Experiment - 10

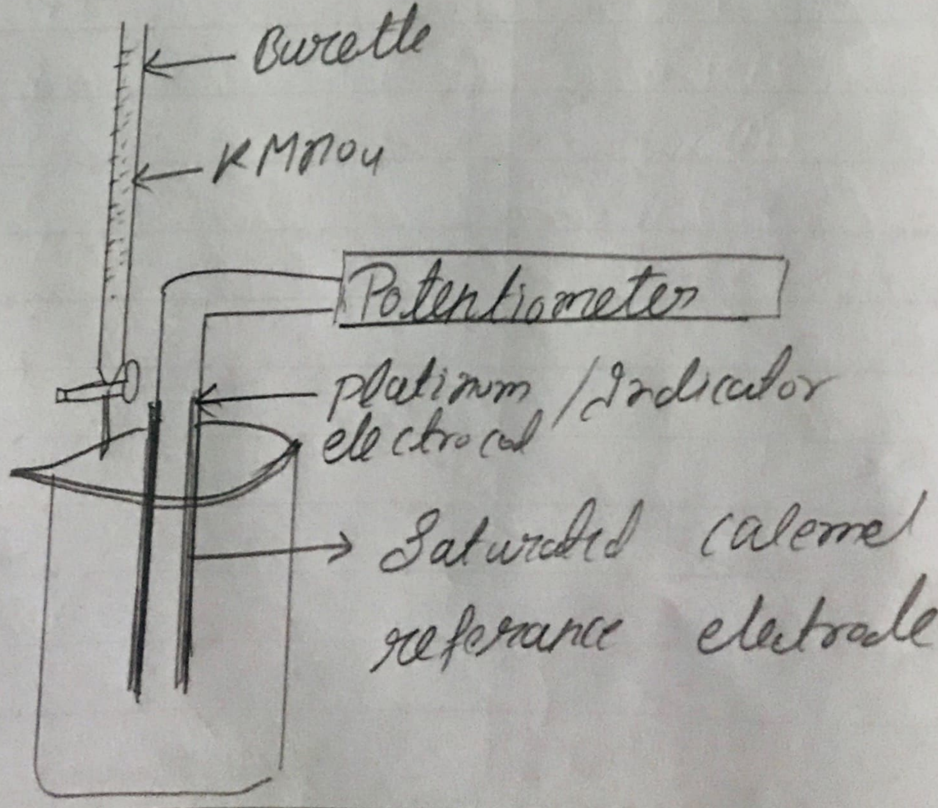
Experiment : To titrate potentionametrically ferrous ammonium sulphate solution against potassium permanganate and to determine the standard electrode potential of ferrous - ferric sulphate.

Apparatus: pipette, burette, beaker, funnel stand, clamp, potentiometer, calomel electrode (or Ag / AgCl electrode and platinum electrode.

Chemical: Mohr's salt solution, $KMnO_4$ and H_2SO_4

Equation : $Fe^{3+} + e^- \rightarrow Fe^{2+}$

$$5Fe^{2+}(aq) + MnO_4^-(aq) + 8H^+ \rightarrow 5Fe^{3+}(aq) + Mn^{2+}(aq) + 4H_2O(l)$$



Experiment - 10

Experiment : To titrate potentiometrically ferrous ammonium sulphate solution against potassium permanganate and to determine the standard electrode potential of ferrous - ferric system.

Apparatus :- Pipette, burette, beaker, funnel, stand, clamp, potentiometer, calomel electrode (or age electrode and platinum electrode.

Chemicals :- Mohr's salt solution (ferrous ammonium sulphate - $FeSO_4 (NH_4)_2SO_4 6H_2O$) potassium permanganate ($KMnO_4$) and sulpheric acid (H_2SO_4)

Theory: An electrochemical cell is a device which establishes measurable electrical potential differences and in which flow of electrical current is accompanied by an overall chemical. A reversible cell is that in which the overall chemical reaction can be reversed on the presence of an opposing external electromotive force of magnitude greater than that is cell itself. An electromagnetic

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Obervations :

(iv) Volume of $KMnO_4$ vs e.m.f of the Soluti

In beaker : FAS Solution (0.01N) + H_2SO_4 (1N)

S.No	Volume of $KMnO_4$ added (ml) (DPB applied)	EMF (volts)
1	0	0.300
2	0.2	0.341
3	0.7	0.361
4	1.2	0.372
5	1.7	0.380
6	2.2	0.388
7	2.7	0.394
8	3.2	0.400
9	3.7	0.405
10	4.2	0.410
11	4.7	0.415
12	5.2	0.419
13	5.7	0.423
14	6.2	0.427
15	6.7	0.432
16	7.2	0.436
17	7.7	0.441
18	8.2	0.446
19	8.7	0.451
20	9.2	0.456
21	9.7	0.464

cell consists of two electrodes or half cells, whose electrolytic solutions are either directly in contact with each other or connected through an interving electrolytic solution. The net chemical change takes place at the individual eletrodes, one of which is oxidation and the other reduction. At the reversible electrodes, the oxidized and reduced status of a system exist in equilibrium in the solution, where an invert metal electrode (like it) is dipped into it eg, Fe^{3+}/Fe^{2+} in which the reaction is.

