CHGN 122 Course Outcomes

Overall Course Level Learning Outcomes:
After fully participating in this course, you should be able to…

1. **Predict physical properties of pure substances & solutions resulting from chemical bonding & intermolecular forces.**
2. **Compare/contrast the direction & spontaneity of chemical & physical processes based on energy & entropy changes.**
3. **Construct chemical rate laws from sequences of molecules collisions; analyze the factors that determine reaction rates.**
4. **Predict the extent & direction of chemical reactions approaching equilibrium; assess the impact of stresses to a system at equilibrium.**
5. **Determine the properties & composition of aqueous systems by applying equilibrium concepts & acid/base theory.**
6. **Design experiments & clearly communicate your findings to scientists & non-scientists.**
7. **Explain how each concept above applies to modern science & engineering challenges.**

Detailed Course Learning Outcomes:
After fully participating in this course, I can…

- explain why intermolecular forces (IMFs) are responsible for the liquid and solid states of covalently bonded substances.
- from molecular structures, identify & distinguish between the types & strengths of IMFs in a pure substance or a mixture: London dispersion forces, dipole-dipole forces, hydrogen bonding and ion- dipole forces.
- explain how Coulomb's Law predicts the energy of attraction in an IMF or chemical bond. o explain how London dispersion forces result from fluctuations of electron distribution as neighboring atoms move closer together.
- predict how shape and sizes of molecules or atoms affects the magnitude of dispersion forces.
- explain why polar molecules attract each other through dipole-dipole forces.
- predict the ability of molecules to exhibit hydrogen bonding.
- recognize hydrogen bonding as the force that holds together liquid water, solid ice, and double- stranded DNA.
- explain how ion-dipole attractions lead to dissolution of ions in water & other polar liquids.
- rank a series of molecular compounds with respect to melting or boiling point based on the strength of IMFs present.
- identify how these properties differ between a solid, liquid and gas: density, molar volume, molecular shape, and net strength of IMFs.
- give examples of surface tension, viscosity, and capillary action, and explain how these are caused by cohesive & adhesive IMFs.

- define these phase change processes, and explain how they result from changes in kinetic and potential energy at the molecular level: fusion and freezing, vaporization and condensation, sublimation and deposition.
- explain how both changes in temperature can affect the phase of a substance, based on the flow of energy among the IMFs.
- use a phase diagram to predict the physical state and/or phase changes of a substance at a given pressure and temperature.
- define triple point, critical point, normal boiling point, and normal melting point.
- memorize the normal melting and boiling point of water.
- calculate the total energy change associated with heating a pure substance (like H_2O) through a series of temperature and phase changes (heating curve).
- explain the process of vaporization in terms of dynamic equilibrium, including the effect of temperature, surface area, and the strength IMFs.
- identify the distribution of thermal energies of molecules or atoms of a substance, and how this changes as a function of temperature (Maxwell-Boltzmann Distribution).
- explain how the vapor pressure of a liquid depends on temperature and how this determines the boiling point of a substance.
- define the heat of vaporization, ΔH_{vap}.
- explain how the Clausius-Clapeyron equation relates vapor pressure to temperature & heat of vaporization.
- define solute, solvent, solution, solubility.
- recognize solute and solvent given solution components.
- state the distinctions between saturated, unsaturated, and supersaturated solutions and the equilibrium nature of a saturated solution.
- given a chemical formula, determine the number of particles produced when it is dissolved in water.
- write an equation that represents a compound/molecule dissolving in water.
- state the major types of intermolecular forces in a solution and their relative strengths.
- state the role of intermolecular forces in solution formation.
- describe the energetics of solution formation.
- understand the enthalpy components of a solution cycle and their effect on ΔH_{soln}.
- understand the meaning of entropy and how the it contributes to the solution process.
- understand how the "like dissolves like" rule depends on intermolecular forces.
- predict relative solubilities from intermolecular forces.
- state the trend of pressure dependence on solubility (solid, liquid and gas solutes equally impacted?).
- state the trend of temperature dependence on solubility (solid, liquid and gas solutes equally impacted?).
- use Henry's Law.
- define molarity, molality, mole fraction, mole percent, percent by mass, parts per million and parts per billion.
- calculate molarity, molality, mole fraction, mole percent, percent by mass, parts per million and parts per billion.

