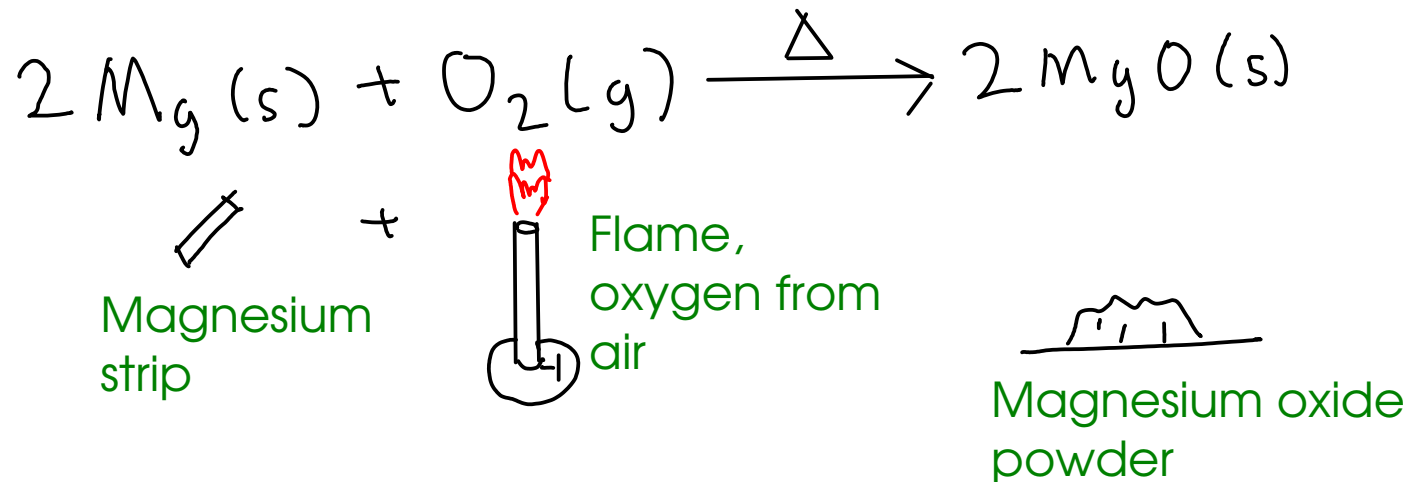


CONCEPT OF LIMITING REACTANT

- When does a chemical reaction STOP?



- 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.

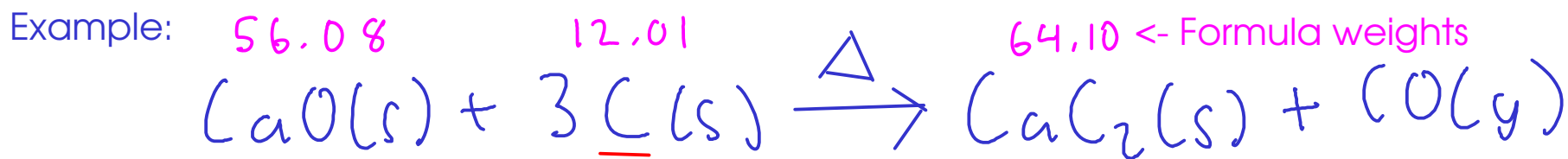
- This reaction is controlled by the amount of available magnesium

- 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 amount of product is the limiting reactant, and the smallest amount of product is the actual amount of product produced.



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

① $56.08 \text{ g CaO} = \text{mol CaO}$ ② $\text{mol CaO} = \text{mol CaC}_2$ ③ $64.10 \text{ g CaC}_2 = \text{mol CaC}_2$

$$100. \text{ g CaO} \times \frac{\text{mol CaO}}{56.08 \text{ g CaO}} \times \frac{\text{mol CaC}_2}{\text{mol CaO}} \times \frac{64.10 \text{ g CaC}_2}{\text{mol CaC}_2} = \boxed{114 \text{ g CaC}_2} < 177 \text{ g}$$

① $12.01 \text{ g C} = \text{mol C}$ ② $3 \text{ mol C} = \text{mol CaC}_2$ ③ $64.10 \text{ g CaC}_2 = \text{mol CaC}_2$

$$100. \text{ g C} \times \frac{\text{mol C}}{12.01 \text{ g C}} \times \frac{\text{mol CaC}_2}{3 \text{ mol C}} \times \frac{64.10 \text{ g CaC}_2}{\text{mol CaC}_2} = 177 \text{ g CaC}_2$$

This reaction should produce 114 grams of calcium carbide. At this point, all the CaO is gone and the reaction will stop. We can say that CaO is LIMITING and 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!

