Estimation of Iron (II) in an iron tablet by using a standard solution of potassium manganate (VII)

THEORY:

To estimate the iron(II) content of an iron tablet, a small number of tablets are first dissolved in dilute sulfuric acid. This solution is then titrated against previously standardised potassium manganate(VII) solution. The reaction is represented by the equation:

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

METHOD

Preparation of tablets:

Find the mass of five iron tablets. (Note that if Feospan tablets are used, each capsule should be opened, and the **contents** weighed.)

Crush the weighed tablets in a mortar and pestle. Transfer all the ground material to a beaker where it is dissolved in about 100 cm^3 of dilute sulfuric acid.

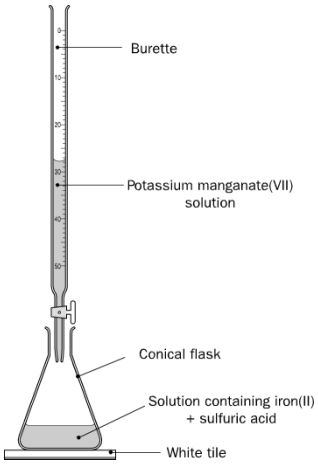
All of this solution (including washings) is transferred to a 250 cm^3 volumetric flask and the solution made up to the mark with deionised water. The volumetric flask should be stoppered and inverted several times. This is the solution containing iron(II) ions.

Titration:

Wash the pipette, burette and conical flask with deionised water. Rinse the burette with the potassium manganate(VII) solution and the pipette with the iron(II) solution.

Using a pipette filler, fill the pipette with the iron(II) solution and transfer the contents of the pipette to the conical flask. Acidify this solution by adding about 10 cm^3 of dilute sulfuric acid.

Using a funnel, fill the burette with potassium manganate(VII) solution, making sure that the part below the tap is filled before adjusting to zero. Because of the intense colour of $KMnO_4$ solution, **readings are taken from the top of the meniscus.**



With the conical flask standing on a white tile, add the solution from the burette to the flask. Swirl the flask continuously and occasionally wash down the walls of the flask with deionised water using a wash bottle.

The end-point of the titration is detected by 'the first persisting pink colour'. Note the burette reading.

Repeat the procedure two or three times, adding the potassium manganate(VII) dropwise approaching the endpoint. These accurate titres should agree to within 0.l cm^3.

Calculate the concentration of the iron(II) solution, and from this calculate the mass of iron in an iron tablet.

Specimen results

Mass of iron tablets = 1.81 g
Rough titre = 17.0 cm^3
Second titre = 16.7 cm^3
Third titre = 16.7 cm^3
Average of accurate titres = 16.7 cm^3
Volume of iron(II) solution used in each titration = 25.0 cm^3
Concentration of potassium manganate(VII) solution = 0.005 M

Calculations

$V_A \times M_A \times n_B = V_B \times M_B \times n_A$
$25.0 \times M_A \times 1 = 16.7 \times 0.005 \times 5$
$M_A = 16.7 \times 0.005 \times 5 / (25.0 \times 1)$
$= 0.0167$ M

Volume of Fe^{2+} solution in total = 250.0 cm^3
Moles of iron in this volume = 0.0167
= 0.004175
Mass of iron in this volume = 0.004175 x 56 g
= 0.2338 g
Percentage of iron in the tablets

$$= \frac{mass\ of\ iron \times 100}{mass\ of\ tablets}$$

$$= \frac{0.2338 \times 100}{1.81}$$

= 12.92%
Mass of iron in each tablet
= 0.2338 / 5 = 46.76 mg.

student questions

In this experiment why is dilute sulfuric acid used rather than deionised water to dissolve the iron tablets?
If deionised water were used, the Fe^{2+} in the tablets would be almost immediately oxidised to Fe^{3+}. The sulfuric acid prevents this occurring.

Why are burette readings taken from the top of the meniscus?
Because the very dark colour of the manganate(VII) solution makes the meniscus difficult to see.

How is the end-point of the titration detected?
When the first permanent pale pink colour forms in the solution in the conical flask.

Why is a rough titration carried out?
To determine the approximate end-point. This can then be used to get accurate results in the subsequent titrations.

Why is more than one titration carried out subsequently?
To reduce experimental error, by getting the mean of the accurate titres.

Prior to the titration, what steps are taken to minimise error?
All glassware is washed with deionised water. The burette and pipette respectively are rinsed with the solution they are to contain. The tap of the burette is opened briefly to fill the part of the burette below the tap.

If a brown precipitate appears during the titration, what does this indicate, and how can it be remedied?
Mn(IV) is formed, because of incomplete reduction of the Mn(VII). This should only happen if there is insufficient sulfuric acid in the conical flask. The remedy is to add more dilute sulfuric acid to the flask, or, preferably, to repeat the experiment with sufficient acid present in the flask.