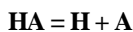


MOLECULAR CELL PHYSIOLOGY Special Lecture - General, Inorganic, & Organic Chemistry

- I. What makes chemistry important?**
1. Metabolism
 2. Growth and development
 3. Environmental interactions
- II. General & inorganic chemistry**
- A. Units of measure**
1. Volume - liter
 2. Mass - gram
 3. Length - meter
- B. Chemical units of measure**
1. Molecular weight - mass of a substance that contains one mole (6.022×10^{23}) of atoms (or molecules)
 2. Mole - amount of a substance that contains 6.022×10^{23} atoms (or molecules)
 3. Molarity - number of moles in one liter of solution
- C. Metric system**
1. kilo - 10^3
 2. centi - 10^{-2}
 3. milli - 10^{-3}
 4. micro - 10^{-6}
- D. Composition of elements**
1. Atoms - smallest portion of an element
 - a) Protons
 - b) Electrons
 - c) Neutrons
- E. Function of atoms**
1. Electron excitation (energy)
 2. Chemical bonds
 - a) Covalent - share electrons
 - b) Ionic - charges attract
 - c) Hydrogen - weak attraction of H & O
- F. Important chemical phenomena**
1. Acid / base pH
 - a) Equilibrium
 - b) Availability & solubility of ions
 - c) Buffering capacity
 2. Oxidation / reduction
 - a) Donate / accept electrons
The reaction, acetaldehyde + $2H^+$ + 2 electrons \rightleftharpoons ethanol, represents a reduction of acetaldehyde
- III. Organic chemistry - carbon chemistry, chemistry of life**
- A. Alkanes, alkenes, and alkynes
 - B. Alcohols, ethers, and amines
 - C. Aldehydes, ketones, and carboxylic acids
 - D. Cyclic compounds and aromatics
- IV. Henderson-Hasselbach equation**



$$K_{eq} = \frac{[H][A]}{[HA]}$$

$$pH = -\log[H]$$

$$[H] = K_{eq} \frac{[HA]}{[A]}$$

$$pH = -\log K_{eq} + \log \frac{[A]}{[HA]}$$

(assume for acid, $K_{eq} = K_a$)

$$-\log[H] = -\log K_{eq} - \log \frac{[HA]}{[A]}$$

$$pH = pK_{eq} + \log \frac{[A]}{[HA]}$$

$$pH = pK_a + \log \frac{[A]}{[HA]}$$

MOLECULAR CELL PHYSIOLOGY - General, Inorganic, & Organic Chemistry (continued)

- V. Thermodynamics
- A. First Law: conservation of energy - energy cannot be created or destroyed (heat is work and work is heat)
 - B. Second Law: Tendency towards entropy - energy tends to follow a path toward disorder (heat cannot on itself pass from a cooler body to a hotter body)
 - C. Gibbs Free Energy (G)
 - 1. Free energy change $\Delta G = \Delta G(\text{prod}) - \Delta G(\text{react})$
 - a) If ΔG is negative, the reaction will occur spontaneously
 - b) If ΔG is positive, the reverse reaction will tend to occur
 - c) If ΔG is zero, the reaction is at equilibrium
 - D. The change in Gibbs Free Energy can also be calculated as $\Delta G = \Delta H - T\Delta S$ where
 - 1. Enthalpy (H) is the bond energy of the system
 - 2. Temperature (T) is in degrees Kelvin
 - 3. Entropy (S) is a measure of randomness

ΔG	ΔH	$T\Delta S$
(-)	$\Delta H < 0$ (exothermic)	$T(\Delta S > 0)$
?	$\Delta H > 0$ (endothermic)	$T(\Delta S > 0)$
? (many biological systems)	$\Delta H < 0$ (exothermic)	$T(\Delta S < 0)$
(+)	$\Delta H > 0$ (endothermic)	$T(\Delta S < 0)$

Note that $\Delta G = \Delta G^\circ + RT\ln Q$ where ΔG° is the value of the change in free energy under conditions of 298 K, 1 atm pressure, pH of 7, and initial concentrations of 1 M for all reactants and products; R is the gas constant of 1.987 cal/(degree)(mol); T is the temperature in Kelvin; and Q is the initial ratio of products to reactants. ΔG° can also be calculated as $\Delta G^\circ = -2.3RT \log K_{eq}$



ATP energy is used to produce heat, electrolytic potential, transportation, cell movement, synthesis of phospholipids, DNA and RNA synthesis, etc.