CHAPTER 11

Chemical Bonds: The Formation of Compounds from Atoms

This colorfully lighted limestone cave reveals dazzling stalactites and stalagmites, formed from calcium carbonate.

Chapter Outline

For centuries we've been aware that certain metals cling to a magnet. We've seen balloons sticking to walls. Why? High-speed levitation trains are heralded to be the wave of the future. How do they function? In each case, forces of attraction and repulsion are at work.

Human interactions also suggest that "opposites attract" and "likes repel." Attractions draw us into friendships and significant relationships, whereas repulsive forces may produce debate and antagonism. We form and break apart interpersonal bonds throughout our lives.

In chemistry, we also see this phenomenon. Substances form chemical bonds as a result of electrical attractions. These bonds provide the tremendous diversity of compounds found in nature.

11.1 Periodic Trends in Atomic Properties

Although atomic theory and electron configuration help us understand the arrangement and behavior of the elements, it's important to remember that the design of the periodic table is based on observing properties of the elements. Before we use the concept of atomic structure to explain how and why atoms combine to form compounds, we need to understand the characteristic properties of the elements and the trends that occur in these properties on the periodic table. These trends allow us to use the periodic table to accurately predict properties and reactions of a wide variety of substances.

Metals and Nonmetals

In Section 3.5, we classified elements as metals, nonmetals, or metalloids. The heavy stair-step line beginning at boron and running diagonally down the periodic table separates the elements into metals and nonmetals. Metals are usually lustrous, malleable, and good conductors of heat and electricity. Nonmetals are just the opposite—nonlustrous, brittle, and poor conductors. Metalloids are found bordering the heavy diagonal line and may have properties of both metals and nonmetals.

Most elements are classified as metals (see Figure 11.1). Metals are found on the left side of the stair-step line, while the nonmetals are located toward the upper right of the table. Note that hydrogen does not fit into the division of metals and nonmetals. It displays nonmetallic properties under normal conditions, even though it has only one outermost electron like the alkali metals. Hydrogen is considered to be a unique element.

It is the chemical properties of metals and nonmetals that interest us most. Metals tend to lose electrons and form positive ions, while nonmetals tend to gain electrons and form negative ions. When a metal reacts with a nonmetal, electrons are often transferred from the metal to the nonmetal.

Atomic Radius

The relative radii of the representative elements are shown in Figure 11.2. Notice that the radii of the atoms tend to increase down each group and that they tend to decrease from left to right across a period.

The increase in radius down a group can be understood if we consider the electron structure of the atoms. For each step down a group, an additional

Metals
Metalloids
Nonmetals

1 H																	2 He
3 Li	4 Be											5 B	6 C	7 N	8 O	9 F	10 Ne
11 Na	12 Mg											13 Al	14 Si	15 P	16 S	17 Cl	18 Ar
19 K	20 Ca	21 Sc	22 Ti	23 V	24 Cr	25 Mn	26 Fe	27 Co	28 Ni	29 Cu	30 Zn	31 Ga	32 Ge	33 As	34 Se	35 Br	36 Kr
37 Rb	38 Sr	39 Y	40 Zr	41 Nb	42 Mo	43 Tc	44 Ru	45 Rh	46 Pd	47 Ag	48 Cd	49 In	50 Sn	51 Sb	52 Te	53 I	54 Xe
55 Cs	56 Ba	57 La*	72 Hf	73 Ta	74 W	75 Re	76 Os	77 Ir	78 Pt	79 Au	80 Hg	81 Tl	82 Pb	83 Bi	84 Po	85 At	86 Rn
87 Fr	88 Ra	89 Ac†	104 Rf	105 Db	106 Sg	107 Bh	108 Hs	109 Mt	110 Ds	111 Rg							

*	58 Ce	59 Pr	60 Nd	61 Pm	62 Sm	63 Eu	64 Gd	65 Tb	66 Dy	67 Ho	68 Er	69 Tm	70 Yb	71 Lu
†	90 Th	91 Pa	92 U	93 Np	94 Pu	95 Am	96 Cm	97 Bk	98 Cf	99 Es	100 Fm	101 Md	102 No	103 Lr

Figure 11.1
The elements are classified as metals, nonmetals, and metalloids.

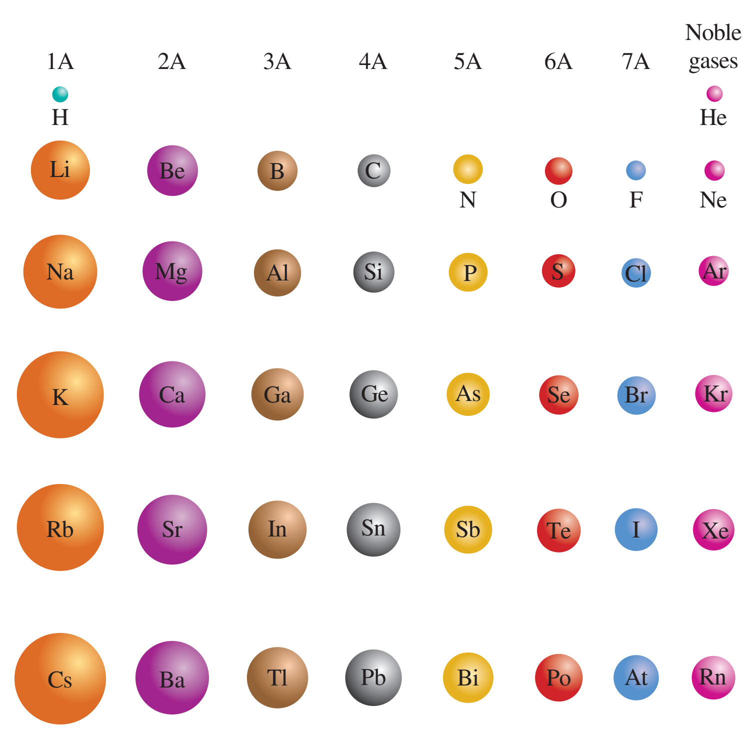


Figure 11.2
Relative atomic radii for the representative elements. Atomic radius decreases across a period and increases down a group in the periodic table.

energy level is added to the atom. The average distance from the nucleus to the outside edge of the atom must increase as each new energy level is added. The atoms get bigger as electrons are placed in these new higher-energy levels.

Understanding the decrease in atomic radius across a period requires more thought, however. As we move from left to right across a period, electrons within the same block are being added to the same principal energy level. Within a given energy level, we expect the orbitals to have about the same size. We would then expect the atoms to be about the same size across the period. But each time an electron is added, a proton is added to the nucleus as well. The increase in positive charge (in the nucleus) pulls the electrons closer to the nucleus, which results in a gradual decrease in atomic radius across a period.

Ionization Energy

ionization energy

The **ionization energy** of an atom is the energy required to remove an electron from the atom. For example,

$$\text{Na} + \text{ionization energy} \longrightarrow \text{Na}^+ + e^-$$

The first ionization energy is the amount of energy required to remove the first electron from an atom, the second is the amount required to remove the second electron from that atom, and so on.

Table 11.1 gives the ionization energies for the removal of one to five electrons from several elements. The table shows that even higher amounts of energy are needed to remove the second, third, fourth, and fifth electrons. This makes sense because removing electrons leaves fewer electrons attracted to the same positive charge in the nucleus. The data in Table 11.1 also show that an extra-large ionization energy (blue) is needed when an electron is removed from a noble gas–like structure, clearly showing the stability of the electron structure of the noble gases.

First ionization energies have been experimentally determined for most elements. Figure 11.3 plots these energies for representative elements in the first four periods. Note these important points:

1. Ionization energy in Group A elements decreases from top to bottom in a group. For example, in Group 1A the ionization energy changes from 520 kJ/mol for Li to 419 kJ/mol for K.
2. Ionization energy gradually increases from left to right across a period. Noble gases have a relatively high value, confirming the nonreactive nature of these elements.

Table 11.1 Ionization Energies for Selected Elements*

	Required amounts of energy (kJ/mol)				
Element	1st e^-	2nd e^-	3rd e^-	4th e^-	5th e^-
H	1,314				
He	2,372	5,247			
Li	520	7,297	11,810		
Be	900	1,757	14,845	21,000	
B	800	2,430	3,659	25,020	32,810
C	1,088	2,352	4,619	6,222	37,800
Ne	2,080	3,962	6,276	9,376	12,190
Na	496	4,565	6,912	9,540	13,355

*Values are expressed in kilojoules per mole, showing energies required to remove 1 to 5 electrons per atom. Blue type indicates the energy needed to remove an electron from a noble gas electron structure.

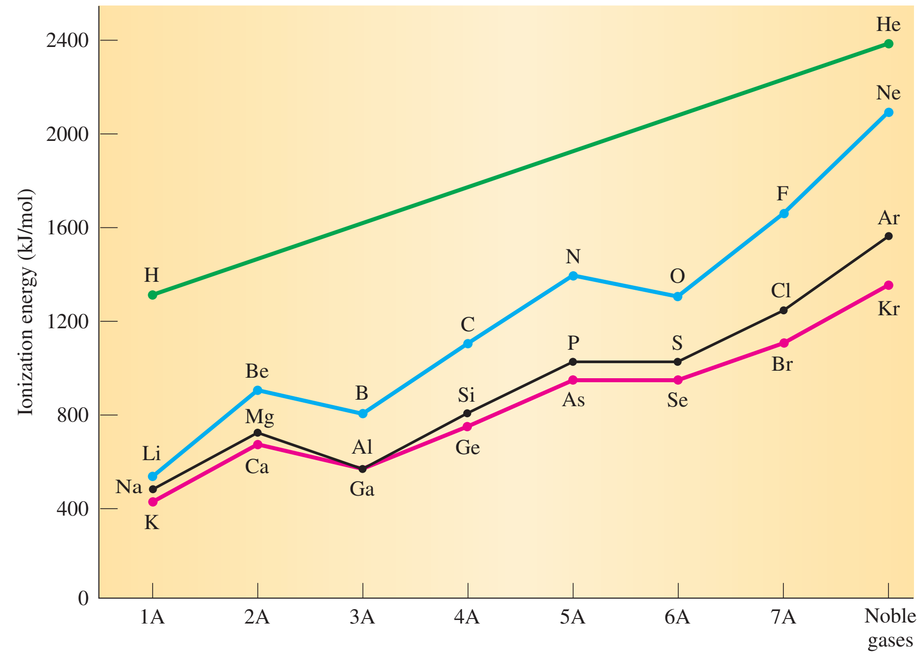


Figure 11.3
Periodic relationship of the first ionization energy for representative elements in the first four periods.

Metals don't behave in exactly the same manner. Some metals give up electrons much more easily than others. In the alkali metal family, cesium gives up its 6*s* electron much more easily than the metal lithium gives up its 2*s* electron. This makes sense when we consider that the size of the atoms increases down the group. The distance between the nucleus and the outer electrons increases and the ionization energy decreases. The most chemically active metals are located at the lower left of the periodic table.

Nonmetals have relatively large ionization energies compared to metals. Nonmetals tend to gain electrons and form anions. Since the nonmetals are located at the right side of the periodic table, it is not surprising that ionization energies tend to increase from left to right across a period. The most active nonmetals are found in the *upper* right corner of the periodic table (excluding the noble gases).

11.2 Lewis Structures of Atoms

Metals tend to form cations (positively charged ions) and nonmetals form anions (negatively charged ions) in order to attain a stable valence electron structure. For many elements this stable valence level contains eight electrons (two *s* and six *p*), identical to the valence electron configuration of the noble gases. Atoms undergo rearrangements of electron structure to lower their chemical potential energy (or to become more stable). These rearrangements are accomplished by losing, gaining, or sharing electrons with other atoms. For example, a hydrogen atom could accept a second electron and attain an electron structure the same as the noble gas helium. A fluorine atom could gain an electron and attain an electron structure like neon. A sodium atom could lose one electron to attain an electron structure like neon.

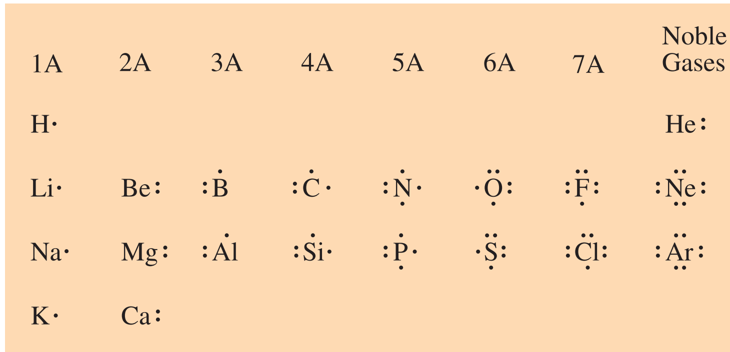
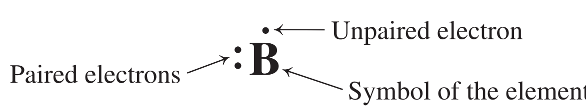


Figure 11.4
Lewis structures of the first 20 elements. Dots represent electrons in the outermost *s* and *p* energy levels only.

Lewis structure

The valence electrons in the outermost energy level of an atom are responsible for the electron activity that occurs to form chemical bonds. The **Lewis structure** of an atom is a representation that shows the valence electrons for that atom. American chemist Gilbert N. Lewis (1875–1946) proposed using the symbol for the element and dots for electrons. The number of dots placed around the symbol equals the number of *s* and *p* electrons in the outermost energy level of the atom. Paired dots represent paired electrons; unpaired dots represent unpaired electrons. For example, **H·** is the Lewis symbol for a hydrogen atom, $1s^1$; **:B** is the Lewis symbol for a boron atom, with valence electrons $2s^22p^1$. In the case of boron, the symbol B represents the boron nucleus and the $1s^2$ electrons; the dots represent only the $2s^22p^1$ electrons.



The Lewis method is used not only because of its simplicity of expression but also because much of the chemistry of the atom is directly associated with the electrons in the outermost energy level. Figure 11.4 shows Lewis structures for the elements hydrogen through calcium.

Example 11.1

Write the Lewis structure for a phosphorus atom.

SOLUTION

First establish the electron structure for a phosphorus atom, which is $1s^22s^22p^63s^23p^3$. Note that there are five electrons in the outermost energy level; they are $3s^23p^3$. Write the symbol for phosphorus and place the five electrons as dots around it.

The $3s^2$ electrons are paired and are represented by the paired dots. The $3p^3$ electrons, which are unpaired, are represented by the single dots.

A quick way to determine the correct number of dots (electrons) for a Lewis structure is to use the Group number. For the A groups on the periodic table, the Group number is the same as the number of electrons in the Lewis structure.

