LIMITING REACTANT CALCULATIONS

- To find the limiting reactant, calculate how much product would be produced from ALL given reactants. Whichever produces the SMALLEST amouht of product is the limiting reactant, and the smallest anount of product is the actual amount of product produced.

Example:

$$\underset{56.08}{CaO(s)} + 3\underset{12.01}{C(s)} \xrightarrow{\Delta} \underset{64.10}{CaC_2(s)} + CO(g)$$

64.10 <- Formula weights

If you start with 100. g of each reactant, how much calcium carbide would be produced?

56.08 g CaO = mol CaO   mol CaO = mol CaC_2 64.10 g CaC_2 = mol CaC_2

$$100\,g\,CaO \times \frac{mol\,CaO}{56.08\,g\,CaO} \times \frac{mol\,CaC_2}{mol\,CaO} \times \frac{64.10\,g\,CaC_2}{mol\,CaC_2} = \boxed{114\,g\,CaC_2}$$

12.01 g C = mol C   3 mol C = mol CaC_2 64.10 g CaC_2 = mol CaC_2

$$100\,g\,C \times \frac{mol\,C}{12.01\,g\,C} \times \frac{mol\,CaC_2}{3\,mol\,C} \times \frac{64.10\,g\,CaC_2}{mol\,CaC_2} = 178\,g\,CaC_2$$

114 g of calcium carbide should be produced. Calcium oxide (the limiting reactant) runs out at that point, and no further reaction is possible.

We call calcium oxide "limiting", and we say carbon is present "in excess".

PERCENT YIELD

- Chemical reactions do not always go to completion! Things may happen that prevent the conversion of reactants to the desired/expected product!

1 SIDE REACTIONS:

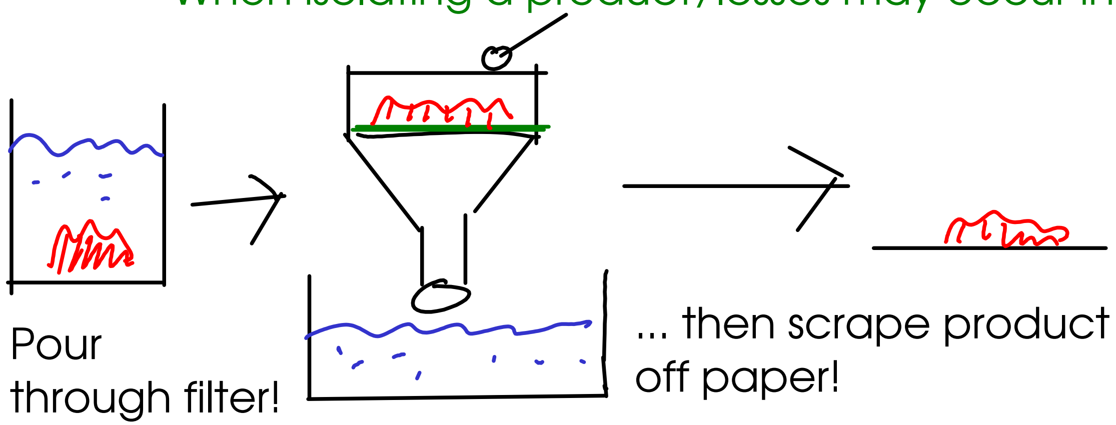
$C + O_2 \longrightarrow CO_2$ | This reaction occurs when there is a large amount of oxygen available

$2C + O_2 \longrightarrow 2CO$ | ... while this reaction is more favorable in low-oxygen environments!

... so in a low-oxygen environment, you may produce less carbon dioxide than expected!

2 TRANSFER AND OTHER LOSSES

- When isolating a product, losses may occur in the process. Example: filtering



During each step of this process, some amount of product loss will occur!

(3) EQUILIBRIUM

- Reactions may reach an equilbrium between prodcuts and reactants. We'll talk more about this in CHM 111. The net results is that the reaction will appear to stop before all reactants have been consumed!

- All of these factors cause a chemical reaction to produce LESS product than calculated. For many reactions, this difference isn't significant. But for others, we need to report the PERCENT YIELD.

