Textbook: *Chemistry: An Atoms-Focused Approach*, 2nd ed., by Gilbert, Kirss, Foster, and Bretz

Chapter 1
Upon completion of this chapter, you should be able to:

- Describe the features of the three states of matter at the macroscopic and molecular/atomic levels.
- Define and distinguish between the following terms: element, atom, compound, molecule, homogeneous and heterogeneous mixtures
- Define and give examples of the following: physical property, extensive property, intensive property, physical change, chemical change
- Know and convert between the metric units and the prefixes specified (Giga- through femto-).
- Convert between the different volume units commonly used in chemistry.
- Understand precision and accuracy.
- Identify the number of significant figures in a given value and determine the proper number of significant figures to include in a calculated value, using proper rounding.
- Quickly perform organized dimensional analysis calculations (with multiple steps) to convert from one set of units to another.
- Use the relationship between mass, volume and density and other simple equations (from geometry, etc.) to solve problems.

Chapter 2, Sections 1-4
Upon completion of these sections, you should be able to:

- Describe the modern view of atomic structure in terms of the three types of subatomic particles. Know where these particles are located in the atom, their charges and relative masses. Understand the relative sizes of the nucleus and the atom.
- Define the following terms: isotope, atomic number, mass number.
- Calculate the average atomic mass of an element and understand the origin of the masses given on the periodic table.
- Identify the number of protons, neutrons and electrons and the overall charge of an atom and write its complete chemical symbol.
- Provide the names corresponding to element symbols and correct symbols for written element names (elements 1-88).
- Understand the arrangement of the periodic table, including the groups discussed in class (alkali metals, alkaline earth metals, halogens, noble gases), and be able to differentiate between metals, nonmetals and metalloids (or semimetals).
- Know what cations and anions are.
- Use the periodic table to predict the charges of common monoatomic ions; determine the numbers of subatomic particles in monoatomic ions and write their complete chemical symbols.

Chapter 3
Upon completion of this chapter, you should be able to:

- Understand the relative placement of the various regions of the electromagnetic spectrum in terms of wavelength, frequency and energy.
- Calculate the wavelength or frequency of light, given the other quantity.
- Understand what is meant by *quantization of energy*.
  - Calculate the energy, wavelength or frequency of a single photon.
- Describe the photoelectric effect.

- o Determine whether an electron can be ejected from a metal with a particular work function and a photon of a given energy, frequency or wavelength.
- Calculate changes in energy associated with the electronic transitions in the H atom (from Bohr's model). Understand where the lines in the H-atom emission spectrum come from.
- Understand the meaning of *wave-particle duality* and its relative importance for macroscopic and atomic-size objects.
    - o Calculate the de Broglie wavelength of a particle.
    - o State the Heisenberg Uncertainty Principle and understand its consequences for electronic structure. (Can we know where an electron is in an atom? Why or why not?)
- Understand, in general, the results of Schrödinger's wave equation:
    - o Understand how quantization of energy applies to the electron wave.
    - o Understand what orbitals are and where quantum numbers come from.
    - o Explain how solutions to Schrödinger's equation can be used to determine electron (probability) density.
- Know the three quantum numbers needed to define an orbital, what values each may have, and what each one tells us about the orbital.
- Draw representations of *s*, *p* and *d* orbitals, indicating orbital shapes, relative signs, and orientations in space.
    - o Locate any nodes present in each orbital, and understand their significance.
- Know the fourth quantum number needed to specify a certain electron in an orbital, what it represents, and what values it may have.
    - o Write a set of four quantum numbers to describe a particular electron in an atom or ion.
- Know the ordering of the energies of atomic orbitals for multi-electron atoms.
- State the Pauli Exclusion Principle and Hund's Rule, and understand their implications for electron configurations.
- Give the lowest energy (ground-state) electron configuration and orbital "box" diagram for any atom or monoatomic ion.
- Understand how orbital size (or "radial extent") changes with increasing (or decreasing) principal quantum number, *n*.
- Understand and estimate the effective nuclear charge experienced by a valence electron in a given atom.
- Use your understanding of effective nuclear charge and radial extent to explain periodic trends in:
    - o Atomic and ionic radii
    - o First ionization energy
- Define first ionization energy and write a chemical reaction to describe the ionization process.
- Explain increases in successive (1st, 2nd, 3rd, 4th, etc.) ionization energies in terms of electronic configurations.
- Define electron affinity and write a chemical equation to describe this process.
- Use your understanding of effective nuclear charge and radial extent to explain periodic trends in electron affinity.

