

OpenStax		Learning Outcomes
10.1		Describe properties of states of matter.
10.1		Identify states of matter.
10.1		Distinguish between properties of states of matter including distance between and movement of molecules.
10.1		Identify the types of forces present between like molecules or between different substances.
10.1		Distinguish between intermolecular (IMF) and intramolecular forces.
10.1		Define dispersion (London) forces, dipole-dipole forces, hydrogen bonding, and ion-dipole forces and what interactions occur for each type of of intermolecular force.
10.1		Identify the types of forces present between like molecules or between different substances.
10.1		Rank properties of substances based on differences in intermolecular forces.
10.1		Describe how molar mass, molecular shape, and polarity affect the magnitude of intermolecular forces.
10.1		Rank substances in order of increasing or decreasing melting point and boiling point.
10.2		Define surface tension.
10.2		Explain differences in surface tension as a function of their intermolecular forces.
10.2		Define viscosity.
10.2		Rank substances in order of increasing viscosity with respect to molar mass, shape, temperature, and intermolecular forces.
10.2		Describe vaporization and complete calculations associated with heat of vaporization.
10.2		Distinguish between vaporization and condensation.
10.3		Distinguish between volatile and nonvolatile.
10.3		Explain how rate of vaporization changes with temperature, surface area, and intermolecular forces.
10.3		Describe how increases in the average kinetic energy of molecules leads to more molecules in the gas phase.
10.3		Define heat of vaporization.
10.3		Describe the relationship between change in enthalpy of condensation and the heat of vaporization.
10.3		Complete calculations involving heat, heat of vaporization, or amount of substance.
10.3		Describe vapor pressure and changes it undergoes.
10.3		Define dynamic equilibrium and vapor pressure.
10.3		Explain how dynamic equilibrium and vapor pressure are affected by changes in pressure as a result of changes in volume.
10.3		Distinguish between boiling point and normal boiling point.
10.3		Describe how intermolecular forces affect boiling points and vapor pressure.
10.3		Explain how dynamic equilibrium and vapor pressures are affected by changes in temperature.
10.3		Use a graph of the natural log of pressure vs inverse temperature to determine the enthalpy of vaporization.
10.3		Explain the different types of phase changes and associated energy changes.
10.3		Describe the phase changes that occur in sublimation, deposition, fusion, and freezing.
10.3		Define melting point and heat of fusion.
10.3		Describe the relationship between change in heat of freezing and the heat of fusion.
10.3		Explain why heat of vaporization has a larger magnitude than heat of fusion.
10.3		Draw and complete calculations for a heating or cooling curve.

10.3		Draw a heating or cooling curve for a substance from the solid phase to the gas phase based on its properties.
10.3		Complete calculations involving heat, heat of fusion, or amount of substance.
10.3		Calculate the total heat change, with the correct units, based on a heating or cooling curve.
10.4	Explain the information provided in a phase diagram.	
10.4		Describe the regions of a phase diagram.
10.4		Identify the variables on the x- and y-axes of a phase diagram.
10.4		Identify the phase changes indicated by the curves in the phase diagram.
10.4		Locate the triple point and the critical point on a phase diagram.
10.4		Recognize the phases in equilibrium that occur on the curves and the triple point of the phase diagram.
10.4		Describe the critical point.
10.4		Explain the meaning of the slope of the fusion curve in a phase diagram and its implications on physical properties.
