



Unit 1

Atomic Structure & Properties

- The **atom**: protons (+), neutrons (0), electrons (-)
- A **mole** relates the mass of an element to the number of particles there are. One mol = **Avogadro's number**
- **Molar mass** is the # of grams in a mole.
- **Isotopes** have a diff # of neutrons → avg. atomic mass
- **Mass spectroscopy**: identifies the compos. of a sample
- **Empirical** (simplest whole # ratio) vs. **Molecular** formulas
- **Election config**: fill electrons in low energy levels first, electrons in same orbital must have opposite spin, unpaired electrons must fill unoccupied orbitals first
- **Photoelectron spectroscopy (PES)**: gives us e- config
- **Periodic trends**: atomic radius greatest on lower left, electronegativity and i.e. greatest on upper right



Unit 2

Molecular & Ionic Bonding

- **Intramolecular forces**: between two atoms in a molecule
 - **Ionic bond**: transfer of e- from metal to nonmetal
 - **Polar covalent**: shared e-; unequal distribution charge
 - **Nonpolar cov**: shared e-; equal distribution of charge
 - Single bonds = longest length, smallest bond energy
- **Ionic solid**: cations & anions arranged in **crystal lattice**
- **Metals**: sea of delocalized electrons, can form **alloys**
 - Interstitial vs. substitutional alloys
- **Lewis diagrams**: showing valence e- and bonds
- **Resonance**: when a molecule can be represented with several Lewis diagrams (because of its bonding structure)
- **Formal charge**: charge assigned to an atom in a molecule
- **VSEPR**: molecular geometry, hybridization, bond angles



Unit 3

Intermolecular Forces & Properties

- **Intermolecular forces**: between two molecules b/c of charge diff
 - **LDFs**: weakest, exist between all molecules, temp. dipole
 - **Dipole-dipole**: permanent dipoles, b/w 2 polar molecules
 - **H Bonding**: strongest in pure samples, H -- F, O, or N
 - **Ion-dipole**: in mixture of ionic compound & polar molecules
- **Solids**: amorphous vs. crystalline (ionic, metallic, cov network, molecular) & their properties
- **Liquids**: surface tension, capillary action, viscosity
- **Gasses**: gas laws, ideal gas law ($PV=nRT$), law of partial press
- **KMT**: gas particles have no attraction, no volume, move in random motion, have elastic collisions, temp related to KE
- **Solutions**: like dissolves like, M, factors that affect solubility
- **Photoelectric effect** ($E=h\nu$, $c=\lambda\nu$) & **Beer's Law** ($A = \epsilon bc$)



Unit 4

Chemical Reactions

- **Physical** vs. **chemical** changes
- **Precipitation reactions**: one product is insoluble
- **Net ionic equations**: show what *really* happens without the expression of spectator ions
- Balancing equations: satisfy law of conservation of mass
- **Stoichiometry**: conversions to predict substance amount
- **Acid-base reactions**: involve the transfer of a proton from one molecule to another (acid → base)
 - **Titration**: determines unknown concentration and finds equivalence point of acid-base reactions
- **Redox reactions**: involve the transfer of electrons from a reducing to an oxidizing agent



Unit 5

Kinetics

- **Rate of reaction**: change of concentration over period of time
- **Rate law**: $R = k[A]^n[B]^m$ where n and m are reaction orders
 - Can only be determined experimentally!
 - The rate constant, k, is temperature specific
- **Integrated rate laws**: determined by [reactant] and time
- **Collision model**: reaction occurs when there is enough energy (**activation energy**) & reactants are in the correct orientation
- **Reaction mechanisms**: shows each step of a chem. Reaction
 - Includes **catalysts** (speeds reaction) and **intermediates**
 - Stoichiometric coefficients of the **rate-determining step** define the rate law of the overall chemical reaction and mechanism



Unit 6

Thermodynamics

- **Kinetic** vs. **potential energy** & open, closed, + isolated systems
- **Enthalpy of Reaction** (ΔH): heat absorbed/released by reaction.
 - **Endothermic** (+ ΔH) vs. **exothermic** (- ΔH)
- Temperature = average kinetic energy
- **Calorimetry**: experimental way to measure ΔH by measuring ΔT
- $q=mC\Delta T$, where C is **specific heat**: energy required to raise 1g 1°C
- Heating and cooling curves: heats of fusion and vaporization
- **Enthalpy of Formation**: The change in enthalpy of forming 1 mole of a compound ($\Delta H = \sum n\Delta H_f(\text{prod}) - \sum n\Delta H_f(\text{reactants})$)
- **Bond Enthalpy** = $\sum \text{energy bonds broken} - \sum \text{energy bonds formed}$
- **Hess's Law**: the total enthalpy of reaction is a sum of the enthalpies for each step.



Unit 7

Equilibrium

- **Equilibrium**: rate of forward reaction equals rate of reverse reaction & concentrations are constant
- Equilibrium can be measured with K_c/K_p , **equilibrium constants** (ratio of products:reactants at equilibrium)
- **Reaction quotient (Q)**: products: reactants at any point
 - Compare Q to K to determine what direction a reaction will shift
- **ICE Tables**: used to calculate equilibrium concentrations
- **Solubility product (Ksp)**: balance between dissolved ions and the solid compound in the solution
- **Common ion effect**: *decreases* the solubility of ionic co.
- **Le Chatelier's principle**: how the reaction responds to a stressor to return to equilibrium



Unit 8

Acids & Bases

- **Acid** (donates a proton) vs. **base** (accepts a proton)
 - Identify acids, conjugate acids, bases, conjugate bases
- $pH = -\log[H_3O^+]$ and $pOH = -\log[OH^-]$
- **Strong** acids and bases completely dissociate in water
- **Weak** acids and bases incompletely dissociate in water, so we need to use ICE tables to determine equilibria concentrations
- **Buffers**: mixture of an acid and its conjugate base that is somewhat resistant to changes in pH (like in your blood!)
 - **Henderson-Hasselbalch Equation** to find the pH
 - Buffer capacity: the more concentrated the acid, the stronger the buffer is at reducing pH
- **Titrations**: equivalence point ($nMaVa=mMbVb$)
- **Indicators**: compounds that change color based on pH



Unit 9

Applications of Thermodynamics

- **Entropy**: disorder, chaos (gases are more disordered than solids)
- **Absolute entropy (S°)**: # of possible states a molecule can take
 - $\Delta S^\circ = \sum S^\circ(\text{products}) - \sum S^\circ(\text{reactants})$
- Spontaneous reactions are thermodynamically favorable.
- **Gibbs Free Energy (G)** describes spontaneity ($\Delta G^\circ = \Delta H^\circ - T\Delta S^\circ$)
 - When ΔG is negative, the reaction is spontaneous.
- Relationship Between ΔG° , ΔG , and Q: $\Delta G = \Delta G^\circ + RT \ln(Q)$
- **Galvanic cells**: spontaneous; electrons travel from the anode → cathode through a salt bridge ($E_{\text{cell}} = E_{\text{cathode}} - E_{\text{anode}}$)
- **Electrolytic cells**: nonspontaneous, need an energy source to drive the reaction (like a battery)
- $\Delta G^\circ = -nFE^\circ_{\text{cell}}$, where F is Faraday's constant and E°_{cell} is standard cell potential & the Nerst equation for E_{cell}