1st Semester Learning Targets

Equilibrium
(1) State Le Châtelier's principle
(2) Predict the shift of an equilibrium from changing concentration, temperature, and pressure/volume
(3) Write an equilibrium constant (K) expression for an equilibrium reaction.
(4) Predict the effect on an equilibrium constant of temperature change for an endothermic and exothermic reaction.
(5) Calculate the value of K_c from the concentrations of chemicals at equilibrium.
(6) Calculate the value of the reaction Quotient (Q) from the initial concentrations of chemicals and predict the shift by comparing Q to K.
(7) Solve an ICE table problem using initial and equilibrium concentrations of chemicals.
(8) Solve an ICE table problem using algebra or approximation. Check the validity of any assumptions used in calculations.
(9) Calculate the value of K_p from the pressures of chemicals at equilibrium.
(10) Solve problems involving mole fraction and partial pressure.
(11) Solve an ICE table problem using initial and equilibrium partial pressures of chemicals, algebra, or approximation.
(12) Convert between K_c and K_p for an equilibrium at a given temperature.
(13) Manipulate an equilibrium constant value for an equilibrium that has been reversed and/or multiplied by a constant. Determine the equilibrium constant for a reaction resulting from adding two equilibria.

Solubility Equilibrium
(1) Write a solubility product constant expression (K_{sp}) for a solubility equilibrium.
(2) Solve an ICE table problem using solubility data.
       (a) Calculate K_{sp} from the solubility of a salt
       (b) Calculate the solubility of a salt from K_{sp}
(3) Rank salts in order of solubility based on K_{sp} values OR with ICE table calculations.
(4) Describe the common ion effect in terms of Le Châtelier's Principle.   Solve ICE table problems involving the common ion effect.
(5) Determine if a precipitate will form when two solutions are mixed by calculating a Q value from the diluted concentrations of ions in solution. Calculate the final concentration of ions in solution if a precipitate does form.
(6) Explain the two ways that ions can be selectively precipitated from one another in solution.
(7) Recognize a complex ion and identity the ligand and state the coordination number for a given compound. Predict the coordination number for a complex ion based on the charge of the central metal ion. Write a balanced chemical equation for the formation of a complex ion.

Acid Base Equilibria
(1) Describe an acid/base according to the Arrhenius definition and give examples of an Arrhenius acid and an Arrhenius base.
(2) List the eight strong acids and the eight strong bases.
(3) Calculate the pH, pOH, $[H^+]$, and $[OH^-]$ for a strong acid or base.
(4) Describe an acid/base according to the Bronsted/Lowry definition and give examples of a Bronsted/Lowry acid and a Bronsted/Lowry base
(5) Write an acid or base equilibrium reaction and identify the conjugate acid base pairs.
(6) Write an acid dissociation constant (K_a) expression for an acid equilibrium and a base dissociation constant (K_b) expression for a base equilibrium.
(7) Compare the relative strength of acids based on their K_a value; compare the relative strength of bases based on their K_b values.
(8) Solve an ICE table problem for an acid or base.
       (a) Calculate pH, pOH, and/or percent dissociation from K_a or K_b
       (b) Calculate K_a or K_b from pH, pOH, or percent dissociation
(9) Convert between K_a and K_b for conjugate acid base pairs by applying $K_w = K_a \times K_b$
(10) Define the term amphoteric and give two equations to show the amphoteric nature of a compound. Calculate K_a or K_b value for each reaction and predict which of the two reactions will dominate in solution (is the solution acidic or basic?).
(11) Write an equilibrium equation for the reaction of a weak acid with a weak base and calculate an equilibrium constant for the equation from K_a and K_b values.
(12) Calculate the pH and pOH of mixtures.
       (a) strong acid with a weak acid OR a strong base with a weak base
       (b) two weak acids OR two weak bases
       (c) a weak acid with a common ion of the conjugate base OR a weak base with a common ion of the conjugate acid
(13) Write successive ionization reactions for diprotic acids. Calculate the equilibrium constant for the complete ionization reaction of a diprotic acid.
(14) Determine if a cation or anion will hydrolyze in order to predict if a salt will be acidic, basic, or neutral. Calculate the pH and pOH for a salt.
(15) Define the role of a buffer and identify pairs of compounds that can act as a buffer in solution.
(16) Explain how a buffer works using Le Châtelier's Principle.
(17) Calculate the pH and pOH for a buffer solution from the concentration of each component. Calculate the pH and pOH that results from adding strong acid and strong base to a buffer solution.
(18) Calculate the pH or pOH of a solution using the Henderson-Hasselbalch Equation.

