Chapter 1

- Scientific method
- Know the terms: macroscopic, microscopic, atomic or molecular level
- Significant Figures (Sig Figs)
- Scientific notation
- Problem solving by unit analysis method
- Metric system
  - Know all prefixes from femto to tera
  - Carry out metric-metric conversions
    - 1 cm³ = 1 mL and 1 dm³ = 1 L
- Length, mass, weight, volume
- Volume by displacement
- density = \( \frac{\text{mass}}{\text{volume}} \)
  - Be able to determine density, mass, or volume given the other two quantities
- Temperature: convert °F-to/from °C and K-to/from °C
- Percentage: ratio of parts per 100 parts
  - Calculate the percentage or use a given percentage to calculate part or whole
- Classification of matter: Classify substances as elements, compounds, and mixtures
- Physical states of matter
  - Determine physical state of substances (solids, liquids, gases) with descriptions of volume, shape, compressibility, attraction between particles, etc.
  - Know terms for transition from one physical state to another (e.g. freezing, condensing, vaporizing, sublimation, deposition, etc.)
- Given examples or molecular-level images, identify elements, compounds, and mixtures, as well as solids, liquids, and gases.
- Classify properties and changes as physical or chemical.

Chapter 2

- KNOW people and discoveries associated with:
  - John Dalton's Model
  - Rutherford's Alpha Scattering Experiment
  - Discovery of protons, neutrons, electrons
- Use law of conservation of mass to solve problems
- Subatomic particles
  - proton (p⁺): +1 charge, inside nucleus, ~1 amu
  - neutron (n): neutral, inside nucleus, ~1 amu
  - electron (e⁻): -1 charge, out of nucleus, 0 amu
- Mass number, atomic number, element symbol
  - Given 2, be able to determine the missing info
- Recognize the term isotope
- Atomic Notation (or Nuclear Symbol):
  \[ \text{atomic number} = Z \quad \text{mass number} = A \quad \text{E} = \text{element symbol} \]
- Determine the # of protons, neutrons, and electrons given any isotope of an atom or ion
- Periodic Table:
  - Know terms: groups, periods, Main-group (Representative) elements, transition metals
  - Group IA: alkali metals
  - Group IIA: alkaline earth metals
  - Group VIIA: halogens
  - Group VIIIA: noble gases
- periodic law: Elements are organized in the Periodic Table so those with similar properties belong to the same group
- Metals, nonmetals, and semimetals:
  - Location on Periodic Table and properties
- Know which elements are solids, liquids, gases at room temperature (25°C)
- Know which elements exist as diatomic molecules at 25°C (H₂, N₂, O₂, F₂, Cl₂, Br₂, I₂)
- molecule: compound of 2 or more nonmetals held together by covalent bonds
- ionic compound: consists of metal + nonmetal(s)
  - formula unit indicates ratio of ions present
Chapter 2 (Continued)

- Know the names and symbols of the first 18 elements of the Periodic Table, all elements on Fig. 2.12 on p. 59, and U. Correct spelling of element names counts!

KNOW all of names and symbols for the POLYATOMIC IONS

Nomenclature:
- Identify a compound as ionic or molecular
- Given formula of a compound, determine name.
- Given name of a compound, determine formula.

Chapter 3

Atomic Mass
- Know that atomic weights reported on periodic table are the weighted average of naturally occurring isotopes

Molar Masses
- Calculate molar mass of element or compound
  1 mole = Avogadro’s # = 6.022×10^{23} units

Mole Conversions
- Convert among masses, molar mass, moles, molecules, atoms

Percent Composition (or Mass Percentages)
- Calculate the percent by mass of each element in a compound

Empirical and Molecular Formulas
- Know empirical formulas are simplest, whole-number ratio of atoms.
- Determine empirical formula given elemental analysis data
- Determine the molecular formula given the empirical formula and molecular weight

Balancing Chemical Equations
1. Balance metals.
2. Balance polyatomic ions - Keep as one unit.
5. Balance oxygens.
6. Balance all other atoms.

