REDOX TITRATIONS – TOPIC TEST 1

QUESTION 1
A 0.0484 mol L^{-1} standard solution of potassium permanganate was titrated against 25.00 mL of an iron(II) sulfate solution. The equivalence point, as indicated by a faint pink colour, was reached when 15.50 mL of potassium permanganate solution had been added.
Calculate the concentration of the iron(II) sulfate solution. (6 marks)

Solution

QUESTION 2

(a) Given the following two half–reactions, construct the **full ionic redox equation** for the reaction of the thiosulfate ion $S_2O_3^{2-}$ and iodine I_2. (1 mark)

$S_4O_6^{2-}{}_{(aq)} + 2e^- \rightarrow 2S_2O_3^{2-}{}_{(aq)}$
$I_{2(aq)} + 2e^- \rightarrow 2I^-{}_{(aq)}$

(b) What mass of iodine reacts with 23.5 cm^3 of 0.0120 mol/L sodium thiosulfate solution?

(c) 25.0 cm^3 of a solution of iodine in potassium iodide solution required 26.5 cm^3 of 0.0950 mol/L sodium thiosulfate solution to titrate the iodine. What is the molarity of the iodine solution and the mass of iodine per L?

Solution

QUESTION 3

A standardised solution of permanganate ion is used to determine the molar concentration of an unknown hydrogen peroxide, H_2O_2, solution as indicated in the net ionic equation below:

$$MnO^-_{4(aq)} + H_2O_{2(aq)} \rightarrow Mn^{2+}_{(aq)} + O_{2(g)}$$

A student places 50.0 mL of permanganate ion solution of known concentration into a burette. 20.0 mL of H_2O_2 solution of unknown concentration is placed into a flask with 20.0 mL of deionised water and 1 mL of $5.0\,M$ H_2SO_4. In the titration, the $KMnO_4$ solution is added to the solution containing H_2O_2.

(a) What is the oxidation number of the oxygen atom in the hydrogen peroxide molecule?

(b) Which is the reducing agent in the reaction above?

(c) Why hasn't an indicator been included in the reaction solution?

(d) How would the calculated concentration of the H_2O_2 solution be affected if 4.0 mL of H_2SO_4 was added to the H_2O_2 solution instead of 1.0 mL? Give a reason for your answer.

Solution

ANSWERS

QUESTION 1

Write the balanced redox reaction:

$$MnO^-_{4(aq)} + 8H^+_{(aq)} + 5e^- \rightarrow Mn^{2+}_{(aq)} + 4H_2O_{(l)}$$
$$Fe^{2+}_{(aq)} \rightarrow Fe^{3+}_{(aq)} + e^-$$

$$MnO^-_{4(aq)} + 8H^+_{(aq)} + 5Fe^{2+}_{(aq)} \rightarrow Mn^{2+}_{(aq)} + 4H_2O_{(l)} + 5Fe^{3+}_{(aq)}$$

Calculate the number of mole of known reactant:

$$n(MnO_4^-) = cV = 0.0484 \times 0.01550 = 7.502 \times 10^{-4}\, mol$$

Use the mole ratio in the balanced equation to determine the number of mole of unknown reacting:

$$n(Fe^{2+}) = 5 \times n(MnO_4^-) = 5 \times 7.502 \times 10^{-4} = 0.003751\, mol$$

Take into account any dilutions performed:

Not applicable

Calculate the required value:

$$c = \frac{n}{V} = \frac{0.003751}{0.02500} = 0.1500\, M$$

QUESTION 2

(a) $2S_2O_3^{2-}{}_{(aq)} \rightarrow S_4O_6^{2-}{}_{(aq)} + 2e^-$

$I_{2(aq)} + 2e^- \rightarrow 2I^-{}_{(aq)}$

$2S_2O_3^{2-}{}_{(aq)} + I_{2(aq)} \rightarrow S_4O_6^{2-}{}_{(aq)} + 2I^-{}_{(aq)}$

(b) $n(S_2O_3^{2-}) = cV = 0.0120 \times 0.0235 = 0.000282\, mol$

$$n(I_2) = \frac{1}{2} \times n(S_2O_3^{2-}) = \frac{1}{2} \times 0.000282 = 0.000141\, mol$$

$$m = n \times M = 0.000141 \times 253.8 = 0.0358\, g$$

(c) $n(S_2O_3^{2-}) = cV = 0.095 \times 0.0265 = 0.002518\,mol$

$n(I_2) = \frac{1}{2} \times n(S_2O_3^{2-}) = \frac{1}{2} \times 0.002518 = 0.001259\,mol$ in $25\,mL$

$c = \frac{n}{V} = \frac{0.001259}{0.0250} = 0.0504\,M$

Mass of iodine per L:

$0.0504\,M = 0.0504\,mol\ per\ litre$ and $n = \frac{m}{M}$ $\therefore m = n \times M$

$0.0504\,M = 0.0504 \times 253.8\,g\ per\ litre$
$= 12.8\,g\ per\ litre$

QUESTION 3

(a) -1

(b) H_2O_2

(c) The reaction is self-indicating i.e. a colour change occurs naturally at the equivalence point of the reaction.

(d) There would be no change to the calculated concentration of H_2O_2. Adding extra H_2SO_4 does not change the number of mole of H_2O_2 present, which means that the volume of permanganate ion delivered from the burette will not change.