Q18.

(a) The data in the following table were obtained in two experiments about the rate of the reaction between substances **B** and **C** at a constant temperature.

Experiment	Initial concentration of **B** / mol dm^{-3}	Initial concentration of **C** / mol dm^{-3}	Initial rate / mol dm^{-3} s^{-1}
1	4.2×10^{-2}	2.6×10^{-2}	8.4×10^{-5}
2	6.3×10^{-2}	7.8×10^{-2}	To be calculated

The rate equation for this reaction is known to be

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

(i) Use the data from Experiment **1** to calculate a value for the rate constant *k* at this temperature and deduce its units.

Calculation _____

_____

_____

_____

_____

Units _____

_____

(3)

(ii) Calculate a value for the initial rate in Experiment **2**.

_____

_____

_____

(1)

(b) The data in the following table were obtained in a series of experiments about the rate of the reaction between substances **D** and **E** at a constant temperature.

Experiment	Initial concentration of **D** / mol dm^{-3}	Initial concentration of **E** / mol dm^{-3}	Initial rate /mol dm^{-3} s^{-1}
3	0.13	0.23	0.26×10^{-3}
4	0.39	0.23	2.34×10^{-3}
5	0.78	0.46	9.36×10^{-3}

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

(1)

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

(1)

(c) The compound $(CH_3)_3CBr$ reacts with aqueous sodium hydroxide as shown in the folfollowing equation.

$$(CH_3)_3CBr + OH^- \longrightarrow (CH_3)_3COH + Br^-$$

This reaction was found to be first order with respect to $(CH_3)_3CBr$ but zero order with respect to hydroxide ions.

The following two-step process was suggested.

Step **1** $(CH_3)_3CBr \longrightarrow (CH_3)_3C^+ + Br^-$

Step **2** $(CH_3)_3C^+ + OH^- \longrightarrow (CH_3)_3COH$

(i) Deduce the rate-determining step in this two-step process.

(1)

(ii) Outline a mechanism for this step using a curly arrow.

(1)

(Total 8 marks)

Q19.

The initial rate of the reaction between two gases **P** and **Q** was measured in a series of experiments at a constant temperature. The following rate equation was determined.

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

(a) 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	$1.8 = 10^{-3}$
2	0.40	0.60	
3	0.60		$5.4 = 10^{-3}$
4		0.90	$12.2 = 10^{-3}$

(Space for working) _____

_____

_____

_____

(3)

(b) Use the data from Experiment **1** to calculate a value for the rate constant *k* and deduce its units.

Calculation _____

_____

_____

_____

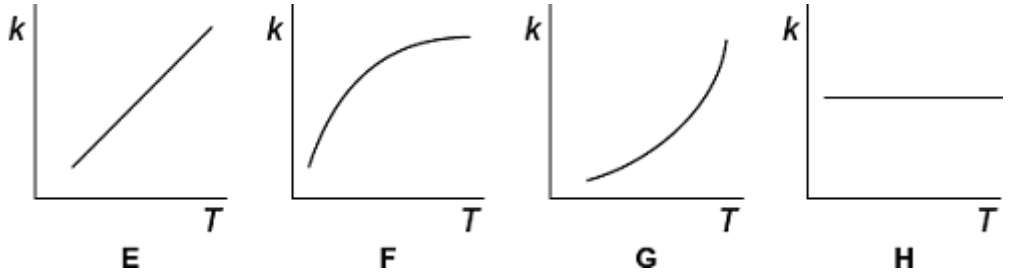
_____

Units _____

_____

(3)

(c) Consider the graphs **E**, **F**, **G** and **H** below.



Write in the box below the letter of the graph that shows how the rate constant *k* varies with temperature.

(1)

(Total 7 marks)

Q20.

(a) In the presence of the catalyst rhodium, the reaction between NO and H_2 occurs according to the following equation.

$$2NO(g) + 2H_2(g) \longrightarrow N_2(g) + 2H_2O(g)$$

The kinetics of the reaction were investigated and the rate equation was found to be

$$\text{rate} = k[NO]^2[H_2]$$

The initial rate of reaction was 6.2×10^{-6} mol dm^{-3} s^{-1} when the initial concentration of NO was 2.9×10^{-2} mol dm^{-3} and the initial concentration of H_2 was 2.3×10^{-2} mol dm^{-3}.