Practice 11.1

Write the Lewis structure for the following elements:
(a) N (b) Al (c) Sr (d) Br

11.3 The Ionic Bond: Transfer of Electrons from One Atom to Another

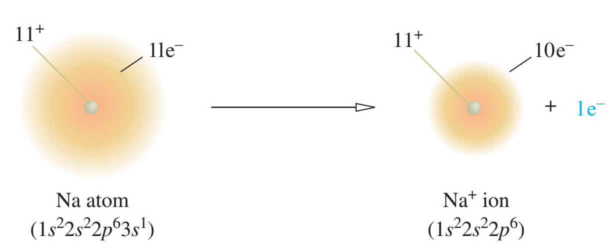
The chemistry of many elements, especially the representative ones, is to attain an outer electron structure like that of the chemically stable noble gases. With the exception of helium, this stable structure consists of eight electrons in the outermost energy level (see Table 11.2).

Table 11.2 Arrangement of Electrons in the Noble Gases*

Noble gas	Symbol	Electron structure					
		$n = 1$	2	3	4	5	6
Helium	He	$\mathbf{1s^2}$					
Neon	Ne	$1s^2$	$\mathbf{2s^22p^6}$				
Argon	Ar	$1s^2$	$2s^22p^6$	$\mathbf{3s^23p^6}$			
Krypton	Kr	$1s^2$	$2s^22p^6$	$3s^23p^63d^{10}$	$\mathbf{4s^24p^6}$		
Xenon	Xe	$1s^2$	$2s^22p^6$	$3s^23p^63d^{10}$	$4s^24p^64d^{10}$	$\mathbf{5s^25p^6}$	
Radon	Rn	$1s^2$	$2s^22p^6$	$3s^23p^63d^{10}$	$4s^24p^64d^{10}4f^{14}$	$5s^25p^65d^{10}$	$\mathbf{6s^26p^6}$

*Each gas except helium has eight electrons in its outermost energy level.

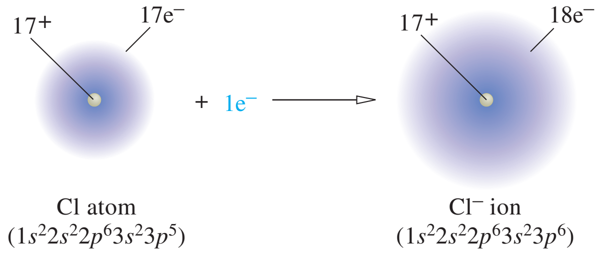
Let's look at the electron structures of sodium and chlorine to see how each element can attain a structure of 8 electrons in its outermost energy level. A sodium atom has 11 electrons: 2 in the first energy level, 8 in the second energy level, and 1 in the third energy level. A chlorine atom has 17 electrons: 2 in the first energy level, 8 in the second energy level, and 7 in the third energy level. If a sodium atom transfers or loses its 3*s* electron, its third energy level becomes vacant, and it becomes a sodium ion with an electron configuration identical to that of the noble gas neon. This process requires energy:



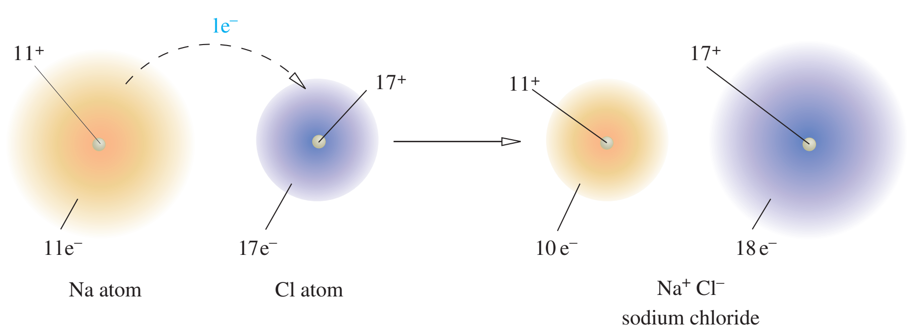
An atom that has lost or gained electrons will have a positive or negative charge, depending on which particles (protons or electrons) are in excess. Remember that a charged particle or group of particles is called an *ion*.

By losing a negatively charged electron, the sodium atom becomes a positively charged particle known as a sodium ion. The charge, +1, results because the nucleus still contains 11 positively charged protons, and the electron orbitals contain only 10 negatively charged electrons. The charge is indicated by a plus sign (+) and is written as a superscript after the symbol of the element (Na^+).

A chlorine atom with seven electrons in the third energy level needs one electron to pair up with its one unpaired 3*p* electron to attain the stable outer electron structure of argon. By gaining one electron, the chlorine atom becomes a chloride ion (Cl^-), a negatively charged particle containing 17 protons and 18 electrons. This process releases energy:



Consider sodium and chlorine atoms reacting with each other. The 3*s* electron from the sodium atom transfers to the half-filled 3*p* orbital in the chlorine atom to form a positive sodium ion and a negative chloride ion. The compound sodium chloride results because the Na^+ and Cl^- ions are strongly attracted to each other by their opposite electrostatic charges. The force holding the oppositely charged ions together is called an ionic bond:



The Lewis representation of sodium chloride formation is

$$Na\cdot + \cdot\ddot{\underset{\cdot\cdot}{Cl}}\!: \longrightarrow [Na]^+\left[:\ddot{\underset{\cdot\cdot}{Cl}}\!:\right]^-$$

The chemical reaction between sodium and chlorine is a very vigorous one, producing considerable heat in addition to the salt formed. When energy is released in a chemical reaction, the products are more stable than the reactants. Note that in NaCl both atoms attain a noble gas electron structure.

Sodium chloride is made up of cubic crystals in which each sodium ion is surrounded by six chloride ions and each chloride ion by six sodium ions, except at the crystal surface. A visible crystal is a regularly arranged aggregate of millions of these ions, but the ratio of sodium to chloride ions is 1 : 1, hence the formula NaCl. The cubic crystalline lattice arrangement of sodium chloride is shown in Figure 11.5.

Figure 11.6 contrasts the relative sizes of sodium and chlorine atoms with those of their ions. The sodium ion is smaller than the atom due primarily to two factors: (1) The sodium atom has lost its outermost electron, thereby

These tiny NaCl crystals (on a penny) show the cubic structure illustrated in Figure 11.5.

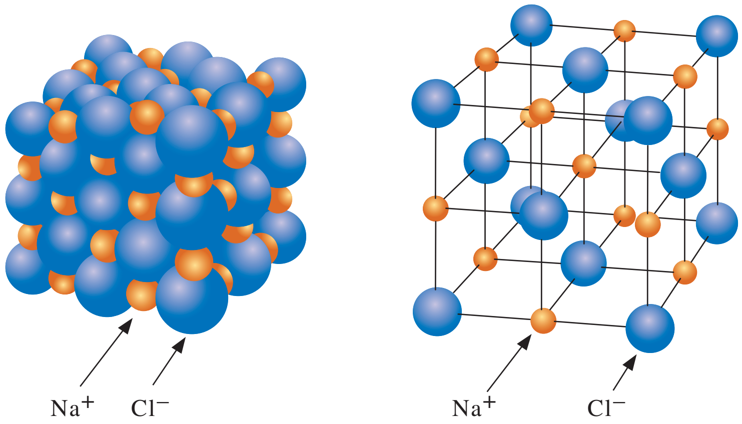


Figure 11.5
Sodium chloride crystal. Diagram represents a small fragment of sodium chloride, which forms cubic crystals. Each sodium ion is surrounded by six chloride ions, and each chloride ion is surrounded by six sodium ions. The tiny NaCl crystals show the cubic crystal structure of salt.

Remember: A cation is always smaller than its parent atom whereas an anion is always larger than its parent atom.

reducing its size; and (2) the 10 remaining electrons are now attracted by 11 protons and are thus drawn closer to the nucleus. Conversely, the chloride ion is larger than the atom because (1) it has 18 electrons but only 17 protons and (2) the nuclear attraction on each electron is thereby decreased, allowing the chlorine atom to expand as it forms an ion.

ionic bond

We've seen that when sodium reacts with chlorine, each atom becomes an ion. Sodium chloride, like all ionic substances, is held together by the attraction existing between positive and negative charges. An **ionic bond** is the attraction between oppositely charged ions.

Ionic bonds are formed whenever one or more electrons are transferred from one atom to another. Metals, which have relatively little attraction for their valence electrons, tend to form ionic bonds when they combine with nonmetals.

It's important to recognize that substances with ionic bonds do not exist as molecules. In sodium chloride, for example, the bond does not exist solely between a single sodium ion and a single chloride ion. Each sodium ion in the crystal attracts six near-neighbor negative chloride ions; in turn, each negative chloride ion attracts six near-neighbor positive sodium ions (see Figure 11.5).

A metal will usually have one, two, or three electrons in its outer energy level. In reacting, metal atoms characteristically lose these electrons, attain

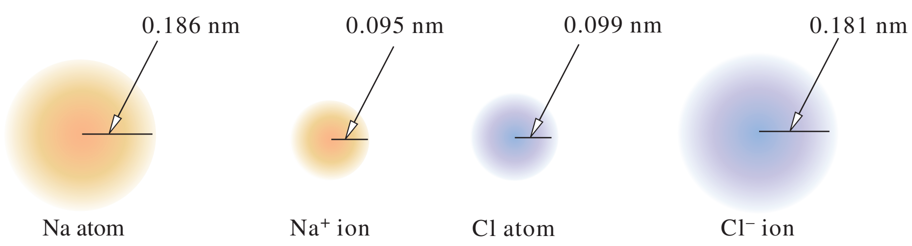


Figure 11.6
Relative radii of sodium and chlorine atoms and their ions.

Table 11.3 Change in Atomic Radii (nm) of Selected Metals and Nonmetals*

Atomic radius		Ionic radius		Atomic radius		Ionic radius	
Li	0.152	Li^+	0.060	F	0.071	F^-	0.136
Na	0.186	Na^+	0.095	Cl	0.099	Cl^-	0.181
K	0.227	K^+	0.133	Br	0.114	Br^-	0.195
Mg	0.160	Mg^{2+}	0.065	O	0.074	O^{2-}	0.140
Al	0.143	Al^{3+}	0.050	S	0.103	S^{2-}	0.184

*The metals shown lose electrons to become positive ions. The nonmetals gain electrons to become negative ions.

the electron structure of a noble gas, and become positive ions. A nonmetal, on the other hand, is only a few electrons short of having a noble gas electron structure in its outer energy level and thus has a tendency to gain electrons. In reacting with metals, nonmetal atoms characteristically gain one, two, or three electrons; attain the electron structure of a noble gas; and become negative ions. The ions formed by loss of electrons are much smaller than the corresponding metal atoms; the ions formed by gaining electrons are larger than the corresponding nonmetal atoms. The dimensions of the atomic and ionic radii of several metals and nonmetals are given in Table 11.3.

Practice 11.2

What noble gas structure is formed when an atom of each of these metals loses all its valence electrons? Write the formula for the metal ion formed.
(a) K (b) Mg (c) Al (d) Ba

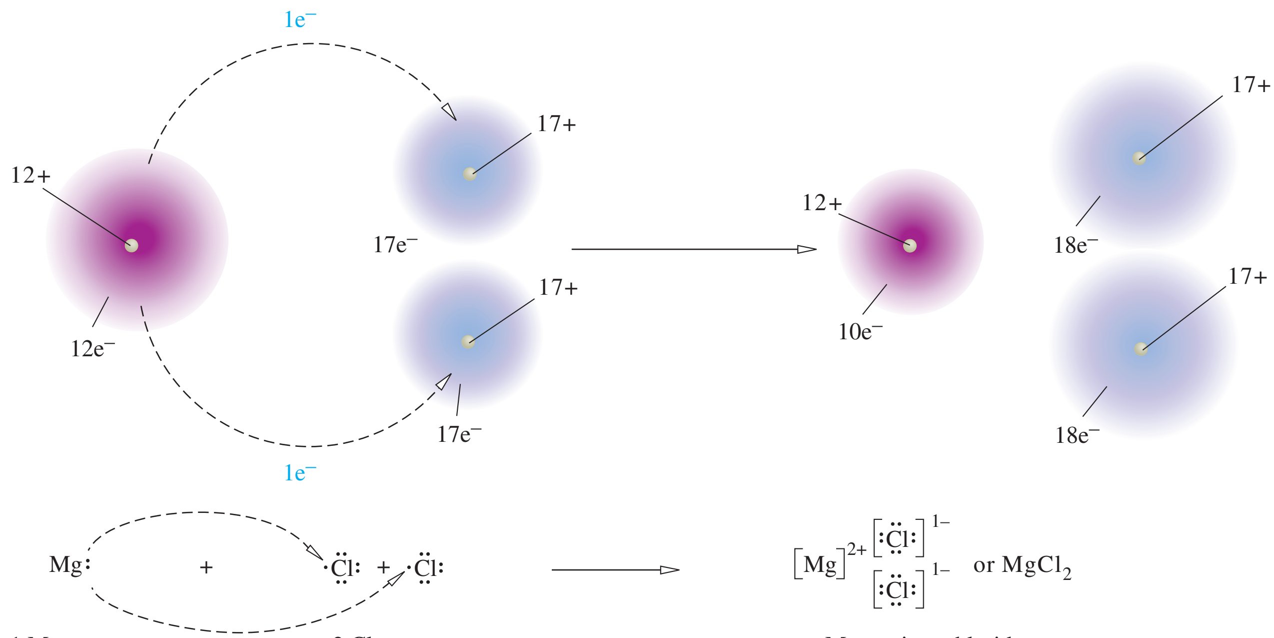
Study the following examples. Note the loss and gain of electrons between atoms; also note that the ions in each compound have a noble gas electron structure.

Example 11.2

Explain how magnesium and chlorine combine to form magnesium chloride, $MgCl_2$.

SOLUTION

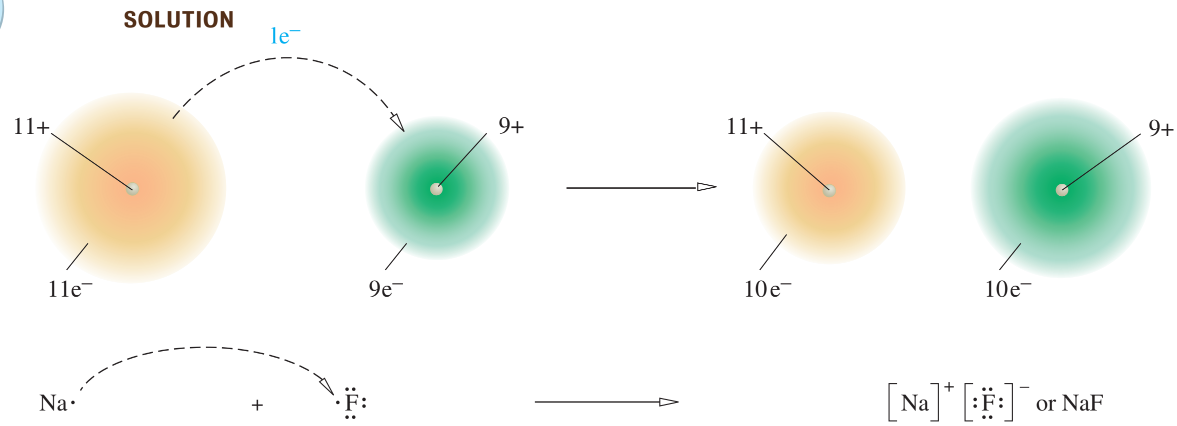
A magnesium atom of electron structure $1s^22s^22p^63s^2$ must lose two electrons or gain six to reach a stable electron structure. If magnesium reacts with chlorine and each chlorine atom can accept only one electron, two chlorine atoms will be needed for the two electrons from each magnesium atom. The compound formed will contain one magnesium ion and two chloride ions. The magnesium atom, having lost two electrons, becomes a magnesium ion with a +2 charge. Each chloride ion will have a −1 charge. The transfer of electrons from a magnesium atom to two chlorine atoms is shown in the following illustration:



1 Mg atom 2 Cl atoms Magnesium chloride

Example 11.3 Explain the formation of sodium fluoride (NaF) from its elements.

