AQA A-LEVEL CHEMISTRY · 7405

CHEMISTRY



EXAM QUESTION PACK

3.1.9 RATE EQUATIONS

SCORE	PERCENTAGE
/412	%

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Q1.

Iodide ions are oxidised to iodine by hydrogen peroxide in acidic conditions.

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

The rate equation for this reaction can be written as

$$rate = k\,[H_2O_2]^a\,[I^-]^b\,[H^+]^c$$

In an experiment to determine the order with respect to $H^+(aq)$, a reaction mixture is made containing $H^+(aq)$ with a concentration of 0.500 mol dm^{-3}

A large excess of both H_2O_2 and I^- is used in this reaction mixture so that the rate equation can be simplified to

$$rate = k_1\,[H^+]^c$$

(a) Explain why the use of a large excess of H_2O_2 and I^- means that the rate of reaction at a fixed temperature depends only on the concentration of $H^+(aq)$.

_____

_____

_____

_____

_____

(2)

(b) Samples of the reaction mixture are removed at timed intervals and titrated with alkali to determine the concentration of $H^+(aq)$.

State and explain what must be done to each sample before it is titrated with alkali.

_____

_____

_____

_____

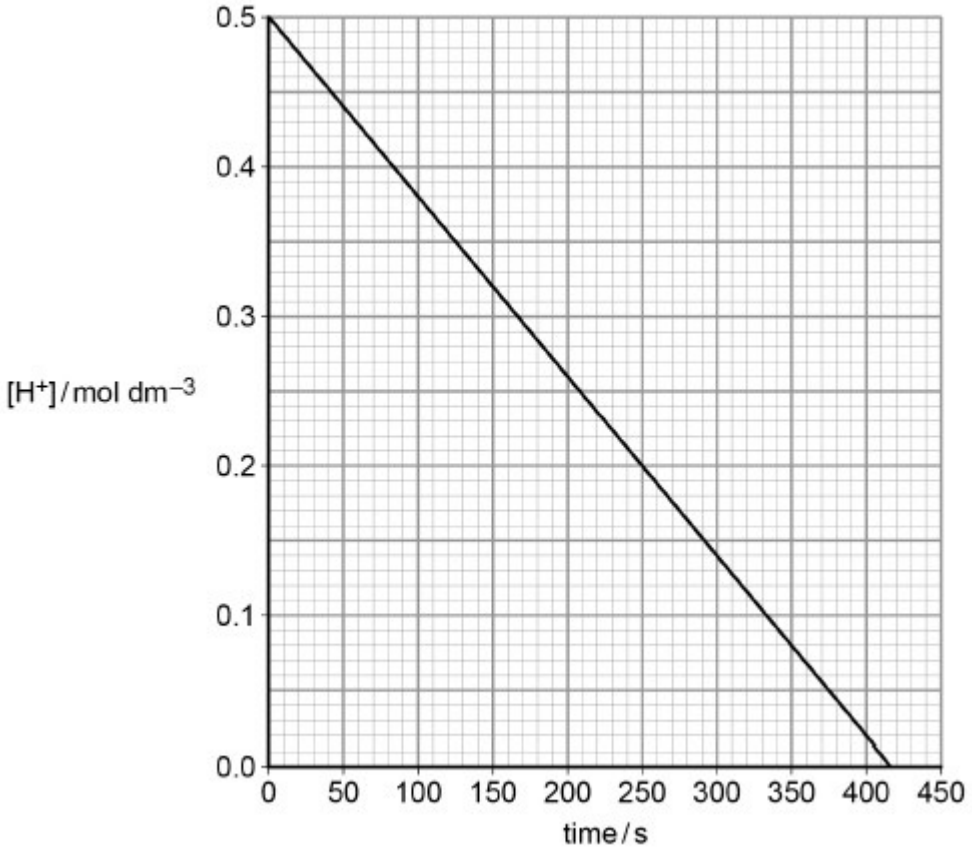
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(2)

(c) A graph of the results is shown in **Figure 1**.

Figure 1

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Explain how the graph shows that the order with respect to $H^+(aq)$ is zero.

_____

_____

_____

_____

(2)

(d) Use the graph in **Figure 1** to calculate the value of k_1

Give the units of k_1

k_1 _____

Units _____

(3)

(e) A second reaction mixture is made at the same temperature. The initial concentrations of $H^+(aq)$ and $I^-(aq)$ in this mixture are both 0.500 mol dm^{-3}

There is a large excess of H_2O_2

In this reaction mixture, the rate depends only on the concentration of $I^-(aq)$.