(i) Calculate the value of the rate constant under these conditions and give its units.

Calculation _____

_____

_____

Units _____

_____

(3)

(ii) Calculate the initial rate of reaction if the experiment is repeated under the same conditions but with the concentrations of NO and of H_2 both doubled from their original values.

_____

_____

(1)

(b) Using the rate equation and the overall equation, the following three-step mechanism for the reaction was suggested. X and Y are intermediate species.

Step **1** $NO + NO \longrightarrow X$

Step **2** $X + H_2 \longrightarrow Y$

Step **3** $Y + H_2 \longrightarrow N_2 + 2H_2O$

Suggest which **one** of the three steps is the rate-determining step.

Explain your answer.

Rate-determining step _____

Explanation _____

_____

_____

(2)

(Total 6 marks)

Q21.

The rate of hydrolysis of an ester **X** ($HCOOCH_2CH_2CH_3$) was studied in alkaline conditions at a given temperature. The rate was found to be first order with respect to the ester and first order with respect to hydroxide ions.

(a) (i) Name ester **X**.

_____

(1)

(ii) Using **X** to represent the ester, write a rate equation for this hydrolysis reaction.

_____

(1)

(iii) When the initial concentration of **X** was 0.024 mol dm^{-3} and the initial concentration of hydroxide ions was 0.035 mol dm^{-3}, the initial rate of the reaction was 8.5×10^{-5} mol dm^{-3} s^{-1}.
Calculate a value for the rate constant at this temperature and give its units.

Calculation _____

_____

_____

_____

Units _____

_____

(3)

(iv) In a second experiment at the same temperature, water was added to the original reaction mixture so that the total volume was doubled.
Calculate the initial rate of reaction in this second experiment.

_____

_____

(1)

(v) In a third experiment at the same temperature, the concentration of **X** was half that used in the experiment in part (a) (iii) and the concentration of hydroxide ions was three times the original value.
Calculate the initial rate of reaction in this third experiment.

_____

_____

(1)

(vi) State the effect, if any, on the value of the rate constant *k* when the temperature is lowered but all other conditions are kept constant. Explain your answer.

Effect _____

Explanation _____

(2)

(b) Compound **A** reacts with compound **B** as shown by the overall equation

$$A + 3B \rightarrow AB_3$$

The rate equation for the reaction is

$$\text{rate} = k[A][B]^2$$

A suggested mechanism for the reaction is

Step **1** $A + B \rightarrow AB$

Step **2** $AB + B \rightarrow AB_2$

Step **3** $AB_2 + B \rightarrow AB_3$

Deduce which one of the three steps is the rate-determining step.

Explain your answer.

Rate-determining step _____

Explanation _____

_____

(2)

(Total 11 marks)

Q22.

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

A graph of $\log_{10}(\frac{1}{t})$ 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	$\log_{10}(\frac{1}{t})$
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 $\log_{10}(\frac{1}{t})$ 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.

_____

_____

_____

(3)

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

_____

(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 _____

_____

Explanation _____

_____

(2)

(Total 11 marks)

Q23.

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.

_____

(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 _____

_____

_____

_____

Units _____

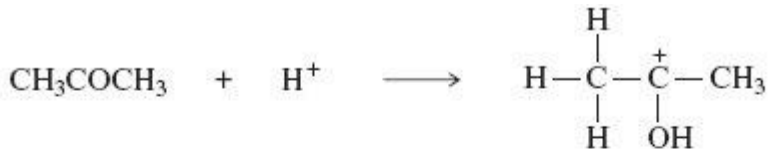
(3)

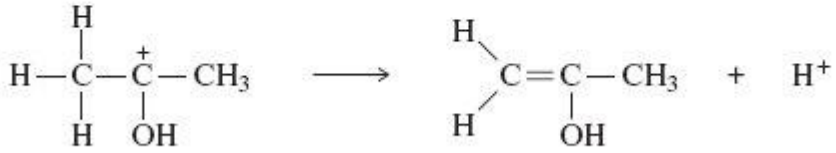
(c) Deduce how the initial rate of reaction changes when the concentration of iodine is doubled but the concentrations of propanone and of hydrochloric acid are unchanged.

