U5LM3B-WS Limiting Reagent & Percent Yield **Name: _____KEY______**

I. Methanol, CH_3OH, is used as a fuel. Recall that combustion reactions produce carbon dioxide and water.

1. Write a balanced chemical equation for the combustion reaction.

   $2\ CH_3OH + 3\ O_2 \rightarrow 2\ CO_2 + 4\ H_2O$

2. How many moles of carbon dioxide are produced in each of the following cases? Include calculations quantifying the number of moles left over for any excess reactants.

   a. 2 mols of CH_3OH react with 3 mols of O_2.

   $2\ \cancel{mol\ CH_3OH} \times \frac{2\ mol\ CO_2}{2\ \cancel{mol\ CH_3OH}} =$ **2 mol CO_2** max

   $3\ \cancel{mol\ O_2} \times \frac{2\ mol\ CO_2}{3\ \cancel{mol\ O_2}} = 2\ mol\ CO_2$ max

   These reagents are in stoichiometric equivalence, so both reagents react completely. **0 moles of excess reagent**

   b. 2 mols of CH_3OH react with 2 mols of O_2

   $2\ \cancel{mol\ CH_3OH} \times \frac{2\ mol\ CO_2}{2\ \cancel{mol\ CH_3OH}} = 2\ mol\ CO_2$ max — CH_3OH excess reagent

   $2\ \cancel{mol\ O_2} \times \frac{2\ mol\ CO_2}{3\ \cancel{mol\ O_2}} =$ **1.33 mol CO_2** max — O_2 llimiting reagent

   $2\ \cancel{mol\ O_2} \times \frac{2\ mol\ CH_3OH}{3\ \cancel{mol\ O_2}} = 1.33\ mol\ CH_3OH$ used — 2-1.33 = **0.67 moles excess methanol**

   c. 3 mols of CH_3OH react with 3 mols of O_2

   Perform hypothetical maximum calculations as above. Another way of determining the limiting reagent is to say…" 3 moles of methanol would require 4.5 moles of oxygen. We only have 3 moles of O_2, so oxygen must be limiting."

   $3\ \cancel{mol\ CH_3OH} \times \frac{3\ mol\ O_2}{2\ \cancel{mol\ CH_3OH}} = 4.5\ mol\ O_2$ required … O_2 is limiting

   All of the oxygen will react…

   $3\ \cancel{mol\ O_2} \times \frac{2\ mol\ CO_2}{3\ \cancel{mol\ O_2}} =$ **2 mol CO_2 produced**

   $3\ \cancel{mol\ O_2} \times \frac{2\ mol\ CH_3OH}{3\ \cancel{mol\ O_2}} = 2\ mol\ CH_3OH$ used; **1 mole of excess methanol**

   d. 88 g of CH_3OH react with 88 g of O_2

   These problems require you to convert grams using the molar mass before performing steps above. Note: methanol and oxygen have different molecular formulas but happen to have similar molar masses.

   $88\ \cancel{g\ CH_3OH} \times \frac{1\ \cancel{mol\ CH_3OH}}{32\ \cancel{g\ CH_3OH}} \times \frac{2\ mol\ CO_2}{2\ \cancel{mol\ CH_3OH}} = 2.75\ mol\ CO_2$ max — CH_3OH excess reagent

   $88\ \cancel{g\ O_2} \times \frac{1\ \cancel{mol\ O_2}}{32\ \cancel{g\ O_2}} \times \frac{2\ mol\ CO_2}{3\ \cancel{mol\ O_2}} =$ **1.83 mol CO_2** max — O_2 limiting reagent

$1.83\ \cancel{mol\ CO_2} \times \frac{2\ mol\ CH_3OH}{2\ \cancel{mol\ CO_2}} = 1.83\ mol\ CH_3OH$ used

$88\ \cancel{g\ CH_3OH} \times \frac{1\ mol\ CH_3OH}{32\ \cancel{g\ CH_3OH}} =$ 2.75 mol CH_3OH available – 1.83 mol used = **0.92 mol excess CH_3OH**

e. 15 g of CH_3OH react with 12 g of O_2

$15g\ CH_3OH \times \frac{1\ mol}{32g}$ = *0.47 mol CH_3OH available* … x(3mol O_2 / 2mol CH_3OH) = 0.71 mol O_2 required

$12g\ O_2 \times \frac{1\ mol}{32g}$ = only *0.375 mol O_2 available*…. So O_2 is limiting.