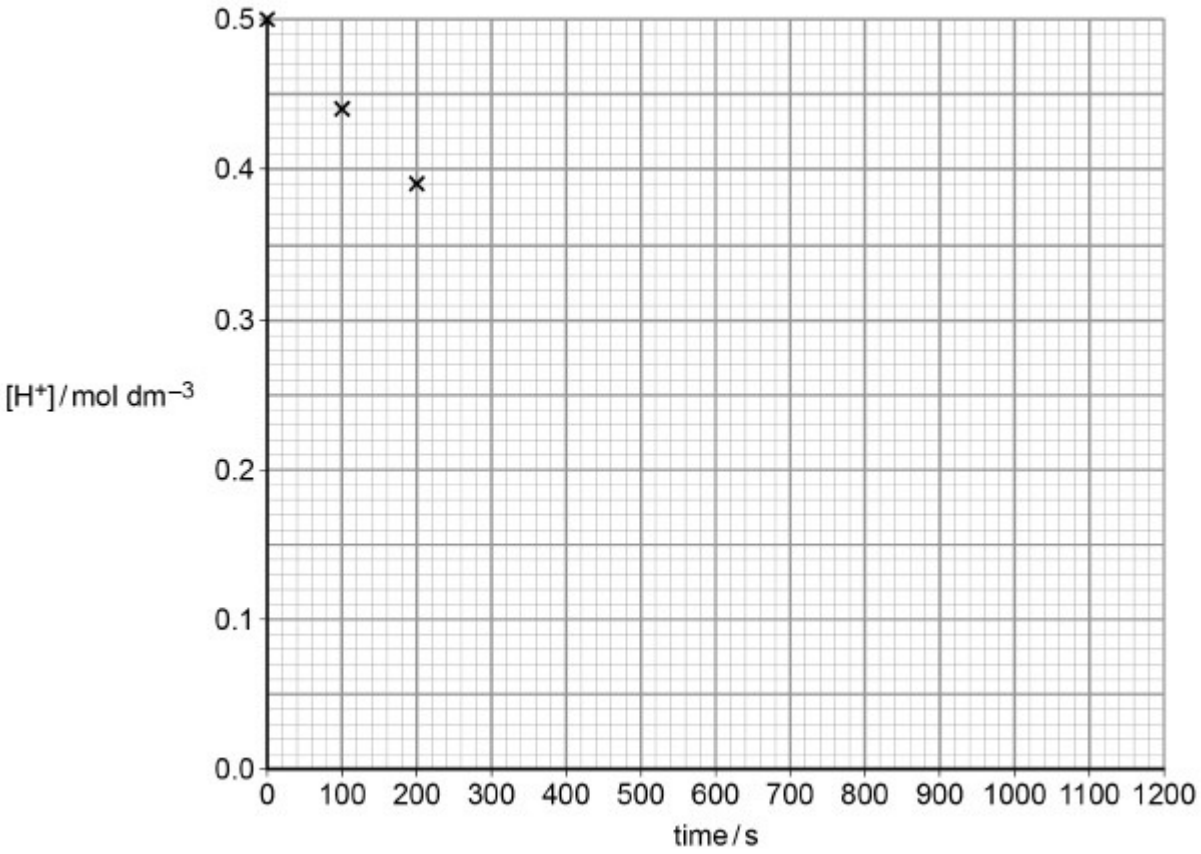
The results are shown in the table.

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Time / s	0	100	200	400	600	800	1000	1200
[H^+] / mol dm^{-3}	0.50	0.44	0.39	0.31	0.24	0.19	0.15	0.12

Plot these results on the grid in **Figure 2**. The first three points have been plotted.

Figure 2



(1)

(f) Draw a line of best fit on the grid in **Figure 2**.

(1)

(g) Calculate the rate of reaction when [H^+] = 0.35 mol dm^{-3}

Show your working using a suitable construction on the graph in **Figure 2**.

Rate _____ mol dm^{-3} s^{-1}

(2)

(h) A general equation for a reaction is shown.

A(aq) + **B**(aq) + **C**(aq) → **D**(aq) + **E**(aq)

In aqueous solution, **A**, **B**, **C** and **D** are all colourless but **E** is dark blue.