Molecule or Molecular Compounds: all nonmetals
- Use Greek prefixes when more than one
- Know ammonia, methane, hydrogen peroxide

Ionic Compounds: metal + nonmetal(s) ions
- Cation(s) and anion(s) with overall charge of zero
- Determine the charges on Main-Group elements using the Periodic Table
- Use charges on anions to determine charges on metals that can multiple charges.
- Have high melting points, all solids at 25°C

Acids: Have H in front, physical state is (aq)
- Be able to name acids given the formula or determine the formula given the name.

Interpreting Chemical Equations
- Reactants and products
- (s), (l), (g), and (aq)
- Coefficients give mole ratio as well ratio of elements and compounds

Stoichiometry: Calculations involving amounts of reactants and products in a chemical reaction
- Coefficients in balanced equation give mole ratios
- Mole-mole relationships
- Mass-mass relationships

Limiting Reagent Problems
- Calculate the mass or volume of product made given amount of reactants and chemical equation
- Smallest amount of product = amount produced
- Reactant producing smallest = limiting reagent
- All other reactants = reactants in excess
- Calculate the amount of reactant in excess that remains after the reaction

Yields of Reactions
theoretical yield: amount of product predicted by balanced equation when limiting reagent is used up
actual yield: amount of product one actually gets

\[
\text{Percent yield} = \frac{\text{actual yield}}{\text{theoretical yield}} \times 100\%
\]
Chapter 4

Aqueous Solutions: liquid or solid solute in H₂O
- solution: mixture of 2 or more substances
- solute: substance present in smaller amount
- solvent: substance present in larger amount

KNOW the Solubility Rules—these will not be provided on the ACS Exam!

Predict products of a reaction given reactants for
- Precipitation Reactions
  - precipitate (ppt): solid forming from 2 solutions
- Acid-Base Neutralization Reactions
  - HX + MOH → water + salt
  - HX + MHCO₃ → water + CO₂ + salt
  - HX + MCO₃ → water + CO₂ + salt
- Oxidation-Reduction (Redox) Reactions
  - Combination reaction
    - metal + nonmetal → ionic cmpd (s)
  - Single-replacement reactions
    - Given Activity Series
    - solid metal + metal solution
    - solid metal + acid
    - active metal + H₂O (l)
  - Combustion reaction
    - CₓHᵧ + O₂ → CO₂ (g) + H₂O (g)
    - CₓHᵧO₂ + O₂ → CO₂ (g) + H₂O (g)

Know Arrhenius and Bronsted-Lowry (B-L) definitions for acids and bases

Electrolyte: substance that dissolves in water to produce ions that conduct electricity
- strong electrolyte: breaks up completely
  → many ions present to conduct electricity
  - e.g. strong acids & bases, aqueous salts
- weak electrolyte: breaks up to small degree
  → only a few ions present to conduct electricity
  - e.g. weak acids & bases, insoluble salts
- nonelectrolyte: a molecular compound that forms molecules in water (e.g sugar)
  → no ions → does not conduct electricity

Acids and Bases as electrolytes
- Know strong acids and strong bases!
- All other acids and bases are weak
- Recognize strong acids and bases ionize completely while weak acids and bases don’t

Chemical Equations and (Net) Ionic Equations
- Chemical Equation: compounds shown intact
- Complete/Total Ionic Equation:
  - shows strong electrolytes as ions
- Spectator Ions: unchanged in reaction
- Net Ionic Equation: Shows substances that change in a chemical reaction

Guidelines for writing Net Ionic Equations:
1. Complete and balance molecular/chem equation
2. Leave solids, liquids, gases, weak electrolytes as compounds; break up strong electrolytes
3. Cancel spectator ions
4. Simplify coefficients if possible
5. If all reactants and products cancel (all spectator ions) → NR

Oxidation-Reduction (Redox) Reactions
- Be able to determine oxidation numbers for all the atoms/elements in a chemical equation.
- Use oxidation numbers to determine which reactant was oxidized (reducing agent) and which reactant was reduced (oxidizing agent)
- Determine the # of electrons gained or lost.

Molarity, Solution Stoichiometry, Mass Percent Concentration, and Volumetric Analysis
- Use molarity and volume to solve for moles
- Use chemical equation given to determine mole-to-mole ratios needed to solve problems
- Solve a variety of problems combining molarity, mass percent concentration, and topics covered from previous chapters.