SOLUTION

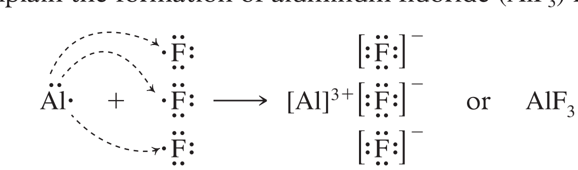


Na atom F atom Sodium fluoride

The fluorine atom, with seven electrons in its outer energy level, behaves similarly to the chlorine atom.

Example 11.4 Explain the formation of aluminum fluoride (AlF_3) from its elements.

SOLUTION



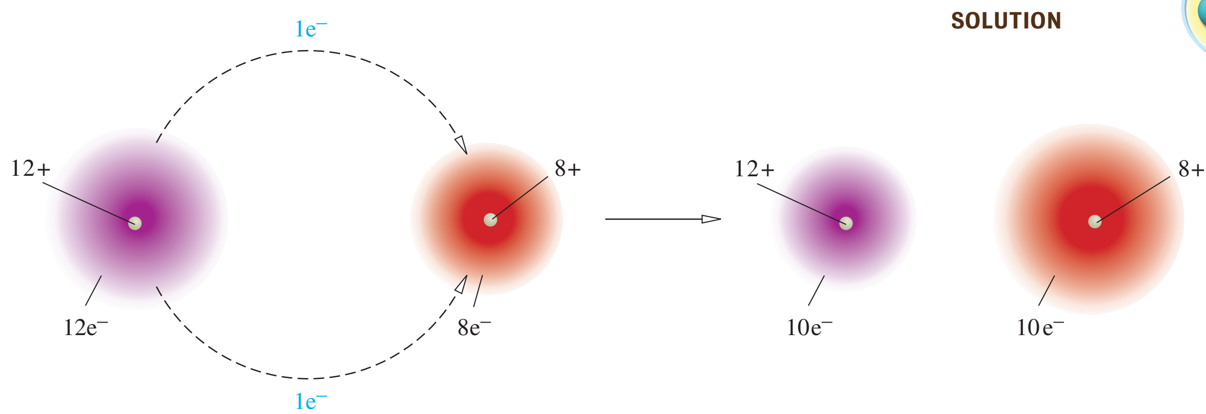
1 Al atom 3 F atoms aluminum flouride

Each fluorine atom can accept only one electron. Therefore, three fluorine atoms are needed to combine with the three valence electrons of one aluminum atom. The aluminum atom has lost three electrons to become an aluminum ion (Al^{3+}) with a +3 charge.

Example 11.5

Explain the formation of magnesium oxide (MgO) from its elements.

SOLUTION



$$Mg\!: \; + \; \cdot\ddot{O}\!: \longrightarrow [Mg]^{2+}[:\ddot{O}:]^{2-} \text{ or } MgO$$

Mg atom   O atom   Magnesium oxide

The magnesium atom, with two electrons in the outer energy level, exactly fills the need of one oxygen atom for two electrons. The resulting compound has a ratio of one magnesium atom to one oxygen atom. The oxygen (oxide) ion has a −2 charge, having gained two electrons. In combining with oxygen, magnesium behaves the same way as when it combines with chlorine—it loses two electrons.

Example 11.6

Explain the formation of sodium sulfide (Na_2S) from its elements.

SOLUTION

$$\begin{matrix} Na\cdot \\ Na\cdot \end{matrix} + \cdot\ddot{S}\!: \longrightarrow \begin{matrix} [Na]^{+} \\ [Na]^{+} \end{matrix}[:\ddot{S}:]^{2-} \quad \text{or} \quad Na_2S$$

2 Na atoms   1 S atom   sodium sulfide

Two sodium atoms supply the electrons that one sulfur atom needs to make eight in its outer energy level.

Example 11.7

Explain the formation of aluminum oxide (Al_2O_3) from its elements.

SOLUTION

$$\begin{matrix} \dot{Al}\cdot \\ \\ \dot{Al}\cdot \end{matrix} + \begin{matrix} \cdot\ddot{O}\!: \\ \cdot\ddot{O}\!: \\ \cdot\ddot{O}\!: \end{matrix} \longrightarrow \begin{matrix} [Al]^{3+} \\ [Al]^{3+} \end{matrix} \begin{matrix} [:\ddot{O}:]^{2-} \\ [:\ddot{O}:]^{2-} \\ [:\ddot{O}:]^{2-} \end{matrix} \quad \text{or} \quad Al_2O_3$$

2 Al atoms   3 O atoms   aluminum oxide

One oxygen atom, needing two electrons, cannot accommodate the three electrons from one aluminum atom. One aluminum atom falls one electron short of the four electrons needed by two oxygen atoms. A ratio of two atoms of aluminum to three atoms of oxygen, involving the transfer of six electrons (two to each oxygen atom), gives each atom a stable electron configuration.

Note that in each of these examples, outer energy levels containing eight electrons were formed in all the negative ions. This formation resulted from the pairing of all the *s* and *p* electrons in these outer energy levels.

11.4 Predicting Formulas of Ionic Compounds

In previous examples, we learned that when a metal and a nonmetal react to form an ionic compound, the metal loses one or more electrons to the nonmetal. In Chapter 6, where we learned to name compounds and write formulas, we saw that Group 1A metals always form +1 cations, whereas Group 2A form +2 cations. Group 7A elements form −1 anions and Group 6A form −2 anions.

It stands to reason, then, that this pattern is directly related to the stability of the noble gas configuration. Metals lose electrons to attain the electron configuration of a noble gas (the previous one on the periodic table). A nonmetal forms an ion by gaining enough electrons to achieve the electron configuration of the noble gas following it on the periodic table. These observations lead us to an important chemical principle:

> In almost all stable chemical compounds of representative elements, each atom attains a noble gas electron configuration. This concept forms the basis for our understanding of chemical bonding.

We can apply this principle in predicting the formulas of ionic compounds. To predict the formula of an ionic compound, we must recognize that chemical compounds are always electrically neutral. In addition, the metal will lose electrons to achieve noble gas configuration and the nonmetal will gain electrons to achieve noble gas configuration. Consider the compound formed between barium and sulfur. Barium has two valence electrons, whereas sulfur has six valence electrons:

Ba $[Xe]6s^2$ S $[Ne]3s^23p^4$

If barium loses two electrons, it will achieve the configuration of xenon. By gaining two electrons, sulfur achieves the configuration of argon. Consequently, a pair of electrons is transferred between atoms. Now we have Ba^{2+} and S^{2-}. Since compounds are electrically neutral, there must be a ratio of one Ba to one S, giving the formula BaS.

The same principle works for many other cases. Since the key lies in the electron configuration, the periodic table can be used to extend the prediction even further. Because of similar electron structures, the elements in a family

Table 11.4 Formulas of Compounds Formed by Alkali Metals

Lewis structure	Oxides	Chlorides	Bromides	Sulfates
Li$\cdot$	Li_2O	LiCl	LiBr	Li_2SO_4
Na$\cdot$	Na_2O	NaCl	NaBr	Na_2SO_4
K$\cdot$	K_2O	KCl	KBr	K_2SO_4
Rb$\cdot$	Rb_2O	RbCl	RbBr	Rb_2SO_4
Cs$\cdot$	Cs_2O	CsCl	CsBr	Cs_2SO_4

generally form compounds with the same atomic ratios. In general, if we know the atomic ratio of a particular compound—say, NaCl—we can predict the atomic ratios and formulas of the other alkali metal chlorides. These formulas are LiCl, KCl, RbCl, CsCl, and FrCl (see Table 11.4).

Similarly, if we know that the formula of the oxide of hydrogen is H_2O, we can predict that the formula of the sulfide will be H_2S, because sulfur has the same valence electron structure as oxygen. Recognize, however, that these are only predictions; it doesn't necessarily follow that every element in a group will behave like the others or even that a predicted compound will actually exist. For example, knowing the formulas for potassium chlorate, bromate, and iodate to be $KClO_3$, $KBrO_3$, and KIO_3, we can correctly predict the corresponding sodium compounds to have the formulas $NaClO_3$, $NaBrO_3$, and $NaIO_3$. Fluorine belongs to the same family of elements (Group 7A) as chlorine, bromine, and iodine, so it would appear that fluorine should combine with potassium and sodium to give fluorates with the formulas KFO_3 and $NaFO_3$. However, potassium and sodium fluorates are not known to exist.

In the discussion in this section, we refer only to representative metals (Groups 1A, 2A, and 3A). The transition metals (Group B) show more complicated behavior (they form multiple ions), and their formulas are not as easily predicted.

Example 11.8

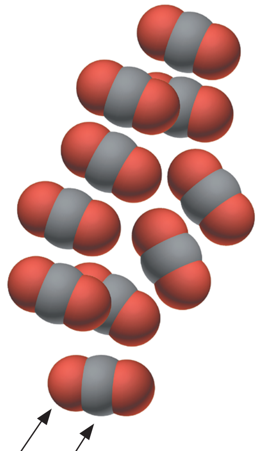
The formula for calcium sulfide is CaS and that for lithium phosphide is Li_3P. Predict formulas for (a) magnesium sulfide, (b) potassium phosphide, and (c) magnesium selenide.

SOLUTION

(a) Look up calcium and magnesium in the periodic table; they are both in Group 2A. The formula for calcium sulfide is CaS, so it's reasonable to predict that the formula for magnesium sulfide is MgS.

(b) Find lithium and potassium in the periodic table; they are in Group 1A. Since the formula for lithium phosphide is Li_3P, it's reasonable to predict that K_3P is the formula for potassium phosphide.

(c) Find selenium in the periodic table; it is in Group 6A just below sulfur. Therefore it's reasonable to assume that selenium forms selenide in the same way that sulfur forms sulfide. Since MgS was the predicted formula for magnesium sulfide in part (a), we can reasonably assume that the formula for magnesium selenide is MgSe.



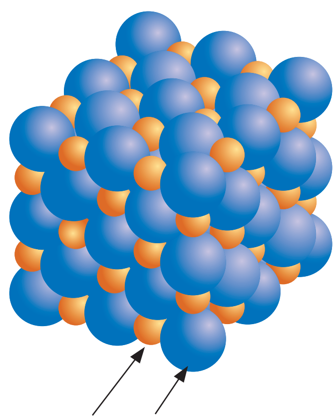


Figure 11.7
Solid carbon dioxide (dry ice) is composed of individual covalently bonded molecules of CO_2 closely packed together. Table salt is a large aggregate of Na^+ and Cl^- ions instead of molecules.

Practice 11.3

The formula for sodium oxide is Na_2O. Predict the formula for
(a) sodium sulfide
(b) rubidium oxide

Practice 11.4

The formula for barium phosphide is Ba_3P_2. Predict the formula for
(a) magnesium nitride
(b) barium arsenide

11.5 The Covalent Bond: Sharing Electrons

Some atoms do not transfer electrons from one atom to another to form ions. Instead they form a chemical bond by sharing pairs of electrons between them. A **covalent bond** consists of a pair of electrons shared between two atoms. This bonding concept was introduced in 1916 by G. N. Lewis. In the millions of known compounds, the covalent bond is the predominant chemical bond.

covalent bond

True molecules exist in substances in which the atoms are covalently bonded. It is proper to refer to molecules of such substances as hydrogen, chlorine, hydrogen chloride, carbon dioxide, water, or sugar (Figure 11.7). These substances contain only covalent bonds and exist as aggregates of molecules. We don't use the term *molecule* when talking about ionically bonded compounds such as sodium chloride, because such substances exist as large aggregates of positive and negative ions, not as molecules (Figure 11.7).

A study of the hydrogen molecule gives us an insight into the nature of the covalent bond and its formation. The formation of a hydrogen molecule (H_2) involves the overlapping and pairing of 1*s* electron orbitals from two hydrogen atoms, shown in Figure 11.8. Each atom contributes one electron of the pair that is shared jointly by two hydrogen nuclei. The orbital of the electrons now includes both hydrogen nuclei, but probability factors show that the most likely place to find the electrons (the point of highest electron density) is between the two nuclei. The two nuclei are shielded from each other by the pair of electrons, allowing the two nuclei to be drawn very close to each other.

The formula for chlorine gas is Cl_2. When the two atoms of chlorine combine to form this molecule, the electrons must interact in a manner similar to that shown in the hydrogen example. Each chlorine atom would be more stable with eight electrons in its outer energy level. But chlorine atoms are identical, and neither is able to pull an electron away from the other. What happens

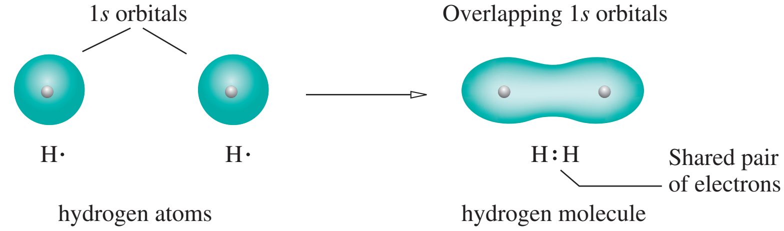


Figure 11.8
The formation of a hydrogen molecule from two hydrogen atoms. The two 1*s* orbitals overlap, forming the H_2 molecule. In this molecule the two electrons are shared between the atoms, forming a covalent bond.

