Lesson 4: Stoichiometry

*Read through the lesson notes. You can write them out, print them or save them.

*Once you have tried to understand the lesson answer the questions that follow and self-evaluate your work by checking the answers.

Learning Intention

-Practice different examples of stoichiometric calculations.

(This lesson does not involve any new theory. It builds on the use of important relationships such as n = c x v and m = n x GFM). It will be set out as a series of worked examples for you to understand and then to practice.

Background

Stoichiometry is the study of mole relationships involved in chemical reactions. We have previous worked through the concept of stoichiometry in National 5 and Higher Chemistry. It is basically those questions that involve understanding balanced equations and using mole ratios to calculate the mass, number of moles or concentration of a substance.

A balanced equation simply informs us how many moles of a substance will react or be produced in relation to another substance.

Example

$Mg + 2HCl \rightarrow MgCl_2 + H_2$

1 mole 2 moles 1 mole 1 mole

For Advanced Higher Chemistry, we are often asked to consider the concentration of a substance present in a commercial or laboratory sample of material. For example, you may be asked to analyse and calculate the concentration of ions in bleach or the mass of metals in alloys. Normally, the analysis is carried out by a volumetric titration and therefore the two key concepts are:

n = c x v

AND

m = n x GFM

Worked Example 1

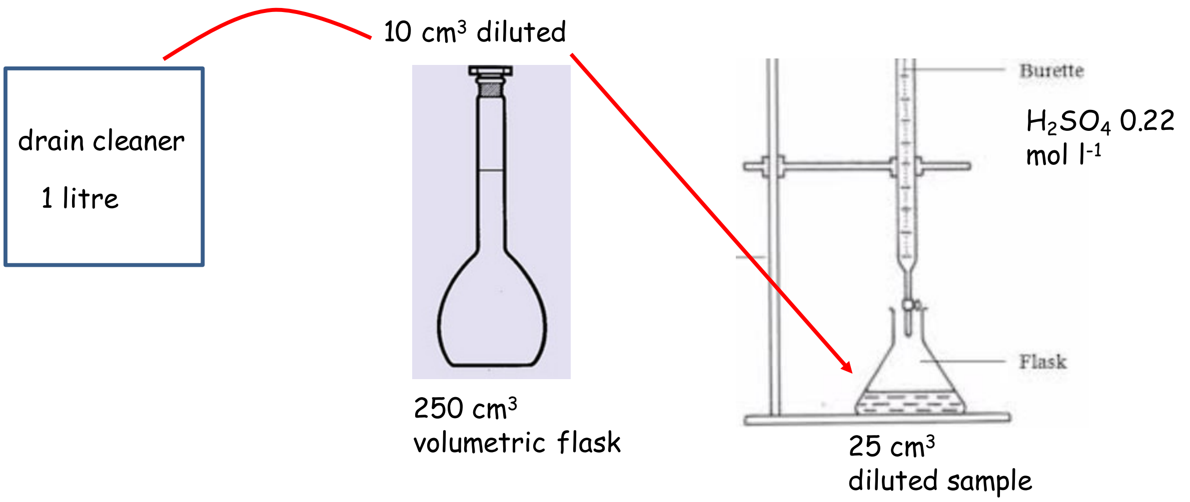
To determine the mass of sodium hydroxide, NaOH, in a commercial drain cleaner the following experiment was set up.

10cm^3 of the liquid drain cleaner containing NaOH was diluted to 250cm^3 in a volumetric flask.
25cm^3 samples of this diluted solution were pipetted into a conical flask and titrated against 0.22mol l^{-1} H_2SO_4. The average of the concordant titres was 17.8cm^3.

Calculate the mass of NaOH in 1 litre of the drain cleaner.

The examples you will come across in Advanced Higher often have more than one step to do when compared with mole ratio questions from National 5 and Higher Chemistry. As there is more reading to the question, it is often a good idea to take the information and draw it as a diagram to have a clearer picture of what you are being asked to do.

The steps involved in the experiment can be summarised in the following way.



1. 10cm^3 of drain cleaner is taken from the container.
2. It is diluted to 250cm^3 in a volumetric flask.
3. 25cm^3 of the diluted sample is titrated against 0.22mol l^{-1} H_2SO_4.

