Chapter 10 – Chemical Reactions

1. Describe four evidences for a chemical reaction.
2. Describe the parts of a chemical equation.
4. Identify the four symbols used to specify physical states for substances in a reaction.
5. Know how to use coefficients to properly balance an equation.
6. Understand the five classifications of chemical reactions.
7. Use the activity series to predict which single replacement reactions are possible.

Chapter 11 – The Mole

1. Define the mole and what it is used for.
2. Identify Avogadro’s number and its significance to chemistry.
3. Convert particles to moles/ moles to particles.
4. Describe how the molar mass of an element and a compound is determined.
5. Convert masses to moles/ moles to masses.
6. Define molar volume and the abbreviation STP.
7. Convert volumes to moles/ moles to volumes for gases (at STP).

Chapter 12 – Stoichiometry

1. Define the term stoichiometry and describe what it is used for.
2. Describe the significance of the coefficients in a balanced equation (what do they stand for?).
3. Show the process of determining the amount of product made/reactant required given a quantity of reactant.
4. Describe the terms limiting reactant and excess reactant and their significance in chemical reactions.
5. Define the term percent yield and what is used in this calculation.
6. Define the term stoichiometric proportions.

Chapter 13/14 – Gases

1. List and explain the five properties of gases.
2. Describe the instruments used to measure pressure of the atmosphere and the pressure of an enclosed gas.
3. Describe what was measured in Boyle’s experiment and Charles’ experiment.
4. State and describe Dalton’s law of partial pressure.
5. List the postulate of the kinetic molecular theory.

Chapter 15 – Solutions

1. Define a solution and name its components.
2. Describe different types of solutions: solid, liquid, and gas.
3. Distinguish between electrolytes and non-electrolysis.
4. Define the terms saturated, unsaturated and supersaturated.
5. Explain how the process of dilution is carried out.

Chapter 16 – Energy

1. Describe the different forms of energy.
2. Explain the effect of making and breaking bonds in terms of energy changes.
3. Characterize reactions as endothermic or exothermic and create a graph from the sign of $\Delta H$. 
4. Explain a calorimeter and how it is used.
5. Describe how energy is involved in these phase changes: fusion, freezing, vaporization, and condensation.

Chapter 18 - Equilibrium

1. Explain how reactions occur using the collision theory
2. Identify the enthalpy change, activation energy, and activated complex on an energy diagram
3. List and explain five factors that determine the rate of a reaction
4. To give two descriptions for a reaction at equilibrium
5. To interpret the meaning of a large and a small equilibrium constant
6. To describe Le Chatelier’s principle and predict how various changes effect the equilibrium position

Chapter 19 – Acids/Bases

1. Discuss properties of acids and bases and give two definitions for each.
2. Explain the role of an indicator.
3. Define the term amphoteric.
4. Define the term diprotic.
5. Distinguish between strong acids and weak acids.
6. Identify the conjugate acid base pair for a solution of nitrous acid, HNO₂, in water.
**CP Chemistry Final Review – Chap. 10-19**

**Calculations**

**Chapter 10 – Chemical Reactions**

1. Write an equation for the following reaction: potassium oxide reacts with water to produce potassium hydroxide
2. Write an equation for the following reaction: carbon dioxide and water are formed when carbonic acid decomposes
3. Balance and classify the reaction: \(\text{NiSO}_4 + \text{Li}_3\text{PO}_4 \rightarrow \text{Ni}_3(\text{PO}_4)_2 + \text{Li}_2\text{SO}_4\)
4. Balance and classify the reaction: \(\text{Al} + \text{S}_4 \rightarrow \text{Al}_2\text{S}_3\)

**Chapter 11 – The Mole**

1. Find the mass of the following: 1 atom of iron, 3 molecules of carbon monoxide, and 1 formula unit of \(\text{CaCl}_2\)
2. Find the mass of the following: 1 mole of iron atoms, 3 moles of carbon monoxide molecules, and 1 mole of \(\text{CaCl}_2\)
3. Determine the molar mass of \(\text{KMnO}_4\), \(\text{Sr(NO}_3)_2\), carbon dioxide, and sodium phosphate
4. Convert 85.3 g of \(\text{NaOH}\) to moles of \(\text{NaOH}\)
5. Convert 27.9 L of \(\text{H}_2\) (g) to grams
6. Convert 101.2 g of \(\text{CH}_3\text{OH}\) to molecules

**Chapter 12 – Stoichiometry**

1. How many moles of nitrogen gas will react with 2.45 moles of hydrogen gas to form ammonia (\(\text{NH}_3\))? How many moles of ammonia will form in the reaction above?
2. What mass of \(\text{NaCl}\) is produced when \(\text{Cl}_2\) reacts with .29 g of \(\text{NaI}\)? \(2\text{NaI} + \text{Cl}_2 \rightarrow 2\text{NaCl} + \text{I}_2\)?
3. What mass of \(\text{Al}\) would produce 1.32 L of \(\text{H}_2\) at STP from this reaction: \(2\text{Al} + 3\text{H}_2\text{SO}_4 \rightarrow \text{Al}_2(\text{SO}_4)_3 + 3\text{H}_2\)?
4. Identify the limiting reactant when 10.0g \(\text{H}_2\text{O}\) reacts with 4.5 g \(\text{Na}\) to produce \(\text{NaOH}\) and \(\text{H}_2\). What mass of \(\text{H}_2\) is formed?
5. How many grams of \(\text{H}_2\text{O}\) and \(\text{NaCl}\) form when 12.5 g \(\text{HCl}\) reacts with 53.2 g of \(\text{NaOH}\)?

