CHEMICAL ANALYSIS

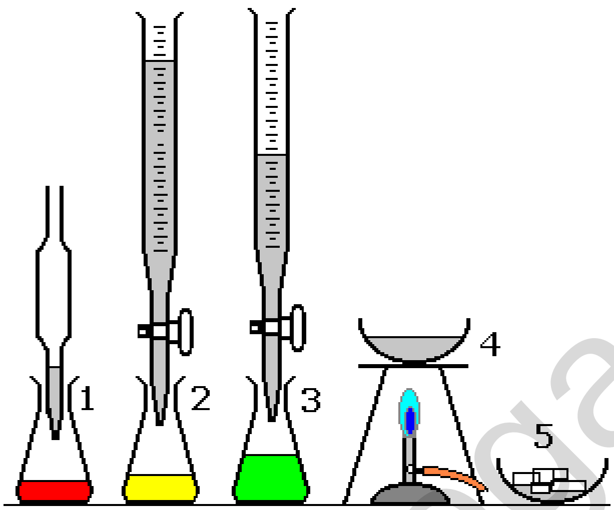
Titration

Titration is a procedure for determining the concentration of a solution by allowing a carefully measured volume to react with a standard (of known concentration) solution of another substance. An acid-base titration is quite common, and is based on the neutralization reaction between an acid and a base. In the laboratory, such a titration can be monitored using an acid-base indicator or an instrument called a pH meter.

(1) A known volume of acid is pipetted into a conical flask and universal **indicator** added. The acid is titrated with the alkali in the burette

(2) until the indicator turns green.

(3). The **volume of alkali needed for neutralisation is then noted**, this is called the **endpoint**. (1-3) are **repeated** with both known volumes mixed together BUT **without the contaminating indicator**.



The volume of standard solution required for complete reaction can be used to calculate the concentration of the unknown solution. Let's consider a titration of HCl with KOH. Results of a titration indicate that 29.7 mL of a 0.100M solution of KOH was required to completely neutralize a 15.0 cm^3 sample of HCl. We begin by writing a balanced equation for the titration reaction.

$$HCl(aq) + KOH(aq) \longrightarrow H_2O(l) + KCl(aq)$$

	HCl	KOH
	15.0mL	29.7 cm^3
	M ?	0.100 M

Because we know the concentration and the volume of the KOH solution, we can calculate the number of moles of KOH used in the reaction.

$$\text{Number of moles of KOH in } 29.7\ cm^3 = \frac{29.7 \times 0.1}{1000}$$

$$= 2.97 \times 10^{-3} \text{ mol KOH}$$

The number of moles of KOH is related to the number of moles of HCl by the stoichiometric coefficients of the balanced chemical equation. In this case, there is a 1:1 ratio. Don't be tempted to omit this step, however, because the ratio is not always 1:1.

$$2.97 \times 10^{-3} \text{ mol KOH } (1 \text{ mol HCl} / 1 \text{ mol KOH}) = 2.97 \times 10^{-3} \text{ mol HCl}$$

We can now calculate the concentration of the HCl solution by dividing moles by volume.

2.1×10^{-3} mol HCl / 0.015 dm^3 = 0.198 mol / dm^3 HCl = 0.198 M HCl

<u>Titration of sulphuric acid with sodium hydroxide.</u>

<u>Problem No. 1</u>

30 cm^3 of 0.1 mol/dm^3 NaOH (aq) reacted completely with 25 cm^3 of H_2SO_4 (aq) in a titration flask. Calculate the concentration of H_2SO_4 in (a) mol/dm^3 and (b) in g/dm^3. The equation for the reaction is

DATA

Concentration Of NaOH	=	0.1 mol/dm^3
Concentration of H_2SO_4	=	unknown
Volume of NaOH	=	30 cm^3
Volume of H_2SO_4	=	25 cm^3

Step 1: First find the number of moles of NaOH use in titration
1 dm3 contain 0.1 mol
0.03dm3 (30cm3) will contain 0.03 x 0.1 = 0.003 moles (in 30 ml)

Step 2:Write the chemical equation for the reaction.

$$2NaOH(aq) + H_2SO_4(aq) \longrightarrow Na_2SO_4(aq) + 2H_2O(l)$$

Step 3: From the equation find the ratio of number of moles of H_2SO_4 to the number of moles of NaOH

NaOH : H_2SO_4
2 : 1

Step 4: Use the ration to find the number of moles of H_2SO_4 that reacted.

NaOH : H_2SO_4
0.003 : 0.0015

Step 5: Find the concentration of H_2SO_4 in moles/ dm^3.

