

INDIAN INSTITUTE OF TECHNOLOGY, BOMBAY

COURSE CH171L	ROLL NO. 23B0912	NAME Avtiya Sanapala
ASSIGNMENT NO. 1	DUE DATE	SUB. DATE 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

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

E_{cell} = electrode potential

E_{cell}° = standard electrode potential

R = Gas Constant - 8.314 J/K mol

F = Faraday's Const = 96500 C/mol

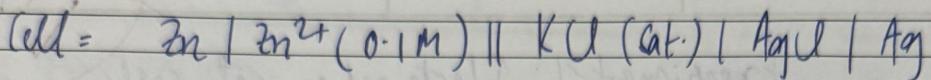
$a_{\text{ox}}, a_{\text{red}}$ = activities

of reagents

Calculation:

Initial reading $E_{\text{cell}} = 0.969 \text{ V}$

① Determination of $E_{\text{Zn}^{2+}/\text{Zn}}^{\circ}$:



$$\text{Cell potential } E_{\text{cell}} = E_{\text{Zn}}^{\circ} - E_{\text{I}}^{\circ} = 0.222 - E_{\text{Zn}^{2+}/\text{Zn}}^{\circ}$$

By Nernst equation,

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

$$\therefore a_{\text{Zn}} = 1 \quad a_{\text{Zn}^{2+}} = \gamma (\text{M}^{2+}) \quad \gamma = \text{activity coeff} = 0.15 \\ = 0.15 \times 0.1 = 0.015$$

$$\Rightarrow E_{\text{Zn}/\text{Zn}^{2+}} = E_{\text{Zn}}^{\circ}/\text{Zn}^{2+} - \frac{0.0592}{2} \log(0.015) \\ = E_{\text{Zn}}^{\circ}/\text{Zn}^{2+} - 0.0296$$

$$\text{Now, } E_{\text{cell}} = (E_{\text{Zn}/\text{Zn}^{2+}}^{\circ} - 0.0296) + 0.222 \text{ V}$$

$$\Rightarrow E_{\text{Zn}/\text{Zn}^{2+}}^{\circ} = E_{\text{cell}} + 0.0296 - 0.222 \\ = 0.969 + 0.0296 - 0.222$$

$$= 0.7766 \text{ V} \quad \Rightarrow E_{\text{Zn}^{2+}/\text{Zn}}^{\circ} = -E_{\text{Zn}/\text{Zn}^{2+}}^{\circ} = -0.7766 \text{ V}$$

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

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② Observations

Sample no. 9.

Vol. of $\text{K}_2\text{Cr}_2\text{O}_7$ (V) (mL)	EMF (V)	ΔE (mV)	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	0.500	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.

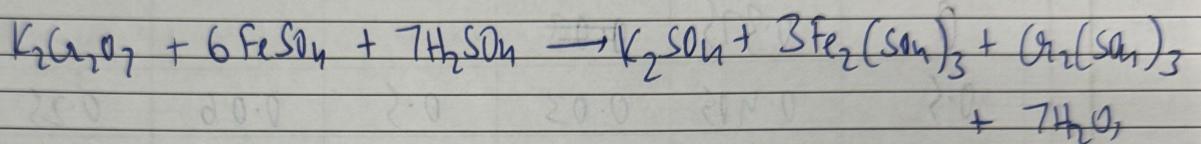
Calculation

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

- ② From the graph, the inflection point occurs at

$$V_{\text{ang}} = 0.31 \text{ mL}$$

for the redox reaction



$$n_f = 6$$

$$N_1 V_1 = N_2 V_2 \quad (\text{Law of equivalents})$$

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

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

Conclusions

- ① The standard electrode potential of the Zn^{2+}/Zn cell is

$$E_{\text{Zn}^{2+}/\text{Zn}}^{\circ} = -0.7766 \text{ Volts.}$$

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

$$N = 0.062$$

SCALE :-

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