Honors Chemistry Chapter 6 Sections 2, 3, and 5.

Chapter 6 Section 2: Covalent bonding and Molecular Compounds pgs. 178-189

Objectives:

1. **Define molecule and molecular formula.**
2. **Explain the relationships among potential energy, distance between approaching atoms, bond length, and bond energy.**
3. **State the octet rule.**
4. **List the six basic steps used in writing Lewis structures.**
5. **Explain how to determine Lewis structures for molecules containing single bonds, multiple bonds, or both.**
6. **Explain why scientists use resonance structures to represent some molecules.**

Vocabulary: Define the following.

1. **molecule--**
2. **molecular compound--**
3. **chemical formula--**
4. **molecular formula--**
5. **bond energy--**
6. **electron-dot notation--**
7. **Lewis structure--**
8. **single bond--**
9. **multiple bond--**
10. **resonance--**

A molecule consists of two or more neutral atoms joined by covalent bonds. A molecule's chemical formula is called a molecular formula. Covalent bonds form between nonmetals and nonmetals.

Diatomic molecules contain two atoms of the same element of which there are seven formed from elements alone. The following seven diatomic molecules should be memorized tonight, they

are: hydrogen, H_2, nitrogen, N_2, oxygen, O_2, fluorine, F_2, chlorine, Cl_2, bromine, Br, and iodine, I_2.

An empirical formula is the lowest possible ratio. In ionic bonds this is the <u>chemical formula</u> but it may not be the molecular formula of a molecular compound. For instance the organic compound ethane has an empirical formula of CH_3 but its molecular formula, as it exists in nature is C_2H_6.

The distance between two bonded atoms at their minimum potential energy is called the <u>bond length</u>. Energy is required for this bond to be broken and form individual atoms. This energy is called <u>bond energy</u>.

Molecular compounds like ionic compounds follow the octet rule which states that compounds tend to form so that each atom has an octet of electrons in the highest level. Smaller elements need only two if they are near helium on the periodic table.

<u>Lewis Dot Diagrams</u>

Valence electrons can be shown by means of a Lewis dot diagram. In this type of diagram only the valence electrons are represented by dots around the element's symbol name. The dots are placed at the 4 compass points around the symbol name. There can be no more than 2 dots at each point with a maximum of 8 dots. For example these are the Lewis dots diagrams of boron, carbon, sodium, and sulfur.

```
         .                   .                               . .
Boron  · B ·       Carbon  · C ·     Sodium  Na ·     Sulfur · S ·
                             .                               . .
```

The octet rule used on ionic bonding is also used in covalent bonding. For example the molecular formula for ammonia is NH_3. Each H wants 1 electron to have the stable number of 2 and the nitrogen wants 3 electrons to attain the stable octet.

When drawing Lewis structures the central atom is usually the least electronegative (except for hydrogen which is never central).

Ammonia can be represented by the Lewis dot as:

```
                                . .                 . .
   H ·,   H ·,   H ·,   +     · N ·  →     H  :  N  :  H
                                 .                  . .
                                                    H
```

Dashes can also be substituted for dots, in which a dash indicates the sharing of two electrons.

```
    . .
H — N — H
    |
    H
```

<u>Multiple Bonds</u>

Covalent bonds can share more than one pair of electrons. The previous examples showed only <u>single bonds</u>.

<u>Double bonds</u> contain 2 pairs of shared electrons and <u>triple bonds</u> contain 3 pairs of shared electrons

<u>Double Bond Example   Formaldehyde---</u>

<u>Double Bond Example   Formaldehyde---H_2CO</u>

```
                .      .                          . .
H ·  H ·      · C ·  · O  :  →     H   : C   :: O   :
                .      . .           . .
                                      H
```

Using dashes

$$\mathrm{H - \overset{\text{}}{C} = \ddot{O}\ :}$$
$$\quad\ |$$
$$\quad\ \mathrm{H}$$

<u>**Triple Bond Example -- Ethyne--C_2H_2**</u>

H · · Ċ · · Ċ · · H → H : C ::: C : H

or

$$\mathrm{H - C \equiv C - H}$$

Triple bonds are usually the upper limit of shared electrons.

