- 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:**

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

**Giving the overall reaction:**

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

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

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

**From the balanced equation, the number of moles of $MnO_4^-(aq)$ 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:**

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

Answer: **0.0206 M**