10.4		Use a phase diagram to explain why a substance may sublime under atmospheric pressure.
10.4		Define a supercritical fluid.
10.4		Define critical temperature.
10.4		Define critical pressure.
10.5	Label examples as molecular, ionic, nonbonding, metallic, or network covalent solids.	
10.5		Identify a solid as crystalline or amorphous.
10.5		Distinguish among molecular, ionic, and atomic solids.
10.5		Describe atomic solids as nonbonding, metallic, or network covalent.
10.5		State the distinguishing properties of molecular, ionic, nonbonding, metallic, and network covalent solids.
10.5		Distinguish between covalent, metallic, and ionic bonding.
10.5		Explain the non-directional features of metallic bonds.
10.6	Describe the geometric properties of cubic unit cells.	
10.6		Define crystalline lattice, unit cell, coordination number, and packing efficiency.
10.6		Calculate the number of atoms per cell in simple, body-centered, and face-centered cubic cells.
10.6		Recognize there are differences in packing efficiencies and coordination number among unit cells.
10.6	Explain the qualitative and quantitative properties of a unit cell.	
10.6		Distinguish among body, edge, face, and corner atoms in a unit cell.
10.6		State the fraction of each atom inside the unit cell for body, edge, face, and corner atoms.
10.6		Calculate the mass, volume, and density of a given unit cell.
10.6		Use mass, volume, or density to identify type of unit cell and molar mass of element.
10.6		Identify an element based on its molar mass.
10.6		Determine the ionic formula from unit cell information.
16.1	Explain the relevance of the first law of thermodynamics.	
16.1		Distinguish between spontaneous and nonspontaneous processes.
16.1		Recognize the spontaneity and speed are NOT related.
16.1		Recognize that chemical potential is comparable to potential energy for a mechanical system.
16.1		Recall the first law of thermodynamics.

16.2	Explain the relevance of the second law of thermodynamics.
16.2	Recognize the enthalpy does not predict spontaneity.
16.2	Define entropy and recognize that it is a state function.
16.2	State the second law of thermodynamics.
16.2	Compare standard state conditions for substances.
16.2	Recall that the ordinal symbol (°) indicates standard conditions.
16.2	State the standard conditions for solids, liquids, gases, and solutions.
16.2	Rank substances in order of the magnitude of entropy values.
16.2	Define standard molar entropies.
16.2	Predict the relative entropy values of solids, liquids, and gases.
16.2	Rank substances in order of magnitude of entropy values based on state of matter.
16.2	Rank substances in order of magnitude of entropy values based on molar mass.
16.2	Rank substances in order of magnitude of entropy values based on allotropic forms.
16.2	Rank substances in order of magnitude of entropy values based on molecular complexity.
16.2	Rank substances in order of magnitude of entropy values for a crystalline solid and dissolving in solution.
16.2	Distinguish between reversible and irreversible reactions.
16.2	Label examples as reversible or irreversible processes.
16.3	Calculate entropy of system, surroundings, and universe.
16.3	Calculate $\Delta S_{\text{universe}}$ as a function of ΔS_{system} and $\Delta S_{\text{surroundings}}$
16.3	Predict the spontaneity of a reaction based on $\Delta S_{\text{universe}}$.
16.3	Calculate $\Delta S_{\text{surroundings}}$ as a function of enthalpy and temperature at constant temperature and pressure.
16.3	Calculate standard entropy change using standard entropies of products and reactants.
16.4	Calculate Gibbs free energy and predict spontaneity of a reaction.
16.4	Define Gibbs Free energy and recognize it is a state function.
16.4	State the equation that relates ΔG to ΔH , temperature, and ΔS .
16.4	Predict the spontaneity of a reaction based on the sign of ΔG .
16.4	Predict the spontaneity of a reaction based on the sign of ΔH , the sign of ΔS , and the magnitude of the temperature.
16.4	Complete calculations with the equation of $\Delta G = \Delta H - T\Delta S$ under non-standard conditions.
16.4	Calculate standard Gibbs free energy of reaction as a function of standard changes of enthalpy and entropy.
16.4	Calculate standard Gibbs free energy of a reaction as a function of ΔG of formation.
16.4	Calculate standard Gibbs free energy based on free energy values for each step in a stepwise reaction.
11.1	Identify solutions and their components.
11.1	Define solute, solvent, and solution.
11.1	Recognize that solutions can be in liquid, solid, or gas phases or a combination of phases.
11.1	Identify the components of an aqueous solution.
11.1	Describe how the relative strengths of the interactions among molecules determine whether or not a solution will form.
