Finding the % of iron in lawnsand

Introduction

Lawnsand is a mixture of iron(II) compounds and sand. It is used to kill moss in lawns. The amount of iron in lawnsand can be found by titration with potassium manganate(VII) solution.

Manganate(VII), MnO_4^-, is a strong oxidising agent. It accepts electrons easily, and is reduced to colourless manganese(II) ions according to the half-equation below:

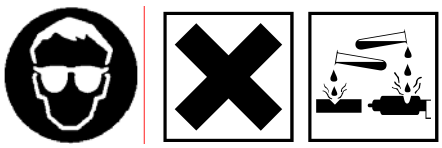
$$\underset{\text{(purple)}}{MnO_4^-(aq)} + 8H^+(aq) + 5e^- \rightleftharpoons \underset{\text{(colourless)}}{Mn^{2+}(aq)} + 4H_2O(l)$$

The electrons are provided by reducing agents such as iron(II) salts: $Fe^{2+}(aq) \rightleftharpoons Fe^{3+}(aq) + e^-$

Overall equation: $MnO_4^-(aq) + 8H^+(aq) + 5Fe^{2+}(aq) \rightarrow Mn^{2+}(aq) + 4H_2O(l) + Fe^{3+}(aq)$

As a result, manganate(VII) can be used in acidic solution to determine the number of moles of reducing agent present. Manganate(VII) is added from a burette to a solution of the reducing agent and is decolourised immediately. As soon as the reducing agent is used up, the next drop of manganate(VII) is not decolourised, and so the solution in the conical flask goes pale pink. The end-point of the titration is the first permanent appearance of this pale pink colour. Manganate(VII) is therefore self-indicating and no other indicator is needed. The acid used to provide $H^+(aq)$ is dilute sulphuric acid; this should always be in excess or else insoluble brown MnO_2 will form.

Methods



1. Use weighing by difference to obtain between 9.50g and 10.50g of the lawnsand provided. Record all your weighings.
2. Thoroughly mix the lawnsand with about $100cm^3$ of sulphuric acid in a beaker.
   Filter into a volumetric flask. Wash the beaker several times with de-ionised water and transfer all the washings to the filter funnel. Make up to the mark with de-ionised water and mix thoroughly.
   Label the volumetric flask with your name, date and details of the contents.
3. Pour some of your solution into a beaker and pipette $25cm^3$ from there into a conical flask.
4. Titrate against 0.010M $KMnO_4$ from the burette until at least two concordant results are obtained.

Results and Analysis

1. Record your weighings, details of the reagents involved and titration readings in an appropriate form.
2. Calculate the mean titre of $KMnO_4$ solution and then the number of moles of $KMnO_4$ you used.
3. Calculate the number of moles of Fe^{2+} in $25cm^3$ of solution, and then the number of moles of Fe^{2+} in the original total quantity of lawnsand.
4. Calculate the mass of iron in your sample of lawnsand, and then the % by mass of iron in it.

Evaluation

1. Discuss the sources of error and their importance in the experiment.
   Include errors that you have calculated, and those that you cannot calculate, in your discussion.
2. Discuss ways of improving the experiment.

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Technicians Notes

Reagents

- Make a 1:1 by mass mixture of ammonium iron(II) sulphate and sand. Label it "Lawnsand". Allow approximately 12g per student.
- 0.010M $KMnO_4$ solution – allow approximately 300cm^3 per student.
- 1M sulphuric acid – allow approximately 120cm^3 per student.
- De-ionised water

Apparatus

In addition to the usual apparatus needed for a titration:

Access is needed to ±0.1g balances and ±0.01g balances

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