However, when answering the question, we set it out in the following order:

step 3 → step 2 → step 1.

Step 3 Titration (calculate number of moles of NaOH)

H_2SO_4 + 2NaOH → Na_2SO_4 + H_2O

n=?

c=0.22 mol l^{-1}

v=0.0178 litres (17.8cm^3)

n= c x v
n= 0.22 x 0.0178
n= 0.003916 moles

H_2SO_4 : 2NaOH
0.003916 : 0.007832moles

Step 2 Volumetric Flask (scale up)

0.007832 moles are present in each 25cm^3 sample
0.07832 moles (x10) in 250cm^3 volumetric flask (as its volume is ten times greater)

Step 1 Drain cleaner from original bottle (scale up)
The original sample taken from the bottle was 10cm^3
0.07832 moles NaOH in 10cm^3
7.832 moles NaOH in 1 litre (1000cm^3, as the volume is one hundred times greater)

Therefore, the mass of NaOH in the original bottle of drain cleaner is:

m = n x GFM
m = 7.832 x 40
m = 313.3g

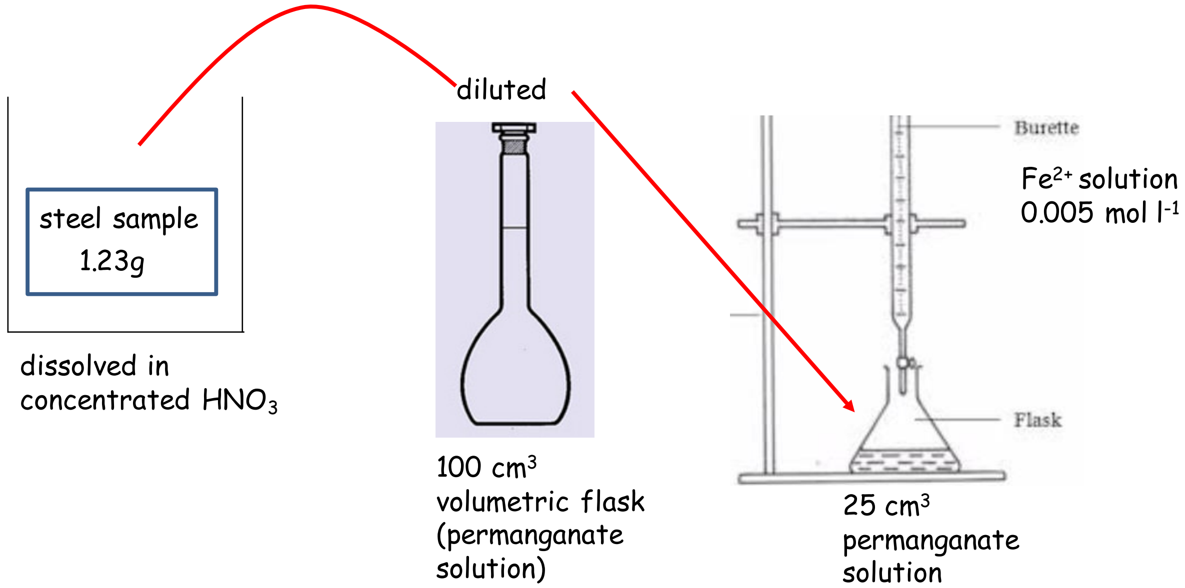
Worked Example 2

A 1.23g sample of steel containing manganese was dissolved in concentrated nitric acid. The Mn^{2+} ions formed were then oxidised to permanganate ions. The resulting purple solution was made up to 100cm^3 in a volumetric flask.

In a titration, a 25cm^3 sample of the permanganate solution was reduced by 34.5cm^3 of 0.005 mol l^{-1} iron (II) solution. The equation for the titration is:

$5Fe^{2+} + MnO_4^- + 8H^+ \rightarrow Mn^{2+} + 5Fe^{3+} + 4H_2O$

Calculate the percentage by mass of manganese in the original sample of steel.



1. Dissolve steel in concentrated nitric acid.
2. Dilute the sample to 100 cm^3 in a volumetric flask.
3. 25 cm3 sample is diluted against 0.005 mol l^{-1} Fe^{2+} solution.

