Core practical 11: Find the amount of iron in an iron tablet using redox titration

Objectives

- To calculate the percentage of iron in an iron tablet
- To perform a redox titration involving Fe^{2+}(aq) and MnO_4^-(aq)

Safety	Specification links
• Use eye protection. • 1.5 mol dm^{-3} sulfuric acid is an irritant.	• Practical techniques 1, 4, 5, 11 • CPAC 1a, 2a, 2b, 3a, 4a, 4b

Procedure	Notes on procedure
1. Crush the iron tablets using the pestle and mortar. 2. Transfer the crushed tablets to a weighing boat and measure their combined mass. Record this mass. 3. Empty the crushed tablets into the small beaker and reweigh the weighing boat. Record this mass. 4. Add 100 cm^3 1.5 mol dm^{-3} of sulfuric acid to the small beaker. Stir to dissolve as much of the tablets as possible. 5. Filter the solution (to remove any undissolved solids) into the volumetric flask. Rinse the beaker with more sulfuric acid and add the washings to the volumetric flask. Make up to the mark with distilled/deionised water. Stopper and shake. 6. Pipette 25.0 cm^3 of this solution into the conical flask. 7. Titrate the iron(II) solution with potassium manganate(VII) solution until the mixture has *just* turned pink. On standing, the pink colour will disappear because there is a secondary reaction between the $KMnO_4$ and another ingredient in the tablet. *Do not add any more* $KMnO_4$. 8. Record your results in an appropriate format. 9. Repeat the titration until concordant results are obtained.	• You may wish to refresh students' memories about titrations by demonstrating this practical. • Remind students that titre values should be recorded to 2 decimal places with the second figure being 0 or 5 only. • Remind students that the titration should be repeated until concordant results are obtained.

Answers to questions

1. $5Fe^{2+} + MnO_4^- + 8H^+ \rightarrow 5Fe^{3+} + Mn^{2+} + 4H_2O$
2. Average titre 21.40 cm^3

   moles = concentration $\times \frac{\text{volume}}{1000}$

   moles = 0.005 $\times \frac{21.40}{1000}$ = 0.000 107

   moles of MnO_4^- = 0.000 107 moles
3. Moles of iron = 0.000 535
4. Moles of iron in the 250 cm^3 graduated flask = 0.00535
5. 0.005 35 moles = 0.3 g (or 300 mg in 5 tablets or 60 mg in 1 tablet)
6. The answer to this question will depend on the students' results but generally the published result for an iron tablet is 65 mg of iron per tablet.
7. Procedural errors:
   - Stirring may not be sufficient to ensure that all the iron dissolves – warming the solution may help.
   - Transfer of the solution and filtering – ensure that the beaker and the filter paper are rinsed with water.
   - The solution may not be mixed – invert the volumetric flask several times to ensure thorough mixing.
   - Glassware measurements may not be read accurately – read glassware marks from the bottom of the meniscus.
   - The end-point may not be clear – use a white tile so that you can see the end-point clearly.
8. $\frac{0.05}{21.40} \times 100 = 0.23\%$

Sample data

Average titre 21.40 cm^3

M_r: Fe = 55.8

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All the maths you need

- Change the subject of an equation.
- Substitute numerical values into algebraic equations using appropriate units for physical quantities.

Equipment

- 5 iron tablets
- 100 cm^3 of 1.5 mol dm^{-3} sulfuric acid
- 100 cm^3 of 0.005 mol dm^{-3} potassium manganate(VII)
- distilled/deionised water
- pestle and mortar
- 100 cm^3 beaker
- 25 cm^3 measuring cylinder
- two 250 cm^3 beakers
- 250 cm^3 volumetric flask and stopper
- spatula, glass rod and dropping pipette
- filter funnel and filter paper
- 50 cm^3 burette and burette stand
- 25 cm^3 pipette and pipette filler
- 250 cm^3 conical flask
- white tile
- mass balance (2 d.p.) and weighing boat

