

Review for the Summer Final**STUDY LIST****Chapter 1 – Matter and Measurement**

- Know the difference between homogeneous and heterogeneous mixtures
- Know the difference between elements and atoms
- Know the difference between ions and molecules
- Know how to distinguish physical properties
- Know the density formula and how to use it to calculate density, mass and volume
- Know the difference between Celsius and Kelvin; convert from one to the other
- Know the difference between intensive and extensive properties
- Know your safety rules
- Know the difference between physical and chemical changes
- Know the SI units
- Know all the prefixes and their values
- Know how to do dimensional analysis
- Know how to do scientific notation
- Know the difference between precision and accuracy, and experimental error
- Know significant figures of measurements and in calculations
- Know how to do percents

Chapter 2 – Atoms and Elements

- Know the locations of Group 1A (Alkali Metals), Group 2A (Alkaline Earth Metals), Group 3A-Group 6A, Group 7A (Halogens) and Group 8A (Noble Gases/inert gases)
- Know the 7 diatomic elements

Chapter 3 – Molecules and Compounds

- Know the terms chemical formula, empirical formula, molecular formula, and structural formula
- Know your ions (monatomic ions and polyatomic ions)

- Know how to name ions in the traditional (-ic, -ous) and Stock system (II, III)
- Know the terms cations and anions
- Write and name ionic compounds
- Write and name nonmetal (molecular) compounds (mono-, di-, tri-, etc.)
- Know common names of binary compounds
- Know definition of mole and Avogadro's Number (6.02×10^{23} molecules in one mole)
- Calculate the molar mass of a substance
- Know how to convert from moles to mass(g), molecules, volume(L) using conversion factor
- Know how to calculate % composition
- Know how to calculate empirical and molecular formulas from mass percent
- Define hydrated compounds
- Know how to determine the formula of a hydrated compound from experimental data

Chapter 4 – Chemical Equations and Stoichiometry

- Define combustion
- Recognize products and reactants
- Be able to write combustion equations
- Know when to label the substances solid (s), gas (g), liquid (l), or aqueous (aq)
- Know how to balance equations
- Be able to find the molecular and empirical formulas and differentiate between the two
- Know how to convert mass and moles (i.e. $1 \text{ mol} = 22.4 \text{ L} = 6.02 \times 10^{23}$)
- Know Molar Mass (grams / mol or $\text{g}\cdot\text{mol}^{-1}$)
- Know how to use the stoichiometric factor (♥ of the problem) (i.e. $1 \text{ mol CO}_2 = 1 \text{ mol C}$)
- Define limiting reactant and excess reactant
- Know how to determine the limiting reactant and excess reactant

- Solve problems involving Limiting Reactants
- Given the actual yield, know how to find the theoretical yield and the percent yield
- Know how to find the mass of each element of a compound
- Know how to find empirical and molecular formula using stoichiometry

Chapter 5 – Reactions in Aqueous Solution

Properties of Aqueous Solutions

- Define **solute**, **solvent**, and **solution**. Give examples
- Give **operational** and **theoretical** definitions of **electrolytes**
- Know that soluble ionic compounds and strong acids are **strong electrolytes**. Ionic compounds of low solubility [e.g. $\text{Mg}(\text{OH})_2$] and weak acids/bases are **weak electrolytes**.
- Know that molecular compounds (except acids) are **non-electrolytes**
- Know that **alcohols** (e.g. CH_3OH) are **not ionic hydroxides** (and are not bases). Bases are usually **metallic hydroxides**
- Know the **solubility rules**. State whether an ionic compound is soluble in water

Precipitation Reactions

- Know that precipitation reactions are **double replacement** reactions that produce an insoluble product
- Given two ionic compounds in solution, correctly **determine the products**. (Know your ions)
- Determine which product(s) is/are **precipitates**. Use **(aq)** and **(s)** symbols
- Correctly write the **ions** in a soluble ionic compound. [e.g. $\text{CaCl}_2(\text{aq})$ becomes $\text{Ca}^{2+} + 2\text{Cl}^-$]
- Identify **spectator ions**
- Write **molecular, detailed ionic, and net ionic** equations for a precipitation reaction

Acids and Bases

- Give **operational** (cabbage juice) and **theoretical** definitions of acids and bases.
- Know that acids increase the **H^+ ion** concentration in an aqueous solution. (Theoretical definition)
- Memorize** the 8 strong acids.
- Know that acids are **molecular compounds** that **form ions** when in aqueous solution.
- Name acids** according to their anion.
- ide** → hydro__ic acid; **ate** → __ic acid;
- ite** → __ous acid;
- sulfur: add “ur”; phosphorus: add “or”]
- Know that bases increase the **OH^- ion** concentration in an aqueous solution. (Theoretical definition)
- Memorize the eight **soluble hydroxides** (except NH_4OH) that are the **strong bases**.
- Understand that **ammonia(aq)**,

$$\text{NH}_3 + \text{H}_2\text{O} \rightleftharpoons \text{NH}_4^+ + \text{OH}^-$$
forms a **weak basic** solution.
- Know that **acids react with bases** to form H_2O and a salt. (**Neutralization**)
- Write **equations** for acid-base reactions including NH_3 (ex on page 199) as the base.
- Know that **strong acids** and **strong bases** are **written as ions** in the ionic equations.

Gas Forming Reactions

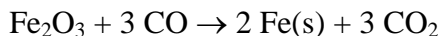
- Recognize the six products that turn into gases. Memorize the gases formed.

