EXPERIMENT

Aim: To prepare 100 ml M/10 solution of oxalic acid solution and titrate it against the given potassium permanganate solution. Also find the molarity and strength of the potassium permanganate solution.

Theory:

Potassium permanganate is a strong oxidising agent and in the presence of sulphuric acid it acts as a powerful oxidising agent. In acidic medium the oxidising ability of $KMnO_4$ is represented by the following equation.

In acidic solution,

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

Solution containing MnO_4^- ions are purple in colour and the solution containing Mn^{2+} ions are colourless and hence permanganate solution is decolourised when added to a solution of a reducing agent. The moment there is an excess of potassium permanganate present the solution becomes purple. Thus, $KMnO_4$ serves as self-indicator in acidic solution.

Potassium permanganate is standardized against pure oxalic acid. It involves a redox reaction. Oxalic acid is oxidised to carbon dioxide by $KMnO_4$, which itself gets reduced to $MnSO_4$.

Chemical reaction:

Reduction Half reaction:-

$$2KMnO_4 + 3H_2SO_4 \rightarrow K_2SO_4 + 2MnSO_4 + 3H_2O + 5[O]$$

Oxidation Half reaction:- $5(COOH)_2 + 5[O] \rightarrow 5H_2O + 10CO_2\uparrow$

Overall reaction:-

$$2KMnO_4 + 3H_2SO_4 + 5(COOH)_2 \rightarrow K_2SO_4 + 2MnSO_4 + 8H_2O + 10CO_2\uparrow$$

The ionic equation involved in the process is given below:

Reduction Half reaction:- $MnO_4^- + 8H^+ + 5e^- \rightarrow Mn^{2+} + 4H_2O$

Oxidation Half reaction:- $C_2O_4^{2-} \rightarrow 2CO_2 + 2e^-$

Overall Ionic reaction:-

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

This titration cannot be carried out in the presence of acids like nitric acid or hydrochloric acid because it is an oxidising agent. So hydrochloric acid chemically reacts with $KMnO_4$ solution forming chlorine which is also an oxidising agent.

Indicator: Here $KMnO_4$ acts as a self - indicator and this titration is called permanganate titration.

Catalyst: Mn^{2+} (auto-catalysis)

Materials Required:

Oxalic acid

Potassium permanganate solution
Sulphuric acid
Chemical balance
Burette
Burette stand
Pipette
Conical flask
Funnel
Measuring flask
Weighing bottle
White tile
Burnet
Wire gauze

Apparatus Setup:

In burette – $KMnO_4$ solution
In Conical flask – 10ml of oxalic acid + half a test tube Sulphuric acid

End Point – Appearance of permanent pale pink colour.

Procedure:

(a) Preparation of 0.1 M standard solution of oxalic acid:

- **The quantity of oxalic acid required for the 100 ml of the solution having a normality of 0.1 M can be calculated as follows.**
- **Molar mass of oxalic acid = 126 g**
- **M= wt x 1000 /molar mass x Volume (in ml)**
- **0.1 = wt x1000/ 126 x 100**
- **Wt = ?**
- **Weigh an empty watch glass using a chemical balance.**
- **Weigh ? g of oxalic acid accurately in the watch glass.**
- **With the help of a funnel transfer the oxalic acid into the measuring flask.**
- **Now wash the funnel with distilled water without removing the funnel from the flask.**
- **Make the solution up to the marked point with distilled water and make sure the oxalic acid is fully dissolved.**
- **This solution is 0.1 M standard solution of oxalic acid.**

(b) Titration of potassium permanganate solution against standard oxalic acid solution:

- **Rinse the burette with the potassium permanganate solution and fill the burette with potassium permanganate solution.**
- **Fix the burette in the burette stand and place the white tile below the burette in order to find the end point correctly.**

- **Pipette out 10ml of 0.1 M standard oxalic acid solution in a conical flask.**
- **Add a test tube full of sulphuric acid in order to prevent oxidation of manganese to form manganese dioxide.**
- **Heat the mixture up to 60°C before titrating with potassium permanganate.**
- **Note down the initial reading in the burette before starting the titration.**
- **The hot solution is titrated against potassium permanganate solution and simultaneously swirl the solution in the flask gently.**
- **Initially the purple colour of $KMnO_4$ is discharged with oxalic acid. The appearance of permanent pink colour reveals the end point.**
- **Repeat the titration until concordant values are obtained.**
- **Note down the upper meniscus on the burette readings. Record the reading in the observation table given below in order to calculate the molarity of $KMnO_4$ given.**

Observations:

S. No.	Volume of FAS	Volume of $KMnO_4$		Volume of $KMnO_4$ used
		Initial reading	Final reading	
1	10 ml			
2	10 ml			
3	10 ml			

Titre value:?

Calculations:

From the balanced chemical equation, it is clear that 2 moles of $KMnO_4$ reacts with 5 moles of oxalic acid.

According to the molarity equation,

$$\frac{\text{Molarity of } KMnO_4 \times \text{Volume of } KMnO_4}{\text{Molarity of oxalic acid} \times \text{Volume of oxalic acid}} = \frac{\text{No. of moles of } KMnO_4}{\text{No. of moles of oxalic acid}} = \frac{2}{5}$$

$$\text{Therefore, Molarity of } KMnO_4 = \frac{\text{Molarity of oxalic acid} \times \text{Volume of oxalic acid} \times 2}{\text{Volume of } KMnO_4 \times 5}$$

= 0.1 x 10 x 2/ ? x 5

= ??

Strength of potassium permanganate in g/L = Molarity x molar mass

= ?? x 158

= ???

Results:

1. **Molarity of $KMnO_4$ is _____M**
2. **The Strength of $KMnO_4$ is _____g/L**

Precautions:

- **Clean all the apparatus with distilled water before starting the experiment and then rise with the solution to be taken in them.**
- **Rinse the pipette and burette before use.**
- **Potassium permanganate is dark in colour, so always read the upper meniscus.**
- **Use dilute sulpuric acid for acidifying the potassium permanganate.**
- **Take accurate readings once it reaches the end point and don't go with average readings.**
- **Do not use rubber cork burette as it is can be attacked by $KMnO_4$**
- **The strength of the unknown solution should be taken up to two decimal places only.**