Step 3 Titration (calculate number of moles of manganese)

$5Fe^{2+} \quad + \quad MnO_4^- \quad + \quad 8H^+ \quad \rightarrow \quad Mn^{2+} \quad + \quad 5Fe^{3+} \quad + \quad 4H_2O$

n=?

c=0.005 mol l^{-1}

v=0.0345 litres (34.5cm^3)

n= c x v
n= 0.005 x 0.0345
n= 0.0001725 moles

$5Fe^{2+}$: MnO_4^-
0.0001725 : 0.0000345moles

Step 2 Volumetric Flask (scale up)

0.0000345 moles are present in each 25cm^3 sample
0.000138 moles (x4) in 100cm^3 volumetric flask (as its volume is four times greater)

Step 1 Calculate mass of manganese in steel sample

m = n x GFM
m = 0.000138 x 54.9g
m = 0.0076g

The percentage of manganese in the steel sample is given by:

$$\frac{0.0076\text{g (mass of manganese)}}{1.23\text{g (mass of steel)}} \times 100$$

0.62%

→Watch the clip on Youtube.

https://www.youtube.com/watch?v=UL1jmJaUkaQ

→ Read Scholar Heriot-Watt/ Researching Chemistry Section 3.1, 3.3.

→ Answer the questions from Sheet 4.4 and check the answers when you have completed them.

4.4 Stoichiometry

1. A bottle of vinegar was analysed to find the concentration of ethanoic acid that it contained. A 25cm^3 sample of the vinegar was taken and diluted to 250cm^3 in a volumetric flask. 10cm^3 samples of the diluted vinegar were than titrated with standardised 0.05mol l^{-1} NaOH. The average volume of NaOH used in the titration was 18.4 cm^3.

   $CH_3COOH\ (aq)\ +\ NaOH\ (aq)\ \rightarrow\ CH_3COONa\ (aq)\ +\ H_2O\ (l)$

   Calculate the concentration of ethanoic acid in the vinegar.

2. Sodium carbonate was used to find the concentration of a sample of hydrochloric acid. A 2.43g sample of Na_2CO_3 was dissolved and made up to 250cm^3 with distilled water in a volumetric flask.10cm^3 portions of the sodium carbonate solution were pipetted into a conical flask and titrated with the hydrochloric acid. The average titre of hydrochloric acid for the reaction was 22.3cm^3. The equation for the reaction is:

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

   Calculate the concentration of the hydrochloric acid.

3. A 2.49g steel bolt was analysed to find the percentage of iron it contained. The bolt was dissolved in an excess of dilute sulfuric acid converting all the iron to iron (II) ions.

   The solution was made up to 250cm^3 with distilled water in a volumetric flask. Several 25cm^3 samples of this solution were titrated with 0.05 mol l^{-1} potassium permanganate and the average titre was 16.4cm^3. The reactions taking place is:

   $5Fe^{2+}\ +\ MnO_4^-\ +\ 8H^+\ \rightarrow\ Mn^{2+}\ +\ 5Fe^{3+}\ +\ 4H_2O$

   a) Calculate the number of moles of permanganate solution present in the average titre.
   b) Calculate the number of moles of Fe^{2+} (aq) in an average 25cm^3 sample.
   c) Calculate the mass of iron in the bolt.
   d) Express this mass as a percentage of the iron in the steel.

4. A 5.82g sample of impure ethanedioic acid $(COOH)_2$, is dissolved in distilled water and made up to 250cm^3 in a volumetric flask. Then 25cm^3 portions of this solution are titrated against standard 0.1 mol l^{-1} potassium permanganate. The average titre to reach the end point was 24.2cm^3. The only reactions taking place is.

$2MnO_4^- \ + \ 5(COOH)_2 \ + \ 6H^+ \ + \ \rightarrow 10CO_2 \ + \ 2Mn^{2+} \ + \ 8H_2O$

Calculate the percentage purity, by mass, of the original ethanedioic acid sample.

