Chem E-1a
Friday Review Notes
Chapter 9: Chemical Bonding I

Another Periodic Trend: Electronegativity

- Electronegativity indicates an atom's attraction for a shared pair of electrons.
- Electronegativity increases as you go up and right across the periodic table. Fluorine has the highest electronegativity.
- A covalent bond between two of the same atoms is a *nonpolar covalent* bond.
- A covalent bond between two atoms with different electronegativities is a *polar covalent* bond because the electrons are shared unequally between the two atoms.
- An ionic bond exists between atoms that have extremely different electronegativities. This will occur between metals and nonmetals.

Drawing Lewis Structures:

- See the handout "Lewis Structures: Rules" from lecture. Follow this procedure for drawing Lewis structures until the process becomes second nature. This handout is also included at the end of this packet.
- Some additional things to consider when drawing Lewis structures:
  - In Oxyacids, the oxygens are connected to the central atom, and hydrogens are connected to the oxygens. e.g.: HNO_3, H_2SO_4, H_2CO_3, H_3PO_4.
  - **Reduced Octets** will typically occur when **Be, B, or Al are the central atoms.**
    - Be will often have only 4 electrons around it in a Lewis structure.
    - B and Al will often have only 6 electrons around them in a Lewis structure.
  - **Reduced Octets** will also occur when you have an **odd total number of electrons.**
    - If you have an odd total number of electrons, it will be impossible for all atoms to follow the octet rule.
    - One atom will need to end up with just 7 electrons.
  - **Expanded Octets** should only be used when necessary to connect the atoms.
    - Do not expand octets simply to reduce formal charges.
    - Expanded octets should only occur on central atoms.
    - Atoms with expanded octets should only have lone pairs and single bonds. They should not have any double or triple bonds.
    - Period 1 and Period 2 elements CANNOT form expanded octets.

Bond Enthalpy Calculations:

- Bond Enthalpies are the energy required to BREAK 1 mole of a given bond in the gas phase.
  - For example, $D_{O=O}$ is the energy necessary to break the O=O double bond in the gas phase: $O_2\,(g) \rightarrow 2\,O\,(g) \quad \Delta H° = D_{O=O} = 499$ kJ/mol
  - If a bond is formed, energy is released, and ΔH° for the reaction is equal to the *negative* of the bond enthalpy: $2\,O\,(g) \rightarrow O_2\,(g) \quad \Delta H° = -D_{O=O} = -499$ kJ/mol
- Breaking bonds is endothermic. Forming bonds is exothermic.
- Bond enthalpies only work in the gas phase.
- Bond enthalpies are approximations—the actual enthalpy of a bond can vary depending on the molecule.
- Always draw Lewis Structures for every molecule so you can see what types of bonds you have.
- For an entirely gas phase reaction (i.e. all reactants and products in the gas phase):

  $$\boldsymbol{\Delta H_{rxn} = \Sigma(\text{Bonds Broken}) - \Sigma(\text{Bonds Formed})}$$
- If the reaction includes solids or liquids, you must convert them to gases before doing bond enthalpy calculations.

Born-Haber Cycles:

•Born-Haber Cycles give us a way to relate the lattice energy of an ionic solid to the enthalpy of formation of that solid and other values (such as electron affinities and ionization energies).

•**Lattice Energy** is defined as the energy required to separate 1 mole of a solid ionic substance into individual gas-phase ions:

