Chemistry Fourth Marking Period Review Sheet
Spring, Mr. Wicks
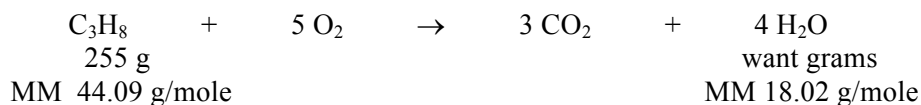
Chapter 11: Chemical Reactions
- Chemical equations use symbols to represent chemical reactions that take place in a laboratory. I can write chemical equations including the states of matter, yield symbol, and symbol for heat.
- I can balance simple chemical equations by inspection and complicated chemical equations by the fraction method.
- I remember that balanced chemical equations have both material balance and charge balance. In addition, the lowest whole-number ratio is used for the set of chemical coefficients.
- I can explain why balanced chemical equations obey the law of conservation of mass.
- I can convert the symbols for a chemical equation into an English sentence and vice versa.
- I can compare and contrast the following types of reactions:
  1. Combination (synthesis) reactions—one product is formed.
  2. Decomposition reactions—one reactant is present.
  3. Single replacement reactions—one metal replaces a less active metal or hydrogen; one halogen replaces a less active halogen. Use an activity series to determine whether or not the reaction will take place.
  4. Double replacement (metathesis) reactions—pairs of ions exchange partners to form new products. Examples include precipitation reactions and acid-base neutralizations.
  5. Combustion—a fuel and oxygen react to form carbon dioxide and water.

Chapter 12: Stoichiometry
- Reaction stoichiometry refers to the set of coefficients relating amounts of reactants and products in a chemical equation. I can write mole ratios from the coefficients to use as conversion factors for stoichiometric calculations.
- I can use stoichiometry to interpret a chemical equation on a microscopic (molecular) level and a macroscopic (molar) level.
- I recall that both mass and number of atoms of each element are conserved in every chemical reaction and equation.
- I can calculate the mass (or moles) of one reactant or product from the mass (or moles) of another reactant or product in a balanced chemical equation.

Example: Determine the grams of water, H_2O, that form when 255 g of propane, C_3H_8, reacts given the following chemical equation.

$$C_3H_8 \quad + \quad 5\,O_2 \quad \rightarrow \quad 3\,CO_2 \quad + \quad 4\,H_2O$$

255 g   want grams
MM 44.09 g/mole   MM 18.02 g/mole

$$\left(255g\,C_3H_8\right)\left(\frac{mole\,C_3H_8}{44.09g\,C_3H_8}\right)\left(\frac{4moles\,H_2O}{1mole\,C_3H_8}\right)\left(\frac{18.02g\,H_2O}{mole\,H_2O}\right)=417g\,H_2O$$

- I can determine which reactant is the limiting reactant in a balanced chemical equation. I can also determine the amounts of products formed and the amount of excess reactant leftover.

- I can distinguish between actual yield, theoretical yield, and percent yield.
- I can calculate theoretical yield and percent yield. $Percent\ Yield = \left(\dfrac{Actual\ Yield}{Theoretical\ Yield}\right)(100)$

