Q1. Hydrogen peroxide is a powerful oxidising agent. Acidified hydrogen peroxide reacts with iodide ions to form iodine according to the following equation.

$$H_2O_2(aq) + 2H^+(aq) + 2I^-(aq) \rightarrow I_2(aq) + 2H_2O(l)$$

The **initial rate** of this reaction is investigated by measuring the time taken to produce sufficient iodine to give a blue colour with starch solution.

A series of experiments was carried out, in which the concentration of iodide ions was varied, while keeping the concentrations of all of the other reagents the same. In each experiment the time taken (t) for the reaction mixture to turn blue was recorded.

The initial rate of the reaction can be represented as $\left(\frac{1}{t}\right)$, and the initial concentration of iodide ions can be represented by the volume of potassium iodide solution used.

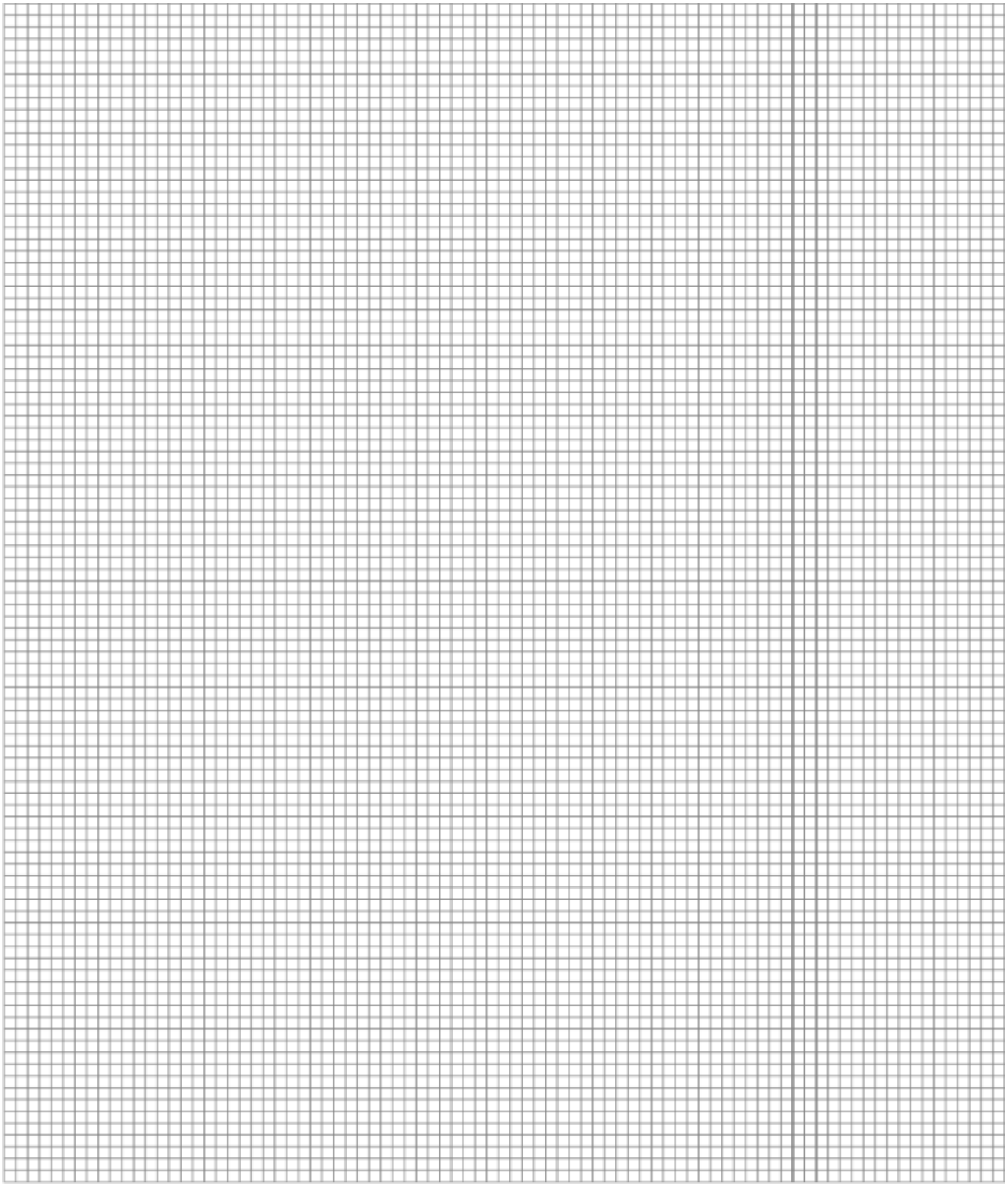
A graph of $\log_{10}\left(\frac{1}{t}\right)$ on the y-axis against $\log_{10}$ (volume of KI(aq)) is a straight line. The gradient of this straight line is equal to the order of the reaction with respect to iodide ions.