Organizing Reactions in Aqueous Solution

- Know the **three examples** of double replacement reactions and each driving force:
- Precipitate reactions form an insoluble product. Acid-Base reactions form water (a very weak electrolyte therefore, a very stable product). Gas-forming reactions form a gas.
- Know that a **driving force** is something that keeps the new combinations of ions from reforming the old combinations of ions.

- Oxidation-Reduction** is a fourth type of reaction driven by the **transfer of electrons**.

Oxidation-Reduction Reactions



- Know that an important type of reaction gets its **name** from atoms that combine with **oxygen**. During the refining of iron, carbon monoxide combines with oxygen (from the iron ore), $\text{CO} \rightarrow \text{CO}_2$ and is **oxidized**. Large masses of iron ore (Fe_2O_3) are **reduced** to a smaller amount of iron metal.
- Understand that since CO helps the iron ore to be **reduced**, CO is called the **reducing agent**. Since Fe_2O_3 causes the C to be **oxidized**, iron ore is called the **oxidizing agent**. Whatever is **oxidized** acts as the **reducing agent**. What-ever is **reduced** acts as the **oxidizing agent**.
- Mnemonics to help: **GROL**; **LeO** the lion says **GeR**; and **OIL RIG**
- Be able to **assign oxidation numbers** to any atom in any substance. Rules on page 207.
- Recognize **redox** reactions because **oxidation numbers change**. (# \uparrow = oxidation / # \downarrow = reduction), electrons are gained or lost.
- Know several common oxidizing agents and reducing agents and what they turn into.

Measuring Concentrations of Solutions

- Know the definition of **molarity, M**, as one way to communicate **concentration** of solute.
- Know that the symbol **[X]** means the concentration of X in moles/Liter ($\text{mol}\cdot\text{L}^{-1}$).
- Be able to determine the concentration of **ions** in an ionic compound.
- For example, in 0.25 M AlCl_3
 $[\text{AlCl}_3] = 0.25 \text{ M}$
 $[\text{Al}^{3+}] = 0.25 \text{ M}$ $[\text{Cl}^-] = 0.75 \text{ M}$
- Use the molarity formula to calculate **moles, mass, volume, or molarity** of a solution.
- Know that **Volume x Molarity = moles** of solute. **Dilution** problems use $V_i M_i = V_f M_f$.

- Describe how to **make a solution** correctly. Know what a **volumetric flask** is.

Stoichiometry of Reactions in Aqueous Solution

- Use **molarity** as a **conversion factor**.
- Know that **titration** is a technique called **quantitative chemical analysis** because you are measuring. It is also called **volumetric analysis** (because you are measuring volumes).
- Understand the terms **indicator, neutralization, and equivalence point**.
- Know common indicators such as **phenolphthalein** for titrations with strong bases.

Chapter 6 – Energy & Chemical Reactions Driving Forces

I can...

- state the sign of ΔH based on observation of warming or cooling of the surroundings.
- correctly apply the terms exothermic and endothermic to situations where the surroundings are warming or cooling.
- draw a PE curve (uphill or downhill) based on information about warming or cooling of the surroundings.

Measuring Heat

- state the units of heat capacity, specific heat, and molar heat capacity as well as their significance.
- convert between the heat units of calories and Joules. ($4.184 \text{ J} = 1 \text{ calorie}$)
- use calorimetry ($q = mC\Delta T$) to calculate heat changes during temperature changes.
- calculate the heat transferred when two objects, at different temperatures, come into contact.

Energy = Heat and Work

- state the difference between work and heat energy.
- state the difference between system and surroundings.
- recognize the system and the surroundings in a chemical or physical system.
- calculate the change in internal energy based on changes in heat absorbed by the system and work done by the system.
- state that ΔH is a more general (and useful) measure of energy than ΔE and that $\Delta H = q$ when a reaction occurs at constant pressure.

Chemical Work = Expanding Gases

- relate physical work ($w = F \cdot d$) and chemical work ($w = P \cdot \Delta V$).
- calculate **PV work** done by an expanding gas.
- state that no work is done in a **constant volume** situation such as a bomb calorimeter.

Calculating ΔH -- Hess's Law

- state the definition of a state function.
- list examples of properties that are and are not state functions.
- write the equation for the **heat of formation** of a substance.
- state that the heat of formation of an **element** under standard conditions has a value of zero.
- use **Hess's Law** to calculate the energy of a chemical or physical change.

Calculating Heat During Phase Changes – Heats of Fusion and Vaporization

- use heats of vaporization or heats of fusion to calculate heat changes during phase changes.
- write an equation showing the heat of fusion or heat of vaporization.
- Sketch a Heating Curve and for each segment, list the energy change, the phase or phases present, and the type of calculation needed to determine the energy change.

Labs & Demonstrations

- Eye Safety Demonstration (denaturation)
- Chromatography Lab
- Density Lab
- Hydrate Demo/Lab
- Handboiler Distillation
- Flame Test Demo
- MicroMole Rockets Lab
- Decomposition of HgO
- Precipitation Lab
- Baggie Lab & Filtration
- Testing for Carbonates with Acid
- Cabbage Lab & Electrolytes
- Introduction to Redox ($\text{CuCl}_2 + \text{Al}^\circ$)
- Molarity Demo/Lab
- Dry Ice & Acid-Base Indicators
- Steel Wool & Vinegar—Exothermic Rxn
- Exothermic & Endothermic Bags
- Heat of Fusion Lab