Bonding

TOPIC 6

What Scientists Know About Bonding

How does the body get energy from food?

We have all learned that we get energy to live from oxidizing foods, often sugar, $C_6H_{12}O_6$. But where does the energy actually come from? The answer is not what most people expect. We don't get the energy from breaking the bonds of sugar, because breaking bonds is an endothermic process. Rather, the energy is released after the body has broken the sugar bonds and made the bonds of the products, carbon dioxide (CO_2) and water (H_2O). The bonds of water and carbon dioxide have less energy than the bonds of sugar, and the difference between the energy of the two is what is used to live.

TOPIC

6 Bonding

Vocabulary

asymmetrical molecule	ionic bond	nonpolar covalent bond
covalent bond	Lewis dot diagram	octet
dipole-dipole forces	London dispersion forces	octet rule
double covalent bond	malleability	polar covalent bond
hydrogen bond	metallic bond	symmetrical molecule
ion	multiple covalent bond	triple covalent bond

Topic Overview

Chemical bonds provide the "glue" that holds all compounds together. There are different types of chemical bonds, and these different bond types account for the different properties of substances. In this section you will learn how the electron structure of atoms helps explain many aspects of chemical bonding. You will also use Lewis dot diagrams to aid in your understanding of electronic structure and its role in bonding.

Energy and Chemical Bonds

Chemical bonds are the forces that hold atoms together in a compound. Energy is required to overcome these attractive forces and separate the atoms in a compound. Thus, the breaking of a chemical bond is an endothermic process. If energy is required to break a bond, then the opposite process of forming a bond must release energy. The formation of a bond is an exothermic process.

When a chemical bond is formed, the resulting compound has less potential energy than the substances from which it was formed. Why? Energy is always released when a bond is formed. The greater the energy released during the formation of a bond, the greater its stability. Consider the following two reactions.

Reaction 1 $\quad B + C \rightarrow BC + 100$ joules
Reaction 2 $\quad K + L \rightarrow KL + 400$ joules

The bond formed in reaction 2 is more stable than the bond formed in reaction 1. It takes 400 joules to break apart compound KL, but only 100 joules to break apart compound BC.

Memory Jogger

In exothermic reactions, heat is a product and is released; in endothermic reactions, heat is a reactant and is absorbed.

Review Questions Set 6.1

1. As energy is released during the formation of a bond, the stability of the chemical system generally
   (1) decreases
   (2) increases
   (3) remains the same

2. Which kind of energy is stored in a chemical bond?
   (1) potential energy
   (2) kinetic energy
   (3) activation energy
   (4) ionization energy

3. Consider the reaction below.

   $$2Na(s) + Cl_2(g) \rightarrow 2NaCl(s)$$

   As the reactants form products, the stability of the chemical system
   (1) decreases
   (2) increases
   (3) remains the same

4. When a chemical bond forms between two hydrogen atoms, the potential energy of the atoms
   (1) decreases
   (2) increases
   (3) remains the same

5. Energy is released when the atoms of two elements bond together to form a compound. Compared to the total potential energy of the atoms before bonding, the total potential energy of the atoms after bonding is
   (1) higher and the compound formed is stable
   (2) higher and the compound formed is unstable
   (3) lower and the compound formed is stable
   (4) lower and the compound formed is unstable

Lewis Electron Dot Structures

In the next sections you will study the two types of bonds that commonly form. These two bond types result from the transfer of electrons from one atom to another and from the sharing of valence electrons between atoms. A simple modeling technique known as a Lewis dot diagram provides an easy method for showing how electrons are transferred or shared during bond formation. Learning to draw and interpret these diagrams is an important skill.

A **Lewis dot diagram,** also called an electron dot diagram, consists of a chemical symbol surrounded by one to eight dots representing valence electrons. See Figure 6-1.

Figure 6-2 shows the Lewis dot diagram of a sodium atom. The symbol of the element (Na) represents the sodium atom's nucleus along with all of its nonvalence electrons; this portion of the diagram is called the kernel. The kernel of every atom is positively charged. The valence electrons surround the kernel and are usually represented by small dots, x's, or o's.

Figure 6-3 shows the relationship between the groups of the periodic table and the electron dot diagrams of the atoms and ions in those groups.

	Group							
Period	**1**	**2**	**13**	**14**	**15**	**16**	**17**	**18**
1	H							He
2	Li	Be	Be	C	N	O	F	Ne
3	Na	Mg	Ai	Si	P	S	Cl	Ar
4	K	Ca	Ga	Ge	As	Se	Br	Kr

Figure 6-1. Electron dot diagrams of elements in Periods 1 through 4

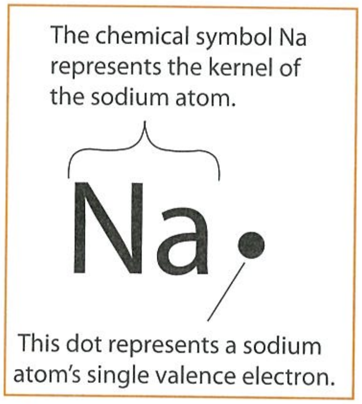


Figure 6-2. The Lewis dot diagram of a sodium atom: The kernel, represented by the symbol Na, contains 11 protons and 10 electrons; the kernel has a 1+ overall charge. The 1+ charge of the kernel is balanced by the 1− charge of the single valence electron. Overall, the sodium atom is neutral.

There are slight variations in how electrons are arranged around the kernel in an electron dot diagram. Often, the first two electrons are placed at the 12 o'clock position. As the next three electrons are added, they are placed singly at the 3, 6, and 9 o'clock positions. The last three electrons are added to make pairs in the same order, at the 3, 6, and 9 o'clock positions.

When atoms gain or lose electrons they become charged particles called **ions.** Notice in Figure 6-3 that ions are represented differently than atoms. When metals react, they do so by losing their valence electrons. Thus, there aren't any dots in the electron dot diagrams of the ions for sodium (Na^{+}), magnesium (Mg^{2+}), or aluminum (Al^{3+}). The square brackets around the kernel and the ionic charge written as a superscript outside the brackets indicate the diagram is that of an ion. Whereas metals lose electrons to form ions, nonmetals gain electrons to form ions. The valence electrons gained when a nonmetal ion forms are shown in its electron dot diagram. It is common to show the added valence electrons with a different symbol than that used to represent the atom's originally present valence electrons.

