CONCEPT OF LIMITING REACTANT
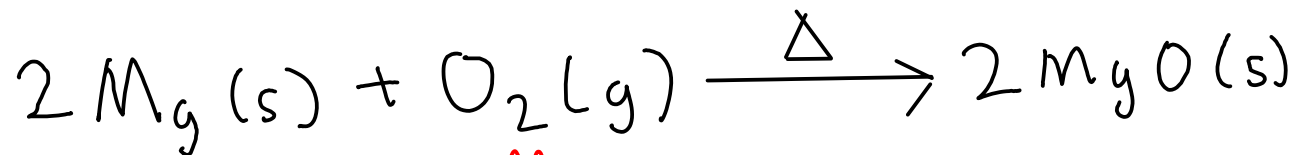
- When does a chemical reaction STOP?

$$2 Mg(s) + O_2(g) \xrightarrow{\Delta} 2 MgO(s)$$

Magnesium
strip

+

Flame,
oxygen from
air

Magnesium oxide
powder

- When does this reaction stop? When burned in open air, this reaction stops when
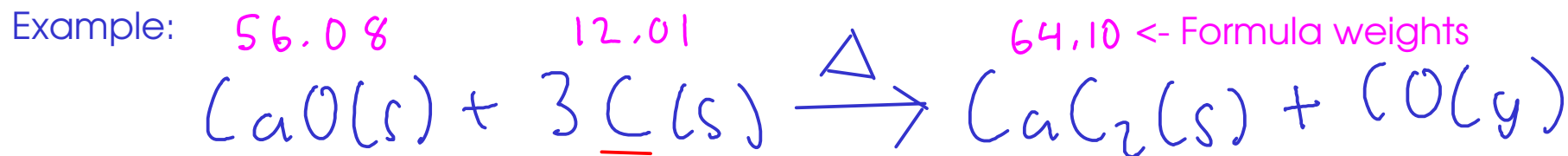all the MAGNESIUM STRIP is gone. We say that the magnesium is LIMITING.

- This reaction is controlled by the amount of available magnesium

- At the end of a chemical reaction, the LIMITING REACTANT will be completely consumed,
but there may be amount of OTHER reactants remaining. We do chemical calculations in
part to minimize these "leftovers".

These are often called "excess" reactants, or reactants present
"in excess"

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:

$$56.08 \qquad 12.01 \qquad\qquad\qquad 64.10 \text{ <- Formula weights}$$

$$CaO(s) + 3C(s) \xrightarrow{\Delta} CaC_2(s) + CO(y)$$

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

$CaO: 56.08g\ C_aO = mol\ C_aO \mid mol\ CaO = mol\ CaC_2 \mid 64.10g\ CaC_2 = mol\ CaC_2$

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

$C: 12.01gC = mol\ C \mid 3mol\ C = mol\ CaC_2 \mid 64.10g\ CaC_2 = mol\ CaC_2$

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

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114 grams of calcium carbide are produced. Calcium oxide runs out when 114g of carbide is formed, so the reaction stops at that point. We xsay that calcium oxide is LIMITING, and carbon is present IN EXCESS (there will be leftover carbon at the end of the reaction).