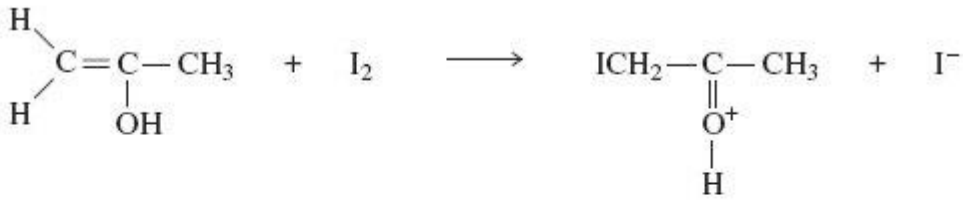
_____

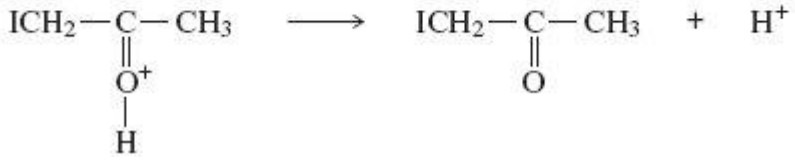
(1)

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

Step 1 

Step 2 

Step 3 

Step 4 

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

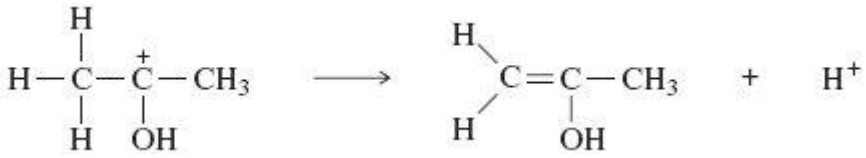
Rate-determining step _____

Explanation _____

_____

(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.

Step 2 

(1)

(Total 8 marks)

Q24.

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**.

_____

_____

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

_____

_____

(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 _____

_____

Units of k _____

_____

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

_____

_____

(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?

_____

(5)

(Total 7 marks)

Q25.

(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.

_____

_____

_____

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

_____

(6)

(Total 8 marks)

Mark Scheme

Q18.

(a) The data in the following table were obtained in two experiments about the rate of the reaction between substances **B** and **C** at a constant temperature.

Experiment	Initial concentration of **B** / mol dm^{-3}	Initial concentration of **C** / mol dm^{-3}	Initial rate / mol dm^{-3} s^{-1}
1	4.2×10^{-2}	2.6×10^{-2}	8.4×10^{-5}
2	6.3×10^{-2}	7.8×10^{-2}	To be calculated

The rate equation for this reaction is known to be

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

(i) Use the data from Experiment **1** to calculate a value for the rate constant k at this temperature and deduce its units.

(3)

(ii) Calculate a value for the initial rate in Experiment **2**.

(1)

(b) The data in the following table were obtained in a series of experiments about the rate of the reaction between substances **D** and **E** at a constant temperature.

Experiment	Initial concentration of **D** / mol dm^{-3}	Initial concentration of **E** / mol dm^{-3}	Initial rate /mol dm^{-3} s^{-1}
3	0.13	0.23	0.26×10^{-3}
4	0.39	0.23	2.34×10^{-3}
5	0.78	0.46	9.36×10^{-3}

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

(1)

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

(1)

(c) The compound $(CH_3)_3CBr$ reacts with aqueous sodium hydroxide as shown in the folfollowing equation.

$$(CH_3)_3CBr + OH^- \longrightarrow (CH_3)_3COH + Br^-$$

This reaction was found to be first order with respect to $(CH_3)_3CBr$ but zero order with respect to hydroxide ions.

The following two-step process was suggested.

Step **1** $(CH_3)_3CBr \longrightarrow (CH_3)_3C^+ + Br^-$

Step **2** $(CH_3)_3C^+ + OH^- \longrightarrow (CH_3)_3COH$

(i) Deduce the rate-determining step in this two-step process.