$$Fe^{3+} + e \longrightarrow Fe^{2+}$$

The experimental cell:

$$|| \; Fe^{3+}/Fe^{2+}/pt \; \text{----} \; Hg, Hg_2Cl_2(s)/KCl \text{ (saturated)} \quad (1)$$

The emf of the cell:

$$E - E^{\circ}(Fe^{2+}/Fe^{2+}) + \frac{2.303RT}{nF} \log \frac{[Fe^{3+}]}{[Fe^{2+}]} - E$$

chemical reaction during potentiometer titration.

$$5Fe^{2+}(aq) + MnO_4(aq) + 8H^{+} \longrightarrow 5Fe^{3+}(aq) + Mn^{2+}(aq) + 4H_2O(l)$$

When potassium permanganate solution is added to Mohr's salt solution, the concentration

S no	Volume of $KMnO_4$ added (ml)	EMF (valume)
22	10.2	0.475
23	10.7	0.483
24	11.2	0.494
25	11.7	0.546
26	12.2	1.004
27	12.7	1.05
28	13.2	1.063
29	13.7	1.065
30	14.7	1.067
31	15.7	1.071
	16.7	

General Calculation:

Emf at half equivalence point

is (A) = 0.423

$E(Fe^{3+}/Fe^{2+}) = E$ (half equivalent) + E (calom

$E(Fe^{3+}/Fe^{2+}) = A + 0.842 = 0.423 + 0.242$

$= 0.665$ ✓

$E(Fe^{3+}/Fe^{2+}) = 0.665$ ✓

of Fe^{2+} ions decreases and that the Fe^{3+} ions increases, and a result the emf of the cell increase slowly. Near to the equivalence point, an inflection in seen due to fall in concentration of Fe^{2+} ions ultimately to 0, resulting in sudden rise in emf of the cell.

At half equivalence, eq 2 becomes
E (half - equivalence) $= E^\circ(Fe^{3+}/Fe^{2+}) - E$ (calomel)

$= E^\circ[Fe^{3+}/Fe^{2+}) - 0.242$ (at 298k)

$\therefore E^\circ(Fe^{3+}/Fe^{2+}) = E$ [half equivalence) $+ 0.242$ ----(3)

producer :- Take 50 ml of 0.1N FAS in beaker and add 5ml of 1N H_2SO_4. Dip the platinum and standard calomel electroted in the solution.

2. Connect the indicator and reference electrodes to black and red terminals of potentiom respectively.

3. Rinse and fill the burette with $KMnO_4$ (0.2N)

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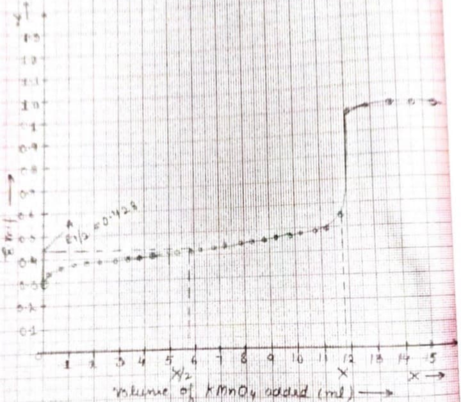

y
1.3
1.2
1.1
1.0
0.9
0.8
0.7
0.6
0.5
0.4
0.3
0.2
0.1
E →
$E_{1/2} = 0.428$
1 2 3 4 5 6 7 8 9 10 11 12 13 14 15
X/2
X
x →
volume of $KMnO_4$ added (ml) →

4. Note the initial emf of cell and start addition of titrant ($KMnO_4$) is portions of 1ml each. Near the equivalence point, pl. decrease the volume of additional titrant to 0.5 ml and later on to 0.2 ml and note the reading after each titration.

5. Continue to take 10-12 readings more, after a change in emf is notice.

6. plot emf in volts against the volume of $KMnO_4$ Solution added (ml) and note from the equivalence and potential at half the equivalence pt

General Calculation:

The emf at equivolence pt (A) observed from graph.

E (half equivalence) $= E^{\circ}[Fe^{3+}/Fe^{2+}] - E$ (calomel)

$E^{\circ}[Fe^{3+}/Fe^{2+}] = E$ (half equivalence $+ E$ (calomel)

Result

The standard half-cell potential of Fe^{2+}/Fe^{3+} couple is 0.665 V

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