A reagent (**X**) is available that reacts rapidly with **E**. This means that, if a small

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amount of **X** is included in the initial reaction mixture, it will react with any **E** produced until all of the **X** has been used up.

Explain, giving brief experimental details, how you could use a series of experiments to determine the order of this reaction with respect to **A**. In each experiment you should obtain a measure of the initial rate of reaction.

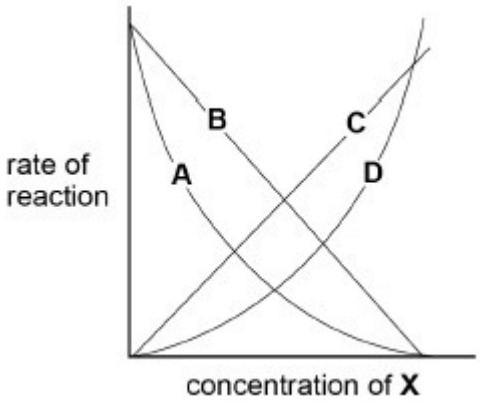
(6)

(Total 19 marks)

Q2.

A series of experiments was carried out to find the order of reaction with respect to reactant **X**. In these experiments, only the concentration of **X** was changed.

Which graph would show that the reaction is second-order with respect to **X**?



A ⬭

B ⬭

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C ⬭

D ⬭

(Total 1 mark)

Q3.

The mechanisms of some reactions can be deduced from kinetic data.

The table shows the results of five experiments involving the reaction between benzaldehyde (C_6H_5CHO) and potassium hydroxide solution.

Experiment	Initial rate /mol dm^{-3} s^{-1}	$[C_6H_5CHO]$ /mol dm^{-3}	$[C_6H_5CHO]^2$	[KOH] /mol dm^{-3}
1	0	0		0
2	2.90×10^{-6}	0.25		0.25
3	1.16×10^{-5}	0.50		0.25
4	2.61×10^{-5}	0.75		0.25
5	4.65×10^{-5}	1.00		0.25

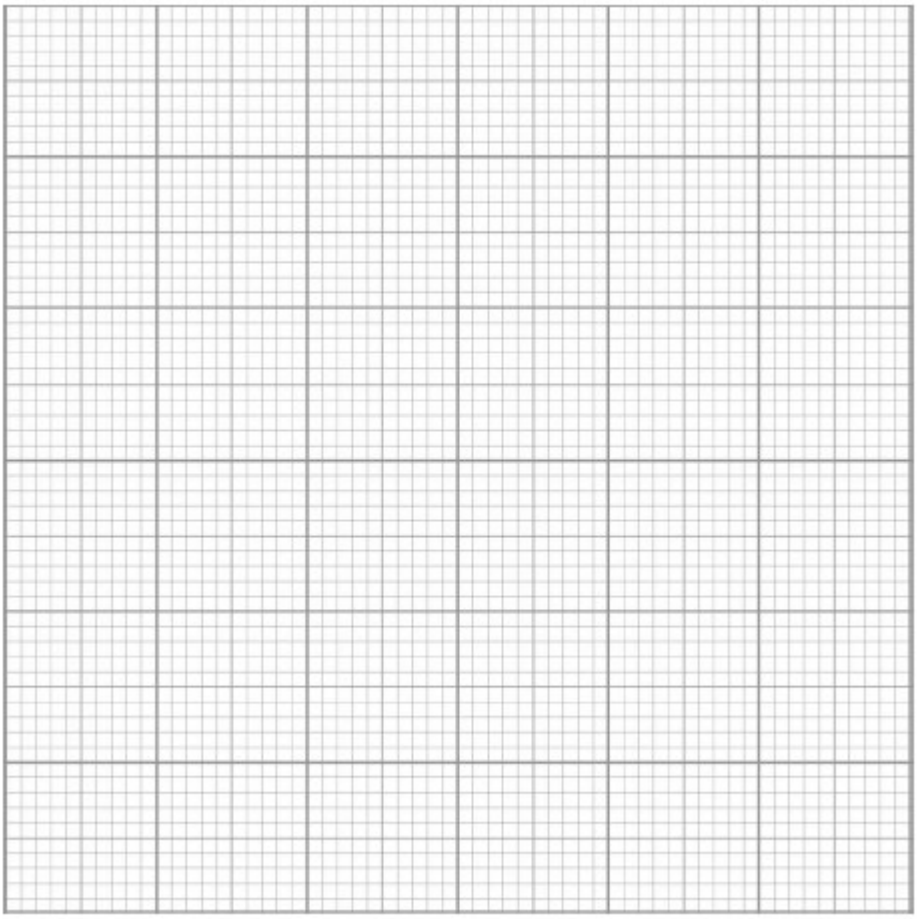
(a) Calculate the value of $[C_6H_5CHO]^2$ for each experiment.

