AP Chemistry Cheat Sheet

[Complete Guide]

Unit 1: Atomic Structure & Properties

- **The Atom:** Consists of protons (+), neutrons (0), electrons (-).
- **Mole Concept:** Relates the mass of an element to the number of particles. One mole equals Avogadro's number (6.022×10^{23}) particles.
- **Molar Mass:** The mass of one mole of a substance in grams.
- **Isotopes:** Atoms of the same element with different numbers of neutrons, leading to different mass numbers.
- **Mass Spectroscopy:** Identifies the composition of a sample by measuring the mass-to-charge ratio of ions.
- **Empirical Formula:** The simplest whole number ratio of elements in a compound.
- **Electron Configuration:** Describes the arrangement of electrons in an atom's orbitals. Electrons fill orbitals from lowest to highest energy.
- **Photoelectron Spectroscopy (PES):** Measures the ionization energies of electrons to deduce electronic structure.
- **Periodic Trends:** Atomic radius, ionization energy, and electronegativity trends across periods and down groups.
- **Quantum Mechanical Model:** Describes electron distribution in atoms. Orbitals (s, p, d, f) with specific shapes and energy levels.
- **Heisenberg Uncertainty Principle:** Impossible to know both the position and momentum of an electron simultaneously.
- **Example Calculation:** Calculate the average atomic mass of an element given the isotopic masses and their abundances.

Additional Notes:

Unit 2: Molecular & Ionic Bonding

- **Intramolecular Forces:** Forces within a molecule (ionic, covalent, and metallic bonds).
  - **Ionic Bond:** Transfer of electrons from a metal to a nonmetal.
  - **Polar Covalent Bond:** Unequal sharing of electrons between atoms.
  - **Nonpolar Covalent Bond:** Equal sharing of electrons.
- **Ionic Solids:** Lattices of cations and anions held together by electrostatic forces.
- **Metallic Bonds:** Delocalized electrons shared among a lattice of metal atoms.
- **Lewis Diagrams:** Represent valence electrons and bonds.
- **Resonance Structures:** Different valid Lewis structures for the same molecule.
- **VSEPR Theory:** Predicts molecular geometry based on electron pair repulsion.
- **Bond Energy:** Energy required to break a bond in a molecule.
- **Example:** Calculate the formal charge of atoms in a molecule to predict the most stable Lewis structure.
- **Hybridization:** Mixing of atomic orbitals to form new hybrid orbitals (sp, sp^2, sp^3).

Additional Notes:

Unit 3: Intermolecular Forces & Properties

- **Intermolecular Forces:** Forces between molecules.
  - **London Dispersion Forces (LDFs):** Weakest, present in all molecules.
  - **Dipole-Dipole Interactions:** Between polar molecules.
  - **Hydrogen Bonding:** Strongest, occurs when H is bonded to F, O, or N.
  - **Ion-Dipole Interactions:** Between ions and polar molecules.
- **Solids:**
  - **Amorphous Solids:** No long-range order (e.g., glass).
  - **Crystalline Solids:** Ordered structures (ionic, metallic, covalent network, molecular).
- **Liquids:**
  - **Properties:** Surface tension, viscosity, capillary action.
  - **Laws:** Ideal gas law ($PV=nRT$), law of partial pressures.
- **Solutions:** Factors affecting solubility (like dissolves like, temperature, pressure).
- **Phase Diagrams:** Graphs showing conditions (temperature and pressure) at which distinct phases occur and coexist.
- **Example:** Explain how intermolecular forces affect boiling and melting points of substances.

Additional Notes:

Unit 4: Chemical Reactions

- **Physical vs. Chemical Changes:** Physical changes do not alter chemical composition, chemical changes do.
- **Precipitation Reactions:** Formation of an insoluble product from soluble reactants.
- **Net Ionic Equations:** Show only the species that undergo a change.
- **Balancing Equations:** Conservation of mass and charge.
- **Stoichiometry:** Calculations based on balanced chemical equations.
- **Acid-Base Reactions:** Transfer of protons between reactants.
  - Titration: Determines the concentration of an unknown acid/base.
- **Redox Reactions:** Transfer of electrons between reactants.
- **Types of Reactions:** Synthesis, decomposition, single replacement, double replacement, and combustion reactions.
- **Example Problem:** Balance a combustion reaction and calculate the amount of product formed.