11.1	Describe how entropy drives the formation of a solution.
11.1	Explain the effect of intermolecular forces on solution formation.
11.1	Define solvent-solute, solvent-solvent, and solute-solute interactions.

11.1		Describe the enthalpies of solute, solvent, solution, and mixing and classify as endothermic or exothermic.
11.3	Describe the	properties of a solution.
11.3		Distinguish between miscible and nonmiscible (immiscible) substances.
11.3		Distinguish among saturated, unsaturated, and supersaturated solutions.
11.3		Explain how to prepare a supersaturated solution.
11.3	Provide qualitative and quantitative explanations of how temperature affects solubility.	
11.3		Recognize that temperature affects the solubility of solids.
11.3		Use a solubility curve and information about a solution to identify a solution as saturated, unsaturated, or supersaturated.
11.3		Describe how temperature and pressure affects the solubility of gases in water.
11.3		State Henry's Law (equation) and identify the variables.
11.3		Use Henry's Law to calculate solubility or pressure given the other and the value of k_H (Henry's law constant).
11.3		Use Henry's Law to calculate solubility or pressure given initial conditions and either solubility or pressure at the final conditions.
11.3		Use given values of Henry's Law constant in problems.
3.3, 3.4, 11.4	Calculate the concentration of a solution.	
3.3, 3.4, 11.4		Define molarity, molality, mole fraction, parts by mass (percent, ppm, ppb), and parts by volume (percent, ppm, ppb).
3.3, 3.4, 11.4		Calculate concentrations in units of molarity, molality, mole fraction, parts by mass (percent, ppm, ppb), and parts by volume (percent, ppm, ppb).
3.3, 3.4, 11.4		Identify the substance for all units in concentrations as solute, solvent, or solution (i.e. moles solute per liter of solution).
3.3, 3.4, 11.4		Solve for the quantity of solute, solvent or solution given a concentration and other needed values.
3.3, 3.4, 11.4		Convert between any two units of concentration (molarity, molality, mole fraction, parts by mass (percent, ppm, ppb), and parts by volume (percent, ppm, ppb).
11.4	Describe qualitative and quantitative effects of solution concentration on properties.	
11.4		Describe colligative properties.
11.4		Define vapor pressure lowering, freezing point depression, boiling point elevation, and osmotic pressure.
11.4		Explain how each colligative property is affected by the concentration of the solute in a solution.
11.4		State Raoult's law (equation) and identify all variables.
11.4		Use Raoult's law to calculate vapor pressure of the solution, vapor pressure of the pure solvent, or the mole fraction of the solvent.
11.4		Calculate the change in vapor pressure as a result of the addition of a solute to form a solution.
11.4		Recognize that both solute and solvent can be volatile.
11.4		Distinguish between solutions with strong and weak solute-solvent interactions.
11.4		State the equations for boiling point elevation and freezing point depression and identify all variables..
11.4		Use the equations to calculate the change in temperature, the molality, the amount of solute, molar mass, or the freezing point depression or boiling point elevation constant for a solution.
11.4	Describe the	flow of solvent particles during osmosis.
11.4		Describe a semipermeable membrane.

11.4		State the equation for osmotic pressure and identify the variables.
11.4		Calculate the osmotic pressure, the molarity, amount of solute, or the temperature of a solution.
11.4		Describe differences between isosmotic (isotonic), hyperosmotic (hypertonic), and hyposmotic (hypotonic) solution.
11.4		Complete calculations involving vapor pressure lowering, freezing point depression, boiling point elevation, and osmotic pressure for strong electrolyte solutions.