All of the oxygen will react…

$0.375\ \cancel{mol\ O_2} \times \frac{2\ mol\ CO_2}{3\ \cancel{mol\ O_2}}$ = **0.25 mol CO_2 produced**

$0.375\ \cancel{mol\ O_2} \times \frac{2\ mol\ CH_3OH}{3\ \cancel{mol\ O_2}}$ = 0.25 mol CH_3OH used; (0.47–0.25) = **0.22 mole excess methanol**

f. 25 g of CH_3OH react with 35 g of O_2

$25g\ CH_3OH \times \frac{1\ mol}{32g}$ = *0.78 mol CH_3OH available* … x(3mol O_2 / 2mol CH_3OH) = 1.17 mol O_2 required

$35g\ O_2 \times \frac{1\ mol}{32g}$ = only *1.09 mol O_2 available*…. So O_2 is limiting.

All of the oxygen will react…

$1.09\ \cancel{mol\ O_2} \times \frac{2\ mol\ CO_2}{3\ \cancel{mol\ O_2}}$ = **0.73 mol CO_2 produced**

$1.09\ \cancel{mol\ O_2} \times \frac{2\ mol\ CH_3OH}{3\ \cancel{mol\ O_2}}$ = 0.73 mol CH_3OH used; (0.78–0.73) = **0.05 mole excess methanol**

II. Propane is by-product of natural gas processing and petroleum refining. It is commonly used as a fuel for engines, oxy-gas torches, barbecues, portable stoves, and residential central heating.

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

Note: C_3H_8 and CO_2 have different molecular formulas but happen to have similar molar masses.

a. What mass of CO_2 is produced when 6.5 g of propane is reacted with 14.2 g of O_2?

$6.5g\ \mathbf{C_3H_8} \times \frac{1\ mol}{44g}$ = *0.15 mol C_3H_8 available* … x(5mol O_2 / 1mol C_3H_8) = 0.74 mol O_2 required

$14.2g\ O_2 \times \frac{1\ mol}{32g}$ = only *0.44 mol O_2 available*…. So O_2 is limiting.

All of the oxygen will react…

$0.44\ \cancel{mol\ O_2} \times \frac{3\ mol\ CO_2}{5\ \cancel{mol\ O_2}} \times \frac{44\ g\ CO_2}{1\ mol\ CO_2}$ = **11.7 g CO_2 produced (theoretical yield)**

b. The actual yield of the reaction described above is 8.0 g of carbon dioxide. What is the percent yield?

Percent yield is how much you got out of how much you could *theoretically* get…

$\frac{8.0\ g\ CO_2}{11.7\ g\ CO_2} \times 100\%$ = **68 % yield**

III. Nitrogen dioxide reacts with hydrogen to produce nitrogen and water. When 125g of nitrogen dioxide are allowed to react with excess hydrogen, the percent yield is 35%. How many grams of each product are actually formed during this process?

$$2\,NO_2 + 4\,H_2 \rightarrow N_2 + 4\,H_2O$$
125g x.s. ?g ?g given 35% yield

Convert limiting NO_2 to moles…

$$125g\ NO_2 \times \frac{1\ mol}{46g} = \textit{2.72 mol}\ NO_2\ \textit{reacts}$$

Calculate theoretical yield of nitrogen, in grams, and take 35% of that value.

$$2.72\ \cancel{mol\ NO_2} \times \frac{1\ \cancel{mol\ N_2}}{2\ \cancel{mol\ NO_2}} \times \frac{28\ g\ N_2}{1\ \cancel{mol\ N_2}} \times 0.35 = \textbf{13.3 g } \mathbf{N_2} \textbf{ produced}$$

Same for water

$$2.72\ \cancel{mol\ NO_2} \times \frac{4\ \cancel{mol\ H_2O}}{2\ \cancel{mol\ NO_2}} \times \frac{18\ g\ H_2O}{1\ \cancel{mol\ H_2O}} \times 0.35 = \textbf{34.3 g } \mathbf{H_2O} \textbf{ produced}$$