The results obtained are given in the table below. The time taken for each mixture to turn blue was recorded on a stopclock graduated in seconds.

Expt.	Volume of KI(aq) / cm^3	$\log_{10}$ (volume of KI(aq))	Time / s	$\log_{10}\left(\frac{1}{t}\right)$
1	5	0.70	71	−1.85
2	8	0.90	46	−1.66
3	10	1.00	37	−1.57
4	15	1.18	25	−1.40
5	20	1.30	19	−1.28
6	25	1.40	14	−1.15

(a) Use the results given in the table to plot a graph of $\log_{10}\left(\frac{1}{t}\right)$ on the y-axis against $\log_{10}$ (volume of KI(aq)).

Draw a straight line of best fit on the graph, ignoring any anomalous points.



(5)

(b) Determine the gradient of the line you have drawn. Give your answer to two decimal places. Show your working.

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(3)

(c) Deduce the order of reaction with respect to iodide ions.

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(1)

(d) A student carried out the experiment using a flask on the laboratory bench. The student recorded the time taken for the reaction mixture to turn blue. State **one** way this method could be improved, other than by repeating the experiment or by improving the precision of time or volume measurements. Explain why the accuracy of the experiment would be improved.

Improvement .....

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Explanation .....

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(2)

(Total 11 marks)

Q2. (a) The initial rate of the reaction between substances **P** and **Q** was measured in a series of experiments and the following rate equation was deduced.

$$\text{rate} = k[\mathbf{P}]^2[\mathbf{Q}]$$

(i) Complete the table of data below for the reaction between **P** and **Q**.

Experiment	Initial [**P**] / mol dm^{-3}	Initial [**Q**] / mol dm^{-3}	Initial rate / mol dm^{-3} s^{-1}
1	0.20	0.30	4.8×10^{-3}
2	0.10	0.10	
3	0.40		9.6×10^{-3}
4		0.60	19.2×10^{-3}

(ii) Using the data from experiment 1, calculate a value for the rate constant, *k*, and deduce its units.

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(6)

(b) What change in the reaction conditions would cause the value of the rate constant to change?

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(1)
(Total 7 marks)

Q3. (a) The following data were obtained in a series of experiments on the rate of the reaction between compounds **A** and **B** at a constant temperature.

Experiment	Initial concentration of **A**/mol dm^{-3}	Initial concentration of **B**/mol dm^{-3}	Initial rate/mol dm^{-3} s^{-1}
1	0.15	0.24	0.45×10^{-5}
2	0.30	0.24	0.90×10^{-5}
3	0.60	0.48	7.20×10^{-5}

(i) Show how the data in the table can be used to deduce that the reaction is first-order with respect to **A**.

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(ii) Deduce the order with respect to **B**.

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(2)

(b) The following data were obtained in a second series of experiments on the rate of the reaction between compounds **C** and **D** at a constant temperature.

Experiment	Initial concentration of **A**/mol dm^{-3}	Initial concentration of **B**/mol dm^{-3}	Initial rate/mol dm^{-3} s^{-1}
4	0.75	1.50	9.30×10^{-5}
5	0.20	0.10	To be calculated

The rate equation for this reaction is

rate = $k[\mathbf{C}]^2[\mathbf{D}]$

(i) Use the data from Experiment 4 to calculate a value for the rate constant, *k*, at this temperature. State the units of *k*.

Value for k .....

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Units of k .....

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(ii) Calculate the value of the initial rate in Experiment 5.

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(4)
(Total 6 marks)

Q4. (a) The following table shows the results of three experiments carried out at the same temperature to investigate the rate of the reaction between compounds **P** and **Q**.

	Experiment 1	Experiment 2	Experiment 3
Initial concentration of **P**/mol dm^{-3}	0.50	0.25	0.25
Initial concentration of **Q**/mol dm^{-3}	0.36	0.36	0.72
Initial rate/mol $dm^{-3} s^{-1}$	7.6×10^{-3}	1.9×10^{-3}	3.8×10^{-3}

Use the data in the table to deduce the order with respect to **P** and the order with respect to **Q**.