The number of valence electrons to use in an electron dot diagram can also be determined by consulting the periodic table. The periodic table often lists the electron configuration for the elements using a modified Bohr model. For example, the electron configuration for sodium (Na) is listed as 2-8-1. This notation indicates that there is only one outermost, or valence electron for a sodium atom. The other 10 electrons belong to the kernel of the atom. The corresponding electron dot diagram for the sodium atom would show this one valence electron as a single dot, outside the kernel.

Group number	1	2
Atoms	Na	Mg
Ions	$[Na]^{+}$	$[Mg]^{2+}$

Group number	13	14	15	16	17	18
Atoms	Al	Si	P	S	Cl	Ar
Ions	$[Al]^{3+}$	*	$[P]^{3-}$	$[S]^{2-}$	$[Cl]^{-}$	*

*Si and Ar do not ordinarily form ions.

Figure 6-3. Electron dot diagrams of some Period 3 elements and ions

Lewis Electron-Dot Diagrams of Compounds

Lewis diagrams can also be used to show how atoms combine to form molecules. When atoms combine to form diatomic molecules, single, double, and triple covalent bonds can be shown.

The Lewis structure of a hydrogen atom is simply

H•

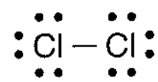
When two hydrogen atoms combine to form the hydrogen molecule (H_2), they share the two electrons between them in a nonpolar covalent bond:

H:H

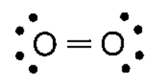
Instead of using dots to show the pair of electrons, a single dash may be used to show the covalent bond.

H—H

The members of column 17, the halogens, also form single bonds between them in their diatomic molecules. In addition to the single covalent bond, each atom has 6 additional valence electrons for a complete octet of 8.

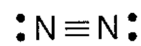


Oxygen has 6 valence electrons with two atoms combining to form O_2. In this case, the two oxygen atoms share two pairs of electrons



Notice that each oxygen atom has a total of 8 valence electrons.

Nitrogen has five valence electrons and forms the diatomic N_2 molecule by forming a triple covalent bond



When drawing Lewis diagrams for compounds, a pair of electrons forming a covalent bond is usually represented by a dashed line, while any unpaired electrons are represented by a pair of dots.

In drawing Lewis diagrams for compounds it is helpful to remember that the octet rule calls for each atom to have 8 valence electrons. Hydrogen is an exception, having a maximum of 2 electrons. Try to follow a pattern when drawing diagrams of compounds.

The following steps are useful in determining the dot diagrams of compounds.

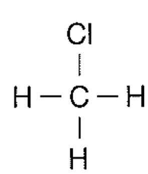
1. Determine the total number of valence electrons of the atoms in the compound.

   Consider the compound CH_3Cl:

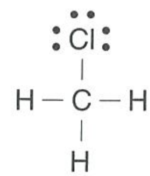
1 carbon with 4 valence electrons	=	4
3 hydrogens with 1 electron each	=	3
1 chlorine with 7 valence electrons	=	7
Total valence electrons	=	14

   The final diagram must contain 14 electrons.

2. Arrange the atoms to show bonds between them. The central atom often has the smallest electronegativity value, and generally appears once in the formula. Remember that hydrogen cannot be the central atom since it can only form one single covalent bond. If more than one atom of an element is present, they generally surround the central atom. Use a dash to show a covalent bond between the atoms. Remember that each dash represents two electrons in a covalent bond.



These 4 bonds account for 8 of the 14 electrons needed. The carbon atom has a complete octet of electrons. Distribute the remaining 6 electrons around the chlorine. Check to see that each atom has an octet of electrons except for hydrogen which needs only two. In this example the six remaining electrons should be distributed around the chlorine to provide it with an octet of electrons.



If each of the atoms does not have an octet of electrons, you will probably also have some "left over" or unassigned electrons. These can be placed as double bonds or triple bonds between atoms that do not have complete octets.

Consider the compound C_2H_2:

2 carbon atoms with 4 electrons	=	8 electrons
2 hydrogen atoms with 1 electron	=	2 electrons
Total	=	10 electrons

Join the carbons and hydrogens with bonds:

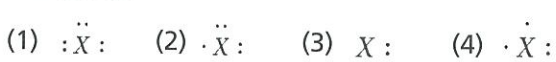
H—C—C—H

The three bonds account for 6 of the 10 electrons. Distributing the four remaining electrons around the carbons will not produce octets of electrons, but forming a triple bond between the two carbons will.

H—C≡C—H

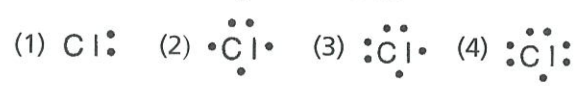
Review Questions — Set 6.2

6. Atom X has an electron configuration of 2-8-2. Which electron dot diagram correctly represents this atom?



7. A) Draw a Lewis dot diagram of water
B) Draw a Lewis dot diagram of carbon dioxide

8. Which electron dot diagram represents an atom of chlorine in the ground state?



9. A) Draw an electron dot diagram of ammonia.
B) Draw an electron dot diagram of the ammonium ion

Metallic Bonds

Metallic atoms have few valence electrons and low ionization energies. The bonds holding metallic atoms together in the solid and liquid phases, however, are apparently strong, as metals have fairly high melting and boiling points. A metallic atom may be considered to have a central portion, or kernel, made up of its nucleus and its nonvalence electrons. The atom's valence electrons surround the kernel. The kernels of the metallic atoms making up a metallic solid are arranged in the fixed positions of a crystalline lattice. The valence electrons move freely throughout the crystal and do not belong to any given atom. The freely moving valence electrons give metals properties of good electrical and thermal conductivity. A **metallic bond** results from the force of attraction of the mobile valence electrons for an atom's positively charged kernel.

