*AP Chemistry*

Chapter 8 Outline

* 1. Chemical Bonds, Lewis Structures, and the Octet Rule
     1. Ionic Bond: electrostatic forces between oppositely charged ions
        1. typically when metals interact with nonmetals
     2. Covalent Bond: results from sharing of electrons between atoms
        1. Typically when nonmetals interact with other nonmetals
     3. [Metallic bond](http://www.drkstreet.com/resources/metallic-bonding-animation.swf): when many metal nuclei share delocalized electrons (Ch. 23)
     4. Lewis electron dot diagrams (aka Lewis symbols) consist of the chemical symbol of the element plus a dot for each valence electron
     5. Octet Rule: Atoms tend to gain, lose or share electrons until they are surrounded by eight valence electrons.
        1. There are many exceptions to the octet rule!
  2. Ionic Bonding
     1. Can be envisioned as an electron transfer
        1. Typically form between element with low ionization energy (which would easily lose an electron) and an element with high electron affinity (which would readily gain an electron)
     2. Energetics of Ionic Bond Formation (Lattice Energy)
        1. Ionization of metal is endothermic, while nonmetal gaining an electron is typically exothermic; need a third energy term to explain why overall bond formation is exothermic
           1. Ions of unlike charge are attracted. This causes most ionic materials to be hard and brittle, with high melting points.
           2. [Lattice energy](http://wps.prenhall.com/wps/media/objects/4678/4790313/images/table6_03.gif) = the energy required to completely separate a mole of a solid ionic compound into its gaseous ions; can use the [Born-Haber cycle](http://wps.prenhall.com/wps/media/objects/4678/4790313/images/aabjuyra.jpg) to calculate

