**Q1.** This question involves the use of kinetic data to deduce the order of a reaction and calculate a value for a rate constant.

The data in **Table 1** were obtained in a series of experiments on the rate of the reaction between compounds **A** and **B** at a constant temperature.

**Table 1**

| **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 | 0.26 | $2.10 \times 10^{-4}$ |
| 2 | 0.36 | 0.26 | $1.89 \times 10^{-3}$ |
| 3 | 0.72 | 0.13 | $3.78 \times 10_{-3}$ |

(a) Show how these data can be used to deduce the rate expression for the reaction between **A** and **B**.

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**(3)**

The data in **Table 2** were obtained in two experiments on the rate of the reaction between compounds **C** and **D** at a constant temperature.

**Table 2**

| Experiment | Initial concentration of C / mol $dm^{-3}$ | Initial concentration of D / mol $dm^{-3}$ | Initial rate / mol $dm^{-3}$ $s^{-1}$ |
|---|---|---|---|
| 4 | $1.9 \times 10^{-2}$ | $3.5 \times 10^{-2}$ | $7.2 \times 10^{-4}$ |
| 5 | $3.6 \times 10^{-2}$ | $5.4 \times 10^{-2}$ | To be calculated |

The rate equation for this reaction is

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

(b) Use the data from experiment **4** to calculate a value for the rate constant, *k*, at this temperature. Deduce the units of *k*.

*k* = ..... Units = .....

**(3)**

(c) Calculate a value for the initial rate in experiment **5**.

Initial rate = ..... mol $dm^{-3}$ $s^{-1}$

**(1)**

(d) The rate equation for a reaction is

$rate = k[\mathbf{E}]$

Explain qualitatively why doubling the temperature has a much greater effect on the rate of the reaction than doubling the concentration of **E**.

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**(3)**

(e) A slow reaction has a rate constant $k = 6.51 \times 10^{-3}$ mol$^{-1}$ dm$^3$ at 300 K.

Use the equation $\ln k = \ln A - E_a / RT$ to calculate a value, in kJ mol$^{-1}$, for the activation energy of this reaction.

The constant $A = 2.57 \times 10^{10}$ mol$^{-1}$ dm$^3$.
The gas constant $R = 8.31$ J K$^{-1}$ mol$^{-1}$.

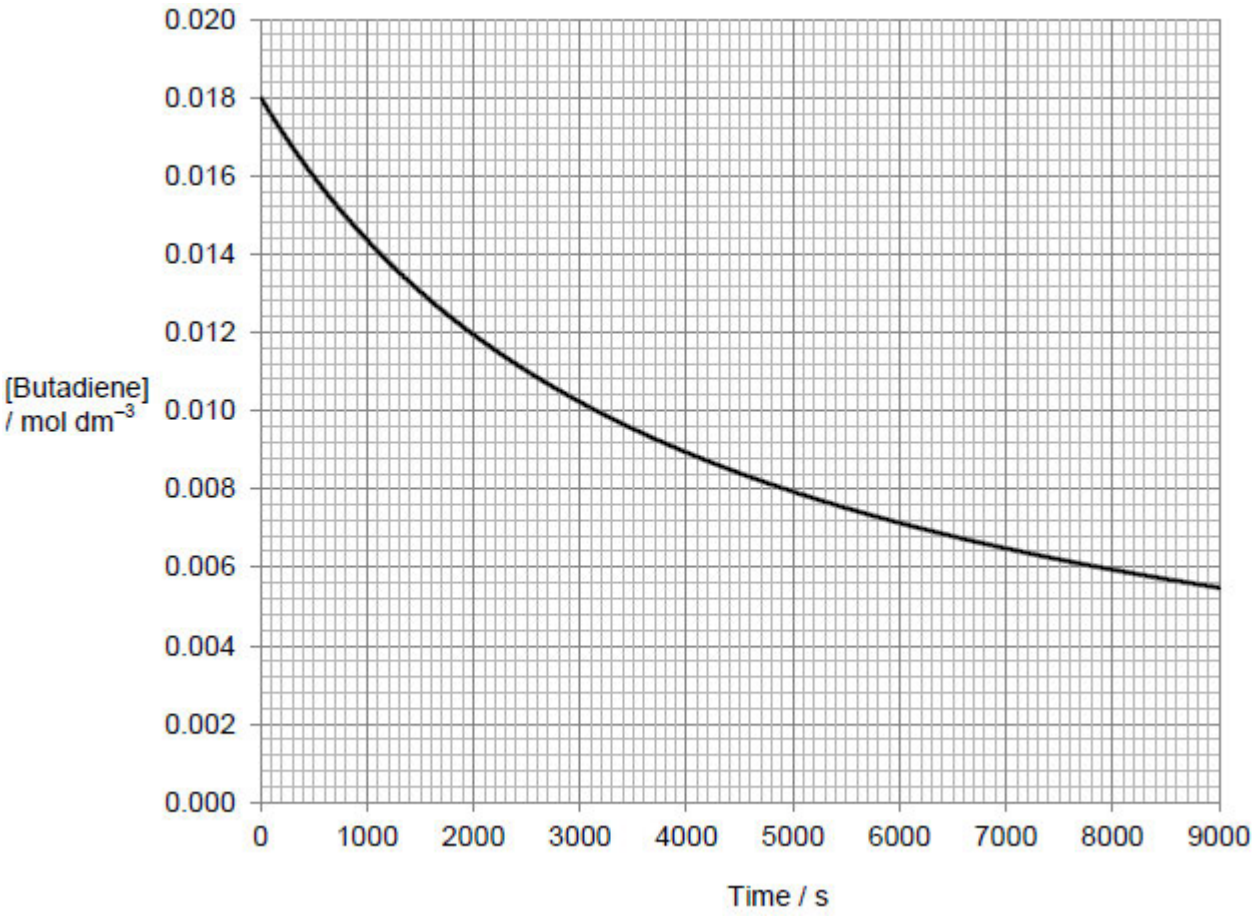
Activation energy = .....

**(2)**
**(Total 12 marks)**

**Q2.**Butadiene dimerises according to the equation

$$2C_4H_6 \longrightarrow C_8H_{12}$$

The kinetics of the dimerisation are studied and the graph of the concentration of a sample of butadiene is plotted against time. The graph is shown below.




(a) Draw a tangent to the curve when the concentration of butadiene is 0.0120 mol dm$^{-3}$.