(19) Perform calculations to determine the mass of a salt required to prepare a buffer with a given pH or pOH
(20) Give the net ionic equation for a neutralization reaction taking place between and acid and a base.
(21) Perform calculations to determine the concentration OR volume of a base required to titrate an acid OR an acid required to titrate a base.
(22) Calculate the pH for adding a given volume of solution for different titration experiments.
> (a) strong acid to strong base
> (b) strong base to strong acid
> (c) strong acid to weak base
> (d) strong base to weak acid

(23) Draw a titration curve from pH values at different volumes of solution added.
(24) Determine K_a for a weak acid from the titration curve with strong base from the initial pH and the volume halfway to the equivalence point.
(25) Compare titration curves for different titration experiments in terms of initial pH, height of the curve, pH at the equivalence point, and final pH.
> (a) strong acid with strong base
> (b) strong base with strong acid
> (c) strong acid with weak base
> (d) strong base with weak acid

(26) Select an appropriate acid base indicator for a titration experiment based on the pH at the equivalence point.
(27) Distinguish between the endpoint and the equivalence point of a titration experiment.
(28) Describe an acid/base according to the Lewis definition and give examples of a Lewis acid and a Lewis Base.
(29) Identify the Lewis acid and Lewis base in a reaction.

Electrochemistry
(1) Assign oxidation numbers to each element in a compound/ion.
(2) Define oxidation, reduction, oxidizing agent, and reducing agent.
(3) Determine if a reaction is a redox reaction from the oxidation numbers of each element. Identify the chemical being oxidized and the chemical being reduced.
(4) Identify the oxidizing and reducing agents in a redox reaction.
(5) Balance a redox reaction in both acid and base using the half reaction method.
(6) Identify and balance a disproportionation reaction.
(7) Perform calculations involving redox titrations including finding the concentration OR volume of a solution required and determining the percent purity of a compound.
(8) Predict the products in a redox reaction and write a balanced net ionic equation for the reaction.
(9) Rank the relative strength of oxidizing and reducing agents from the voltage of each half reaction found on a table of standard reduction potentials.
(10) Calculate the voltage of a reaction by determining the voltage of each half reaction from a table of standard reduction potentials.
(11) Determine if a reaction will be spontaneous or non spontaneous based on the overall voltage.
(12) Determine the overall reaction in an electrochemical cell and calculate the overall voltage.
(13) Describe the functions of each component of an electrochemical cell including the anode, cathode, voltmeter, and salt bridge and identify each on a diagram.
(14) Determine the direction of electron flow and ion migration for an electrochemical cell.
(15) Describe how the mass of the anode and the cathode change as an electrochemical cell operates.
(16) Give the line notation for an electrochemical cell.
(17) Determine the overall reaction for an electrolytic cell and the minimum voltage that must be applied to operate the cell.
(18) Perform current calculations for an electrochemical cell to relate the mass of a substance deposited on an electrode in a given amount of time for the level of current applied.

Nuclear Chemistry
(1) Write nuclide symbols for a given isotope.
(2) Describe alpha, beta, and gamma radiation and give the appropriate symbol for each.
(3) Define a transmutation and state what types of radiation can lead to a transmutation.
(4) Define penetrating power and rank alpha, beta, and gamma radiation according to their strength.
(5) Complete nuclear reactions including those involving alpha, beta or gamma radiations as well as neutrons and protons
(6) Define half-life.
(7) Perform calculations involving half-life in order to solve for mass, time, original mass, and half life.

Organic Chemistry
(1) Name and draw hydrocarbons including alkanes, cycloalkanes, alkenes, and alkynes
(2) Name and draw molecules with functional groups including alkyl halides, alcohols, ethers, aldehydes, ketones, carboxylic acids, esters, amines, and amides.
(3) Complete ester condensation reactions and combustion reactions.

Thermochemistry

(1) Define temperature and heat and give appropriate units for each. Relate temperature to the amount of kinetic energy and describe heat transfer in terms of kinetic energy.

(2) Describe and graph the temperature changes for a heating or cooling curve, and label each part of the curve with the appropriate phase(s). Determine the melting/freezing point and boiling/condensing point from a heating or cooling curve.

(3) Describe how the bonds between particles change during phase changes and identify endothermic vs. exothermic processes. Define heat of vaporization and heat of fusion and give the appropriate units for each. Perform calculations involving heat of vaporization and heat of fusion for phase changes of substances.

(4) Describe how the speed of particles changes during heating/cooling. Define specific heat capacity and give the appropriate units. Perform calculations involving specific heat capacity for heating and cooling of substances.

(5) Calculate the total amount of energy from the heat and work. Identify an endothermic process as one that involves a transfer of heat energy from the surroundings to the system (q is positive) and an exothermic process as one that involves a transfer of heat energy from the system to the surroundings (q is negative). Calculate work for the expansion of a gas (where the system does work on the surroundings, w is negative) or compression of a gas (where the surroundings do work on a system, w is positive).

(6) Use data from a calorimetry experiment to determine the heat change of a process/reaction. Predict the expected temperature change in a calorimetry experiment; for an endothermic process/reaction, the temperature of the water in a calorimeter will decrease and for an exothermic process/reaction, the temperature of the water in a calorimeter will increase. Identify the main source of error in calorimetry experiments.

(7) Define enthalpy/heat of reaction and give the appropriate units. Discuss the transfer in potential energy and kinetic energy for an exothermic and endothermic reaction.

(8) Determine if a reaction is exothermic or endothermic from the chemical equation, ΔH value, or potential energy diagram. Determine the heat change given the moles or mass of a reactant or product.

(9) Calculate the heat of a reaction from tabulated bond energies. Calculate a bond energy from the heat of a reaction.

(10) Calculate the heat of a reaction from tabulated heats of formation. Calculate a heat of formation for a reactant or product from the heat of a reaction.

(11) Calculate the heat of a reaction from Hess' Law.

(12) Define entropy and give the appropriate units. Predict if entropy increases or decreases when temperature changes (entropy increases when temperature is increased and decreases when temperature is decreases) or during a phase change (entropy increases during melting and boiling and decreases during freezing and condensing). Relate the entropy value of a substance to its phases (solids have the lowest entropy and gases have the highest entropy). Predict if entropy increases or decreases in a reaction (generally, the side of the reaction with more gas molecules will have a higher entropy).

(13) Calculate the entropy change of a reaction from tabulated standard entropy values. Calculate the standard entropy of a reactant or product from the entropy change of a reaction.

(14) Define free energy (the energy available to do useful work) and give the appropriate units.

(15) Calculate free energy from tabulated free energies of formation (and calculate a standard free energy for a reactant or product from the free energy change of a reaction) and the Hess' Law method.

(16) Determine if a reaction is spontaneous (thermodynamically favoured) based on the sign of the free energy change (if the free energy change for a reaction is negative, it is thermodynamically favoured; of the free energy change for a reaction is positive, it is not thermodynamically favoured).

(17) Calculate free energy from enthalpy, entropy, and temperature. Predict the spontaneity of a reaction from the sign of enthalpy and entropy. Calculate the temperature conditions at which a reaction could be spontaneous.

(18) Convert between the free energy of a reaction and the value of the equilibrium constant. Relate the sign of the free energy change to the magnitude of the equilibrium constant (a negative free energy change corresponds to an equilibrium constant greater than one; a positive free energy change corresponds to an equilibrium constant less than one).

(19) Calculate a reaction quotient (from given pressures or concentrations) for a reaction that is not at equilibrium and determine the free energy for the system.

(20) Convert between the free energy change of an electrochemical reaction and the value of standard potential.

(21) Relate the sign of the free energy change of an electrochemical reaction to the sign of the standard potential (a negative free energy change corresponds to a positive standard reduction potential; a positive free energy change corresponds to a negative standard reduction potential).

Reaction Kinetics

(1) List two factors required for a chemical reaction to occur according to collision theory (sufficient energy, correct geometry).

(2) Describe the factors that affect the rate of a chemical reaction including concentration, temperature, surface area, and catalyst and relate these factors to collision theory.

(3) Determine the rate of a reaction from experimental data/graphs. Relate the rate of reaction of one substance in a reaction to rate of reaction of another substance in a reaction based on the coefficients in the balanced chemical equation. Describe how reaction rate changes as a reaction progresses.

(4) Perform calculations to determine a differential rate law from experimental data using the method of initial rates and calculate the value and give the units of the rate constant. Use a differential rate law to calculate the rate of a reaction from the initial concentration of each reactant. Determine the overall order for a reaction from a differential rate law. Predict the effect on reaction rate of changing the concentration of a reactant given the differential rate law equation.

(5) Determine the order of a reaction and the appropriate integrated rate law from experimental data using graphical methods for first, second, and zeroth order reactions. Determine the half-life for a first, second, and zeroth order reaction. Use an integrated rate law to determine the concentration of a chemical at a given time and the time required for the concentration to equal a given value.

(6) Classify nuclear reactions as first order reactions. Use a given half-life to determine the specific rate constant. Perform calculations using the specific rate constant, to determine the percent of nuclear material remaining after a given amount of time or the amount of time for a given percent of nuclear material to remain.

(7) Determine the rate constant for a reaction at a given temperature from the activation energy and frequency factor of the reaction. Describe how the rate constant for a reaction changes with temperature (increase temperature gives a higher rate constant).

(8) Interpret a potential energy diagram to determine heat of reaction and activation energy for forward and reverse reactions as well as catalyzed and uncatalyzed reactions.

(9) Recognize that reactions often occur through a series of steps called a reaction mechanism. Determine the overall reaction given the steps in a reaction mechanism or find a missing step given the overall reaction.

(10) Identify reaction intermediates and catalysts in a reaction mechanism. Distinguish the rate-determining step in a reaction mechanisms as the slowest step.

(11) Sketch a potential energy diagram given the steps in a reaction mechanism and the enthalpy of the overall reaction. Distinguish the rate-determining step on a potential energy diagram as the step with highest activation energy.

(12) Predict the rate law for a reaction given the steps in a reaction mechanism with the rate-determining step identified or identify the rate-determining step from a proposed rate law.

(13) Define a catalyst and describe its role in increasing the rate of a reaction. Identify three classes of catalysts (acid-base, surface, biological) and give examples of each.

Quantum Mechanics, Periodic Trends, and Bonding

(1) Define and give examples of electromagnetic radiation (radio waves, infrared, visible, UV, X-rays, and gamma rays).

(2) Define wavelength, frequency, and energy and give the appropriate units for each quantity. Perform calculations involving wavelength, frequency, and energy.

(3) Describe the experiment used to show the photoelectric effect and the significance of the findings by defining a photon.

(4) Define the quantum numbers: principal quantum number, angular momentum quantum number, magnetic quantum number, and spin quantum number and the atomic properties they define.

(6) Give the orbital notation for an atom/ion by applying the Aufbau Principle, Hund's Rule, and the Pauli Exclusion Principle. Write the electronic configuration for an atom/ion. Write the noble gas configuration for an atom/ion.

(7) Interpret a photoelectron spectrum (PES) to compare the energy of different electrons in an atom and the relative number of electrons in each orbital. Compare the PES spectra for two different elements.

(8) Use Coulomb's law to calculate the electrostatic force between two charged particles from the values of the charges and the distance separating them. Use Coulomb's law to explain periodic trends as they relate to atomic structure in that a stronger force will result from higher charges or a smaller distance between them.

(9) Define atomic radius and explain periodic trends in this property as they relate to atomic structure.

(10) Define ionization energy and explain periodic trends in this property as they relate to atomic structure. List elements that are exceptions to the general periodic trend (i.e. B, O, Al, and S) and use electron configuration to show why they are exceptions. Write equations to show successive ionizations for a given element and explain the trends observed for each successive ionization and predict where the largest jump in ionization will be observed for a given element (using the number of valence electrons for an element).

(13) Define electronegativity and explain periodic trends in this property as they relate to atomic structure.

(14) Define ionic radius and relate the size of an anion to a neutral atom of the same element (an anion is larger) and a cation to a neutral atom of the same element (a cation is smaller). Compare the size of ions for elements in the same group and for elements in the same row (for isoelectronic ions).

(15) Draw Lewis structures for molecules and ions. Recognize elements that can be electron deficient and those that have an expanded octet (recall, that elements in the third row and below can accommodate more than eight electrons).

(16) Assign bond orders for a molecule from the Lewis structure.

(17) Calculate (and label) the formal charge on each atom for an ion. Draw possible resonance structures for an ion.

(18) Use formal charge to determine the most favoured Lewis structure for a molecule/ion.

(19) Calculate the electronegativity difference between two atoms. Classify a bond as covalent (0-1.7) or ionic (1.8^+) and as non-polar (0-0.4) or polar (0.5^+) based on the result.

(20) Draw VSEPR diagrams for molecules and ions and give the name the VSEPR shape. Determine the approximate bond angle(s) in a molecule based on its VSEPR shape. Relate the bond angle the number of lone pairs on the central atom (generally the more lone pairs, the smaller the bond angle). Predict if a molecule is polar or non-polar based on VSEPR shape. Define dipole moment and give the appropriate units for this quantity. Determine the steric number and hybridization of an atom based on its molecular geometry.

(21) Discuss the basis for the formation of sigma and pi bonds and relate to bond order. Determine the total number of sigma bonds and pi bonds in a molecule from its Lewis structure.

(22) Discuss the relationship between atomic radius and bond length (larger atoms result in longer bonds). Discuss the relationship between orbital overlap and bond energy (greater overlap produces higher bond energy). Interpret a bond energy profile to determine the bond length and bond energy.

(23) Discuss the relationship between bond order and bond energy/length (higher bond order produces shorter bonds with higher energy).

(24) Relate the strength of a binary acid to the bond strength (stronger bonds result in a weaker acid). Relate the strength of an oxyacid to the electronegativity of elements (higher electronegativity results in a stronger acid) and the number of oxygen atoms (more oxygen atoms results in a stronger acid).

(25) Describe common spectroscopy techniques including infrared spectroscopy (used to detect the presence of different types of bonds in a molecule) and UV/Vis spectroscopy (used to detect transitions in electronic energy levels as well as to detect the concentration of a substance in solution using Beer's Law).

Intermolecular Forces, Phases, and Solutions

(1) Identify the type of intermolecular force(s) present between atoms/molecule as London dispersion forces (found between all molecules but is the only IMF for non-polar molecules), dipole interactions (between polar molecules), dipole-induced dipole interactions (between a polar molecule and a non-polar molecule), ion-dipole interactions (between an ion and a polar molecule), and hydrogen bonds (for molecules with hydrogen bonded to nitrogen, oxygen, or fluorine).

(2) Predict relative the strength of London dispersion forces between two atoms based on their size (larger atoms are more polarizable and have stronger LDF's). Predict the relative strength of London dispersion forces between two molecules based on their size (i.e. molar mass/number of electrons- larger molecules have stronger LDF's).

(3) Predict the relative strength of dipole interactions between two molecules based on the magnitude of the dipole moment (a higher dipole moment results in a stronger dipole interaction).

(4) Predict the relative strength of dipole-induced dipole interactions based on the dipole moment of the polar molecule and the polarizability of the non-polar molecule (a polar molecule with a higher dipole moment with more polarizable non-polar molecule results in a stronger dipole-induced dipole interaction).

(5) Predict the relative strength of an ion-dipole interaction based on the size and charge of the ion (a higher charged ion and a smaller ion results in a stronger ion-dipole interaction).

(6) Compare the relative strength of London dispersion forces, dipole interactions, and hydrogen bonds and relate to boiling point/heat of vapourization (H-bonds are stronger than dipole which are stronger than LDF's; stronger IMF's give a higher boiling point/heat of vapourization).

(7) Relate intermolecular forces to solubility (molecules that can form H-bonds with water will be more soluble).

(8) Compare the three phases/states of matter (solid, liquid, and gas) in terms of molecular speed and molecular spacing, and be able to draw/identify a simple diagram for each state

(9) Classify a solid as crystalline or amorphous. Classify crystalline solids as atomic, molecular, or ionic. Classify atomic solids as metallic or network. Describe the distinguishing properties of each type of solid and give examples for each.

(10) Discuss two factors (charge and size of the ions) related to the magnitude of lattice energy for an ionic compound. Relate lattice energy to melting point (higher lattice energy gives a higher melting point).

(11) Distinguish between substitutional alloys and interstitial alloys based on their composition and properties. Give examples for each type of alloy and draw a diagram to represent each type of alloy.

(12) Define surface tension, viscosity, and vapour pressure and relate the these properties in a liquid to the strength of the intermolecular forces present. Explain what happens to vapour pressure when temperature is increased.

(13) Calculate the root mean square velocity (i.e. speed) of a gas molecule given the temperature.

(14) Solve problems involving Graham's law for the relative rate of effusion of two gases.

(15) Describe the experimental procedure for gas collection over water and perform stoichiometric calculations from experimental data.

(16) Write the net ionic equation for the formation of common gases in a chemical reaction i.e. H_2S, CO_2 (from H_2CO_3), SO_2 (from H_2SO_3), and NH_3 (from NH_4OH).

(17) List the assumptions of the Ideal gas Law and describe how the real behaviour of a gas differs from ideal behaviour. Recognize that non-ideal behaviour is more apparent for polar molecules (since they have stronger IMF's) and for gases at low temperatures and high pressures.

(18) Define evaporation/boiling, condensation, melting, freezing, sublimation, and deposition by stating the starting and ending phase for each change. Give the sign of enthalpy and entropy for each of these changes and predict the temperature conditions at which each phase change would be thermodyamically favoured.

(19) Express the concentration of a solution in terms of molarity. Describe how to prepare a solution from a solid or by diluting a solution of a higher concentration (including the safe procedure for diluting an acid).

(20) List the factors than can affect solubility (structure/polarity, temperature, and pressure). Describe the effect each factor has on solubility (i.e. "like dissolves like", increasing temperature increases the solubility of a solid and decreases the solubility of a gas, increasing pressure increases the solubility of a gas).

(21) Relate the solubility of a substance to intermolecular forces and thermochemistry.

(22) Classify a solute as an electrolyte or a non-electrolyte. Classify an electrolyte as a strong electrolyte or a weak electrolyte. Describe how electrolytes and non-lectrolytes can be distinguished from one another experimentally.

(23) Describe the basis for chromatography experiments in terms of affinity between the solutes in a solution and solvent. Calculate the retention factor (R_f) value for a solute from experimental data.

(24) Describe a distillation experiment that could be used to separate liquid components of a solution based on their differing boiling points.