

OpenStax	Homework	Learning Outcomes
8.4	#1	Describe the basic idea of molecular orbital (MO) theory.
8.4		Explain why MO theory is needed.
8.4		Distinguish between bonding and antibonding orbitals.
8.4		Draw the MO diagram for homonuclear diatomics.
8.4		Calculate the bond order based on the MO diagram.
8.4		Relate the bond order to the stability of the bond.
8.4		Determine if a diatomic molecule is diamagnetic or paramagnetic based on the MO diagram.
8.4		Relate the concept of delocalized molecular orbitals to the concept of resonance.
8.4		Describe the linear combination of atomic orbitals that form bonding and anti-bonding molecular orbitals.
8.4		Describe how destructive and constructive interference affects molecular orbitals.
8.4		Label the components of the molecular orbital diagram of H <sub>2</sub> .
10.5		Explain the non-directional features of metallic bonds.
10.5		Describe how mobility of electrons affects the physical properties of metals (conductivity of electricity and heat, malleability, and ductility).
10.5		Define conductivity, malleability, and ductility.
8.4		Differentiate among atomic, bonding molecular orbital, antibonding molecular orbital, valence band, and conduction band.
8.4		Define band gap.
8.4		Relate the size of the band gap to its conductivity (conductor, semiconductor, insulator)
8.4		Describe a doped semiconductor.
8.4		Identify n- and p-type semiconductors.
8.4		Recognize the importance of p-n junctions, diodes, and amplifiers in modern electronics.
		Recognize that temperature is proportional to conductivity in semiconductors but not conductors due to band gap.
10.6	#2	Define crystalline lattice, unit cell, coordination number, and packing efficiency.
10.6		Describe the geometric properties of cubic unit cells.
10.6		Calculate the number of atoms per cell in simple, body-centered, and face-centered cubic cells.
10.6		State the coordination number and packing efficiency for each type of cubic unit cell.
10.6		Calculate edge length of a unit cell or radius of an atom given the appropriate formula.
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, edge length of cell, radius of atom in 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.
10.6		Recognize that cubic closest packing is the same structure as a face-centered cubic.
10.6		Distinguish between the packing patterns of hexagonal closest packing and cubic closest packing.
5.1	#3	Define specific heat and heat capacity.
5.2		Define state functions.
5.2		Recall how thermal energy is transferred between the system and surrounding.

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5.2		State the equation relating heat, mass, specific heat, and temperature and identify all variables.
5.2		Use the equation relating heat, mass, specific heat, and temperature in calculations.
5.2		State the equation relating heat, heat capacity, and temperature and identify all variables.
5.2		Use the equation relating heat, heat capacity, and temperature and define all variables.
5.2		Complete calculations for the transfer of heat between two substances.
5.2		Sketch energy diagram for endothermic and exothermic processes.
5.2		Explain the sign convention used for endothermic and exothermic processes.
5.2		Use the principles of calorimetry to measure heat transfer between objects.
5.2		Use thermochemical equations to convert between quantity of a substance and heat.
5.2		Use measurement values from a constant volume calorimeter to calculate unknown values such as enthalpy change or heat capacity.
5.2		Define calorimetry.
5.2		Determine the heat capacity of a bomb calorimeter from experimental data.
5.2		Recognize the components of a constant volume (bomb) calorimeter.
5.2		Explain connections between change in internal energy and change in enthalpy.
5.2		Recognize the components of a constant pressure calorimeter.
10.3		Draw a heating curve for a substance from the solid phase to the gas phase based on its properties.
10.3		Draw a cooling curve for a substance from the gas phase to the solid 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 based on a heating or cooling curve.
10.3		Use correct units in calculations involving the heating or cooling curve.
10.3		Determine the appropriate equations to calculate heat for each segment of the heating or cooling curve.
10.3		Identify the correct sign for the heat change given a heating or cooling curve.
10.3		Calculate the final temperature when two substances reach thermal equilibrium through heat transfer and/or phase change.
5.3	#4	Recall the first law of thermodynamics.
5.3		Recall the sign convention of heat.
5.3		Recall the sign convention of work.
5.3		Define internal energy.
5.3		Calculate change in internal energy given work and heat.
5.3		Explain how the sign of $\Delta U$ (internal energy) indicates the flow of energy.
5.3		Determine PV work.
5.3		State the equation for the relationship between pressure, volume, and work and identify all variables.
5.3		Define enthalpy, endothermic and exothermic.
5.3		Use measurement values from a constant pressure calorimeter to calculate unknown values such as enthalpy change or heat capacity.
5.3		Describe the relationships between the chemical equation and $\Delta H$ of the reaction as the reaction is modified.
5.3		Explain the concept of Hess's Law.
5.3		Use Hess's Laws to determine the enthalpy change of a reaction.
5.3		Define standard state for gas, liquid, solid, or solution.
5.3		Define enthalpy of formation.

