EXP6: Determination of Ferrous ion (Fe^{2+}) in a supplied solution of Iron salt by standard Potassium Dichromate ($K_2Cr_2O_7$) solution.

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Theory:

Method: Oxidation-reduction titration

Reaction:

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

Redox Half Reactions:

(a) $6Fe^{2+} \rightarrow 6Fe^{3+} + 6e$

(b) $Cr_2O_7^{2-} + 14H^+ + 6e \rightarrow 2Cr^{3+} + 7H_2O$

Indicator:

Diphenyl amine $(C_6H_5)_2NH$

Experimental Data:

The strength of $K_2Cr_2O_7$ solution $= \frac{\text{Weight taken (in gm)} \times 0.1}{0.49}$ (N)

$$= \frac{0.65 \times 0.1}{0.49} N$$
$$= 0.13 N$$

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 (in mL)	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	$\frac{5.00+4.90+5.10+5.10}{4}$
2	10	5.00	9.90	4.90	
3	10	9.90	15.00	5.10	
4	10	15.00	20.10	5.10	$= 5.025$

Calculations:

1 ml 1N $K_2Cr_2O_7 \equiv 0.05589$ gm of Fe^{2+}

Amount of iron in 10 ml of iron salt solution $= 0.05589 \times V \times S$ gm

$= 0.05589 \times 5.025 \times 0.13$ gm

$= 0.036$ gm

Amount of iron in 500 ml of iron salt solution $= 0.05584 \times V \times 5 \times 50$ gm

$= 0.05584 \times 5.025 \times 0.13 \times 50$ gm

$= 1.824$ gm

Observe value of Fe^{2+} (in 500 ml solution) $= 1.824$ gm

known value of Fe^{2+} (in 500 ml solution) $= \frac{55.84 \times 8.90}{392.14}$ gm

$= 1.267$ gm

Results:

The amount of ferrous ions in 500 ml of iron salt solution is 1.824 gm.

Percentage of error: $\frac{\text{known value} - \text{observed value}}{\text{known value}} \times 100$

$= \frac{1.267 - 1.824}{1.267} \times 100$

$= -0.4396 \times 100$

$= -43.96\%$