<u>**Basics of Chemical Bonding**</u>

AP Chemistry Lecture Outline

Properties of substances are largely dependent on the bonds holding the material together.

Basics of Bonding

A <u>chemical bond</u> occurs when atoms or ions are **strongly** attached to each other.

<u>Ionic bonds</u> involve the transfer of e^- and the subsequent electrostatic attractions.

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<u>Covalent bonds</u> involve the sharing of e^- between two atoms.

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<u>metallic bonds</u>: each metal atom is bonded to several neighboring atoms
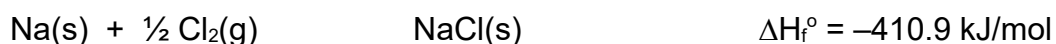
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<u>Lewis symbols</u> show ONLY the valence e^- (i.e., the ones involved in bonding).

<u>octet rule</u>:

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Ionic Bonding   "Salts" are brittle solids with high melting points.

Na(s) + ½ Cl₂(g)   NaCl(s) ΔH_f° = –410.9 kJ/mol

<u>lattice energy</u>:   the energy required to separate 1 mole of solid ionic compound into gaseous ions

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NaCl(s)   Na⁺(g) + Cl⁻(g) ΔH_{latt} = +788 kJ/mol

In general, ionically bonded substances have...

Because lattice energies are electrostatic in nature, two variables are involved in how big they are: 1.

2.

EX.   Put the following in order of increasing lattice energy: LiBr, FeN, CdO

Now these: MgS, MgCl₂, MgO

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With transition-metal ions, the first e⁻ lost come from the subshell with the largest value of n.

　　　e.g.,　　Ni = [Ar] $4s^2\ 3d^8$　　　　　　Ni^{2+} =　　　　　　　Ni^{3+} =

Recall that <u>polyatomic ions</u> are groups of atoms that stay together and have a net charge.

　　-- e.g.,　NO_3^-　　　CH_3COO^-

　　-- their atoms are held to each other by…

　　Cations are *smaller* than the neutral atoms from which they are derived.

　　　e.g.,　　Li　$1s^2\ 2s^1$

　　　　　　　Li^+　$1s^2$
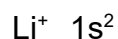
　　　Fe　　　　　　　　　Fe^{2+}　　　　　　　Fe^{3+}

　　Anions are *larger* than the neutral atoms from which they are derived

　　　　　　Cl　　　　　　　Cl^-

　　An <u>isoelectronic series</u> is a list of species having an identical electron configuration.

　　　e.g.,

EX.　　Which has the largest radius?　　　　Rb^+　　　　Sr^{2+}　　　　Y^{3+}

Covalent Bonding
　　-- atoms share e-
　　-- covalent (molecular) compounds
　　　tend to be solids with low melting
　　　points, or liquids or gases

　　　-- one shared pair of e⁻ (i.e., 2 e⁻) = a single covalent bond
　　　-- two shared pairs of e⁻ (i.e., 4 e⁻) = a double covalent bond
　　　-- three shared pairs of e⁻ (i.e., 6 e⁻) = a triple covalent bond

<u>bond polarity</u>:
　　<u>nonpolar covalent bond</u>:
　　<u>polar covalent bond</u>:
<u>electronegativity (EN)</u>: the ability of an atom in a molecule to attract e⁻ to itself

-- A bonded atom ^{w}/a large EN has a great ability to attract e$^-$.

-- A bonded atom ^{w}/a small EN does not attract e$^-$ very well.

-- EN values have been tabulated.

The ΔEN between bonded atoms approximates the type of bond between them.

As ΔEN increases, bond polarity...

Dipole Moments   Polar covalent molecules have a partial (–) and a partial (+) charge and are said to have a <u>dipole moment</u>.

H–F   H–F

.   big ΔEN = _____ polarity = _____ dipole moment

Polar molecules tend to align themselves with each other and with ions.

H–F   H–F

** Nomenclature tip: *For binary compounds, the <u>less electronegative</u> element comes first.*

-- Compounds of metals ^{w}/high ox. #'s (e.g., 4+ or higher) tend to be molecular rather than ionic.

e.g.,

Lewis Structures

1. Sum the valence e$^-$ for all atoms. If the species is an ion, add one e$^-$ for every (–); subtract one e$^-$ for every (+).
2. Write the element symbols and connect the symbols with single bonds.
3. Complete octets for the atoms on the exterior of the structure, but NOT for H.
4. Count up the valence e$^-$ on your L.S. and compare that to the # from Step 1.
   -- If your LS doesn't have enough e$^-$, place as many e$^-$ as needed on central atom.
   -- If LS has too many e$^-$ OR if central atom doesn't have an octet, use multiple bonds.

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EX.   Draw Lewis structures for the following species.

PCl₃   HCN PO_4^{3-}

H₂   CH₃CH₂OH   CO₂
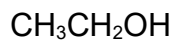
As the # of bonds between two atoms increases, the distance between the atoms…

formal charge: the charge a bonded atom would have if all the atoms had the same
                 electronegativity; to find it, you must first draw the Lewis structure

        When several Lewis structures are possible, the most stable is the one in which:
                (1) the atoms have the smallest formal charges, and
                (2) the (−) charges reside on the most electronegative atoms

EX.   Find the formal charge on each atom in the following species

        SO_3^{2-}

        NCS^-

resonance structures: the two or more Lewis structures that are equally correct for a species
        e.g., For NO_3^-…

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Resonance structures are a blending of two or more Lewis structures.

    <u>Aromatic compounds</u> are based on resonance

    structures of the <u>benzene</u> molecule.

Exceptions to the Octet Rule

There are a few cases (other than for H) in which the octet rule is violated. These are:

        1. particles with an odd number of valence e^-

          -- e.g.,

        2. atoms with less than an octet

          -- e.g.,

        3. atoms with more than an octet

          -- this occurs when an atom gains an expanded valence shell

          -- other e.g.,

              Expanded valence shells occur only for atoms in periods > 3.

                --

                -- large central atom =

                -- small exterior atoms =

The strengths of a molecule's covalent bonds are related to the molecule's stability and the amount of energy required to break the bonds.

<u>bond enthalpy</u>: the ΔH req'd to break 1 mol of a particular bond in a gaseous substance

        -- big ΔH =

        -- e.g.,

<u>atomization</u>: the process of breaking a molecule into its individual atoms

      -- ΔH for a given bond (e.g., the C–H bond) varies little between compounds.

         e.g., C–H bonds in CH_4 vs. those in $CH_3CH_2CH_3$ have about the same ΔH

      -- Typical values of bond enthalpies for specific bonds have been tabulated.

      -- To find bond enthalpy for atomization, add up bond enthalpies for each bond broken.

EX.   Calculate the bond enthalpy for the atomization of dichloromethane.

You can approximate reaction enthalpy using Hess's law and tabulated bond enthalpies.

EX.   Approximate the reaction enthalpy for the combustion of propane.

<u>bond length</u>: the center-to-center distance between two bonded atoms

      -- fairly constant for a given bond (e.g., the C–H bond), no matter the compound

         e.g., C–H bonds in CH_4 are about the same length as those in $CH_3CH_2CH_3$

      -- Average bond lengths have been tabulated for many bonds.

      -- As the number of bonds between two atoms increases,

          bond length...   and bond enthalpy...

        e.g.,   C–C   C=C