WKS
Limiting & Excess Reactants

NAME Answer Key
Period _______ **Date** _________________

1) What is meant by the limiting reactant? Excess reactant? Why is it necessary to identify the limiting reactant when you want to know how much product will form in a chemical reaction?
The limiting reactant is completely consumed during a chemical reaction. The excess reactant has some remaining. The limiting reactant determines the amount of product that can theoretically be produced

2) The unbalanced equation representing the production of tetraphosphorus trisulfide (P_4S_3), a substance used in some match heads, is

$$8\ P_4(s) + 3\ S_8(s) \rightarrow 8\ P_4S_3(s).$$

a. When 6.00 mol P_4 and 6.00 mol S_8 react, which is the limiting reactant?
Calculate and compare mole equivalents:

$$\text{equiv. } P_4 = \frac{6.00 \text{ mol } P_4}{8 \text{ mol } P_4} = 0.75;\ \text{equiv. } S_8 = \frac{6.00 \text{ mol } S_8}{3 \text{ mol } S_8} = 2.00;\ 0.75 < 2.00, \text{ so } \boxed{P_4 \text{ is limiting.}}$$

b. How many moles of the excess reactant are used when the reaction is complete?

$$?\text{ moles } S_8 \text{ used} = 6.00\ \cancel{\text{mol } P_4} \times \frac{3\ \cancel{\text{mol } S_8}}{8\ \cancel{\text{mol } P_4}} = \boxed{2.25 \text{ mol } S_8 \text{ used}}$$

c. How many moles of the excess reactant remain after the reaction is complete?

$$\text{Excess } S_8 = (6.00 \text{ mol } S_8 \text{ present} - 2.25 \text{ mol } S_8 \text{ used}) = \boxed{3.75 \text{ mol } S_8 \text{ remaining}}$$

d. How many *grams* of P_4S_3 are produced?

$$?\text{ g } P_4S_3 = 6.00\ \cancel{\text{mol } P_4} \times \frac{8\ \cancel{\text{mol } P_4S_3}}{8\ \cancel{\text{mol } P_4}} \times \frac{220.1 \text{ g } P_4S_3}{1\ \cancel{\text{mol } P_4S_3}} = \boxed{1320 \text{ g } P_4S_3}$$

3) Trisilicon tetranitride, a ceramic that has orthopedic applications, is formed by reacting silicon with nitrogen gas at high temperature:

$$\underline{\ 3\ }\ Si(s) + \underline{\ 2\ }\ N_2(g) \rightarrow \underline{\quad}\ Si_3N_4(s)$$

a. Balance this synthesis reaction using the lowest whole-number coefficients.
b. What is the limiting reactant if 3.17 g of Si and 2.55 g of N_2 react?

$$3.17 \text{ g Si} \times \frac{1 \text{ mol Si}}{28.09 \text{ g Si}} = 0.113 \text{ mol Si};\ 2.55 \text{ g } N_2 \times \frac{1 \text{ mol } N_2}{28.02 \text{ mol } N_2} = 0.0910 \text{ mol } N_2$$

$$\text{Si: } \frac{0.113 \text{ mol Si}}{3 \text{ mol Si}} = 0.0377;\ N_2: \frac{0.0910 \text{ mol } N_2}{2 \text{ mol } N_2} = 0.0455;\ 0.0377 < 0.0455 \text{ so } \boxed{\text{Si is limiting.}}$$

c. What is the excess reactant, and how many grams remain after the reaction is completed?

$$?\text{ g } N_2 \text{ used} = 0.113\ \cancel{\text{mol Si}} \times \frac{2\ \cancel{\text{mol } N_2}}{3\ \cancel{\text{mol Si}}} \times \frac{28.02 \text{ g } N_2}{1\ \cancel{\text{mol } N_2}} = \boxed{2.11 \text{ g } N_2 \text{ used}}$$

d. What mass of excess reactant (in g), remains after the reaction is completed?

$$\text{Excess } N_2 = (2.55 \text{ g } N_2 \text{ present} - 2.11 \text{ g } N_2 \text{ used}) = \boxed{0.44 \text{ g } N_2 \text{ remaining}}$$

e. How many grams of Si_3N_4 can be produced from these reactants?

$$\text{mass } Si_3N_4 = 0.113\ \cancel{\text{mol Si}} \times \frac{1\ \cancel{\text{mol } Si_3N_4}}{3\ \cancel{\text{mol Si}}} \times \frac{140.31 \text{ g } Si_3N_4}{1\ \cancel{\text{mol } Si_3N_4}} = \boxed{5.28 \text{ g } Si_3N_4}$$