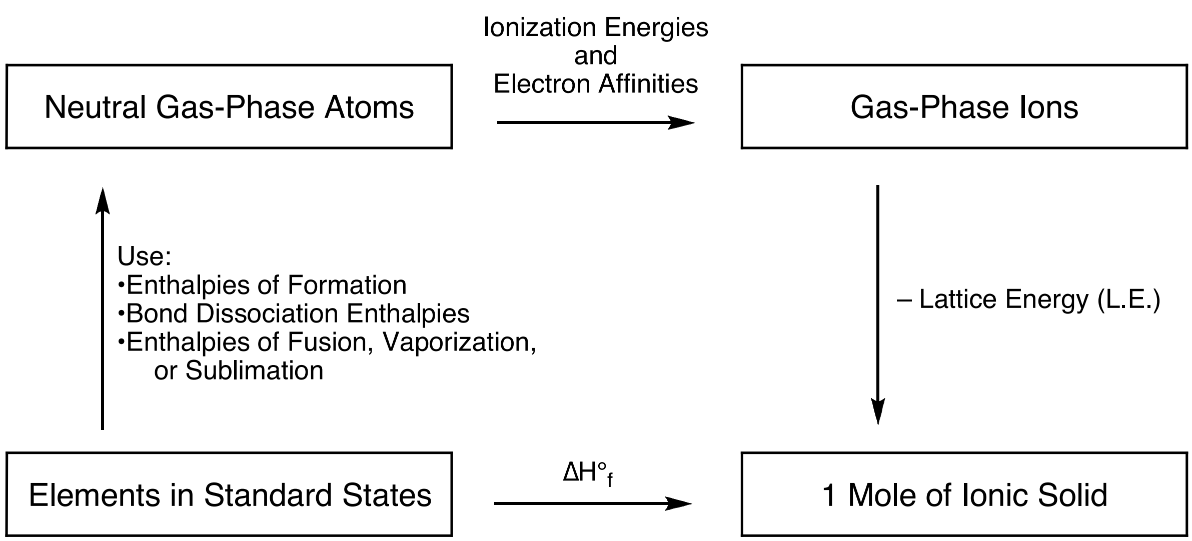
$$NaCl\ (s) \rightarrow Na^{+}\ (g) + Cl^{-}\ (g) \qquad \Delta H° = \text{Lattice Energy of NaCl}$$

•Forming an ionic solid from ions in the gas phase (as we do in the Born-Haber Cycle) is the reverse of this process, so it is the *negative* of the Lattice Energy:

$$Na^{+}\ (g) + Cl^{-}\ (g) \rightarrow NaCl\ (s) \qquad \Delta H° = -\text{ L.E. of NaCl}$$

•Born-Haber Cycles will follow the following *general* template. However, variations to this standard template may be necessary depending on the information you are given and the particular ionic solid. (For example, an ionic solid containing a polyatomic ion will require formation of a gaseous polyatomic ion, and alterations to this standard template may be necessary.)

Born-Haber Cycle Template:



Solving Numerically:

$\Delta H°_f$ = Sum of all terms around the other path

Lewis Structures: Rules

These rules are helpful guidelines. Practice using them until they become second-nature.

1. **Count** the total number of valence electrons that must appear in the structure.
   - •Be sure to account for the extra or fewer electrons in a polyatomic ion!

2. **Arrange** the atoms in the correct arrangement for the molecule or ion.
   - •A few guidelines:
     - •Molecules are often written in the order that the atoms are connected.
     - •The least electronegative atom is usually the central atom.
     - •Hydrogen and fluorine are always terminal atoms.
   - •In oxyacids (H_2SO_4, HNO_3, etc.), the oxygen atoms are bonded to the central atom, and the hydrogen atoms are bonded to the oxygens. (Note you already know the oxyacids H_2SO_4, $HClO_4$, HNO_3, H_3PO_4 and CH_3COOH, all other oxyacid we will identify for you)

3. **Connect** each adjacent atom with a single bond. (Use lines for bonds.)

4. **Complete** the octets of terminal atoms by filling in lone pairs of electrons. (Use dots!)
   - •Remember that hydrogen should only have two electrons.

5. **Add leftover electrons** to the central atom (even if that gives it more than an octet).

6. **Try multiple bonds** if the central atom lacks an octet.
   - •As a guide for where to place multiple bonds in order to minimize formal charges, the total number of bonds to any **uncharged** atom is usually given by its position on the periodic table. This is a guideline, not a hard-and-fast rule!

Group #:	1A	2A	3A	4A	5A	6A	7A
Example:	H	Be	B	C	N	O	F
# of bonds:	**1**	**2**	**3**	**4**	**3**	**2**	**1**

7. **Check** formal charges on each atom, and write in any non-zero formal charges.
   - •Formal charge = (# of valence e^- in free atom) – (# of dots and lines around atom)
   - •The sum of the formal charges on all atoms will equal zero for a neutral molecule, and will equal the charge on the ion for a polyatomic ion.
   - •Lewis structures with octets should be drawn to minimize formal charges
   - •In general, negative formal charge should go on the most electronegative atoms.

8. **Check** for resonance, and indicate it if necessary. Note when drawing resonance structures do not change how the atoms are connected to each other or move the atoms to different locations.

9. **Check** your Lewis structure.
   - •Have you drawn the correct number of valence electrons?
   - •Do all atoms have octets? (Except for examples of reduced or expanded octets) Note: only expand the octet if that is the only possible option to draw the structure.
   - •Have you minimized formal charges? Note: structures with an octet of electrons and formal charge are better than those drawn minimizing formal charge but with an expanded octet.