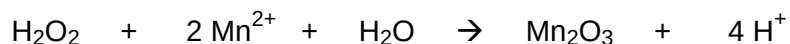# ACID BASE PRACTICE QUESTIONS 6

Question 4 **(15 marks)**

A jar containing a pale pink powder is labelled *commercial grade manganese (II) sulfate MnSO$_4$*. A chemist needs to know its percentage by mass purity. He decides to analyse it by utilizing the reaction between hydrogen peroxide and manganese ion. The manganese ions are converted into a black precipitate of manganese (III) oxide. The black oxide quickly settles to the bottom of the conical flask. The equation for the reaction is

$$H_2O_2 \ + \ 2\,Mn^{2+} \ + \ H_2O \ \rightarrow \ Mn_2O_3 \ + \ 4\,H^+$$

The end point is taken to be when the final drop of hydrogen peroxide no longer produced a black precipitate.

The chemist dissolved 2.000 g sample of the impure manganese (II) sulfate in water in a 100 mL volumetric flask. He then pipetted 25.00 mL of this solution and diluted it to 250 mL in another volumetric flask.

Next, he titrated 20.00 mL aliquots of the diluted manganese (II) sulfate solution against 0.002211 mol L^{-1} hydrogen peroxide solution. The average titre required was 46.55 mL.

(a) How many moles of hydrogen peroxide were consumed in an average titration? (2 marks)

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(b) How many moles of manganese (II) ions were oxidised in an average titration? (2 marks)

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(c) How many moles of manganese (II) sulfate were present in the impure sample? (3 marks)

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(d)   What was the percentage purity of the commercial manganese (II) sulfate?   (3 marks)

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(e)   The chemist could also analyse the impure manganese sulfate by dissolving a sample in water, then adding excess hydrogen peroxide solution and finally performing a titration to determine the excess hydrogen peroxide.

(i)   Suggest what reagent he could use for the titration.
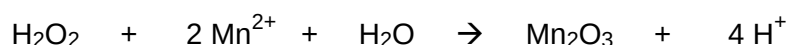Include an equation to justify your answer.   (4 marks)

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(ii) Suggest how the end point of this titration would be determined. (1 mark)

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ANSWER

Question 43 **(15 marks)**

A jar containing a pale pink powder is labelled *commercial grade manganese (II) sulfate MnSO₄*. A chemist needs to know its percentage by mass purity. He decides to analyse it by utilizing the reaction between hydrogen peroxide and manganese ion. The manganese ions are converted into a black precipitate of manganese (III) oxide. The black oxide quickly settles to the bottom of the conical flask. The equation for the reaction is

$$H_2O_2 \quad + \quad 2\ Mn^{2+} \quad + \quad H_2O \quad \rightarrow \quad Mn_2O_3 \quad + \quad 4\ H^+$$

The end point is taken to be when the final drop of hydrogen peroxide no longer produced a black precipitate.

The chemist dissolved 2.000 g sample of the impure manganese (II) sulfate in water in a 100 mL volumetric flask. He then pipetted 25.00 mL of this solution and diluted it to 250 mL in another volumetric flask.

Next, he titrated 20.00 mL aliquots of the diluted manganese (II) sulfate solution against 0.002211 mol L^{-1} hydrogen peroxide solution. The average titre required was 46.55 mL.

(a) How many moles of hydrogen peroxide were consumed in an average titration? (2 marks)

n (H_2O_2) = c V = (0.002211)(0.04655)   ☐**1**

            = 0.000102922 mol [1.03 x 10⁻⁴]   ☐**1**

(b) How many moles of manganese (II) ions were oxidised in an average titration? (2 marks)

n (Mn^{2+}) = 2 x n (H_2O_2)   ☐**1**

           = (2)(0.000102922) = 0.000205844 mol [2.06 x 10⁻⁴]   ☐**1**

(c) How many moles of manganese (II) sulfate were present in the impure sample? (3 marks)

                                                         ☐**1**

$n \, (MnSO_4) \;=\; n \, (Mn^{2+}) \;\times\; 250 \, / \, 20 \;\times\; 100 \, / \, 25$

$\qquad\qquad = \; 0.000205844 \;\times\; 250 \, / \, 20 \;\times\; 100 \, / \, 25$ 1

$\qquad\qquad = \; 0.0102922 \; mol \qquad [1.03 \times 10^{-2}]$ 1

(d) What was the percentage purity of the commercial manganese (II) sulfate? (3 marks)

$m \, (MnSO_4) \;=\; n \, M \;=\; (0.0102922)(151)$ 1

$\qquad\qquad\qquad = \; 1.554 \; g \qquad [1.55 \; g]$

Percentage purity $=$ mass $MnSO_4$ / sample mass $\times$ 100 2

$\qquad\qquad\qquad = \; (1.554 \, / \, 2.000) \times 100$

$\qquad\qquad\qquad = \; 77.7 \, \%$ 1

(e) The chemist could also analyse the impure manganese sulfate by dissolving a sample in water, then adding excess hydrogen peroxide solution and finally performing a titration to determine the excess hydrogen peroxide.

(i) Suggest what reagent he could use for the titration.
Include an equation to justify your answer. (4 marks)

The excess peroxide could be titrated against acidified permanganate 2

$5 \, H_2O_2 \;+\; 2 \, MnO_4^- \;+\; 6 \, H^+ \;\rightarrow\; 2 \, Mn^{2+} \;+\; 5 \, O_2 \;+\; 8 \, H_2O$ 2

Accept reductants, such as chromium (III), bromide, iodide, iron (II) and oxalic acid

$3 \, H_2O_2 \;+\; 2 \, Cr^{3+} \;+\; H_2O \;\rightarrow\; Cr_2O_7^{2-} \;+\; 8 \, H^+$

$H_2O_2 \;+\; 2 \, Br^- \;+\; 2 \, H^+ \;\rightarrow\; Br_2 \;+\; 2 \, H_2O$

$$H_2O_2 \;+\; 2\,I^- \;+\; 2\,H^+ \rightarrow I_2 \;+\; 2\,H_2O$$

$$H_2O_2 \;+\; HOOCCOOH \rightarrow 2\,CO_2 \;+\; 2\,H_2O$$

$$H_2O_2 \;+\; 2\,Fe^{2+} \;+\; 2\,H^+ \rightarrow 2\,Fe^{3+} \;+\; 2\,H_2O$$

(ii) Suggest how the end point of this titration would be determined.   (1 mark)

For permanganate in the burette the end point will be when the final drop added from the burette turns the peroxide solution in the flask permanently pink.

For permanganate in the flask the end point will be when the final drop of peroxide solution added from the burette permanently decolorizes the purple colour in the flask.

For the other reagent students may suggest valid end point determinations.