Chapter 14: The Behavior of Gases
- I can describe the characteristics of ideal gases and how real gases are different from them. Ideal gas molecules are assumed to have no volume and to not interact with one another by intermolecular forces. Deviations from ideal behavior are greatest at high pressure and low temperature.
- Ideal gas characteristics:
  1. Gases are composed of very tiny molecules with lots of empty space between them.
  2. Gas molecules move rapidly, move in straight lines, and travel in random directions.
  3. Gas molecules do not attract each other.
  4. Gas molecules have elastic collisions. In other words, the total kinetic energy of the molecules before and after a collision is the same.
  5. The average kinetic energy of gas molecules is proportional to the absolute (Kelvin) temperature.
- I can apply the kinetic-molecular theory of gas behavior at the molecular level.
- In kinetic theory, the molecules of a gas are imagined to be a large number of points bouncing off the walls of a container. Gas pressure is related to the number of collisions that occur with the walls per unit time.
- I understand how pressure measurements are made.
- I can convert between various pressure units: 1atm = 760 mm Hg = 760 torr = 101.3 kPa = 14.7 psi
- I know and can use the standard conditions of temperature and pressure (STP): 0^oC (273 K) and 1 atm.
- I recall that all temperature-related gas laws use Kelvin rather than Celsius temperature, and I can convert between Kelvin and Celsius temperatures.
- I know that absolute zero on the Kelvin scale is the temperature at which all atomic and molecular motion stops.
- I can compare and contrast Boyle's law, Charles' law, Gay-Lussac's law, Avogadro's law, the combined gas law, the ideal gas law, Dalton's law, and Graham's law, which are presented in Table 1 along with values for the gas law constant, R.
- I can use the standard molar volume of a gas at STP to solve problems. The volume occupied by one mole of an ideal gas at standard conditions of temperature and pressure (STP, 0^oC and 1 atm) is 22.4 L/mole.
- I can calculate gas densities and molar masses by substituting $n = \dfrac{m}{MM}$ into the ideal gas law, $PV = nRT$, to get $d = \dfrac{m}{V} = \dfrac{PMM}{RT}$ and $MM = \dfrac{mRT}{PV} = \dfrac{dRT}{P}$.
- I can apply the gas laws when solving stoichiometry problems involving gaseous materials.

Table 1: Gas Laws		
Gas Law Name	*Equation*	*Quantities Held Constant*
Boyle's Law:	$P_1V_1 = P_2V_2$	Temperature and moles
Charles' Law: (Absolute zero, 0 K, can be determined by extrapolating a line to zero volume on a V-T diagram.)	$\dfrac{V_1}{T_1} = \dfrac{V_2}{T_2}$	Pressure and moles
Gay-Lussac's Law: (Absolute zero, 0 K, can be determined by extrapolating a line to zero pressure on a P-T diagram.)	$\dfrac{P_1}{T_1} = \dfrac{P_2}{T_2}$	Volume and moles
Avogadro's Law: (Stoichiometric coefficients can be used to compare volumes of different gases in a chemical equation,)	$\dfrac{V_1}{n_1} = \dfrac{V_2}{n_2}$	Temperature and pressure
Combined Gas Law: (Combines Boyle's, Charles', and Gay-Lussac's Laws)	$\dfrac{P_1V_1}{T_1} = \dfrac{P_2V_2}{T_2}$	Moles
Ideal Gas Law: (Combines Boyle's, Charles', and Avogadro's Laws)	$PV = nRT$ where $R =$ 0.0821 L-atm/(mole-K) $=$ 8.314 J/(mole-K) $=$ Gas Law Constant	None
Dalton's Law: Total pressure is equal to the sum of the partial pressures.	$P_{Total} = P_1 + P_2 + P_3 + ...$	Not applicable

Chapters 13, 15, and 16: Liquids, Solids, and Solutions
- I can compare and contrast the properties of gases, liquids, and solids.
- I can apply the Kinetic Molecular Theory of Matter to explain the physical properties of gases, liquids, and solids. Focus on volume, shape, density, and molecular motion.
- I can compare and contrast the intermolecular forces of attraction for molecular compounds. See Table 2.
- I can use intermolecular forces of attraction to explain relative melting points, boiling points, surface tensions, viscosities, specific heats, and vapor pressures for a given series of molecular compounds.

Table 2: Intermolecular Forces of Attraction ("van der Waals forces")			
Type	*Strength*	*Attraction*	*Examples*
Hydrogen bonding	Strongest	"H" on one molecule is attracted to "F," "O," or "N" on an adjacent molecule.	H_2O, NH_3, HF, base pairs in DNA
Dipole-dipole forces	Intermediate	Polar molecules attract each other like tiny magnets.	HCl, H_2S
London dispersion forces	Weakest	Induced dipoles attract each other for a fraction of a second. All covalent compounds have London dispersion forces.	CH_4, CCl_4

- I can explain why water has "surprising" physical properties
  - Water is polar because it has polar bonds and bent geometry. It is also capable of hydrogen bonding, which results in several unusual properties.
    - Water has a higher melting point and boiling point than expected.
    - Water has considerable surface tension.
    - Water has a high capacity to absorb heat.
    - Ice is less dense than liquid water and will therefore float.
  - In addition to these properties, polar and ionic substances will dissolve in water.
- I can compare and contrast solutions, colloids, and suspensions. See Table 3.
- I can give examples of colloids from everyday life.
  - Emulsions—cheese, butter
  - Foams—whipped cream, shaving cream
  - Gels—gelatin
  - Aerosols—smoke, fog, mist
- I can describe the Tyndall effect—when a beam of light is passed through a substance, light scattering occurs. Ex. The beam of light from auto headlights is visible during a fog but not during other conditions.