**(1)**

(b) The initial rate of reaction in this experiment has the value $4.57 \times 10^{-6}$ mol dm$^{-3}$ s$^{-1}$.

Use this value, together with a rate obtained from your tangent, to justify that the order of the reaction is 2 with respect to butadiene.

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**(5)**
**(Total 6 marks)**

**Q3.** 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.

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**(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
$8.5 \times 10^{-5}$ mol $dm^{-3}$ $s^{-1}$.
Calculate a value for the rate constant at this temperature and give its units.

Calculation .....

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Units .....

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**(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.

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**(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.

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**(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)**

**Q4.** (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 $6.2 \times 10^{-6}$ mol $dm^{-3}$ $s^{-1}$ when the initial concentration of NO was $2.9 \times 10^{-2}$ mol $dm^{-3}$ and the initial concentration of $H_2$ was $2.3\times 10^{-2}$ mol $dm^{-3}$.

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

Calculation .....

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Units .....

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**(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.

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**(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 .....

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*(Extra space)* .....

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**(2)**
**(Total 6 marks)**

**Q5.** This question involves the use of kinetic data to calculate the order of a reaction and also a value for a rate constant.

(a) The data in this table were obtained in a series of experiments on the rate of the reaction between compounds **E** and **F** at a constant temperature.

| Experiment | Initial concentration of **E** / mol $dm^{-3}$ | Initial concentration of **F** / mol $dm^{-3}$ | Initial rate of reaction / mol $dm^{-3}$ $s^{-1}$ |
|---|---|---|---|
| **1** | 0.15 | 0.24 | $0.42 \times 10^{-3}$ |
| **2** | 0.45 | 0.24 | $3.78 \times 10^{-3}$ |
| **3** | 0.90 | 0.12 | $7.56 \times 10^{-3}$ |

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

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*(Space for working)* .....

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**(1)**

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

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*(Space for working)* .....

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**(1)**

(b) The data in the following table were obtained in two experiments on the rate of the reaction between compounds **G** and **H** at a constant temperature.

| Experiment | Initial concentration of **G** / mol $dm^{-3}$ | Initial concentration of **H** / mol $dm^{-3}$ | Initial rate of reaction / mol $dm^{-3}$ $s^{-1}$ |
|---|---|---|---|
| **4** | $3.8 \times 10^{-2}$ | $2.6 \times 10^{-2}$ | $8.6 \times 10^{-4}$ |
| **5** | $6.3 \times 10^{-2}$ | $7.5 \times 10^{-2}$ | To be calculated |

The rate equation for this reaction is

$$rate = \mathbf{k}[\mathbf{G}]^2[\mathbf{H}]$$

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

Calculation .....

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Units .....

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**(3)**

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

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**(1)**
**(Total 6 marks)**

**Q6.**(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** | $4.2 \times 10^{-2}$ | $2.6 \times 10^{-2}$ | $8.4 \times 10^{-5}$ |
| **2** | $6.3 \times 10^{-2}$ | $7.8 \times 10^{-2}$ | To be calculated |

The rate equation for this reaction is known to be

rate = $k$[**B**]$^2$[**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 .....

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Units .....

.....

*(Extra space)* .....

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**(3)**

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

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**(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 | $0.26 \times 10^{-3}$ |
| **4** | 0.39 | 0.23 | $2.34 \times 10^{-3}$ |
| **5** | 0.78 | 0.46 | $9.36 \times 10^{-3}$ |

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

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**(1)**

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

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**(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)**

**Q7.** 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)* .....

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**(3)**

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

Calculation .....

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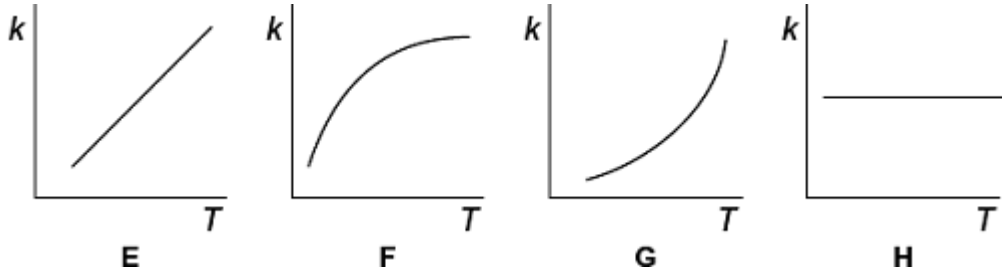
.....