0.0015 moles are present in 0.025 dm^3
x moles are present in 1 dm^3

$$\frac{0.0015 \times 1}{0.025} = 0.06 \text{ moles / } dm^3$$

Step 6: Find the concentration of H_2SO_4 in g/ dm^3

$0.06 \times 98 = 5.88\ g/dm^3$

Titration of iron (II) sulphate with potassium manganate (VII).

25.0 cm^3 of FeSO4(aq), acidified with sulphuric acid, required 27.5 cm^3 of 0.0200 mol/dm^3 $KMnO_4$ (aq) for reaction in a titration. Calculate the concentration of $FeSO_4$ (aq).

Step1: Find the number of moles of $KMnO_4$ used in titration

0.02 moles are present in 1 dm3
x moles are present in 0.0275 dm3 (27.5 cm_3)

$$\frac{0.02 \times 0.0275}{1} = 0.00055 \text{ molers in } 27.5 \text{ cm3}$$

Step 2: Write the chemical equation for the reaction.

$$2KMnO_4\ (aq) + 10FeSO_4\ (aq) + 8\ H_2SO_4\ (aq) \text{ -----------}\rightarrow K_2SO_4\ (aq) + 2MnSO_4\ (aq) + 5Fe_2(SO_4)_3(aq) + 8H_2O\ (l)$$

Step 3: From the equation find the ratio of number of moles of $FeSO_4$ to the number of moles of $KMnO_4$.

$KMnO_4$	:	$FeSO_4$
1	:	5

Step 4: Use the ratio to find the number of moles of $FeSO_4$ that reacted in the titration.

$KMnO_4$	:	$FeSO_4$
0.00055	:	00275

Step 5: Find the concentration of $FeSO_4$ (aq) in mol/ dm^3.

0.025 dm^3 contains 0.00275 moles
1 dm^3 contains

$$\frac{0.00275 \times 0.025}{1} = 0.11 \text{ moles / dm}^3$$

USES OF TITRATION IN ANALYSIS

1. Identification of Acids and Alkalis.

An acid has the formula H_2XO_4. One mole of H_2XO_4 reacts with two moles of NaOH. A solution of the acid contains 5.0 g/ dm^3 of H_2XO_4. In a titration, 25.0 cm^3 of the acid reacted with 25.5 cm^3 of 0.1 mol/ dm^3 NaOH (aq). Calculate the concentration of the acid in mol/ dm^3 and hence calculate the relative molecular mass of the acid.

Solution:

Concentration Of NaOH = 0.1 mol/dm^3
Concentration of H_2XO_4 = 50 g /dm^3
Volume of NaOH = 25 cm^3
Volume of H_2SO_4 = 25cm^3

Step 1: Write the balanced equation

$$H_2XO_4 + 2NaOH \longrightarrow Na_2XO_4 + 2H_2O$$

Step 2 : Find the numbers of moles of NaOH used in titration

No. of moles of NaOH used in titration = concentration x vol. in dm^3

$$= 0.1 \times \frac{25.5}{1000}$$

No. of moles of H_2XO_4 = ½ x no. of moles of NaOH

$$= \frac{1}{2} \times 0.1 \times \frac{25.5}{1000}$$

Concentration of the acid $= \frac{\text{no. of moles of } H_2XO_4}{\text{Vol. of acid in } dm^3}$

$$= \frac{\frac{1}{2} \times 0.1 \times \frac{25.5}{1000}}{\frac{25.0}{1000}}$$

$$= 0.051 \text{ mol/ } dm^3$$

1dm^3 acid solution contains 0.051 mol and 5.0 g of H_2XO_4.
So 0.051 mol of H_2XO_4 has a mass of 5.0 g of H_2XO_4.
And 1 mol. Of H_2XO_4 has a mass of 5.0/0.051 = 98g.

Hence the relative molecular mass of H_2XO_4 is 98.

The relative atomic mass of X = 98 – 66 =32.
So X is sulphur and the acid is H_2SO_4.

2. Percentage Purity of Compounds

Solution of X contains 5.00 g of impure sulphuric acid dissolved in 1 dm^3 of solution. 25.0 cm^3 of solution X required 23.5 cm^3 of 0.100 mol/ dm^3 NaOH for the reaction in titration. Calculate the percentage purity of acid.

$$\text{Percentage purity} = \frac{\text{mass of actual acid / dm}^3}{\text{Mass of impure acid/ dm}^3} \times 100$$

Solution:

No. of moles of NaOH used in titration $= \frac{23.5}{1000} \times 0.100$ mol.

The equation is

$$H_2SO_4 + 2NaOH \longrightarrow Na_2SO_4 + 2H_2O$$

From the equation,

No. of moles of H_2SO_4 = ½ x no. of moles of NaOH

$= \frac{1}{2} \times \frac{23.5}{1000} \times 0.100$ mol.

