

## 1. Chemical Bonding

- a. Elements use their valence electrons by losing, gaining or sharing these electrons.
- b. The rest of the atom not including the valence electrons is the kernel.
- c. Two main types of bonds
  - i. Ionic - transfer of electrons; the bond is the electrostatic attraction between the cation and anion.
  - ii. Covalent - sharing valence electrons
    1. Three types
      - a. pure covalent - totally equal sharing
      - b. non polar covalent - almost equal sharing
      - c. polar covalent - unequal sharing

## 2. Ionic Bonding a closer look

- a. Both atoms usually attain a noble gas configuration (octet)
- b. Electronegativity difference is 1.9 and higher
- c. Orbital notation
  - i. Orbital notations of each atom is drawn
  - ii. Valence electrons that are to be transferred are circled and an arrow is drawn to open space orbital in receiving atom.
  - iii. The cation ion and anion orbital notations are then drawn including charges
- d. Lewis dot structures
  - i. Proper dot notations are drawn then follow 2 & 3 above.

- ii. Be certain that the cation's dot notation only represents the valence electrons. If these have been transferred then there are **no** dots.

e. Lattice Energy of Ionic Compounds

- i. Lattice Energy - energy needed to separate one mole of a solid ionic compound into gaseous ions.
- ii. Can be determined using Coulomb's law
  - 1. See page 361 Chang 9<sup>th</sup> Ed Chemistry.
- iii. We will use the Born-Haber Cycle
  - 1. Relates lattice energy to ionization energy, electron affinity, and other atomic and molecular properties. It is based on Hess's law.
  - 2. Basically, the overall change in energy equals the sum of all the individual steps to make the crystal. This may involve many steps.
    - a. Convert solid to gas; dissociation of molecules into single atoms; removing and adding electrons; combining ions to form crystal
- iv. The greater the lattice energy the more stable the ionic compound.
- v. Lattice energy is always positive because the separation of the ionic crystal from solid to gas is, by Coulomb's Law, endothermic.
- vi. There is more attraction between ions with larger charges than between ions of lesser charge.
- vii. Lattice Energy can help explain formulas of ionic compounds. Basically, if the lattice energy is greater than the energy required to make the formula then the formula is stable.

- viii. Usually crystalline solids at room temp.
- ix. High melting points
- x. Conduct electric current when molten
- xi. If soluble in water they will also conduct electricity – called electrolytes

### 3. Covalent Bonding – the friendly bond

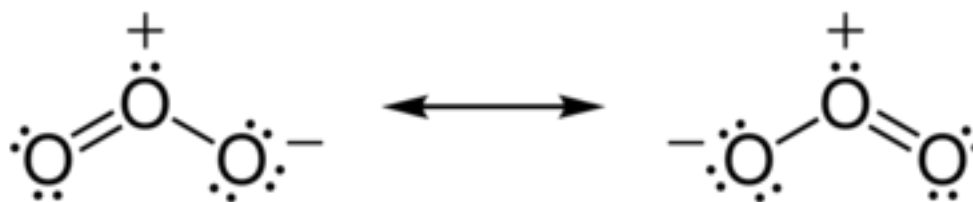
- a. Valence electrons are shared by two atoms
- b. The two shared electrons (with opposite spins) effectively fill an orbital in each atom. This makes a covalent electron pair.
- c. Atoms bond until a stable configuration is established. The most preferable is octet. The Octet Rule (ominous music plays).
- d. One bond between two atoms is a single bond (sigma,  $\sigma$ ), two = double bond (sigma,  $\sigma$  and pi,  $\pi$ ) and three = triple bond ( $\sigma$  and 2  $\pi$ 's)
- e. Single bonds are longer than double which are longer than triple. Single bonds are weaker than double bonds which are weaker than triple bonds.
- f. There are three types of covalent bonds which depend on the difference in electronegativity.
  - i. **Pure covalent** – electronegativity difference of 0.0 – completely equal sharing of electrons. Exist mainly between two of the same atoms i.e. the diatomic elements.
  - ii. **Non-Polar Covalent** – electronegativity difference of 0.1 – 0.4 – almost equal sharing electrons.
  - iii. **Polar Covalent** - electronegativity difference of 0.5 – 1.8 – unequal sharing of electrons.
- g. Special covalent bond – **coordinate covalent bond** – formed when shared electrons come from the same atom to achieve octet rule (ominous music plays).
- h. Properties – Usually gas, liquid or low melting solid since intermolecular forces are weaker than ionic compounds.
- i. Orbital notation
  - i. Electrons that are shared are boxed.
  - ii. Each box represents one covalent bond.
- j. Electron Dot Notation
  - i. Shared electrons are placed in between atoms sharing those electrons. These can be replaced with a dash.
  - ii. Left over valence electron pairs are placed around the corresponding atom, evenly spaced.

### 4. Lewis Dot Structures

- a. The best directions I can give as to how to write Lewis Dot Structures is to use your common sense and knowledge about electron stability.
- b. To start determine which atom is going to bond the most number of times. This will be the central atom, the atom most everything else is connected to. If carbon is in the formula it will probably be the central atom.
- c. Place the other atoms around the central atom and bond them to the central atom so that the octet rule is satisfied. If the octet rule cannot be satisfied, try another arrangement. If that still doesn't work count the total valence electrons and start to fill them in to satisfy the octet rule for as many atoms as possible. This may mean using coordinate covalent bonds.
  - i. When counting valence electrons present remember to add or subtract electrons for anions(add) or cations(subtract).
- d. Official Lewis Rules – starting after (b.) above
  - i. Sum the valence electrons
  - ii. Draw single bonds to the central atom
  - iii. Complete the octets of the atoms bonded to the central atom
  - iv. Place any leftover electrons on the central atom
  - v. Does the central atom have an octet? If yes then done. If no, then add a multiple bond (double first) to the central atom until an octet is achieved.
  - vi. Try nitrogen trifluoride, nitric acid, and the carbonate ion. (See Chang pg. 373 – 374)
  - vii. Now try carbon disulfide, formic acid (HCOOH), and the nitrite ion.

#### 5. Formal Charge and Lewis Structure

- a. An atom's formal charge is the electrical charge difference between the valence electrons in an isolated atom and the number of electrons assigned to that atom in a Lewis structure.
- b. In ozone,  $O_3$ , the O on the left with double bond has 6 valence electrons, as all O's have and 6 assigned electrons which gives a difference of zero. The middle O has 6 valence electrons and

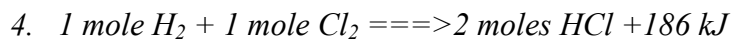
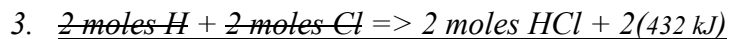
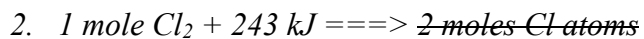
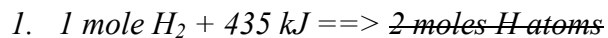


only 5 assigned giving a difference of +1, the O on the right has 6 valence electrons and 7 assigned which gives a difference of -1. (The diagram is showing the resonance structure of ozone)

#### 6. Resonance

- a. Using two or more Lewis structures to represent a particular molecule that cannot be represented by one structure. One Lewis Structure doesn't fully explain the actual structure and in reality neither do the resonance structures. The actual structure lies somewhere between.
  - b. The reason we have resonance structures is that through experimentation the bond lengths in molecules don't match the proposed Lewis structure. Looking at the Ozone structure above with the double bond and single bond. The double bond should be shorter than the single bond however, from experimentation it has been found that the bonds in ozone are shorter than single bonds but longer than double bonds.
  - c. The resonance structures are sometimes replaced with one structure replacing the floating double bond with a single bond and dashed other bond.
7. Exceptions to the Octet Rule – We LOVE exceptions!!!
- a. The Incomplete Octet – Sometimes the atoms don't want eight electrons.
  - b. Odd – Electron Molecules – molecules contain an odd number of electrons, nitric oxide (NO) and nitrogen dioxide (NO<sub>2</sub>). These molecules are called radicals and are HIGHLY reactive. Can wreak havoc in body, causes aging (Baylor Theory)
  - c. The Expanded Octet - Some atoms just like to break the rules and bond as many times as they have electrons floating around. These atoms use inner electrons to bond and form hybrid orbitals (combination of orbitals to make one that fits our theories)
8. Bond Enthalpy – the enthalpy change required to break a particular bond in 1 mole of gaseous molecules. (Enthalpy and energy are just about the same thing.)
- a. It is possible to predict the enthalpy of a reaction using the average bond enthalpies. Remember in order to break a bond energy is required, endothermic, and when bonds form energy is released, exothermic. Also remember that endothermic energy is added, is on the reactant side of a reaction, is listed as positive. Exothermic energy is released, is on the product side of a reaction, is listed as negative.
  - b. One way to look at this – the change in enthalpy of a reaction is equal to the total energy needed to break bonds minus the total energy released. Using this formula there is no need for assigning exothermic energy as negative.
  - c. Another way – incorporates the idea of endo = + energy and exo = - energy. The change in enthalpy of a reaction equals the product enthalpy minus the reactant enthalpy.
  - d. But Wait....There's more!! – we can add the actual equations like algebraic equations to get the overall reaction.

- i. Breaking bond(s) of one reactant  
Breaking bond(s) of other reactant  
Combine both broken reactants to form product  
Overall reaction



This an exothermic reaction since the overall energy is on the product side and would be reported as a negative quantity.  $\Delta H^\circ = -186 \text{ kJ}$