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)**

**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.

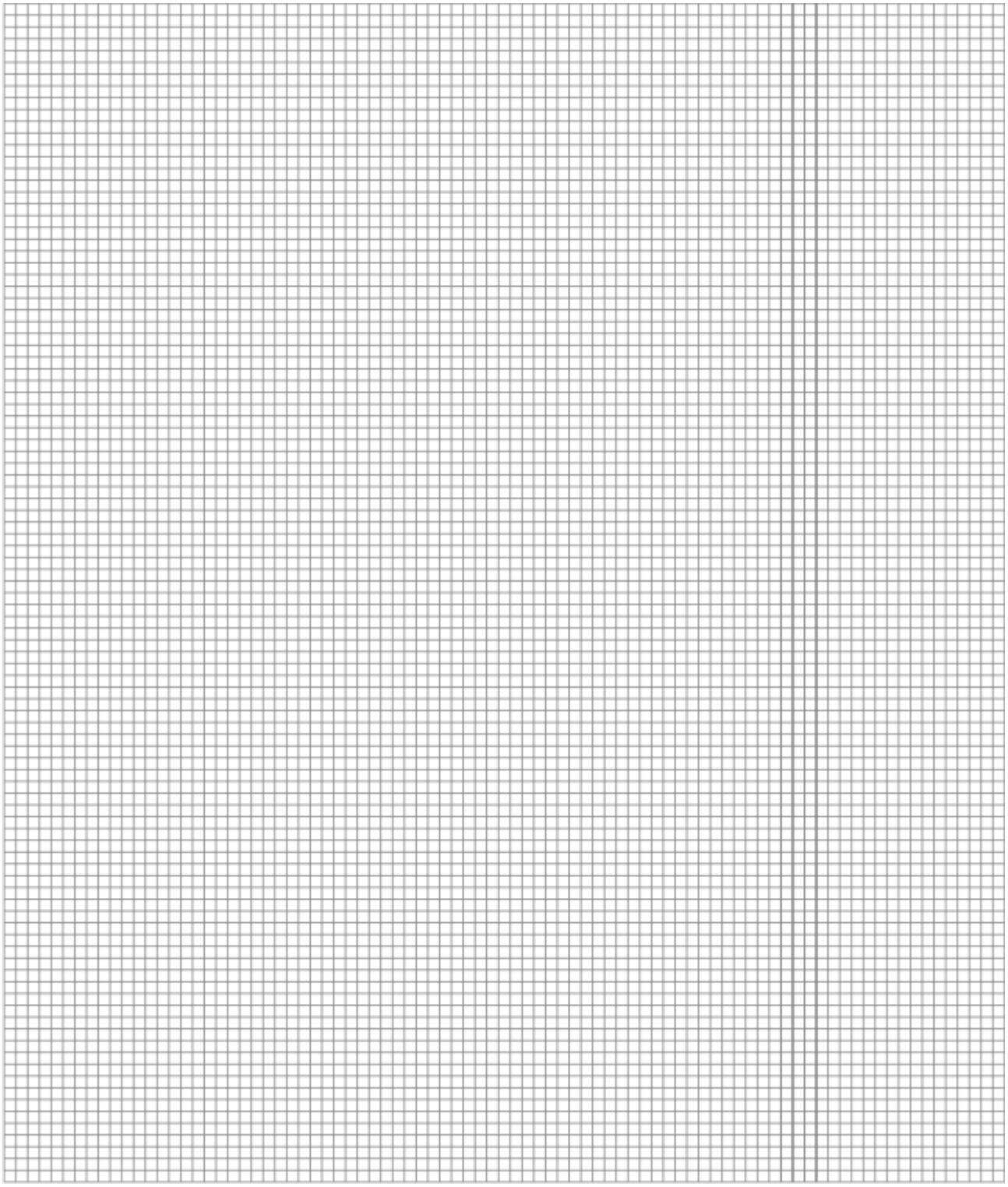
A graph of $\log_{10}\left(\frac{1}{t}\right)$ on the $y$-axis against $\log_{10}$ (volume of KI(aq)) is a straight line. The gradient of this straight line is equal to the order of the reaction with respect to iodide ions.

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.

.....

**(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)**

**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 | $4.8 \times 10^{-3}$ |
| 2 | 0.10 | 0.10 | |
| 3 | 0.40 | | $9.6 \times 10^{-3}$ |
| 4 | | 0.60 | $19.2 \times 10^{-3}$ |

(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 | $0.45 \times 10^{-5}$ |
| 2 | 0.30 | 0.24 | $0.90 \times 10^{-5}$ |
| 3 | 0.60 | 0.48 | $7.20 \times 10^{-5}$ |

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

.....

.....

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

.....

.....

**(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 | $9.30 \times 10^{-5}$ |
| 5 | 0.20 | 0.10 | To be calculated |

The rate equation for this reaction is

rate = $k$[**C**]$^2$[**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* .....

.....

.....

*Units of k* .....

.....

(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 $dm^{-3} s^{-1}$ | $7.6 \times 10^{-3}$ | $1.9 \times 10^{-3}$ | $3.8 \times 10^{-3}$ |

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 $4.2 \times 10^{-4}$ $mol^{-2}$ $dm^6$ $s^{-1}$.

(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 $8.1 \times 10^{-5}$ 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.

.....

**(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.

.....

**(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* .....

.....

**(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.

.....

**(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 $1.24 \times 10^{-4}$ mol $dm^{-3}$ $s^{-1}$.

| | initial concentration / mol $dm^{-3}$ |
|---|---|
| $CH_3COCH_3$ | 4.40 |
| $I_2$ | $5.00 \times 10^{-3}$ |
| $H^+$ | 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 .....

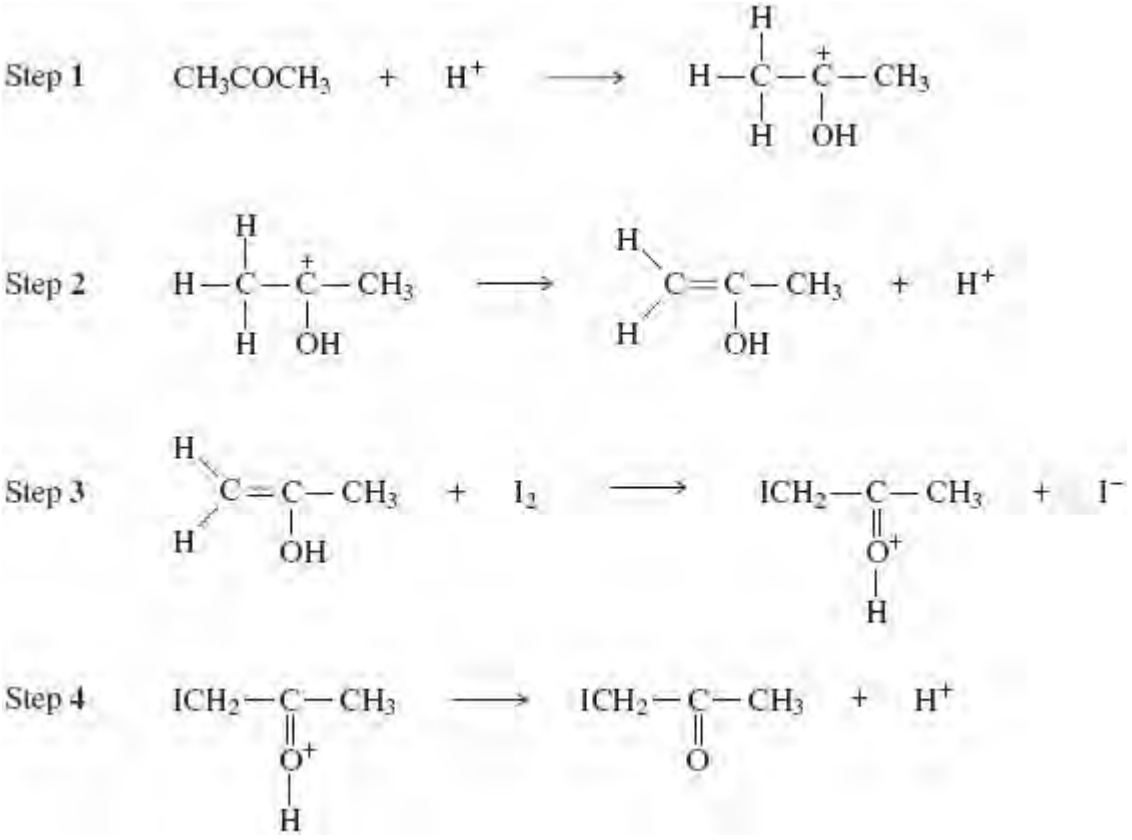
**(3)**

(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.

