# INDIAN INSTITUTE OF TECHNOLOGY, BOMBAY

| COURSE | ROLL NO. | NAME | |
|---|---|---|---|
| CH117L | 23B0912 | Aditya Sanapala | |
| **ASSIGNMENT NO.** | | **DUE DATE** | **SUB. DATE** |
| 1 | | | |

9/10

Experiment-1: Electrochemical cells

## Objectives

① To measure the standard electrode potential of the $Zn^{2+}/Zn$ Couple.

② To determine the concentration of $Fe^{2+}$ by a potentiometric titration.

## Introduction

→ Electrochemical techniques can be broadly classified into two categories which are

1. Equilibrium techniques.
2. Dynamic techniques.

→ Both of these techniques are based on the Nernst equation

Nernst eqn: $E_{cell} = E^{0}_{cell} + \frac{2.303RT}{nF} \log \frac{a_{ox}}{a_{red}}$

$E_{cell}$ = electrode potential

$E^{0}_{cell}$ = standard electrode potential

$a_{ox}, a_{red}$ = activities of reagents

R = Gas Constant - 8.314 J/K mol

F = Faraday's Const = 96500 C/mol

## Calculation:

Initial reading $E_{cell} = 0.969\,V$.

① Determination of $E^0_{Zn^{2+}/Zn}$ :

Cell = $Zn \,|\, Zn^{2+}(0.1M) \,||\, KCl\,(sat.) \,|\, AgCl \,|\, Ag$

Cell potential $E^0_{cell} = E^0_R - E^0_L = 0.222 - E^0_{Zn^{2+}/Zn^{2+}}$

By Nernst equation,

$$E_{Zn/Zn^{2+}} = E^0_{Zn/Zn^{2+}} - \frac{RT}{2F}\ln\left(\frac{a_{Zn^{2+}}}{a_{Zn}}\right)$$

$\because a_{Zn} = 1$ & $a_{Zn^{2+}} = \gamma\,[Zn^{2+}]$ $\quad \gamma$ = activity coeff = 0.15

$= 0.15 \times 0.1 = 0.015$

$$\Rightarrow E_{Zn/Zn^{2+}} = E^0_{Zn/Zn^{2+}} - \frac{0.0592}{2}\log(0.015)$$

$$= E^0_{Zn/Zn^{2+}} - 0.0296$$

Now, $E_{cell} = \left(E^0_{Zn/Zn^{2+}} - 0.0296\right) + 0.222\,V$

$$\Rightarrow E^0_{Zn/Zn^{2+}} = E_{cell} + 0.0296 - 0.222$$

$$= 0.969 + 0.0296 - 0.222$$

$= 0.7766\,V$ $\quad \Rightarrow E^0_{Zn^{2+}/Zn} = -E^0_{Zn/Zn^{2+}} = -0.7766\,V$

Note: $E_{cell}$ was measured to be 0.969V for 0.1 M $ZnSO_4$ & the same cell.

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2 Observations

Sample no. 9.

| Vol. of $K_2Cr_2O_7$ (V) (mL) | EMF (V) | ΔE (volts) | ΔV (mL) | $\frac{\Delta E}{\Delta V}$ (V/mL) | $V_{avg}$ (mL) |
|---|---|---|---|---|---|
| 0 | 0.375 | ~~0.03~~ | | | |
| 0.5 | 0.405 | 0.03 | 0.5 | 0.06 | 0.25 |
| 1 | 0.427 | 0.022 | 0.5 | 0.044 | 0.75 |
| 1.5 | 0.444 | 0.017 | 0.5 | 0.034 | 1.25 |
| 2 | 0.462 | 0.018 | 0.5 | 0.036 | 1.75 |
| 2.5 | 0.484 | 0.022 | 0.5 | 0.044 | 2.25 |
| 3 | 0.600 | 0.116 | 0.5 | 0.232 | 2.75 |
| 3.2 | 0.673 | 0.073 | 0.2 | 0.365 | 3.1 |
| 3.4 | 0.681 | 0.008 | 0.2 | 0.04 | 3.3 |
| 3.6 | 0.685 | 0.004 | 0.2 | 0.02 | 3.5 |
| 3.8 | 0.689 | 0.004 | 0.2 | 0.02 | 3.7 |
| 4 | 0.692 | 0.003 | 0.2 | 0.015 | 3.9 |
| 4.2 | 0.694 | 0.002 | 0.2 | 0.01 | 4.1 |

Data set for the titration.

## Calculation

Vol. of $Fe^{2+}$ soln taken = 25 mL

② From the graph, the inflexion point occurs at

$V_{avg}$ = 3.1 mL.

for the redox reaction

$$K_2Cr_2O_7 + 6FeSO_4 + 7H_2SO_4 \longrightarrow K_2SO_4 + 3Fe_2(SO_4)_3 + Cr_2(SO_4)_3 + 7H_2O$$

$n_f = 6$

$N_1V_1 = N_2V_2$ (Law of equivalents)

$\Rightarrow \cancel{(N_{Fe^{2+}}) \times (25)}$ $\Rightarrow (N_{Fe^{2+}})(25\,mL) = (0.5) \times (3.1\,mL)$

$$\Rightarrow N_{Fe^{2+}} = \frac{0.5 \times 3.1}{25} = 0.062N.$$

## Conclusions

① The standard electrode potential of the $Zn^{2+}|Zn$ cell is

$\boxed{E^{o}_{Zn^{2+}/Zn} = -0.7766 \text{ Volts.}}$

② The normality of the $Fe^{2+}$ solution given was

$\boxed{N = 0.062}$

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SCALE :-