11.4		Define van't Hoff factor.
11.4		Determine the ideal van't Hoff factor for an ionic substance.
11.4		Describe the ion pairing that leads to non-ideal van't Hoff factors.
11.4		Calculate molality and molarity of particles in solution (im or iM).
11.4		Use colligative property measurements to determine the molar mass of an unknown solute.
12.1		Calculate rate of reaction or change in concentration of substance based on given data.
12.1		Define rate of reaction.
12.1		Write rate of reaction in terms of change of concentration of reactants or products with a change in time.
12.1		Distinguish between instantaneous and average rate of reaction.
12.3		Write rate law and determine the order of reaction.
12.3		Define rate constant and rate law,
12.3		Recognize that the rate law must be based on experimental data.
12.3		Write the rate law based on experimental data.
12.3		Identify the overall order and the order with respect to reactants given a rate law.
12.4		Calculate rate constant based on experimental data and the rate law.
12.4		Calculate rate based on changes in concentration (i.e. twice A or half of B).
12.4		Label a graph of reactant concentration vs time as zero, first or second order.
12.4		Label a graph of rate vs reactant concentration as zero, first, or second order.
12.4		Complete calculations using the integrated rate laws for zero-, first-, and second-order reactions.
12.4		Define integrated rate law.
12.4		State the integrated rate laws (equations) for zero-, first-, and second-order reactions and identify all variables .
12.4		Know the units of the rate constant, k, for zero-, first-, and second-order reactions.
12.4		Label the axes, y-intercept, and slope of the straight-line plot for zero-, first-, and second-order reactions.
12.4		Complete calculations using the half-life equations for zero-, first-, and second-order reactions.
12.4		State the half-life equations for zero-, first-, and second-order reactions and identify all variables.
12.4		Recognize the nuclear decay follows first order kinetics.
12.5		Determine and use activation energy in calculations.
12.5		Define activation energy, activated complex, and transition state.
12.5		Recognize that the value of the rate constant is dependent on temperature.
12.5		Describe how temperature affects the rate of a reaction.
12.5		Use Arrhenius plot (ln k vs. 1/T) to calculate the activation energy.
12.5		State the two-point Arrhenius equation and define all variables.
12.5		Complete calculations using the two-point Arrhenius equation.
12.6		Explain reaction mechanisms.
12.6		Define reaction mechanism, elementary step, and reaction intermediate.
12.6		Define molecularity, unimolecular, bimolecular, and termolecular.

12.6		Identify reaction intermediates based on the reaction mechanism.
12.6		Write rate laws for elementary steps.
12.6		Define rate-determining step.
12.6		Identify the rate-determining step in a mechanism.
12.6		Describe the two conditions for a valid reaction mechanism.
12.6		Label activation energies of each step, transition states, reactants, products, and enthalpy of reaction in an energy diagram (energy vs reaction progress).
12.6		Write the overall rate law for mechanisms when the first step is slow.
12.6		Recognize that when the second or later step in the mechanism is slow, the overall rate law must be derived algebraically.
12.7		Describe the effect of a catalyst on the energetics of a reaction and the reaction rate.
12.7		Define catalyst.
12.7		Identify the catalyst in a mechanism.
12.7		Distinguish between homogeneous and heterogeneous catalysts.
12.7		Describe the function of an enzyme as a biological catalyst.
13.2		Write the law of mass action for an equilibrium reaction and explain its relevance.
13.1		Explain what is constant for a reaction at equilibrium.
13.1		Explain what changes in a system to reach equilibrium.
13.1		Explain how the concentrations of reactants and products change with time as a reaction achieves equilibrium.
13.2		Define equilibrium constant, K .
13.2		Recognize that concentrations must be in units of molarity.
13.2		Identify whether products, reactants, or neither are favored based on the magnitude of K .
13.2		Calculate the equilibrium constant when reactions are combined or manipulated.
13.2		Calculate K for the reverse reaction when given K of the forward reaction.
13.2		Calculate K when the coefficients of a reaction are multiplied by a factor given the original value of K .
13.2		Calculate K when two or more equations are combined given K for the individual equations.
13.2		Convert between K_p and K_c .
13.2		Write the law of mass action using partial pressures of substances in units of atm.
13.2		Calculate K_p based on equilibrium partial pressures.
13.2		Calculate the change in moles of gas based on a chemical equation.
13.2		State the equation that relates K_p and K_c .
13.2		Calculate the equilibrium constant using the law of mass action for a heterogeneous system.
13.2		Identify an equilibrium system that is heterogeneous.
13.2		Write the law of mass action for a heterogeneous system.
13.2		Recognize that pure liquids and solids are excluded from the equilibrium calculation.
13.3		Predict the shift in a reaction according to Le Chatelier's principle.
13.3		Define Le Chatelier's principle.
13.3		Explain what is meant by a "shift" in a reaction (i.e. which direction is favored).
13.3		Predict the shift in a reaction due to change in concentration of a reactant or product.
13.3		Recognize that changing the amount of solids and pure liquids does not shift the reaction.
13.3		Describe how the concentration of each substance changes as the shift occurs.
13.3		Predict the shift in a reaction due to change in volume (which results in a change in pressure) of the reaction container.

