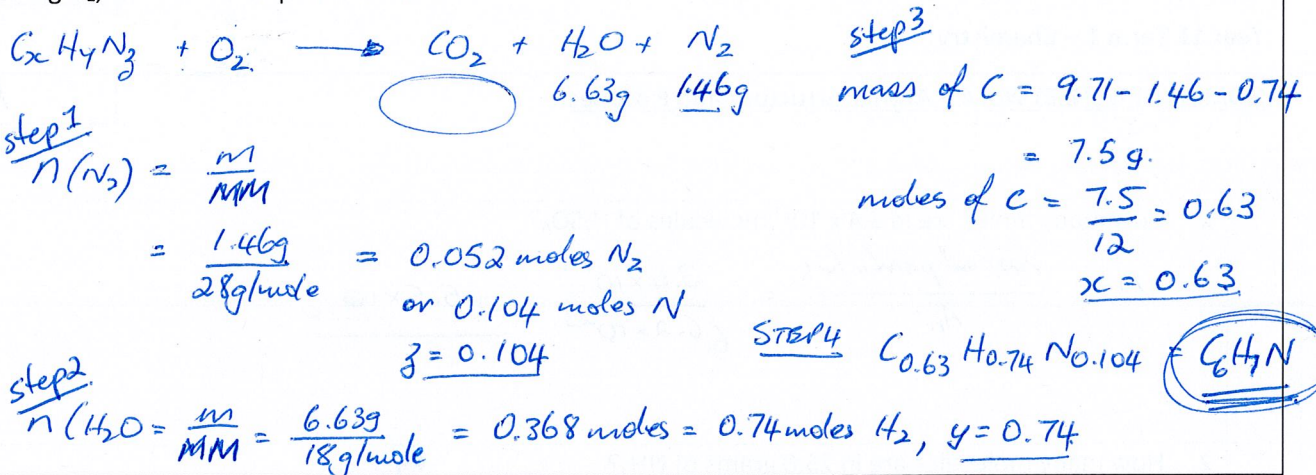
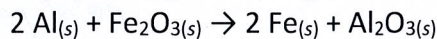


/1	<p>1. How many moles are in 3.4×10^{23} molecules of H_2SO_4?</p> $n = \frac{\text{no. of particles}}{A_n} = \frac{3.4 \times 10^{23}}{6.02 \times 10^{23}} = \underline{0.56 \text{ moles}}$
/2	<p>2. How many molecules are in 25.0 grams of NH_3?</p> $n = \frac{m}{MM} = \frac{25\text{g}}{17\text{g/mole}} = 1.47 \text{ moles}$ $\text{no of particles} = n \times A_n = 1.47 \times 6.02 \times 10^{23} = \underline{8.85 \times 10^{23} \text{ molecules}}$
/2	<p>3. How many grams are in 8.200×10^{22} molecules of N_2I_6?</p> $n = \frac{\text{no. of particles}}{A_n} = \frac{8.2 \times 10^{22}}{6.02 \times 10^{23}} = 0.136 \text{ moles}$ $m = n \times MM = 0.136 \times 789.4 = 107.4 \text{ g.}$
/2	<p>4. What's the empirical formula of a molecule containing 18.7% lithium, 16.3% carbon, and 65.0% oxygen?</p> <p>assuming 100g = 18.7g Li, 16.3g C, 65g O</p> $n(\text{Li}) = \frac{m}{MM} = \frac{18.7\text{g}}{6.94\text{g/mol}} = 2.7 \text{ moles}$ $n(\text{C}) = \frac{m}{MM} = \frac{16.3\text{g}}{12\text{g/mole}} = 1.36 \text{ moles}$ $n(\text{O}) = \frac{m}{MM} = \frac{65\text{g}}{16\text{g/mole}} = 4.06$ <p>$\text{Li}_{2.7} \text{C}_{1.36} \text{O}_{4.06}$</p> <p>÷ subscripts by lowest value (1.36)</p> <p><u>$\text{Li}_2 \text{C}_1 \text{O}_3$</u></p>
/1	<p>5. If the molar mass of the compound in question 4 is 73.8 grams/mole, what's the molecular formula?</p> <p>Empirical formula = Li_2CO_3 $M.M = 2 \times 6.94 + 12 + 3 \times 16 = 73.88\text{g.}$</p> <p><u>$\text{Li}_2\text{CO}_3$</u></p> <p>is approx equal to MM of the compound therefore molecular formula is the same as the empirical formula</p>

6. ** Aniline, a starting material for urethane plastic foams, consists of C, H, and N. Combustion of such compounds yields: CO_2 , H_2O , and N_2 as products. If the combustion of 9.71g of aniline yields 6.63 g H_2O and 1.46g N_2 , what is its empirical formula?

