

AP Chemistry
Chapter 8 Outline

- a) Chemical Bonds, Lewis Structures, and the Octet Rule
- i) Ionic Bond: electrostatic forces between oppositely charged ions
 - (1) typically when metals interact with nonmetals
 - ii) Covalent Bond: results from sharing of electrons between atoms
 - (1) Typically when nonmetals interact with other nonmetals
 - iii) [Metallic bond](#): when many metal nuclei share delocalized electrons (Ch. 23)
 - iv) Lewis electron dot diagrams (aka Lewis symbols) consist of the chemical symbol of the element plus a dot for each valence electron
 - v) 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!
- b) Ionic Bonding
- i) 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)
 - ii) 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
 - (a) Ions of unlike charge are attracted. This causes most ionic materials to be hard and brittle, with high melting points.
 - (b) [Lattice energy](#) = the energy required to completely separate a mole of a solid ionic compound into its gaseous ions; can use the [Born-Haber cycle](#) to calculate
 - (i) Lattice energy can't be determined directly by experiment
 - (ii) Can be considered an application of Hess' Law
 - (2) Magnitude of lattice energy depends on 2 factors, governed by Coulomb's Law
$$E_{el} \propto \frac{Q_1 Q_2}{d}$$
 where E_{el} = the potential energy of two interacting charged particles
 - (a) Ionic charge: As the charge on the ions increases, the lattice energy increases.
 - (b) Ionic radius: As the radii on the ions decreases, the lattice energy increases.
 - (i.e., as the distance between the nuclei decreases)
 - iii) 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
 - (a) Group 15 elements typically form -3 ions
 - (b) Group 16 elements typically form -2 ions
 - (c) Group 17 elements typically form -1 ions

- iv) 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.
 - v) Properties of Ionic Compounds
 - (1) Brittle, with high melting points
 - (2) Crystalline structure—rigid, well defined 3D arrangement
- c) Covalent Bonding
- i) Properties of Covalent Materials
 - (1) Low melting points, may vaporize readily; may be pliable in their solid forms
 - ii) Covalent bond = chemical bond formed by sharing a pair of electrons
 - (1) Both repulsive & attractive forces 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
 - iii) 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
 - iv) Bond length = the distance between the nuclei of the atoms involved in a bond
- d) Bond Polarity and Electronegativity
- i) 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
 - ii) 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
 - iii) 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 δ^+ and δ^- regions of bond
 - (3) Ionic bond = large difference in EN results in electron transfer
 - iv) 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: draw δ^+ , δ^- or an arrow pointing towards the negative end
 - (3) Dipole moment $\mu = Qr$
 - (a) As magnitude of separated charge increases, dipole moment increases.
 - (b) As the distance separating the charge increases, dipole moment increases.
 - v) 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!
- e) Drawing Lewis Structures ESSENTIAL SKILL! PRACTICE!
- i) 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.
 - (a) N, O, C, and F always follow the octet rule.
 - (b) 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.
 - ii) Formal Charge = the charge the atom would have if all the atoms in the molecule had the same electronegativity
 - (1) $FC = \# \text{ valence electrons} - \# \text{ lone pair electrons} - \# \text{ bonds}$ *Be able to calculate!*
 - (2) We use formal charge to evaluate alternative Lewis structures.
 - (a) In general, choose the Lewis structure in which the atoms have formal charges closest to zero.
 - (b) 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!
- f) Resonance Structures
- i) 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, and they are separated by double-headed arrows.
 - ii) The “true” structure is somewhere in the middle—a blend—of the equivalent structures.
 - iii) 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.
 - iv) Resonance in benzene, C_6H_6
 - (1) Aromatic molecule
 - (2) Not alternating single/double bonds—all C-C bond lengths are identical with a bond order of 1.5

g) Exceptions to the Octet Rule

- i) Some molecules have an odd number of electrons.
 - (1) Examples: NO, ClO₂, NO₂, O₂⁻
 - (2) Lewis structures do not handle this well!
- ii) Electron-Deficient Molecules
 - (1) Less than an octet around the central atom
 - (2) Often associated with compounds of B, Be
 - (3) Example: BF₃
- iii) Expanded Octets
 - (1) More than an octet around the central atom
 - (2) Often found with elements of P, S
 - (a) 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.
 - (a) 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.
- iv) 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.

h) Strengths of Covalent Bonds

- i) Bond enthalpy = the enthalpy change for the breaking of a particular bond in one mole of a gaseous substance
 - (1) $\Delta H_{(\text{bond type})}$ information is tabulated.
 - (2) Often, average bond enthalpies are used.
 - (a) 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
- ii) [Bond enthalpies](#) can be used as a means to calculate enthalpies of reactions
 - (1) An application of Hess' Law!
 - (2) $\Delta H_{\text{rxn}} = \Sigma(\text{enthalpies of bonds broken}) - \Sigma(\text{enthalpies of bonds formed})$
- iii) As the number of bonds between two atoms increases, the bond becomes shorter.
 - (1) Single bonds are longer than double bonds, etc.