Acid-Base Titrations
- Know the definitions for standard solution, acid-base indicators, titration, endpoint
- Solve problems using titration data.
- Carry out error analysis (how different kinds of experimental error will impact experimental results)
Chapter 5: Thermochemistry

energy: capacity to do work or to produce heat
heat: energy is transferred from a hotter substance to a cooler substance
kinetic energy (KE): KE = \frac{1}{2}mv^2
- energy associated with an object's motion
potential energy: energy associated with an object's position or its chemical bonds

joule (J): 1 J = \frac{1 \text{ kg} \cdot \text{m}^2}{\text{s}^2} and 1 kJ=1000 J
calorie (cal): 1 cal = 4.184 J (This is exact!)

- Distinguish between system and surroundings

heat (q):
- recognize the meaning of signs for heat
  - q is + for heat flow into system
  - q is - for heat flow out of system

heat of reaction (q_{reaction}):
- endothermic reaction: q_{reaction} is +
  - energy of reactants < energy of products,
    → surroundings feel cooler after reaction
- exothermic reaction: q_{reaction} is -
  - energy of reactants > energy of products,
    → surroundings feel hotter after reaction

- Classify different chemical or physical changes as exothermic or endothermic.

- 1^{st} Law of Thermodynamics: The energy of the universe is constant.
  - Law of Conservation of Energy: energy is neither created nor destroyed in a chemical reaction, so the total E of reactants equals total E of the products.

- Use following equations to solve problems:
  heat = (c_s=specific heat)(mass)(\Delta T)
  heat = (c_p=molar heat capacity)(n)(\Delta T)

- The specific heat of water (4.184 J/g·°C or 1.00 cal/g·°C) will be provided.

Calorimetry Problems
- Use the heat absorbed/lost by water/solution to determine heat lost by a metal or associated with reaction to determine \Delta H or specific heat of metal (e.g. Calorimetry experiment)

Know the regions and features of a Heating/Cooling Curve

Under constant pressure (like atmospheric pressure), q_{reaction} = \Delta H
- for endothermic reactions, \Delta H is +
- for exothermic reactions, \Delta H is -

Thermochemical Equations indicate the enthalpy changes as well as reactants and products
- Be able to solve stoichiometry problems involving enthalpy changes using a thermochemical equation.

Guidelines for Thermochemical Equations
1. The magnitude of \Delta H is proportional to the amount of reactants and products in the equation.
2. \Delta H is the opposite sign for the reverse reaction.
3. If the coefficients in an equation are multiplied by a factor n, \Delta H must able be multiplied by n.

Hess' Law: The value of \Delta H for a reaction is the same whether it occurs in one step or several steps: \Delta H = \Delta H_1 + \Delta H_2 + \Delta H_3 + ...

Standard Enthalpy of Reaction (\Delta H_{rxn}^\circ):
- Enthalpy change accompanying a chemical reaction at standard state conditions

Standard Enthalpy of Formation (\Delta H_f^\circ):
- Enthalpy change for the formation of 1 mole of a compound from its elements in their standard states.

Solve problems using \Delta H_{rxn}^\circ and \Delta H_f^\circ:
- Be able to calculate \Delta H_{rxn}^\circ given \Delta H_f^\circ for all the reactants and products
- Be able to calculate \Delta H_f^\circ for one reactant or product given \Delta H_{rxn}^\circ and \Delta H_f^\circ for the other reactants and products.
Chapter 6: Gases

Know the physical properties of gases

Gas pressure and Atmospheric pressure
- Standard atmospheric pressure: 760 torr at 0°C
- Convert between units of pressure:
  \[1 \text{ atm} = 760 \text{ torr} = 760 \text{ mmHg}\]

Know how a manometer is used to measure gas pressure for open systems (p. 2 of lecture notes).

Solve for a variety of problems involving gases
- Given 2 sets of conditions, use
  \[\frac{P_1 V_1}{T_1} = \frac{P_2 V_2}{T_2}\]
  cancel variables that stay the same to simplify, and temperature must be in Kelvins
- Use ideal gas law (PV = nRT) to solve for P, V, n, or T. R = \(\frac{0.08206 \text{ L} \cdot \text{atm}}{\text{mol} \cdot \text{K}}\) will be given.
- Recognize that STP is standard temperature and pressure, defined at 0.00°C and 1.00 atm.
- Use the ideal gas law and unit analysis to solve for density, molar mass, and other properties.
- Solve stoichiometry problems using 22.4L/mol at STP and \(n = \frac{PV}{RT}\) for other conditions.

