151.90 g/mol

$$10\,FeSO_4 + 2\,KMnO_4 + 8\,H_2SO_4 \rightarrow 5\,Fe_2(SO_4)_3 + 2\,MnSO_4 + K_2SO_4 + 8\,H_2O$$

How many mL of 0.250M potassium permangenate are needed to react with 3.36 g of iron(II) sulfate?

1 - Convert 3.36 grams of iron(II) sulfate to moles. Use FORMULA WEIGHT.
2 - Convert moles iron(II) sulfate to moles potassium permangenate. Use CHEMICAL EQUATION.
3 - Convert moles potassium permangenate to volume. Use MOLARITY (0.250M)

1 $151.90\,g\ FeSO_4 = mol\ FeSO_4$ 2 $10\ mol\ FeSO_4 = 2\,mol\ KMnO_4$

3 $0.250\ mol\ KMnO_4 = L$

$$3.36\,g\ FeSO_4 \times \underset{1}{\frac{mol\ FeSO_4}{151.90\,g\ FeSO_4}} \times \underset{2}{\frac{2\,mol\ KMnO_4}{10\ mol\ FeSO_4}} \times \underset{3}{\frac{L}{0.250\ mol\ KMnO_4}} = 0.0177\ L \text{ of } 0.250\,M\ KMnO_4$$

Let's convert L to mL because the problem specifically asks for mL.

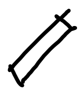
$mL = 10^{-3}\,L$

$$0.0177\ L \times \frac{mL}{10^{-3}\,L} = \boxed{17.7\ mL \text{ of } 0.250\,M\ KMnO_4}$$

CONCEPT OF LIMITING REACTANT

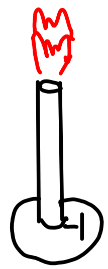
- When does a chemical reaction STOP?

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



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

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

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) $64.10\text{ g } CaC_2 = \text{mol } CaC_2$

$$100.\text{ g } CaO \times \frac{\text{mol } CaO}{56.08\text{ g } CaO} \times \frac{\text{mol } CaC_2}{\text{mol } CaO} \times \frac{64.10\text{ g } CaC_2}{\text{mol } CaC_2} = 114\text{ g } CaC_2 \longleftarrow \text{ANSWER}$$

(1) $12.01\text{ g } C = \text{mol } C$ (2) $3\text{ mol } C = \text{mol } CaC_2$ (3) $64.10\text{ g } CaC_2 = \text{mol } CaC_2$

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

The reaction will stop when we reach 114 grams of calcium carbide. At that point, there is still leftover carbon, but nothing for that carbon to react with as all the CaO is gone! We say that CaO is LIMITING, and C 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

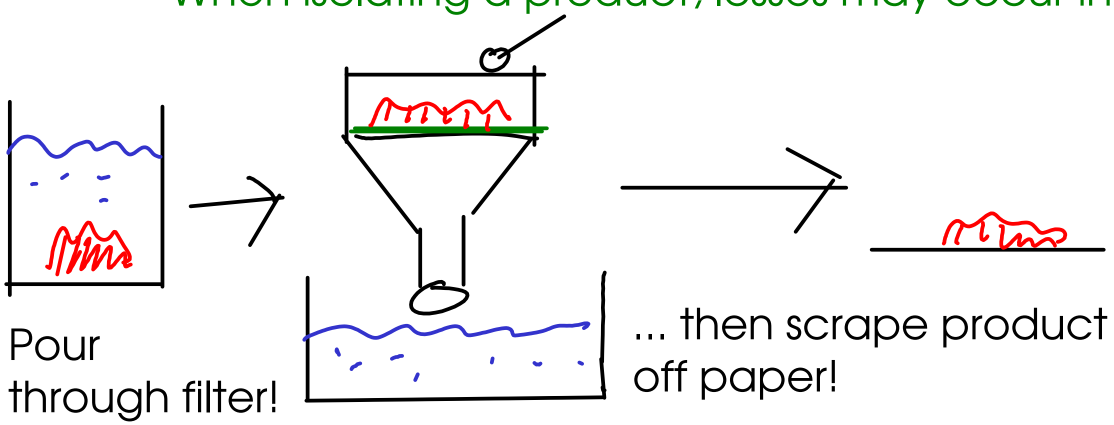
$$2C + O_2 \longrightarrow 2CO$$

... 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!

$78.114\ g/mol$ $123.111\ g/mol$ <- 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?

Start by calculating the amount of nitrobenzene we could make if all the 22.4 grams of benzene were reacted. This is the THEORETICAL YIELD.

① $78.114\ g\ C_6H_6 = mol\ C_6H_6$ ② $mol\ C_6H_6 = mol\ C_6H_5NO_2$

③ $123.111\ g\ C_6H_5NO_2 = mol\ C_6H_5NO_2$

$$22.4\,g\ C_6H_6 \times \underset{①}{\frac{mol\ C_6H_6}{78.114\,g\ C_6H_6}} \times \underset{②}{\frac{mol\ C_6H_5NO_2}{mol\ C_6H_6}} \times \underset{③}{\frac{123.111\,g\ C_6H_5NO_2}{mol\ C_6H_5NO_2}} = 35.3\,g\ C_6H_5NO_2$$

(theoretical)

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