

Appendix 2

3. How much iron is there in an iron tablet? – Student Sheet

In this practical you will have the opportunity to perform a quantitative analysis using the technique of titration. You are going to analyse an iron tablet to find out how much iron is actually present in it.

Titration involving potassium manganate(VII) may form part of your Practical Assessment.

Intended lesson outcomes

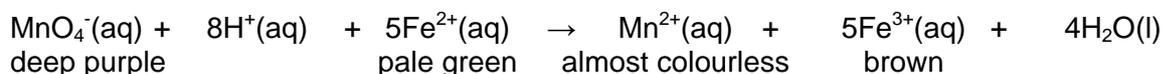
By the end of this practical you should be able to:

- perform a titration involving potassium manganate(VII);
- read a burette and use a pipette;
- use a volumetric flask;
- record your titration results appropriately in tables you have drawn yourself;
- use and understand an ionic equation;
- use the mole concept to perform calculations.

Background information

Iron performs a vital role in our bodies. It is present in red blood cells and forms part of the haemoglobin molecule, which combines with oxygen from the lungs. The oxygen is then transported all round the body. When young people are growing rapidly, the body may not have enough iron and this causes anaemia. This can be remedied by a course of 'ferrous sulphate' tablets, often known as iron tablets.

The iron in iron tablets is in the form of hydrated iron(II) sulphate (sometimes called ferrous sulphate). As the name iron(II) suggests, the Fe^{2+} ion is present. To determine just how much Fe^{2+} is in each tablet, we can react the Fe^{2+} ions with manganate(VII) ions, which have a formula MnO_4^- .



Although this ionic equation may appear complicated at this point in your course, you can see from the colours that the deep purple of the $\text{MnO}_4^-(\text{aq})$ will disappear when the reaction is complete.

The end point is when the addition of one extra drop of potassium manganate(VII) solution turns the solution in the conical flask to a pale pink colour.

Question

- 1 Explain why this volumetric analysis does not require an indicator.

Safety

	You must wear eye protection throughout this experiment.
	1.0 mol dm ⁻³ sulphuric acid is irritant . Wash all spillages with plenty of water.

Always use a pipette filler, or other suitable safety device. Never be tempted to use your mouth to draw liquid into the pipette.

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Procedure

1. Accurately weigh **two** iron tablets and record their mass.
2. Grind up the tablets with about 5 cm³ of 1.0 mol dm⁻³ sulphuric acid, using a pestle and mortar. Transfer this into a 100 cm³ volumetric flask. Use further small volumes of the dilute sulphuric acid to rinse the ground-up tablets into the flask, until no traces of the iron tablets are left in the mortar.
3. Add more 1.0 mol dm⁻³ sulphuric acid to make the volume in the volumetric flask exactly 100.0 cm³. Stopper the flask and shake it thoroughly to mix the solution. Not all of the outer coating of the tablets will dissolve. This does not matter as it does not contain any Fe²⁺ ions.
4. Use a pipette and pipette filler to withdraw 10.0 cm³ of the tablet solution and transfer it into a clean conical flask.
5. Wash a 100 cm³ beaker with deionised water and then twice with small volumes of the 0.0050 mol dm⁻³ potassium manganate(VII) solution you are going to use. Now half fill the beaker with the potassium manganate(VII) solution and use this to fill the burette. Remember that the burette does not have to be filled to the 0.00 cm³ mark. Make sure that you run some of the solution back into the beaker to ensure that the tip of the burette is full. Read the volume of potassium manganate (VII) solution in the burette. In this case the colour of the potassium manganate(VII) solution is so intense that you cannot see the bottom of the meniscus so you must use the top of the meniscus to measure the volume.
6. Draw up a table in which you can record your titration results. Label the rows and columns, using the appropriate units. Think about which value you will need to subtract from which when deciding on the order of your rows.
7. Perform one rough titration, followed by two accurate titrations that agree within 0.05 cm³ and record the results in your table.

Questions

- 2 (a) How many moles of MnO₄⁻(aq) were added from the burette?
(b) How many moles of Fe²⁺(aq) were present in 10.0 cm³ of the iron tablet solution?
(c) How many moles of Fe²⁺(aq) were present in 100.0 cm³ of the tablet solution?
(d) What is the mass of iron in the two tablets?
(e) Find out the mass of iron that the manufacturer states is in one tablet. How does your result compare?
Suggest why there are differences.
(f) If you were to repeat this experiment, how would you make sure that your result was as accurate as possible?
- 3 The reaction you have performed can be written as two half equations.
(a) Use only the ion symbols for iron(II) and iron(III) to write a half equation showing the oxidation reaction. Use e⁻ for the electron that is lost.
(b) The other half equation is the reaction of the manganate(VII) ion. What type of reaction will this be?
- 4 In air, an aqueous solution of iron(II) ions will oxidise easily to iron(III) ions. The addition of dilute sulphuric acid prevents this. However, there is another reason for adding dilute sulphuric acid. What is the other reason?

