**Basics of Chemical Bonding** Name: \_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_

*AP Chemistry Lecture Outline*

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

**Basics of Bonding**

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

 Ionic bonds involve the transfer of e– and the subsequent electrostatic attractions.

 --

 Covalent bonds involve the sharing of e– between two atoms.

 --

 metallic bonds: each metal atom is bonded to several neighboring atoms

 --

 Lewis symbols show ONLY the valence e– (i.e., the ones involved in bonding).

 octet rule:

 --

**Ionic Bonding**  “Salts” are brittle solids with high melting points.

 Na(s) + ½ Cl2(g) NaCl(s) Hfo = –410.9 kJ/mol

 lattice energy: the energy required to separate 1 mole of solid ionic compound into

 gaseous ions

 --

 NaCl(s) Na+(g) + Cl–(g) Hlatt = +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

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

 e.g., Ni = [Ar] 4s2 3d8 Ni2+ = Ni3+ =

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

 -- e.g., NO3– CH3COO–

 -- their atoms are held to each other by…

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

 e.g., Li 1s2 2s1

--

--

 Li+ 1s2

 Fe Fe2+ Fe3+

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

 --

 Cl Cl–

 An isoelectronic series is a list of species having an identical electron configuration.

 e.g.,

EX. Which has the largest radius? Rb+ Sr2+ Y3+

**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

bond polarity:

 nonpolar covalent bond:

 polar covalent bond:

electronegativity (EN): 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 EN between bonded atoms approximates the type of bond between them.

   

As EN increases, bond polarity...

**Dipole Moments** Polar covalent molecules have a partial (–) and a partial (+) charge

and are said to have a dipole moment.

 H–F H–F

 . big EN = \_\_\_\_\_\_ 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 less electronegative 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.

EX. Draw Lewis structures for the following species.

 PCl3 HCN PO43–

 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

formal

charge

# of v.e– in

= the isolated –

atom

# of e– assigned

to the atom in 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

 SO32–

 NCS–

resonance structures: the two or more Lewis structures that are equally correct for a species

 e.g., For NO3–...

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

 Aromatic compounds are based on resonance

structures of the benzene 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.

bond enthalpy: the H req’d to break 1 mol of a particular bond in a gaseous substance

 --

 -- big H =

 -- e.g.,

atomization: the process of breaking a molecule into its individual atoms

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

 e.g., C–H bonds in CH4 vs. those in CH3CH2CH3 have about the same H

 -- 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.

bond length: 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 CH4 are about the same length as those in CH3CH2CH3

 -- 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 