Name _______________ **Chemistry** **___/___/___**

Limiting Reagent/Percent Yield Answers

1. $4Fe(s) + 3O_2(g) \rightarrow 2Fe_2O_3(s)$

Iron reacts with oxygen to produce iron(III) oxide, as represented by the equation above. A 75.0 g sample of Fe(*s*) is mixed with 11.5 L of $O_2(g)$ at STP.
(a) Identify the limiting reactant when the mixture is heated to produce $Fe_2O_3(s)$. Support your answer with calculations.
(b) Determine the amount of excess that remains.
(c) Calculate the mass of $Fe_2O_3(s)$ produced when the reaction proceeds to completion.
(d) If 42.3 grams of Fe_2O_3 are actually produced, what is the percent yield?

(a)
$$\frac{75.0\text{ g Fe}}{1} \times \frac{1\text{ mole Fe}}{55.8\text{ g Fe}} \times \frac{2\text{ moles }Fe_2O_3}{4\text{ mole Fe}} \times \frac{159.6\text{ g }Fe_2O_3}{1\text{ mol }Fe_2O_3} = \mathbf{107\text{ g }Fe_2O_3}$$

$$\frac{11.5\text{ L }O_2}{1} \times \frac{1\text{ mole }O_2}{22.4\text{ L }O_2} \times \frac{2\text{ mole }Fe_2O_3}{3\text{ moles }O_2} \times \frac{159.6\text{ g }Fe_2O_3}{1\text{ mol }Fe_2O_3} = \mathbf{54.6\text{ g }Fe_2O_3}$$

Limiting Reagent: $\mathbf{O_2}$ Excess Reagent: **Fe**

(b)
$$\frac{11.5\text{ L }O_2}{1} \times \frac{1\text{ mole }O_2}{22.4\text{ L }O_2} \times \frac{4\text{ mole Fe}}{3\text{ moles }O_2} \times \frac{55.8\text{ g Fe}}{1\text{ mol Fe}} = \mathbf{38.2\text{ g Fe needed}}$$

75.0 – 38.2 = **36.8 grams excess**

(c)
$$\frac{11.5\text{ L }O_2}{1} \times \frac{1\text{ mole }O_2}{22.4\text{ L }O_2} \times \frac{2\text{ mole }Fe_2O_3}{3\text{ moles }O_2} \times \frac{159.6\text{ g }Fe_2O_3}{1\text{ mol }Fe_2O_3} = \mathbf{54.6\text{ g }Fe_2O_3}$$

(d) Percent Yield $= \frac{42.3}{54.6} \times 100 = \mathbf{77.5\%}$

2. $N_2 + 3H_2 \rightarrow 2NH_3$

Nitrogen gas reacts with hydrogen gas to form ammonia, NH_3, gas. 50.0 liters of hydrogen gas and 18.0 liters of nitrogen gas react.
(a) Identify the limiting reactant when the mixture is heated to produce $NH_3(g)$. Support your answer with calculations.
(b) How many liters of excess reagent remain?
(c) Calculate the volume of NH_3 produced.
(d) Calculate the percent yield if 20.0 liters of NH_3 is actually produced?

(a)
$$\frac{50.0\text{ L }H_2}{1} \times \frac{1\text{ mole }H_2}{22.4\text{ L }H_2} \times \frac{2\text{ moles }NH_3}{3\text{ moles }H_2} \times \frac{22.4\text{ L }NH_3}{1\text{ mole }N_2} = \mathbf{33.3\text{ L }NH_3}$$

$$\frac{18.0\text{ L }N_2}{1} \times \frac{1\text{ mole }N_2}{22.4\text{ L }N_2} \times \frac{2\text{ mole }NH_3}{1\text{ mole }N_2} \times \frac{22.4\text{ L }NH_3}{1\text{ mole }N_2} = \mathbf{36.0\text{ L }NH_3}$$

Limiting Reagent: $\mathbf{H_2}$ Excess Reagent: $\mathbf{N_2}$

(b)
$$\frac{50.0\text{ L }H_2}{1} \times \frac{1\text{ mole }H_2}{22.4\text{ L }H_2} \times \frac{1\text{ mole }N_2}{3\text{ moles }H_2} \times \frac{22.4\text{ L }N_2}{1\text{ mole }N_2} = \mathbf{16.7\text{ L }N_2\text{ needed}}$$