*Order with respect to **P*** .....

*Order with respect to **Q*** .....

(2)

(b) In a reaction between **R** and **S**, the order of reaction with respect to **R** is one, the order of reaction with respect to **S** is two and the rate constant at temperature T_1 has a value of 4.2×10^{-4} mol^{-2} dm^6 s^{-1}.

(i) Write a rate equation for the reaction. Calculate a value for the initial rate of reaction when the initial concentration of **R** is 0.16 mol dm^{-3} and that of **S** is 0.84 mol dm^{-3}.

Rate equation .....

Calculation .....

(ii) In a second experiment performed at a different temperature, T_2, the initial rate of reaction is 8.1×10^{-5} mol $dm^{-3} s^{-1}$ when the initial concentration of **R** is 0.76 mol dm^{-3} and that of **S** is 0.98 mol dm^{-3}. Calculate the value of the rate constant at temperature T_2.

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(iii) Deduce which of T_1 and T_2 is the higher temperature.

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(6)
(Total 8 marks)

Q5. The rate of the reaction between substance **A** and substance **B** was studied in a series of experiments carried out at the same temperature. In each experiment the initial rate was measured using different concentrations of **A** and **B**. These results were used to deduce the order of reaction with respect to **A** and the order of reaction with respect to **B**.

(a) What is meant by the term *order of reaction* with respect to **A**?

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(1)

(b) When the concentrations of **A** and **B** were both doubled, the initial rate increased by a factor of 4. Deduce the **overall** order of the reaction.

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(1)

(c) In another experiment, the concentration of **A** was increased by a factor of three and the concentration of **B** was halved. This caused the initial rate to increase by a factor of nine.

(i) Deduce the order of reaction with respect to **A** and the order with respect to **B**.

*Order with respect to **A*** .....

*Order with respect to **B*** .....

(ii) Using your answers from part (c)(i), write a rate equation for the reaction and suggest suitable units for the rate constant.

Rate equation .....

Units for the rate constant .....

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(4)
(Total 6 marks)

Q6. Propanone and iodine react in acidic conditions according to the following equation.

$$CH_3COCH_3 + I_2 \rightarrow ICH_2COCH_3 + HI$$

A student studied the kinetics of this reaction using hydrochloric acid and a solution containing propanone and iodine. From the results the following rate equation was deduced.

$$\text{rate} = k[CH_3COCH_3][H^+]$$

(a) Give the overall order for this reaction.

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(1)

(b) When the initial concentrations of the reactants were as shown in the table below, the initial rate of reaction was found to be 1.24×10^{-4} mol dm^{-3} s^{-1}.

	initial concentration / mol dm^{-3}
CH_3COCH_3	4.40
I_2	5.00×10^{-3}
H^+	0.820

Use these data to calculate a value for the rate constant, *k*, for the reaction and give its units.

Calculation .....

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Units .....

(3)

(c) Deduce how the initial rate of reaction changes when the concentration of iodine is doubled but the concentrations of propanone and of hydrochloric acid are unchanged.

