

Chapter 7

XI. Chemical Bonding

- A. Substances bond in order to become more stable; that is to acquire an octet or noble gas configuration.
- B. Elements use their valence electrons by losing, gaining or sharing these electrons.
- C. The rest of the atom not including the valence electrons is the kernel.
- D. Two main types of bonds
 - 1. Ionic - transfer of electrons; the bond is the attraction between the cation and anion.
 - 2. Covalent - sharing valence electrons
 - a. Three types
 - 1) pure covalent - totally equal sharing
 - 2) non polar covalent - almost equal sharing
 - 3) polar covalent - unequal sharing
- E. Ionic Bonding a closer look
 - 1. Both atoms usually attain a noble gas configuration (octet)
 - 2. Electronegativity difference is 1.8 and higher
 - 3. Orbital notation
 - a. Orbital notations of each atom is drawn
 - b. Valence electrons that are to be transferred are circled and an arrow is drawn to open space orbital in receiving atom.
 - c. The cation ion and anion orbital notations are then drawn including charges
 - 4. Lewis dot structures
 - a. Proper dot nots are drawn then follow b & c above.
 - b. Be certain that the cation's dot not only represents the valence electrons. If these have been transferred then there are no dots .
 - 5. Energy involved in ionic bonding
 - a. Ionic bonding consists of three different steps
 - (1) removal of an electron using ionization energy (endothermic)
(a) $\text{Na} + 496\text{kJ} \rightarrow \text{Na}^+ + \text{e}^-$
 - (2) addition of the electron considering electron affinity (usually exothermic)
(a) $\text{Cl} + \text{e}^- \rightarrow \text{Cl}^- + 348 \text{kJ}$
 - (3) the addition of the cation and anion considering bond energy - the amount of energy given off when a bond is formed or the amount of energy needed to break a bond- (in this case exothermic)
(a) $\text{Na}^+ + \text{Cl}^- \rightarrow \text{NaCl} + 792 \text{kJ}$
 - (4) these three reactions can be added together to get an overall reaction combining like terms
(a) $\text{Na} + \text{Cl} \rightarrow \text{NaCl} + 644 \text{kJ}$ (exothermic)
 - b. Ionic bonding is always exothermic
 - 6. Properties of ionic compounds
 - a. Usually crystalline solids at room temp.
 - b. High melting points
 - c. Conduct electric current when molten

d. If soluble in water they will also conduct electricity - called electrolytes

F. Metallic Bonding

1. Vacant orbitals in valence shell overlap allowing outer electrons to flow freely throughout the entire metal.
2. Electrons are said to be delocalized, not belonging to any one atom, like a commune.
3. These free floaters form a *sea of electrons* around the metal.
4. The chemical bonding that results from the attraction between metal atoms and the surrounding sea of electrons is called metallic bonding.
5. Metals are malleable and ductile because of this. Since metallic bonding is the same in all directions throughout the solid, layers are able to slide past one another without resisting or breaking bonds.
6. Metals also conduct electricity because the electrons are free to flow from one end to another thus carrying charge.
7. The strength of metallic bonds can be determined by vaporizing the metals and using the heat of vaporization (found on a table pg. 182). The more heat required to vaporize the metal the stronger the metallic bond.

G. Covalent Bonding

1. Valence electrons are being shared by the bonded atoms
2. These two shared electrons (with opposite spin) effectively fill an orbital in each atom. They make up a covalent electron pair.
3. There are three types of covalent bonds which depends on the difference in electronegativity
 - a. Pure Covalent - electronegativity diff of 0.0 completely equal sharing
(1) exists mainly between diatomic elements (H₂, N₂, O₂, ALL halogens)
 - b. Non-polar Covalent - electronegativity diff of 0.1 - 0.4 almost equal sharing
 - c. Polar Covalent - electronegativity diff of 0.5 - 1.8 unequal sharing
4. Orbital Notation
 - a. Electrons that are being shared are boxed.
 - b. Each box corresponds to one covalent electron pair
5. Dot Structures
 - a. Shared electrons are placed in between the atoms
 - b. Each pair of electrons may be replaced with a dash to form dash notation
6. Resonance
 - a. Not every molecule can be represented by a dot structure that adheres to the octet rule. Sometimes the actual structure exists somewhere in between a dot structure we can represent on paper. The structure may actually flip flop back and forth between two structures, resonating like sound waves making a perfect pitch.--
 - b. An example is SO₂
 - c. When this structure is drawn there are two possibilities. One with a coordinate covalent bond from sulfur on the top and one with the coordinate covalent bond on the bottom with a double bond opposite. O=S-O or O-S=O. This is also a bent molecule. Actually through

experimentation it has been determined that both bonds from sulfur are the same thus our representation is not accurate. The true structure must lie somewhere between these , a resonance structure.

- d. The concept of resonance is an attempt to make up for the deficiencies in writing the electron dot structures of certain molecules.
- e. Another example is the aromatic ring structure of benzene, C_6H_6 .