So the concentration of H_2SO_4

$$= \frac{\text{no of moles}}{\text{vol. in dm}^3}$$

$$= \frac{\frac{1}{2} \times \frac{23.5}{1000} \times 0.100}{\frac{25.5}{1000}}$$

$= 0.047$ mol/ dm^3.

Hence the no. of grams of H_2SO_4 in 1 dm^3

= 0.047 x Mr of H_2SO_4

= 0.047 x 98

= 4.61 g.

Hence the percentage purity $= \frac{4.61}{5.00} \times 100 = 92.2\%$

3. *Formulas of compounds*

Solution Y contains 30.0 g of $FeSO_4 . xH_2O$. In a titration 25.0 cm^3 of solution Y ($FeSO_4 . xH_2O$)reacted with 27.0 cm^3 of 0.02 mol/dm^3 $KMnO_4$. In the reaction 5 moles of Y reacts with one mole of $KMnO_4$. Calculate the concentration of Y in mol/dm^3 and hence find the value of x.

No. of moles of $KMnO_4$ used in the titration $= \frac{27.0}{1000} \times 0.020$ mol = 0.00054 mol

No. of moles of $FeSO_4 . xH_2O$ that reacted with $KMnO_4$ in titration.
= 5 x 0.00054 = 0.0027 mol

Conc. of $FeSO_4 . xH_2O$ $= \frac{0.0027}{25} \times 1000$

= 0.108 mol/dm3

Hence 0.108 mol of $FeSO_4.xH_2O$ = 30 g.
Therefore 1 mole of $FeSO_4 . xH_2O$ will be equal to 30/ 0.108 = 278 g.

Mr. Of $FeSO_4$ = 152
Mr. Of $FeSO_4.xH_2O$ = 278
Therefore $x.H_2O$ = 278- 152 = 126
Mr. of H_2O = 18
Mr. of xH_2O = 126
Therefore x = 126 / 18 = 7

Therefore formula will be $FeSO_4.7H_2O$

4. *Numbers of Reacting Moles in an equation*.

$xH_2O_2 + yKMnO_4 + \text{acid} \text{ ------------------}\rightarrow \text{Product}$

In a titration, 25.0 cm^3 of 0.0400mol/dm^3 H_2O_2 reacted with 20.0 cm^3 of 0.0200 mol/dm^3 $KMnO_4$. Find the value of x and y in the outline equation above.

No. of moles of H_2O_2 used in the titration $= \frac{25.0}{1000} \times 0.0400$ mol = 0.001 mol.

No. of moles of $KMnO_4$ used in the titration $= \frac{20.0}{1000} \times 0.0200$ mol = 0.0004 mol.

So we can say that 0.001 moles of H_2O_2 reacts with 0.004 moles of $KMnO_4$.

So 1 mole of $KMnO_4$ would react with $\frac{0.001}{0.0004} = 2.5$ mole

Hence the ratio of x: y is 1: 2.5 = 2: 5.

So x = 2 and y = 5.

Precipitation Reactions and Solubility Rules

To predict whether a precipitation reaction will occur upon mixing aqueous solutions, you must know the solubility of each of the potential products. A substance that has a low solubility in water will likely form a precipitate in aqueous solution. A substance with a high solubility in water will not precipitate in solution.

The following solubility guidelines will be helpful in predicting precipitates:

1. A compound containing one of the following cations is probably soluble:
   Group 1A cation: Li^+, Na^+, K^+, Rb^+, Cs^+ Ammonium ion: NH_4^+
2. A compound that contains one of the following anions is probably soluble:
   Halide: Cl^-, Br^-, I^-
   Except Ag^+, Hg^{2+}, Pb^{2+} compounds
   Nitrate (NO_3^-), perchlorate (ClO_4^-), acetate (CH_3COO^-), sulfate (SO_4^{2-})
   Except Ba^{2+} , Hg^{2+}, Pb^{2+} sulfates
3. Most compounds that contain the following anions are insoluble unless they contain a Group 1A cation, ammonium ion:

Hydroxide (OH^-), oxide (O^{2-}), carbonate (CO_3^{2-}), phosphate (PO_4^{3-}), chromate(CrO_4^{2-}), sulfide(S^{2-})

To predict the outcome when combining aqueous solutions of ionic compounds:

1. Write the complete molecular equation.
2. Determine whether the products will be soluble or insoluble by consulting the solublity guidelines.
3. Write the complete ionic equation, separating the soluble products into their component ions.
4. Cancel and remove the spectator ions. The resulting net ionic equation must show the formation of an insoluble solid product in a precipitation reaction.

DONE