Determined EXPERIMENTALLY!

$$\text{PERCENT YIELD} = \frac{\text{ACTUAL YIELD}}{\text{THEORETICAL YIELD}} \times 100\%$$

Calculated based on the limiting reactant. (The chemical calculations you've done up to now have been theoretical yields!)

... the percent yield of a reaction can never be greater than 100% due to conservation of mass! If you determine that a percent yield is greater than 100%, then you've made a mistake somewhere - either in a calculation or in the experiment itself!

$78.114\ g/mol$ $\qquad$ $123.111\ g/mol$ <- Formula weights

22.4 g $\qquad$ 31.6 g ACTUAL

$$C_6H_6 + HNO_3 \longrightarrow C_6H_5NO_2 + H_2O$$

benzene $\qquad$ nitric acid $\qquad$ nitrobenzene

22.4 grams of benzene are reacted with excess nitric acid. If 31.6 grams of nitrobenzene are collected from the reaction, what is the percent yield?

To find the PERCENT YIELD, we need to calculate the THEORETICAL YIELD of nitrobenzene. Start with the 22.4 g of benzene.

$78.114\ g\ C_6H_6 = mol\ C_6H_6 \qquad mol\ C_6H_6 = mol\ C_6H_5NO_2 \qquad 123.111\ g\ C_6H_5NO_2 = mol\ C_6H_5NO_2$

$$22.4\ g\ C_6H_6 \times \frac{mol\ C_6H_6}{78.114\ g\ C_6H_6} \times \frac{mol\ C_6H_5NO_2}{mol\ C_6H_6} \times \frac{123.111\ g\ C_6H_5NO_2}{mol\ C_6H_5NO_2} =$$

$$= 35.3\ g\ C_6H_5NO_2 \text{ = THEORETICAL YIELD}$$

$$\%\,yield = \frac{\text{actual yield}}{\text{theoretical yield}} \times 100\% = \frac{31.6\ g}{35.3\ g} \times 100\% = \boxed{89.5\%}$$

25.0 mL of acetic acid solution requires 37.3 mL of 0.150 M sodium hydroxide for complete reaction. The equation for this reaction is:

$$NaOH + HC_2H_3O_2 \rightarrow NaC_2H_3O_2 + H_2O$$

What is the molar concentration of the acetic acid?

$$\frac{\text{mol } HC_2H_3O_2}{\text{L solution}} \leftarrow = 25.0\text{ mL or } 0.0250\text{ L}$$

Start with the 37.3 mL of NaOH. Convert to moles

$0.150 \text{ mol NaOH} = \text{L}$ $\quad$ $\text{mol NaOH} = \text{mol } HC_2H_3O_2$ $\quad$ $\text{mL} = 10^{-3}\text{ L}$

$$37.3\text{ mL} \times \frac{10^{-3}\text{ L}}{\text{mL}} \times \frac{0.150\text{ mol NaOH}}{\text{L}} \times \frac{\text{mol } HC_2H_3O_2}{\text{mol NaOH}} = 0.005595 \text{ mol } HC_2H_3O_2$$

To get molarity, divide the moles acetic acid by the volume (in L) of acetic acid.

$$\frac{\text{mol } HC_2H_3O_2}{\text{L solution}} = \frac{0.005595\text{ mol } HC_2H_3O_2}{0.0250\text{ L}} = \boxed{0.224\text{ M } HC_2H_3O_2}$$

Shortcut: Use MILLIMOLES and MILLILITERS to save some unit conversion steps

$$37.3\text{ mL} \times \frac{0.150\text{ mol NaOH}}{\text{L}} \times \frac{\text{mol } HC_2H_3O_2}{\text{mol NaOH}} = 5.595\text{ mmol } HC_2H_3O_2$$

$$M = \frac{\text{mol } HC_2H_3O_2}{\text{L solution}} = \frac{5.595\text{ mmol } HC_2H_3O_2}{25.0\text{ mL}} = \boxed{0.224\text{ M } HC_2H_3O_2}$$