- define colligative properties.
- state the four colligative properties and their dependence on the number of dissolved particles.
- approach the four colligative properties from both a qualitative and quantitative approach.
- use Raoult's Law.
- calculate freezing point depression and boiling point elevation for a given solution.
- recognize how the phase diagrams of a solution differs from that of the pure solvent.
- use a colligative property to calculate the molar mass of a solute.
- write a rxn rate in terms of changing reactant or product concentrations over time based on the balanced chemical equation.
- explain the distinction between average and instantaneous rate, and why the instantaneous rate changes during a rxn.
- explain why rxn rate depends on reactant concentration.
- calculate average rates from concentration vs time data, or instantaneous rates from the slope of a tangent to a concentration vs. time plot.
- explain why rate laws cannot be written directly from (overall) balanced chemical equations.
- determine rxn order from a known rate law.
- identify the differences between 0th, 1st and 2nd order rxns.
- explain why concentrations are multiplied in 2nd order (or greater) rate laws.
- explain the experimental basis of the rate law and the information it contains (rxn orders & rate constant).
- calculate rxn order and rate constant from initial rates at different reactant concentrations.
- identify the proper units for a rate constant based on a known rate law.
- explain how the integrated rate laws show the dependence of concentration on time.
- use the integrated rate law to calculate an unknown concentration or solve for time, given concentration values.
- distinguish between 0th, 1st and 2nd order rxns graphically based on the differences between their integrated rate laws.
- explain how the half-life differs between 0th, 1st and 2nd order reactions.
- calculate the half-life of a 0th, 1st or 2nd order rxn based on known concentration values and rate constants.
- define activation energy.
- interpret reaction energy diagrams to quantify energy changes during a rxn.
- explain how the transition state represents the temporary, higher energy species between reactants and products.
- explain the effect of temperature on the rate constant.
- explain the effects of collision alignment & frequency on the rate constant.
- use the Arrhenius equation to calculate the activation energy or the rate constant from a known rate constant.

- explain how an elementary step represents a single molecular event, and that its molecularity equals the number of colliding particles.
- determine the molecularity of a given elementary step.
- explain how a rxn mechanism consists of several elementary steps, with the slowest step determining the overall rate.
- identify the criteria for a valid reaction mechanism.
- predict the rate law that is consistent given reaction mechanism if the slow and fast steps are known.
- explain how a catalyst speeds a reaction by lowering the activation energy.
- distinguish between homogeneous and heterogeneous catalysis.
- distinguish between rate and extent of a reaction.
- explain why a system attains dynamic equilibrium when forward & reverse reaction rates are equal.
- write the equilibrium constant (K) expression as a ratio of rate constants & concentration terms.
- explain why the expression for K is based exactly on the balanced equation as written.
- explain how the magnitude of K refers to the extent of the reaction.
- explain how the same equilibrium condition is reached regardless of the starting concentrations of the reactants.
- convert between K_c and K_p based on the ideal gas law and Δn_{gas}.
- calculate an unknown concentration, pressure, or K value from a reaction table (commonly called an ICE or RICE chart) – including cases where you need to
  - solve a quadratic equation to find an unknown equilibrium quantity.
  - simplify equilibrium calculations by using the assumption that the change in [reactant] is relatively small, when applicable.
- write the reaction quotient (Q) expression from a balanced equation.
- from given concentration values, calculate Q and compare with K to determine the direction a reaction will proceed.
- explain why pure solids & liquids do not appear in the Q or K expressions, and write the correct expression for a heterogeneous chemical equilibrium.
- write Q or K for a reverse reaction or for a reaction multiplied by a factor.
- write the overall Q or K for the sum of several reaction steps.
- explain why a change in temperature affects K.
- explain why addition of a catalyst does not affect K.
- compare Q with K to predict the effect of a change in concentration on the equilibrium position.
- use LeChatelier's principle to explain the effect of a change in concentration, pressure (volume), or temperature on a system at equilibrium.
- state that the proton in an aqueous acid-base reaction exists bound to a water molecule as the hydronium ion, H_3O^+.
- identify a substance as an acid or base using the Bronsted-Lowry, Arrhenius, or Lewis definitions.
- identify conjugate acid base pairs.