(1)

(ii) Outline a mechanism for this step using a curly arrow.

(1)

(Total 8 marks)

Q19.

The initial rate of the reaction between two gases **P** and **Q** was measured in a series of experiments at a constant temperature. The following rate equation was determined.

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

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

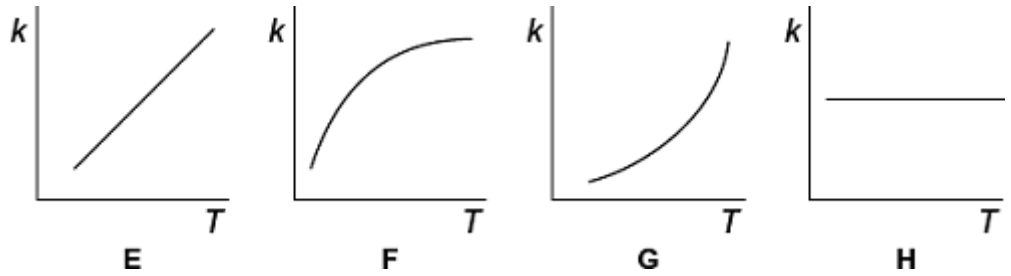
Experiment	Initial [**P**] /mol dm^{-3}	Initial [**Q**] /mol dm^{-3}	Initial rate /mol dm^{-3} s^{-1}
1	0.20	0.30	$1.8 = 10^{-3}$
2	0.40	0.60	
3	0.60		$5.4 = 10^{-3}$
4		0.90	$12.2 = 10^{-3}$

(3)

(b) Use the data from Experiment **1** to calculate a value for the rate constant *k* and deduce its units.

(3)

(c) Consider the graphs **E**, **F**, **G** and **H** below.



Write in the box below the letter of the graph that shows how the rate constant *k* varies with temperature.

(1)

(Total 7 marks)

Q20.

(a) In the presence of the catalyst rhodium, the reaction between NO and H_2 occurs according to the following equation.

$$2NO(g) + 2H_2(g) \longrightarrow N_2(g) + 2H_2O(g)$$

The kinetics of the reaction were investigated and the rate equation was found to be

$$\text{rate} = k[NO]^2[H_2]$$

The initial rate of reaction was 6.2×10^{-6} mol dm^{-3} s^{-1} when the initial concentration of NO was 2.9×10^{-2} mol dm^{-3} and the initial concentration of H_2 was 2.3×10^{-2} mol dm^{-3}.

(i) Calculate the value of the rate constant under these conditions and give its units.

(3)

(ii) Calculate the initial rate of reaction if the experiment is repeated under the same conditions but with the concentrations of NO and of H_2 both doubled from their original values.

(1)

(b) Using the rate equation and the overall equation, the following three-step mechanism for the reaction was suggested. X and Y are intermediate species.

Step **1** $NO + NO \longrightarrow X$

Step **2** $X + H_2 \longrightarrow Y$

Step **3** $Y + H_2 \longrightarrow N_2 + 2H_2O$

Suggest which **one** of the three steps is the rate-determining step.

Explain your answer.

(2)

(Total 6 marks)

Q21.

The rate of hydrolysis of an ester **X** ($HCOOCH_2CH_2CH_3$) was studied in alkaline conditions at a given temperature. The rate was found to be first order with respect to the ester and first order with respect to hydroxide ions.

(a) (i) Name ester **X**.

(1)

(ii) Using **X** to represent the ester, write a rate equation for this hydrolysis reaction.

(1)

(iii) When the initial concentration of **X** was 0.024 mol dm^{-3} and the initial concentration of hydroxide ions was 0.035 mol dm^{-3}, the initial rate of the reaction was 8.5×10^{-5} mol dm^{-3} s^{-1}.
Calculate a value for the rate constant at this temperature and give its units.

(3)

(iv) In a second experiment at the same temperature, water was added to the original

reaction mixture so that the total volume was doubled.
Calculate the initial rate of reaction in this second experiment.