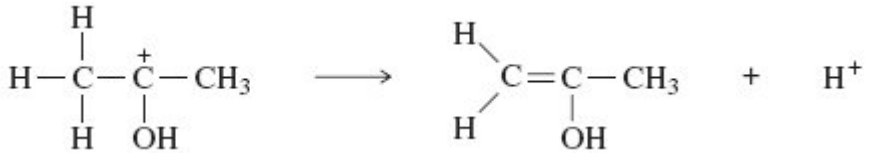
Rate-determining step .....

Explanation .....

.....

**(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)**

**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 | $0.32 \times 10^{-3}$ |
| **2** | 0.36 | 0.15 | $2.88 \times 10^{-3}$ |
| **3** | 0.72 | 0.30 | $11.52 \times 10^{-3}$ |

(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** | $5.0 \times 10^{-2}$ | $2.0 \times 10^{-2}$ | $6.5 \times 10^{-4}$ |
| **5** | $6.5 \times 10^{-2}$ | $3.4 \times 10^{-2}$ | 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)**

**Q1.** Gases **A** and **B** react as shown in the following equation.

$$2\mathbf{A}(g) + \mathbf{B}(g) \longrightarrow \mathbf{C}(g) + \mathbf{D}(g)$$

The initial rate of the reaction was measured in a series of experiments at a constant temperature. The following rate equation was determined.

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

An incomplete table of data for the reaction between **A** and **B** is shown in the table.

| **Experiment** | **Initial [A] / mol dm$^{-3}$** | **Initial [B] / mol dm$^{-3}$** | **Initial rate / mol dm$^{-3}$ s$^{-1}$** |
|---|---|---|---|
| **1** | $4.2 \times 10^{-3}$ | $2.8 \times 10^{-3}$ | $3.3 \times 10^{-5}$ |
| **2** | $7.9 \times 10^{-3}$ | $2.8 \times 10^{-3}$ | |
| **3** | | $5.6 \times 10^{-3}$ | $1.8 \times 10^{-4}$ |

(a) Use the data from Experiment **1** to calculate a value for the rate constant, *k*, at this temperature.
Deduce the units of *k*.

Calculation .....

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.....

.....

.....

Units .....

.....

**(3)**

(b) Use your value of *k* from (a) to complete the table for the reaction between **A** and **B**. (If you have been unable to calculate an answer for (a), you may assume a value of 2.3. This is **not** the correct answer.)

**(2)**

(c) The reaction is zero order with respect to **B**.

State the significance of this zero order for the mechanism of the reaction.

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**(1)**
**(Total 6 marks)**

**Q2.** (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 $6.2 \times 10^{-6}$ mol $dm^{-3}$ $s^{-1}$ when the initial concentration of NO was $2.9 \times 10^{-2}$ mol $dm^{-3}$ and the initial concentration of $H_2$ was $2.3\times 10^{-2}$ 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 .....

.....

.....

*(Extra space)* .....

.....

**(2)**
**(Total 6 marks)**

**Q3.** This question involves the use of kinetic data to deduce the order of a reaction and calculate a value for a rate constant.

The data in **Table 1** were obtained in a series of experiments on the rate of the reaction between compounds **A** and **B** at a constant temperature.

**Table 1**

| **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 | 0.26 | $2.10 \times 10^{-4}$ |
| 2 | 0.36 | 0.26 | $1.89 \times 10^{-3}$ |
| 3 | 0.72 | 0.13 | $3.78 \times 10_{-3}$ |

(a) Show how these data can be used to deduce the rate expression for the reaction

between **A** and **B**.

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**(3)**

The data in **Table 2** were obtained in two experiments on the rate of the reaction between compounds **C** and **D** at a constant temperature.

**Table 2**

| **Experiment** | **Initial concentration of C / mol dm$^{-3}$** | **Initial concentration of D / mol dm$^{-3}$** | **Initial rate / mol dm$^{-3}$ s$^{-1}$** |
|---|---|---|---|
| 4 | $1.9 \times 10^{-2}$ | $3.5 \times 10^{-2}$ | $7.2 \times 10^{-4}$ |
| 5 | $3.6 \times 10^{-2}$ | $5.4 \times 10^{-2}$ | To be calculated |

The rate equation for this reaction is

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

(b) Use the data from experiment **4** to calculate a value for the rate constant, *k*, at this temperature. Deduce the units of *k*.

$k$ = ..... Units = .....

**(3)**

(c) Calculate a value for the initial rate in experiment **5**.

Initial rate = ..... mol $dm^{-3}$ $s^{-1}$

**(1)**

(d) The rate equation for a reaction is

$$rate = k[\mathbf{E}]$$

Explain qualitatively why doubling the temperature has a much greater effect on the rate of the reaction than doubling the concentration of **E**.

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**(3)**

(e) A slow reaction has a rate constant $k = 6.51 \times 10^{-3}$ $mol^{-1}$ $dm^3$ at 300 K.

Use the equation $\ln k = \ln A - E_a / RT$ to calculate a value, in kJ $mol^{-1}$, for the activation energy of this reaction.

The constant $A = 2.57 \times 10^{10}$ $mol^{-1}$ $dm^3$.
The gas constant $R = 8.31$ J $K^{-1}$ $mol^{-1}$.

Activation energy = .....

**(2)**
**(Total 12 marks)**

**Q4.** 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)* .....

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**(3)**

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

Calculation .....

.....