18.0– 16.7 = **1.3 L of excess remain**

(c)
$$\frac{50.0\text{ L }H_2}{1} \times \frac{1\text{ mole }H_2}{22.4\text{ L }H_2} \times \frac{2\text{ moles }NH_3}{3\text{ moles }H_2} \times \frac{22.4\text{ L }NH_3}{1\text{ mole }N_2} = \mathbf{33.3\text{ L }NH_3}$$

(d) Percent Yield $= \frac{20.0}{33.3} \times 100 = \mathbf{60.1\%}$

3. $2Cu + S \rightarrow Cu_2S$

Solid copper reacts with solid sulfur to form solid cuprous sulfide. 80.0 grams of copper and 25.0 grams of sulfur are reacted.

(a) Identify the limiting reactant when the mixture is heated to produce Cu_2S. Support your answer with calculations.
(b) How many grams of excess reagent remain?
(c) Calculate the mass of Cu_2S produced.
(d) Calculate the percent yield if 87.3 grams of Cu_2S is actually produced?

(a)

$$\frac{80.0\text{ g Cu}}{1} \times \frac{1\text{ mole Cu}}{63.55\text{ g Cu}} \times \frac{1\text{ mole }Cu_2S}{2\text{ moles Cu}} \times \frac{159.16\text{ g }Cu_2S}{1\text{ mole }Cu_2S} = \mathbf{100.\text{ g }Cu_2S}$$

$$\frac{25.0\text{ g S}}{1} \times \frac{1\text{ mole S}}{32.06\text{ g S}} \times \frac{1\text{ mole }Cu_2S}{1\text{ mole S}} \times \frac{159.16\text{ g }Cu_2S}{1\text{ mole }Cu_2S} = \mathbf{124\text{ g }Cu_2S}$$

Limiting Reagent: **Cu** Excess Reagent: **S**

(b)

$$\frac{80.0\text{ g Cu}}{1} \times \frac{1\text{ mole Cu}}{63.55\text{ g Cu}} \times \frac{1\text{ mole S}}{2\text{ moles Cu}} \times \frac{32.06\text{ g S}}{1\text{ mole S}} = \mathbf{20.2\text{ g S needed}}$$

25.0 – 20.2 = **4.8 grams of excess**

(c)

$$\frac{80.0\text{ g Cu}}{1} \times \frac{1\text{ mole Cu}}{63.55\text{ g Cu}} \times \frac{1\text{ mole }Cu_2S}{2\text{ moles Cu}} \times \frac{159.16\text{ g }Cu_2S}{1\text{ mole }Cu_2S} = \mathbf{100.\text{ g }Cu_2S}$$

(d)

$$\text{Percent Yield} = \frac{87.3}{100.} \times 100 = \mathbf{87.3\%}$$

4. $2Al + 3CuSO_4 \rightarrow Al_2(SO_4)_3 + 3Cu$

Solid aluminum reacts with aqueous copper(II) sulfate in a single displacement reaction. 1.87 grams of aluminum reacts with 2.50 grams of $CuSO_4$.

(a) Identify the limiting reactant when the mixture is heated to produce Cu. Support your answer with calculations.
(b) How many grams of excess reagent remain?
(c) Calculate the mass of Cu produced.
(d) Calculate the percent yield if 0.65 grams of Cu actually produced?

(a)

$$\frac{1.87\text{ g Al}}{1} \times \frac{1\text{ mole Al}}{26.98\text{ g Al}} \times \frac{3\text{ moles Cu}}{2\text{ moles Al}} \times \frac{63.55\text{ g Cu}}{1\text{ mole Cu}} = \mathbf{6.61\text{ g Cu}}$$

$$\frac{2.50\text{ g }CuSO_4}{1} \times \frac{1\text{ mole }CuSO_4}{159.61\text{ g }CuSO_4} \times \frac{3\text{ moles Cu}}{3\text{ moles }CuSO_4} \times \frac{63.55\text{ g Cu}}{1\text{ mole Cu}} = \mathbf{0.995\text{ g Cu}}$$