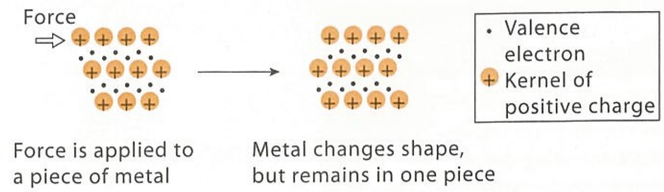


Figure 6-4. Metallic bonding: When hammered, the positively charged kernels are moved, but they are still attracted by the valence electrons.

As shown in Figure 6-4, metals can be hammered into shapes, which is a property known as **malleability.** Although hammering forces the kernels to move to new locations, they are still surrounded by the valence electrons. The freedom of the valence electrons often leads to the use of the term "sea of mobile electrons" to describe metallic bonding.

Review Questions — Set 6.3

10. Metallic bonding occurs between metal atoms that have
- (1) full valence orbitals, low ionization energies
- (2) full valence orbitals, high ionization energies
- (3) vacant valence orbitals, low ionization energies
- (4) vacant valence orbitals, high ionization energies

11. Which element has a crystalline lattice through which electrons flow freely?
- (1) bromine
- (2) calcium
- (3) carbon
- (4) sulfur

12. Which element has properties of good electrical conductivity and luster and exists as a liquid at STP?

(1) Hg (2) Br (3) I (4) C

The Octet Rule

You may recall that the noble gases of Group 18 are extremely stable and undergo very few chemical reactions. Although there are a few stable compounds of argon and xenon combined with fluorine, these compounds are some of the rare exceptions to the noble gases' lack of reactivity. What is it about the structure of the noble gases that leads to their stability? The answer lies in their number of valence electrons. Except for helium (which only has two valence electrons) all of the noble gas atoms have eight valence electrons. The configuration of eight valence electrons is known as an **octet.** An octet represents the maximum number of valence electrons that an atom can have.

Ordinary chemical reactions result in changes to the valence electron configurations of the atoms involved. A complete octet of eight valence electrons results in an exceptionally stable electron configuration. This stable configuration is the reason why noble gases are so unreactive. Note that the noble gas helium requires only two electrons to fill its valence shell. The **octet rule** states that atoms generally react by gaining,

Digging Deeper

The composition of all substances cannot be explained by the octet rule. For example, boron hydride (BH_3) and sulfur hexafluoride (SF_6) do not obey the octet rule because their stable electron configurations do not contain eight valence electrons. Chemists use concepts such as resonance and expanded valence shells to explain some of the compounds that do not conform to the octet rule.

losing, or sharing electrons in order to achieve a complete octet of eight valence electrons—the configuration of a noble gas.

Covalent Bonds

When two atoms approach one another, their electrons repel each other, tending to push the atoms apart. Their positive nuclei also repel each other. An attractive force between the atoms comes from the attraction of the positively charged nucleus of one atom for the negatively charged electrons of the other atom. Chemical bonds occur when the attractive forces between atoms are greater than the repulsive forces.

A **covalent bond** is formed when two nuclei share electrons in order to achieve a stable arrangement of electrons. Covalent bonds are often formed between two nonmetal atoms of the same element. The diatomic chlorine molecule (Cl_2) is an example of a covalent bond. Atoms of different elements may also combine to form covalent bonds. Sulfur and oxygen combine covalently to form sulfur dioxide (SO_2).

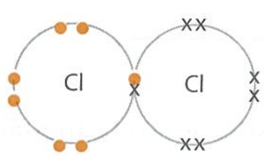


Figure 6-5. A diatomic chlorine molecule: Each chlorine atom has an octet of valence electrons.

Nonpolar Covalent Bonds When one chlorine atom forms a bond with another chlorine atom, they share a pair of electrons between them and form diatomic Cl_2. See Figure 6-5. In this way, the shared electrons of the bond are counted as part of each atom's stable octet. This bond is a **nonpolar covalent bond** because the attraction of the two chlorine nuclei for the shared electrons is equal, causing the pair of electrons to be shared equally. Nonpolar covalent bonds are formed between atoms having equal or close electronegativity values. Similar nonpolar bonds exist between each of the other Group 17 atoms in their diatomic molecules, namely F_2, Br_2, and I_2.

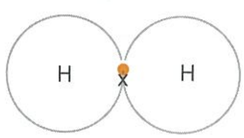


Figure 6-6. A diatomic hydrogen molecule: Each hydrogen atom has two valence electrons in its completely filled valence level.

In a similar way, two hydrogen atoms share a pair of electrons to form diatomic H_2. See Figure 6-6. In this case there is not an octet of electrons because hydrogen only needs two electrons to fill its valence level.

Multiple Covalent Bonds Atoms may share more than one pair of electrons, resulting in the formation of a **multiple covalent bond.** Oxygen atoms combine with other oxygen atoms to form O_2 and achieve a stable octet configuration. In achieving this arrangement, the oxygen nuclei must share two pairs of electrons. The sharing of two pairs of valence electrons results in the type of multiple covalent bond called a **double covalent bond.** See Figure 6-7.

Diatomic nitrogen (N_2) forms when two nitrogen atoms share three pairs of valence electrons. The sharing of three pairs of valence electrons results in a **triple covalent bond.** See Figure 6-8.

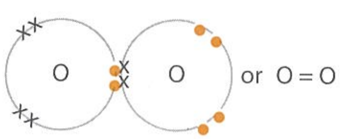


Figure 6-7. A diatomic oxygen molecule with a double covalent bond: Each oxygen atom shares two pairs of valence electrons.

Polar Covalent Bonds The different atoms involved in a bond are likely to have unequal electronegativity values. As you may recall, electronegativity is a measure of an atom's tendency to attract bonded electrons. When the electronegativity values of the two atoms in a covalent bond are different, the sharing of electrons in the bond is unequal. The unequal sharing of electrons in a covalent bond results in a **polar covalent bond.** The element with the higher electronegativity value attracts the shared electrons more strongly, causing that portion of the molecule to acquire a partially negative charge. Likewise, the other end of the polar covalent bond acquires a partially positive charge. Hydrogen chloride (HCl) and hydrogen iodide (HI) are examples of polar covalent compounds. Look at the electron distribution for different types of bonds, as shown in Figure 6-11.