End of Material for Exam I

Chapter 4
Upon completion of this chapter, you should be able to:

- Know the names, formulas and charges for the common polyatomic ions provided on the handout.
- Write a formula for a given ionic compound; give the correct name for a compound formula. Name binary molecular compounds and provide formulas for names.
- Describe ionic and covalent bonding:
  - Predict whether a particular bond will be ionic or covalent (i.e., whether a particular compound is ionic or molecular)
  - Understand how ionic bonds are formed (starting with neutral atoms) and the driving force for this process; compare lattice energies for different ionic compounds.
  - Understand covalent bond formation.
- Draw the Lewis structure for a given atom, molecule or ion. Understand the "octet rule" and its exceptions.
- Draw equivalent or non-equivalent resonance structures for a given molecule or ion and understand their implications for the actual depiction of bonding in the molecule/ion.
- Define electronegativity; predict and explain periodic trends in electronegativities.
- Assign formal charges to atoms in Lewis structures; use formal charges (and electronegativities) to evaluate Lewis structures.
- Determine bond order and understand trends in bond length and strength among single, double, and triple bonds.
- Understand the continuum between 100% ionic and 100% covalent bonding, including polar covalent bonds and covalent character in ionic compounds.

Chapter 5, Sections 1-6
Upon completion of these sections, you should be able to:

- Understand the basis of VSEPR Theory and use it to predict electron domain geometries and molecular geometries.
- Sketch and name molecular geometries and provide estimated bond angles.
- Identify polar covalent bonds and use molecular geometries to determine whether molecules/ions are polar or nonpolar.
- Describe Valence Bond Theory and explain why it (alone) is not an acceptable bonding model.
- Explain how hybrid orbitals are formed, starting from valence electron configurations.
  - Understand why hybridization is used as a modification to Valence Bond Theory and what its limitations are.
- Predict the types of valence and/or hybrid orbitals used by atoms in molecules/ions.
- Understand what sigma and pi bonds are and in which situations each type is formed.
- Explain how atomic and/or hybrid orbitals overlap to form the bonds (both sigma and pi) in a molecule or polyatomic ion; also, be able to specify the type of orbital holding each lone pair.

Chapter 7, Sections 2.5-2.6, Section 6.3, and Sections 8.1-8.3
Upon completion of these sections, you should be able to:

- Know what a mole is and where Avogadro's number came from.
- Calculate the molar mass for a given compound.

- Use Avogadro's number and the concept of molar mass to interconvert between mass, moles, and number of atoms of an element (or number of molecules of a compound).
- Write and balance chemical equations.
- Write and balance (complete) combustion reactions, given only the formula of the compound to be burned.
- Understand and apply balanced chemical equations (stoichiometry) to:
  - Determine the mass (or number of moles) of a desired reactant or product given the mass (or number of moles) of another.
  - Identify limiting reactants and reactants in excess.
  - Calculate the mass of remaining excess reactant.
  - Determine percent yields.
- Define the following terms: *solution, solvent, solute, molarity, strong electrolyte.*
- Calculate the molar concentration of a solution.
  - Interconvert between molarity, moles solute, and volume solution.
  - Use molar concentration to work with aqueous solutions in stoichiometry problems.
- Complete calculations related to dilution, determining an unknown volume or concentration.

End of Material for Exam II

Chapter 10
Upon completion of this chapter, you should be able to:

- Understand what pressure is and what causes it.
  - Use and convert between the different units used to express pressure.
- Understand how changes in pressure, volume, temperature and moles of gas affect properties of gases (what happens to volume when pressure increases, etc.).
- Use the ideal gas equation to determine an unknown pressure, volume, temperature or amount (moles) of gas, given the other three quantities.
- Use the ideal gas equation (incorporating d and M) to determine densities and molar masses of gases.
- Solve ideal-gas problems involving initial and final P-V-T-n conditions.
- Solve stoichiometric problems involving gaseous reactants or products, starting with the ideal gas law.
- Understand Dalton's law of partial pressures.
  - Calculate the partial pressures of gases in a mixture given moles (or mole fractions) of each and total pressure.
- Understand the main points of Kinetic Molecular Theory (KMT).
  - Use KMT to explain the properties of gases (P and T) on a molecular level.
  - Understand relationships between temperature, average kinetic energy and average speed.
  - Use KMT to explain the effects on pressure of: 1) increased V at constant T; 2) increased T and constant V; and 3) increased n at constant V and T.

