COURSE TITLE : **PHYSICAL CHEMISTRY PRACTICALS**

COURSE CODE : **15U6PRCHE06**

SESSION 2 : Potentiometric Titrations

POTENTIOMETRIC TITRATIONS

- **Potentiometric titration is quantitative analysis used to estimate the concentration of a given analyte solution.**
- **Potentiometric titration involves the continuous addition of a titrant to an analyte solution and measuring the EMF of the cell.**
- **The equivalence point of the titration reaction will be indicated by the sudden change in the potential.**
- **The potential of an electrode varies with ion concentration in accordance with Nernst equation:**

$$E = E^0 + \frac{2.303RT}{nF}\log C$$

E is the electrode potential at the concentration C

$KMnO_4$ *Vs* Fe^{2+} soln

AIM:

To determine the strength of a given ferrous ion solution by potentiometric titration with the given $KMnO_4$ solution.

PRINCIPLE

The redox reaction for the $KMnO_4$ *vs* Fe^{2+} titration:

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

A platinum electrode - ferrous salt solution - calomel electrode

Calomel || KCl || Fe^{3+}, Fe^{2+} | Pt

Indicator electrode – Platinum electrode

Reference electrode – Calomel electrode

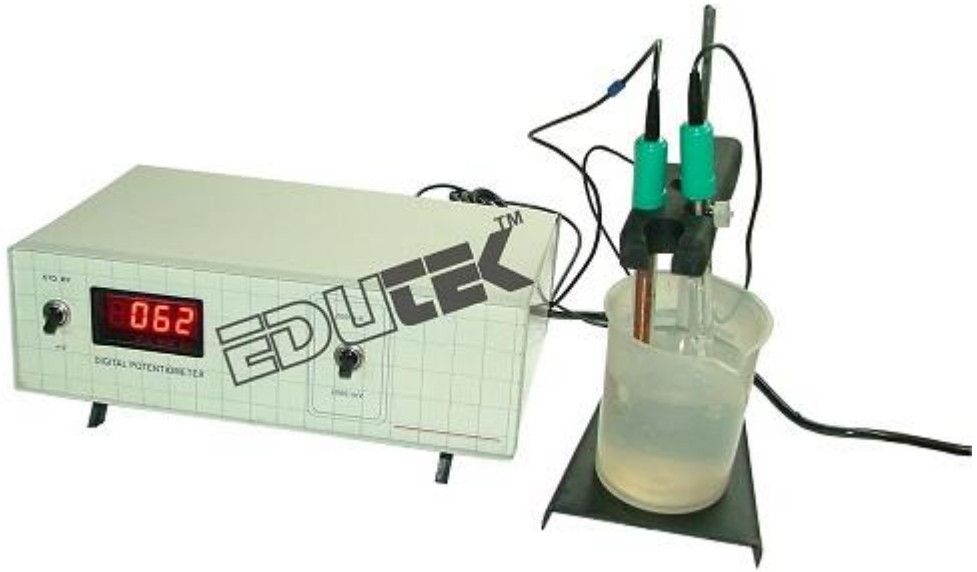
The potential acquired by the indicator electrode:

$$E = E^0 + \frac{0.0591}{n} \ln \frac{\left[Fe^{3+} \right]}{\left[Fe^{2+} \right]}$$

- The EMF of the cell depends on the ratio $[Fe^{3+}]/[Fe^{2+}]$.
- In the beginning, the potential depends on the concentration of Fe^{2+} state.
- After the equivalence point, the potential depends on the concentration of Fe^{3+} state.
- The potential shows a sharp increase at the equivalence point.

MATERIALS REQUIRED:

1. Potentiometer, Platinum and Calomel electrodes



2. 0.1 N $KMnO_4$
3. Unknown Ferrous ion solution
4. 4N Sulphuric acid
5. Burette
6. Pipette – 10 mL
7. Standard Flasks – 100 mL (2 Nos)

PROCEDURE:

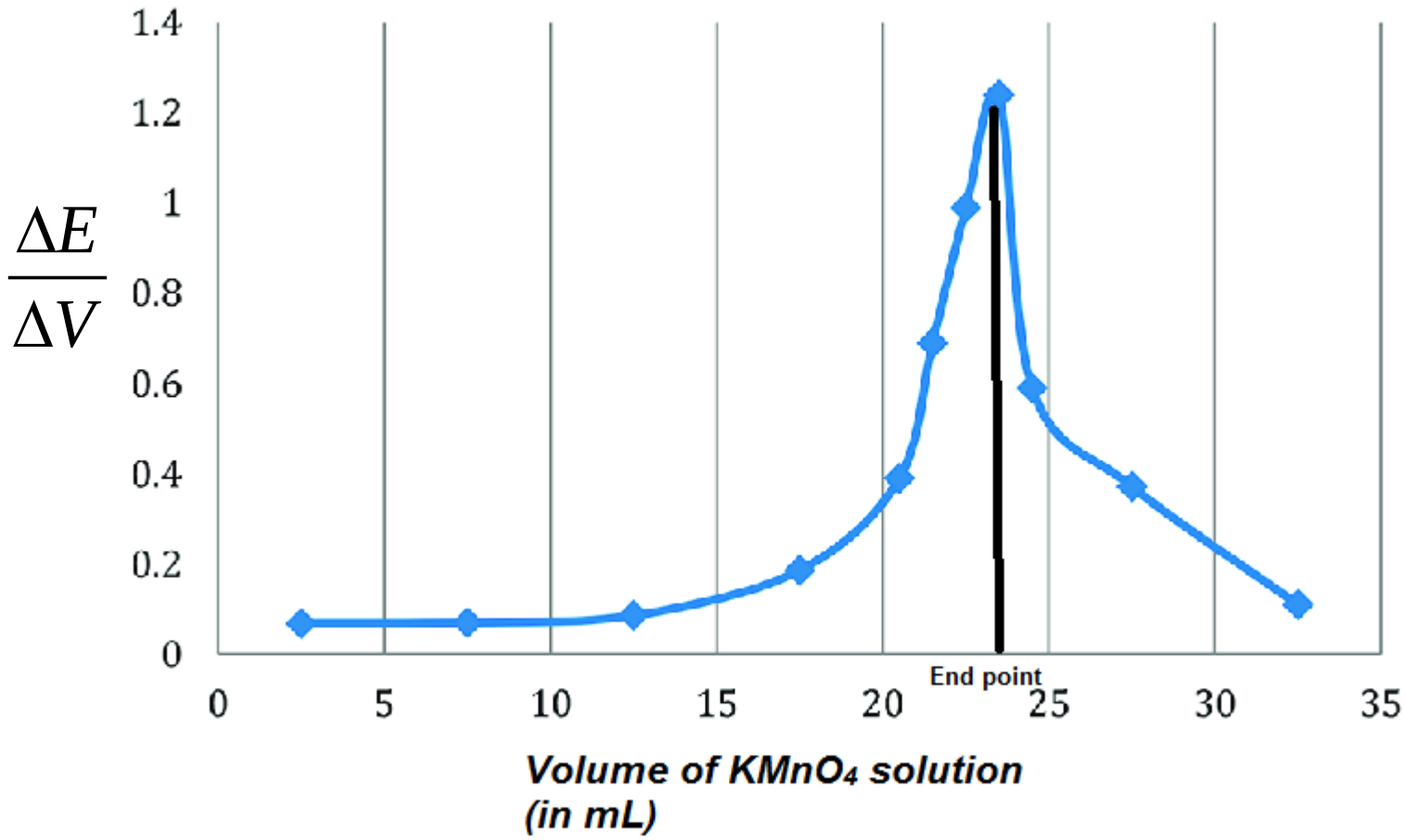
- ❑ The unknown Ferrous ion solution is made upto 100 mL
- ❑ The titrant (0.1 N $KMnO_4$ solution) is filled in the burette.
- ❑ Pipette out 20 mL of unknown Ferrous ion solution into a 250 mL beaker.
- ❑ Add 40 mL of 6N sulphuric acid to the ferrous ions solution.
- ❑ Dip the platinum and calomel electrodes in the solution and connected to the potentiometer.
- ❑ Record the EMF of the solution.
- ❑ Add 1 mL of 0.1 N $KMnO_4$ solution from the burette, stir the solution and note the EMF.
- ❑ The addition of titrant is repeated in amounts of 1 mL and the EMF is measured each time.

$\frac{\Delta E}{\Delta V}$

Volume of $KMnO_4$ (V) added (in mL)	EMF (E) (in Volts)	ΔE	$\frac{\Delta E}{\Delta V}$
0	330	-	-
1	350	20	20
2	360	10	10
3	380	20	20
4			
5			
6			
7			
8			
9			

GRAPH

$\frac{\Delta E}{\Delta V}$ (*y-axis*) and Volume of titrant added (*x-axis*).



CALCULATION

Volume of unknown Ferrous ion soln (V_1) = 20 mL

Strength of $KMnO_4$ soln (N_2) = 0.1 N

Volume of unknown Ferrous ion soln (V_2) = -------- mL (determined from graph)

Normality of Fe^{2+} solution (N_1) = $\frac{N_2 \times V_2}{V_1} = \frac{0.1 \times V_2}{20}$

RESULT:

The strength of the given Fe^{2+} solution is …….N.