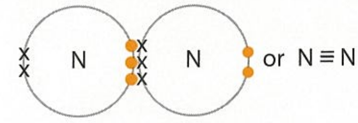


Figure 6-8. A diatomic nitrogen molecule with a triple covalent bond: Each nitrogen atom shares three pairs of valence electrons.

SAMPLE PROBLEM

Which of the following bonds is the most polar in nature? (a) O_2 (b) HCl (c) NH_3 (d) HBr

SOLUTION: Consult the *Reference Tables for Physical Setting/Chemistry* to determine the electronegativity values of each element. For each bond, determine the difference of the two electronegativity values. The bond with the smallest difference in electronegativity is the least polar, whereas the bond with the largest difference in electronegativity is the most polar.

(a) Because both oxygen atoms have the same electronegativity value (3.4), the difference is 0. This is a nonpolar covalent bond.

(b) Chlorine has an electronegativity of 3.2, and hydrogen has an electronegativity of 2.2. The electronegativity difference is 1.0, and the bond is polar covalent.

(c) Nitrogen has an electronegativity of 3.0, and hydrogen has an electronegativity of 2.2. The electronegativity difference is 0.8, and the bond is polar covalent.

(d) Bromine has an electronegativity of 3.0, and hydrogen has an electronegativity of 2.2. The electronegativity difference is 0.8, and the bond is polar covalent.

The largest electronegativity difference occurs in the bonds of HCl, hence it is the most polar in nature. The bond in diatomic oxygen is the least polar; diatomic oxygen has a nonpolar covalent bond.

Review Questions — Set 6.4

13. The correct electron dot diagram for hydrogen chloride is

(1) $H:Cl$
(2) $:\underset{\cdot\cdot}{\overset{\cdot\cdot}{H}}:Cl$
(3) $H:\underset{\cdot\cdot}{\overset{\cdot\cdot}{Cl}}:$
(4) $:H:\overset{\cdot\cdot}{Cl}:$

14. A covalent bond forms when

(1) two nuclei share electrons in order to achieve a complete octet of electrons
(2) atoms form ions and then electrostatic forces of attraction bond the ions together
(3) repulsive forces between atoms are greater than the attractive forces
(4) a metal atom combines with a nonmetal atom

15. Which of the following bonds is the most polar in nature?

(1) Cl_2
(2) HCl
(3) HBr
(4) HI

16. Polar covalent bonds are caused by

(1) unbalanced ionic charges
(2) unequal electronegativity values
(3) the transfer of electrons from one atom to another
(4) equally shared valence electrons

17. The bond in a diatomic nitrogen molecule (N_2) is best described as

(1) polar
(2) polar double covalent
(3) nonpolar triple covalent
(4) polar ionic

Molecular Substances

A <u>molecule</u> is the smallest discrete particle of an element or compound formed by covalently bonded atoms. Each atom in a molecule usually has the electron configuration of a noble gas. Molecular substances may exist as solids, liquids, or gases, depending on the strength of the forces of attraction between the molecules.

Molecules generally have properties associated with covalent bonding. Molecules are generally soft, are poor conductors of heat and electricity, and have relatively low melting and boiling points.

Water (H_2O), carbon dioxide (CO_2), and ammonia (NH_3) are examples of molecular compounds. The diatomic gases, such as oxygen (O_2) and nitrogen (N_2), are also molecular substances. Some common molecular substances are represented in Figure 6-9.

Polar Molecules Although a molecule may contain polar bonds, it does not necessarily follow that the molecule itself is polar. Molecules such as hydrogen chloride and water have polar bonds and are also polar molecules. Carbon dioxide and carbon tetrachloride, however, contain polar bonds but are not polar molecules. To understand why these molecules are polar or nonpolar, you must examine their molecular shapes.

Memory Jogger

Symmetrical molecules have identical parts on each side of an axis; **asymmetrical molecules** lack identical parts on each side of an axis.

Study the shapes of the molecules in Figure 6-10. The water, hydrogen chloride, and ammonia molecules are asymmetrical, whereas the carbon dioxide and carbon tetrachloride molecules are symmetrical. The symmetrical shape of the carbon dioxide and carbon tetrachloride molecules causes the pull of the various polar bonds to be offset by other bonds—the net result is a nonpolar molecule.

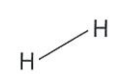
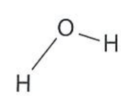
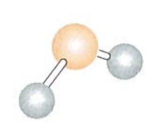
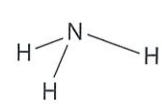
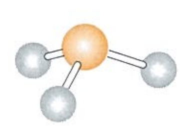
Name	Molecular Formula	Structural Formula	Ball-and-Stick Model	Space-Filling Model
Hydrogen	H_2			
Water	H_2O			
Ammonia	NH_3			

Figure 6-9. **Models of some common molecular substances**

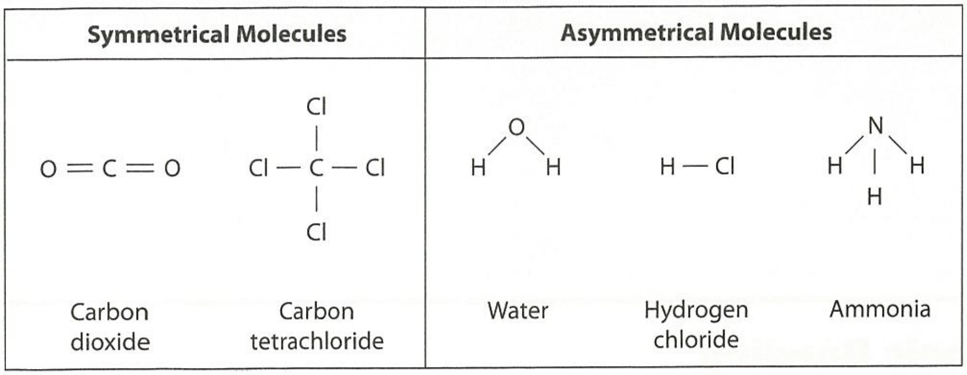


Figure 6-10. Five compounds with polar bonds: The symmetrical molecules are nonpolar; the asymmetrical molecules are polar.