**Chapter 13/14 – Gases**

1. Convert 4.2 atm into mmHg, torr, Pa, and PSI.
2. The side of a manometer open to the atmosphere is 100 mm higher than the side open to a gas sample. Assuming that atmospheric pressure 780 mm Hg, determine the pressure of the gas sample.
3. The gas pressure in a 20-L tank is 4.8 atm. What is the new pressure if the temp. is raised from 100°C to 350°C?
4. What is the temperature of a 2.4 mole gas sample at a pressure of 1.5 atm and a volume of 35.0 L?
5. What is the total pressure of 4.5 mol He and 12.7 mol \(\text{O}_2\) in an 8.0 L tank at a temperature of 10°C?
6. What is the pressure of \(\text{CO}_2\) when 13.7 g of \(\text{CaCO}_3\) is decomposed, assuming the gas is in a 500 mL container at 25°C? (The other product is \(\text{CaO}\)).

**Chapter 15 – Solutions**

1. 125g of acetic acid is added to enough water to prepare a 1.50 L solution. What is the molarity?
2. What is the molality of a solution containing 0.125g of chromium and 81.3g of iron?
3. A gas mixture contains 45.6g \(\text{CO}\), and 899g \(\text{CO}_2\). What is the mass percent of carbon monoxide?
4. What is the concentration of the solution formed by adding 10.0 mL of a 0.582 M solution to water to make 250 mL solution?
5. What mass of \(\text{PbS}\) is formed by adding 25 mL of 0.125 M \(\text{Pb(NO}_3)_2\) to excess sulfide ions? \(\text{Pb(NO}_3)_2 + \text{S}^{2-} \rightarrow \text{PbS}\)
Chapter 16 – Energy

1. How much heat is produced when 14 g of methane is burned? \( \text{CH}_4 + \text{O}_2 \rightarrow \text{CO}_2 + \text{H}_2\text{O} \ \Delta H = -486 \text{ kJ/mol}. \)

2. How much heat is transferred when 14g of NO\(_2\) is dissolved in 100.g water.

\[ 3\text{NO}_2 + \text{H}_2\text{O} \rightarrow 2\text{HNO}_3 + \text{NO} \ \Delta H^\circ = -138\text{kJ}? \]

3. What is the specific heat of aluminum if the temperature of 28.4g Al increases by 8.1˚C when 207 J of heat is added?

4. The temperature drops from 23.4˚C to 19.7˚C when 13.7g of Pb(NO\(_3\)_2) dissolves in 85.0g water. What is \( \Delta H \) for the process?

5. How much heat is required to melt 50. g of ice? \( \text{H}_2\text{O} (s) \rightarrow \text{H}_2\text{O} (l) \ \Delta H = +6.02 \text{ kJ/mol}. \)

Chapter 18 - Equilibrium

1. Write the equilibrium expression for this reaction: \( 2\text{NO}(g) + \text{Br}_2 (g) \rightleftharpoons 2\text{NOBr}(g) \).

2. \( \text{H}_2(g) + \text{I}_2(g) \rightleftharpoons 2\text{HI}(g) \ K_{eq} = 50.5 \). Calculate Q when equilibrium concentrations are: \([\text{H}_2]=0.150\text{M}, [\text{I}_2]=.175\text{M}\) and \([\text{HI}]=.950\text{M} \). Which way will the reaction proceed?

3. Write the expression for the solubility product for SrSO\(_4\). Write the expression for the solubility product for MgCl\(_2\).

4. GaF\(_3\) is added to pure water. The concentration of Ga\(^{3+}\) is .049M at equilibrium. What is the value of \( K_{sp} \) for GaF\(_3\)?

5. What are the equilibrium concentrations of dissolved ions in a saturated solution of Ag\(_2\)SO\(_4\) \( (K_{sp} = 1.2 \times 10^{-5}) \)

Chapter 19 – Acids/Bases

1. What is the concentration of H\(_3\)O\(^+\) ions formed by 1.5 \( M \) HNO\(_2\) \( (K_a = 4.0 \times 10^{-4}) \) in water?

2. What is the concentration of OH\(^-\) ions formed when 40.0 g of NaOH is dissolved in 250 mL H\(_2\)O?

3. Write an equation showing how the weak acid HF reacts with the water it is dissolved in. Label acid/base.

4. An aqueous solution has a [H\(_3\)O\(^+\)] = 4.8 \times 10^{-5}. What is [OH\(^-\)]? Is the solution acidic or basic? \( (K_w = 1.00 \times 10^{-14}) \)

5. What is the pH of a solution that has a [H\(_3\)O\(^+\)] = 7.3 \times 10^{-8}? What is the pOH of that solution?