Table 3: Contrasting Solutions, Colloids, and Suspensions			
	Solution	*Colloid*	*Suspension*
Example:	NaCl dissolved in water	Milk	Sand mixed with water
Settles out?	No	No	Yes
Can be separated by filtration?	No	No	Yes
Shows the Tyndall effect?	No	Yes	Yes

- I can distinguish between solute, solvent, solution, and solubility.
- I remember that solubility is defined as the maximum amount of a solute that will dissolve in a given quantity of solvent at a specific temperature.

- I can describe general solubility trends:
  - As temperature increases, the solubility of a solid in water increases.
  - As temperature increases, the solubility of a gas in water decreases.
  - As pressure increases, the solubility of a gas in water increases. Ex. Soda pop effervescence.
- I can use both words and graphs to compare and contrast saturated, unsaturated, and supersaturated solutions.
- I can distinguish between miscible and immiscible liquids.
- I can clearly explain and apply the solubility rule "like dissolves like," which means that the solute and solvent are alike in polarity:
  - Nonpolar solutes will dissolve in nonpolar solvents.
  - Polar and ionic solutes will dissolve in polar solvents.
- I can explain how soap works. Soap molecules have ionic "heads" that dissolve in water and nonpolar "tails" that dissolve in grease. Thus, soap and water can remove grease despite the fact that water and grease, which are polar and nonpolar compounds, respectively, will not dissolve in each other.
- I can calculate the molar concentration of ionic substances dissolved in water using

$$Molarity = \left(\frac{Moles\ of\ Solute}{Liters\ of\ Solution} \right)$$

- I can safely prepare solutions of various molar concentrations:
  1. from solids.
  2. by diluting liquids. (Recall $C_1V_1 = C_2V_2$ where C = concentration and V = volume.)
- I can use the solubility rules given in Table 4 to predict the solubility of ionic compounds in water.
- I can write the ions formed when an ionic compound dissolves in water.
- I can predict products for precipitation reactions (double replacement reactions).
- I can write molecular, ionic, and net ionic equations, and I can identify spectator ions.
- I can compare and contrast the four colligative properties vapor-pressure lowering, boiling-point elevation, freezing-point depression, and osmotic pressure.

Table 4: Solubility Rules for Common Ionic Compunds in Water at 25°C	
General Rule	*Exceptions to the Rule*
Almost all compounds containing alkali metal ions (Li^+, Na^+, K^+, etc.) and NH_4^+ are soluble	
Almost all compounds containing nitrates (NO_3^-), bicarbonates (HCO_3^-), and chlorates (ClO_3^-) are soluble	
Most compounds containing halides (Cl^-, Br^-, and I^-) are soluble	Halides of Ag^+, Hg_2^{2+}, and Pb^{2+}
Most sulfates (SO_4^{2-}) are soluble	Sulfates of Ag^+, Ca^{2+}, Sr^{2+}, Ba^{2+}, Hg_2^{2+}, and Pb^{2+}
Carbonates (CO_3^{2-}), phosphates (PO_4^{3-}), chromates (CrO_4^{2-}), and sulfides (S^{2-}) are usually insoluble	Compounds containing alkali metal ions (Li^+, Na^+, K^+, etc.) and NH_4^+
Hydroxides (OH^-) are usually insoluble	Hydroxides containing alkali metal ions (Li^+, Na^+, K^+, etc.) and Ba^{2+}

Chapter 19: Acids, Bases, and Salts
- I can compare and contrast the properties and general characteristics of acids and bases.
- I can give examples of common acids and bases.
- I can write equations for acid ionization and base dissociation.
- I can write molecular, ionic, and net ionic equations for acid-base neutralizations (double replacement reactions).
- I know that when acid-base reactions form "salts," this does NOT mean that they all form table salt, NaCl. The word "salt" in this context refers to an inorganic compound whose cation comes from a base and whose anion comes from an acid.