13.3		Recognize that a change in volume will not shift the equilibrium if there are no moles of gas in the reaction or equal moles of gas in products and reactants.
13.3		Identify an inert gas in a reaction container.
13.3		Recognize that the addition of an inert gas does not result in a shift in the equilibrium.
13.3		Recall the difference between endothermic and exothermic reactions.
13.3		Predict the shift in a reaction due to changes in temperature in both endothermic and exothermic reactions.
13.3		Predict whether K will increase or decrease with a change in temperature.
13.4		Determine equilibrium concentrations based on initial concentrations and a single equilibrium concentration or a change in concentration.
13.4		Explain how to use an ICE (initial-change-equilibrium) table.
13.4		Predict the direction of a reaction based on comparison of Q and K.
13.4		Define reaction quotient, Q.
13.4		Calculate reaction quotient in terms of concentration or partial pressure.
13.4		Calculate equilibrium concentrations.
13.4		Determine equilibrium concentration (or partial pressure) when given K and all but one equilibrium concentration (or partial pressure).
13.4		State the quadratic formula and identify the variables.
13.4		Calculate equilibrium concentrations (or partial pressures) given K and initial concentrations (or partial pressures) using the quadratic equation.
13.4		Describe when a simplifying assumption can be made.
13.4		Calculate equilibrium concentrations (or partial pressures) given K and initial concentrations (or partial pressures) using the simplifying assumption.
13.4		Calculate equilibrium concentrations (or partial pressures) given K and initial concentrations (or partial pressures) using the method of perfect squares.
16.4		Calculate the free energy change of a reaction under nonstandard conditions.
16.4		Describe how Gibbs free energy is the energy available to do work in a system.
16.4		State the equation that relates ΔG to ΔG°
16.4		Recognize that ΔG under nonstandard conditions is zero at equilibrium (when $Q = K$).
16.4		Complete calculations between ΔG° and K.
16.4		State the equation that relates ΔG° to the equilibrium constant.
16.4		Explain how the sign of ΔG° is determined by the value of K.
16.4		State the value of ΔG° when $K = 1$.
16.4		Predict the direction of spontaneity based on the value of ΔG° or K.
16.4		Explain the relationship between the temperature and the magnitude of the equilibrium constant.
14.1		Describe Arrhenius and Bronsted-Lowry acids and bases and how they behave in solution.
14.1		List the general properties of acids and bases.
14.1		Define Arrhenius acids and bases.
14.1		Write equations to represent the behavior of Arrhenius acids and bases in aqueous solution.
14.1		Define Bronsted-Lowry acids and bases.
14.1		Label substances in a reaction as Bronsted-Lowry acids and bases.
14.1		Define conjugate acid and conjugate base.
14.1		Generate the formulas for conjugate acid-base pairs.
14.1		Define amphoteric.
14.1		Explain the meaning of K_w .
14.1		Define autoionization.