Sections 6.1-6.2 and 11.3 (Intermolecular Forces)
Upon completion of these sections, you should be able to:

- Differentiate between the states of matter based on the relative strengths of their intermolecular forces (IMF) and relative speeds of their particle motions.
- Understand the five types of IMF discussed (ion-dipole, dipole-dipole, hydrogen bonding, dipole-induced dipole, and induced dipole-induced dipole, a.k.a. London dispersion), including:
  - Conditions under which each IMF is active (e.g., in polar versus non-polar molecules)
  - Diagram specifically depicting each type of IMF
  - Factors influencing strengths of each IMF
  - Relative strengths of different types of IMF (e.g., which two types are weakest and why?)
- Define *viscosity, surface tension, melting point and boiling point.* Understand and predict how these will vary with varying strengths of IMF.
- Recognize the different classifications of solids and the types of interactions holding atoms, ions, or molecules together in each.
  - Compare and predict relative melting points of different types of solids based on the types of forces holding particles together.

Chapter 9
Upon completion of this chapter, you should be able to:

- Understand the following terms: *energy, system, surroundings, heat, work, potential energy, kinetic energy.*
- State and use the First Law of Thermodynamics.
  - Understand the conversion between potential and kinetic energy.
  - Calculate the change in internal energy (ΔE) of a system resulting from inputs and outputs of work and heat.
  - Understand and use sign conventions for heat, work and changes in energy.
- Define and give examples of state and path functions.
- Understand what enthalpy is and use the sign convention for enthalpy changes.
- Describe endothermic and exothermic processes.
- Understand how to use thermochemical equations; calculate ΔH for reactions involving consumption of a particular amount of reactant (or production of a particular amount of product).
- State Hess's Law and use it to determine the enthalpy change for a reaction (given reactions of known ΔH).
- Use standard enthalpies of formation (ΔH_f° values) to determine ΔH°_{rxn}.
  - Understand what is meant by the standard enthalpy of formation of a substance.
  - Write the chemical equation corresponding to the standard enthalpy of formation for a given compound.
- Understand the relationship between heat, mass (or moles), heat capacity and temperature change and use it to determine any unknown quantity (given the others).
  - Calculate the heat gained or lost during heating within a phase.
  - Calculate the heat gained or lost during phase changes.
  - Deal with heat transfer between different components of a system ($q_{gained} = -q_{lost}$).
- Relate q to the change in enthalpy for a reaction.

- Understand and describe how to conduct simple constant-pressure ("coffee cup") calorimetry experiments.
    - Use calorimetry data to calculate q (in kJ) and ΔH (in kJ/mol) for a particular reaction.

Chapter 12
Upon completion of this chapter, you should be able to:

- Clearly distinguish between *spontaneous, non-spontaneous* and *reversible* processes.
- Define entropy and understand what causes it on a molecular level.
    - List factors affecting the entropy of molecules and use your understanding to compare entropies of different substances.
- Calculate $\Delta S°$ for a chemical reaction from standard molar entropies in Appendix 4.
- Determine whether a process is spontaneous:
    - Based on $\Delta S_{universe}$
    - Based on ΔG_{system}
- Predict the sign of ΔG and spontaneity or non-spontaneity based on estimates of ΔH and ΔS.
- Calculate $\Delta G°_{rxn}$:
    - From $\Delta G_f°$ data in Appendix 4.
    - From calculated values of $\Delta H°$ and $\Delta S°$.

