

## REDOX TITRATIONS – TOPIC TEST 1

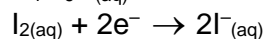
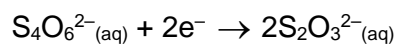
### QUESTION 1

A  $0.0484 \text{ mol L}^{-1}$  standard solution of potassium permanganate was titrated against  $25.00 \text{ mL}$  of an iron(II) sulfate solution. The equivalence point, as indicated by a faint pink colour, was reached when  $15.50 \text{ mL}$  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  $\text{S}_2\text{O}_3^{2-}$  and iodine  $\text{I}_2$ . (1 mark)

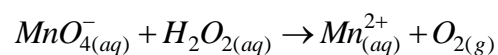


- (b) What mass of iodine reacts with  $23.5 \text{ cm}^3$  of  $0.0120 \text{ mol/L}$  sodium thiosulfate solution?
- (c)  $25.0 \text{ cm}^3$  of a solution of iodine in potassium iodide solution required  $26.5 \text{ cm}^3$  of  $0.0950 \text{ mol/L}$  sodium thiosulfate solution to titrate the iodine. 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:



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  $H_2O_2$ .

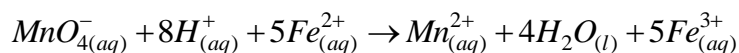
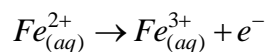
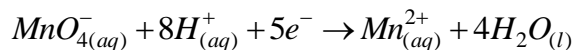
- What is the oxidation number of the oxygen atom in the hydrogen peroxide molecule?
- Which is the reducing agent in the reaction above?
- Why hasn't an indicator been included in the reaction solution?
- How would the calculated concentration of the  $H_2O_2$  solution be affected if 4.0 mL of  $H_2SO_4$  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:



Calculate the number of mole of known reactant:

$$n(\text{MnO}_4^-) = cV = 0.0484 \times 0.01550 = 7.502 \times 10^{-4} \text{ mol}$$

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

$$n(\text{Fe}^{2+}) = 5 \times n(\text{MnO}_4^-) = 5 \times 7.502 \times 10^{-4} = 0.003751 \text{ mol}$$

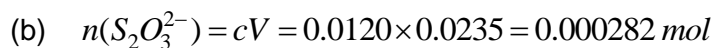
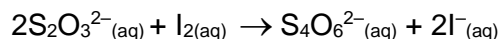
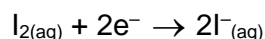
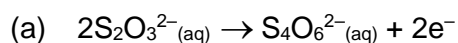
Take into account any dilutions performed:

Not applicable

Calculate the required value:

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

### QUESTION 2



$$n(\text{I}_2) = \frac{1}{2} \times n(\text{S}_2\text{O}_3^{2-}) = \frac{1}{2} \times 0.000282 = 0.000141 \text{ mol}$$

$$m = n \times M = 0.000141 \times 253.8 = 0.0358 \text{ g}$$

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

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

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

Mass of iodine per L:

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

$$0.0504 \text{ M} = 0.0504 \times 253.8 \text{ g per litre} \\ = 12.8 \text{ g per litre}$$

### QUESTION 3

(a)  $-1$



(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.