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5.3		Express the reaction represented by a given enthalpy of formation.
5.3		Use values of enthalpy of formation to determine the enthalpy change for a given reaction.
5.3		Determine the enthalpy of formation for a substance given the enthalpy change of a reaction.
7.5	#5	Describe the steps of the Born-Haber cycle.
7.5		Write the equation for the energies associated with the Born-Haber cycle.
7.5		Calculate lattice energy using the Born-Haber cycle.
7.5		Calculate the unknown given bond energy values and/or the enthalpy of a reaction.
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		Recognize the enthalpy does not predict spontaneity.
16.2		Define entropy.
16.2		Recognize that entropy is a state function.
16.2		Describe the flow of heat from warmer object to a cooler object using the second law of thermodynamics.
16.2	#6	Predict the relative entropy values of solids, liquids, and gases.
16.2		Recall that the ordinal symbol (°) indicates standard conditions.
16.2		State the standard conditions for solids, liquids, gases, and solutions.
16.2		Define standard molar entropies.
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		State the second law of thermodynamics
16.3		Calculate $DS_{\text{universe}}$ as a function of $DS_{\text{system}}$ and $DS_{\text{surroundings}}$
16.3		Predict the spontaneity of a reaction based on $DS_{\text{universe}}$ .
16.3		Explain the magnitude of $DS$ of the surroundings as a function of temperature.
16.3		Calculate $DS_{\text{surroundings}}$ 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		Define Gibbs Free energy.
16.4		State the equation that relates $DG$ to $DH$ , temperature, and $DS$ .
16.4		Predict the spontaneity of a reaction based on the sign of $DG$ .
16.4		Predict the spontaneity of a reaction based on the sign of $DH$ , the sign of $DS$ , and the magnitude of the temperature.
16.4		Complete calculations with the equation of $DG = DH - TDS$ under non-standard conditions.
16.4		Recognize that Gibbs free energy is a state function.
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 $DG$ of formation.

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16.4		Calculate standard Gibbs free energy based on free energy values for each step in a stepwise reaction.
11.1	#7	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		Define entropy.
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 how the relative strengths of the interactions determine whether or not a solution will form.
11.1		Define the enthalpies of solute, solvent, solution, and mixing.
11.1		Describe the enthalpies of solute, solvent, and mixing as endothermic or exothermic.
11.1		Determine if enthalpy of solution is endothermic or exothermic using Hess' Law.
11.3		Define miscible.
11.3		Distinguish between miscible and nonmiscible (immiscible) substances.
11.3		Define saturated, unsaturated, and supersaturated solutions.
11.3		Explain how to prepare a supersaturated solution.
11.3		Recognize the difference between saturated and supersaturated solutions.
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.
11.4	#8	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 ideal and non-ideal solutions.
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 a semipermeable membrane.
11.4		Describe the flow of solvent particles during osmosis.

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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		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		Complete calculations involving vapor pressure lowering, freezing point depression, boiling point elevation, and osmotic pressure for strong electrolyte solutions.
11.4		Describe differences between isosmotic (isotonic), hyperosmotic (hypertonic), and hyposmotic (hypotonic) solution.
11.4		Use colligative property measurements to determine the molar mass of an unknown solute.
12.1	#9	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		Calculate rate of reaction or change in concentration of substance based on given data.
12.1		Distinguish between instantaneous and average rate of reaction.
12.3		Define rate constant and rate law,
12.3		Write the rate law based on experimental data.
12.3		Recognize that the rate law must be based on experimental data.
12.3		Describe reaction order and overall order.
12.3		Identify the overall order and the order with respect to reactants given a rate law.
12.3		Determine order with respect to a reactant 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		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		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		Complete calculations using the integrated rate laws for zero-, first-, and second-order reactions.
12.4		Derive or memorize the units of the rate constant, k, 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		Complete calculations using the half-life equations 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.
12.5	#10	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		Define reaction mechanism, elementary step, and reaction intermediate.