14.1		Write the reaction for the autoionization of water.
14.1		Write the K_w expression for the autoionization of water.
14.1		Recall that K_w at 25°C equals 1.0E-14 and varies with temperature.
14.2		Calculate among pH, pOH, $[H_3O^+]$, and $[OH^-]$.
14.2		Recognize that $pH + pOH = pK_w$.
14.1		Calculate among the values for K_w , $[OH^-]$, and $[H^+]$ at stated temperature.
14.2		Describe neutral, acidic, and basic solutions in terms of pH and pOH or $[H_3O^+]$ and $[OH^-]$.
4.2		Identify the relative amounts of each species for a strong acid or strong base in water.
3.3		Calculate the $[OH^-]$ of strong bases, considering chemical formula.
14.2		Calculate the pH of a strong acid or strong base solution.
14.3		Distinguish between strong and weak acids and strong and weak bases.
14.3		Write an equation representing the behavior of a weak acid in water.
14.3		Identify substances that are weak acid or bases.
14.3		Use K_a values to qualitatively and quantitatively describe weak acids.
14.3		Recognize that K_a is an equilibrium constant for a weak acid ionization reaction.
14.3		Write the K_a expression for the ionization reaction of a weak acid.
14.3		Rank weak acids in order of strength based on K_a values.
14.3		Calculate the pH of a weak acid solution using the K_a value.
14.3		Find the value of K_a based on pH and initial concentration of the acid.
14.3		Use percent ionization of a weak acid in calculations.
14.3		Calculate the percent ionization of a weak acid.
14.3		Calculate the pH of a weak acid given based on percent ionization.
14.3		Calculate the pH for mixtures of acids.
14.3		Calculate the pH for a mixture of strong acids.
14.3		Calculate the pH for a mixture of strong and weak acids.
14.3		Calculate the pH for a mixture of weak acids.
14.3		Use K_b values to to qualitatively and quantitatively describe weak bases.
14.3		Write the reaction for the ionization of a weak base in water.
14.3		Recognize weak bases include NH_3 , related amine compounds (i.e. CH_3NH_2), and conjugate bases of weak acids.
14.3		Calculate the pH of a weak base solution using the K_b value.
14.3		Find the value of K_b based on pH and initial concentration of the base.
14.3		Rank acids in order of strength.
14.3		Rank binary acids in order of strength based on polarity of the bond.
14.3		Rank binary acids in order of strength based on bond strength.
14.3		Rank oxyacids in order of strength based on the electronegativity of the central atom when the number of oxygen atoms in each acid are equal.
14.3		Rank oxyacids in order of strength based on the number of oxygen atoms when the central atom is the same.
14.3		Identify the ionizable protons in an oxyacid.
14.4		Classify salt solutions as acidic, basic, or neutral.
14.4		Recognize that anions can be weakly basic or neutral.
14.4		Write the reaction for an anion in water.
14.4		Recognize that the conjugate base of a strong acid is pH neutral.
14.4		Recognize that the conjugate base of a weak acid is a weak base.
14.4		Rank conjugate bases in order of strength according to the strength of the acid.
14.4		Recognize that cations can be weakly acidic or neutral.

14.4		Write the reaction for a cation in water.
14.4		Recognize that the conjugate acid of a strong base is pH neutral.
14.4		Recognize that the conjugate acid of a weak base is a weak acid.
14.4		Rank conjugate acids in order of strength according to the strength of the base.
14.4	Explain the relationship between K_a , K_b and K_w .	
14.4		State the relationship that K_a times K_b equals K_w .
14.4		Complete calculations using K_a times $K_b = K_w$.
14.4		Calculate the appropriate K value for a conjugate acid or base.
14.4		Complete calculations using $pK_a + pK_b = pK_w$
14.4	Describe qualitative and quantitative information about neutralization reactions.	
14.4		Predict the products of a neutralization reaction.
14.4		Calculate the amount of each substance remaining after a neutralization reaction occurs.
14.5	Describe types of acids.	
14.5		Identify the carboxylic acid group.
14.5		Define monoprotic, diprotic, and triprotic.
14.5		Label acids as monoprotic, diprotic, or triprotic.
14.5		Define polyprotic acid.
14.5		Write the successive ionization reactions for a polyprotic acid.
14.5		Recognize that $K_{a1} > K_{a2} > K_{a3}$.
14.5		Calculate the pH of a polyprotic acid solution (other than sulfuric acid).
14.5		Recognize that only the first ionization step is used in determining pH (except for sulfuric acid).
14.5		Recognize that for a dilute sulfuric acid solution, the second ionization step contributes significant $[H_3O^+]$ to the solution which lowers the pH.