Write your answers in the table above.

Plot a graph of initial rate against $[C_6H_5CHO]^2$ on the grid below.

Label the axes with units.

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(4)

(b) Deduce the order of the reaction with respect to C_6H_5CHO from the graph.

Explain your answer.

Order _____

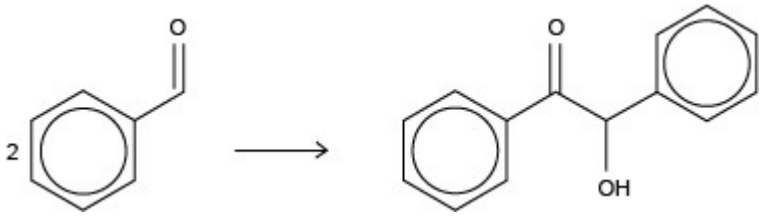
Explanation _____

_____

_____

(2)

(c) Benzaldehyde (C_6H_5CHO) undergoes a dimerisation reaction as shown:



The rate equation for the reaction is

$$rate = k\,[C_6H_5CHO]\,[CN^-]$$

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Suggest the role of the cyanide ion in the reaction.

Explain your answer.

Role of CN^- _____

Explanation _____

_____

_____

(2)

(Total 8 marks)

Q4.

Iodine reacts slowly with propanone in the presence of an acid catalyst according to the equation

$$CH_3COCH_3 + I_2 \longrightarrow CH_3COCH_2I + HI$$

The rate of this reaction can be followed by preparing mixtures in which only the initial concentration of propanone is varied. At suitable time intervals, a small sample of the mixture is removed and titrated with sodium thiosulfate solution. This allows determination of the concentration of iodine remaining at that time.The rate of this reaction can be followed by preparing mixtures in which only the initial concentration of propanone is varied. At suitable time intervals, a small sample of the mixture is removed and titrated with sodium thiosulfate solution. This allows determination of the concentration of iodine remaining at that time.

Five mixtures, **A**, **B**, **C**, **D** and **E**, are prepared as shown in **Table 1**.

Table 1

Mixture	**A**	**B**	**C**	**D**	**E**
Volume of 0.0200 mol dm^{-3} $I_2(aq)/cm^3$	40.0	40.0	40.0	40.0	40.0
Volume of 0.100 mol dm^{-3} $H_2SO_4(aq)/cm^3$	25.0	25.0	25.0	25.0	25.0
Volume of 1.00 mol dm^{-3} $CH_3COCH_3(aq)/cm^3$	25.0	20.0	15.0	10.0	6.5
Volume of distilled water/cm^3	0.0	5.0	10.0	15.0	18.5

(a) Calculate the initial concentration, in mol dm^{-3}, of the propanone in mixture **A**.

Concentration = _____ mol dm^{-3}

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(2)

(b) State and explain why different volumes of water are added to mixtures **B**, **C**, **D** and **E**.

_____

_____

_____

_____

(2)

(c) Calculate the volume of 0.0100 mol dm^{-3} sodium thiosulfate solution required to react with all of the iodine in a 10.0 cm^3 sample of mixture **E**, before the iodine reacts with propanone.

The equation for the reaction in the titration is

$$2Na_2S_2O_3(aq) + I_2(aq) \longrightarrow Na_2S_4O_6(aq) + 2NaI(aq)$$

Volume = _____ cm^3

(4)