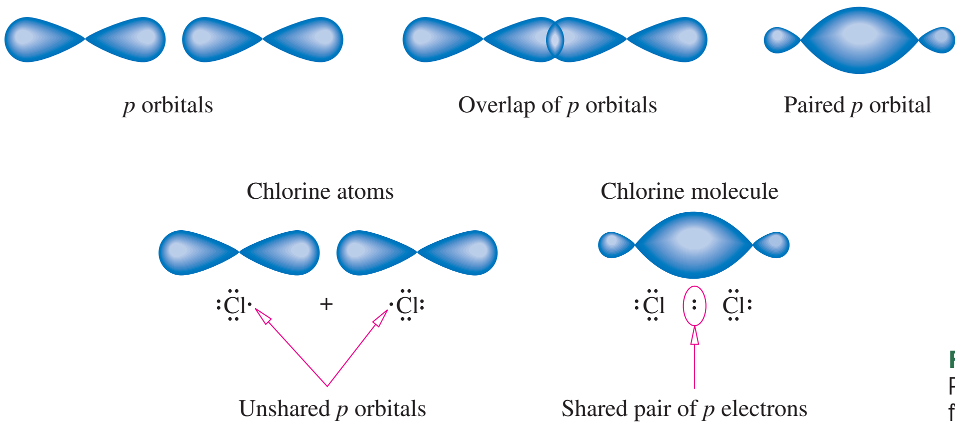
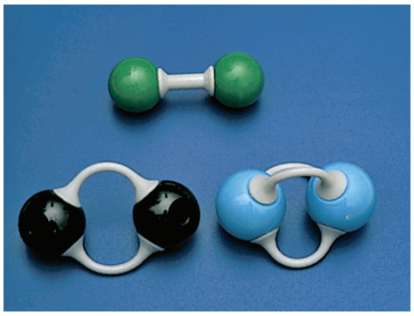


Figure 11.9
Pairing of *p* electrons in the formation of a chlorine molecule.

is this: The unpaired 3*p* electron orbital of one chlorine atom overlaps the unpaired 3*p* electron orbital of the other atom, resulting in a pair of electrons that are mutually shared between the two atoms. Each atom furnishes one of the pair of shared electrons. Thus, each atom attains a stable structure of eight electrons by sharing an electron pair with the other atom. The pairing of the *p* electrons and the formation of a chlorine molecule are illustrated in Figure 11.9. Neither chlorine atom has a positive or negative charge, because both contain the same number of protons and have equal attraction for the pair of electrons being shared. Other examples of molecules in which electrons are equally shared between two atoms are hydrogen (H_2), oxygen (O_2), nitrogen (N_2), fluorine (F_2), bromine (Br_2), and iodine (I_2). Note that more than one pair of electrons may be shared between atoms:

H:H	:F:F:	:Br:Br:	:I:I:	:O::O:	:N:::N:
hydrogen	fluorine	bromine	iodine	oxygen	nitrogen



Molecular models for F_2 (green, single bond), O_2 (black, double bond), and N_2 (blue, triple bond).

The Lewis structure given for oxygen does not adequately account for all the properties of the oxygen molecule. Other theories explaining the bonding in oxygen molecules have been advanced, but they are complex and beyond the scope of this book.

In writing structures, we commonly replace the pair of dots used to represent a shared pair of electrons with a dash (—). One dash represents a single bond; two dashes, a double bond; and three dashes, a triple bond. The six structures just shown may be written thus:

H—H :F—F: :Br—Br: :I—I: :O=O: :N≡N:

Remember: A dash represents a shared pair of electrons.

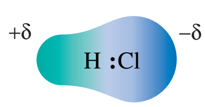
The ionic bond and the covalent bond represent two extremes. In ionic bonding the atoms are so different that electrons are transferred between them, forming a charged pair of ions. In covalent bonding, two identical atoms share electrons equally. The bond is the mutual attraction of the two nuclei for the shared electrons. Between these extremes lie many cases in which the atoms are not different enough for a transfer of electrons but are different enough that the electron pair cannot be shared equally. This unequal sharing of electrons results in the formation of a **polar covalent bond**.

polar covalent bond

11.6 Electronegativity

When two *different* kinds of atoms share a pair of electrons, a bond forms in which electrons are shared unequally. One atom assumes a partial positive charge and the other a partial negative charge with respect to each other. This difference in charge occurs because the two atoms exert unequal attraction for the pair of shared electrons. The attractive force that an atom of an element has for shared electrons in a molecule or polyatomic ion is known as its **electronegativity**. Elements differ in their electronegativities. For example, both hydrogen and chlorine need one electron to form stable electron configurations. They share a pair of electrons in hydrogen chloride (HCl). Chlorine is more electronegative and therefore has a greater attraction for the shared electrons than does hydrogen. As a result, the pair of electrons is displaced toward the chlorine atom, giving it a partial negative charge and leaving the hydrogen atom with a partial positive charge. Note that the electron is not transferred entirely to the chlorine atom (as in the case of sodium chloride) and that no ions are formed. The entire molecule, HCl, is electrically neutral. A partial charge is usually indicated by the Greek letter delta, δ. Thus, a partial positive charge is represented by $\delta+$ and a partial negative charge by $\delta-$.

electronegativity



$\overset{\delta+}{\mathrm{H}} : \overset{\delta-}{\ddot{\mathrm{Cl}}}:$

hydrogen chloride

The pair of shared electrons in HCl is closer to the more electronegative chlorine atom than to the hydrogen atom, giving chlorine a partial negative charge with respect to the hydrogen atom.

A scale of relative electronegativities, in which the most electronegative element, fluorine, is assigned a value of 4.0, was developed by the Nobel laureate (1954 and 1962) Linus Pauling (1901–1994). Table 11.5 shows that the relative electronegativity of the nonmetals is high and that of the metals is low. These electronegativities indicate that atoms of metals have a greater tendency to lose electrons than do atoms of nonmetals and that nonmetals have a greater tendency to gain electrons than do metals. The higher the electronegativity value, the greater the attraction for electrons. Note that electronegativity generally increases from left to right across a period and decreases down a group for the representative elements. The highest electronegativity is 4.0 for fluorine, and the lowest is 0.7 for francium and cesium. It's important to remember that the higher the electronegativity, the stronger an atom attracts electrons.

The polarity of a bond is determined by the difference in electronegativity values of the atoms forming the bond (see Figure 11.10). If the electronegativities are the same, the bond is **nonpolar covalent** and the electrons are shared equally. If the atoms have greatly different electronegativities, the bond is very *polar*. At the extreme, one or more electrons are actually transferred and an ionic bond results.

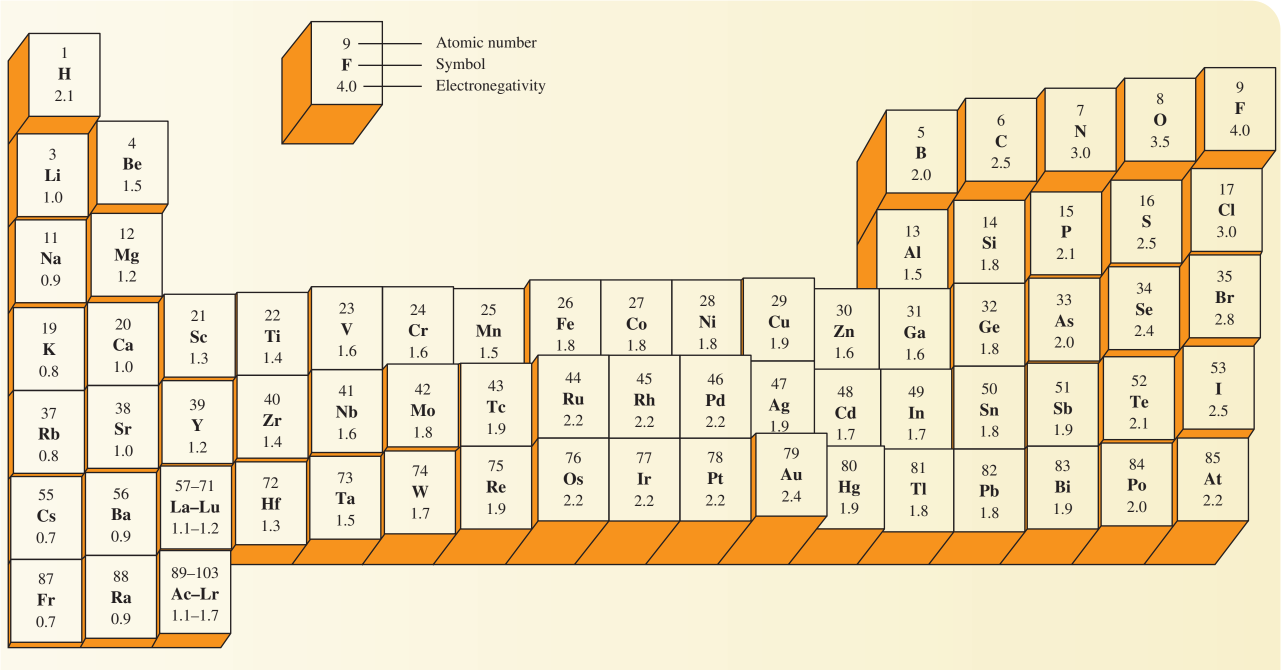
nonpolar covalent bond

dipole

A **dipole** is a molecule that is electrically asymmetrical, causing it to be oppositely charged at two points. A dipole is often written as $(+-)$. A hydrogen chloride molecule is polar and behaves as a small dipole. The HCl dipole may be written as $\mathrm{H} \longmapsto \mathrm{Cl}$. The arrow points toward the negative end of the dipole. Molecules of H_2O, HBr, and ICl are polar:

$\mathrm{H} \longmapsto \mathrm{Cl} \qquad \mathrm{H} \longmapsto \mathrm{Br} \qquad \mathrm{I} \longmapsto \mathrm{Cl} \qquad \mathrm{H} \nearrow \mathrm{O} \nwarrow \mathrm{H}$

Table 11.5 Three-Dimensional Representation of Electronegativity



How do we know whether a bond between two atoms is ionic or covalent? The difference in electronegativity between the two atoms determines the character of the bond formed between them. As the difference in electronegativity increases, the polarity of the bond (or percent ionic character) increases.

If the electronegativity difference between two bonded atoms is greater than 1.7–1.9, the bond will be more ionic than covalent.

If the electronegativity difference is greater than 2.0, the bond is strongly ionic. If the electronegativity difference is less than 1.5, the bond is strongly covalent.

Practice 11.5

Which of these compounds would you predict to be ionic and which would be covalent?

(a) $SrCl_2$
(b) PCl_3
(c) NH_3
(d) RbBr
(e) LiCl
(f) CS_2

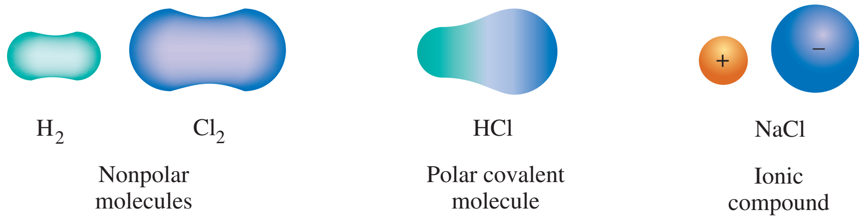


Figure 11.10
Nonpolar, polar covalent, and ionic compounds.

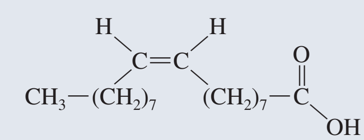
CHEMISTRY IN ACTION • *Trans*-forming Fats

Trans fats are virtually everywhere in the American diet. These fats remain when vegetable oils are converted into solid substances that are used in many processed foods. In fact, beginning January 1, 2006, manufacturers must label products to show their *trans* fat content. So just what is a *trans* fat?

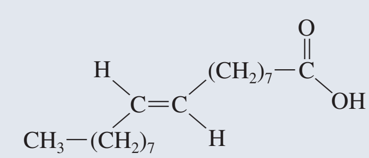
Different categories of fat can be identified from the pattern of bonds and hydrogen atoms in the molecule. Fatty acids are one component of fats and contain long chains of carbon atoms with hydrogen atoms bonded to some or all of the carbon atoms. Unsaturated fats (such as corn or soybean oil) contain double bonds between some of the carbon atoms in the chains. A carbon that is doubly bonded to another carbon usually also bonds to a hydrogen atom. If the fatty acid has only one double bond, it is monounsaturated (such as olive oil). If the fatty acid contains more than one double bond, it is polyunsaturated. All fat with no double bonds is saturated—all the carbon atoms have the maximum possible hydrogen atoms bonded to them.

Trans fats contain a particular kind of unsaturated fatty acid. When there is a double bond between carbon atoms, the molecule bends in one of two ways: the *cis* or the *trans* direction. In *cis* configuration, the carbon chain on both sides of the double bond bends in to the same side of the double bond (see structure). In the *trans* configuration the chain on either side of the double bond bends toward opposite sides of the double bond (see structure). Most *trans* fats come from processing oils for prepared foods and from solid fats such as margarine.

When an oil is converted into a solid fat, some of the double bonds are converted to single bonds by adding hydrogen (hydrogenation). This process is easier at *cis* double bonds, and therefore the remaining double bonds are mainly in the *trans* configuration. These *trans* fatty acids tend to stack together, making a solid easier than the *cis* forms. Studies have linked diets high in *trans* fats to poor health, high cholesterol, heart disease, and diabetes. Food producers are working on ways to lower the *trans* fat content of foods. Gary List at USDA in Peoria, Illinois, has used high-pressure hydrogen gas on soybean oil at 140°–170°C to hydrogenate the oil, producing a soft margarine containing 5%–6% *trans* fat instead of the ~40% from the standard hydrogenation techniques. This could lead to a product that qualifies for a label of 0 g *trans* fat. Look for lots of new products and labels in your grocery stores as manufacturers *trans* -form products into ones with less *trans* fat.



cis fatty acid

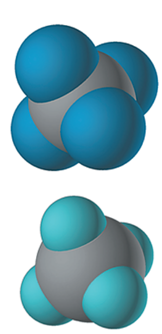


trans fatty acid

Care must be taken to distinguish between polar bonds and polar molecules. A covalent bond between different kinds of atoms is always polar. But a molecule containing different kinds of atoms may or may not be polar, depending on its shape or geometry. Molecules of HF, HCl, HBr, HI, and ICl are all polar because each contains a single polar bond. However, CO_2, CH_4, and CCl_4 are nonpolar molecules despite the fact that all three contain polar bonds. The carbon dioxide molecule O=C=O is nonpolar because the carbon–oxygen dipoles cancel each other by acting in opposite directions.

$$\overset{\longleftarrow\!\!+}{\mathrm{O}=}\overset{+\!\!\longrightarrow}{\mathrm{C}=\mathrm{O}}$$

dipoles in opposite directions



Spacefilling molecular model CCl_4 (top) and methane (CH_4).

Carbon tetrachloride (CCl_4) is nonpolar because the four C—Cl polar bonds are identical, and since these bonds emanate from the center to the corners of a tetrahedron in the molecule, their polarities cancel one another. Methane has the same molecular structure and is also nonpolar. We will discuss the shapes of molecules later in this chapter.

We have said that water is a polar molecule. If the atoms in water were linear like those in carbon dioxide, the two O—H dipoles would cancel each other, and the molecule would be nonpolar. However, water is definitely polar and has a nonlinear (bent) structure with an angle of 105° between the two O—H bonds.