<u>**Exceptions to the Octet Rule**</u>

There are exceptions to the octet rule that will not be covered in this course. Read about these for your own information.

<u>**Resonance Structures**</u>

Resonance structures are used when molecules or ions cannot be correctly represented by a single Lewis structure. A resonance structure has a double-headed arrow placed between the different Lewis structures indicating either or.

Example:

$$\mathrm{\ddot{\underset{..}{O}} = \ddot{O} - \ddot{\underset{..}{O}}: \quad \leftrightarrow \quad :\ddot{\underset{..}{O}} - \ddot{O} = \ddot{\underset{..}{O}}}$$

Answer the following questions.

1. **How many pairs of electrons are shared in the following types of covalent bonds?**
   a. **a single bond**

   b. **a double bond**

   c. **a triple bond**

2. **Draw Lewis structures for the following.**
   a. **IBr**

   b. **CH_3Br**

   c. **C_2HCl**

   d. **$SiCl_4$**

3. What determines bond length?

4. In covalent bonding, what is meant by an unshared or lone pair of electrons?

5. When drawing Lewis structures, which atom is usually the central atom?

Chapter 6 Section 3 pgs. 190-194.
Students should read this section on their own. Most of the material has already been covered in chapter 7 and the following notes, but still must be read for background information.
Vocabulary: Define the following.
1. ionic compound--

2. formula unit--

3. lattice energy--

4. polyatomic ion--

Characteristics of Ionic Bonding
In ionic crystals, ions reduce their potential by combining in an orderly arrangement called a crystal lattice. The strength of these bonds is compared by using the amounts of energy released when separated ions in a gas come together. The energy released when one mole of an ionic crystalline compound is formed is called the lattice energy. A negative value indicates energy is released. The higher the lattice energy the stronger the bond.
In ionic compounds the metallic element loses electron(s), while the nonmetallic element gains electron(s).
********The atoms of an ionic compound are bonded more strongly than in a covalent compound. This stronger bond accounts for the following properties of ionic compounds as compared to covalent compounds.**

Ionic Compound Properties

1) High melting points.
2) Brittle.
3) Easily dissolve in water.
4) Good conductors of heat and good conductors of electricity when dissolved in water or melted.

Answer the following questions.

1. **What types of elements, metals or nonmetals form ionic compounds?**

2. **Compound A has lower melting and boiling points than compound B. compound B vapor- rizes faster than compound B. Which compound is molecular?**

3. **Why are ionic compounds brittle?**

4. **Why do ionic crystals conduct electric current in the liquid phase or when dissolved in water but not in the solid phase?**

5. **Explain why each of the following are not likely to form an ionic bond.**
   a. **chlorine and bromine**

   b. **potassium and helium**

   c. **sodium and lithium**

<u>Chapter 6 Section 5 Molecular Geometry pgs. 197-207.</u>
<u>Objectives:</u>
1. **Explain VSEPR theory.**
2. **Predict the shape of molecules or polyatomic ions using VSEPR theory.**
3. **Explain how the shapes of molecules are accounted for by hybridization theory.**
4. **Describe dipole-dipole forces, hydrogen bonding, induced dipoles, and London dispersion forces and their effects on properties such as boiling and melting points.**
5. **Explain what determines molecular polarity.**

<u>Vocabulary</u>: Define the following.
1. **VSEPR theory--**

2. **hybridization--**

3. **hybrid orbitals--**

4. **dipole--**

5. hydrogen bonding--

6. London dispersion forces--

<u>VSEPR Theory</u>

VSEPR theory states that repulsion between the sets of valence level electrons surrounding an atom cause these electrons to be as far apart as possible. The repulsion of these electrons explains the 3-dimensional <u>shape</u> of molecules. The shape of the molecule is determined by the number of shared and unshared pairs of electrons around the central atom. Shared pairs of electrons are farther away from each other.
VSEPR stands for valence shell, electron pair repulsion.

```
                                   ..        ..
The Lewis structure for BeF2 is  : F : Be:  F :
                                   ..        ..
```

The distance between the shared pairs will be the greatest if these pairs are on opposite sides of the beryllium atom, or 180º apart. This makes the molecular shape linear.

```
                                              ..            ..
The Lewis structure for boron trifluoride is : F            F :
                                              ..  \      /  ..
                                                     B
                                                     |
                                                   : F :
                                                     ..
```

The three bonds are maximized if the molecule has angles of 120º just as in an equilateral triangle.
There are many other different shapes that can be assumed as shown on page 200 of your text.