Limiting Reagent: **$CuSO_4$** Excess Reagent: **Al**

(b)

$$\frac{2.50\text{ g }CuSO_4}{1} \times \frac{1\text{ mole }CuSO_4}{159.61\text{ g }CuSO_4} \times \frac{2\text{ moles Al}}{3\text{ moles }CuSO_4} \times \frac{26.98\text{ g Al}}{1\text{ mole Cu}} = \mathbf{0.282\text{ g Al}}$$

1.87 – 0.282 = **1.59 grams of excess**

(c)

$$\frac{2.50\text{ g }CuSO_4}{1} \times \frac{1\text{ mole }CuSO_4}{159.61\text{ g }CuSO_4} \times \frac{3\text{ moles Cu}}{3\text{ moles }CuSO_4} \times \frac{63.55\text{ g Cu}}{1\text{ mole Cu}} = \mathbf{0.995\text{ g Cu}}$$

(d)

$$\text{Percent Yield} = \frac{0.65}{0.995} \times 100 = \mathbf{65\%}$$

5. $2CH_3OH + O_2 \rightarrow 2H_2CO + 2H_2O$

Methanol, CH_3OH, reacts with oxygen to produce formaldehyde, H_2CO and water according to the following equation: The above reaction is carried out in a laboratory and 270.0 grams of methanol and 48.6 liters of oxygen at STP.

(a) Identify the limiting reactant. Support your answer with calculations.

(b) How much of the excess reagent remains?

(c) Calculate the mass of H_2CO produced.

(d) If 82.34 grams of formaldehyde are actually produced, what is the percent yield?

(a)

$$\frac{270.0 \text{ g } CH_3OH}{1} \times \frac{1 \text{ mole } CH_3OH}{32.05 \text{ g } CH_3OH} \times \frac{2 \text{ moles } H_2CO}{2 \text{ moles } CH_3OH} \times \frac{30.03 \text{ g } H_2CO}{1 \text{ mole } H_2CO} = \mathbf{253.0 \text{ g } H_2CO}$$

$$\frac{48.6 \text{ L } O_2}{1} \times \frac{1 \text{ mole } O_2}{22.4 \text{ L } O_2} \times \frac{2 \text{ moles } CH_3OH}{1 \text{ mole } O_2} \times \frac{30.03 \text{ g } H_2CO}{1 \text{ mole } H_2CO} = \mathbf{130. \text{ g } H_2CO}$$

Limiting Reagent: **O_2** Excess Reagent: **CH_3OH**

(b)

$$\frac{48.6 \text{ L } O_2}{1} \times \frac{1 \text{ mole } O_2}{22.4 \text{ L } O_2} \times \frac{2 \text{ moles } CH_3OH}{1 \text{ mole } O_2} \times \frac{32.05 \text{ g } CH_3OH}{1 \text{ mole } CH_3OH} = \mathbf{139 \text{ g } CH_3OH \text{ needed}}$$

270.0 – 139 = **131 grams of excess CH_3OH**

(c)

$$\frac{48.6 \text{ L } O_2}{1} \times \frac{1 \text{ mole } O_2}{22.4 \text{ L } O_2} \times \frac{2 \text{ moles } CH_3OH}{1 \text{ mole } O_2} \times \frac{30.03 \text{ g } H_2CO}{1 \text{ mole } H_2CO} = \mathbf{130. \text{ g } H_2CO}$$

(d)

$$\text{Percent Yield} = \frac{82.34}{130.} \times 100 = \mathbf{63.3\%}$$

6. $C_6H_6 + HNO_3 \rightarrow C_6H_5NO_2 + H_2O$

Nitrobenzene is used in the production of perfumes. Nitrobenzene, $C_6H_5NO_2$, along with water, is produced by reacting benzene, C_6H_6, with nitric acid, HNO_3. 109.5 grams of C_6H_6 and 102.2 grams of HNO_3 react.

(a) Identify the limiting reactant. Support your answer with calculations.

(b) How much of the excess reagent remains?

(c) Calculate the mass of Nitrobenzene, $C_6H_5NO_2$ produced.

(d) If 159.7 grams of $C_6H_5NO_2$ are actually produced, what is the percent yield?

