Practical question 1 (Paper 3 practice) 17 marks - 25 minutes

Hydrogen peroxide reacts with potassium iodide to produce iodine:

$$H_2O_2 + 2KI + H_2SO_4 \rightarrow 2H_2O + I_2 + K_2SO_4$$

The rate of this reaction can be followed by using the *iodine clock* technique. A fixed amount of sodium thiosulfate ($Na_2S_2O_3$) solution and a small amount of starch solution are added to the reaction mixture. The sodium thiosulfate reacts with the iodine as soon as it is formed but when all the sodium thiosulfate is used up the iodine reacts with the starch to produce a blue-black colour. The experiment thus measures the time taken for a fixed amount of iodine to be produced.

The equation for the reaction between thiosulfate ions and iodine is: $I_2 + 2S_2O_3^{2-} \rightarrow 2I^- + S_4O_6^{2-}$

A student carried out a series of experiments to investigate how changing the concentration of potassium iodide changes the rate of the reaction. They made up the reaction mixtures shown in the table below. The hydrogen peroxide solution was added last and the timer started. The timer was stopped when the reaction mixture became blue-black.

The student recorded the following data

Expt. No.	Volume 1.00 mol dm^{-3} H_2SO_4 /cm^3 ±0.1	Volume 0.100 mol dm^{-3} KI /cm^3 ±0.1	Volume H_2O /cm^3 ±0.1	Volume of 0.00500 mol dm^{-3} $Na_2S_2O_3$ /cm^3 ±0.1	Volume of 1.2 volume hydrogen peroxide ±0. 1	Time / s ±1
1	10.0	25.0	0.0	10.0	5.0	32
2	10.0	20.0	5.0	10.0	5.0	40
3	10.0	15.0	12.0	10.0	5.0	
4	10.0	10.0	15.0	10.0	5.0	80
5	10.0	5.0	20.0	10.0	5.0	160

1.0±0.1 cm^3 of starch solution was added to each reaction mixture.

(a) Other than the concentrations and volumes of the reactants explain one other variable that it is important to control in this experiment. [1]

Temperature – because any change in temperature will have a large effect on the rate of reaction

(b) (i) Calculate the absolute uncertainty on the volume of the reaction mixture in experiment 1 [1]

0.1+0.1+0.1+0.1+0.1+0.1 = ±0.6 cm^3

Volumes are added together, therefore add the absolute uncertainties (including the starch)

(ii) Calculate the concentration of potassium iodide in the reaction mixture in experiment 1. [2]

$\frac{25.0}{51.0} \times 0.100 = 0.0490$ mol dm^{-3}

(iii) Determine the absolute uncertainty of the concentration of potassium iodide in experiment 1 [2]

$\frac{25.0 \pm 0.1}{51.0 \pm 0.6} \times 0.100$ dividing values with uncertainties, therefore add % uncertainties

% uncertainty on volume of KI = 0.1/25.0 x 100 = 0.4%
% uncertainty on total volume = 0.6/51.0 x 100 = 1.2%
Total % uncertainty on concentration of KI is 0.4 + 1.2 = 1.6%
Absolute uncertainty = 1.6/100 x0.0490 = ±8x10^{-4} mol dm^{-3}

(c) (i) Calculate the number of moles of iodine required to react with the sodium thiosulfate present in experiment 1. [2]

Moles of $Na_2S_2O_3$ = 10.0/1000 x 0.00500 = 5.00x10^{-5} mol
Moles of I_2 = ½ x 5.00x10^{-5} = 2.50x10^{-5} mol

(ii) Calculate the change in concentration of iodine required to cause the reaction mixture to become blue-black and hence the rate of reaction in experiment 1 [2]

2.50×10^{-5} mol I_2 must be produced to react with all the sodium thiosulfate, then any excess iodine will react with the starch
Total volume = 51.0 cm^3
Concentration of iodine = $(2.50 \times 10^{-5})/(51.0/1000) = 4.90 \times 10^{-4}$ mol dm^{-3}
Initial concentration of iodine = 0.00 mol dm^{-3}
Change in concentration of iodine = 4.90×10^{-4} mol dm^{-3}

$$\text{Rate of reaction} = \frac{\text{change in concentration}}{\text{Time}}$$

2 significant figures now because time was to 2 significant figures

$$\text{Rate of reaction} = \frac{4.90 \times 10^{-4}}{32} = 1.5 \times 10^{-5} \text{ mol dm}^{-3} \text{ s}^{-1}$$

(d) The student made a mistake in determining the volumes to use in experiment 3. Explain what the mistake is and the effect this would have on the time recorded. [2]

The volume of H_2O should have been 10.0 and not 12.0/volume of water was too large;
The time recorded will be longer because the solutions will be more dilute and the rate of reaction will be slower;

(e) Use the data in the table to deduce the order of reaction with respect to potassium iodide. [2]

From e.g. experiments 2 and 4, halving the [KI] doubles the reaction time, i.e. halves the rate.
Therefore if [KI] is multiplied by 2, the rate increases by a factor of 2^1
First order with respect to KI

(f) Hydrogen peroxide decomposes to form water and oxygen.

(i) Write an equation for the decomposition of hydrogen peroxide. [1]

$$2H_2O_2 \rightarrow 2H_2O + O_2$$

(ii) The concentration of hydrogen peroxide solutions is often expressed as a volume concentration, where the number refers to the volume of oxygen gas (measured at STP) obtained when 1 cm^3 of solution decomposes. Thus 1 cm^3 of *20 volume* hydrogen peroxide solution produces 20 cm^3 of oxygen gas when it decomposes fully. Calculate the concentration in mol dm^{-3} of a 1.2 volume hydrogen peroxide solution. [2]

1 cm^3 of solution produces 1.2 cm^3 of O_2

$$\text{Moles of } O_2 = \frac{(1.2/1000)}{22.7} = 5.29 \times 10^{-5} \text{ mol}$$

Moles of $H_2O_2 = 2 \times 5.29 \times 10^{-5} = 1.06 \times 10^{-4}$ mol

Rounded to 2 significant figures now because 1.2 was 2 significant figures (1 cm^3 assumed to be exact)

$$\text{Concentration of } H_2O_2 = \frac{1.06 \times 10^{-4}}{(1/1000)} = 0.11 \text{ mol dm}^{-3}$$