KMnO4 Reaction With Fe(NO2)2

https://gemini.google.com/app/d5e646fc40275c01

Response: To find the number of moles of $\mathrm{KMnO_4}$ required to react completely with 1 mole of $\mathrm{Fe(NO_2)_2}$ in an acidic medium, we can analyze the oxidation and reduction half-reactions using the mole/electron balance method.

Step 1: Identify the Oxidation Half-Reaction

In $\mathrm{Fe(NO_2)_2}$, both the iron ion ($\mathrm{Fe^{2+}}$) and the nitrite ions ($\mathrm{NO_2^-}$) undergo oxidation:

1. **Oxidation of $\mathrm{Fe^{2+}}$ to $\mathrm{Fe^{3+}}$:**

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

(1 electron released per $\mathrm{Fe^{2+}}$ ion)

2. **Oxidation of $\mathrm{NO_2^-}$ to $\mathrm{NO_3^-}$:**

$$\mathrm{NO_2^-} + \mathrm{H_2O} \rightarrow \mathrm{NO_3^-} + 2\mathrm{H^+} + 2e^-$$

(2 electrons released per $\mathrm{NO_2^-}$ ion)

Since 1 mole of $\mathrm{Fe(NO_2)_2}$ contains 1 mole of $\mathrm{Fe^{2+}}$ and 2 moles of $\mathrm{NO_2^-}$:

- Total electrons lost per mole of $\mathrm{Fe(NO_2)_2} = 1(1e^-) + 2(2e^-) = 5 \text{ electrons}$

Step 2: Identify the Reduction Half-Reaction

In acidic medium, permanganate ($\mathrm{MnO_4^-}$) is reduced to $\mathrm{Mn^{2+}}$:

$$\mathrm{MnO_4^-} + 8\mathrm{H^+} + 5e^- \rightarrow \mathrm{Mn^{2+}} + 4\mathrm{H_2O}$$

(5 electrons gained per $\mathrm{MnO_4^-}$ ion)

Step 3: Mole Ratio Calculation

By equating the total electrons lost in oxidation to the total electrons gained in reduction:

$$\text{Moles of } e^- \text{ released by 1 mole of } \mathrm{Fe(NO_2)_2} = 5 \text{ moles}$$

$$\text{Moles of } e^- \text{ consumed by 1 mole of } \mathrm{KMnO_4} = 5 \text{ moles}$$

$$\text{Moles of } \mathrm{KMnO_4} \text{ required} = \frac{\text{Moles of } e^- \text{ lost}}{\text{Electrons gained per } \mathrm{KMnO_4}} = \frac{5}{5} = 1$$

Correct Option

(3) 1