PERCENT YIELD

- Chemical reactions do not always go to completion! Things may happen that prevent the conversion of reactants to the desired/expected product!
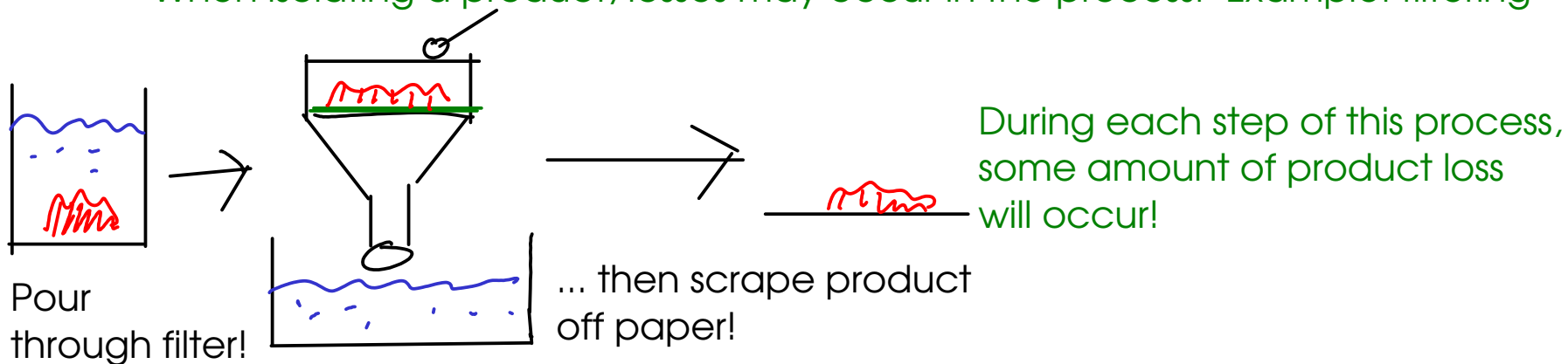
① SIDE REACTIONS:

$$C + O_2 \longrightarrow CO_2$$ 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

Pour through filter!

... then scrape product off paper!

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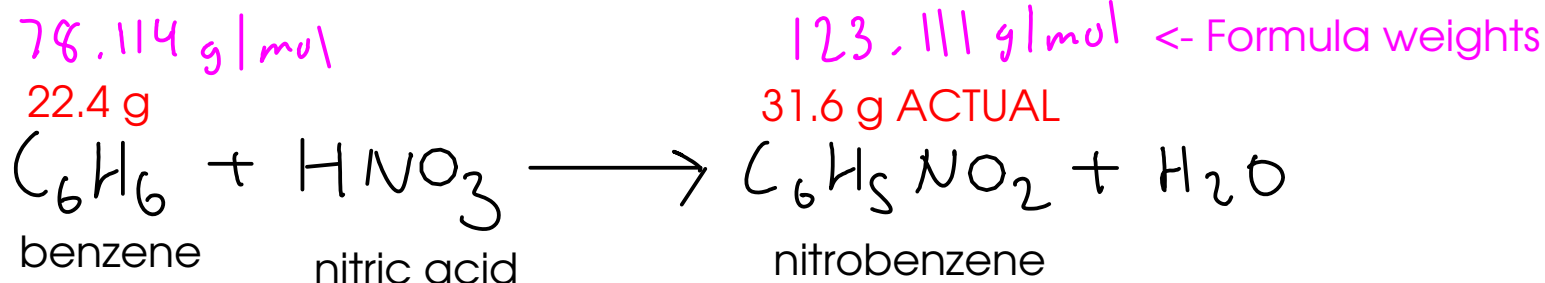
- All of these factors cause a chemical reaction to produce LESS product than calculated. For many reactions, this difference isn't significant. But for others, we need to report the PERCENT YIELD.

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 123.111 g/mol <- Formula weights

22.4 g 31.6 g ACTUAL

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

benzene   nitric acid   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 determine the percent yield, we first need to determine the THEORETICAL YIELD of nitrobenzene. This is the amount of nitrobenzene that could be produced from 22.4 grams of benzene under ideal conditions.

$$78.114 \text{ g } C_6H_6 = \text{mol } C_6H_6 \quad | \quad \text{mol } C_6H_6 = \text{mol } C_6H_5NO_2$$

$$123.111 \text{ g } C_6H_5NO_2 = \text{mol } C_6H_5NO_2$$

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

$$= 35.3 \text{ g } C_6H_5NO_2 \text{ .. theoretical yield}$$

PERCENT YIELD = $\dfrac{\text{ACTUAL YIELD}}{\text{THEORETICAL YIELD}}$ X 100% $= \dfrac{31.6 \text{ g}}{35.3 \text{ g}} \times 100\% = \boxed{89.5\%}$