- write the chemical equation for dissociation of acids and bases in water alone, showing how water can act as acid or base (amphiprotic) depending on what substance was dissolved.
- explain how the strength of an acid or base relates to the extent of its dissociation equilibrium in water.
- classify acids or bases as strong or weak from their names or formulas.
- rank acids from strongest to weakest based on their acid-dissociation constants, K_a.
- identity that weak bases in water accept a proton rather than dissociate, and rank base strength based on relative K_b values.
- explain how a polyprotic acid dissociates in two or more steps, and why only the first step supplies significant $[H_3O^+]$.
- write the chemical equation for how amines and weak-acid anions act as weak bases in water.
- explain why $[H_3O^+]$ is inversely related to $[OH^-]$ in any aqueous solution based on the autoionization of water.
- explain how the relative magnitudes of $[H_3O^+]$ and $[OH^-]$ define whether a solution is acidic, basic or neutral.
- define the relationship of the K_a and K_b of a conjugate acid-base pair to K_W.
- use K_W to calculate $[H_3O^+]$ and $[OH^-]$ in an aqueous solution.
- interconvert pH, pOH, $[H_3O^+]$ and $[OH^-]$.
- calculate K_a from pH.
- find K_b of A^- from K_a of HA and K_W.
- explain why relative concentrations of HA and A^- determine the acidity of their solution.
- apply the quadratic equation to find unknown acid or base concentrations from known K_a or K_b values.
- calculate $[H_3O^+]$ (and thus pH) from K_a and initial [HA], including for polyprotic acids.
- calculate pH from K_b and initial [B] OR from A^- and $[A^-]$ initial.
- explain how percent dissociation of a weak acid increases as its concentration decreases, based on the equilibrium constant expression.
- calculate the percent dissociation of a weak acid or base using its known K_a or K_b value.
- calculate $[H_3O^+]$ and other concentrations for a polyprotic acid.
- use relative acid strengths to predict the net direction of an acid-base reaction.
- identify the various combinations of cations and anions that lead to acid, basic or neutral salt solution.
- explain why all reactions of a strong acid and a strong base have the same ΔH_{rxn} based on the net ionic equation.
- explain why a stronger acid and base react ($K_c > 1$) to form a weaker base and acid.
- explain the effects of electronegativity, bond polarity, and bond energy on acid strength.
- state why aqueous solutions of small, highly charged metals ions are acidic.
- predict the relative acid strength of nonmetal hydrides and oxoacids.

- predict the relative acidity of a salt solution from the nature of the cation and anion.
- state how the presence of a common ion suppresses a reaction from which is formed.
- state why the concentrations of buffer components must be high to minimize the change in pH from addition of small amounts of H_3O^+ or OH^-.
- state how buffer capacity depends on buffer concentration and on the pK_a of the acid component; why buffer range is within ± 1 pH unit from the pK_a.
- state the nature of an acid-base indicator as a conjugate acid-base pair with differently colored acidic and basic forms.
- state the distinction between equivalence point and end point in an acid-base titration.
- state why the shapes of strong acid-strong base, weak acid-strong base, and strong acid-weak base titration curves differ.
- state how the pH at the equivalence point is determined by the species present; why the pH at the midpoint of the buffer region equals the pK_a of the acid.
- state how the titration curve of a polyprotic acid has a buffer region and equivalence point for each ionizable proton.
- state how a slightly soluble ionic compound reaches equilibrium in water, expressed by an equilibrium (ion product) constant K_{sp}.
- state why incomplete dissociation of an ionic compound leads to approximate values calculated for K_{sp} and solubility.
- state why a common ion in solution decreases the solubility of its compounds.
- state how pH affects the solubility of a compound that contains a weak-acid anion.
- state how precipitate formation depends on the relative values of Q_{sp} and K_{sp}.
- state how complex-ion formation is stepwise and expressed by an overall equilibrium (formation) constant K_f.
- state why addition of a ligand increases the solubility of a compound whose metal ion forms a complex ion.
- state the aqueous chemistry of amphoteric hydroxides involves precipitation, complex-ion, and acid-base equilibria.
- state how selective precipitation and simultaneous equilibria are used to separate ions.
- state how qualitative analysis is used to separate and identify the ions in a mixture.
- use stoichiometry and equilibrium problem-solving techniques to calculate the effect of added H_3O^+ or OH+- on buffer pH.
- use the Henderson-Hasselbalch equation to calculate buffer pH.
- choose the components of a buffer of a given pH and calculate their quantities.
- calculate the pH at any point in an acid-base titration.
- choose an appropriate indicator based on the pH at various points in a titration.
- write K_{sp} expressions for slightly soluble ionic compounds.
- calculate K_{sp} from solubility data and calculate solubility from K_{sp}.
- use K_{sp} values to compare solubilities for compounds with the same total number of ions.
- calculate the decrease in solubility caused by the presence of a common ion.
- predict the effect of H_3O^+ on solubility.
- use ion concentrations to calculate Q_{sp} and compare it with K_{sp} to predict whether a precipitate forms.