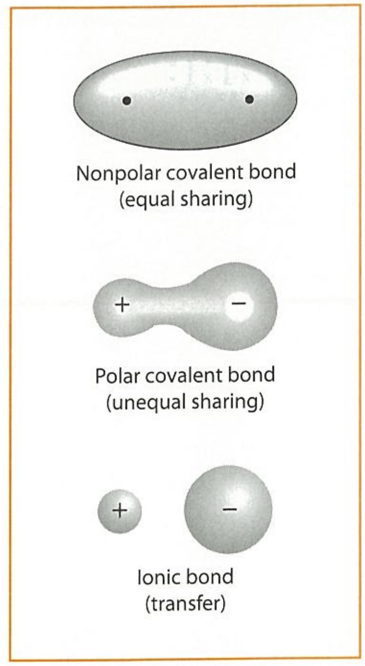


Figure 6-11. Three bond types that are dependent on electron distribution

The charge distribution within a bond affects the nature of the bond. Figure 6-11 shows the changes in bond types with different electron distributions. The nonpolar covalent bond has an even electron distribution. As the electron distribution becomes unequal, a polar covalent bond is formed. When the electron is actually transferred, an ionic bond is formed. You will learn about ionic bonds in the next section.

Review Questions — Set 6.5

18. Which electron dot diagram represents a polar molecule?

(1) H : H

(2) H : C̈l :

(3)
```
   H
   ..
H : C : H
   ..
   H
```

(4) : F̈ : F̈ :

19. Which electron dot diagram represents a polar molecule?

(1) :Ö::C::Ö:

(2)
```
  H
H:C:H
  H
```

(3)
```
H:Ö:
  H
```

(4)
```
   :C̈l:
:C̈l:C:C̈l:
   :C̈l:
```

20. Which electron dot diagram represents H_2?

(1) H • H

(2) H : H

(3) :H • H:

(4) :H : H:

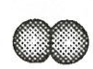
21. The diagram below represents a hydrogen fluoride molecule.

H : F̈ :

This molecule is best described as

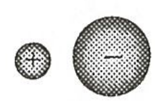
(1) polar with polar covalent bonds
(2) polar with nonpolar covalent bonds
(3) nonpolar with polar covalent bonds
(4) nonpolar with nonpolar covalent bonds

22. Which diagram best represents a polar molecule?

(1) 
Cl_2

(2) 
H_2

(3) 
HCl

(4) 

NaCl

23. Which is the correct electron dot formula for a chlorine molecule?

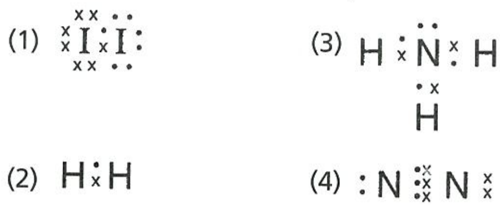
(1) Cl • Cl

(2) •C̣l:C̣l•

(3) :C̈l•C̈l:

(4) :C̈l:C̈l:

24. Which molecule contains a polar covalent bond?



25. Which electron dot diagram represents the atom in Period 4 with the highest first ionization energy?

Ionic Bonding

Recall that when atoms gain or lose electrons they become charged particles called ions. An ionic bond is formed when ions bond together because of the electrostatic attraction of oppositely charged ions. Figure 6-12 shows the formation of an ionic bond by the transfer of an electron.

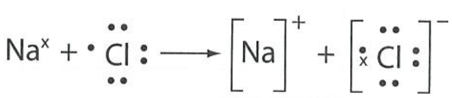


Figure 6-12. The formation of an ionic bond by the transfer of an electron

Ion Formation in Metals As atoms form ions, they tend to do so in a manner that results in the formation of ions with noble gas electron configurations. Consider the metals of Group 1 with their single valence electron. When a Group 1 metallic atom reacts, it loses its electron and forms an ion with a 1+ charge. By losing its valence electron, the atom acquires the octet arrangement of a noble gas. The resulting ion is quite stable. The loss of the valence electron from the atom's outer energy level results in the ions having a smaller radius than the atom from which it was formed.

Ion Formation in Nonmetals Now consider a nonmetal atom from Group 17 with its seven valence electrons. This nonmetal atom reacts by gaining a single electron to form an ion with a 1− charge. By gaining an electron, the atom has acquired the noble gas electron configuration (eight valence electrons). The resulting ion is quite stable. The addition of the valence electron to the atom's outer energy level results in the ions having a larger radius than the atom from which it was formed. See Table 6-1 that summarizes the reactions of metals and nonmetals.

Ion Formation and the Octet Rule Look at Figure 6-13. Each of the atoms has a different number of electrons. When each atom reacts, it does so by losing or gaining electrons. Notice that as ions, each has an octet of valence electrons—the same number of electrons as the noble gas neon.

When metallic atoms lose electrons to form positive ions, they acquire the electron configuration of the noble gas preceding them on the

Table 6–1. Reactions of Metals and Nonmetals

When metals react they:	When nonmetals react they:
1. Lose electrons	1. Gain electrons
2. Become positively charged	2. Become negatively charged
3. Have smaller radii	3. Have larger radii
4. Acquire the electron configuration of a noble gas	4. Acquire the electron configuration of a noble gas