CHEMISTRY IN ACTION • Goal! A Spherical Molecule

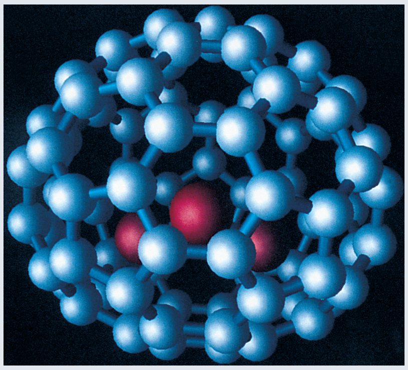
One of the most diverse elements in the periodic table is carbon. Graphite and diamond, two well-known forms of elemental carbon, both contain extended arrays of carbon atoms. In graphite the carbon atoms are arranged in sheets, and the bonding between the sheets is very weak. This property makes graphite useful as a lubricant and as a writing material. Diamond consists of transparent octahedral crystals in which each carbon atom is bonded to four other carbon atoms. This three-dimensional network of bonds gives diamond the property of hardness for which it is noted. In the 1980s, a new form of carbon was discovered in which the atoms are arranged in relatively small clusters.

When Harold Kroto of the University of Sussex, England, and Richard Smalley of Rice University, Texas, discovered a strange carbon molecule of formula C_{60}, they deduced that the most stable arrangement for the atoms would be in the shape of a soccer ball. In thinking about possible arrangements, the scientists considered the geodesic domes designed by R. Buckminster Fuller in the 1960s. This cluster form of carbon was thus named *buckminsterfullerene* and is commonly called buckyballs.

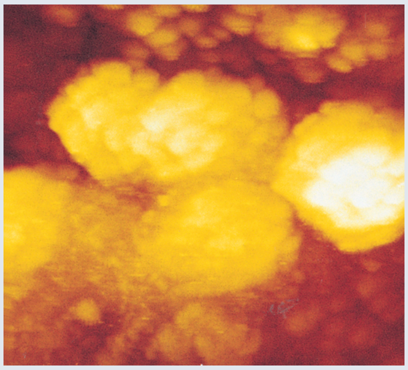
Buckyballs have captured the imagination of a variety of chemists. Research on buckyballs has led to a host of possible applications for these molecules. If metals are bound to the carbon atoms, the fullerenes become superconducting; that is, they conduct electrical current without resistance, at very low temperature. Scientists are now able to make buckyball compounds that superconduct at temperatures of 45 K. Other fullerenes are being used in lubricants and optical materials.

Chemists at Yale University have managed to trap helium and neon inside buckyballs. This is the first time chemists have ever observed helium or neon in a compound of any kind. They found that at temperatures from 1000°F to 1500°F one of the covalent bonds linking neighboring carbon atoms in the buckyball breaks. This opens a window in the fullerene molecule through which a helium or neon atom can enter the buckyball. When the fullerene is allowed to cool, the broken bond between carbon atoms re-forms, shutting the window and trapping the helium or neon atom inside the buckyball. Since the trapped helium or neon cannot react or share electrons with its host, the resulting compound has forced scientists to invent a new kind of chemical formula to describe the compound. The relationship between the "prisoner" helium or neon and the host buckyball is shown with an @ sign. A helium fullerene containing 60 carbon atoms would thus be $He@C_{60}$.

Buckyballs can be tailored to fit a particular size requirement. Raymond Schinazi of the Emory University School of Medicine, Georgia, made a buckyball to fit the active site of a key HIV enzyme that paralyzes the virus, making it noninfectious in human cells. The key to making this compound was preparing a water-soluble buckyball that would fit in the active site of the enzyme. Eventually, scientists created a water-soluble fullerene molecule that has two charged arms to grasp the binding site of the enzyme. It is toxic to the virus but doesn't appear to harm the host cells.



Scandium atoms trapped in a buckyball.



"Raspberries" of fullerene lubricant.

The relationships among types of bonds are summarized in Figure 11.11. It is important to realize that bonding is a continuum; that is, the difference between ionic and covalent is a gradual change.

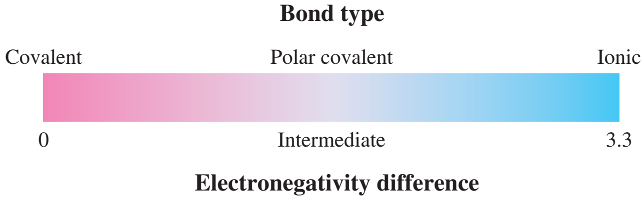


Figure 11.11
Relating bond type to electronegativity difference between atoms.

11.7 Lewis Structures of Compounds

As we have seen, Lewis structures are a convenient way of showing the covalent bonds in many molecules or ions of the representative elements. In writing Lewis structures, the most important consideration for forming a stable compound is that the atoms attain a noble gas configuration.

The most difficult part of writing Lewis structures is determining the arrangement of the atoms in a molecule or an ion. In simple molecules with more than two atoms, one atom will be the central atom surrounded by the other atoms. Thus, Cl_2O has two possible arrangements, Cl—Cl—O or Cl—O—Cl. Usually, but not always, the single atom in the formula (except H) will be the central atom.

Although Lewis structures for many molecules and ions can be written by inspection, the following procedure is helpful for learning to write them:

Remember: The number of valence electrons of Group A elements is the same as their group number in the periodic table.

Step 1 Obtain the total number of valence electrons to be used in the structure by adding the number of valence electrons in all the atoms in the molecule or ion. If you are writing the structure of an ion, add one electron for each negative charge or subtract one electron for each positive charge on the ion.

Step 2 Write the skeletal arrangement of the atoms and connect them with a single covalent bond (two dots or one dash). Hydrogen, which contains only one bonding electron, can form only one covalent bond. Oxygen atoms are not normally bonded to each other, except in compounds known to be peroxides. Oxygen atoms normally have a maximum of two covalent bonds (two single bonds or one double bond).

Step 3 Subtract two electrons for each single bond you used in Step 2 from the total number of electrons calculated in Step 1. This gives you the net number of electrons available for completing the structure.

Step 4 Distribute pairs of electrons (pairs of dots) around each atom (except hydrogen) to give each atom a noble gas structure.

Step 5 If there are not enough electrons to give these atoms eight electrons, change single bonds between atoms to double or triple bonds by shifting unbonded pairs of electrons as needed. Check to see that each atom has a noble gas electron structure (two electrons for hydrogen and eight for the others). A double bond counts as four electrons for each atom to which it is bonded.

Example 11.9 How many valence electrons are in each of these atoms: Cl, H, C, O, N, S, P, I?

SOLUTION You can look at the periodic table to determine the electron structure, or, if the element is in Group A of the periodic table, the number of valence electrons is equal to the group number:

Atom	Group	Valence electrons
Cl	7A	7
H	1A	1
C	4A	4
O	6A	6
N	5A	5
S	6A	6
P	5A	5
I	7A	7

Example 11.10

Write the Lewis structure for water (H_2O).

SOLUTION

Step 1 The total number of valence electrons is eight, two from the two hydrogen atoms and six from the oxygen atom.

Step 2 The two hydrogen atoms are connected to the oxygen atom. Write the skeletal structure:

```
H O      or      H O H
  H
```

Place two dots between the hydrogen and oxygen atoms to form the covalent bonds:

```
H:O      or      H:O:H
  ..
  H
```

Step 3 Subtract the four electrons used in Step 2 from eight to obtain four electrons yet to be used.

Step 4 Distribute the four electrons in pairs around the oxygen atom. Hydrogen atoms cannot accommodate any more electrons:

```
   ..                 ..
H—O:      or      H—O—H
   |                  ..
   H
```

These arrangements are Lewis structures because each atom has a noble gas electron structure. Note that the shape of the molecule is not shown by the Lewis structure.

Example 11.11

Write Lewis structures for a molecule of methane (CH_4).

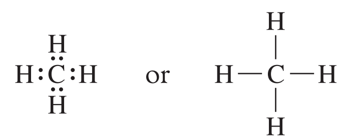
SOLUTION

Step 1 The total number of valence electrons is eight, one from each hydrogen atom and four from the carbon atom.

Step 2 The skeletal structure contains four H atoms around a central C atom. Place two electrons between the C and each H.

```
  H          H
H C H        ..
  H        H:C:H
             ..
             H
```

Step 3 Subtract the eight electrons used in Step 2 from eight (obtained in Step 1) to obtain zero electrons yet to be placed. Therefore the Lewis structure must be as written in Step 2:



Example 11.12 Write the Lewis structure for a molecule of carbon tetrachloride (CCl_4).

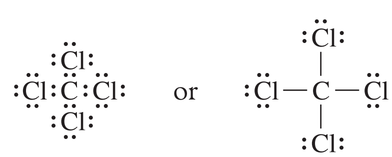
SOLUTION

Step 1 The total number of valence electrons to be used is 32, 4 from the carbon atom and 7 from each of the four chlorine atoms.

Step 2 The skeletal structure contains the four Cl atoms around a central C atom. Place 2 electrons between the C and each Cl:

Step 3 Subtract the 8 electrons used in Step 2 from 32 (obtained in Step 1) to obtain 24 electrons yet to be placed.

Step 4 Distribute the 24 electrons (12 pairs) around the Cl atoms so that each Cl atom has 8 electrons around it:



This arrangement is the Lewis structure; CCl_4 contains four covalent bonds.

Example 11.13 Write a Lewis structure for CO_2.

SOLUTION

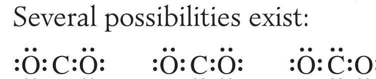
Step 1 The total number of valence electrons is 16, 4 from the C atom and 6 from each O atom.

Step 2 The two O atoms are bonded to a central C atom. Write the skeletal structure and place 2 electrons between the C and each O atom.

O:C:O

Step 3 Subtract the 4 electrons used in Step 2 from 16 (found in Step 1) to obtain 12 electrons yet to be placed.

Step 4 Distribute the 12 electrons (six pairs) around the C and O atoms. Several possibilities exist:



I II III

Step 5 Not all the atoms have 8 electrons around them (noble gas structure). Remove one pair of unbonded electrons from each O atom in structure I and place one pair between each O and the C atom, forming two double bonds:

$:\ddot{O}::C::\ddot{O}:$ or $:\ddot{O}=C=\ddot{O}:$

Each atom now has 8 electrons around it. The carbon is sharing four pairs of electrons, and each oxygen is sharing two pairs. These bonds are known as double bonds because each involves sharing two pairs of electrons.

Practice 11.6

Write the Lewis structures for the following:

(a) PBr_3 (b) $CHCl_3$ (c) HF (d) H_2CO (e) N_2

Although many compounds attain a noble gas structure in covalent bonding, there are numerous exceptions. Sometimes it's impossible to write a structure in which each atom has 8 electrons around it. For example, in BF_3 the boron atom has only 6 electrons around it, and in SF_6 the sulfur atom has 12 electrons around it.

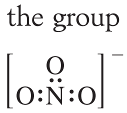
Although there are exceptions, many molecules can be described using Lewis structures where each atom has a noble gas electron configuration. This is a useful model for understanding chemistry.

11.8 Complex Lewis Structures

Most Lewis structures give bonding pictures that are consistent with experimental information on bond strength and length. There are some molecules and polyatomic ions for which no single Lewis structure consistent with all characteristics and bonding information can be written. For example, consider the nitrate ion, NO_3^-. To write a Lewis structure for this polyatomic ion, we use the following steps.

Step 1 The total number of valence electrons is 24, 5 from the nitrogen atom, 6 from each oxygen atom, and 1 from the −1 charge.

Step 2 The three O atoms are bonded to a central N atom. Write the skeletal structure and place two electrons between each pair of atoms. Since we have an extra electron in this ion, resulting in a −1 charge, we enclose the group of atoms in square brackets and add a − charge as shown.



Step 3 Subtract the 6 electrons used in Step 2 from 24 (found in Step 1) to obtain 18 electrons yet to be placed.

CHEMISTRY IN ACTION • Cleaner Showers through Chemistry

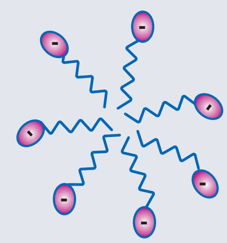
Keeping the shower area sparkling clean and free of mildew is a job none of us enjoy. Now thanks to chemist Bob Black of Jacksonville, Florida, there is a product that cleans the shower without any scrubbing! Black struggled with his home shower until he finally decided he really needed a new product that would solve the mildew and scrubbing problem.

His search was based on the following needs:

1. A molecule to lift deposits off the walls of the shower
2. A way to prevent hard-water deposits from forming
3. A wetting agent to wet the walls and rinse off deposits

In all cases, he limited his search to substances nontoxic and environmentally safe.

Black used a molecule called a glycol ether to lift deposits off the shower wall. This molecule is a long chain with a polar end and a nonpolar end. Substances that are nonpolar (such as grease, oils, and organic material) are attracted to the nonpolar end of the molecules. The molecules cluster together to form micelles (with the polar end pointed out). The polar sphere dissolves in the polar water from the shower, washing off the organic deposits.



A micelle

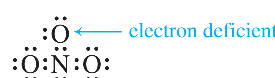
Preventing the hard-water deposits from forming on the shower walls required use of a molecule called EDTA. This molecule bonds to ions (like Ca^{2+}, Mg^{2+}, or Fe^{3+}) and prevents the formation of soap scum or hard-water deposits on the walls.

Lastly, Black added isopropyl alcohol to wet the shower wall and also to disturb the mildew fungus. He mixed all the ingredients in just the right proportions and found he no longer needed to work so hard to clean the shower.

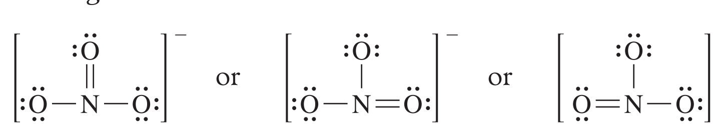
Fortunately for all of us, Black shared his solution with friends, who also liked it. Black patented his new product and began mass production. You can now find Clean Shower in your local grocery store!

Simply spraying your shower regularly with Clean Shower will keep it free of deposits and mildew.

Step 4 Distribute the 18 electrons around the N and O atoms:



Step 5 One pair of electrons is still needed to give all the N and O atoms a noble gas structure. Move the unbonded pair of electrons from the N atom and place it between the N and the electron-deficient O atom, making a double bond.



Are these all valid Lewis structures? Yes, so there really are three possible Lewis structures for NO_3^-.

A molecule or ion that has multiple correct Lewis structures shows *resonance*. Each of these Lewis structures is called a **resonance structure**. In this book, however, we will not be concerned with how to choose the correct resonance structure for a molecule or ion. Therefore any of the possible resonance structures may be used to represent the ion or molecule.

resonance structure

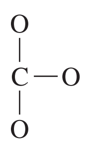
Example 11.14

Write the Lewis structure for a carbonate ion (CO_3^{2-}).

SOLUTION

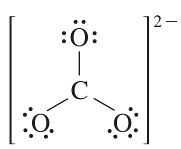
Step 1 These four atoms have 22 valence electrons plus 2 electrons from the −2 charge, which makes 24 electrons to be placed.