- explain that when energy moves in a physical process or chemical reaction, some of it is always lost to the surroundings as heat. For example:
  - The energy released from gasoline combustion is used in part to move a car forward, but most of the energy is dissipated as heat.
  - Energy is lost as heat in charging rechargeable batteries and also in the normal operation (discharge) of the batteries.
  - Heating with natural gas is more efficient than heating with electricity because there are fewer places for the energy to be lost, and heat is the desired product.
  - Reference made to "heat tax" is due to heat loss to the surroundings as well as the inefficiency of real-world processes.
- define spontaneous and explain the difference between spontaneous and nonspontaneous processes.
- explain why spontaneity and speed are not the same.
- explain that spontaneity is not solely dependent on the enthalpy change for a reaction or process. *For instance, some endothermic processes, like ice melting, are spontaneous under certain conditions.*
- define entropy in words.
- define entropy in terms of Boltzmann's equation.
- define precisely what is meant by the terms "order and disorder" when they are used to explain entropy. *For example, we say that the arrangements of molecules or ions becomes more disordered when ice melts, liquid water evaporates, or NaCl dissolves in water. These phenomena occur because the energy is spreading out into more possible places, i.e. entropy is increasing.*
- explain difference between a macrostate and a microstate.
- explain what is meant by a change in entropy, ΔS.
- define the second law of thermodynamics.
- predict the sign of ΔS for phase changes or mixing processes.
- write an equation for ΔS for the universe in terms of ΔS for the system and ΔS for the surroundings.
- explain why exothermic processes *increase* the entropy of the surroundings, while endothermic processes *decrease* it.
- calculate ΔS of the surroundings from ΔH of the system and temperature.
- calculate the temperature at which a given phase change is spontaneous or nonspontaneous.
- calculate ΔS for a chemical reaction from standard thermodynamic quantities.
- define Gibbs free energy.
- explain how the change in Gibbs free energy, ΔG, allows us to predict the spontaneity of a process.
- calculate ΔG and predict spontaneity from given values for T, ΔH, and ΔS.
- calculate ΔG from standard thermodynamic quantities.
- understand the difference between standard and non-standard conditions.
- explain why some product forms in an equilibrium even when a reaction is non-spontaneous.

- use Le Chatlier's principle to explain why non-spontaneous reactions can be driven to completion.
- explain how the value of the equilibrium constant is related to the sign of the standard Gibbs free energy change and determine whether a reaction is spontaneous from its equilibrium constant, K.
- relate the standard Gibbs free energy to the equilibrium constant, K.
- relate the Gibbs free energy of reaction to the reaction quotient, Q.
- calculate the Gibbs free energy of a reaction under non-standard conditions from Q, the reaction quotient.
- calculate the equilibrium constant
  - from the standard enthalpy and entropy changes of the reaction at a given temperature
  - at a new temperature from the known equilibrium constant at another temperature