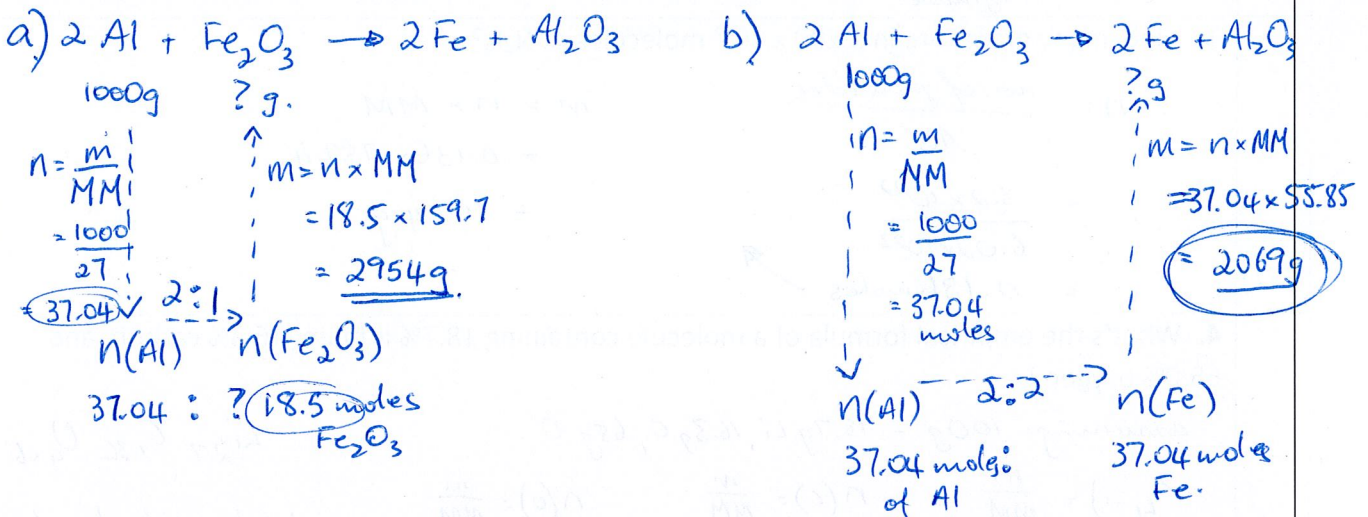


7. The reaction of aluminium and iron(III) oxide gives off a great deal of heat and light:

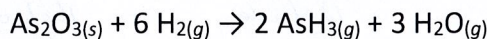


a. What mass of $\text{Fe}_2\text{O}_3(s)$ is required to react completely with 1.0 kg of $\text{Al}(s)$?

b. What mass of $\text{Fe}(s)$ will be produced based on the quantities in a)?

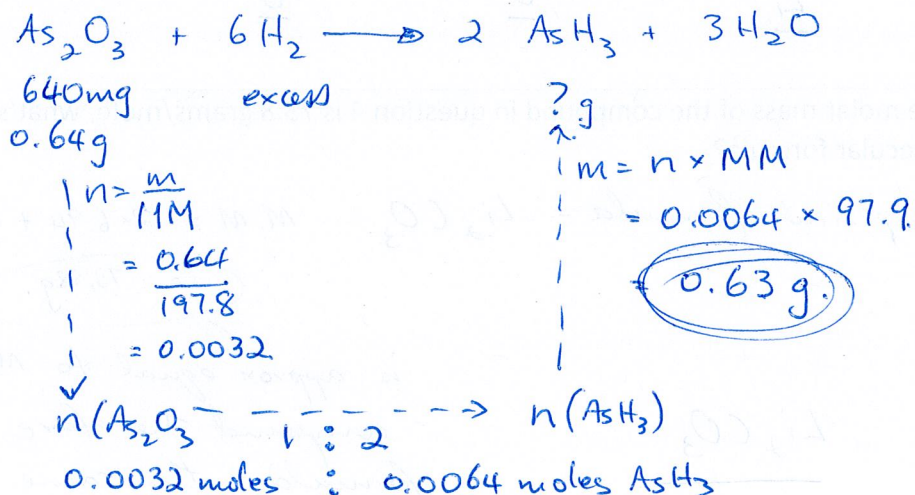


8. The Marsh test was used historically to detect arsenic in cases of suspected poisoning:

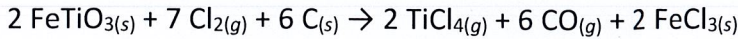


A sample contains 640 mg of As_2O_3 . What mass of AsH_3 will be isolated from the above reaction?

