[AQA A2 Specimen Paper 1 (set 2)]

A student dissolved 1980mg of iron tablets in an excess of dilute sulphuric acid. The solution was titrated with 0.0200 moldm^{-3} potassium manganate(VII) solution.

A 32.50 cm^3 volume of potassium manganate(VII) solution was required to reach the end point in the titration.

a) Calculate the percentage of iron in the sample of iron tablets. Give your answer to the appropriate number of significant figures.

① Write the redox reaction between Fe^{2+} and MnO_4^{2-} ions:

$$MnO_4^- + 5Fe^{2+} + 8H^+ \longrightarrow Mn^{2+} + 5Fe^{3+} + 4H_2O$$

$\Rightarrow$ molar ratio $= MnO_4^- : Fe^{2+}$

$\boxed{1 : 5}$

② Calculate the number of moles of MnO_4^- used:

$$\text{moles} = \frac{0.02 \times 32.50}{1000}$$

$$= 6.5 \times 10^{-4} \text{ moles}$$

$\text{moles} = \frac{\text{conc.} \times \text{vol.}}{1000}$

③ Use the molar ratio to find the moles of Fe^{2+} ions that react:

$$\Rightarrow 5 \times 6.5 \times 10^{-4} = 3.25 \times 10^{-3} \text{ moles}$$

④ Calculate the mass of iron in the tablet:

M_r of $Fe^{2+} = 55.8$

$\Rightarrow 3.25 \times 10^{-3} \times 55.8$

$= 0.18135 g$

mass = moles × Mr

5 Express this mass as a percentage of whole tablet:

mass of tablet = 1980mg

= 1.98g

$\Rightarrow \% = \dfrac{0.18135}{1.98} \times 100$

$= 9.159...$

$\Rightarrow \underline{9.16\%}$