

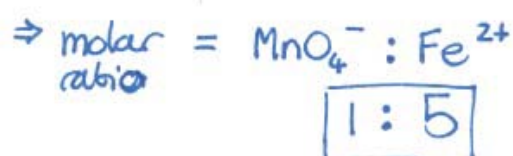
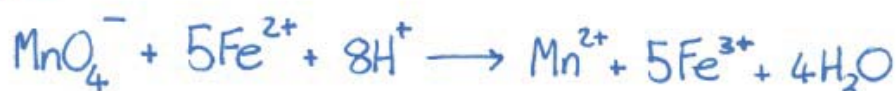
[AQA A2 Specimen Paper 1 (set 2)]

A student dissolved 1980mg of iron tablets in an excess of dilute sulphuric acid. The solution was titrated with $0.0200 \text{ mol dm}^{-3}$ 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:



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

$$\begin{aligned} \text{moles} &= \frac{0.02 \times 32.50}{1000} \\ &= 6.5 \times 10^{-4} \text{ moles} \end{aligned}$$

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:

$$\text{Mr of } \text{Fe}^{2+} = 55.8$$

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

$$= 0.18135 \text{ g}$$

mass = moles × Mr





⑤ Express this mass as a percentage of whole tablet:

$$\begin{aligned}\text{mass of tablet} &= 1980\text{mg} \\ &= 1.98\text{g}\end{aligned}$$

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

$$= 9.159\dots$$

$$\Rightarrow \underline{9.16\%}$$