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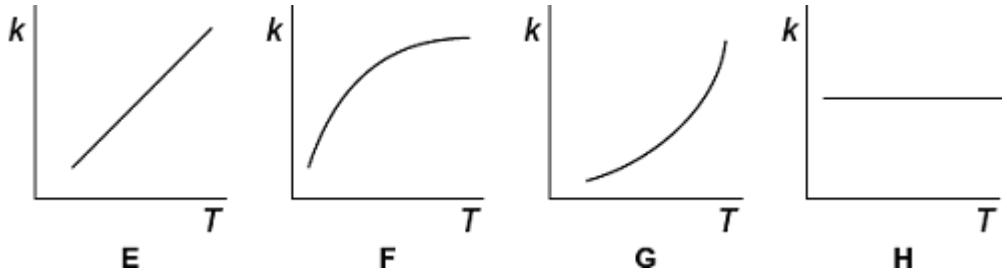
.....

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)**

**Q5.** Gases **P** and **Q** react as shown in the following equation.

$$2P(g) + 2Q(g) \longrightarrow R(g) + S(g)$$

The initial rate of the reaction 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 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 | $2.5 \times 10^{-2}$ | $1.8 \times 10^{-2}$ | $5.0 \times 10^{-5}$ |
| 2 | $7.5 \times 10^{-2}$ | $1.8 \times 10^{-2}$ | |
| 3 | $5.0 \times 10^{-2}$ | | $5.0 \times 10^{-5}$ |
| 4 | | $5.4 \times 10^{-2}$ | $4.5 \times 10^{-4}$ |

*(Space for working)* .....

.....

.....

.....

**(3)**

(b) Use the data from Experiment **1** to calculate a value for the rate constant (*k*) at this temperature. Deduce the units of *k*.

Calculation .....

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Units .....

.....

**(3)**
**(Total 6 marks)**

**Q6.**(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** | $4.2 \times 10^{-2}$ | $2.6 \times 10^{-2}$ | $8.4 \times 10^{-5}$ |
| **2** | $6.3 \times 10^{-2}$ | $7.8 \times 10^{-2}$ | To be calculated |

The rate equation for this reaction is known to be

rate = $k$[**B**]$^2$[**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 .....

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.....

.....

.....

Units .....

.....

*(Extra space)* .....

.....

**(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 | $0.26 \times 10^{-3}$ |
| **4** | 0.39 | 0.23 | $2.34 \times 10^{-3}$ |
| **5** | 0.78 | 0.46 | $9.36 \times 10^{-3}$ |

(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)**

**Q7.**(a) The table shows the results of three experiments to investigate the rate of reaction between compounds **A** and **B** dissolved in a given solvent.
All three experiments were carried out at the same temperature.

| | **Experiment 1** | **Experiment 2** | **Experiment 3** |
|---|---|---|---|
| **Initial concentration of A / mol dm$^{-3}$** | $1.60 \times 10^{-2}$ | $2.40 \times 10^{-2}$ | $3.60 \times 10^{-2}$ |
| **Initial concentration of B / mol dm$^{-3}$** | $4.20 \times 10^{-2}$ | $6.30 \times 10^{-2}$ | $6.30 \times 10^{-2}$ |
| **Initial rate /mol dm$^{-3}$ s$^{-1}$** | $8.00 \times 10^{-5}$ | $1.80 \times 10^{-4}$ | $4.05 \times 10^{-4}$ |

(i) Deduce the order of reaction with respect to **A**.
Tick (✓) **one** box.

| **Order of reaction with respect to A** | **Tick (✓)** |
|---|---|
| 0 | |
| 1 | |
| 2 | |

**(1)**

(ii) Deduce the order of reaction with respect to **B**.
Tick (✓) **one** box.

| **Order of reaction with respect to B** | **Tick (✓)** |
|---|---|

| | |
|---|---|
| 0 | |
| 1 | |
| 2 | |

**(1)**

(b) The reaction between two different compounds, **C** and **D**, is studied at a given temperature.
The rate equation for the reaction is found to be

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

(i) When the initial concentration of **C** is $4.55 \times 10^{-2}$ mol $dm^{-3}$ and the initial concentration of **D** is $1.70 \times 10^{-2}$ mol $dm^{-3}$, the initial rate of reaction is $6.64 \times 10^{-5}$ mol $dm^{-3}$ $s^{-1}$.

Calculate the value of the rate constant at this temperature and deduce its units.

Calculation .....

.....

.....

.....

Units of rate constant .....

.....

**(3)**

(ii) The experiment in part (i) is repeated at the same temperature but after the addition of extra solvent so that the total volume of the mixture is doubled.

Deduce the new initial rate of reaction.

.....

.....

**(1)**
**(Total 6 marks)**

**Q8.** This question involves the use of kinetic data to calculate the order of a reaction and also a value for a rate constant.

(a) The data in this table were obtained in a series of experiments on the rate of the reaction between compounds **E** and **F** at a constant temperature.

| Experiment | Initial concentration of **E** / mol $dm^{-3}$ | Initial concentration of **F** / mol $dm^{-3}$ | Initial rate of reaction / mol $dm^{-3}$ $s^{-1}$ |
|---|---|---|---|
| **1** | 0.15 | 0.24 | $0.42 \times 10^{-3}$ |
| **2** | 0.45 | 0.24 | $3.78 \times 10^{-3}$ |
| **3** | 0.90 | 0.12 | $7.56 \times 10^{-3}$ |

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

.....

*(Space for working)* .....

.....

.....

.....

**(1)**

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

.....

*(Space for working)* .....

.....

.....

.....

**(1)**

(b) The data in the following table were obtained in two experiments on the rate of the reaction between compounds **G** and **H** at a constant temperature.

| Experiment | Initial concentration of **G** / mol $dm^{-3}$ | Initial concentration of **H** / mol $dm^{-3}$ | Initial rate of reaction |
|---|---|---|---|

| | | | / mol $dm^{-3}$ $s^{-1}$ |
|---|---|---|---|
| **4** | $3.8 \times 10^{-2}$ | $2.6 \times 10^{-2}$ | $8.6 \times 10^{-4}$ |
| **5** | $6.3 \times 10^{-2}$ | $7.5 \times 10^{-2}$ | To be calculated |

The rate equation for this reaction is

*rate* = **k**[**G**]$^2$[**H**]

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

Calculation .....

.....

.....

.....

Units .....

.....

**(3)**

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

.....

.....

**(1)**
**(Total 6 marks)**

**Q9.** The rate equation for the hydrogenation of ethene

$$C_2H_4(g) + H_2(g) \longrightarrow C_2H_6(g)$$

is *Rate* = $k[C_2H_4][H_2]$

At a fixed temperature, the reaction mixture is compressed to triple the original pressure.