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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		Write the overall rate law for mechanisms when the second or later step is slow.
12.7		Define catalyst.
12.7		Describe the effect of a catalyst on the energetics of a reaction and the reaction rate.
12.7		Identify the catalyst in a mechanism.
12.7		Distinguish between homogeneous and heterogeneous catalysts.
12.7		Describe the function of an enzyme.
12.7		Label the substrate, enzyme, enzyme-substrate complex, and products in a biological reaction diagram.
13.1	#11	Describe reversible reaction.
13.1		Define dynamic equilibrium.
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		Write the law of mass action for an equilibrium reaction.
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 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		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		Convert between $K_p$ and $K_c$ .
13.2		Identify a 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		Calculate the equilibrium constant using the law of mass action and equilibrium concentrations.
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.

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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	#12	Explain how to use an ICE (initial-change-equilibrium) table.
13.4		Determine equilibrium concentrations based on initial concentrations and a single equilibrium concentration or a change in concentration.
13.4		Define reaction quotient, Q.
13.4		Calculate reaction quotient in terms of concentration or partial pressure.
13.4		Predict the direction of a reaction based on comparison of Q and K.
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.
14.1	#13	List the general properties of acids.
14.1		List the general properties of 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 definition of acids and bases.
14.1		Define conjugate acid and conjugate base.
14.1		Generate the formulas for conjugate acid-base pairs.
14.1		Label substances in a reaction as Bronsted-Lowry acids and bases.
14.1		Define amphoteric.
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		Describe neutral, acidic, and basic solutions in terms of $[OH^-]$ and $[H_3O^+]$ .
14.1		Recall that $K_w$ at 25°C equals 1.0E-14 and varies with temperature.
14.1		Calculate among the values for $K_w$ , $[OH^-]$ , and $[H^+]$ at stated temperature.
14.2		Calculate pH using the given data ( $[H^+]$ , $[OH^-]$ , $K_w$ ).
14.2		Calculate pOH using the given data ( $[H^+]$ , $[OH^-]$ , $K_w$ ).
14.2		Given pH or pOH, find $[H^+]$ and $[OH^-]$ .
14.2		Describe neutral, acidic, and basic solutions in terms of pH and pOH.

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14.2		Recognize that $\text{pH} + \text{pOH} = \text{p}K_w$ .
4.2		Identify the relative amounts of each species for a strong acid or strong base in water.
3.3		Calculate the $[\text{OH}^-]$ of strong bases, considering chemical formula.
14.2		Calculate the pH of a strong acid or strong base solution.
14.2		Determine the pOH and $[\text{OH}^-]$ of basic solutions.
14.3	#14	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		Classify a substance as a weak acid.
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		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 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		Write the reaction for the ionization of a weak base in water.
14.3		Recognize weak bases include $\text{NH}_3$ , related amine compounds (i.e. $\text{CH}_3\text{NH}_2$ ), 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 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		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		State the relationship that $K_a \text{ times } K_b \text{ equals } K_w$ .
14.4		Complete calculations using $K_a \text{ times } K_b = K_w$ .
14.4		Calculate the appropriate K value for a conjugate acid or base.
14.4		Complete calculations using $\text{p}K_a + \text{p}K_b = \text{p}K_w$
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		Classify salt solutions as acidic, basic, or neutral.
14.4		Determine the acidity or basicity of a salt solution when it contains a weakly acidic cation and a weakly basic anion.
14.4		Predict the products of a neutralization reaction.



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14.4		Calculate the amount of each substance remaining after a neutralization reaction occurs.
14.5	#15	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		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		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		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.
14.6		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		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.
14.7	#16	Recall the properties of and how to predict the products of an acid-base titration.
14.7		Define indicator.
		Interpret indicator chart for titrations.
14.7		Distinguish between equivalence point and end point.

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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 strong acid-strong base titration under the following conditions: initial, before equivalence point, at equivalence point, and after equivalence point.
14.7		Calculate the pH in a weak acid-strong base titration under the following conditions: initial, before equivalence point, halfway to the equivalence point, at equivalence point, and after equivalence point.
14.7		Calculate the pH in a weak base-strong acid titration 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	#17	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		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		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.
16.4	#18	Describe how Gibbs free energy is the energy available to do work in a system.
16.4		State the equation that relates DG to $\Delta G^\circ$

