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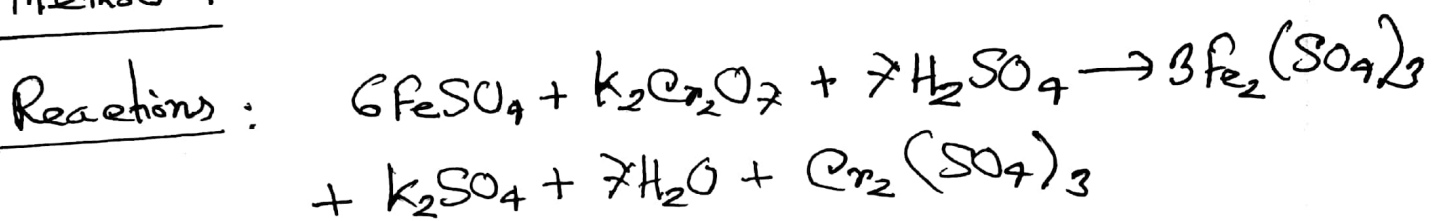
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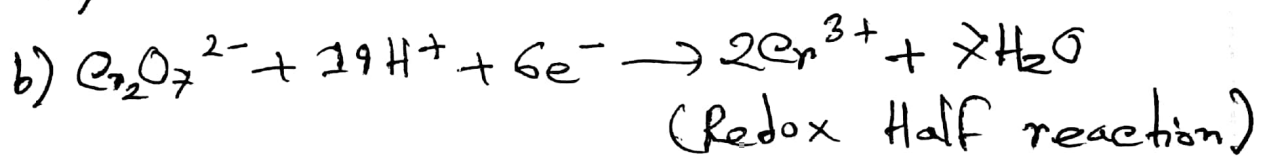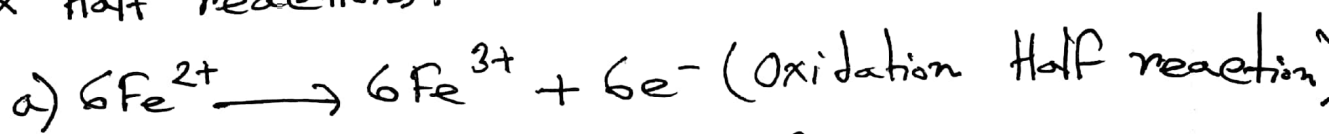Sec : R

Experiment - 6 : Determination of ferrous Ion (Fe^{2+}) in a supplied of Iron salt by standard Potassium Dichromate ($K_2Cr_2O_7$) Solution.

Experimental data :

Method : Oxidation - Reduction titration

Reactions :

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

Redox Half reactions :

a) $6Fe^{2+} \longrightarrow 6Fe^{3+} + 6e^-$ (Oxidation Half reaction)

b) $Cr_2O_7^{2-} + 19H^+ + 6e^- \longrightarrow 2Cr^{3+} + 7H_2O$

(Redox Half reaction)

Weight taken (in gm) = 0.66 g

The strenght of $K_2Cr_2O_7$ Solution

$$= \frac{\text{weight taken (in gm)} \times 0.1}{0.49} \ (N)$$

$$= \frac{0.66 \times 0.1}{0.49} \ (N)$$

$$= 0.1346 \ (N)$$

$$= 0.13 \ (N)$$

Table : Determination of the amount of iron in Mohr's salt solution using standard $K_2Cr_2O_7$ Solution.

No of reading	Vol. of Mohr's Salt Solution	Vol of $K_2Cr_2O_7$ (burette reading) (in ML)			Mean (in mL) (v)
		Initial	Final	Difference	
1	10	0.00	5.00	5.00	
2	10	5.00	9.90	4.90	5.03
3	10	9.90	15.00	5.1	
4	10	15.00	20.10	5.1	

Calculations : 1 mL 1N $K_2Cr_2O_7$ = 0.05584 gm of Fe^{2+}

Amount of iron in 10mL of iron salt solution

$$0.05584 \times V \times S \ (gm)$$

$$= 0.05584 \times 5.03 \times 0.13$$

$$= 0.0365 \ gm$$
$$= 0.04 \ gm$$

Amount of iron in 500 mL of iron salt solution

$$0.05584 \times V \times S \times 50 \ gm$$

$$= 0.05584 \times 5.03 \times 0.13 \times 50$$

$$= 1.8256 \ gm \qquad = 1.83 \ gm$$

Observe value of Fe^{2+} (in 500mL Solution)

$$= 1.83 \ gm$$

known value of Fe^{2+} (in 500ml solution)

$$= \frac{55.84 \times 8.55}{392.14} \ (gm)$$

$$= 1.2175 \ (gm)$$

$$= 1.22 \ gm.$$

Result : The amount of ferrous ions in 500mL of iron salt solution is 0.

Percentage of error :

$$\frac{known \ value - observed \ value}{known \ value} \times 100$$

$$= \frac{1.22 - 1.83}{1.22} \times 100$$

$$= -0.28 \, \%.$$