You can combine all three steps on one line if you like!

$$25.0\,g\,Br_2 \times \frac{1\,mol\,Br_2}{159.80\,g\,Br_2} \times \frac{2\,mol\,Al}{3\,mol\,Br_2} \times \frac{26.98\,g\,Al}{1\,mol\,Al} = \boxed{2.81\,g\,Al}$$

1 2 3

You can solve the second part of the question using CONSERVATION OF MASS - since there's only a single product and you already know the mass of all reactants.

$$\begin{array}{r} 25.0\,g\,Br_2 \\ +\ 2.81\,g\,Al \\ \hline 27.8\,g\,AlBr_3 \end{array}$$

But ...

...what would you have done to calculate the mass of aluminum bromide IF you had NOT been asked to calculate the mass of aluminum FIRST?

$$25.0\,g\,Br_2 \times \frac{1\,mol\,Br_2}{159.80\,g\,Br_2} \times \frac{2\,mol\,AlBr_3}{3\,mol\,Br_2} \times \frac{266.694\,g\,AlBr_3}{1\,mol\,AlBr_3} = 27.8\,g\,AlBr_3$$

1 convert mass bromine to moles

2 convert moles bromine to moles aluminum bromide

3 convert moles aluminum bromide to mass

101 Example:

How many milliliters of 6.00M hydrochloric acid is needed to completely react with 25.0 g of sodium carbonate?

$$2HCl(aq) + Na_2CO_3(s) \longrightarrow H_2O(l) + CO_2(g) + 2NaCl(aq)$$

1 - Convert 25.0 grams sodium carbonate to moles. Use FORMULA WEIGHT.
2 - Convert moles sodium carbonate to moles HCl. Use CHEMICAL EQUATION.
3 - Convert moles HCl to volume. Use MOLAR CONCENTRATION (6.00 moles/L)

1 Na_2CO_3 : Na: 2 x 22.99
C: 1 x 12.01
O: 3 x 16.00

105.99 g Na_2CO_3 = mol Na_2CO_3

$$25.0\text{ g } Na_2CO_3 \times \frac{\text{mol } Na_2CO_3}{105.99\text{ g } Na_2CO_3} = 0.2358713086\text{ mol } Na_2CO_3$$

2 2 mol HCl = mol Na_2CO_3

$$0.2358713086\text{ mol } Na_2CO_3 \times \frac{2\text{ mol HCl}}{\text{mol } Na_2CO_3} = 0.4717426172\text{ mol HCl}$$

Example:

How many milliliters of 6.00M hydrochloric acid is needed to completely react with 25.0 g of sodium carbonate?

$$2HCl(aq) + Na_2CO_3(s) \longrightarrow H_2O(l) + CO_2(g) + 2NaCl(aq)$$

1 - Convert 25.0 grams sodium carbonate to moles. Use FORMULA WEIGHT.
2 - Convert moles sodium carbonate to moles HCl. Use CHEMICAL EQUATION.
3 - Convert moles HCl to volume. Use MOLAR CONCENTRATION (6.00 moles/L)

3 $6.00 \text{ mol HCl} = \text{L}$

$$0.4717426172 \text{ mol HCl} \times \frac{\text{L}}{6.00 \text{ mol HCl}} = 0.0786 \text{ L of } 6.00 \text{ M HCl}$$

... but the problem asks us for mL, not L. Not a problem, though - just convert L to mL.

$\text{mL} = 10^{-3}\text{L}$

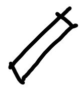
$$0.0786 \text{ L} \times \frac{\text{mL}}{10^{-3}\text{L}} = \boxed{78.6 \text{ mL of } 6.00 \text{ M HCl}}$$

Tip: In most chemical calculation problems, we start by converting a given AMOUNT of substance to MOLES. We usually can't start with a conversion factor (a ratio, like molarity or formula weight).

CONCEPT OF LIMITING REACTANT

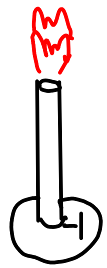
- When does a chemical reaction STOP?

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



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 12.01 64.10 <- Formula weights

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

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

CaO: 56.08 g CaO = mol CaO | mol CaO = mol CaC_2 | 64.10 g CaC_2 = mol CaC_2

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

C: 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 grams of calcium carbide could be produced by the reaction. When we reach 114 grams of calcium carbide, we've consumed all of the CaO present, and the reaction stops. We say that CaO is LIMITING, and C is present in excess. There will still be C left over after the reaction stops.

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

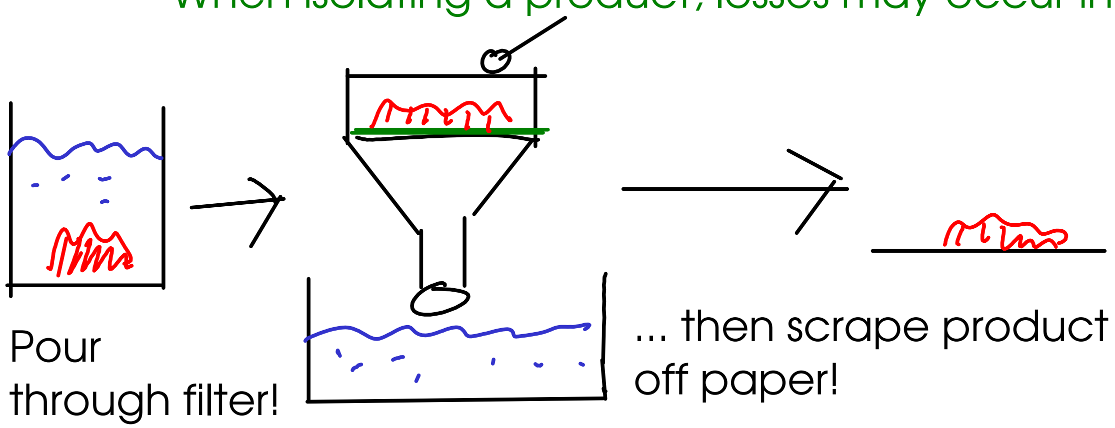
$$2C + O_2 \longrightarrow 2CO$$

... 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 determine the percent yield, we need to calculate the THEORETICAL YIELD based on the amount of starting material - the 22.4 grams of benzene.

78.114 g C_6H_6 = mol C_6H_6 | mol C_6H_6 = mol $C_6H_5NO_2$

123.111 g $C_6H_5NO_2$ = mol $C_6H_5NO_2$

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

$= 35.3\,g\,C_6H_5NO_2$ (THEORETICAL YIELD)

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

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?

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

Since we already know the volume of acetic acid, all we realy have to find is the moles of acetic acid. We can't calculate that directly, but we CAN relate it to the amount of NaOH. So, start by finding out how many moles NaOH were reacted.

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

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

To get concentration, divide moles acid and volume OF THE ACID SOLUTION

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