Additional Notes:

Unit 5: Kinetics

- **Rate of Reaction:** Change in concentration of reactants/products over time.
- **Rate Laws:** Express reaction rate as a function of reactant concentrations.
  - **Rate Law Formula:** Rate = $k[A]^m[B]^n$
- **Reaction Mechanisms:** Step-by-step sequence of elementary reactions.
- **Collision Model:** Reactions occur when particles collide with sufficient energy and correct orientation.
- **Catalysts:** Lower the activation energy and speed up reactions without being consumed.
- **Factors Affecting Reaction Rate:** Concentration, temperature, surface area, and catalysts.
- **Arrhenius Equation:** $k = Ae^{-E_a/RT}$, relates temperature and rate constant.
- **Example Calculation:** Determine the activation energy given the rate constants at different temperatures.

Additional Notes:

Unit 6: Thermodynamics

- **Kinetic vs. Potential Energy:** Kinetic energy relates to motion, potential energy to position or composition.
- **Enthalpy (ΔH):** Heat absorbed or released in a reaction.
- **Endothermic:** Absorbs heat ($\Delta H > 0$).
- **Exothermic:** Releases heat ($\Delta H < 0$).
- **Calorimetry:** Measures heat changes in reactions.
- **Hess's Law:** Total enthalpy change is the sum of the enthalpy changes for individual steps.
- **Standard Enthalpy of Formation ($\Delta H^\circ f$):** Enthalpy change when one mole of a compound forms from its elements.
- **Second Law of Thermodynamics:** Entropy of an isolated system always increases.
- **Example:** Use Hess's Law to calculate the enthalpy change of a reaction from known enthalpy changes of related reactions.

Additional Notes:

Unit 7: Equilibrium

- **Dynamic Equilibrium:** Rate of the forward reaction equals the rate of the reverse reaction.
- **Equilibrium Constant (K):** Ratio of product concentrations to reactant concentrations at equilibrium.
- **Le Chatelier's Principle:** A system at equilibrium will adjust to relieve any applied stress.
- **ICE Tables:** Used to calculate changes in concentrations of reactants/products.
- **Le Chatelier's Principle Examples:** Predict the shift in equilibrium when concentration, temperature, or pressure changes.
- **Equilibrium Expressions:** Write Kc and Kp expressions for given reactions.

Additional Notes:

Unit 8: Acids & Bases

- **Acid-Base Definitions:** Bronsted-Lowry (proton transfer), Lewis (electron pair).
- **pH and pOH:** Measures of acidity/basicity.
  - pH = $-\log[H_3O^+]$
  - pOH = $-\log[OH^-]$
- **Titrations:** Determine the concentration of an acid/base by neutralization.
- **Buffers:** Solutions that resist changes in pH when small amounts of acid or base are added.
- **Strength of Acids and Bases:** Strong acids/bases dissociate completely, weak acids/bases do not.
- **pKa and pKb:** Measures of acid and base strength.
- **Example Calculation:** Calculate the pH of a weak acid solution using the acid dissociation constant (Ka).

Additional Notes:

Unit 9: Applications of Thermodynamics

- **Entropy (S):** Measure of disorder.
- **Gibbs Free Energy (ΔG):** Determines spontaneity of a reaction.
    - $\Delta G = \Delta H - T\Delta S$
    - Spontaneous when $\Delta G < 0$.
- **Electrochemistry:** Study of redox reactions and their applications.
    - **Galvanic Cells:** Convert chemical energy to electrical energy.
    - **Electrolytic Cells:** Use electrical energy to drive non-spontaneous reactions.
- **Standard Electrode Potentials:** Determine cell potential from standard reduction potentials.
- **Nernst Equation:** $E = E^\circ - \frac{RT}{nF}\ln Q$, relates cell potential to reaction quotient.
- **Example:** Calculate the Gibbs free energy change for a redox reaction from the cell potential.

Additional Notes: