IGC	Learning Outcomes
1.1	**Describe properties of states of matter.**
1.1	Identify states of matter.
7.1	Distinguish between macroscopic properties and the particulate view of states of matter.
12.1	**Identify the types of forces present between like molecules or between different substances.**
12.1	Distinguish between intermolecular (IMF) and intramolecular forces.
12.1	Define dispersion (London) forces, dipole-dipole forces, hydrogen bonding, and ion-dipole forces and what interactions occur for each type of of intermolecular force.
12.2	**Rank properties of substances based on differences in intermolecular forces.**
12.2	Describe how molar mass, molecular shape, and polarity affect the magnitude of intermolecular forces.
12.2	Rank substances in order of increasing or decreasing melting point and boiling point.
12.2	Define surface tension.
12.2	Rank substances in order of increasing surface tension with respect to molar mass, shape, temperature, and intermolecular forces.
12.2	Define viscosity.
12.2	Rank substances in order of increasing viscosity with respect to molar mass, shape, temperature, and intermolecular forces.
12.3	**Explain the different types of phase changes and associated energy changes.**
12.3	Describe the phase changes that occur in fusion, freezing, vaporization, condensation, sublimation, and deposition.
12.3	Define melting point and enthalpy (heat) of fusion.
	Define boiling point and enthalpy (heat) of vaporization.
12.3	Describe the relationship between change in enthalpies (heats) of freezing and fusion, as well as enthalpies (heats) of vaporization and condensation.
12.3	Explain the relative magnitudes of the enthalpies associated with each phase change.
12.3	**Complete calculations associated with phase changes at the melting or boiling point.**
12.3	**Draw and complete calculations for a heating or cooling curve.**
12.3	Calculate the total heat change, with the correct units, based on a heating or cooling curve.
12.4	**Describe properties related to vaporization**
12.4	Distinguish between volatile and nonvolatile.
12.4	Use particulate level models to explain how and why the rate of vaporization changes with temperature, surface areas, and intermolecular forces.
12.4	**Define dynamic equilibrium and vapor pressure.**
12.4	Explain how dynamic equilibrium and vapor pressure are affected by changes in pressure as a result of changes in volume.
12.4	Distinguish between boiling point and normal boiling point.
12.4	Use particulate level modeling to explain why and how intermolecular forces and temperature affect vapor pressure.
12.4	**Use the Clausius-Clapeyron equaiton to calculate vapor pressure, temperature, or enthalpy of vaporizaiton.**
12.4	Use the two-point Clausius-Clapeyron equation to calculate the enathalpy of vaporization, temperature, or vapor pressure for a change in conditions.
12.4	Use graphical data to determine the relationship between temperature, pressure, and enthalpy of vaporization.
12.5	**Explain the information provided in a phase diagram.**
12.5	Identify the variables on the x- and y-axes of a phase diagram.
12.5	Describe the stable phase in each region of a phase diagram.
12.5	Identify the phase changes indicated by the curves in the phase diagram.
12.5	Explain the meaning of the slope of the fusion curve in a phase diagram and its implications on the density of the solid vs the liquid.
12.5	Locate the triple point on a phase diagram.
12.5	Recognize the phases in equilibrium on the curves and the triple point of the phase diagram.
12.5	Locate the critical point, and describe the critical temperature and pressure.
12.5	Define a supercritical fluid.
12.6	**Identify a solid as crystalline or amorphous.**
12.6	**Classify crystalline solids as molecular, ionic, nonbonding, metallic, or network covalent solids.**
12.6	Distinguish bewteen the physical properties of each type of crystalline solid.

12.6	Describe the bonding between ions, molecules, or atoms for each type of crystalline solid.
12.7	**Describe the geometric properties of cubic unit cells.**
12.7	Define crystalline lattice and unit cell.
12.7	Calculate the number of atoms per cell in simple, body-centered, and face-centered cubic cells.
12.7	**Explain the qualitative and quantitative properties of a unit cell.**
12.7	Distinguish among body, edge, face, and corner positions for atoms in a unit cell.
12.7	State the fraction of each atom inside the unit cell for body, edge, face, and corner positions.
12.7	Use given unit cell information to calculate the dimensions of the cell, density, and/or molar mass of the substance.
12.7	Determine the ionic formula of a compound from unit cell information.
13.1	**Identify solutions and their components.**
13.1	Define solute, solvent, and solution.
13.1	Recognize that solutions may be in liquid, solid, or gas phases, and that multiple phases may be present.
13.1	Recognize that water is the solvent in aqueous solutions.
13.1	**Describe how the relative strengths of the interactions among molecules determine whether or not a solution will form.**
13.1	Define solvent-solute, solvent-solvent, and solute-solute interactions.
13.1	Explain the effect of intermolecular forces on solution formation.
13.1	Describe the enthalpies of solute, solvent, solution, and mixing and classify as endothermic or exothermic.
13.1	Define entropy simplistically as a measure of randomness or disorder in a system. Note: a more-detailed definition will be given in Chapter 18.
13.1	Explain that an increase in entropy is favorable for solution formation even when the enthalpy of solution is positive.
13.2	**Describe the properties of a solution.**
13.2	Distinguish between miscible and immiscible substances.
13.2	Distinguish among saturated, unsaturated, and supersaturated solutions.
13.2	Explain how to prepare a supersaturated solution.
13.2	**Provide qualitative and quantitative explanations of how temperature affects solubility.**
13.2	Recognize that temperature affects the solubility of solids.
13.2	Use a solubility curve and information about a solution to identify a solution as saturated, unsaturated, or supersaturated.
13.2	Describe how temperature and pressure affect the solubility of a gases in a liquid.
13.2	Use Henry's law to calculate the Henry's law constant, pressure, or solubility of a gas in a liquid.
13.2	Use Henry's law to calculate solubility or pressure when one condition is varied.
13.3	**Recognize that concentration can be expressed in a variety of units.**
13.3	Define molarity, molality, mole fraction, parts by mass (percent, ppm, ppb), and parts by volume (percent, ppm, ppb).
13.3	Use the definitions of the concentration units to determine concentrations, amount of solute, solvent, or solution.
13.3	Convert between any two units of concentration.
13.4	**Describe qualitative and quantitative effects of concentration on solution properties.**
13.4	Describe colligative properties.
13.4	Define vapor pressure lowering, freezing point depression, boiling point elevation, and osmotic pressure.
13.4	Explain how each colligative property is affected by the concentration of the solute in a solution.
13.4	Use Raoult's law to calculate vapor pressure of the solution, vapor pressure of the pure solvent, mole fraction of the solvent (or solute), or change in vapor pressure.
13.4	Recognize that both solute and solvent can be volatile.
13.4	Use the boiling point elevation or freezing point depression equation to solve for any variable, or the
13.4	**Describe the flow of solvent particles during osmosis.**
13.4	Describe a semipermeable membrane.
13.4	Describe differences between isosmotic (isotonic), hyperosmotic (hypertonic), and hyposmotic (hypotonic) solution.
13.4	Use the osmotic pressure equation to solve for any variable.
13.4	Calculate the molar mass of a compound using colligative property information.
13.5	**Complete calculations involving for electrolyte solutions.**

13.5	Define van't Hoff factor.
13.5	Determine the ideal van't Hoff factor for a substance.
13.5	Describe the ion pairing that leads to non-ideal van't Hoff factors.
14.1	**Calculate the rate of reaction or change in concentration of a substance from experimental data.**
14.1	Define rate of reaction.
14.2	Distinguish between instantaneous and average rate of reaction.
14.1	Write the overall rate of reaction in terms of change in concentration of each reactant or product per change in time.
14.2	**Write the rate law for and determine the order of a reaction.**
14.2	Define rate constant and rate law.
14.2	Write the rate law based on experimental data for reactant concentration and reaction rate.
14.2	Identify the overall order of the reaction and the order with respect to each reactant from the rate law.
14.2	Calculate the rate constant based on experimental data and the rate law.
14.2	Calculate relative rates based on changes in concentration (i.e. twice A or half of B).
14.2	Label a graph of rate vs reactant concentration as zero, first, or second order.
14.3	**Complete calculations using the integrated rate laws for zero-, first-, and second-order reactions.**
14.3	Define integrated rate law.
14.3	Use the integrated rate laws (equations) for zero-, first-, and second-order reactions to solve for any variable.
14.3	Understand and explain why the units of the rate constant are different for different reaction orders.
14.2	Label a graph of reactant concentration vs time as zero, first or second order.
14.3	Recognize the linear relationship for plotting concentration data vs. time for each reaction order.
14.3	Complete calculations using the half-life equations for zero-, first-, and second-order reactions.
14.4	**Explain concepts of collision theory.**
14.4	Understand that a chemical reaction can occur when reactants collide with one another with property energy and orientation.
14.4	Define activation energy, activated complex, and transition state.
14.4	Describe how temperature affects the rate of a reaction.
14.4	Recognize that the value of the rate constant is dependent on temperature.
14.4	Label activation energies of each step, transition states, reactants, products, and enthalpy of reaction in an energy diagram (energy vs reaction progress).
14.4	**Perform calculations using the Arrhenius equation.**
14.4	Use Arrhenius plot (ln k vs. 1/T) to calculate the activation energy or freqency factor.
14.4	Complete calculations using the two-point Arrhenius equation to solve any variable.
14.5	**Explain reaction mechanisms.**
14.5	Define reaction mechanism, elementary step, and reaction intermediate.
14.5	Identify reaction intermediates based on the reaction mechanism.
14.5	Write rate laws for elementary steps.
14.5	Define molecularity, unimolecular, bimolecular, and termolecular collisions.
14.5	Define rate-determining step.
14.5	Identify the rate-determining step in a mechanism.
14.5	Write the overall rate law for mechanisms when the first step is slow.
14.5	Recognize that when the second or later step in the mechanism is slow, the overall rate law must be derived algebraically.
14.5	Describe the two conditions for a valid reaction mechanism: the overall rate law must be consistent with experimental data, and the elementary steps must sum to the overall reaction.
14.5	**Describe the effect of a catalyst on the energetics of a reaction and the reaction rate.**
14.5	Define catalyst.
14.5	Identify the catalyst in a reaction mechanism.
14.5	Distinguish between homogeneous and heterogeneous catalysts.
14.5	Describe the function of an enzyme as a biological catalyst.
15.1	**Define and explain the conditions of chemical dynamic equilibrium.**
15.1	Explain what is constant for a reaction at equilibrium.

15.1	Explain how the concentrations of reactants and products change with time as a reaction achieves equilibrium.
15.2	**Write the equilibrium constant expression (law of mass action) for an equilibrium reaction in terms of concentrations or partial pressures and explain its relevance.**
15.2	Define equilibrium constant, K.
15.2	Identify whether products, reactants, or neither are favored at equilibrium based on the magnitude of K.
15.2	Calculate K_C or K_P using the equilibrium constant expression and equilibrium quantities, recognizing that
15.2	Use the equation to convert between K_C or K_P recognizing that Δn represents the change in the moles of gas for the
15.2	**Write the equilbirum constant expression (law of mass action) for a heterogeneous system.**
15.2	Identify an equilbrium system that is heterogeneous.
15.2	Recognize that pure liquids and solids are excluded from the equilibrium calculation.
15.2	Calculate K_C or K_P for a heterogenous system.
15.3	**Calculate the equilbrium constant when reactions are combined or manipulated.**
15.3	Calculate K for the reverse reaction when given K of the forward reaction.
15.3	Calculate K when the coefficients of a reaction are multiplied by a factor given the original value of K.
15.3	Calculate K when two or more equations are combined given K for the individual equations.
15.4	**Predict the direction of a reaction based on comparison of Q and K.**
15.4	Define reaction quotient, Q.
15.4	Calculate reaction quotient in terms of concentration or partial pressure.
15.5	**Calculate equilibrium concentrations.**
15.5	Determine the missing equilibrium concentration (or partial pressure) when given K and all but one equilibrium concentration (or partial pressure).
15.5	Explain how to use an ICE (initial-change-equilibrium) table.
15.5	Calculate equilibrium concentrations given initial concentrations and a single equilibrium concentration or a change
15.5	Calculate equilibrium concentrations given only initial concentrations using a variety of problem solving techniques (perfect squares, the quadratic equation, or by using a simplifying assumption.)
15.5	Describe when a simplifying assumption can be made (x is small).
15.6	**Predict the shift in a reaction according to Le Chatelier's principle.**
15.6	Define Le Chatelier's principle.
15.6	Explain what is meant by a "shift" in a reaction (i.e., which direction a reaction proceeds).
15.6	Predict the shift in a reaction due to change in concentration of a reactant or product.
15.6	Describe how the concentration of each substance changes as the shift occurs.
15.6	Recognize that changing the amount of solids and pure liquids does not affect the reaction.
15.6	Predict the shift in a reaction due to a change volume of the reaction container (which results in a change in pressure) for reactions involving gases.
15.6	Recognize that the addition of an inert gas does not result in a shift in the equilibrium.
15.6	Predict the shift in a reaction due to a change in temperature for both endothermic and exothermic reactions.
15.6	Predict whether K will increase or decrease with a change in temperature.
16.1, 16.2	**Describe Arrhenius and Bronsted-Lowry acids and bases and how they behave in solution.**
16.1	List the general properties of acids and bases.
16.1	Define Arrhenius acids and bases.
16.1	Write equations to represent the behavior of Arrhenius acids and bases in aqueous solution.
16.2	Define Bronsted-Lowry acids and bases.
16.2	Label substances in a reaction as Bronsted-Lowry acids and bases.
16.2	Define conjugate acid and conjugate base.
16.2	Generate the formulas for conjugate acid-base pairs and recognize them in a reaction.
16.1	**Distinguish between strong and weak acids and strong and weak bases.**
16.1	Memorize the strong acids and strong bases.
16.1	Recognize that all other acids and bases are weak.
16.1	Write an equation representing the behavior of a weak acid or a weak base in water.
16.3	**Explain the meaning of K_w.**
16.3	Define autoionization and write the reaction for the autoionization of water.
16.3	Write the K_w expression for the autoionization of water.

16.3	Recall that K_w at 25°C equals 1.0E-14 and varies with temperature.
16.4	**Calculate among pH, pOH, $[H_3O^+]$, and $[OH^-]$.**
16.4	Recognize and use the p-function as the -log of a quantity.
16.4	Calculate among the values for K_w, $[OH^-]$, and $[H_3O^+]$ at a stated temperature.
16.4	Describe neutral, acidic, and basic solutions in terms of pH and pOH or $[H_3O^+]$ and $[OH^-]$.
16.5	**Use K_a values to qualitatively and quantitatively describe weak acids.**
16.5	Recognize that K_a is an equilibrium constant for a weak acid ionization reaction.
16.5	Write the K_a expression for the ionization reaction of a weak acid.
16.5	Rank weak acids in order of strength based on K_a values.
16.5	Calculate the K_a value for an acid using the K_a expression and equilibrium concentrations.
16.5	Calculate the concentration(s) of species in a weak acid soutlion given the K_a value.
16.5	Find the value of K_a given the pH and initial acid concentration.
16.5	Calculate the pH of a weak acid solution using the K_a value.
16.5	**Use percent ionization of a weak acid in calculations.**
16.5	Calculate the percent ionization of a weak acid.
16.5	Calculate the pH of a weak acid given based on percent ionization.
16.5	Calculate the K_a value for a weak acid given percent ionization and initial acid concentration.
16.5	**Calculate the pH of acid mixtures.**
16.5	Calculate the pH of a mixture of strong acids.
16.5	Calculate the pH of a mixture of strong and weak acids.
16.5	Calculate the pH for a mixture of weak acids when one is significantly stronger than the other(s).
16.5	**Use K_b values to qualitatively and quantitatively describe weak bases.**
	Recognize that K_b is an equilibrium constant for a weak base ionization reaction.
16.5	Write the K_b expression for the ionization reaction of a weak base.
16.5	Recognize weak bases include NH_3, and related nitrogen-containing compounds (for example, CH_3NH_2),
16.5	Calculate the K_b value for an base using the K_b expression and equilibrium concentrations.
16.5	Calculate the concentration(s) of species in a weak base soutlion given the K_b value.
16.6	**Describe types of acids acording to the number of ionizable protons.**
16.6	Define and label monoprotic, diprotic, triprotic and polyprotic acids.
16.6	Write the successive ionization reactions for a polyprotic acid.
16.6	Recongize that $K_{a1} > K_{a2} > K_{a3}$.
16.7	**Explain the relationship between K_a, K_b and K_w.**
16.7	Calculate the K_b for a conjugate base given the K_a value of the acid, or calculate the K_a of a conjugate base given the
16.7	**Classify salt solutions as acidic, basic, or pH neutral.**
16.7	Recognize that anions can be weakly basic or neutral.
16.7	Recognize that the counter ion (conugate base) of a strong acid is pH neutral.
16.7	Recognize that the counter ion (conjugate base) of a weak acid is a weak base.
16.7	Rank conjugate bases in order of strength according to the strength of the acid.
16.7	Write the ionization reaction for a basic anion in water.
16.7	Recognize that the conjugate acid of a weak base is a weak acid.
16.7	Recognize that cations can be weakly acidic or neutral.
16.7	Recognize that cations of IA and IIA metals are neutral. That is, the counter ion of a soluble strong base is pH neutral.
16.7	Recognize that small highly charged metal cations are acidic.
16.7	Write the ionization reaction for a acidic cation in water.
16.7	Calculate the pH of a salt containing either an acidic cation or a basic anion.
16.9	**Describe Lewis acids and bases.**
16.9	Define Lewis acid and Lewis base.
16.9	Define a Lewis adduct as the single product resulting from a Lewis acid-base reaction, while distinguishing an adduct from a salt.
16.9	Classify reactants as a Lewis acid or Lewis base given a balanced chemical equation.
17.1	**Discuss the purpose and function of a buffer.**
17.2	Describe the common ion effect in a buffer solution.
17.1	Describe the components of a buffer solution.
17.1	Describe how a buffer neutralizes added strong acid or strong base.

17.1	Explain the purpose of a buffer solution.
17.11	Predict the qualitative change in pH of a buffer solution after the addition of acid or base.
17.1	Identify which combination of substances will result in a buffer solution.
17.2	**Complete buffer calculations.**
17.2	Calculate the pH of a buffer using equilbrium reactions and K expressions or the Henderson-Hasselbalch equation.
17.2	Recognize that Henderson-Hasselbalch equation can only be used when the "x is small" approximation is valid.
17.2	Describe the relative amounts of each substance remaining after the addition to a buffer solution.
17.2	Calculate the change in pH after an acid or base is added to a buffer solution.
17.2	Select the acid-base pair needed to prepare a buffer solution with the given pH.
17.2	Calculate the amounts of acid and conjugate base (or base and conjugate acid) needed to prepare a buffer solution with the given pH.
17.2	**Describe buffer capacity and effectiveness.**
17.2	Define buffer capacity.
17.2	Explain how changing absolute concentrations of components changes the buffer capacity.
17.2	Explain how changing relative concentrations of components changes the buffer capacity.
17.2	Describe how the relative (qualitatitive) amounts of acid and base affect buffer effectiveness.
17.2	Describe how the absolute (quantitative) amounts of acid and base affect buffer effectiveness.
5.9/17.3	**Review the concepts an acid-base titration.**
5.9	Define indicator and explain its use in a titration.
5.9	Distinguish between equivalence point and end point.
17.3/17.4	**Interpret titration curves.**
17.3, 17.4	Predict the approximate pH at the equivalence point of a titration (less than, greater than, or approximately 7).
17.3, 17.4	Identify a titration curve as a strong acid-strong base, strong base-strong acid, weak acid-strong base, or weak base-strong acid.
17.3, 17.4	Identify the equivalence point on a titration curve.
17.3, 17.4	List the substances present in the solution under each condition for each type of titration.
17.4	Identify the "buffer region" on a titration curve in a titration involving a weak acid or weak base.
17.3, 17.4	**Complete titration calculations.**
17.3, 17.4	Calculate the pH in a titration (SA-SB, WA-SB, WB-SA) under the following conditions: initial, before equivalence point, halfway to the equivalence point, at equivalence point, and after equivalence point.
17.3, 17.4	Calculate the volume of titrant needed to reach the equivalence point of a titration.
17.3, 17.4	Determine the point in the titration given volumes and concentrations of solutions.
17.4	Explain the significance of the pH at halfway to the equivalence point for a titration involving a weak acid or a weak base.
17.6	**Explain the concepts surrounding the equilibrium of an insoluble salt in water.**
17.6	Write the chemical equation representing the dissolution of a slightly soluble/insoluble salt.
17.6	Write the K_{sp} expression based on the chemical equation for the dissolution of a slightly soluble/insoluble salt.
17.6	Define molar solubility.
17.6	**Use K_{sp} value in calculations.**
17.6	Calculate the K_{sp} of a salt given molar solubility.
17.6	Calculate molar solubility given a chemical formula and a K_{sp} value.
17.6	For salts that have the same cation to anioin ratio, rank ionic compoiunds in order of solubility using Ksp only.
17.6	For salts that do not have the same cation to anion ratio, rank ionic compounds in order of solubility by calculating the molar solubility.
17.7	**Determine how changes in solution affect solubility.**
17.7	Calculate the molar solubility of a salt in the presence of a common ion.
17.7	Desribe how changes in pH affects the solubility of a salt in solution.
17.8	Determine saturation, unsaturation, and supersaturation based on comparison of Q and K_{sp}.
17.1	**Use K_f in calculations.**
17.1	Define complex ion.
17.1	Define ligand.

22.3	Define chelate, chelation, and chelating agent.
22.3	Define monodentate, bidentate, and polydentate ligands.
22.3	Define coordination compound.
22.3	Define coordinate covalent bond.
17.1	Write the reaction associated with complex ion formation equilibrium.
17.1	Define K_f.
17.1	Use K_f in calculations for equilibrium concentrations of a desired species.
18.1	**Explain the relevance of the first law of thermodynamics.**
18.1	Distinguish between spontaneous and nonspontaneous processes.
18.1	Recognize the spontaneity and speed are NOT related.
18.1	Recognize that chemical potential is comparable to potential energy for a mechanical system.
18.1	Recall the first law of thermodynamics.
18.1	**Explain the relevance of the second law of thermodynamics.**
18.1	Recognize the enthalpy does not predict spontaneity.
18.1	Define entropy and recognize that it is a state function.
18.1	State the second law of thermodynamics.
18.1	**Rank substances in order of the magnitude of entropy values.**
18.1	Predict the relative entropy values of solids, liquids, and gases.
18.1	Rank substances in order of magnitude of entropy values based on state of matter.
18.1	Rank substances in order of magnitude of entropy values based on molar mass.
18.1	Rank substances in order of magnitude of entropy values based on molecular complexity.
18.1	Rank substances in order of magnitude of entropy values for a crystalline solid and dissolving in solution.
18.2	**Explain the relevance of the third law of thermodynamics.**
18.2	State the third law of thermodynamics.
18.2	Predict the relative entropy change for chemical reactions.
18.2	Calculate the entropy change for reactions given standard entropy values.
18.2	**Compare standard state conditions for substances.**
18.2	Define standard molar entropies.
18.2	Recall that the ordinal symbol (°) indicates standard conditions.
18.2	State the standard conditions for solids, liquids, gases, and solutions.
15.1	**Distinguish between reversible and irreversible reactions.**
15.1	Label examples as reversible or irreversible processes.
18.3	**Calculate entropy of system, surroundings, and universe.**
18.3	Calculate $\Delta S_{universe}$ as a function of ΔS_{system} and $\Delta S_{surroundings}$
18.3	Predict the spontaneity of a reaction based on $\Delta S_{universe}$.
18.3	Calculate $\Delta S_{surroundings}$ as a function of enthalpy and temperature at constant temperature and pressure.
18.3	Calculate standard entropy change using standard entropies of products and reactants.
18.4	**Predict spontaneity of a reaction using Gibbs free energy.**
18.4	Define Gibbs Free energy and recognize it is a state function.
18.5	Describe how Gibbs free energy is the energy available to do work in a system.
18.4	Predict the spontaneity of a reaction based on the sign of ΔG.
18.5	**Calculate Gibb's free energy.**
18.5	Complete calculations with the equation of $\Delta G = \Delta H - T\Delta S$ under standard and non-standard conditions.
18.4	Predict the spontaneity of a reaction based on the sign of ΔH, the sign of ΔS, and the magnitude of the temperature.
18.4	Calculate standard Gibbs free energy of a reaction as a function of standard ΔG of formation.
18.4	Calculate Gibbs free energy based on free energy values for each step in a stepwise reaction.
18.5	Calculate Gibb's free energy change of a reaction under nonstandard conditions given the standard Gibb's free energy and concentration or partial pressure of species in the reaction.
18.5	Recongize that ΔG under nonstandard conditions is zero at equilibrium (when $Q = K$).
18.6	**Complete calculations between $\Delta G°$ and K.**
18.6	Explain how the sign of $\Delta G°$ is determined by the value of K.
18.6	Predict the direction of spontaneity based on the value of $\Delta G°$ or K.
18.6	Explain the relationship between the temperature and the magnitude of the equilbrium constant.

4.6, 19.1	**Identify the oxidation and reduction half-reactions.**
4.6	Recall rules for assigning oxidation numbers.
4.6	Recall definitions for oxidation, reduction, oxidizing agent, and reducing agent.
4.6	Assign oxidation states to all elements in a reaction.
4.6	Identify the oxidizing and reducing agents.
19.4	**Describe the components of an electrochemical cell and their purpose.**
19.4	Define electrical current.
19.4	Define amperes (amps) and volts and state their meaning in terms of other SI units (C/s; J/C).
19.4	Define electrochemical cell, voltaic (galvanic) cell, and electrolytic cell.
19.4	Define half-cell and electrodes.
19.4	Define anode, cathode, and salt bridge.
19.4	Label an electrochemical cell to indicate anode, cathode, salt bridge, anodic solution, cathodic solution, and the flow of electrons.
19.4	Describe an inert electrode and when it is used.
19.4	Interpret the electrochemical cell notation.
19.4	**Identify the reactions in an electrochemical cell.**
19.4	Write the half-reactions that occur at the anode and cathode.
19.5	**Determine the cell potential.**
19.5	Recognize that voltage is an intensive property.
19.5	Define potential difference and electromotive force (emf).
19.5	Define cell potential and standard cell potential.
19.5	Relate the sign of the cell potential to the spontaneity of the forward reaction.
19.5	Recognize that the magnitude of the cell potential indicates the likelihood that the reaction will occur.
19.5	Describe the standard hydrogen electrode (SHE) and state its potential.
19.5	Predict the cathode and anode of an electrochemical cell given the standard reduction potentials.
19.5	Calculate the standard cell potential given the standard reduction potentials.
19.5	Predict the spontaneous redox reaction.
19.6	**Complete calculations using the equations that relate K, $\Delta G°$, and E°.**
19.6	Determine the number of moles of electrons (n) transferred in an electrochemical reaction.
19.6	Determine E°, $\Delta G°$, and K using standard reduction potentials.
19.7	**Complete calculations using the Nernst equation for nonstandard conditions.**
19.7	Distinguish between standard and nonstandard conditions of an electrochemical cell.
19.9	**Describe electrolytic cells.**
19.9	Define electrolysis in terms of spontaneity.
19.9	Recognize that an electrolytic cell requires an external source of voltage.
20.1	**Describe types of radioactive decay.**
20.1	Identify alpha, beta minus, positron, and gamma particles.
20.1	Write reactions representing, alpha, beta minus, positron, and gamma decay.
20.1	Define nucleon.
20.7	**Use binding energy arguments to explain why nuclei are stable or radioactive.**
20.7	Define binding energy.
20.7	**Use half-life and first order kinetics in calculations for radioisotopes.**
20.3	Recognize a plot of exponential decay and identify the half-life.
20.3	Describe radioactive decay with first-order kinetics.
20.4	Understand radiocarbon dating in terms of first-order kinetics.