CONCEPTS TO KNOW

GENERAL CHEMISTRY

ELECTRONIC STRUCTURE AND PERIODIC TABLE

A. Electronic Structure

1. Orbital structure of hydrogen atom, principal quantum number *n*, number of electrons per orbital
2. Ground state, excited states
3. Absorption and emission spectra
4. Quantum numbers *l*, *m*, *s*, and number of electrons per orbital
5. Common names and geometric shapes for orbitals *s*, *p*, *d*
6. Conventional notation for electronic structure
7. Bohr atom
8. Effective nuclear charge

B. The Periodic Table: Classification of Elements into Groups by Electronic Structure; Physical and Chemical Properties of Elements

1. Alkali metals
2. Alkaline earth metals
3. Halogens
4. Noble gases
5. Transition metals
6. Representative elements
7. Metals and nonmetals
8. Oxygen group

C. The Periodic Table: Variations of Chemical Properties with Group and Row

1. Electronic structure
    a. representative elements
    b. noble gases
    c. transition metals
2. Valence electrons
3. First and second ionization energies
    a. definition
    b. prediction from electronic structure for elements in different groups or rows
4. Electron affinity
    a. definition
    b. variations with group and row
5. Electronegativity
    a. definition
    b. comparative values for some representative elements and important groups
6. Electron shells and the sizes of atoms

BONDING

A. The Ionic Bond (Electrostatic Forces Between Ions)

1. Electrostatic energy $\propto q_1q_2/r$
2. Electrostatic energy $\propto$ lattice energy
3. Electrostatic force $\propto q_1q_2/r^2$

B. The Covalent Bond

1. Sigma and pi bonds
    a. hybrid orbitals (sp^3, sp^2, sp, and respective geometries)
    b. valence shell electron-pair repulsion (VSEPR) theory, predictions of shapes of molecules (e.g., NH_3, H_2O, CO_2)
2. Lewis electron dot formulas
    a. resonance structures
    b. formal charge
    c. Lewis acids and bases
3. Partial ionic character
    a. role of electronegativity in determining charge distribution
    b. dipole moment

PHASES AND PHASE EQUILIBRIA

A. Gas Phase

1. Absolute temperature, K
2. Pressure, simple mercury barometer
3. Molar volume at 0°C and 1 atm = 22.4 L/mol
4. Ideal gas
    a. definition
    b. ideal gas law ($PV = nRT$)
        i. Boyle's law
        ii. Charles's law
        iii. Avogadro's law
4. Kinetic theory of gases
5. Deviation of real-gas behavior from ideal gas law
    a. qualitative
    b. quantitative (van der Waals equation)
6. Partial pressure, mole fraction
7. Dalton's law relating partial pressure to composition

B. Intermolecular Forces

1. Hydrogen bonding
2. Dipole interactions
3. London dispersion forces

C. Phase Equilibria

1. Phase changes, phase diagrams
2. Freezing point, melting point, boiling point, condensation point
3. Molality
4. Colligative properties
    a. vapor pressure lowering (Raoult's law)
    b. boiling point elevation ($\Delta T_b = K_b m$)
    c. freezing point depression ($\Delta T_f = K_f m$)
    d. osmotic pressure
5. Colloids
6. Henry's law

STOICHIOMETRY

1. Molecular weight
2. Empirical formula versus molecular formula
3. Metric units commonly used in the context of chemistry
4. Description of composition by percent mass
5. Mole concept, Avogadro's number
6. Definition of density
7. Oxidation number
    a. common oxidizing and reducing agents
    b. disproportionation reactions
    c. redox titration
8. Description of reactions by chemical equations
    a. conventions for writing chemical equations
    b. balancing equations including redox equations
    c. limiting reactants
    d. theoretical yields

THERMODYNAMICS AND THERMOCHEMISTRY

A. Energy Changes in Chemical Reactions: Thermochemistry

1. Thermodynamic system, state function
2. Endothermic and exothermic reactions
    a. enthalpy *H*, standard heats of reaction and formation
    b. Hess's law of heat summation
3. Bond dissociation energy as related to heats of formation
4. Measurement of heat changes (calorimetry), heat capacity, specific heat capacity (specific heat capacity of water = 4.184 J/g·K)
5. Entropy as a measure of "disorder," relative entropy for gas, liquid, and crystal states
6. Free energy *G*
7. Spontaneous reactions and ΔG^o

B. Thermodynamics

1. Zeroth law (concept of temperature)
2. First law ($\Delta E = q + w$, conservation of energy)
3. Equivalence of mechanical, chemical, electrical, and thermal energy units
4. Second law (concept of entropy)
5. Temperature scales, conversions
6. Heat transfer (conduction, convection, radiation)
7. Heat of fusion, heat of vaporization
8. *PV* diagram (work done = area under or enclosed by curve)

RATE PROCESSES IN CHEMICAL REACTIONS: KINETICS AND EQUILIBRIUM

1. Reaction rates
2. Rate law, dependence of reaction rate on concentrations of reactants
    a. rate constant
    b. reaction order
3. Rate-determining step
4. Dependence of reaction rate on temperature
    a. activation energy
        i. activated complex or transition state
        ii. interpretation of energy profiles showing energies of reactants and products, activation energy, ΔH for the reaction
    b. Arrhenius equation
5. Kinetic control versus thermodynamic control of a reaction
6. Catalysts, enzyme catalysis
7. Equilibrium in reversible chemical reactions
    a. law of mass action
    b. the equilibrium constant
    c. application of Le Châtelier's principle
8. Relationship of the equilibrium constant and ΔG°

SOLUTION CHEMISTRY

A. Ions in Solution

1. Anion, cation (common names, formulas, and charges for familiar ions; e.g., NH_4^+, ammonium; PO_4^{3-}, phosphate; SO_4^{2-}, sulfate)
2. Hydration, the hydronium ion

B. Solubility

1. Units of concentration (e.g.,molarity)
2. Solubility product constant, the equilibrium expression
3. Common-ion effect, its use in laboratory separations
4. Complex ion formation
5. Complex ions and solubility
6. Solubility and pH

ACIDS AND BASES

A. Acid–Base Equilibria

1. Brønsted–Lowry definition of acids and bases
2. Ionization of water
    a. K_w, its approximate value ($K_w = [H_3O^+][OH^-] = 10^{-14}$ at 25°C)
    b. pH definition, pH of pure water
3. Conjugate acids and bases
4. Strong acids and bases (common examples; e.g., nitric, sulfuric)
5. Weak acids and bases (common examples; e.g., acetic, benzoic)
    a. dissociation of weak acids and bases with or without added salt
    b. hydrolysis of salts of weak acids or bases
    c. calculation of pH of solutions of weak acids or bases
6. Equilibrium constants K_a and K_b (pK_a and pK_b)
7. Buffers
    a. definition, concepts (common buffer systems)
    b. influence on titration curves

B. Titration

1. Indicators
2. Neutralization
3. Interpretation of titration curves

ELECTROCHEMISTRY

1. Electrolytic cell
    a. electrolysis
    b. anode, cathode
    c. electrolytes
    d. Faraday's law relating amount of elements deposited (or gas liberated) at an electrode to current
    e. electron flow, oxidation and reduction at the electrodes
2. Galvanic (voltaic) cell
    a. half-reactions
    b. reduction potentials, cell potential
    c. direction of electron flow