What is the factor by which the rate of reaction changes?

**A** 6 

**B** 9 ⬭

**C** 12 ⬭

**D** 27 ⬭

**(Total 1 mark)**

**Q1.** 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
$8.5 \times 10^{-5}$ 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.

.....

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**(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)**

**Q2.** A reaction mechanism is a series of steps by which an overall reaction may proceed. The reactions occurring in these steps may be deduced from a study of reaction rates. Experimental evidence about initial rates leads to a rate equation. A mechanism is then proposed which agrees with this rate equation.
Ethanal dimerises in dilute alkaline solution to form compound **X** as shown in the following equation.

$$2CH_3CHO \rightarrow CH_3CH(OH)CH_2CHO$$
**X**

A chemist studied the kinetics of the reaction at 298 K and then proposed the following rate equation.

$$\text{Rate} = k\,[CH_3CHO][OH^-]$$

(a) Give the IUPAC name of compound **X**.

.....

**(1)**

(b) The initial rate of the reaction at 298K was found to be $2.2 \times 10^{-3}$ mol $dm^{-3}$ $s^{-1}$ when the initial concentration of ethanal was 0.10 mol $dm^{-3}$ and the initial concentration of sodium hydroxide was 0.020 mol $dm^{-3}$.
Calculate a value for the rate constant at this temperature and give its units.

Calculation .....

.....

.....

Units .....

**(3)**

(c) The sample of **X** produced consists of a racemic mixture (racemate). Explain how this racemic mixture is formed.

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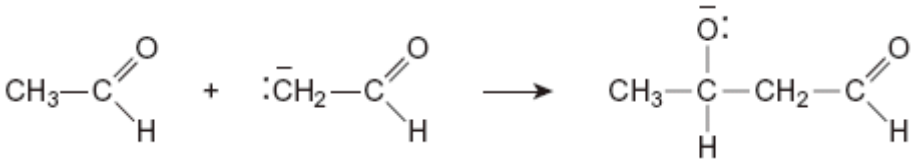
**(2)**

(d) A three-step mechanism has been proposed for this reaction according to the following equations.

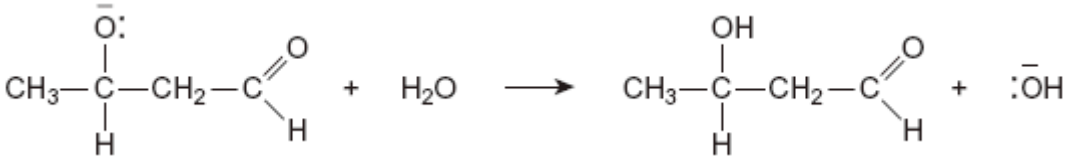
Step **1**



Step **2**



Step **3**



(i) Using the rate equation, predict which of the three steps is the rate-determining step. Explain your answer.

Rate-determining step .....

Explanation .....

.....

**(2)**

(ii) Deduce the role of ethanal in Step **1**.

.....

**(1)**

(iii) Use your knowledge of reaction mechanisms to deduce the type of reaction occurring in Step **2**.

.....

**(1)**

(iv) In the space below draw out the mechanism of Step **2** showing the relevant curly arrows.

**(2)**

(e) In a similar three-step mechanism, one molecule of **X** reacts further with one molecule of ethanal. The product is a trimer containing six carbon atoms.

Deduce the structure of this trimer.

**(1)**
**(Total 13 marks)**

**Q3.** (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 $dm^{-3}$ $s^{-1}$ | $7.6 \times 10^{-3}$ | $1.9 \times 10^{-3}$ | $3.8 \times 10^{-3}$ |

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 $4.2 \times 10^{-4}$ 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 $8.1 \times 10^{-5}$ 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)**

**Q4.** The initial rate of the reaction between the gases NO and $H_2$ was measured in a series of experiments at a constant temperature and the following rate equation was determined.

$$\text{rate} = k[\text{NO}]^2[\text{H}_2]$$

(a) Complete the table of data below for the reaction between NO and $H_2$

| Experiment | Initial [NO] / mol $dm^{-3}$ | Initial [$H_2$] / mol $dm^{-3}$ | Initial rate / mol $dm^{-3}$ $s^{-1}$ |
|---|---|---|---|
| 1 | $3.0 \times 10^{-3}$ | $1.0 \times 10^{-3}$ | $1.8 \times 10^{-5}$ |
| 2 | $3.0 \times 10^{-3}$ | | $7.2 \times 10^{-5}$ |
| 3 | $1.5 \times 10^{-3}$ | $1.0 \times 10^{-3}$ | |
| 4 | | $0.50 \times 10^{-3}$ | $8.1 \times 10^{-5}$ |

**(3)**

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

.....

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.....

.....

**(3)**
**(Total 6 marks)**

**Q5.** The hydrolysis of methyl propanoate was studied in acidic conditions at 25°C and the rate equation was found to be

$$\text{rate} = k[CH_3CH_2COOCH_3][H^+]$$

(a) Use the data below to calculate the value of the rate constant, *k*, at this temperature.
Deduce its units.

| Initial rate of reaction / mol $dm^{-3}$ $s^{-1}$ | Initial concentration of methyl propanoate / mol $dm^{-3}$ | Initial concentration of hydrochloric acid / mol $dm^{-3}$ |
|---|---|---|
| $1.15 \times 10^{-4}$ | 0.150 | 0.555 |

*Rate constant* .....

.....

.....

*Units* .....

.....

**(3)**

(b) The reaction in part (a) was repeated at the same temperature, but water was added so that the volume of the reaction mixture was doubled. Calculate the initial rate of reaction under these conditions.

.....

.....

**(1)**

(c) A third experiment was carried out at a different temperature. Some data from this experiment are shown in the table below.

| Initial rate of reaction / mol $dm^{-3}$ $s^{-1}$ | Value of rate constant at this different temperature | Initial methyl propanoate / mol $dm^{-3}$ |
|---|---|---|
| $4.56 \times 10^{-5}$ | $8.94 \times 10^{-4}$ | 0.123 |

Calculate the initial pH of the reaction mixture. Give your answer to two decimal places.

.....

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.....

.....