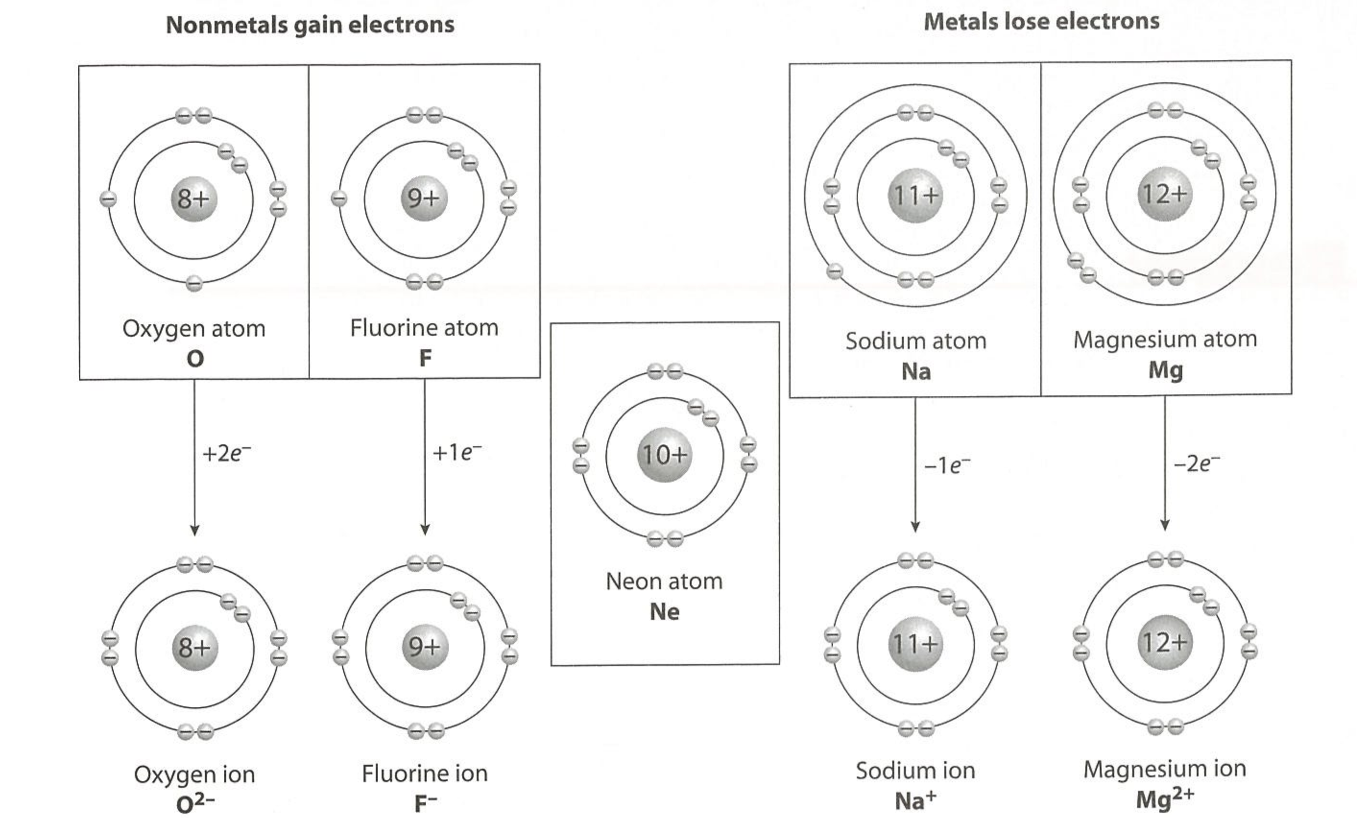


Figure 6-13. **Atoms tend to gain or lose electrons and acquire the electron configuration of the nearest noble gas: All of the ions and neon are isoelectronic, that is, they have the same electron configuration.**

periodic table. When nonmetals gain electrons, they form negative ions and acquire the electron configuration of the noble gas that follows them on the table.

Electronegativity difference between atoms in the bond	Bond type	Covalent character	Ionic character
zero	nonpolar covalent	↑ Increases	↓ Increases
↓ intermediate	↓ polar covalent		
↓ large	↓ ionic		

Figure 6-14. **Electronegativity and bond type**

Electronegativity As you just learned, as the electronegativity difference between two atoms in a bond increases, the bond becomes more polar in nature. Figure 6-14 summarizes the relationship between electronegativity and bond type. As shown, as the electronegativity difference increases, the bond becomes more ionic in character. At some point the bond can no longer be considered as a sharing of electrons, but rather as a bond in which one or more electrons have actually been transferred from one atom to another. The transfer of electrons from one atom to another results in an ionic bond.

If the electronegativity difference between the bonding atoms is 1.7 or greater, the bond is generally considered to be ionic. Ionic bonds are generally formed between metals and nonmetallic atoms. Refer to Table S in the *Reference Tables for Physical Setting/Chemistry* for Ⓡ electronegativity values.

Polyatomic Ions Any compound containing a polyatomic ion must also contain an ionic bond. Consider the ionic compound ammonium carbonate $(NH_4)_2CO_3$. The ammonium ion is attracted to the carbonate ion by an ionic bond. Within the ammonium and carbonate ions, there is a second type of bond. The nitrogen atom and hydrogen atoms of the ammonium ion and the carbon atom and oxygen atoms of the carbonate ion are held together by covalent bonds. That is, the bonds holding the atoms within the ion are themselves covalent. All compounds with polyatomic ions contain both ionic and covalent bonds.

Review Questions — Set 6.6

26. When a calcium atom loses its valence electrons, the ion formed has an electron configuration that is the same as an atom of

(1) Cl (2) Ar (3) K (4) Sc

27. The bond between which pair of elements is the least ionic in character?

(1) H-F (2) H-Cl (3) H-I (4) H-O

28. Which compound has the greatest degree of ionic character?

(1) NaF (2) MgF_2 (3) AlF_3 (4) SiF_4

29. Which compound contains a bond with the least ionic character?

(1) CO (2) CaO (3) K_2O (4) Li_2O

30. Look at the electron dot formula shown below.

$$\begin{array}{c} \mathrm{H} : \overset{..}{\underset{..}{\mathrm{X}}} : \\ \mathrm{H} \end{array}$$

The attraction of X for the bonding electrons would be greatest when X represents an atom of

(1) S (2) O (3) Se (4) Te

31. Which substance contains a bond with the greatest ionic character?

(1) KCl (2) HCl (3) Cl_2 (4) F_2

32. Which compound contains both ionic and covalent bonds?

(1) HBr (2) CBr_4 (3) NaBr (4) NaOH

33. Which element forms an ionic bond with fluorine?

(1) fluorine
(2) carbon
(3) potassium
(4) oxygen

34. Which compound is described correctly?

(1) $BaCl_2$ is covalent and molecular.
(2) H_2O_2 is covalent and empirical.
(3) H_2O is ionic and molecular.
(4) NaCl is ionic and empirical.