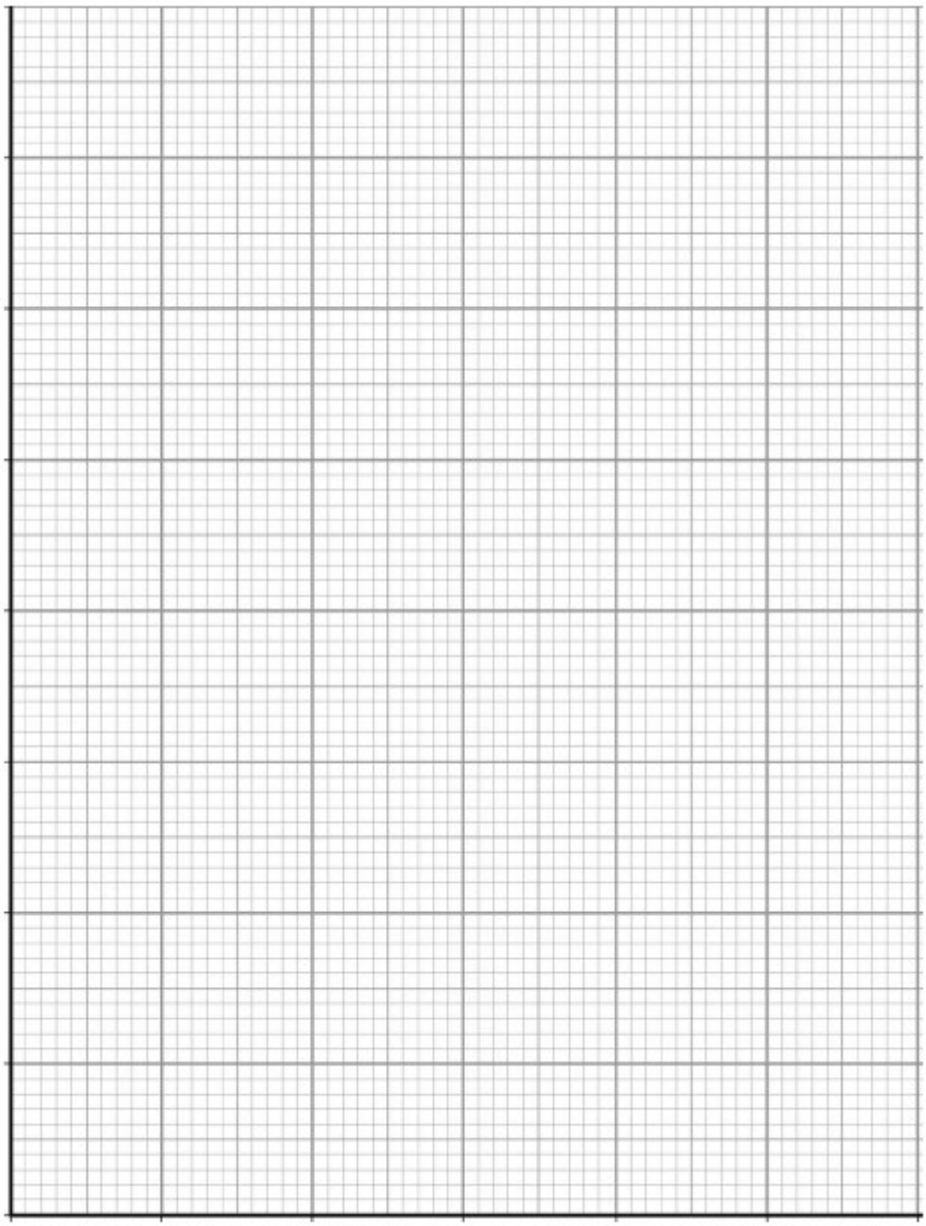
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(d) The results for mixture **E** are shown in **Table 2**.
V is the volume of 0.0100 mol dm^{-3} sodium thiosulfate solution needed, at different times, **t**, to react with the iodine in a 10.0 cm^3 sample of **E**.

Table 2

t/min	5	10	20	30
V/cm^3	17.5	17.2	16.6	16.0

Use these data and your answer to part **(c)** to plot a graph of **V** (y-axis) against **t** (x-axis) for mixture **E**.
Draw a best-fit straight line through your points and calculate the gradient of this line.



gradient = _____ $cm^3\ min^{-1}$

(5)

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(e) The gradients for similar graphs produced by mixtures **A**, **B**, **C** and **D** are shown in **Table 3**.
Each gradient is a measure of the rate of the reaction between iodine and propanone.

Table 3

Mixture	**A**	**B**	**C**	**D**
Gradient / cm^3 min^{-1}	−0.24	−0.20	−0.15	−0.10

Use information from **Table 1** and **Table 3** to deduce the order with respect to propanone. Explain your answer.