7. Energy Involved

- a. Again there are essentially four steps
 - (1) Break the bonds of one molecule using bond energy
 - (a) $1 \text{ mole } H_2 + 435 \text{ kJ} \implies 2 \text{ moles } H \text{ atoms}$
 - (2) Break the bonds of another molecule using its bond energy
 - (a) $1 \text{ mole } Cl_2 + 243 \text{ kJ} \implies 2 \text{ moles } Cl \text{ atoms}$
 - (3) Combine the free atoms which releases new bond energy
 - (a) $2 \text{ moles } H + 2 \text{ moles } Cl \implies 2 \text{ moles } HCl + 2(432 \text{ kJ})$
 - (4) Add all three reactions together subtracting and combining like terms
 - (a) $1 \text{ mole } H_2 + 1 \text{ mole } Cl_2 \implies 2 \text{ moles } HCl + 186 \text{ kJ}$

8. Hybridization

- a. The combining of two or more orbitals of nearly the same energy into new orbitals of equal energy.
- b. Common occurrence in methane
- c. The 2s and the 2p orbitals of the carbon combine and the electrons evenly distribute themselves so that there are four free electrons to bond. This new orbital is called $2sp^3$ orbital
- d. Hybrids can be sp^3 , sp^2 , sp^1 .

9. Coordinate Covalent Bonds

- a. Atoms will use two of their own electrons to form a bond with another to achieve octet.
- b. Carbon monoxide is an example. Oxygen uses two of its unshared electrons to form the third bond between C and O .

10. Properties of molecular substances

- a. Covalent bonds
- b. Usually nonmetals
- c. Exist as covalent network (giant molecules- atoms are covalently bonded in crystal lattice structure) or covalent molecular crystals (molecules are held together by van der Waals forces)
- d. Relatively low melting points
- e. Solubility in water depends on polarity
- f. Usually poor to nonconducting.

H. Molecular Polarity and Shapes

- 1. Non-polar molecules
 - a. All bonds are pure or non-polar covalent
 - b. Equal distribution of valence electrons and charge
- 2. Polar molecules
 - a. Molecules that have slightly positive and slightly negative ends
 - b. Called dipoles (two poles)

- c. Polarity depends on molecular geometry
 - (1) if molecule is symmetrical then it is usually non-polar
 - (2) if molecule is not symmetrical then it is probably polar
 - (3) to determine shape we use the VSEPR theory
- d. VSEPR theory
 - (1) Valence Shell Electron Pair Repulsion Theory
 - (2) basically, unshared pairs of electrons will push and other electron pairs as far from themselves as possible
 - (3) molecules must have at least 3 atoms for polarity to depend on geometry
 - (4) SHAPES

(a)	linear	3 atoms	180° bond angle
(b)	bent	3 atoms	105° bond angle
(c)	triangular planar	4 atoms	120° bond angle
(d)	tetrahedral	5 atoms	109.5°
(e)	triangular pyramidal	4 atoms	107° bond angle
(f)	triangular bipyramidal	6 atoms	90° & 120°
(g)	octahedral	7 atoms	all 90°
 - (5) of these - the bent and triangular pyramidal are the most likely to be polar

I. Intermolecular Attractive Forces

- 1. Johannes van der Waals
 - a. Tried to explain why something condenses. As the molecules slow down in a gas being cooled something must pull these molecules together to change phase. Intermolecular forces. The higher the temperature at which the gas condenses the stronger the intermolecular forces. There are two basic types
 - b. London dispersion forces (Dispersion interaction or momentary dipole forces) {Named for Fritz London}
 - (1) This type exists between all molecules. The larger the molecule (number of electrons) the greater the attraction. This is the weakest van der Waals force. This can also be viewed as a momentary dipole. Since the electrons are constantly moving about the molecule, it is logical to assume at a particular time most of the electrons will be clustered about one end of a molecule thus creating a momentary unequal distribution of charge or a dipole. This creates attractive forces between two molecules, momentarily.
 - c. Dipole - Dipole attractive force.
 - (1) This only exists between dipoles (polar molecules with two distinct regions of opposite charge). This is a stronger force than dispersion. The slightly positive end of a molecule is attracted to the slightly negative end of another.
 - (2) A special type of Dipole - Dipole attraction is a hydrogen bond. This is the strongest type of dipole - dipole attraction

since the hydrogen only has one valence electron and its proton is easily exposed. This exposed proton attracts other negatively charged regions of other dipoles. This has 5% of the strength of a covalent bond. It is not a true chemical bond but is important in determining properties.

J. Hybrid Orbitals

1. sp - In BeF_2 we look at Be orbital notation and notice it has 2 electrons paired in the 2s. To accommodate bonds it combines the 2s with one of its 2p to form a 2sp. This will correspond to a linear shape.
2. sp^2 - In BF_3 we look at B and notice it has 2 electrons paired in the 2s and 1 electron in the 2p, thus it must combine the 2s with 2 2p's to form a $2sp^2$. This corresponds to a triangular (trigonal) planar.
3. sp^3 - This occurs when an s orbital combines with three p orbitals. This hybrid is also present when the central atom has 4 pairs of electrons with no multiple bonds. Thus CH_4 has sp^3 orbitals as well as H_2O and NH_3 . This corresponds to the tetrahedral, bent (angular), and pyramidal shapes.