Procedure

1. Crush the iron tablets using the pestle and mortar.
2. Transfer the crushed tablets to a weighing boat and measure their combined mass. Record this mass.
3. Empty the crushed tablets into the small beaker and reweigh the weighing boat. Record this mass.
4. Add 100 cm^3 1.5 mol dm^{-3} of sulfuric acid to the small beaker. Stir to dissolve as much of the tablets as possible.
5. Filter the solution (to remove any undissolved solids) into the volumetric flask. Rinse the beaker with more sulfuric acid and add the washings to the volumetric flask. Make up to the mark with distilled/deionised water. Stopper and shake.
6. Pipette 25.0 cm^3 of this solution into the conical flask.
7. Titrate the iron(II) solution with potassium manganate(VII) solution until the mixture has *just* turned pink. On standing, the pink colour will disappear because there is a secondary reaction between the $KMnO_4$ and another ingredient in the tablet. *Do not add any more $KMnO_4$*.
8. Record your results in an appropriate format.
9. Repeat the titration until concordant results are obtained.

Analysis of results

	Rough	1	2	3	4
Initial reading/cm^3					
Final reading/cm^3					
Titre/cm^3					

Use concordant results to calculate the average titre and then answer the questions below to calculate the mass of iron in the tablet.

Learning tips

- You need two equations:

  number of moles = concentration $\times$ $\dfrac{\text{volume}}{1000}$

  number of moles = $\dfrac{\text{mass}}{M_r}$

- Show all working carefully in a titration calculation and explain what you are doing in each step. That way you can still gain marks in an exam – even if you get the final answer wrong.

Questions

1. Combine the two half-equations given below to write the equation for the reaction:

   $Fe^{2+}(aq) \rightarrow Fe^{3+} + e^-$

   $MnO_4^-(aq) + 8H^+(aq) + 5e^- \rightarrow Mn^{2+}(aq) + 4H_2O(l)$

2. Use your average titre to calculate the number of moles of manganate(VII) ions that were used in the titration.
3. Use the equation to calculate the number of moles of iron(II) ions in the 25 cm^3 sample of iron(II) sulfate from the iron tablet.
4. Calculate the number of moles of iron(II) ions in the 250 cm^3 graduated flask at the start of the experiment.
5. Calculate the mass of iron in the original five iron tablets, and hence the mass of iron in one iron tablet. M_r Fe = 55.8 g mol^{-1}
6. Compare your value for the mass of iron with the information from the supplier about the composition of each iron tablet.
7. Make a list of any procedural errors. Suggest ways in which these errors can be avoided.
8. Calculate the percentage measurement uncertainty for the burette.

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Safety

- Use eye protection when handling acid.
- 1.5 mol dm^{-3} sulfuric acid is an irritant.
- Consult CLEAPSS Hazcards® 81 and 98A. Perform a risk assessment using up-to-date information before this practical is carried out.

Equipment per student/group	Notes on equipment
5 iron tablets	The iron tablets should contain iron(II) sulfate. They can be purchased by asking a pharmacist for this particular type. Iron tablets sold off the shelf usually contain iron(II) fumarate (rather than iron(II) sulfate). **These do not readily dissolve**. If you are not able to obtain tablets of this type then substitute with ammonium iron(II) sulfate.
100 cm^3 of 1.5 mol dm^{-3} sulfuric acid solution	Irritant
100 cm^3 of 0.005 mol dm^{-3} potassium manganate(VII) solution	Low hazard
distilled/deionised water	
pestle and mortar	
100 cm^3 beaker	
25 cm^3 measuring cylinder	
two 250 cm^3 beakers	
250 cm^3 volumetric flask, stoppered	
spatula, glass rod and dropping pipettes	
filter funnel and filter paper	
50 cm^3 burette and burette stand	
25 cm^3 pipette and pipette filler	
250 cm^3 conical flask	
white tile	
mass balance (2 d.p.) and weighing boat	

Notes