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:

$$NaOH + HC_2H_3O_2 \longrightarrow NaC_2H_3O_2 + H_2O$$
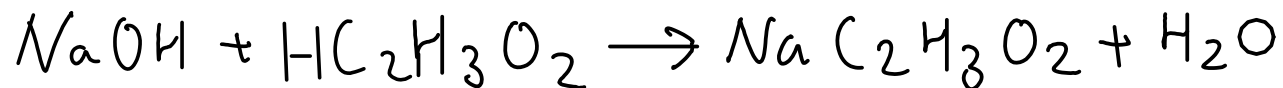
What is the <u>molar concentration</u> of the acetic acid?

$$\frac{\lfloor \text{ mol } HC_2H_3O_2}{\text{L Solution}} \longleftarrow = 25.0\,mL \text{ or } 0.025OL$$

Since we already know the colume of acetic acid, we just have to find the moles of acetic acid that were reacted - then we can find concentration. To do that, we start with the NaOH - since we know both how much of it we used AND its concentration.

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

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

$$M = \frac{\text{mol } HC_2H_3O_2}{\text{L Solution}} = \frac{0.005595 \text{ mol } HC_2H_3O_2}{0.025OL} = \boxed{0.224 \text{ M } HC_2H_3O_2}$$

$$42.081 \ g/mol$$
$$53.064 \ g/mol$$

$$4 C_3H_6 + 6 NO \longrightarrow 4 C_3H_3N + 6 H_2O + N_2$$

propylene $\qquad$ acrylonitrile
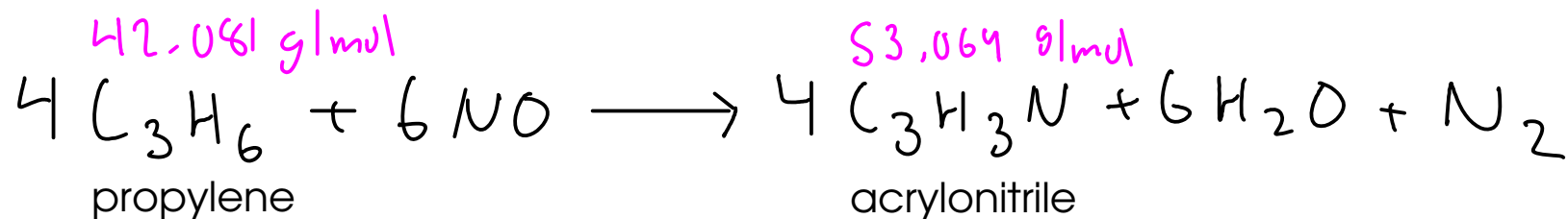
Calculate how many grams of acrylonitrile could be obtained from 651 kg of propylene, assuming there is excess NO present.

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1 - Convert mass propylene to moles using formula weight (and kg to g conversion)
2 - Convert moles propylene to moles acrylonitrile using chemical equation.
3 - Convert moles acrylonitrile to mass using formula weight.

$$42.081 \ g \ C_3H_6 = mol \ C_3H_6 \quad | \quad 4 \ mol \ C_3H_6 = 4 \ mol \ C_3H_3N$$

$$53.064 \ g \ C_3H_3N = mol \ C_3H_3N \quad | \quad kg = 10^3 \ g$$

$$651 \ kg \ C_3H_6 \times \frac{10^3 \ g}{kg} \times \frac{mol \ C_3H_6}{42.081 \ g \ C_3H_6} \times \frac{4 \ mol \ C_3H_3N}{4 \ mol \ C_3H_6} \times \frac{53.064 \ g \ C_3H_3N}{mol \ C_3H_3N} =$$

$$\underbrace{\qquad\qquad}_{①} \qquad \underbrace{\qquad}_{②} \qquad \underbrace{\qquad}_{③}$$

$$= \boxed{821000 \ g \ C_3H_3N} \quad (821 \ kg)$$