) for the two elements involved in the bond. You can calculate the electronegativity difference for any two elements by finding the electronegativity of each element in a table, such as the one in **Figure 1.22** on page 36, and then subtracting the smaller electronegativity from the larger electronegativity.

Applying Electronegativity Difference

As shown in **Figure 2.18**, bonds in which the electronegativity difference of the atoms is greater than 1.7 are classified as *mostly ionic*. The term "mostly" is used because there is always some attraction between the nucleus of one atom and the electrons of the other atom involved in the bond. If the electronegativity difference of two atoms that are bonded together is between 0.4 and 1.7, the bond is classified as *polar covalent*. If the electronegativity difference is less than 0.4, the bond is classified as *slightly polar covalent*. It is only when the electronegativity difference is zero that the bond can be classified as a *non-polar covalent* bond. The images on the right of **Figure 2.18** are electron cloud models of atoms bonded together. The image at the top shows a positive ion and a negative ion beside each other, indicating that the bond is mostly ionic. The next image shows atoms joined by a polar covalent bond. Chemists often use an arrow, like the one above this image, to show a polar bond. The tail of the arrow above this image looks like a plus sign, to signify the slightly positively charged end of the bond. The arrow points in the direction in which the electrons spend more time. The bottom image shows two atoms equally sharing electrons in a non-polar covalent bond.

The following examples show you how to use the information in **Figure 2.18** to calculate the electronegativity difference for two atoms. The first example involves the bond between a potassium atom and a fluorine atom. The electronegativity of fluorine is 4.0, and the electronegativity of potassium is 0.8. The electronegativity difference is calculated by subtracting the smaller number from the larger number, as shown below. Because 3.2 is much larger than 1.7, the bond between potassium and fluorine is mostly ionic.

$$\Delta EN = EN_{\mathrm{F}} - EN_{\mathrm{K}}$$
$$\Delta EN = 4.0 - 0.8$$
$$\Delta EN = 3.2$$

Next, consider the bond between two oxygen atoms in an oxygen molecule. The electronegativity of oxygen is 3.4. The electronegativity difference, as shown below, is zero. Therefore, the bond between two oxygen atoms is non-polar covalent.

$$\Delta EN = EN_{\mathrm{O}} - EN_{\mathrm{O}}$$
$$\Delta EN = 3.4 - 3.4$$
$$\Delta EN = 0.0$$

Finally, consider the bond between a carbon atom and a chlorine atom, discussed on the previous page. The electronegativity of carbon is 2.6, and the electronegativity of chlorine is 3.2. The electronegativity difference is 0.6, as shown below. This value is between 1.7 and 0.4, indicating that the bond is a polar covalent bond.

$$\Delta EN = EN_{\mathrm{Cl}} - EN_{\mathrm{C}}$$
$$\Delta EN = 3.2 - 2.6$$
$$\Delta EN = 0.6$$

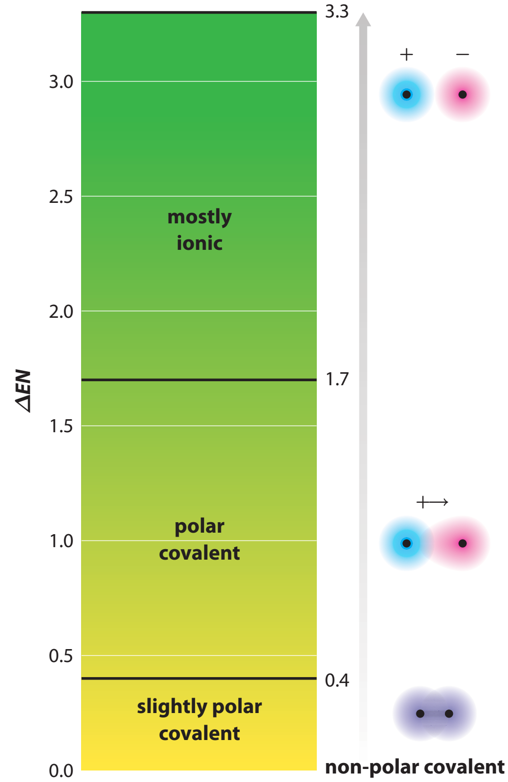


Figure 2.18 The shading in the diagram indicates that the character of bonds changes gradually from mostly ionic at the top to non-polar covalent at the bottom. The electronegativity difference values on the right are the transition points that separate the types of bonds. The images on the far right are models of compounds with the bond character in the different ranges of electronegativity difference.

Percent Ionic and Covalent Character

Chemists have devised another approach for describing the bond character, using percentages of either ionic or covalent character. **Table 2.1** relates electronegativity differences to *percent ionic character* and *percent covalent character*. In the following activity, you will analyze the relationship between electronegativity differences and percent ionic character.

Table 2.1 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

Classifying bond type is not always simple. The bond between a hydrogen atom and a chlorine atom provides a good example of overlap in ionic character and covalent character. The electronegativity difference for hydrogen and chlorine is 1.0, placing it in the polar covalent category. As a gas, the compound behaves as a polar molecule. When the compound is dissolved in water, however, the atoms become separate ions, both surrounded by water molecules. Thus, the bond type of this compound varies, depending on whether it is a gas or dissolved in water. **Figure 2.19** shows Lewis diagrams for the two states.

Figure 2.19 The electronegativity difference for hydrogen and chlorine indicates that a bond between these atoms results in a polar covalent molecule when HCl is in a gaseous state **(A)**. Its interaction with water molecules causes HCl to behave as an ionic compound **(B)**.