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3. How much iron is there in an iron tablet? – Teachers' Notes

If you have used Practicals 1 and 2 in Appendix 1 of this booklet, students should have developed good techniques in titration. This volumetric analysis introduces the potassium manganate(VII) titration. In the practical scheme in Appendix 1 this is placed very early in the course and is intended to cover mole calculations and further develop skills that will be tested in the practical examination. It is not, at this stage, appropriate to spend time on a detailed explanation of the redox reaction as this comes up in the A2 part of the course. However, Question 3 on the Student Sheet should be accessible to students with the knowledge they have gained from their pre-A level course.

Learning outcomes

These are printed on the Student Sheet.

A suggested approach

Set the scene by explaining the context. Iron deficiency, or anaemia, affects many people and particularly the young. 'Ferrous sulphate' tablets, often just called iron tablets, rectify this deficiency. Iron deficiency may be the result of not eating enough foods containing iron and a change in diet to include more iron-rich food is often recommended to patients.

When drug companies manufacture iron tablets they put a set amount of iron into each tablet. This is usually written on the packet. One way a drug company can check the iron content of its tablet is to analyse it using this titration.

The Student Sheet gives full information about how to do the practical. Students could prepare for the practical session by reading the sheet in advance and answering Question 1.

The students should finish the titration in a 1-hour session.

Technical information

Requirements per student/group:

100 cm³ beaker

100.0 cm³ volumetric flask

Filter funnel

10.0 cm³ pipette

Pipette filler

50.0 cm³ burette

100 cm³ conical flask

Pestle and mortar

Sheet of white paper or a white tile

Wash bottle and deionised water

0.0050 mol dm⁻³ KMnO₄(aq) – about 100 cm³ * See notes 2 and 3 below

Approximately 1 mol dm⁻³ H₂SO₄(aq) – about 120 cm³

Two 'ferrous sulphate' tablets

Access to the packet showing the mass of iron in each tablet

Access to a balance weighing to 0.01 g

Notes

1. If brown or red colours are seen during the titrations, this will produce inaccurate results. The remedy is to add more dilute sulphuric acid.
2. The packet of iron tablets often describes the iron tablets as '200 mg tablets'.
3. If two '200 mg tablets' are used, then the expected volume of 0.0050 mol dm⁻³ KMnO₄(aq) will be 28.60 cm³ per titration. If the mass of iron is 200 mg, then the volume will be 14.30 cm³.

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Safety

The main points are included on the Student Sheet but it is the teacher's responsibility to ensure that a full risk assessment is carried out prior to the practical session(s). MSDS sheets should be consulted so that the correct action can be taken in event of a spillage and/or accident.

Answers to questions on the Student Sheet

- 1 The reaction is self-indicating. The end point of the titration is shown when one extra drop of $\text{KMnO}_4(\text{aq})$ does not go colourless but turns the reacting solution a pale pink colour.
- 2 **(a)** Moles of MnO_4^- depend on the volume of solution run in from the burette.
Moles $\text{MnO}_4^- = \text{concentration } \text{MnO}_4^-(\text{aq}) \text{ (i.e. } 0.0050 \text{ mol dm}^{-3}\text{) } \times \text{volume delivered from burette}/1000$
(b) Moles of $\text{Fe}^{2+}(\text{aq})$ in $10.0 \text{ cm}^3 = 5 \times$ the answer given in **1(a)**.
(c) Since a 100 cm^3 solution was made up from the two tablets: moles $\text{Fe}^{2+}(\text{aq}) = 10 \times$ answer in **(b)**.
(d) The moles given in **2(c)** are multiplied by $A_r \text{ Fe (56)}$.
(e) Look for an appreciation of the limitations of the experiment. The reading of the burette is not likely to be the most significant error, as the scale can be read to $\pm 0.05 \text{ cm}^3$. However, it may be that some of the Fe^{2+} was oxidised by the air. Perhaps the crushing operation did not allow all of the contents of the tablet to dissolve. Look for sensible evaluations of technique.
(f) This may depend on what is said in **(e)** but also look for sensible comments about using more tablets in case there is a variation of mass between tablets, or to make the weighing more accurate. Students may suggest that a balance that reads to 0.001 g should be used. As the usual mass of iron in a tablet is 200 mg , this is a sensible suggestion.
- 3 **(a)** $\text{Fe}^{2+}(\text{aq}) \rightarrow \text{Fe}^{3+}(\text{aq}) + \text{e}^-$
(b) Reduction.
- 4 $\text{H}^+(\text{aq})$ is an essential reactant as it appears in the reaction equation.