OpenStax	Homework	Learning Outcomes
16.4		Calculate the free energy change of a reaction under nonstandard conditions.
16.4		Recognize that $\Delta G$ under nonstandard conditions is zero at equilibrium ( $Q = K$ ).
16.4		State the equation that relates $\Delta G^\circ$ to the equilibrium constant.
16.4		Explain how the sign of $\Delta G^\circ$ is determined by the value of $K$ .
16.4		State the value of $\Delta G^\circ$ when $K = 1$ .
16.4		Predict the direction of spontaneity based on the value of $\Delta G^\circ$ or $K$ .
16.4		Complete calculations between $\Delta G^\circ$ and $K$ .
16.4		Explain the relationship between the temperature and the magnitude of the equilibrium constant.
17.0		Recall rules for assigning oxidation numbers.
17.0		Recall definitions for oxidation, reduction, oxidizing agent, and reducing agent.
17.0		Assign oxidation to all elements in a reaction.
17.1		Identify the oxidation and reduction half-reactions.
17.1		Identify the oxidizing and reducing agents.
17.1		Balance a redox reaction under acidic conditions.
17.1		Balance a redox reaction under basic conditions.
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		Write the half-reactions that occur at the anode and cathode.
17.2		Explain the purpose of a salt bridge.
17.2		Interpret the electrochemical cell notation.
17.2		Describe an inert electrode and when it is used.
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.
17.3		Predict whether a metal will dissolve in acid given the standard reduction potentials.
17.4		State the qualitative relationships between $\Delta G$ , $K$ , and $E^\circ$ of the cell.
17.4		Recognize Faraday's constant.
17.4		State the equation to relate $\Delta G^\circ$ and $E^\circ$ and identify all variables.
17.4		Complete calculations using the equation that relates $\Delta G^\circ$ and $E^\circ$ .
17.4		State the equation that relates $E^\circ$ and the equilibrium constant, $K$ , and identify all variables.

OpenStax	Homework	Learning Outcomes
17.4		Complete calculations using the equation that relates $E^\circ$ and $K$ .
17.4		Determine the moles of electrons ( $n$ ) in an electrochemical reaction.
17.4		Determine $E^\circ$ , $DG^\circ$ , and $K$ based on the standard reduction potentials of an electrochemical reaction.
17.4		Distinguish between standard and nonstandard conditions of an electrochemical cell.
17.4		State the Nernst equation and identify all variables.
17.4		Complete calculations using the Nernst equation for nonstandard conditions.
17.4		Recall the reaction quotient, $Q$ , and how to calculate the value of $Q$ .
17.4		Describe a concentration cell.
17.4		Complete calculations using the Nernst equation for concentration cells.
17.4		Recognize that $Q$ for a concentration cell equals $[\text{lower ion concentration}]/[\text{higher ion concentration}]$
17.4		Describe the flow of electrons in a concentration cell from lower concentration of ions to higher concentration of ions.
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.
17.7		Explain the process that occurs in electroplating a metal onto a surface.
17.7		Recognize that the identity of the surface does not affect the electrochemical reaction.
17.7		Predict the products from the electrolysis of a molten salt.
17.7		Predict the products from the electrolysis of a mixture of molten salts.
17.7		Recognize that the standard reduction potentials from aqueous solutions follow the same trends but have different magnitudes in molten salts.
17.7		Predict the products from the electrolysis of an aqueous salt solution.
17.7		Calculate between time and amount of substance plated using given amps for an electrolytic reaction.
21.0		Recall the isotopic symbol notation.
21.1		Define radioactivity, nuclear equation, nuclide, parent nuclide, and daughter nuclide.
21.1		Describe characteristics that make an isotope unstable.
21.1		Describe how the "strong force" holds nucleons together.
21.2		State the five modes of radioactive decay.
21.2		Write balanced nuclear equations for each type of radioactive decay.
21.2		List the symbols used for the radioactive particles.
21.2		Distinguish between ionizing and penetrating radiation.
21.3		Recall the equations and graphs for first order kinetics of chemical reactions.
21.3		Recognize that nuclear reactions follow first order kinetics.
21.3		Complete calculations for nuclear reactions considering that different quantities can be used in place of concentration.
21.3		Define radiocarbon dating.
21.3		Describe the types of samples for which radiocarbon dating can be used.
21.3		Describe the upper time limit for which radiocarbon dating can be used.
21.3		Describe uranium/lead dating to estimate the age of a rock.
21.3		Recognize that uranium/lead dating can date objects that are billions of years old.
21.3		Calculate the age of a rock using the lead-to-uranium mass ratio.
21.4		Define fission.
21.4		Define chain reaction and critical mass.
21.4		Distinguish how fission is used in a nuclear bomb vs nuclear reactor.

<b>OpenStax</b>	<b>Homework</b>	<b>Learning Outcomes</b>
21.4		Define fusion.
21.4		Distinguish how fusion could be used in a nuclear bomb or as a potential energy source.
21.4		Distinguish between fission and fusion as an energy source in terms of amount of energy produced and products.
21.6		Recognize that radioactivity is measured in different ways with varying units.