(a)

$$\frac{109.5 \text{ g } C_6H_6}{1} \times \frac{1 \text{ mole } C_6H_6}{78.12 \text{ g } C_6H_6} \times \frac{1 \text{ mole } C_6H_5NO_2}{1 \text{ mole } C_6H_6} \times \frac{123.12 \text{ g } C_6H_5NO_2}{1 \text{ mole } C_6H_5NO_2} = \mathbf{172.6 \text{ g } C_6H_5NO_2}$$

$$\frac{102.2 \text{ g } HNO_3}{1} \times \frac{1 \text{ mole } HNO_3}{63.02 \text{ g } HNO_3} \times \frac{1 \text{ mole } C_6H_5NO_2}{1 \text{ mole } HNO_3} \times \frac{123.12 \text{ g } C_6H_5NO_2}{1 \text{ mole } C_6H_5NO_2} = \mathbf{199.7 \text{ g } C_6H_5NO_2}$$

Limiting Reagent: **$\underline{C_6H_6}$** Excess Reagent: **HNO_3**

(b)

$$\frac{109.5 \text{ g } C_6H_6}{1} \times \frac{1 \text{ mole } C_6H_6}{78.12 \text{ g } C_6H_6} \times \frac{1 \text{ mole } HNO_3}{1 \text{ mole } C_6H_6} \times \frac{63.02 \text{ g } HNO_3}{1 \text{ mole } HNO_3} = \mathbf{88.33 \text{ g } HNO_3 \text{ needed}}$$

102.2 – 88.33 = **13.9 grams of excess HNO_3**

(c)

$$\frac{109.5 \text{ g } C_6H_6}{1} \times \frac{1 \text{ mole } C_6H_6}{78.12 \text{ g } C_6H_6} \times \frac{1 \text{ mole } C_6H_5NO_2}{1 \text{ mole } C_6H_6} \times \frac{123.12 \text{ g } C_6H_5NO_2}{1 \text{ mole } C_6H_5NO_2} = \mathbf{172.6 \text{ g } C_6H_5NO_2}$$

(d)

$$\text{Percent Yield} = \frac{159.7}{172.6} \times 100 = \mathbf{92.54\%}$$

7. $2Al + 3CrO \rightarrow Al_2O_3 + 3Cr$

Aluminum oxide is produced in a single replacement reaction between aluminum and chromium(II) oxide. 225.0 grams of chromium(II) oxide reacted with 125.0 grams of aluminum.

(a) Identify the limiting reactant. Support your answer with calculations.

(b) How much of the excess reagent remains?

(c) Calculate the mass of aluminum oxide produced.

(d) If 100.0 grams of aluminum oxide are actually produced, what is the percent yield?

(a) $\frac{225.0 \text{ g CrO}}{1} \times \frac{1 \text{ mole CrO}}{68.00 \text{ g CrO}} \times \frac{1 \text{ mole } Al_2O_3}{3 \text{ moles CrO}} \times \frac{101.96 \text{ g } Al_2O_3}{1 \text{ mole } Al_2O_3} = \mathbf{112.5\ g\ Al_2O_3}$

$\frac{125.0 \text{ g Al}}{1} \times \frac{1 \text{ mole Al}}{26.98 \text{g Al}} \times \frac{1 \text{ mole } Al_2O_3}{2 \text{ moles Al}} \times \frac{101.96 \text{g } Al_2O_3}{1 \text{ mole } Al_2O_3} = \mathbf{236.2\ g\ Al_2O_3}$

Limiting Reagent: **CrO** Excess Reagent: **Al**

(b) $\frac{225.0 \text{ g CrO}}{1} \times \frac{1 \text{ mole CrO}}{68.00 \text{ g CrO}} \times \frac{2 \text{ mole Al}}{3 \text{ moles CrO}} \times \frac{26.98 \text{ g Al}}{1 \text{ mole Al}} = \mathbf{59.51\ g\ Al}$

125.0 – 59.51 = **65.5 grams of excess Al**

(c) $\frac{225.0 \text{ g CrO}}{1} \times \frac{1 \text{ mole CrO}}{68.00 \text{ g CrO}} \times \frac{1 \text{ mole } Al_2O_3}{3 \text{ moles CrO}} \times \frac{101.96 \text{ g } Al_2O_3}{1 \text{ mole } Al_2O_3} = \mathbf{112.5\ g\ Al_2O_3}$

(d) Percent Yield = $\frac{100.0}{112.5} \times 100 = \mathbf{88.89\%}$