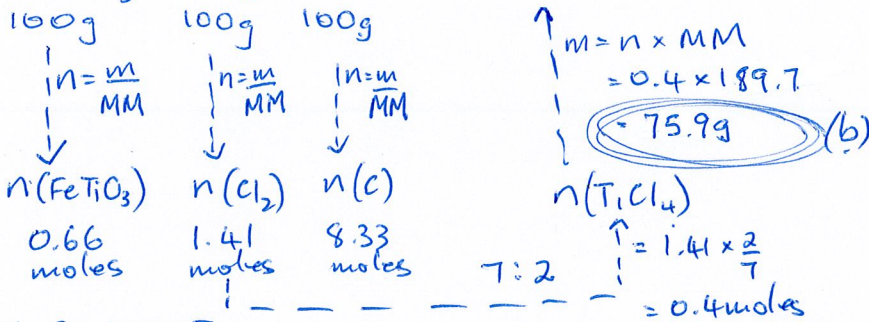
Assume that there is enough (excess) H_2 to react with all the As_2O_3 .



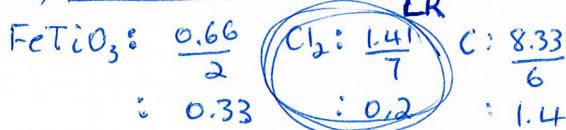
9. Titanium (IV) chloride can be made by the following reaction:



- If you start with 100.0 g of each reactant, which one limits the reaction?
- What is the theoretical yield of TiCl_4 based on the amounts in a)?
- If only 70.0 g of TiCl_4 is recovered, what is the percent yield of the reaction?



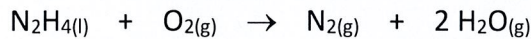
a) LIMITING REAGENT



(c) % yield = $\frac{\text{actual}}{\text{theoretical}} \times 100$

$= \frac{70.0}{75.9} \times 100 = 92\%$

10. Hydrazine is used in rocket fuel. It reacts with Oxygen according to the equation below



In a particular rocket engine, 2.29 g of hydrazine and 3.14 g of Oxygen are available to react.

- Identify the limiting reagent and show your calculations
- Determine the mass of unreacted reagent that will be left after the reaction and show your calculations
- Calculate the mass of water produced from the amounts stated in a).



a) L. Reagent.

$$n(\text{N}_2\text{H}_4) = \frac{m}{MM} = \frac{2.29}{32} = 0.072$$

$$n(\text{O}_2) = \frac{m}{MM} = \frac{3.14}{32} = 0.098$$

coeff are 1:1

∴ $\text{N}_2\text{H}_4 = \text{LR}$

(b)

$$n(\text{O}_2 \text{ left over}) = 0.098 - 0.072 = 0.026 \text{ moles}$$

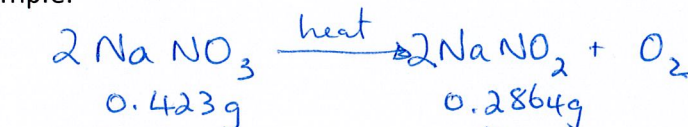
$$m(\text{O}_2) = n \times MM = 0.026 \times 32 = 0.83\text{g}$$

(c) $n(\text{H}_2\text{O}) = 2 \times n(\text{N}_2\text{H}_4)$

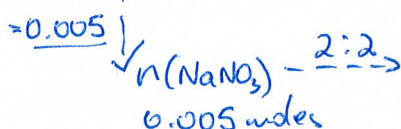
$$= 2 \times 0.072 = 0.144$$

$$m(\text{H}_2\text{O}) = n \times MM = 0.144 \times 18 = 2.6\text{g}$$

11. ** A 0.423 g sample of impure Sodium Nitrate was heated, converting all the Sodium Nitrate to 0.2864 g of Sodium Nitrite and Oxygen gas. Determine the percent of sodium Nitrate in the original sample.



$$n = \frac{m}{MM} = \frac{0.423}{85} = 0.005$$



$$m = n \times MM = 0.005 \times 69 = 0.345$$

$$\% \text{ yield} = \frac{\text{actual}}{\text{theoretical}} \times 100 = \frac{0.2864}{0.345} \times 100 = 83\%$$

this indicates the % purity of the original sample.