Step 2 In the carbonate ion, the carbon is the central atom surrounded by the three oxygen atoms. Write the skeletal structure and place 2 electrons (or a single line) between each C and O:



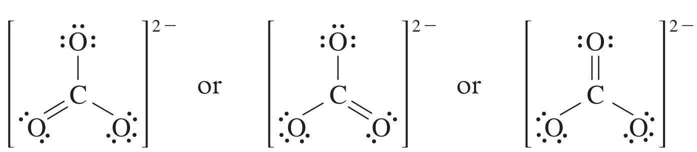
Step 3 Subtract the 6 electrons used in Step 2 from 24 (from Step 1) to give 18 electrons yet to be placed.

Step 4 Distribute the 18 electrons around the three oxygen atoms and indicate that the carbonate ion has a −2 charge:



The difficulty with this structure is that the carbon atom has only six electrons around it instead of a noble gas octet.

Step 5 Move one of the nonbonding pairs of electrons from one of the oxygens and place them between the carbon and the oxygen. Three Lewis structures are possible:



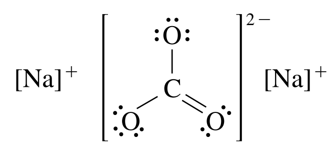
Practice 11.7

Write the Lewis structure for each of the following:

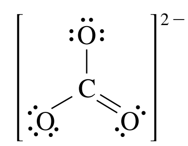
(a) NH_3 (b) H_3O^+ (c) NH_4^+ (d) HCO_3^-

11.9 Compounds Containing Polyatomic Ions

A polyatomic ion is a stable group of atoms that has either a positive or a negative charge and behaves as a single unit in many chemical reactions. Sodium carbonate, Na_2CO_3, contains two sodium ions and a carbonate ion. The carbonate ion (CO_3^{2-}) is a polyatomic ion composed of one carbon atom and three oxygen atoms and has a charge of −2. One carbon and three oxygen atoms have a total of 22 electrons in their outer energy levels. The carbonate ion contains 24 outer electrons and therefore has a charge of −2. In this case, the 2 additional electrons come from the two sodium atoms, which are now sodium ions:



sodium carbonate



carbonate ion

Sodium carbonate has both ionic and covalent bonds. Ionic bonds exist between each of the sodium ions and the carbonate ion. Covalent bonds are present between the carbon and oxygen atoms within the carbonate ion. One important difference between the ionic and covalent bonds in this compound can be demonstrated by dissolving sodium carbonate in water. It dissolves in water, forming three charged particles—two sodium ions and one carbonate ion—per formula unit of Na_2CO_3:

$$\underset{\text{sodium carbonate}}{Na_2CO_3(s)} \xrightarrow{\text{water}} \underset{\text{sodium ions}}{2\,Na^+(aq)} + \underset{\text{carbonate ion}}{CO_3^{2-}(aq)}$$

The CO_3^{2-} ion remains as a unit, held together by covalent bonds; but where the bonds are ionic, dissociation of the ions takes place. Do not think, however, that polyatomic ions are so stable that they cannot be altered. Chemical reactions by which polyatomic ions can be changed to other substances do exist.

11.10 Molecular Shape

So far in our discussion of bonding we have used Lewis structures to represent valence electrons in molecules and ions, but they don't indicate anything regarding the molecular or geometric shape of a molecule. The three-dimensional arrangement of the atoms within a molecule is a significant feature in understanding molecular interactions. Let's consider several examples illustrated in Figure 11.12.

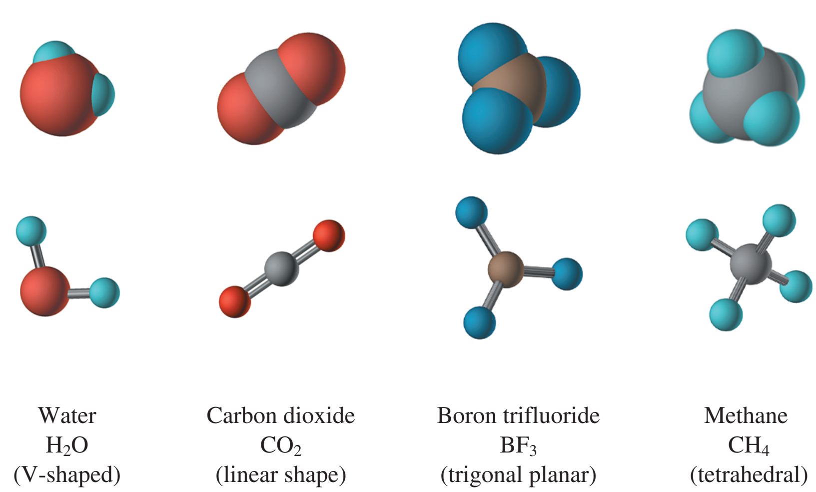
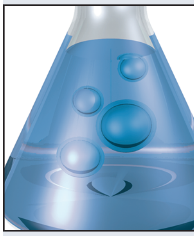


Figure 11.12
Geometric shapes of common molecules. Each molecule is shown as a ball-and-stick model (showing the bonds) and as a spacefilling model (showing the shape).



CHEMISTRY IN ACTION • Strong Enough to Stop a Bullet?

What do color-changing pens, bullet-resistant vests, and calculators have in common? The chemicals that make each of them work are liquid crystals. These chemicals find numerous applications; you are probably most familiar with liquid crystal displays (LCDs) and color-changing products, but these chemicals are also used to make superstrong synthetic fibers.

Molecules in a normal crystal remain in an orderly arrangement, but in a liquid crystal the molecules can flow *and* maintain an orderly arrangement at the same time. Liquid crystal molecules are linear and polar. Since the atoms tend to lie in a relatively straight line, the molecules are generally much longer than they are wide. These polar molecules are attracted to each other and are able to line up in an orderly fashion, without solidifying.

Liquid crystals with twisted arrangements of molecules give us novelty color-changing products. In these liquid crystals the molecules lie side by side in a nearly flat layer. The next layer is similar, but at an angle to the one below. The closely packed flat layers have a special effect on light. As the light strikes the surface, some of it is reflected from the top layer and some from lower layers. When the same wavelength is reflected from many layers, we see a color. (This is similar to the rainbow of colors formed by oil in a puddle on the street or the film of a soap bubble.) As the temperature is increased, the molecules move faster, causing a change in the angle and the space between the layers. This results in a color change in the reflected light. Different compounds change color within different temperature ranges, allowing a variety of practical and amusing applications.

Liquid crystal (nematic) molecules that lie parallel to one another are used to manufacture very strong synthetic fibers. Perhaps the best example of these liquid crystals is Kevlar, a synthetic fiber used in bullet-resistant vests, canoes, and parts of the space shuttle. Kevlar is a synthetic polymer, like nylon or polyester, that gains strength by passing through a liquid crystal state during its manufacture.

In a typical polymer, the long molecular chains are jumbled together, somewhat like spaghetti. The strength of the material is limited by the disorderly arrangement. The trick is to get the molecules to line up parallel to each other. Once the giant molecules have been synthesized, they are dissolved in sulfuric acid. At the proper concentration the molecules align, and the solution is forced through tiny holes in a nozzle and further aligned. The sulfuric acid is removed in a water bath, thereby forming solid fibers in near-perfect alignment. One strand of Kevlar is stronger than an equal-sized strand of steel. It has a much lower density as well, making it a material of choice in bullet-resistant vests.

Kevlar is used to make protective vests for police.

Water is known to have the geometric shape known as "bent" or "V-shaped." Carbon dioxide exhibits a linear shape. BF_3 forms a third molecular shape called *trigonal planar* since all the atoms lie in one plane in a triangular arrangement. One of the more common molecular shapes is the tetrahedron, illustrated by the molecule methane (CH_4).

How do we predict the geometric shape of a molecule? We will now study a model developed to assist in making predictions from the Lewis structure.

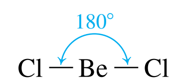
11.11 The Valence Shell Electron Pair Repulsion (VSEPR) Model

The chemical properties of a substance are closely related to the structure of its molecules. A change in a single site on a large biomolecule can make a difference in whether or not a particular reaction occurs.

Instrumental analysis can be used to determine exact spatial arrangements of atoms. Quite often, though, we only need to be able to predict the approximate structure of a molecule. A relatively simple model has been developed to allow us to make predictions of shape from Lewis structures.

Nonbonding pairs of electrons are not shown here, so you can focus your attention on the shapes, not electron arrangement.

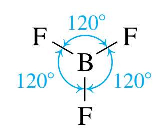
The VSEPR model is based on the idea that electron pairs will repel each other electrically and will seek to minimize this repulsion. To accomplish this minimization, the electron pairs will be arranged around a central atom as far apart as possible. Consider $BeCl_2$, a molecule with only two pairs of electrons surrounding the central atom. These electrons are arranged 180° apart for maximum separation:



linear structure

This molecular structure can now be labeled as a **linear structure**. When only two pairs of electrons surround a central atom, they should be placed 180° apart to give a linear structure.

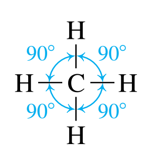
What occurs when there are only three pairs of electrons around the central atom? Consider the BF_3 molecule. The greatest separation of electron pairs occurs when the angles between atoms are 120°:



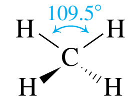
trigonal planar structure

This arrangement of atoms is flat (planar) and, as noted earlier, is called **trigonal planar structure**. When three pairs of electrons surround an atom, they should be placed 120° apart to show the trigonal planar structure.

Now consider the most common situation (CH_4), with four pairs of electrons on the central carbon atom. In this case the central atom exhibits a noble gas electron structure. What arrangement best minimizes the electron pair repulsions? At first it seems that an obvious choice is a 90° angle with all the atoms in a single plane:



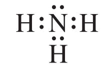
However, we must consider that molecules are three-dimensional. This concept results in a structure in which the electron pairs are actually 109.5° apart:

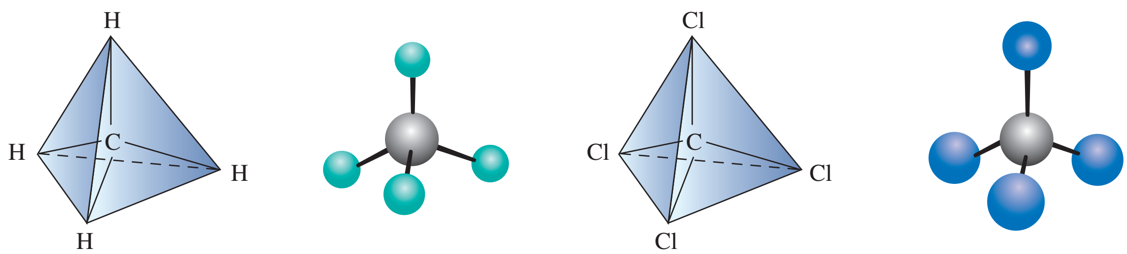


tetrahedral structure

In this diagram the wedged line seems to protrude from the page whereas the dashed line recedes. Two representations of this arrangement, known as **tetrahedral structure**, are illustrated in Figure 11.13. When four pairs of electrons surround a central atom, they should be placed 109.5° apart to give them a tetrahedral structure.

The VSEPR model is based on the premise that we are counting electron pairs. It's quite possible that one or more of these electron pairs may be nonbonding (lone) pairs. What happens to the molecular structure in these cases? Consider the ammonia molecule. First we draw the Lewis structure to determine the number of electron pairs around the central atom:



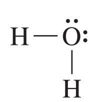


Methane, CH_4 CH_4 Carbon tetrachloride, CCl_4 CCl_4

Figure 11.13
Ball-and-stick models of methane and carbon tetrachloride. Methane and carbon tetrachloride are nonpolar molecules because their polar bonds cancel each other in the tetrahedral arrangement of their atoms. The carbon atoms are located in the centers of the tetrahedrons.

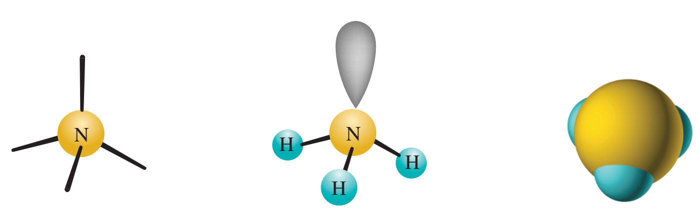
Since there are four pairs of electrons, the arrangement of electrons around the central atom will be tetrahedral (Figure 11.14a). However, only three of the pairs are bonded to another atom, so the molecule itself is pyramidal. It is important to understand that the placement of the electron pairs determines the shape but the name for the molecule is determined by the position of the atoms themselves. Therefore, ammonia is pyramidal. See Figure 11.14c.

Now consider the effect of two unbonded pairs of electrons in the water molecule. The Lewis structure for water is



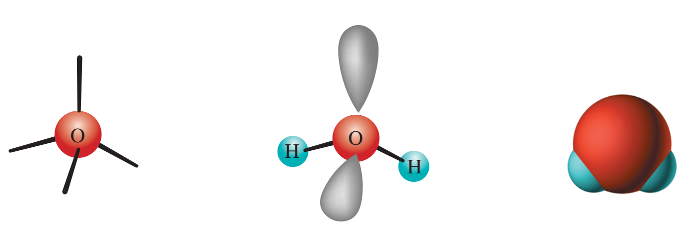
The four electron pairs indicate a tetrahedral electron arrangement is necessary (see Figure 11.15a). The molecule is not called tetrahedral because two of the electron pairs are unbonded pairs. The water molecule is "bent," as shown in Figure 11.15c.

The arrangement of electron pairs around an atom determines its shape, but we name the shape of molecules by the position of the atoms.



(a) (b) (c)

Figure 11.14
(a) The tetrahedral arrangement of electron pairs around the N atom in the NH_3 molecule. (b) Three pairs are shared and one is unshared. (c) The NH_3 molecule is pyramidal.



(a) (b) (c)

Figure 11.15
(a) The tetrahedral arrangement of the four electron pairs around oxygen in the H_2O molecule. (b) Two of the pairs are shared and two are unshared. (c) The H_2O molecule is bent.

Let's summarize the VSEPR model. To determine the molecular shape for a substance, follow these steps:

Step 1 Draw the Lewis structure for the molecule.

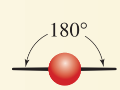
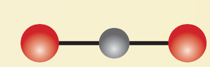
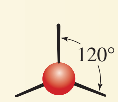
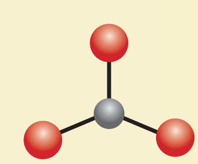
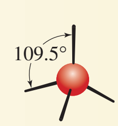
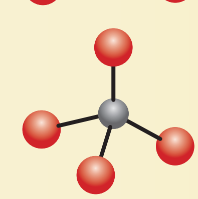
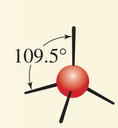
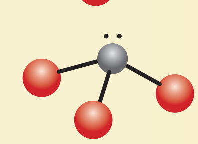
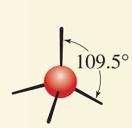
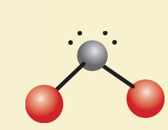
Step 2 Count the electron pairs and arrange them to minimize repulsions (as far apart as possible). This determines the electron pair arrangement.

Step 3 Determine the positions of the atoms.

