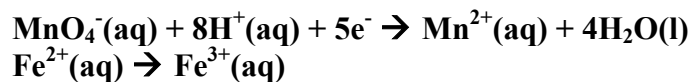


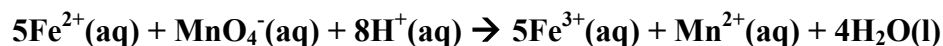
- A 20.0 mL sample of 0.121 M Fe^{2+} in an acid solution was used to titrate 23.5 mL of a KMnO_4 solution of unknown concentration. Write the balanced redox reaction that occurs in solution upon titration, and calculate the molarity of the KMnO_4 solution.

Marks
4

From the standard reduction potentials, the two relevant half cells are:



Giving the overall reaction:



The number of moles of Fe^{2+} used in the titration is:

$$\begin{aligned}\text{number of moles of Fe}^{2+} &= \text{concentration} \times \text{volume} \\ &= 0.121 \text{ mol L}^{-1} \times (0.0200 \text{ L}) = 2.42 \times 10^{-3} \text{ mol}\end{aligned}$$

From the balanced equation, the number of moles of $\text{MnO}_4^-(\text{aq})$ is therefore:

$$\text{number of moles of MnO}_4^- = 1/5 \times 2.42 \times 10^{-3} \text{ mol} = 4.84 \times 10^{-4} \text{ mol}$$

This amount is present in 23.5 mL so its concentration must be:

$$\begin{aligned}\text{concentration} &= \text{number of moles} / \text{volume} \\ &= 4.82 \times 10^{-4} \text{ mol} / 0.0235 \text{ L} = 0.0206 \text{ M}\end{aligned}$$

Answer: **0.0206 M**