**(3)**
**(Total 7 marks)**

**Q6.** 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 | $0.32 \times 10^{-3}$ |
| **2** | 0.36 | 0.15 | $2.88 \times 10^{-3}$ |
| **3** | 0.72 | 0.30 | $11.52 \times 10^{-3}$ |

(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** | $5.0 \times 10^{-2}$ | $2.0 \times 10^{-2}$ | $6.5 \times 10^{-4}$ |
| **5** | $6.5 \times 10^{-2}$ | $3.4 \times 10^{-2}$ | 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)**

**Q7.** 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 $1.24 \times 10^{-4}$ mol $dm^{-3}$ $s^{-1}$.

| | initial concentration / mol $dm^{-3}$ |
|---|---|
| $CH_3COCH_3$ | 4.40 |
| $I_2$ | $5.00 \times 10^{-3}$ |
| $H^+$ | 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 .....

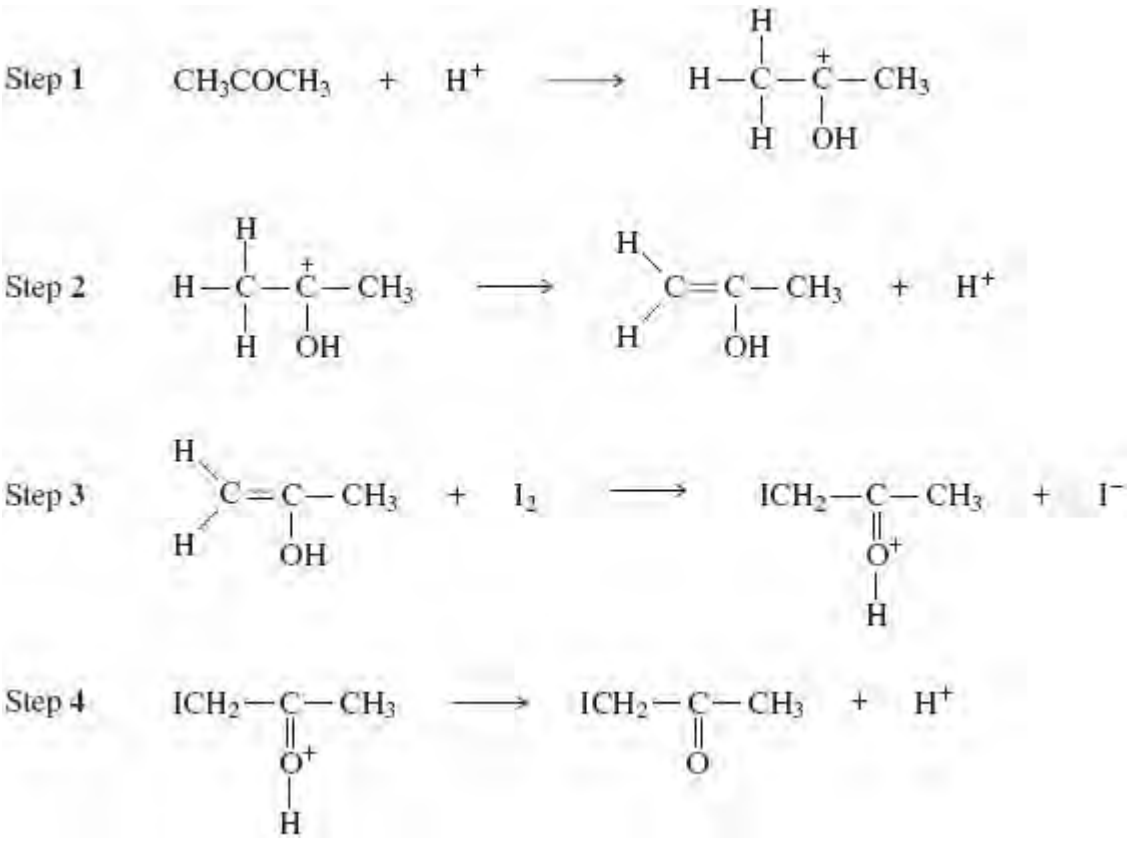
**(3)**

(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.

.....

**(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 .....

.....

**(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)**

**Q1.** (a) Compound **A**, $HCOOCH_2CH_2CH_3$, is an ester. Name this ester and write an equation for its reaction with aqueous sodium hydroxide.

*Name* .....

*Equation* .....

**(2)**

(b) The initial rate of reaction between ester **A** and aqueous sodium hydroxide was measured in a series of experiments at a constant temperature. The data obtained are shown below.

| Experiment | Initial concentration of NaOH / mol $dm^{-3}$ | Initial concentration of **A** / mol $dm^{-3}$ | Initial rate / mol $dm^{-3}$ $s^{-1}$ |
|---|---|---|---|
| 1 | 0.040 | 0.030 | $4.0 \times 10^{-4}$ |
| 2 | 0.040 | 0.045 | $6.0 \times 10^{-4}$ |
| 3 | 0.060 | 0.045 | $9.0 \times 10^{-4}$ |
| 4 | 0.120 | 0.060 | to be calculated |

Use the data in the table to deduce the order of reaction with respect to **A** and the order of reaction with respect to NaOH. Hence calculate the initial rate of reaction in Experiment 4.

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

*Order with respect to NaOH* .....

*Initial rate in Experiment 4* .....

.....

**(3)**

(c) In a further experiment at a different temperature, the initial rate of reaction was found to be $9.0 \times 10^{-3}$ mol $dm^{-3}$ $s^{-1}$ when the initial concentration of **A** was 0.020 mol $dm^{-3}$ and the initial concentration of NaOH was 2.00 mol $dm^{-3}$.
Under these new conditions with the much higher concentration of sodium hydroxide, the reaction is first order with respect to **A** and appears to be zero order with respect to sodium hydroxide.

(i) Write a rate equation for the reaction under these new conditions.

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(ii) Calculate a value for the rate constant under these new conditions and state its units.

*Calculation* .....

.....

.....

Units .....

(iii) Suggest why the order of reaction with respect to sodium hydroxide appears to be zero under these new conditions.

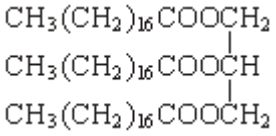
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**(6)**

(d) A naturally-occurring triester, shown below, was heated under reflux with an excess of aqueous sodium hydroxide and the mixture produced was then distilled. One of the products distilled off and the other was left in the distillation flask.



(i) Draw the structure of the product distilled off and give its name.

*Structure*

*Name* .....

(ii) Give the formula of the product left in the distillation flask and give a use for it.