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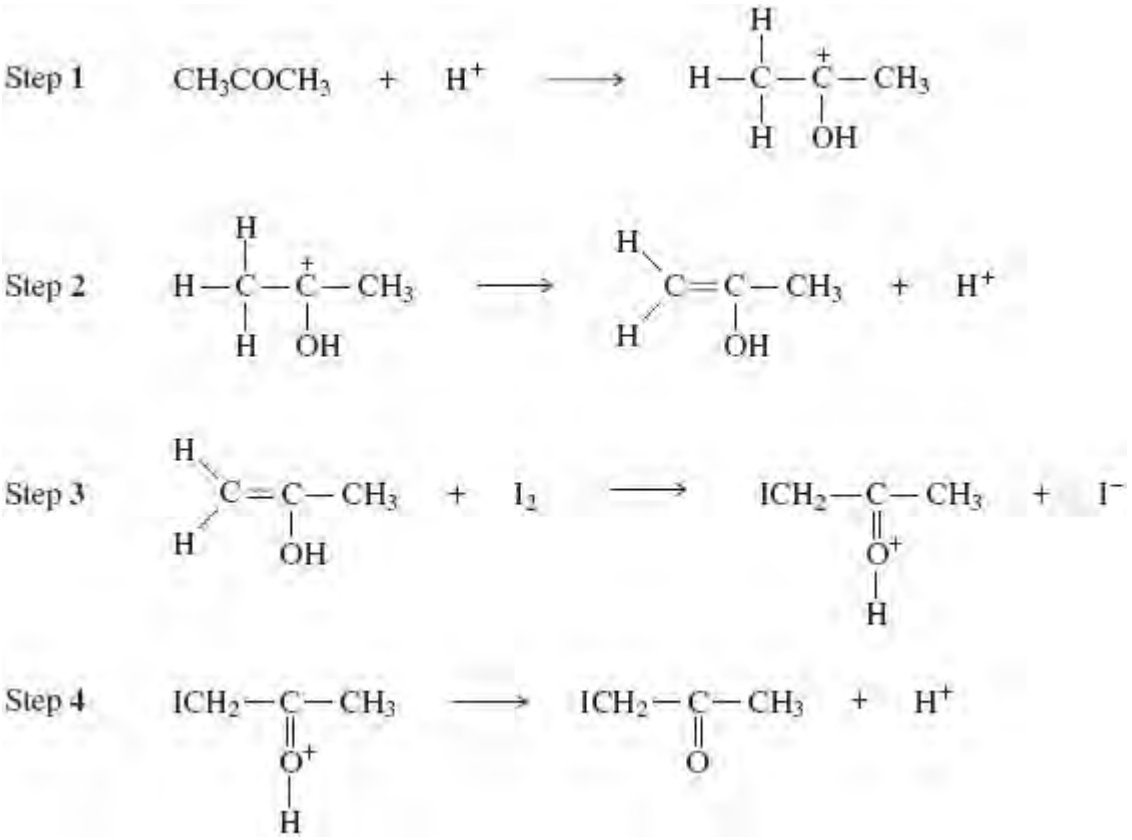
(1)

(d) The following mechanism for the overall reaction has been proposed.



Use the rate equation to suggest which of the four steps could be the rate-determining step. Explain your answer.

Rate-determining step .....

Explanation .....

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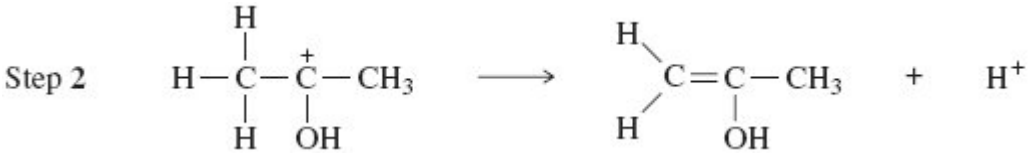
(2)

(e) Use your understanding of reaction mechanisms to predict a mechanism for Step **2** by adding one or more curly arrows as necessary to the structure of the carbocation below.



(1)
(Total 8 marks)

Q7. Kinetic studies enable chemists to suggest mechanisms for reactions.

(a) The following data were obtained in a series of experiments on the rate of the reaction between compounds **A** and **B** at a constant temperature.

Experiment	Initial concentration of **A**/mol dm^{-3}	Initial concentration of **B**/mol dm^{-3}	Initial rate/ mol dm^{-3} s^{-1}
1	0.12	0.15	0.32×10^{-3}
2	0.36	0.15	2.88×10^{-3}
3	0.72	0.30	11.52×10^{-3}

(i) Deduce the order of reaction with respect to **A**.

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(ii) Deduce the order of reaction with respect to **B**.

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(2)

(b) The following data were obtained in a series of experiments on the rate of the reaction between NO and O_2 at a constant temperature.

Experiment	Initial concentration of NO/mol dm^{-3}	Initial concentration of O_2/mol dm^{-3}	Initial rate/ mol dm^{-3} s^{-1}
4	5.0×10^{-2}	2.0×10^{-2}	6.5×10^{-4}
5	6.5×10^{-2}	3.4×10^{-2}	To be calculated

The rate equation for this reaction is

$$\text{rate} = k[NO]^2[O_2]$$

(i) Use the data from Experiment **4** to calculate a value for the rate constant, *k*, at this temperature, and state its units.

Value of k .....

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Units of k .....

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(ii) Calculate a value for the initial rate in Experiment **5**.

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(iii) Using the rate equation, a scientist suggested a mechanism for the reaction which consisted of the two steps shown below.

Step 1 $NO + NO \rightarrow N_2O_2$

Step 2 $N_2O_2 + O_2 \rightarrow 2NO_2$

Which did the scientist suggest was the rate–determining step?

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(5)
(Total 7 marks)