Step 4 Name the molecular structure from the position of the atoms.

It is important to recognize that the placement of the electron pairs determines the structure but the name of the molecular structure is determined by the position of the atoms. Table 11.6 shows the results of this process. Note that when the number of electron pairs is the same as the number of atoms, the electron pair arrangement and the molecular structure are the same. But when the number of atoms and the number of electron pairs are not the same, the molecular structure is different from the electron pair arrangement. This is illustrated when the number of electron pairs is four (a tetrahedral arrangement) in Table 11.6.

Table 11.6 Arrangement of Electron Pairs and Molecular Structure

Number of electron pairs	Electron pair arrangement	Ball-and-stick model	Bonds	Molecular structure	Molecular structure model
2	Linear		2	Linear	
3	Trigonal planar		3	Trigonal planar	
4	Tetrahedral		4	Tetrahedral	
4	Tetrahedral		3	Trigonal pyramidal	
4	Tetrahedral		2	Bent	

Practice 11.8

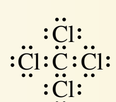
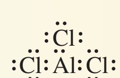
Predict the shape for CF_4, NF_3, and BeI_2.

Example 11.15

Predict the molecular shape for these molecules: H_2S, CCl_4, $AlCl_3$.

SOLUTION

1. Draw the Lewis structure.
2. Count the electron pairs around the central atom and determine the electron arrangement that will minimize repulsions.
3. Determine the positions of the atoms and name the shape of the molecule.

Molecule	Lewis structure	Number of electron pairs	Electron pair arrangement	Molecular shape
H_2S		4	tetrahedral	bent
CCl_4		4	tetrahedral	tetrahedral
$AlCl_3$		3	trigonal planar	trigonal planar

Chapter 11 Review

11.1 Periodic Trends in Atomic Properties

KEY TERM

Ionization energy

- Metals and nonmetals
- Atomic radius:
  - Increases down a group
  - Decreases across a row
- Ionization energy:
  - Energy required to remove an electron from an atom
  - Decreases down a group
  - Increases across a row

11.2 Lewis Structures of Atoms

KEY TERM

Lewis structure

- A Lewis structure is a representation of the atom where the symbol represents the element and dots around the symbol represent the valence electrons.
- To determine a Lewis structure for representative elements, use the Group number as the number of electrons to place around the symbol for the element.

11.3 The Ionic Bond: Transfer of Electrons from One Atoms to Another

KEY TERM

Ionic bond

- The goal of bonding is to achieve stability:
  - For representative elements, this stability can be achieved by attaining a valence electron structure of a noble gas.
- In an ionic bond stability is attained by transferring an electron from one atom to another:
  - The atom that loses an electron becomes a cation:
    - Positive ions are smaller than their parent atoms.
    - Metals tend to form cations.

  - The atom gaining an electron becomes an anion:
    - Negative ions are larger than their parent atoms.
    - Nonmetals tend to form anions.
  - Ionic compounds do not exist as molecules:
    - Ions are attracted by multiple ions of the opposite charge to form a crystalline structure.

11.4 Predicting Formulas of Ionic Compounds

- Chemical compounds are always electrically neutral.
- Metals lose electrons and nonmetals gain electrons to form compounds.
- Stability is achieved (for representative elements) by attaining a noble gas electron configuration.

11.5 The Covalent Bond: Sharing Electrons

KEY TERMS

Covalent bond
Polar covalent bond

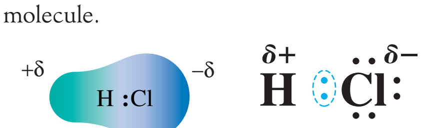
- Covalent bonds are formed when two atoms share a pair of electrons between them:
  - This is the predominant type of bonding in compounds.
  - True molecules exist in covalent compounds.
  - Overlap of orbitals forms a covalent bond.
- Unequal sharing of electrons results in a polar covalent bond.

11.6 Electronegativity

KEY TERMS

Electronegativity
Nonpolar covalent bond
Dipole

- Electronegativity is the attractive force an atom has for a shared electrons in a molecule or polyatomic ion.
- Electrons spend more time closer to the more electronegative atom in a bond forming a polar bond.
- The polarity of a bond is determined by the electronegativity difference between the atoms involved in the bond:
  - The greater the difference, the more polar the bond is.
  - At the extremes:
    - Large differences result in ionic bonds.
    - Tiny differences (or no difference) result(s) in a nonpolar covalent bond.
  - A molecule that is electrically asymmetrical has a dipole, resulting in charged areas within the molecule.



hydrogen chloride

  - If the electronegativity difference between two bonded atoms is greater than 1.7–1.9, the bond will be more ionic than covalent.
  - Polar bonds do not always result in polar molecules.

11.7 Lewis Structures of Compounds

- To write a Lewis structure for a compound:
  - Add all the valence electrons in all the atoms in the compound. For ions, adjust the number accordingly.
  - Write the skeletal arrangement for the atoms and connect them with single bonds (2 electrons per bond).
  - Subtract the electrons used in the bonds from the total valence electrons.
  - Distribute pairs of electrons around each atom to give each atom a noble gas structure.
  - If there are not enough electrons, convert single bonds between atoms to multiple bonds to attain noble gas structures.

11.8 Complex Lewis Structures

KEY TERM

Resonance structure

- When a single unique Lewis structure cannot be drawn for a molecule, resonance structures (multiple Lewis structures) are used to represent the molecule

11.9 Compounds Containing Polyatomic Ions

- Polyatomic ions behave like a single unit in many chemical reactions.
- The bonds within a polyatomic ion are covalent.

11.10 Molecular Shape

- Lewis structures do not indicate the shape of a molecule.

11.11 The Valence Shell Electron Pair Repulsion (VSEPR) Model

KEY TERMS

Linear structure
Trigonal planar structure
Tetrahedral structure
Bent structure

- VSEPR model:
  - Electron pairs around an atom tend to orient themselves in space as far apart as possible to minimize repulsive forces.
  - The arrangement of electron pairs around an atom determines its structure, but the molecular shape is determined by the position of the atoms in the molecule

Review Questions

All questions with blue numbers have answers in the appendix of the text.

1. Rank these elements according to the radii of their atoms, from smallest to largest: Na, Mg, Cl, K, and Rb. (Figure 11.2)
2. Explain why much more ionization energy is required to remove the first electron from neon than from sodium. (Table 11.1)
3. Explain the large increase in ionization energy needed to remove the third electron from beryllium compared with that needed for the second electron. (Table 11.1)
4. Does the first ionization energy increase or decrease from top to bottom in the periodic table for the alkali metal family? Explain. (Figure 11.3)
5. Does the first ionization energy increase or decrease from top to bottom in the periodic table for the noble gas family? Explain. (Figure 11.3)
6. Why does barium (Ba) have a lower ionization energy than beryllium (Be)? (Figure 11.3)
7. Why is there such a large increase in the ionization energy required to remove the second electron from a sodium atom as opposed to the first? (Table 11.1)
8. Which element in the pair has the larger atomic radius? (Figure 11.2)
   (a) Na or K
   (b) Na or Mg
   (c) O or F
   (d) Br or I
   (e) Ti or Zr
9. In Groups 1A–7A, which element in each group has the smallest atomic radius? (Figure 11.2)
10. Why does the atomic size increase in going down any family of the period table?
11. Why are only valence electrons represented in a Lewis structure?
12. Why do metals tend to lose electrons and nonmetals tend to gain electrons when forming ionic bonds?
13. State whether the elements in each group gain or lose electrons in order to achieve a noble gas configuration. Explain.
    (a) Group 1A
    (b) Group 2A
    (c) Group 6A
    (d) Group 7A
14. What is the difference between a polar bond and a polar molecule?
15. What is the difference between electron pair arrangement and molecular shape?
16. What is the purpose of a Lewis structure?
17. In a polar covalent bond, how do you determine which atom has a partial negative charge ($\delta-$) and which has a partial positive charge ($\delta+$)?
18. In a Lewis structure, what do the dots represent and what do the lines represent?
19. All the atoms within each Group A family of elements can be represented by the same Lewis structure. Complete the following table, expressing the Lewis structure for each group. (Use E to represent the elements.) (Figure 11.4)

Group	1A	2A	3A	4A	5A	6A	7A
	E·						

20. Draw the Lewis structure for Cs, Ba, Tl, Pb, Po, At, and Rn. How do these structures correlate with the group in which each element occurs?
21. In which general areas of the periodic table are the elements with (a) the highest and (b) the lowest electronegativities located?
22. What are valence electrons?
23. Explain why potassium usually forms a K^+ ion, but not a K^{2+} ion.
24. Why does an aluminum ion have a +3 charge?

Paired Exercises

All exercises with blue numbers have answers in the appendix of the text.

1. Which one in each pair has the larger radius? Explain.
   (a) a calcium atom or a calcium ion
   (b) a chlorine atom or a chloride ion
   (c) a magnesium ion or an aluminum ion
   (d) a sodium atom or a silicon atom
   (e) a potassium ion or a bromide ion
2. Which one in each pair has the larger radius? Explain.
   (a) Fe^{2+} or Fe^{3+}
   (b) a potassium atom or a potassium ion
   (c) a sodium ion or a chloride ion
   (d) a strontium atom or an iodine atom
   (e) a rubidium ion or a strontium ion

3. Using the table of electronegativity values (Table 11.5), indicate which element is more positive and which is more negative in these compounds:
(a) H_2O (b) NaF (c) NH_3 (d) PbS (e) NO (f) CH_4

4. Using the table of electronegativity values (Table 11.5), indicate which element is more positive and which is more negative in these compounds:
(a) HCl (b) LiH (c) CCl_4 (d) IBr (e) MgH_2 (f) OF_2

5. Classify the bond between these pairs of elements as principally ionic or principally covalent (use Table 11.5):
(a) sodium and chlorine
(b) carbon and hydrogen
(c) chlorine and carbon
(d) calcium and oxygen

6. Classify the bond between these pairs of elements as principally ionic or principally covalent (use Table 11.5):
(a) hydrogen and sulfur
(b) barium and oxygen
(c) fluorine and fluorine
(d) potassium and fluorine

7. Explain what happens to the electron structures of Mg and Cl atoms when they react to form $MgCl_2$.

8. Write an equation representing each of the following:
(a) the change of a fluorine atom to a fluoride ion
(b) the change of a calcium atom to a calcium ion

9. Use Lewis structures to show the electron transfer that enables these ionic compounds to be formed:
(a) MgF_2 (b) K_2O

10. Use Lewis structures to show the electron transfer that enables these ionic compounds to be formed:
(a) CaO (b) NaBr

11. How many valence electrons are in each of these atoms? H, K, Mg, He, Al

12. How many valence electrons are in each of these atoms? Si, N, P, O, Cl

13. How many electrons must be gained or lost for the following to achieve a noble gas electron structure?
(a) a calcium atom
(b) a sulfur atom
(c) a helium atom

14. How many electrons must be gained or lost for the following to achieve a noble gas electron structure?
(a) a chloride ion
(b) a nitrogen atom
(c) a potassium atom

15. Determine whether the following atoms will form an ionic compound or a molecular compound and give the formula of the compound.
(a) sodium and chlorine
(b) carbon and 4 hydrogen
(c) magnesium and bromine
(d) 2 bromine
(e) carbon and 2 oxygen

16. Determine whether each of the following atoms will form a nonpolar covalent compound or a polar covalent compound, and give the formula of the compound.
(a) 2 oxygen
(b) hydrogen and bromine
(c) oxygen and 2 hydrogen
(d) 2 iodine

17. Let E be any representative element. Following the pattern in the table, write formulas for the hydrogen and oxygen compounds of the following:
(a) Na (b) Ca (c) Al (d) Sn

Group 1A	2A	3A	4A	5A	6A	7A
EH	EH_2	EH_3	EH_4	EH_3	H_2E	HE
E_2O	EO	E_2O_3	EO_2	E_2O_5	EO_3	E_2O_7

18. Let E be any representative element. Following the pattern in the table, write formulas for the hydrogen and oxygen compounds of the following:
(a) Sb (b) Se (c) Cl (d) C

Group 1A	2A	3A	4A	5A	6A	7A
EH	EH_2	EH_3	EH_4	EH_3	H_2E	HE
E_2O	EO	E_2O_3	EO_2	E_2O_5	EO_3	E_2O_7

19. The formula for sodium sulfate is Na_2SO_4. Write the names and formulas for the other alkali metal sulfates.

20. The formula for calcium bromide is $CaBr_2$. Write the names and formulas for the other alkaline earth metal bromides.

21. Write Lewis structures for the following:
(a) Na (b) Br^- (c) O^{2-}

22. Write Lewis structures for the following:
(a) Ga (b) Ga^{3+} (c) Ca^{2+}

23. Classify the bonding in each compound as ionic or covalent:
(a) H_2O (c) MgO
(b) NaCl (d) Br_2

24. Classify the bonding in each compound as ionic or covalent:
(a) HCl (c) NH_3
(b) $BaCl_2$ (d) SO_2

25. Predict the type of bond that would be formed between the following pairs of atoms:
(a) Na and N
(b) N and S
(c) Br and I

26. Predict the type of bond that would be formed between the following pairs of atoms:
(a) H and Si
(b) O and F
(c) Ca and I

27. Draw Lewis structures for the following:
(a) H_2 (b) N_2 (c) Cl_2

28. Draw Lewis structures for the following:
(a) O_2 (b) Br_2 (c) I_2

29. Draw Lewis structures for the following:
(a) NCl_3 (c) C_2H_6
(b) H_2CO_3 (d) $NaNO_3$

30. Draw Lewis structures for the following:
(a) H_2S (c) NH_3
(b) CS_2 (d) NH_4Cl

31. Draw Lewis structures for the following:
(a) Ba^{2+} (d) CN^-
(b) Al^{3+} (e) HCO_3^-
(c) SO_3^{2-}

32. Draw Lewis structures for the following:
(a) I^- (d) ClO_3^-
(b) S^{2-} (e) NO_3^-
(c) CO_3^{2-}

33. Classify these molecules as polar or nonpolar:
(a) H_2O
(b) HBr
(c) CF_4

34. Classify these molecules as polar or nonpolar:
(a) F_2
(b) CO_2
(c) NH_3

35. Give the number and arrangement of the electron pairs around the central atom:
(a) C in CCl_4
(b) S in H_2S
(c) Al in AlH_3

36. Give the number and arrangement of the electron pairs around the central atom:
(a) Ga in $GaCl_3$
(b) N in NF_3
(c) Cl in ClO_3^-

37. Use VSEPR theory to predict the structure of these polyatomic ions:
(a) sulfate ion
(b) chlorate ion
(c) periodate ion

38. Use VSEPR theory to predict the structure of these polyatomic ions:
(a) ammonium ion
(b) sulfite ion
(c) phosphate ion

39. Use VSEPR theory to predict the shape of these molecules:
(a) SiH_4 (b) PH_3 (c) SeF_2

40. Use VSEPR theory to predict the shape of these molecules:
(a) SiF_4 (b) OF_2 (c) Cl_2O

41. Element X reacts with sodium to form the compound Na_2X and is in the second period on the periodic table. Identify this element.

