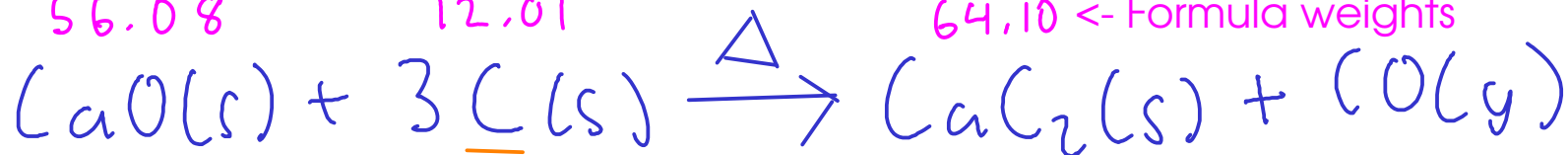


LIMITING REACTANT CALCULATIONS

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

Example: 56.08 12.01 64.10 <- Formula weights



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

$$\text{CaO: } 56.08 \text{ g CaO} = \text{mol CaO} \quad | \quad \text{mol CaO} = \text{mol CaC}_2 \quad | \quad 64.10 \text{ g CaC}_2 = \text{mol CaC}_2$$

$$100. \text{ g CaO} \times \frac{\text{mol CaO}}{56.08 \text{ g CaO}} \times \frac{\text{mol CaC}_2}{\text{mol CaO}} \times \frac{64.10 \text{ g CaC}_2}{\text{mol CaC}_2} = 114 \text{ g CaC}_2$$

$$\text{C: } 12.01 \text{ g C} = \text{mol C} \quad | \quad 3 \text{ mol C} = \text{mol CaC}_2 \quad | \quad 64.10 \text{ g CaC}_2 = \text{mol CaC}_2$$

$$100. \text{ g C} \times \frac{\text{mol C}}{12.01 \text{ g C}} \times \frac{\text{mol CaC}_2}{3 \text{ mol C}} \times \frac{64.10 \text{ g CaC}_2}{\text{mol CaC}_2} = 178 \text{ g CaC}_2$$

114 g of calcium carbide should be produced. Calcium oxide runs out when this amount of the carbide is made, and there's nothing left for the remaining carbon to react with. No further product can be made.

We say that calcium oxide is "limiting". We say that 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!

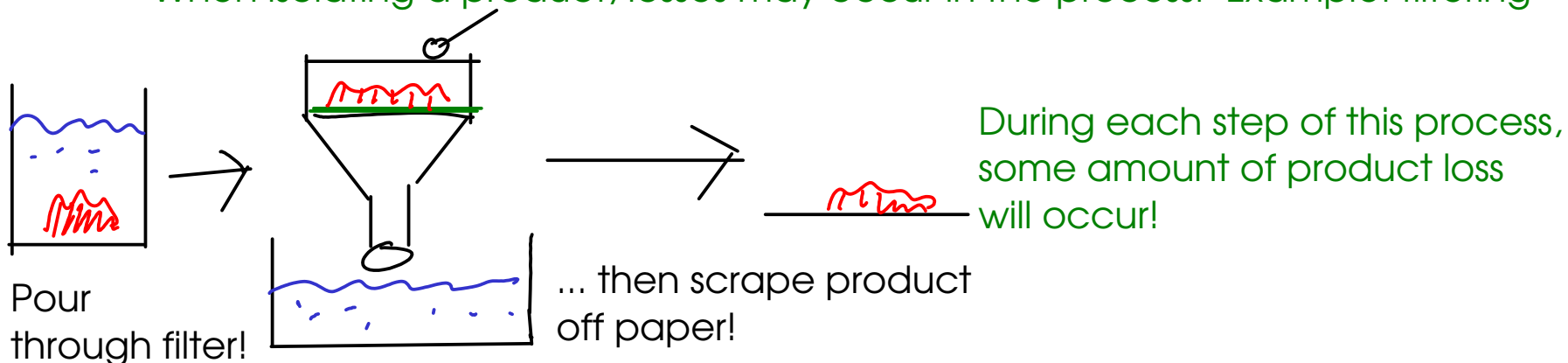
① SIDE REACTIONS:



... 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



③ EQUILIBRIUM

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

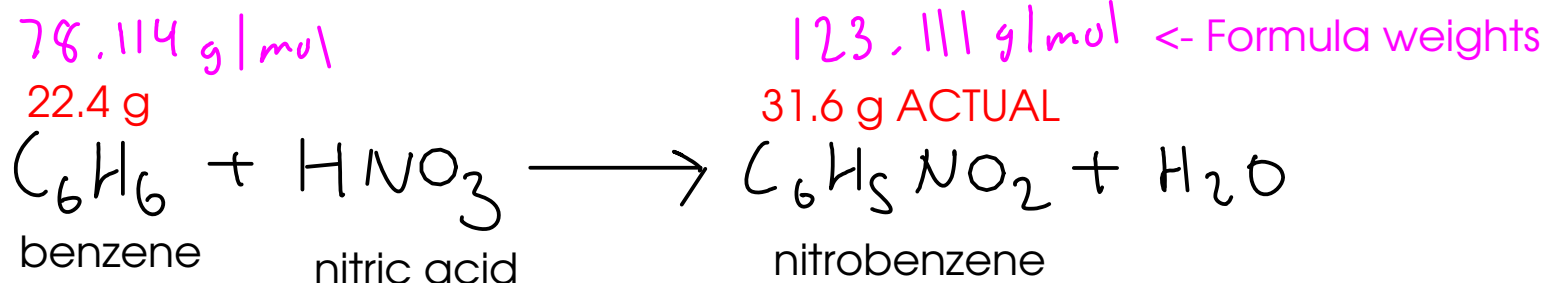
- All of these factors cause a chemical reaction to produce LESS product than calculated. For many reactions, this difference isn't significant. But for others, we need to report the PERCENT YIELD.

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

A blue arrow points from the text "Determined EXPERIMENTALLY!" to the "ACTUAL YIELD" term in the numerator of the equation.

↑ 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!



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 get the percent yield, we need to know the THEORETICAL yield ... how much nitrobenzene we could produce if all 22.4 grams of benzene were converted to product.

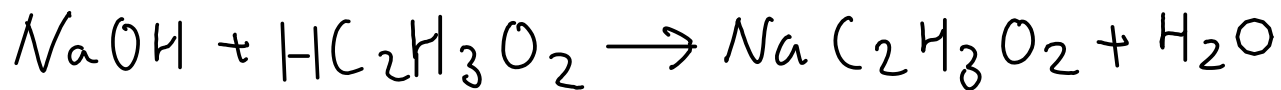
$$78.114 \text{ g C}_6\text{H}_6 = \text{mol C}_6\text{H}_6 \quad | \quad \text{mol C}_6\text{H}_6 = \text{mol C}_6\text{H}_5\text{NO}_2$$

$$123.111 \text{ g C}_6\text{H}_5\text{NO}_2 = \text{mol C}_6\text{H}_5\text{NO}_2$$

$$\begin{aligned}
 & 22.4 \text{ g C}_6\text{H}_6 \times \frac{\text{mol C}_6\text{H}_6}{78.114 \text{ g C}_6\text{H}_6} \times \frac{\text{mol C}_6\text{H}_5\text{NO}_2}{\text{mol C}_6\text{H}_6} \times \frac{123.111 \text{ g C}_6\text{H}_5\text{NO}_2}{\text{mol C}_6\text{H}_5\text{NO}_2} \\
 & = 35.3 \text{ g C}_6\text{H}_5\text{NO}_2 \quad \text{THEORETICAL YIELD}
 \end{aligned}$$

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

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:



What is the molar concentration of the acetic acid?

$$\frac{\text{L mol HC}_2\text{H}_3\text{O}_2}{\text{L solution}} \leftarrow = 25.0 \text{ mL or } 0.0250 \text{ L}$$

Since we already know the volume of the acid, all we really have to find out is the moles of acetic acid! (Once we know that, we can simply divide to get the concentration)

$$0.150 \text{ mol NaOH} = \text{L} \quad | \quad \text{mL} = 10^{-3} \text{ L} \quad | \quad \text{mol NaOH} = 2 \text{ mol HC}_2\text{H}_3\text{O}_2$$

$$37.3 \text{ mL} \times \frac{10^{-3} \text{ L}}{\text{mL}} \times \frac{0.150 \text{ mol NaOH}}{\text{L}} \times \frac{\text{mol HC}_2\text{H}_3\text{O}_2}{\text{mol NaOH}} =$$

$$= 0.005595 \text{ mol HC}_2\text{H}_3\text{O}_2$$

To get molarity, we need to divide by the volume of ACETIC ACID SOLUTION:

$$M = \frac{\text{mol HC}_2\text{H}_3\text{O}_2}{\text{L solution}} = \frac{0.005595 \text{ mol HC}_2\text{H}_3\text{O}_2}{0.0250 \text{ L}} = \boxed{0.224 \text{ M HC}_2\text{H}_3\text{O}_2}$$