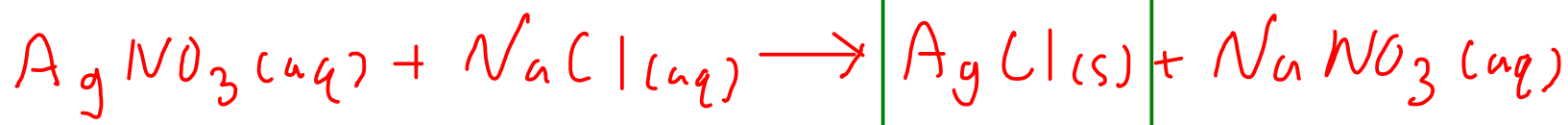
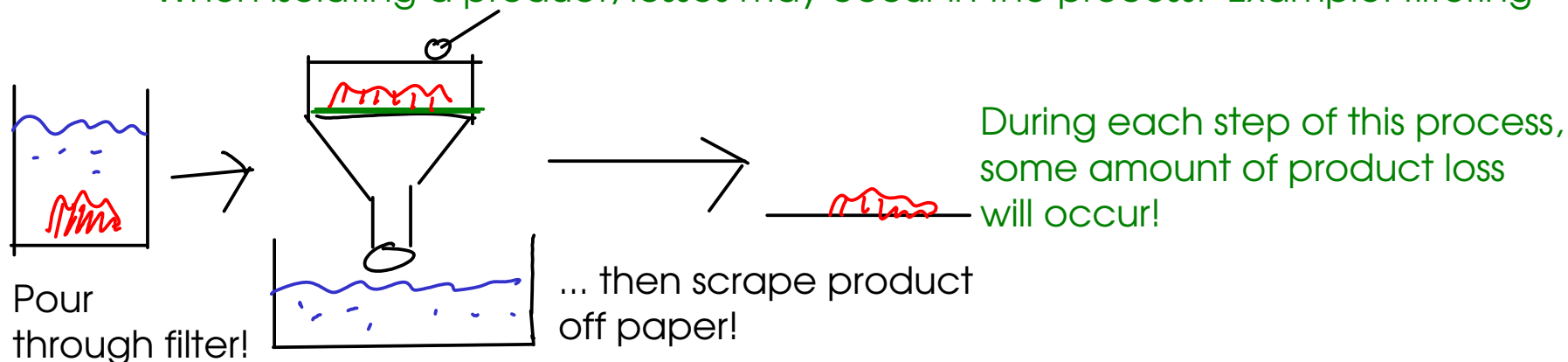
① SIDE REACTIONS:



... 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!

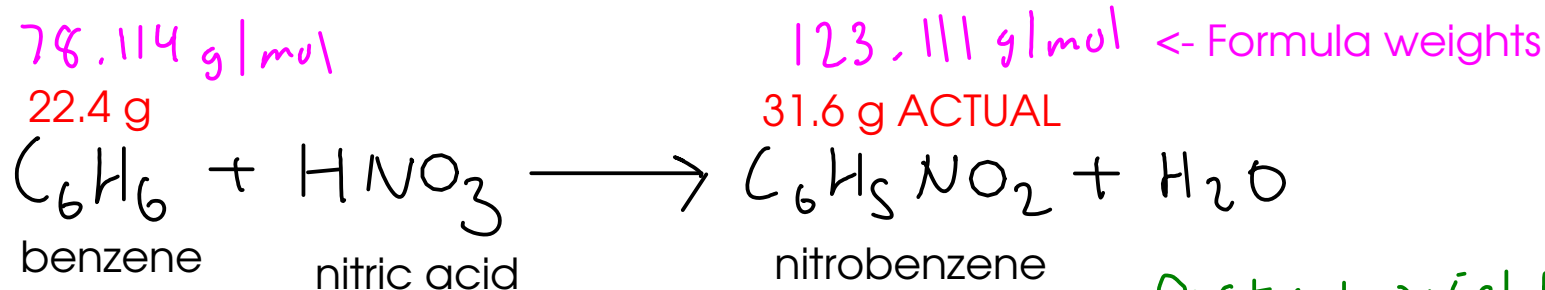
- All of these factors cause a chemical reaction to produce LESS product than calculated. For many reactions, this difference isn't significant. But for others, we need to report the PERCENT YIELD.

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

↙ Determined EXPERIMENTALLY!

↑ 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?

Calculate the THEORETICAL YIELD based on 22.4 grams of BENZENE (our starting material!)

$$\textcircled{1} 78.114 \text{ g C}_6\text{H}_6 = \text{mol C}_6\text{H}_6 \quad \textcircled{2} \text{ mol C}_6\text{H}_6 = \text{mol C}_6\text{H}_5\text{NO}_2$$

$$\textcircled{3} 123.111 \text{ g C}_6\text{H}_5\text{NO}_2 = \text{mol C}_6\text{H}_5\text{NO}_2$$

$$22.4 \text{ g C}_6\text{H}_6 \times \frac{\text{mol C}_6\text{H}_6}{78.114 \text{ g C}_6\text{H}_6} \times \frac{\text{mol C}_6\text{H}_5\text{NO}_2}{\text{mol C}_6\text{H}_6} \times \frac{123.111 \text{ g C}_6\text{H}_5\text{NO}_2}{\text{mol C}_6\text{H}_5\text{NO}_2} =$$

$$= 35.3 \text{ g C}_6\text{H}_5\text{NO}_2 \text{ (THEORETICAL YIELD)}$$

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