Activity 2.1 Electronegativity Difference versus Percent Ionic Character

Why did chemists choose the electronegativity differences of 0.4 and 1.7 for the transition points for slightly polar covalent, polar covalent, and mostly ionic bonds? Analyzing the relationship between electronegativity difference and percent ionic character in this activity will help you understand the reasons behind the choice of these values.

Materials

- graph paper
- ruler
- pencil

Procedure

1. Construct a graph using the data in the first two rows of **Table 2.1**. Put electronegativity difference on the *x*-axis and percent ionic character on the *y*-axis. Choose scales for the axes that will make the graph take up more than half of a sheet of graph paper.
2. After you plot all the points, draw a smooth curved line of best fit through the points.
3. Draw a straight, vertical line on the graph through the point where the electronegativity difference is 1.7. At the point at which the vertical line crosses the curve, draw a horizontal line across the graph. Record the value of the percent ionic character at the point where your horizontal line touches the axis.
4. Repeat step 3 for the point where the electronegativity difference is 0.4.

Questions

1. What is the percent ionic character when the electronegativity difference is 1.7? Do you think this is a reasonable value for the transition point between polar covalent and mostly ionic bonds? Explain your reasoning.
2. What is the percent ionic character when the electronegativity difference is 0.4? Do you think this is a reasonable value for the transition point between polar covalent and slightly polar covalent bonds? Explain your reasoning.
3. Why do you think it was important to make your graph spread out to more than half of the sheet of graph paper?
4. Imagine that you were to draw a graph of percent covalent character versus electronegativity difference. Predict the values of percent covalent character that you would find when the electronegativity differences are 0.4 and 1.7. Explain why you think you would find these results.

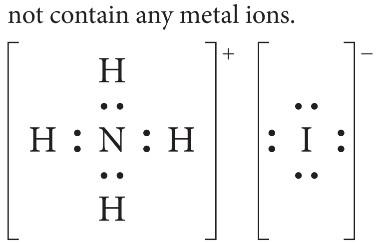
Section 2.1 REVIEW

Section Summary

- The octet rule can be used to predict how bonds will form.
- An ionic bond forms when a negatively charged ion and a positively charged ion are attracted to each other.
- A covalent bond forms when two atoms share one or more pairs of electrons.
- A polyatomic ion consists of two or more atoms that are covalently bonded together and carry a charge. A polyatomic ion can form an ionic compound with a simple ion or another polyatomic ion of the opposite charge.
- A chemical bond can be non-polar covalent, slightly polar covalent, polar covalent, or mostly ionic, depending on the electronegativity difference between the two atoms that are bonded together.

Review Questions

1. K/U What property of the noble gases led to the octet rule? Explain.
2. K/U Explain why metal atoms tend to lose electrons to form ions and why non-metal atoms tend to gain electrons to form ions.
3. C Draw Lewis diagrams of calcium and bromine. Use these diagrams to show how ionic bonds form between these atoms. Explain how these structures satisfy the octet rule.
4. T/I For each of the following, use Lewis diagrams to predict the number of atoms of each element that will be present in an ionic compound formed by the two elements.
   a. calcium and fluorine
   b. sodium and oxygen
   c. magnesium and nitrogen
5. T/I Draw Lewis diagrams of two oxygen atoms. Use your diagrams to show how an oxygen molecule forms from two oxygen atoms. Explain why there must be a double bond between the two oxygen atoms.
6. K/U How many electrons make up a triple bond?
7. K/U Draw a Lewis structure of a hydrogen atom covalently bonded to a fluorine atom. Identify all the bonding pairs and all the lone pairs.
8. K/U Assume that you are shown a Lewis structure with one nitrogen atom and three oxygen atoms. How would you determine whether the structure represented a neutral molecule or a polyatomic ion?
9. T/I Explain why the following compound can be considered an ionic compound, even though it does not contain any metal ions.



10. T/I Predict whether the bond between each pair of atoms will be non-polar covalent, slightly polar covalent, polar covalent, or mostly ionic.
    a. carbon and fluorine
    b. oxygen and nitrogen
    c. chlorine and chlorine
    d. copper and oxygen
    e. silicon and hydrogen
    f. sodium and fluorine
    g. iron and oxygen
    h. manganese and oxygen
11. T/I For each polar and slightly polar covalent bond in question 10, indicate the locations of the partial positive and partial negative charges. Explain how you made each decision.
12. T/I Arrange the bonds in each group below in order of increasing polarity.
    a. hydrogen bonded to chlorine, oxygen bonded to nitrogen, carbon bonded to sulfur, sodium bonded to chlorine
    b. carbon bonded to chlorine, magnesium bonded to chlorine, phosphorus bonded to oxygen, nitrogen bonded to nitrogen
13. C Make a sketch that shows the relationship between electronegativity difference and percent ionic character of a chemical bond. Why do you think that the transition points between types of chemical bonds are reported in electronegativity difference rather than percent ionic character?
14. K/U Explain the meaning of the symbol above these chemical symbols. $\overset{+\!\!\longrightarrow}{\mathrm{NO}}$
15. A Toward the beginning of this section, you read that metals are usually found in combination with non-metals in nature, and that these compounds are solid. From what you now know, how would you classify these compounds? Give an example.
16. A The atmosphere consists mostly of nitrogen and oxygen, along with small amounts of carbon dioxide and trace amounts of hydrogen. Does the atmosphere consist almost entirely of polar compounds or non-polar compounds? Explain your reasoning.