5. A bottle of hydrogen peroxide solution, H_2O_2 (aq), was analysed as follows. A 5cm^3 sample of the solution was diluted to 250cm^3 in a volumetric flask. Several 25cm^3 samples of the diluted solution were titrated with standard 0.02 mol l^{-1} potassium dichromate solution. The reactions taking place were:

$Cr_2O_7^{2-}$ (aq) + $14H^+$ (aq) + 6e $\rightarrow$ $2Cr^{3+}$ (aq) + $7H_2O$ (l)

H_2O_2 (aq) $\rightarrow$ O_2 (g) + $2H^+$ (aq) + 2e

The average endpoint was observed when 35.8cm^3 of the dichromate solution was added.

a) Give the overall balanced redox equation.

b) Calculate the number of moles of hydrogen peroxide in the average 25cm^3 diluted sample.

c) Calculate the number of moles of hydrogen peroxide in the 5cm^3 sample which was taken from the original bottle.

d) Calculate the concentration of the hydrogen peroxide solution in the original bottle.

6. 25·0 cm^3 of an acidified solution of potassium oxalate, $K_2C_2O_4$, was heated to 80 °C and titrated with a standard solution of 0·020 mol l^{-1} potassium permanganate, $KMnO_4$.
The end-point was reached when 22·5 cm^3 of $KMnO_4$ solution had been added.

The equation for the reaction is:

$5C_2O_4^{2-} \ + \ 2MnO_4^- \ + \ 16H^+ \ + \ \rightarrow 10CO_2 \ + \ 2Mn^{2+} \ + \ 8H_2O$

Calculate the concentration of the potassium oxalate solution used in the titration.

7. 1·80 g of iron(II) ammonium sulfate, $Fe(NH_4)_2(SO_4)_2.6H_2O$, was dissolved in 35 cm^3 of distilled water. The solution was then diluted to 50 cm^3 using dilute sulfuric acid. The final solution was titrated against potassium permanganate solution. 10 cm^3 of the iron(II) ammonium sulfate solution required 30 cm^3 of the permanganate solution to reach the end point at which all the iron(II) had been converted to iron(III).

   The equation for the titration is:

   $$5Fe^{2+} + MnO_4^- + 8H^+ \rightarrow Mn^{2+} + 5Fe^{3+} + 4H_2O$$

   a) Calculate the number of moles of iron(II) ions present in 1·80 g of iron(II) ammonium sulfate.
   b) How many moles of iron (II) ions are present in the 10 cm^3 portion of iron(II) ammonium sulfate used in the titration.
   c) Calculate the concentration of permanganate solution used in the titration.

8. The concentration of a thiosulfate solution can be determined by following 2 steps.

   Step 1: Excess Iron(III) nitrate was added to 40·0 cm^3 of 0·10 mol l^{-1} of potassium iodide solution producing iron(II) ions and iodine.

   $$2Fe^{3+} + 2I^- \rightarrow 2Fe^{2+} + I_2$$

   Step 2: The iodine formed was titrated against sodium thiosulfate solution. 15·0 cm^3 of sodium thiosulfate solution was required for complete reaction. The reaction can be represented by the equation:

   $$I_2(aq) + 2S_2O_3^{2-}(aq) \text{ ---> } 2I^-(aq) + S_4O_6^{2-}(aq)$$

   a) Calculate the number of moles of iodine produced from the potassium iodide.
   b) Calculate the concentration of the sodium thiosulfate solution.

9. Brass is an alloy consisting mainly of copper and zinc. To determine the percentage of copper in a sample of brass, 3·13 g of the brass was dissolved in concentrated nitric acid and the solution diluted to 100 cm^3 in a volumetric flask. Excess potassium iodide was added to 10·0 cm^3 of this solution, iodine being produced according to the equation:

   Equation 1 $2Cu^{2+}(aq) + 4I^-(aq) \rightarrow 2CuI(s) + I_2(aq)$

   The iodine formed was titrated with 0·10 mol l^{-1} thiosulfate solution, $S_2O_3^{2-}$. The volume of thiosulfate solution required for complete reaction was 24·8 cm^3.

   Equation 2 $I_2(aq) + 2S_2O_3^{2-}(aq) \rightarrow 2I^-(aq) + S_4O_6^{2-}(aq)$

   Calculate the percentage by mass of copper in the sample of brass.