Dalton's Law of Partial Pressure:
- Use Dalton's Law \(P_{\text{total}} = P_1 + P_2 + P_3 + \ldots\) to solve for total pressure or the partial pressure of one gas in a mixture

Chapter 7: Atomic Theory

wavelength (\(\lambda\)): distance between successive peaks; generally given in m, cm, or nm
- Know 1 m = 10⁹ nm

Use equation: \(c = \lambda v\)

speed of light, \(c\): 2.998×10⁸ m/s (will be given)

frequency (v): number of crests passing by a given point in 1 s; given in \(\frac{1}{s}\) = hertz (Hz)

wavelength (\(\lambda\)): distance between successive peaks; generally given in m, cm, or nm
- Know 1 m = 10⁹ nm

Use equation: \(c = \lambda v\)

speed of light, \(c\): 2.998×10⁸ m/s (will be given)

frequency (v): number of crests passing by a given point in 1 s; given in \(\frac{1}{s}\) = hertz (Hz)

Relate energy, frequency, and wavelength:
\(E = hv = \frac{hc}{\lambda}\)

Dual Nature of the Electron
- Electron can behave like a wave or a particle
Chapter 7 (Continued)

Planck and Quantum Theory
- proposed energy is absorbed and emitted as bundles = quanta
- single bundle of energy = quantum

Einstein and the Photoelectric Effect
- Be able to describe the Photoelectric Effect
- Know this provided experimental evidence for light existing as particles = photons

Bohr Theory of the Atom
- Electrons move in quantized orbits called “energy levels” around nucleus
  - ground state: electron in lowest energy level
  - excited state: electron in higher energy level
  - When atom absorbs energy, electron jumps from lower energy to higher energy level.
  - When electron drops from a higher energy to a lower energy level, it releases energy, in some cases as light → atomic emission spectra
  - Know limitations of the Bohr Model
  - Recognize energy levels further from the nucleus are closer together
  - Calculate the change in energy $\Delta E$ for an electron moving from one energy level ($n_{\text{initial}}$) to another energy level ($n_{\text{final}}$):
    \[
    \Delta E = -2.178 \times 10^{-18} \text{ J} \left( \frac{1}{n_{\text{final}}^2} - \frac{1}{n_{\text{initial}}^2} \right)
    \]

Chapter 8: Bonding

valence electrons: outermost electrons
- the group number for each element is equal to the number of valence electrons it has

electronegativity (EN): ability of an atom in a bond to draw electrons to itself
  - Know F is most electronegative, further away from F, less electronegative an atom.

ionic bond: electrostatic attraction between cations and anions, holding the ions together in an ionic compound

be able to write electron configurations for any neutral atom or ion
- Be able to write using full notation and core notation (Noble Gas abbreviation)
- ground state: electrons fill lowest energy orbitals first
- excited state: electrons in higher energy orbitals before lower levels filled
- Recognize extra stability gained with filled and half-filled d orbitals (Cr, Mo, Cu, Ag)
- Account for electrons gained or lost for ions
- Know Representative Elements usually form ions that are isoelectronic with a Noble Gas
- Note that transition metals, Sn, Pb lose their s electrons before their d electrons.

be able to draw atomic orbital diagrams
- Know Pauli Exclusion Principle
- Know Hund’s Rule
- Show that each sublevel is slightly higher in energy than the previous sublevel

Know definitions and periodic trends for atomic radius, ionization energy

Know Ionic Radius trend:
\[ r_{\text{cation}} < r_{\text{neutral atom}} < r_{\text{anion}} \]

Quantum Numbers ($n$, $l$, $m_l$, and $m_s$)
- Know the restrictions on each and the orbital described by a set of quantum numbers.

covalent bond: sharing of electrons between two nonmetal atoms
- polar covalent: unequal sharing of electrons by 2 atoms with different EN values
- nonpolar covalent: equal sharing of electrons by two atoms with equal EN

bond length: distance between nuclei of 2 bonded atoms; shorter the bond → stronger the bond.
- single bonds are the longest and weakest
- double bonds are shorter and stronger
- triple bonds are the shortest and strongest
Chapter 8: Bonding (Continued)

octet rule: atoms bond such that each has 8 electrons, except H only needs 2 electrons.