Lattice energy can’t be determined directly by experiment

Can be considered an application of Hess’ Law

* + - 1. Magnitude of lattice energy depends on 2 factors, governed by Coulomb’s Law   where Eel = the potential energy of two interacting charged particles
         1. Ionic charge: As the charge on the ions increases, the lattice energy increases.
         2. Ionic radius: As the radii on the ions decreases, the lattice energy increases. (i.e, as the distance between the nuclei decreases)
    1. Electron Configurations of Ions of the Representative Elements
       1. Group 1 elements form +1 ions because loss of additional electrons from an inner shell requires a very large amount of energy;
       2. Group 2 elements form +2 ions; group 13 elements form +3 ions
       3. Group 14 metals usually form +2 ions
          1. Group 15 elements typically form -3 ions
          2. Group 16 elements typically form -2 ions
          3. Group 17 elements typically form -1 ions
    2. Transition Metal ions
       1. Transition metals generally do not form ions that have a noble-gas configuration.
       2. Transition metals form the valence shell s electrons first, than as many d electrons as are required to reach the charge of the ion.
    3. Properties of Ionic Compounds
       1. Brittle, with high melting points
       2. Crystalline structure—rigid, well defined 3D arrangement
  1. Covalent Bonding
     1. Properties of Covalent Materials
        1. Low melting points, may vaporize readily; may be pliable in their solid forms
     2. [Covalent bond](http://ch301.cm.utexas.edu/svg/s-sigma-formation.svg) = chemical bond formed by sharing a pair of electrons
        1. Both [repulsive & attractive forces](http://people.uwplatt.edu/~sundin/114/image/l1432a.gif) are present
        2. Attractions between the nuclei & the electrons cause the electron density to concentrate between the nuclei leads to a net attractive force
        3. If the nuclei are too close, the repulsions will be stronger and the atoms will separate
     3. Covalent bonds can be represented as Lewis structures, with lines to represent pairs of shared electrons (i.e., bonds)
        1. Single bond = sharing of a pair of electrons
        2. Double bond = when 2 pairs of electrons are shared
        3. Triple bond = sharing of 3 pairs of electrons
     4. Bond length = the distance between the nuclei of the atoms involved in a bond
  2. Bond Polarity and Electronegativity
     1. Bond polarity describes the sharing of electrons between atoms.
        1. Nonpolar covalent = electrons are shared equally between two atoms
        2. Polar covalent = one of the atoms exerts a greater attraction for the bonding electrons than the other
     2. Electronegativity = the ability of an atom in a molecule to attract electrons to itself
        1. Period trend: generally increases from left to right
        2. Group tend: typically decreases as you go down a column
     3. We can use the difference in electronegativity values to predict bond polarity.
        1. Nonpolar bond = very small difference in EN
        2. Polar bond = small to moderate differences in EN result in + and - regions of bond
        3. Ionic bond = large difference in EN results in electron transfer
     4. Polar bonds have a “dipole moment” that is non-zero
        1. the positive and negative regions of the bond do not coincide. (Non-polar bonds have zero dipole moment!)
        2. [Two ways to indicate:](http://classconnection.s3.amazonaws.com/311/flashcards/1406311/jpg/polar_-_water1334534718519.jpg) draw +, - or an arrow pointing towards the negative end
        3. Dipole moment  = Qr
           1. As magnitude of separated charge increases, dipole moment increases.
           2. As the distance separating the charge increases, dipole moment increases.
     5. Remember to name molecules properly!
        1. Ionic compounds: use Stock notation NO PREFIXES!
        2. Covalent molecules: use prefixes
        3. Note that compounds containing metals with high oxidation numbers (usually above 3+) often have properties more similar to molecular compounds than ionic compounds!
  3. Drawing Lewis Structures ESSENTIAL SKILL! PRACTICE!
     1. Simple rules:
        1. Sum all the valence electrons. Add one for each negative charge; remove one for each positive charge.
        2. Write a skeleton structure. Place the least electronegative element (but never hydrogen) in the central position.
        3. Use dashes to connect bonded atoms. Each dash represents a shared pair.
        4. Complete the octets around all the atoms bonded to the central atom.
        5. Place any leftover electrons on the central atom.
           1. N, O, C, and F always follow the octet rule.
           2. P and S frequently have expanded octets, while compounds containing B and Be may be electron deficient around the central atom.
        6. If there are not enough electrons to give the central atom an octet, try multiple bonds.
     2. Formal Charge = the charge the atom would have if all the atoms in the molecule had the same electronegativity
        1. FC = # valence electrons - # lone pair electrons - # bonds *Be* *able to calculate!*
        2. We use formal charge to evaluate alternative Lewis structures.
           1. In general, choose the Lewis structure in which the atoms have formal charges closest to zero.
           2. Choose the Lewis structure which places any negative formal charges on the most electronegative atoms.
        3. Formal charges DO NOT represent real charges on atoms!
  4. Resonance Structures
     1. Sometimes a single Lewis structure is inadequate for describing a molecule or polyatomic ion.
        1. Sometimes it is possible to draw several correct, equivalent structures for the same molecule, in which only the placement of the electrons differs.
        2. These structures are called [resonance structures](http://www.chem.ucla.edu/harding/tutorials/resonance/resonance_str03.GIF), and they are separated by double-headed arrows.
     2. The “true” structure is somewhere in the middle—a blend—of the equivalent structures.
     3. In some cases, the possible Lewis structures for a species may not be equivalent to each other; instead, one or more may represent a more stable arrangement than the others.
     4. Resonance in benzene, C6H6
        1. Aromatic molecule
        2. Not alternating single/double bonds—all C-C bond lengths are identical with a bond order of 1.5
  5. Exceptions to the Octet Rule
     1. Some molecules have an odd number of electrons.
        1. Examples: NO, ClO2, NO2, O2-
        2. Lewis structures do not handle this well!
     2. Electron-Deficient Molecules
        1. Less than an octet around the central atom
        2. Often associated with compounds of B, Be
        3. Example: BF3
     3. Expanded Octets
        1. More than an octet around the central atom
        2. Often found with elements of P, S
           1. Elements must be in period 3 or lower in order to have expanded octets
        3. Examples: PCl5, SF4, AsF6-, ICl4-
        4. Old viewpoint:The central atoms seem to be using their empty d orbitals to accommodate additional electrons.
           1. Current thinking: computational data don’t seem to support this idea of involving d orbitals. Other Since there is no consensus, this is no longer part of the AP Chem curriculum.
        5. The larger the central atom, the larger the number of atoms that can surround it. Small surrounding atoms (F, Cl, O) are frequently found in expanded octet situations.
     4. Sometimes Lewis structures are written with an expanded octet even though structures can be written with an octet.
        1. Rule of thumb: Go with an expanded octet if it reduces formal charges, but in general try to satisfy the octet rule if possible.
  6. Strengths of Covalent Bonds
     1. Bond enthalpy = the enthalpy change for the breaking of a particular bond in one mole of a gaseous substance
        1. ΔH(bond type) information is tabulated.
        2. Often, average bond enthalpies are used.
           1. Aka bond strength or bond energy
        3. Bond enthalpy is always a positive quantity. (Energy is always required to break chemical bonds.)
        4. The greater the bond enthalpy, the stronger the bond.
        5. A molecule with strong chemical bonds is less reactive than one with weak bonds
     2. [Bond enthalpies](http://chemwiki.ucdavis.edu/Physical_Chemistry/Thermodynamics/State_Functions/Enthalpy/Bond_Enthalpies) can be used as a means to calculate enthalpies of reactions
        1. An application of Hess’ Law!
        2. 
     3. As the number of bonds between two atoms increases, the bond becomes shorter.
        1. Single bonds are longer than double bonds, etc.