Name: **Honors Chemistry**

Section: Stoichiometry Review Sheet

After studying Chapter 9-2, you should be able to:

- Interpret balanced chemical equations in terms of interacting moles, representative particles, masses, and volumes (at STP).
- Construct mole ratios from balanced chemical equations and apply these ratios in calculating mole-mole stoichiometric quantities.
- Calculate stoichiometric quantities from balanced chemical equations using units of mass.
- Calculate stoichiometric quantities from balanced chemical equations using units of moles, mass, representative particles, and volume (gases at STP).
- Identify the limiting reagent in a reaction and use it to calculate stoichiometric quantities and the amount of excess reagent(s).
- Calculate the theoretical yield, actual yield, and/or percent yield for a chemical reaction.
- Construct equations that show the heat changes for chemical and physical processes.
- Determine the heat of reaction for a chemical reaction in which a specified amount of substance is involved.

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1. What is meant by the term "mole ratio"? Give an example of a mole ratio. How is the mole ratio used in solving problems?

2. Consider the reaction represented by the unbalanced equation: $NH_3 + O_2 \rightarrow NO + H_2O$
   a) For every molecule of NH_3 that reacts, ____ molecules of O_2 are required.

   b) For every 1.00 mol of NH_3 that reacts, ____ mol of O_2 is required.

3. The compound freon-12 (CCl_2F_2) was formerly used as a coolant in refrigerators and air conditioners. It is produced through the following *unbalanced* equation:
   $CC\ell_4 + SbF_3 \rightarrow CC\ell_2F_2 + SbC\ell_3$
   If 1800 grams of freon need to be produced, calculate the mass required for each reactant.

4. Zinc oxide can be prepared industrially by treating zinc sulfide with oxygen. The by-product is sulfur dioxide. An engineer expects to obtain a 78% yield of zinc oxide by this process. How much zinc sulfide should the chemical plant have on hand to prepare 2.0×10^4 kg of zinc oxide? Start by writing a balanced chemical equation for the reaction. (3.1×10^4 kg ZnS)

5. A 60.0 mL sample of aqueous $Ca(OH)_2$ requires 38.44 mL of 0.0975 M nitric acid, HNO_3, for neutralization.
    a) Write a balanced chemical equation for this reaction.

    b) Calculate the concentration of the original solution of calcium hydroxide.

6. When a limiting reactant is present, in what way is the reaction "limited"? What happens to a reaction when the limiting reactant is used up?

7. Consider the reaction $Mg(s) + I_2(s) \rightarrow MgI_2(s)$
Identify the limiting reagent in each of the reaction mixtures below:
    a) 100 atoms of Mg and 100 molecules of I_2
    b) 150 atoms of Mg and 100 molecules of I_2
    c) 0.16 mol Mg and 0.25 mol I_2
    d) 0.12mol Mg and 0.08 mol I_2
    e) 6.078 g Mg and 53.455 g I_2
    f) 1.00 g Mg and 2.00 g I_2

8. Chlorine gas is a very reactive substance and will combine with most metals. For example,
$2\ K(s) + Cl_2(g) \rightarrow 2\ KCl\ (s)$
$Ca(s) + Cl_2(g) \rightarrow CaCl_2(s)$
$2\ Al(s) + 3\ Cl_2(g) \rightarrow 2\ AlCl_3(s)$

    a) Suppose individual 25.0 g samples of these three metals are reacted with separate 50.0 g samples of $Cl_2(g)$. In each case, determine whether the metal or chlorine is the limiting reactant, and calculate the theoretical yield for each process.

    b) Suppose you run the reaction between potassium and chlorine and you collect 31.2 g of potassium chloride. Determine your percent yield. (65.4% yield)

9. Your teacher gives you 5.00 g of a mixture of two salts—silver nitrate and potassium nitrate—and asks you to determine the percent silver nitrate by mass in the mixture. You dissolve the mixture in water and add an excess of aqueous sodium chloride. You collect and dry the white solid that precipitates and find that it has a mass of 1.48 g. Provide balanced equations for all reactions that occur in this process and determine the percent silver nitrate by mass in the original mixture.

10. The dissociation of ammonia into its elements is an endothermic reaction, absorbing 92.2 kJ of energy according to the equation:

    $2\ NH_3\,(g) + 92.2\ kJ \rightarrow 3\ H_2(g) + N_2(g)$

    How much energy will be required to decompose 85.0 grams of ammonia? (+231 kJ)

11. Phosphorus burns in air to produce dense white clouds of P_4O_{10} gas. When this gas is dissolved in rain water, phosphoric acid is produced. How much energy is released when 14.2 g of P_4O_{10} reacts? (-21.2 kJ)

    $P_4O_{10}(g) + 6\ H_2O(l) \rightarrow 4\ H_3PO_4(aq) + 424\ kJ$

12. For the following balanced chemical reaction:

    $2H_2S + 3O_2 \rightarrow 2SO_2 + 2H_2O$

    If the reaction of 53.2 grams of H_2S and 72.95 grams of O_2 produces 20.03 grams of H_2O, what is the % yield? (73.21% yield)

13. What mass of solid AgBr is produced when 100.0 mL of 0.150 M $AgNO_3$ is added to 20.0 mL of 1.00M NaBr? (2.82 g)

14. Hydrogen cyanide gas is prepared commercially by the reaction of methane, $CH_4(g)$, ammonia, $NH_3(g)$, and oxygen, $O_2(g)$; the other product is gaseous water.
    a) Write a chemical equation for the reaction.

    b) What volume of HCN can be can be obtained from 20.0 L methane, 20.0 L ammonia, and 20.0 L oxygen? (13.3L HCN)