Draw Lewis Structure (Electron Dot Formulas) for Molecules & Polyatomic Ions

Exceptions to the Octet Rule
- Incomplete octet
  - Be can have 4 e⁻s, B can have 6 e⁻s, and N can have an 7 e⁻s
- Elements in the third period and lower can have an expanded octet (>8 electrons).

Lewis structures for ternary oxyacids
- Oxygens around the central atom, and one hydrogen each bonded to an oxygen

Chapter 9

Valence-Shell Electron Pair Repulsion (VSEPR) Model
- repulsions between electrons give rise to the shapes or molecular geometries of molecules

Be able to predict the shape of a molecule using its Lewis electron dot formula.
- If there are only two atoms in the molecule, then it’s linear.
- For molecules with 3 or more atoms, determine the general formula and match that with the corresponding shape.

\[
\begin{array}{c}
\text{A} = \text{central atom}, \ B = \text{outer atom(s)}, \\
\text{E} = \text{lone pair(s) on central atom}
\end{array}
\]

AB₂ → linear, 180° bond angle
AB₃ → trigonal planar, 120° bond angle
AB₄ → tetrahedral, 109.5° bond angle
AB₅ → trigonal bipyramidal, 90° & 120° angles
AB₆ → octahedral, 90° bond angle

Variation on trigonal planar shape
AB₂E → bent, <120° bond angle

Variations on tetrahedral shape
AB₃E → trigonal pyramidal, <109.5° bond angle
AB₂E₂ → bent, <109.5° bond angle

Formal charges: hypothetical charge an atom would have if bonding electrons are shared equally and lone pairs belong solely to a single atom

\[
\text{Formal charge} = \frac{\text{total} \# \text{of valence electrons in free atom}}{2} - \frac{\text{total} \# \text{of nonbonding electrons}}{2} - \frac{\text{total} \# \text{of bonds}}{2}
\]

- Use formal charges to determine the most plausible structure.
- Minimize formal charges for molecules where the central atom can have an expanded octet

resonance structures: two or more structures representing a single molecule that cannot be described fully with only one Lewis structure
- Recognize which molecules require resonance structures by drawing electron dot formulas for given examples.

Variations on trigonal bipyramidal shape
AB₄E → see-saw, <90° & <120° bond angles
AB₃E₂ → T-shaped, <90° bond angles
AB₂E₃ → linear, 180° bond angle

Variations on octahedral shape
AB₅E → square pyramidal, <90° bond angles
AB₄E₂ → square planar, 90° bond angles

Draw Lewis structures for molecules with multiple central atoms.
- Determine shape and bond angle around each central atom
- Recognize that in oxyacids (e.g. HNO₃, H₂SO₄) the H atoms are bonded to the O atoms.

bond energy: energy required to break a bond
- always positive
- Use bond energies given to calculate ΔH

Polarity
- Use the shape and electronegativity differences to draw dipoles and determine if the molecule is polar or nonpolar
Chapter 9

Valence Bond Theory
- Recognize that two atoms form a bond because they share electrons to become more stable
- The bond length (distance between nuclei) is achieved by balancing the electron-nuclei attractions and electron-electron and nucleus-nucleus repulsions

Steric Number (SN)
\[
SN = \left( \frac{\text{# of outer atoms bonded to central atom}}{\text{central atom}} \right) + \left( \frac{\text{# of lone pairs on central atom}}{\text{central atom}} \right)
\]