_____

_____

_____

_____

_____

_____

(2)

(f) Each sample taken from the reaction mixtures is immediately added to an excess of sodium hydrogencarbonate solution before being titrated with sodium thiosulfate solution.

Suggest the purpose of this addition.
Explain your answer.

_____

_____

_____

_____

_____

_____

(2)

(Total 17 marks)

Q5.

The rate equation for the acid-catalysed reaction between iodine and propanone is:

$$\text{rate} = k\,[H^+]\,[C_3H_6O]$$

The rate of reaction was measured for a mixture of iodine, propanone and sulfuric acid at pH = 0.70

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In a second mixture the concentration of the sulfuric acid was different but the concentrations of iodine and propanone were unchanged. The new rate of reaction was a quarter of the original rate.

What was the pH of the second mixture?

A 1.00 ○

B 1.30 ○

C 1.40 ○

D 2.80 ○

(Total 1 mark)

Q6.

A rate investigation was carried out on a reaction involving three reactants, **X**, **Y** and **Z**. The concentrations of the reactants were varied and the relative rate for each mixture determined.

Experiment	[**X**]/mol dm^{-3}	[**Y**]/mol dm^{-3}	[**Z**]/mol dm^{-3}	Relative rate
1	1×10^{-3}	1×10^{-3}	2×10^{-3}	1
2	2×10^{-3}	2×10^{-3}	2×10^{-3}	4
3	5×10^{-4}	2×10^{-3}	4×10^{-3}	0.5

The reaction is zero order with respect to **Y**.

What is the overall order of reaction?

A 0 ○

B 1 ○

C 2 ○

D 3 ○

(Total 1 mark)

Q7.

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**	**Initial**	**Initial rate**

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	concentration of A / mol dm^{-3}	concentration of B / mol dm^{-3}	/ mol dm^{-3} s^{-1}
1	0.12	0.26	2.10×10^{-4}
2	0.36	0.26	1.89×10^{-3}
3	0.72	0.13	3.78×10^{-3}

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

(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	1.9×10^{-2}	3.5×10^{-2}	7.2×10^{-4}
5	3.6×10^{-2}	5.4×10^{-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.

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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**.

_____

_____

_____

_____

_____

_____

(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}.

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Activation energy = _____

(2)

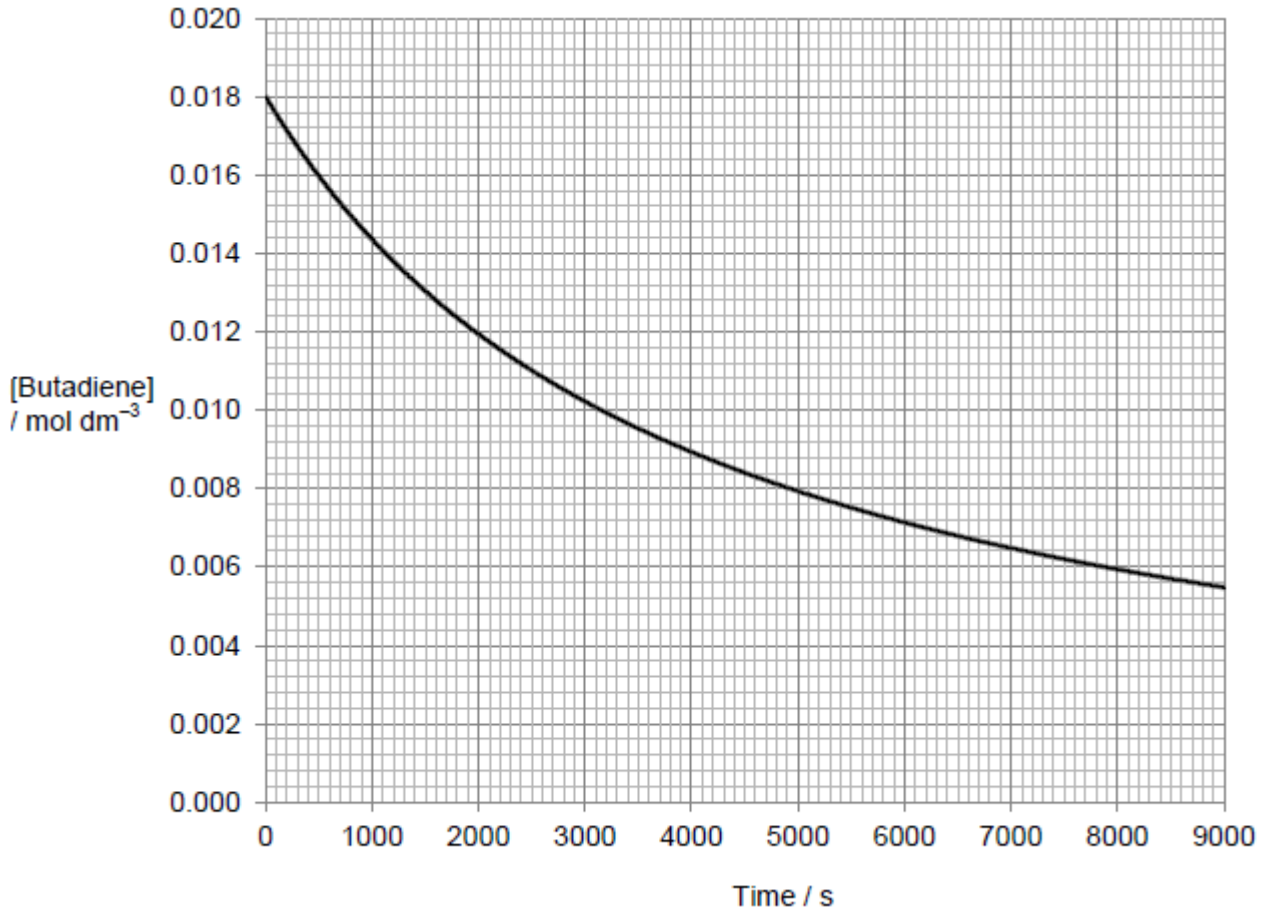
(Total 12 marks)