35. The elements Li and F combine to form an ionic compound. The electron configurations within this compound are the same as the electron configurations of atoms of Group

(1) 1 (2) 14 (3) 17 (4) 18

36. An atom of which element has the same electron configuration as O^{2-}?

(1) Li (2) Na (3) Ar (4) Ne

37. Which ion has the electron configuration of a noble gas?

(1) Cu^{2+} (2) Fe^{2+} (3) Ca^{2+} (4) Hg^{2+}

Distinguishing Bond Types

Metallic, covalent, and ionic bonds have different properties that can be used to distinguish among them. Metals generally have high melting points. Mercury, which is a liquid at standard conditions, is an exception, as are the metals of Group 1. Ionic compounds also have high melting points, while covalently bonded molecules have relatively low melting points.

Of course, melting point alone will not accurately differentiate between bond types. Other properties must also be considered. Metallic bonds are the only type that result in good thermal and electrical conductivity. Neither ionic solids nor molecular solids are good conductors. Ionic substances become conductors when they are melted (fused) or dissolved in aqueous solutions.

Molecular substances are held together by covalent bonds. Thus, there are no charged particles to conduct an electric current. Molecular substances are poor conductors, whether they are in the solid state, in the liquid state, or in aqueous solution.

The three bond types also differ in hardness. Ionic and metallic solids are generally hard, while a majority of covalently bonded solids are soft. Table 6-2 summarizes some of the properties of the various bond types.

Table 6-2. Properties of Metallic, Ionic, and Covalent Bonds

Bond Type	Melting and Boiling Points	Hardness	Conductivity		
			Solid	Liquid	Aqueous
Metallic	High	Hard	Yes	Yes	Yes
Covalent	Low	Soft	No	No	No
Ionic	High	Hard	No	Yes	Yes

Digging Deeper

Why do ionic substances conduct electricity when melted or when in aqueous solution, but not when in the solid state? In solids, electrical current is carried by electrons. In liquids, however, electrical current is carried by charged particles (ions). The ions in an ionic solid are not able to move from their fixed positions and conduct electrical current. When the ionic solid is melted or dissolved in water, the ions are free to move and conduct electrical current.

Intermolecular Forces

Dipole-Dipole Forces Just as atoms are held together by chemical bonds, molecules also have attractive forces acting on them in the solid and liquid states. It is fairly easy to understand how polar molecules are attracted to each other. Because polar molecules have positive and negative ends, polar molecules are also called <u>dipoles</u>. The positive area of one dipole molecule is attracted to the negative portion of an adjacent dipole molecule. These attractive forces are known as **dipole-dipole forces.**

Hydrogen Bonds **A hydrogen bond** is an intermolecular bond between a hydrogen atom in one molecule and a nitrogen, oxygen, or fluorine atom in another molecule. See Figure 6-15. Perhaps the most important example of hydrogen is found in water.

Water is a polar molecule, so you might expect it to be held to other water molecules by dipole-dipole attractions. However, if only dipole-dipole forces acted to hold water molecules together, all of the water on earth

Memory Jogger

You should recall the difference between *intra-* and *inter-*. When you play *intra*mural sports, you play against students from within your own school. In chemistry, *intra*molecular forces refer to the covalent bonds between atoms within a molecule. When you play *inter*scholastic sports, you play against teams from other schools. In chemistry, *inter*molecular forces refer to attractions between and among molecules.

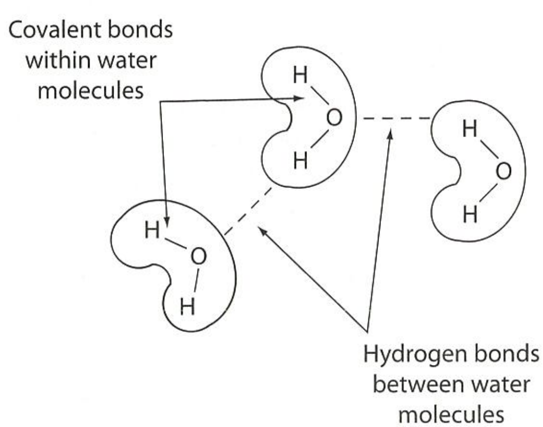


Figure 6-15. **Covalent and hydrogen bonding in water**

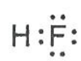
Part B–1

16 Which statement is true concerning the reaction $N(g) + N(g) \rightarrow N_2(g) +$ energy?
(1) A bond is broken and energy is absorbed.
(2) A bond is broken and energy is released.
(3) A bond is formed and energy is absorbed.
(4) A bond is formed and energy is released.

17 The bond between which pair of elements is the least ionic in character?
(1) H—F
(2) H—Br
(3) H—S
(4) H—O

18 Which compound has the greatest degree of ionic character?
(1) $NaCl$
(2) $MgCl_2$
(3) $AlCl_3$
(4) $SiCl_4$

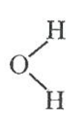
19 Look at the electron dot diagram below.

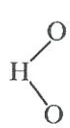


The electrons in the bond between hydrogen and fluorine are more strongly attracted to the atom of
(1) hydrogen, which has the higher electronegativity
(2) hydrogen, which has the lower electronegativity
(3) fluorine, which has the higher electronegativity
(4) fluorine, which has the lower electronegativity

20 Which type of bond is formed between the two chlorine atoms in a chlorine molecule?
(1) polar covalent
(2) nonpolar covalent
(3) metallic
(4) ionic

21 Which diagram best represents the structure of a water molecule?

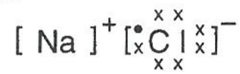
(1) 

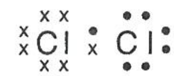
(2) 

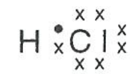
(3) O—H—O

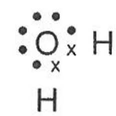
(4) H—H—O

22 Which electron dot formula represents a substance that contains a nonpolar covalent bond?

(1) 

(2) 

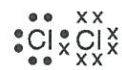
(3) 

(4) 

23 Which electron dot diagram represents a molecule that has a polar covalent bond?

(1) 

(2) 

(3) 

(4) 

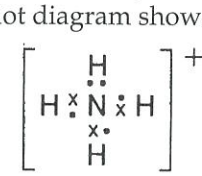
24 When a sodium atom reacts with a chlorine atom to form a compound, the electron configuration of the ions forming the compound are the same as those in which noble gases?
(1) krypton and neon
(2) krypton and argon
(3) neon and helium
(4) neon and argon

25 Which atom will form an ionic bond with a Br atom?
(1) N
(2) Li
(3) O
(4) C

26 A white crystalline salt conducts electricity when it is melted and when it is dissolved in water. Which type of bond does this salt contain?
(1) ionic
(2) metallic
(3) nonpolar covalent
(4) polar covalent

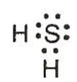
27 In which molecule is hydrogen bonding the strongest?
(1) HF
(2) HCl
(3) HBr
(4) HI

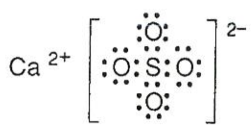
28 Which statement correctly describes the bonds in the electron dot diagram shown below?