<u>Hybridization</u>

Hybridization occurs when two or more orbitals of equal energy combine to form a combination or hybrid orbital. This mixing of orbitals of similar energies on the same atom is illustrated by the carbon atom in methane, CH_4.

The orbital diagram of C is shown as ↑↓ ↓↑ ↑ ↑ ___
$1s^2$ $2s^2$ 2ps

The $2s^2$ and the two 2ps can combine to forma hybrid orbital called $2sp^3$. These four sp^3 orbitals all have the same energy, which is greater than that of the 2s but less than that of the 2p orbitals. This hybrid orbital explains why carbon can form 4 equal tetrahedral bonds with 4 hydrogen atoms in CH_4.

In summary, the two theories of VSEPR and hybridization explain bond angles (VSEPR) and the description of orbitals of valence electrons (hybridization)

<u>Intermolecular Forces</u>

Intramolecular forces describe the <u>strong</u> bonds of in ionic, covalent, or metal compounds. These bonds are <u>within</u> the compound formed. Intermolecular forces in contrast, describe the relatively <u>weak</u> bonds that develop <u>between</u> adjacent molecules.

Both boiling and melting points are much higher for ionic and metals compounds than covalent compounds.

Molecular Polarity and Dipole-Dipole Forces

The strongest intermolecular force occurs between polar molecules. Polar molecules are also referred to as dipoles, meaning two different ends of slightly positive or negative charge. A dipole will attract another dipole. This dipole to dipole attraction accounts for the relatively high boiling points of dipoles as compared to weak bonds between nonpolar molecules. Diagrams of a dipole are drawn by an arrow with the head pointing toward the more negative pole with the arrow's tail crossed over the more positive pole as in hydrogen chloride.

This dipole-dipole attraction can be more than one atom to atom bond as in ammonia, NH_3. In this instance all three hydrogen atoms are to one side of the nitrogen creating an additive effect and making the ammonia highly polar.
In some molecules such as carbon tetrachloride, CCl_4 the chlorine are equally distributed around the carbon atom cancelling any dipole effect and making it nonpolar.
Polar molecules can induce a dipole in a nonpolar molecule by temporarily attracting its electrons, making a normally nonpolar molecule polar when in closed proximity to a dipole molecule. This induced dipole effect explains why the normally nonpolar oxygen molecule O ─ O is soluble in water.

Hydrogen Bonding

Another very important type of intermolecular bonding involves hydrogen when it forms a compound with strongly electronegative elements such as fluorine, oxygen, and nitrogen.
In these compounds the lone electron is pulled so strongly towards the highly electronegative atom that a very strong dipole is formed. The hydrogen end is attracted to an unshared pair of electrons of the electronegative atom in the adjacent molecule. This is the main reason that water has such a relatively high boiling point for a molecule of such low molecular weight.

London Dispersion Forces

London dispersion forces are temporary dipoles created as electrons move about an atom. At some point the electrons may be more to one side of the atom inducing a temporary dipole in an adjacent atom even if that atom is strongly nonpolar such as in noble gases. In fact this is the only known intermolecular force found in noble gases.
London forces increase with increasing atomic mass and can be seen in the boiling points of many substances, such as hydrogen (H_2) at ─ 253°C, oxygen (O_2) at ─183°C and chlorine (Cl_2), at ─ 34°C.
Answer the following questions.

1. **What type of intermolecular force contributes to the high boiling point of water.**

2. **What is an intermolecular force?**

3. **Which is a type of dipole-dipole force, a hydrogen bond or London force?**

4. What is a temporary dipole?

5. Which is stronger, a hydrogen bond or London force?

6. Explain why oxygen, nitrogen, and fluorine are elements in molecules that form strong hydrogen bonds.

7. How is the strength of London dispersion forces related to the number of electsrons?

8. a. Which is nonpolar: CF_4 or CH_2F_2?

b. Which substance likely has a higher boiling point? Explain your answer.

9. Why do electron pairs repel each other as stated in VSEPR?