Hybrid Atomic Orbitals
- Identify for the central atom(s) in a molecule:
  - \(AB_2\), linear (180°) \(\rightarrow\) \(SN=2\) \(\rightarrow\) \(sp\)
  - \(AB_3\), trigonal planar (120°) or \(AB_2E\), bent (120°) \(\rightarrow\) \(SN=3\) \(\rightarrow\) \(sp^2\)
  - \(AB_4\) (109.5°), \(AB_3E\), \(AB_2E_2\) (109.5°) \(\rightarrow\) \(SN=4\) \(\rightarrow\) \(sp^3\)
  - \(AB_5\) (90° & 120°), \(AB_4E\) (90° & 120°), \(AB_3E_2\) (120°), \(AB_2E_3\) (90°) \(\rightarrow\) \(SN=5\) \(\rightarrow\) \(sp^3d\)
  - \(AB_6\) or \(AB_4E_2\) (90°), \(AB_5E\) (90°) \(\rightarrow\) \(SN=6\) \(\rightarrow\) \(sp^3d^2\)

Chapter 10

Describe solids, liquids, and gases in terms of kinetic energy of their particles and molecular motion.

Ionic Compounds
- Coulomb's Law
  - Relative strength of an ionic bond is given by the following:
    - Charges of ions:
      Higher the charges \(\rightarrow\) stronger the bond
    - Distance between two ions:
      Shorter distance \(\rightarrow\) stronger the bond
    - Lattice energy: energy released when gaseous ions come together to form an ionic compound
  - Recognize how melting point and lattice energy are related to the strength of ionic bonds

Ion-Dipole Forces: attraction between an ion and polar molecule (e.g. Na⁺ and H₂O)

• sigma bond (σ): bond formed from direct overlap of two orbitals
  - formed by single bonds and first bond for double and triple bonds
• pi bond (π): bond formed from sideways or indirect overlap of two orbitals on a separate plane from the sigma-bonded atoms
  - formed by the second and third bonds in double and triple bonds.
Identify bonds as sigma (σ) and pi (π) bonds using the Lewis structure of a molecule or polyatomic ion.

Molecular Orbital Model
- Electrons in molecules occupy molecular orbitals
  - electrons delocalized over entire molecule, not just bound between 2 atoms
- Explains resonance, bond strength, and paramagnetism of molecules more effectively than Localized Electron Model

Intermolecular Forces (IMF's): attraction between 2 different molecules in a liquid or solid
- Identify IMF's experienced by molecules.
- Also called van der Waals forces

London/Dispersion (Induced-Dipole) Forces: attraction between molecules with instantaneous dipole resulting from electrons shifting
- only type of IMF's between nonpolar molecules
- experienced by all molecules
- becomes stronger with more polar molecules
  - the more electrons in the molecule (higher molar mass) \(\rightarrow\) stronger the London forces

Dipole-dipole Forces: attraction between polar molecules without H-O, H-N, or H-F bond
- Stronger than London forces because polar molecules have permanent dipoles
Chapter 10 (Continued)

Hydrogen Bonds: attraction between polar molecules containing H–O, H–N, or H–F bond
- Strongest type of intermolecular force

Ionic and covalent bonds are stronger than any intermolecular force, even hydrogen bonds.

Intermolecular Forces (IMFs) and Physical Properties
- Know IMF’s influence boiling point, vapor pressure.
- Use IMFs to rank substances in terms of increasing boiling or melting points.
- normal melting point (m.p.) or boiling point (b.p.) always at 1 atm

Intermolecular Forces vs. Ionic/Covalent Bonds
- Distinguish between IMF’s between different molecules and the chemical bonds (ionic, polar covalent, or nonpolar covalent) holding atoms or ions together in a compound
- Identify type of bond made or broken for a chemical reaction or a change in physical state.

Structure and Properties of Water
- Because of hydrogen bonds,
  - density of ice is lower than the density of liquid.
  - water has relatively high m.p. and b.p.

Crystalline Solids: w/ regular, geometric pattern
- Identify a solid as ionic, molecular, or network covalent (graphite and diamond) given its formula.

YOU WILL BE GIVEN ALL RELEVANT CONSTANTS AND A PERIODIC TABLE WITH ELEMENT SYMBOLS, ATOMIC NUMBERS, AND ATOMIC MASSES.

THE EXAM IS MULTIPLE-CHOICE SCANTRON FORMAT, SO BRING A BASIC SCANTRON FORM, A #2 PENCIL, AND A BASIC SCIENTIFIC CALCULATOR.