*Formula* .....

*Use* .....

**(4)**
**(Total 15 marks)**

**Q2.** The equation and rate law for the reaction of substance P with substance Q are given below.

$2P + Q \rightarrow R + S$

$rate = k[P]^2[H^+]$

Under which one of the following conditions, all at the same temperature, would the rate of reaction be slowest?

| | [P] / mol $dm^{-3}$ | pH |
|---|---|---|
| **A** | 0.1 | 0 |
| **B** | 1 | 2 |
| **C** | 3 | 3 |
| **D** | 10 | 4 |

**(Total 1 mark)**

**Q3.** Rate = k $[A]^2$ [B]

Correct units for the rate constant in the rate equation above are

**A** mol $dm^{-3}$ $s^{-1}$

**B** $mol^{-1}$ $dm^{-3}$ $s^{-1}$

**C** $mol^2$ $dm^{-6}$ $s^{-1}$

**D** $mol^{-2}\ dm^{6}\ s^{-1}$

**(Total 1 mark)**

**Q4.** (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 | $0.45 \times 10^{-5}$ |
| 2 | 0.30 | 0.24 | $0.90 \times 10^{-5}$ |
| 3 | 0.60 | 0.48 | $7.20 \times 10^{-5}$ |

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

.....

.....

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

.....

.....

**(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 | $9.30 \times 10^{-5}$ |
| 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* .....

.....

.....

*Units of k* .....

.....

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

.....

.....

.....

**(4)**
**(Total 6 marks)**

**Q5.** Iodine and propanone react in acid solution according to the equation

$$I_2 + CH_3COCH_3 \rightarrow CH_3COCH_2I + HI$$

The rate equation for the reaction is found to be

rate = $k\,[CH_3COCH_3][H^+]$

(a) Deduce the order of reaction with respect to iodine and the overall order of reaction.

*Order with respect to iodine* .....

*Overall order* .....

**(2)**

(b) At the start of the experiment, the rate of reaction was found to be

$2.00 \times 10^{-5}$ mol $dm^{-3}$ $s^{-1}$ when the concentrations of the reactants were as shown below.

| Reactant | Concentration / mol $dm^{-3}$ |
|---|---|
| $CH_3COCH_3$ | 1.50 |
| $I_2$ | $2.00 \times 10^{-2}$ |
| $H^+$ | $3.00 \times 10^{-2}$ |

Use these data to calculate a value for the rate constant and deduce its units.

*Rate constant* .....

.....

.....

*Units* .....

**(3)**

(c) How can you tell that $H^+$ acts as a catalyst in this reaction?

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**(2)**

(d) Calculate the initial rate of reaction if the experiment were to be repeated at the same temperature and with the same concentrations of iodine and propanone as in part (b) but at a pH of 1.25

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**(3)**
**(Total 10 marks)**

**Q6.** (a) The initial rate of the reaction between compounds **A** and **B** was measured in a series of experiments at a fixed temperature. The following rate equation was deduced.

rate = $k$[**A**][**B**]$^2$

(i) Complete the table of data below for the reaction between **A** and **B**.

| Expt | Initial [**A**] /mol dm$^{-3}$ | Initial [**B**] /mol dm$^{-3}$ | Initial rate /mol dm$^{-3}$ s$^{-1}$ |
|---|---|---|---|
| 1 | $4.80 \times 10^{-2}$ | $6.60 \times 10^{-2}$ | $10.4 \times 10^{-3}$ |
| 2 | $4.80 \times 10^{-2}$ | $3.30 \times 10^{-2}$ | |
| 3 | | $13.2 \times 10^{-2}$ | $5.20 \times 10^{-3}$ |
| 4 | $1.60 \times 10^{-2}$ | | $10.4 \times 10^{-3}$ |

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

*Calculation* .....

.....

*Units* .....

**(6)**

(b) State how the value of the rate constant, $k$, would change, if at all, if the concentration of **A** were increased in a series of experiments.

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**(1)**

**(Total 7 marks)**

**Q7.** (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 | $4.8 \times 10^{-3}$ |
| 2 | 0.10 | 0.10 | |
| 3 | 0.40 | | $9.6 \times 10^{-3}$ |
| 4 | | 0.60 | $19.2 \times 10^{-3}$ |

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

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**(6)**

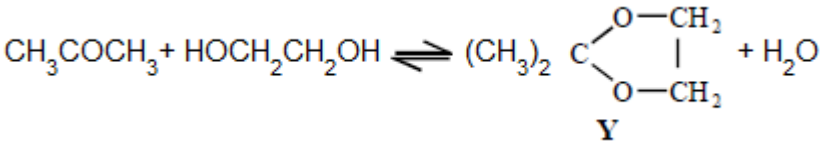
(b) What change in the reaction conditions would cause the value of the rate constant to change?

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**(1)**

**(Total 7 marks)**

**Q8.** This question is about the reaction between propanone and an excess of ethane-1,2-diol, the equation for which is given below.



**Y**

In a typical procedure, a mixture of 1.00 g of propanone, 5.00 g of ethane-1,2-diol and 0.100 g of benzenesulphonic acid, $C_6H_5SO_3H$, is heated under reflux in an inert solvent. Benzenesulphonic acid is a strong acid.

When the concentration of benzenesulphonic acid is doubled, the rate of the reaction doubles. It can be deduced that

**A** the reaction is first order overall.

**B** the reaction is third order overall.

**C** the reaction is acid-catalysed.

**D** units for the rate constant, $k$, are $mol^{-2}\ dm^6\ s^{-1}$.

**(Total 1 mark)**

**Q9.** (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 | $0.32 \times 10^{-3}$ |
| 2 | 0.36 | 0.15 | $2.88 \times 10^{-3}$ |
| 3 | 0.72 | 0.30 | $11.52 \times 10^{-3}$ |

(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 | $5.0 \times 10^{-2}$ | $2.0 \times 10^{-2}$ | $6.5 \times 10^{-4}$ |
| 5 | $6.5 \times 10^{-2}$ | $3.4 \times 10^{-2}$ | To be calculated |

The rate equation for this reaction is

$rate = k[NO]^2[O_2]$

(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.

.....

.....

.....

**(4)**
**(Total 6 marks)**

**Q10.** 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**?

.....

.....

**(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.

.....

**(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* .....

.....

**(4)**
**(Total 6 marks)**