Finding the % of Fe^{2+} in $Fe(NH_4)_2(SO_4)_2.xH_2O$ and the value of x

Introduction

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:

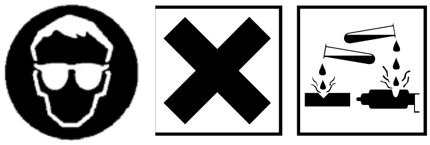
$$\underset{\text{(purple)}}{MnO_4^-(aq)} + 8H^+(aq) + 5e^- \rightleftharpoons \underset{\text{(colourless)}}{Mn^{2+}(aq)} + 4H_2O(l)$$

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 (from the above, one mole of manganate(VII) will react with 5 moles of Fe^{2+}). 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 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. Weigh between 9.5g and 10.5g of the iron(II) salt provided.
   Record your weighings to the nearest 0.1g.
2. Dissolve the salt in about 150cm^3 of water in a beaker.
   Transfer to a volumetric flask, mix thoroughly and make up to 250cm^3.
3. Pour the solution into a beaker and pipette 25cm^3 from there into a conical flask.
4. Using a measuring cylinder, add 25cm^3 of dilute sulphuric acid.
   Titrate against 0.02M $KMnO_4$ from the burette.

Results and calculations

1. Record your weighings and titration readings in the most appropriate form.
2. Calculate the average titre of $KMnO_4$ solution.
3. Calculate the number of moles of $KMnO_4$ you used.
4. Calculate the number of moles of Fe^{2+} in 25cm^3 of solution, and then the number of moles of Fe^{2+} in your sample of crystals.
5. Calculate the mass of Fe^{2+} in your sample of crystals, and then the % by mass of Fe^{2+} in them.
6. Calculate the mass of anhydrous $Fe(NH_4)_2(SO_4)_2$ present in your original sample of crystals, and then the mass of water in those crystals.
7. What is the value of x in $Fe(NH_4)_2(SO_4)_2.xH_20$?
8. Discuss the sources of error and their importance in the experiment, and ways of improving it.