End of Material for Exam III

Chapter 13
Upon completion of this chapter, you should be able to:

- Understand reaction energy profiles, and the information they provide, including:
    - The sign of ΔG.
    - The magnitude of the activation energy, and its effect on reaction rate.
    - The effect of adding a catalyst. (How does this change the profile and the rate?)
- Use your understanding of collision frequency to explain the effects of reactant concentrations, increased temperature, and gaseous state on reaction rates.
- Use stoichiometry to express reaction rates in terms of changing reactant and/or product concentrations.
- Understand the difference between average and instantaneous rates, and explain how each is determined.
- Use experimental data to determine the rate law for reaction, including the order in each reactant, the overall order, and the value of the rate constant:
    - Using the method of initial rates.
    - Using integrated rate laws and graphical analysis.
- Use integrated rate laws to determine the fraction of initial reactant remaining at a given time point, or the time required for reactant concentration to reach a specified fraction of its initial value.
- Understand the concept of half-life, and use the relationship between $t_{1/2}$ and k to find a desired value.
- Use the Arrhenius equation to explain relationships between reaction rate constant, activation energy, and temperature, and to calculate any one of these quantities.

Chapter 14
Upon completion of this chapter, you should be able to:

- Define *equilibrium*; understand the implications for rates of forward and reverse processes and the value(s) of ΔG.
- Write expressions for equilibrium constants of provided reactions:
  - K_P, using partial pressures of gases
  - K_C, using molar concentrations of solutions (and/or gases)
- Interpret K_P and K_C values to determine what species dominate at equilibrium (reactants vs. products).
- Determine the reaction quotient, Q_P or Q_C for a given reaction.
  - Use Q to determine whether a reaction is at equilibrium and, if not, in which direction it must proceed to reach equilibrium.
- Use the relationship between $\Delta G°$ and K to determine one quantity from the other.
- Calculate ΔG_{rxn} under non-standard conditions from $\Delta G°$ and Q.

- Create and use equilibrium tables (containing initial conditions, changes in conditions and equilibrium conditions for reactants and products) to solve problems involving chemical equilibria:
  - Determining equilibrium conditions from initial value(s) and K_P or K_C.
  - Solving for K_P or K_C given certain initial and equilibrium conditions.
- Use your understanding of Le Chatelier's Principle to predict how a reaction at equilibrium will respond to disturbances caused by: addition/removal of a reactant or product; compression/expansion; changes in temperature; addition of a catalyst.
  - Apply your understanding of Le Chatelier's Principle (and/or Q) to quantitative equilibrium calculations, determining amounts of reactants and products that will ultimately result from disturbing the equilibrium in a particular way.

Chapter 15 and Section 8.4
Upon completion of this material, you should be able to:

- Define *acid* and *base* according to both the Brønsted-Lowry and Lewis definitions.
- Write acid- and base-ionization reactions in aqueous solution; identify conjugate acid-base pairs.
- Understand what is meant by *strong* and *weak* acids and bases.
  - List the common strong acids and bases in water.
  - Compare acid or base strengths given K_a or K_b values.
  - Understand the relationship between strengths of conjugate acid-base pairs; calculate K_b from K_w and K_a (or K_a from K_w and K_b).
- Write the equilibrium for the autoionization of water and the expression corresponding to K_w.
- Understand how pH and pOH are related to the concentrations of H_3O^+ and OH^- ions, respectively; convert between concentration and pH and/or pOH.
- Determine the pH and/or pOH of a solution:
  - For a strong acid or base (using stoichiometry)
  - For a weak acid or base (using an equilibrium table)

- Write the balanced neutralization reaction for any acid and base, including:
  - The balanced overall ("molecular") reaction
  - The complete ionic equation, including spectator ions
  - The net ionic equation

End of Material for Exam IV

The cumulative final exam will cover all material listed above, plus the following:

Chapter 17 and Section 8.6
Upon completion of these sections, you should be able to:

- Define the following terms: *oxidation, reduction, oxidizing agent, reducing agent.*
- Determine the oxidation number of each atom in a chemical reaction; identify which species are oxidized and reduced.
- Balance oxidation-reduction reactions using the method of half-reactions.
  - Use the appropriate species to balance H and O in acidic or basic conditions.
- Understand galvanic or voltaic cells:
  - Describe what processes occur at the anode and cathode and know which directions electrons, cations and anions spontaneously flow.
  - Draw pictures to represent galvanic/voltaic cells (containing anode and cathode half-cells).
  - Determine $E°_{cell}$ and spontaneity or non-spontaneity for balanced oxidation-reduction reactions from tables of standard reduction potentials (as in Appendix 6).
  - Write the spontaneous reaction that will take place between two half-reactions, given their standard reduction potentials.
  - Relate $E°_{cell}$ to $\Delta G°$.