(1)

(v) In a third experiment at the same temperature, the concentration of **X** was half that used in the experiment in part (a) (iii) and the concentration of hydroxide ions was three times the original value.
Calculate the initial rate of reaction in this third experiment.

(1)

(vi) State the effect, if any, on the value of the rate constant *k* when the temperature is lowered but all other conditions are kept constant. Explain your answer.

(2)

(b) Compound **A** reacts with compound **B** as shown by the overall equation

$$A + 3B \rightarrow AB_3$$

The rate equation for the reaction is

$$\text{rate} = k[A][B]^2$$

A suggested mechanism for the reaction is

Step **1** $A + B \rightarrow AB$

Step **2** $AB + B \rightarrow AB_2$

Step **3** $AB_2 + B \rightarrow AB_3$

Deduce which one of the three steps is the rate-determining step.

Explain your answer.

(2)
(Total 11 marks)

Q22.

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

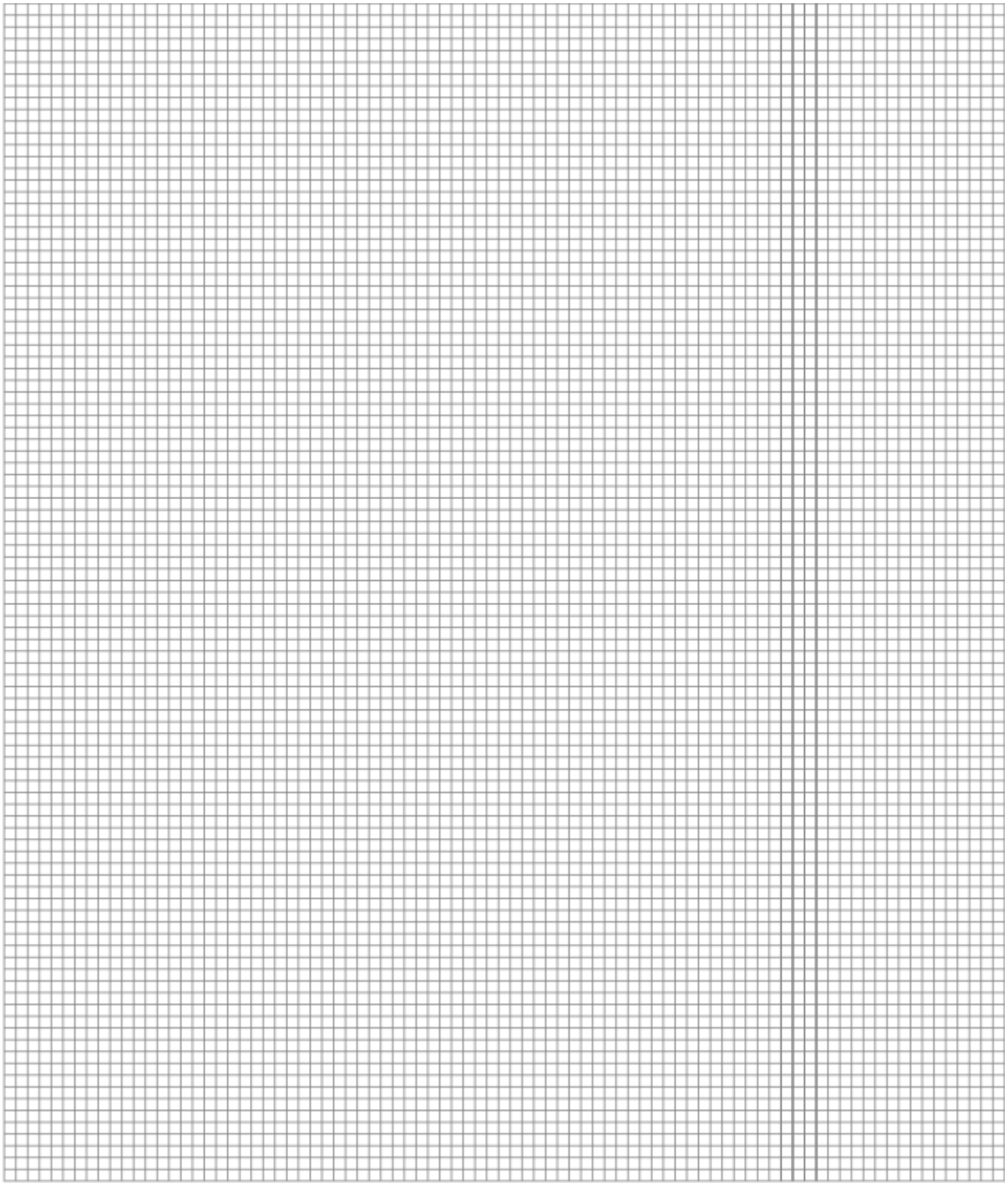
A graph of $\log_{10}\left(\frac{1}{t}\right)$ on the *y*-axis against $\log_{10}$ (volume of KI(aq)) is a straight line. The gradient of this straight line is equal to the order of the reaction with respect to iodide ions.