4) In blast furnaces, iron(III) oxide reacts with (is reduced by) gaseous carbon monoxide to form liquid iron and gaseous carbon dioxide

a. Write and balance the equation with the lowest whole-number coefficients.

$$Fe_2O_3 + 3\ CO \rightarrow 2\ Fe + 3\ CO_2$$

b. If 100. g of iron(III) oxide and 56.0 g of carbon monoxide are reacted, which reactant is limiting?

$$100.\text{ g } Fe_2O_3 \times \frac{1 \text{ mol } Fe_2O_3}{159.7 \text{ g } Fe_2O_3} = 0.626 \text{ mol } Fe_2O_3;\ 56.0 \text{ g CO} \times \frac{1 \text{ mol CO}}{28.01 \text{ mol CO}} = 2.00 \text{ mol CO}$$

$$Fe_2O_3: \frac{0.626 \text{ mol } Fe_2O_3}{1 \text{ mol } Fe_2O_3} = 0.626;\ \text{CO: } \frac{2.00 \text{ mol CO}}{3 \text{ mol CO}} = 0.667;\ 0.626 < 0.667, \text{ so } \boxed{Fe_2O_3 \text{ is limiting}}$$

c. How much excess reactant (in g) remain after the reaction is completed?

$$?\text{ g CO used} = 0.626 \cancel{\text{mol } Fe_2O_3} \times \frac{3 \cancel{\text{mol CO}}}{1 \cancel{\text{mol } Fe_2O_3}} \times \frac{28.01 \text{ g CO}}{1 \cancel{\text{mol CO}}} = \boxed{52.6 \text{ g CO used}}$$

d. How much excess reactant (in g) remains after the reaction is completed?

$$\text{Excess CO} = (56.0 \text{ g CO present} - 52.6 \text{ g CO used}) = \boxed{3.4 \text{ g CO remaining}}$$

e. How many grams of iron are produced in this reaction?

$$?\text{ g Fe} = 0.626 \cancel{\text{mol } Fe_2O_3} \times \frac{2 \cancel{\text{mol Fe}}}{\cancel{\text{mol } Fe_2O_3}} \times \frac{55.85 \text{ g Fe}}{1 \cancel{\text{mol Fe}}} = \boxed{69.9 \text{ g Fe}}$$

5) One industrial process for producing nitric acid begins by combusting ammonia gas in oxygen gas to form nitrogen monoxide gas and liquid water.

a. Write and balance the equation with the lowest whole-number coefficients.

$$4\ NH_3\ (g) + 5\ O_2\ (g) \rightarrow 4\ NO\ (g) + 6\ H_2O\ (l)$$

b. Which reactant is limiting if 4.20×10^4 g of NH_3 and 1.31×10^5 g of O_2 are available?

$$4.20\times10^4 \text{ g } NH_3 \times \frac{1 \text{ mol } NH_3}{17.04 \text{ g } NH_3} = 2460 \text{ mol } NH_3;\ 1.31\times10^5 \text{ g } O_2 \times \frac{1 \text{ mol } O_2}{32.00 \text{ mol } O_2} = 4090 \text{ mol } O_2$$

$$\text{equiv. } NH_3 = \frac{2460 \text{ mol } NH_3}{4 \text{ mol } NH_3} = 615;\ \text{equiv. } O_2 = \frac{4090 \text{ mol } O_2}{5 \text{ mol } O_2} = 818;\ 615 < 818, \text{ so } \boxed{NH_3 \text{ is limiting}}$$

c. Which reactant is in excess, and how many grams remain after the reaction is completed?

$$?\text{ g } O_2 \text{ used} = 2460 \cancel{\text{mol } NH_3} \times \frac{5 \cancel{\text{mol } O_2}}{4 \cancel{\text{mol } NH_3}} \times \frac{32.00 \text{ g } O_2}{1 \cancel{\text{mol } O_2}} = \boxed{9.86\times10^4 \text{ g } O_2 \text{ used}}$$

d. How many grams of the excess reactant remain after the reaction is completed?

$$\text{Excess } O_2 = (1.31\times10^5 \text{ g } O_2 \text{ present} - 9.86\times10^4 \text{ g } O_2 \text{ used}) = \boxed{3.2\times10^4 \text{ g } O_2 \text{ remaining}}$$

e. How many grams of NO will be produced in this reaction?

$$?\text{ g NO} = 2460 \cancel{\text{mol } NH_3} \times \frac{4 \cancel{\text{mol NO}}}{4 \cancel{\text{mol } NH_3}} \times \frac{30.01 \text{ g NO}}{1 \cancel{\text{mol NO}}} = \boxed{7.38\times10^4 \text{ g NO}}$$