14.6	Discuss the purpose and function of a buffer.	
14.6		Calculate between pK_a and K_a .
14.6		Describe the components of a buffer solution.
14.6		Describe how a buffer neutralizes added strong acid or strong base.
14.6		Explain the purpose of a buffer solution.
14.6		Predict the qualitative change in pH of a buffer solution after the addition of acid or base.
14.6		Identify which combination of substances will result in a buffer solution.
14.6	Complete buffer calculations.	
14.6		Calculate the pH of a buffer solution formed from an acid and its conjugate base.
14.6		Describe the common ion effect in a buffer solution.
14.6		State the Henderson-Hasselbalch equation and define all variables.
14.6		Recognize that Henderson-Hasselbalch equation can only be used when the "x is small" approximation is valid.
14.6		Use the Henderson-Hasselbalch equation in calculations with a buffer solution formed from an acid and its conjugate base.
14.6		Use the Henderson-Hasselbalch equation in calculations involving additions to a buffer solution.
14.6		Describe the relative amounts of each substance remaining after the addition to a buffer solution.
14.6		Calculate the pH of a buffer solution formed from a base and its conjugate acid.
14.6		Use the Henderson-Hasselbalch equation in calculations involving a buffer formed from a base and its conjugate acid.
14.6		Explain the range of a buffer solution in terms of pK_a of the acid.

14.6		Calculate the amounts of acid and conjugate base (or base and conjugate acid) needed to prepare a buffer solution with the given pH.
14.6		Select the acid-base pair needed to prepare a buffer solution with the given pH.
14.6	Describe buffer capacity and effectiveness.	
14.6		Define buffer capacity.
14.6		Explain how changing absolute concentrations of components changes the buffer capacity.
14.6		Explain how changing relative concentrations of components changes the buffer capacity.
14.6		Describe how the relative (qualitative) amounts of acid and base affect buffer effectiveness.
14.6		Describe how the absolute (quantitative) amounts of acid and base affect buffer effectiveness.
14.7	Choose appropriate indicator for a titration.	
14.7		Define indicator.
14.7		Interpret indicator chart for titrations.
14.7	Complete titration calculations.	
14.7		Distinguish between equivalence point and end point.
14.7		Identify the titration as strong acid-strong base, weak acid-strong base, or weak base-strong acid.
14.7		Predict the approximate pH at the equivalence point of a titration (less than, greater than, or approximately 7).
14.7		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.
14.7		Identify a titration curve as a strong acid-strong base, strong base-strong acid, weak acid-strong base, or weak base-strong acid.
14.7		Identify the equivalence point on a titration curve.
14.7		Identify the "buffer region" on a titration curve in a titration involving a weak acid or weak base.
14.7		Calculate the volume of titrant needed to reach the equivalence point of a titration.
14.7		List the substances present in the solution under each condition for each type of titration.
14.7		Determine the point in the titration given volumes and concentrations of solutions.
14.7		Explain the significance of the pH at halfway to the equivalence point for a titration involving a weak acid or a weak base.
14.7		Identify titration curve involving the titration of a polyprotic acid.
14.7		Identify the first and successive equivalence points on the titration curve in a titration with a polyprotic acid.
14.7		Calculate the amount of titrant needed to reach the first and successive equivalence points in a titration with a polyprotic acid.
15.1	Use K_{sp} value in calculations.	
15.1		Define solubility product constant (K_{sp}).
15.1		Write the chemical equation representing the dissolution of a slightly soluble salt.
15.1		Write the K_{sp} expression based on the chemical equation for the dissolution of a slightly soluble salt.
15.1		Define molar solubility.
15.1		Calculate molar solubility given a chemical formula and a K_{sp} value.