42. Element Y reacts with oxygen to form the compound Y_2O and has the lowest ionization energy of any Period 4 element on the periodic table. Identify this element.

Additional Exercises

All exercises with blue numbers have answers in the appendix of the text.

43. Write Lewis structures for hydrazine (N_2H_4) and hydrazoic acid (HN_3).

44. Draw Lewis structures and give the molecular or ionic shape of each of the following compounds:
(a) NO_2^- (c) $SOCl_2$
(b) SO_4^{2-} (d) Cl_2O

45. Draw Lewis structures for each of the following compounds:
(a) ethane (C_2H_6)
(b) ethylene (C_2H_4)
(c) acetylene (C_2H_2)

46. Identify the element on the periodic table that satisfies each description:
(a) transition metal with the largest atomic radius
(b) alkaline earth metal with the greatest first ionization energy
(c) least dense member of the nitrogen family
(d) alkali metal with the greatest ratio of neutrons to protons
(e) most electronegative transition metal

47. Choose the element that fits each description:
(a) the lower electronegativity: As or Zn
(b) the lower chemical reactivity: Ba or Be
(c) the fewer valence electrons: N or Ne

48. Identify two reasons why fluorine has a much higher electronegativity than neon.

49. When one electron is removed from an atom of Li, it has two left. Helium atoms also have two electrons. Why is more energy required to remove the second electron from Li than to remove the first from He?

50. Group 1B elements (see the periodic table on the inside cover of your book) have one electron in their outer energy level, as do Group 1A elements. Would you expect them to form compounds such as CuCl, AgCl, and AuCl? Explain.

51. The formula for lead(II) bromide is $PbBr_2$; predict formulas for tin(II) and germanium(II) bromides.

52. Why is it not proper to speak of sodium chloride molecules?

53. What is a covalent bond? How does it differ from an ionic bond?

54. Briefly comment on the structure Na:Ö:Na for the compound Na_2O.

55. What are the four most electronegative elements?

56. Rank these elements from highest electronegativity to lowest: Mg, S, F, H, O, Cs.

57. Is it possible for a molecule to be nonpolar even though it contains polar covalent bonds? Explain.

58. Why is CO_2 a nonpolar molecule, whereas CO is a polar molecule?

59. Estimate the bond angle between atoms in these molecules:
(a) H_2S
(b) NH_3
(c) NH_4^+
(d) $SiCl_4$

60. Consider the two molecules BF_3 and NF_3. Compare and contrast them in terms of the following:
(a) valence-level orbitals on the central atom that are used for bonding
(b) shape of the molecule
(c) number of lone electron pairs on the central atom
(d) type and number of bonds found in the molecule

61. With respect to electronegativity, why is fluorine such an important atom? What combination of atoms on the periodic table results in the most ionic bond?

62. Why does the Lewis structure of each element in a given group of representative elements on the periodic table have the same number of dots?

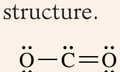
63. A sample of an air pollutant composed of sulfur and oxygen was found to contain 1.40 g sulfur and 2.10 g oxygen. What is the empirical formula for this compound? Draw a Lewis structure to represent it.

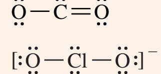
64. A dry-cleaning fluid composed of carbon and chlorine was found to have the composition 14.5% carbon and 85.5% chlorine. Its known molar mass is 166 g/mol. Draw a Lewis structure to represent the compound.

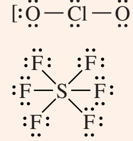
Challenge Exercises

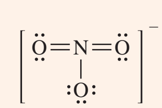
All exercises with blue numbers have answers in the appendix of the text.

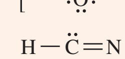
***65.** Determine whether the following Lewis structures are correct. If they are incorrect, state why and provide the correct Lewis structure.

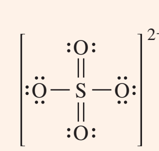
(a) CO_2 

(b) ClO_2^- 

(c) SF_6 

(d) NO_3^- 

(e) HCN 

(f) SO_4^{2-} 

***66.** The first ionization energy for lithium is 520 kJ/mol. How much energy would be required to change 25 g of lithium atoms to lithium ions? (Refer to Table 11.1.)

***67.** What is the total amount of energy required to remove the first two electrons from 15 moles of sodium atoms? (Refer to Table 11.1.)

Answers to Practice Exercises

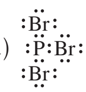
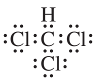
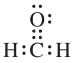
11.1 (a) 
(b) 
(c) 
(d) 

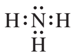
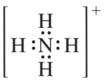
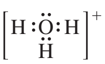
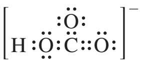
11.2 (a) Ar; K^+
(b) Ar; Mg^{2+}
(c) Ne; Al^{3+}
(d) Xe; Ba^{2+}

11.3 (a) Na_2S
(b) Rb_2O

11.4 (a) Mg_3N_2
(b) Ba_3As_2

11.5 ionic: (a), (d), (e)
covalent: (b), (c), (f)

11.6 (a)  (c)  (e) 
(b)  (d) 

11.7 (a)  (c) 
(b)  (d) 

11.8 CF_4, tetrahedral; NF_3, pyramidal; BeI_2, linear

PUTTING IT TOGETHER: Review for Chapters 10–11

Multiple Choice:

Choose the correct answer to each of the following.

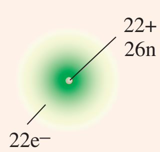
1. The concept of electrons existing in specific orbits around the nucleus was the contribution of
(a) Thomson (c) Bohr
(b) Rutherford (d) Schrödinger

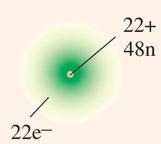
2. The correct electron structure for a fluorine atom (F) is
(a) $1s^22s^22p^5$ (c) $1s^22s^22p^43s^1$
(b) $1s^22s^22p^23s^23p^1$ (d) $1s^22s^22p^3$

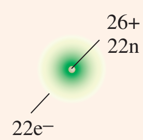
3. The correct electron structure for $_{48}Cd$ is
(a) $1s^22s^22p^63s^23p^64s^23d^{10}$
(b) $1s^22s^22p^63s^23p^64s^23d^{10}4p^65s^24d^{10}$
(c) $1s^22s^22p^63s^23p^64s^23d^{10}4p^64d^4$
(d) $1s^22s^22p^63s^23p^64s^24p^64d^{10}5s^25d^{10}$

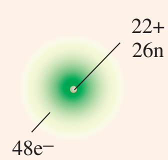
4. The correct electron structure of $_{23}V$ is
(a) $[Ar]4s^23d^3$ (c) $[Ar]4s^24d^3$
(b) $[Ar]4s^24p^3$ (d) $[Kr]4s^23d^3$

5. Which of the following is the correct atomic structure for $^{48}_{22}Ti$?

(a)


(b)


(c)


(d)


6. The number of orbitals in a *d* sublevel is
(a) 3 (c) 7
(b) 5 (d) no correct answer given

7. The number of electrons in the third principal energy level in an atom having the electron structure $1s^22s^22p^63s^23p^2$ is
(a) 2 (c) 6
(b) 4 (d) 8

8. The total number of orbitals that contain at least one electron in an atom having the structure $1s^22s^22p^63s^23p^2$ is
(a) 5 (c) 14
(b) 8 (d) no correct answer given

9. Which of these elements has two *s* and six *p* electrons in its outer energy level?
(a) He (c) Ar
(b) O (d) no correct answer given

10. Which element is not a noble gas?
(a) Ra (c) He
(b) Xe (d) Ar

11. Which element has the largest number of unpaired electrons?
(a) F (c) Cu
(b) S (d) N

12. How many unpaired electrons are in the electron structure of $_{24}Cr$, $[Ar]4s^13d^5$?
(a) 2 (c) 5
(b) 4 (d) 6

13. Groups 3A–7A plus the noble gases form the area of the periodic table where the electron sublevels being filled are
(a) *p* sublevels (c) *d* sublevels
(b) *s* and *p* sublevels (d) *f* sublevels

14. In moving down an A group on the periodic table, the number of electrons in the outermost energy level
(a) increases regularly
(b) remains constant
(c) decreases regularly
(d) changes in an unpredictable manner

15. Which of the following is an incorrect formula?
(a) NaCl (c) AlO
(b) K_2O (d) BaO

16. Elements of the noble gas family
(a) form no compounds at all
(b) have no valence electrons
(c) have an outer electron structure of ns^2np^6 (helium excepted), where *n* is the period number
(d) no correct answer given

17. The lanthanide and actinide series of elements are
(a) representative elements
(b) transition elements
(c) filling in *d*-level electrons
(d) no correct answer given

18. The element having the structure $1s^22s^22p^63s^23p^2$ is in Group
(a) 2A (c) 4A
(b) 2B (d) 4B

19. In Group 5A, the element having the smallest atomic radius is
(a) Bi (b) P (c) As (d) N

20. In Group 4A, the most metallic element is
(a) C (b) Si (c) Ge (d) Sn

21. Which group in the periodic table contains the least reactive elements?
(a) 1A (b) 2A (c) 3A (d) noble gases

22. Which group in the periodic table contains the alkali metals?
(a) 1A (b) 2A (c) 3A (d) 4A

23. An atom of fluorine is smaller than an atom of oxygen. One possible explanation is that, compared to oxygen, fluorine has
(a) a larger mass number
(b) a smaller atomic number
(c) a greater nuclear charge
(d) more unpaired electrons

24. If the size of the fluorine atom is compared to the size of the fluoride ion,
(a) they would both be the same size.
(b) the atom is larger than the ion.
(c) the ion is larger than the atom.
(d) the size difference depends on the reaction.

25. Sodium is a very active metal because
(a) it has a low ionization energy.
(b) it has only one outermost electron.
(c) it has a relatively small atomic mass.
(d) all of the above

26. Which of the following formulas is not correct?
(a) Na^+ (b) S^- (c) Al^{3+} (d) F^-

27. Which of the following molecules does not have a polar covalent bond?
(a) CH_4 (b) H_2O (c) CH_3OH (d) Cl_2

28. Which of the following molecules is a dipole?
(a) HBr (b) CH_4 (c) H_2 (d) CO_2

29. Which of the following has bonding that is ionic?
(a) H_2 (b) MgF_2 (c) H_2O (d) CH_4

30. Which of the following is a correct Lewis structure?
(a) $:\underset{..}{\ddot{O}}:C:\underset{..}{\ddot{O}}:$
(b) $:\underset{..}{\ddot{Cl}}:\overset{:\ddot{Cl}:}{\underset{:\underset{..}{Cl}:}{C}}:\underset{..}{\ddot{Cl}}:$
(c) $\ddot{Cl}::\ddot{Cl}$
(d) $:\ddot{N}:\ddot{N}:$

31. Which of the following is an incorrect Lewis structure?
(a) $H:\underset{..}{N}:H$
(b) $:\ddot{O}:H$ with H below O
(c) $H:\overset{H}{\underset{H}{C}}:H$
(d) $:N:::N:$

32. The correct Lewis structure for SO_2 is
(a) $:\underset{..}{\ddot{O}}:\ddot{S}:\ddot{O}:$
(b) $:\underset{..}{\ddot{O}}:\ddot{S}::\ddot{O}:$
(c) $:\ddot{O}::S::\ddot{O}:$
(d) $:\underset{..}{\ddot{O}}:\underset{..}{\ddot{S}}:\underset{..}{\ddot{O}}:$

33. Carbon dioxide (CO_2) is a nonpolar molecule because
(a) oxygen is more electronegative than carbon
(b) the two oxygen atoms are bonded to the carbon atom
(c) the molecule has a linear structure with the carbon atom in the middle
(d) the carbon-oxygen bonds are polar covalent

34. When a magnesium atom participates in a chemical reaction, it is most likely to
(a) lose 1 electron
(b) gain 1 electron
(c) lose 2 electrons
(d) gain 2 electrons

35. If X represents an element of Group 3A, what is the general formula for its oxide?
(a) X_3O_4 (b) X_3O_2 (c) XO (d) X_2O_3

36. Which of the following has the same electron structure as an argon atom?
(a) Ca^{2+} (b) Cl^0 (c) Na^+ (d) K^0

37. As the difference in electronegativity between two elements decreases, the tendency for the elements to form a covalent bond
(a) increases
(b) decreases
(c) remains the same
(d) sometimes increases and sometimes decreases

38. Which compound forms a tetrahedral molecule?
(a) NaCl (b) CO_2 (c) CH_4 (d) $MgCl_2$

39. Which compound has a bent (V-shaped) molecular structure?
(a) NaCl (b) CO_2 (c) CH_4 (d) H_2O

40. Which compound has double bonds within its molecular structure?
(a) NaCl (b) CO_2 (c) CH_4 (d) H_2O

41. The total number of valence electrons in a nitrate ion, NO_3^- is

(a) 12 (c) 23
(b) 18 (d) 24

42. The number of electrons in a triple bond is

(a) 3 (c) 6
(b) 4 (d) 8

43. The number of unbonded pairs of electrons in H_2O is

(a) 0 (c) 2
(b) 1 (d) 4

44. Which of the following does not have a noble gas electron structure?

(a) Na (c) Ar
(b) Sc^{3+} (d) O^{2-}

Free Response Questions:

Answer each of the following. Be sure to include your work and explanations in a clear, logical form.

1. An alkaline earth metal, M, combines with a halide, X. Will the resulting compound be ionic or covalent? Why? What is the Lewis structure for the compound?

2. "All electrons in atoms with even atomic numbers are paired." Is this statement true or false? Explain your answer using an example.

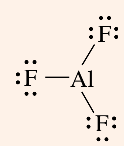
3. Discuss whether the following statement is true or false: "All nonmetals have two valence electrons in an *s* sublevel with the exception of the noble gases, which have at least one unpaired electron in a *p* sublevel."

4. The first ionization energy (IE) of potassium is lower than the first IE for calcium, but the second IE of calcium is lower than the second IE of potassium. Use an electron configuration or size argument to explain this trend in ionization energies.

5. Chlorine has a very large first ionization energy, yet it forms a chloride ion relatively easily. Explain.

6. Three particles have the same electron configuration. One is a cation of an alkali metal, one is an anion of the halide in the third period, and the third particle is an atom of a noble gas. What are the identities of the three particles (including charges)? Which particle should have the smallest atomic/ionic radius, which should have the largest, and why?

7. Why is the Lewis structure of $AlCl_3$ not written as



What is the correct Lewis structure and which electrons are shown in a Lewis structure?

8. Why does carbon have a maximum of four covalent bonds?

9. Both NCl_3 and BF_3 have a central atom bonded to three other atoms, yet one is pyramidal and the other is trigonal planar. Explain.

10. Draw the Lewis structure of the atom whose electron configuration is $1s^22s^22p^63s^23p^64s^23d^{10}4p^5$. Would you expect this atom to form an ionic, nonpolar covalent, or polar covalent bond with sulfur?