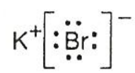


(1) One of the bonds is ionic.
(2) One of the bonds is metallic.
(3) All of the bonds are covalent.
(4) None of the bonds are covalent.

29 Which electron dot diagram best represents a compound that contains both ionic and covalent bonds?

(1) 

(2) 

(3) 

(4) 

30 Which statement explains why Br_2 is a liquid at STP and I_2 is a solid at STP?

(1) Molecules of Br_2 are polar, and molecules of I_2 are nonpolar.

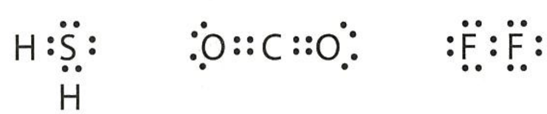
(2) Molecules of Br_2 are nonpolar, and molecules of I_2 are polar.

(3) Molecules of Br_2 have stronger intermolecular forces than molecules of I_2.

(4) Molecules of I_2 have stronger intermolecular forces than molecules of Br_2.

Parts B-2 and C

Base your answers to questions 31 through 33 on your knowledge of chemical bonding and on the Lewis electron-dot diagrams of H_2S, CO_2, and F_2 below.



31 Which atom, when bonded as shown, has the same electron configuration as an atom of argon?

32 Explain, in terms of *structure* and/or *distribution of charge*, why CO_2 is a nonpolar molecule.

33 Explain, in terms of *electronegativity*, why a C=O bond in CO_2 is more polar than the F—F bond in F_2.

Base your answers to questions 34 through 36 on the following information.

In 1864, the Solvay process was developed to make soda ash. One step in the process is represented by the balanced equation below.

$$NaCl + NH_3 + CO_2 + H_2O \rightarrow NaHCO_3 + NH_4Cl$$

34 Write the chemical formula for *one* compound in the equation that contains *both* ionic bonds and covalent bonds.

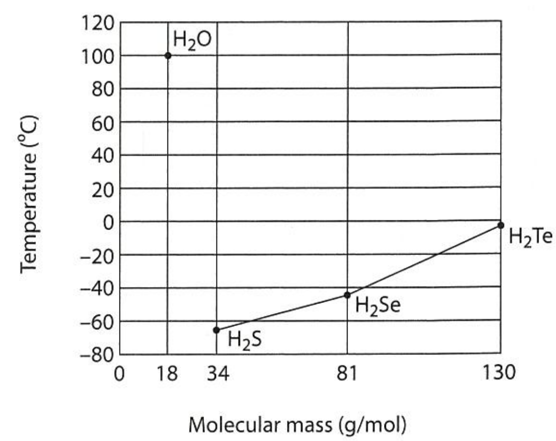
35 Explain, in terms of electronegativity difference, why the bond between hydrogen and oxygen in a water molecule is more polar than the bond between hydrogen and nitrogen in an ammonia molecule.

36 Draw a Lewis electron-dot diagram for the reactant containing nitrogen in the equation.

Base your answers to questions 37 through 39 on the following information.

The hydrides of Group 16 are all dipoles. The graph below shows the boiling points of H_2S, H_2Se, and H_2Te.

Boiling Points of the Group 16 Hydrides



37 Based on the plotted boiling points, write the formula of the compound that shows the greatest intermolecular forces of attraction.

38 As molecular mass is considered from 34 g/mol to 130 g/mol, what general trend is observed in the boiling points of the Group 16 compounds?

39 Construct a Lewis electron-dot structure to illustrate the compound with the lowest boiling point.

Base your answers to questions 40 and 41 on the information below.

Physical Properties of CF_4 and NH_3 at Standard Pressure

Compound	Melting Point (°C)	Boiling Point (°C)	Solubility in Water at 20.0°C
CF_4	−183.6	−127.8	insoluble
NH_3	−77.7	−33.3	soluble

40 State evidence that indicates NH_3 has stronger intermolecular forces than CF_4.

41 Draw a Lewis electron-dot formula for CF_4.

Use the following information to answer questions 42 through 45.

Element X is a solid metal that reacts with chlorine to produce a water-soluble binary compound.

42 What is the most likely type of bond formed between the metal and chlorine?

43 The metal combines with chlorine to form a compound with the formula MCl_2. What Group of the Periodic Table is M a member of?

44 After reacting, with which noble gas are the metal and chlorine isoelectronic?

45 Explain, in terms of particles, why an aqueous solution of the binary compound conducts an electric current.

Use the following information to answer questions 46 through 48.

Metal X is recovered from a rock sample and found to combine with iodine in a 1:1 ratio according to the equation

$$2X + I_2 \rightarrow 2XI$$

46 What is the most likely type of bond formed between element X and iodine?

47 Explain, in terms of electrons, why element X is a good conductor of electricity.

48 Once bonded, the iodine atom will have an electron configuration resembling that of what element?

Base your answers to questions 49 through 51 on the following information.

Ozone, O_3(g), is produced from oxygen, O_2(g) by electrical discharge during thunder-storms. The following chart provides a description of each gas that includes some basic physical properties.

Table of Properties for Oxygen and Ozone

	Oxygen gas	Ozone gas
Molecular Formula	O_2	O_3
Appearance	transparent	bluish colored
Melting Point	54.36 K	80.7 K
Boiling Point	90.20 K	161.3 K

49 Identify the type of bonding between the atoms in an oxygen molecule.

50 Explain, in terms of electron configuration, why an oxygen molecule is more stable than an oxygen atom.

51 Explain, in terms of intermolecular forces of attraction, the difference in boiling points between the two substances.