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## **Access Chemistry Practical Assessment**

An investigation of iron tablets.

This is an experiment to find out the amount of ferrous sulphate in milligrammes in iron tablets of the type given to patients suffering from anaemia.

This can be evaluated by oxidising the iron with a known solution of potassium manganate 7.

### **Apparatus**

Burette, 10cl<sup>3</sup> pipette, Filter, Filter paper, Clamp stand, Graduated flask, Beakers, Conical Flask, Pestle and Mortar, Hotplate.

### **Method**

Two iron tablets are crushed using the pestle and mortar, while 50cm<sup>3</sup> of sulphuric acid is heated in a beaker on the hotplate. 25cm<sup>3</sup> is added to the mortar and is mixed and filtered into a graduated flask. The mortar is swilled out with the remaining 25cm<sup>3</sup> of warm acid, this is also poured into the filter. The contents of the graduated flask are made up to 100cm<sup>3</sup> with distilled water. The filter should have removed the chalk which comprised the mass of the tablet.

10cm<sup>3</sup> of the solution is transferred into the conical flask using the pipette. The burette is filled up to the 0 mark with potassium manganate 7 solution. The potassium manganate solution is added to the iron solution slowly until the iron solution has a slight pink coloration. The volume level on the burette is recorded. This is repeated until two consistent results have been recorded.

<u>Results</u>	1 <sup>st</sup> Titration	2 <sup>nd</sup> Titration	3 <sup>rd</sup> Titration	4 <sup>th</sup> Titration
1 <sup>st</sup> Reading	0	0	0	0
2 <sup>nd</sup> Reading	2.4	2.9	2.5	2.4

### Calculations

1. Calculate the number of moles of potassium manganate 7 in average burette volume:

$$\text{KMnO}_4 = 0.019 \text{ mol dm}^{-3}$$

$$\text{Number of moles of KMnO}_4 \text{ used} = 2.55 / 1000 * 0.019 \text{ mol dm}^{-3} = 48.45 \times 10^{-6} \text{ moles}$$

2. Knowing the reacting ratio of  $1 \text{ MnO}_4 : 5\text{Fe}^{2+}$  deduce the number of moles of  $\text{Fe}^{2+}$  in  $10\text{cm}^3$  of solution.

$$48.45 \times 10^{-6} * 5 = 242.25 \times 10^{-6} \text{ moles in } 10\text{cm}^3 \text{ of solution.}$$

3. Hence calculate the moles of  $\text{Fe}^{2+}$  (which is the same as moles of  $\text{FeSO}_4$ ) in the whole  $100\text{cm}^3$  of solution.

$$= 2.42 \times 10^{-3} \text{ moles}$$

4. Hence mass of  $\text{FeSO}_4$  in 2 tablets.

Iron 56 molecular mass. Sulphate 96 molecular mass.

RMM = 152

0.38g mass of  $\text{FeSO}_4$  in 2 tablets

5. Finally calculate final answer of milligrams of  $\text{FeSO}_4$  in one tablet.  
190mg in one tablet

### Evaluation:

There were some discrepancies due to dissolving and transferring technique, there was possibly some of the tablet left in the mortar which was not transferred completely into the flask.

They may have been some parallax error in the readings off of the burette.