15.1		Rank substances in order of solubility based on molar solubility values.
15.1		Explain why K_{sp} cannot always be used to determine relative solubilities of salts.
15.1		Calculate the K_{sp} of a salt given molar solubility.
15.1	Determine how changes in solution affect solubility.	
15.1		Describe how a common ion effects solubility of a salt using Le Chatelier's principle.
15.1		Calculate the molar solubility of a salt in the presence of a common ion.
15.1		Describe how changes in pH affects the solubility of a salt in solution.
15.1		Recall definitions of soluble and insoluble.
15.1		Distinguish between Q and K_{sp} .
15.1		Determine saturation, unsaturation, and supersaturation based on comparison of Q and K_{sp} .
15.1	Complete calculations for selective precipitation.	
15.1		Define selective precipitation.
15.1		Calculate the minimum required reagent concentration needed for selective precipitation.
15.1		Calculate concentration of ions remaining in solution after selective precipitation.
15.2	Describe Lewis acids and bases.	
15.2		Define Lewis acid and Lewis base.
15.2		Explain how small, highly charged (3+ or greater) metal cations behave as an acid.
15.2		Explain how a Lewis acid can function as an electron pair acceptor.
15.2		Explain how a Lewis base can function as an electron pair donor.
15.2		Classify as a Lewis acid or Lewis base given a balanced chemical equation.
17.1	Identify the oxidation and reduction half-reactions.	
4.2		Recall rules for assigning oxidation numbers.
4.2		Recall definitions for oxidation, reduction, oxidizing agent, and reducing agent.
4.2		Assign oxidation to all elements in a reaction.
17.1		Identify the oxidizing and reducing agents.
17.1	Describe the components of an electrochemical cell and their purpose.	
17.1		Define electrical current.
17.1		Define amperes (amps) and volts and state their meaning in terms of their fundamental units (C/s; J/C).
17.2		Define electrochemical cell, voltaic (galvanic) cell, and electrolytic cell.
17.2		Define half-cell and electrodes.
17.2		Define anode, cathode, and salt bridge.
17.2		Label an electrochemical cell to indicate anode, cathode, salt bridge, anodic solution, cathodic solution, and the flow of electrons.
17.2		Explain the purpose of a salt bridge.
17.2		Describe an inert electrode and when it is used.
17.2	Identify the reactions in an electrochemical cell.	
17.2		Write the half-reactions that occur at the anode and cathode.
17.2		Interpret the electrochemical cell notation.
17.2	Determine the cell potential.	
17.2		Recognize that voltage is an intensive property.
17.2		Define potential difference and electromotive force (emf).
17.3		Define cell potential and standard cell potential.
17.3		Relate the sign of the cell potential to the spontaneity of the forward reaction.
17.3		Recognize that the magnitude of the cell potential indicates the likelihood that the reaction will occur.

17.3		Describe the standard hydrogen electrode (SHE) and state its potential.
17.3		State the equation to calculate the standard cell potential.
17.3		Predict the cathode and anode of an electrochemical cell given the standard reduction potentials.
17.3		Predict the stronger oxidizing agent and stronger reducing agent of an electrochemical cell given the standard reduction potentials.
17.3		Calculate the standard cell potential given the standard reduction potentials.
17.3		Predict the spontaneous reaction of a redox reaction.
17.4	Complete calculations using the equations that relate K , ΔG° , and E° .	
17.4		State the qualitative relationships between ΔG , K , and E° of the cell.
17.4		Recognize Faraday's constant.
17.4		State the equation to relate ΔG° and E° and identify all variables.
17.4		State the equation that relates E° and the equilibrium constant, K , and identify all variables.
17.4		Determine the moles of electrons (n) in an electrochemical reaction.
17.4		Determine E° , ΔG° , and K based on the standard reduction potentials of an electrochemical reaction.
17.4	Complete calculations using the Nernst equation for nonstandard conditions.	
17.4		Distinguish between standard and nonstandard conditions of an electrochemical cell.
17.4		State the Nernst equation and identify all variables.
17.4		Recall the reaction quotient, Q , and how to calculate the value of Q .
17.7	Describe electrolytic cells.	
17.7		Define electrolysis in terms of spontaneity.
17.7		Recognize that an electrolytic cell requires an external source of voltage.
17.7		Recognize that water can be oxidized or reduced.