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.

(3)

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

(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.

(2)

(Total 11 marks)

Q23.

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.

(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

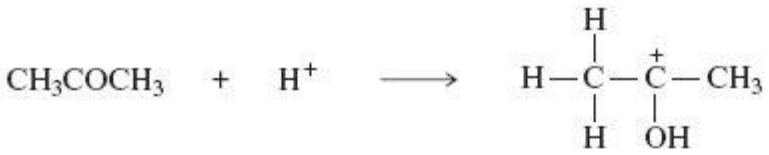
Use these data to calculate a value for the rate constant, *k*, for the reaction and give its units.

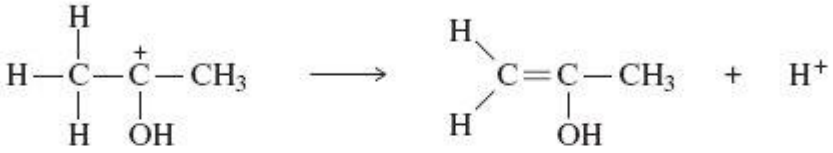
(3)

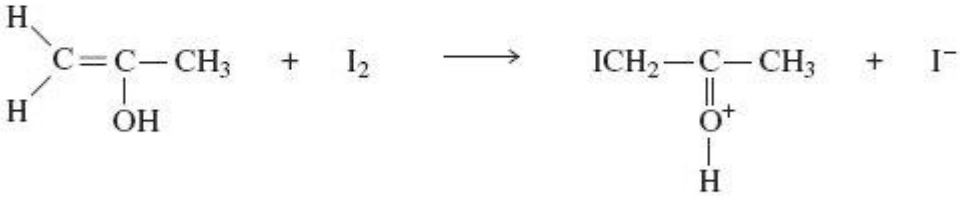
(c) Deduce how the initial rate of reaction changes when the concentration of iodine is doubled but the concentrations of propanone and of hydrochloric acid are unchanged.

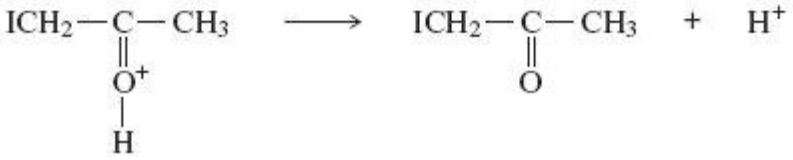
(1)

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

Step 1 

Step 2 

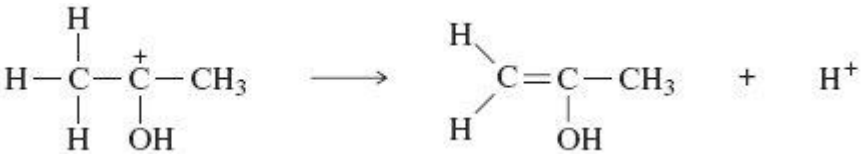
Step 3 

Step 4 

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

(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.

Step 2 

(1)

(Total 8 marks)

Q24.

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**.

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

(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.

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

(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?

(5)

(Total 7 marks)

Q25.

(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**.

(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}.

(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.

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

(6)

(Total 8 marks)