Q8.

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 4.57×10^{-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)

Q9.

The initial rate of the reaction between gases **D** and **E** was measured in a series of experiments at a constant temperature. The results are shown in the table.

Expt	Initial [D] / mol dm^{-3}	Initial [E] / mol dm^{-3}	Initial rate / mol dm^{-3} s^{-1}
1	1.25×10^{-2}	5.81×10^{-1}	1.16×10^{-2}
2	1.88×10^{-2}	8.73×10^{-1}	3.92×10^{-2}
3	1.88×10^{-2}	1.75	1.57×10^{-1}

(a) Deduce the order of reaction with respect to **D** and the order with respect to **E**.

Order with respect to **D** _____

Order with respect to **E** _____

Space for working _____

(2)

(b) Suggest why initial rates of reaction are used to determine these orders rather than rates of reaction at other times during the experiments.

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_____

(1)

(c) State how the initial rate is obtained from a graph of the concentration of the product against time.

_____

_____

_____

_____

(2)

(Total 5 marks)

Q10.

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	0.42×10^{-3}
2	0.45	0.24	3.78×10^{-3}
3	0.90	0.12	7.56×10^{-3}

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

_____

(1)

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

_____

(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	3.8×10^{-2}	2.6×10^{-2}	8.6×10^{-4}
5	6.3×10^{-2}	7.5×10^{-2}	To be calculated

The rate equation for this reaction is

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

(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 _____

_____

_____

_____

Units _____

_____

(3)

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

_____

_____

(1)

(Total 6 marks)

Q11.

The reaction between propanone and iodine in the presence of hydrochloric acid was studied at a constant temperature.

$$CH_3COCH_3 + I_2 \longrightarrow CH_3COCH_2I + HI$$

The following rate equation was deduced.

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

(a) Suggest why the order with respect to iodine is zero.

_____

_____

_____

(1)

(b) In an experiment the initial concentrations of propanone, iodine and hydrochloric acid were as shown in the table. The initial rate of reaction in this experiment was 8.64×10^{-7} mol dm^{-3} s^{-1}.

	Initial concentration / mol dm^{-3}
CH_3COCH_3	5.82×10^{-2}
I_2	1.78×10^{-3}
H^+	4.76×10^{-1}

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Use the data in the table and the rate equation to calculate a value for the rate constant at this temperature.
Give units with your answer.

(2)

(c) A series of experiments was carried out using concentrations of propanone approximately 100 times the concentrations of iodine and hydrochloric acid.

Suggest the rate equation under these conditions.
Explain your answer.

(2)

(Total 5 marks)

Q12.

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		$5.4 = 10^{-3}$

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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 _____

_____

_____

_____

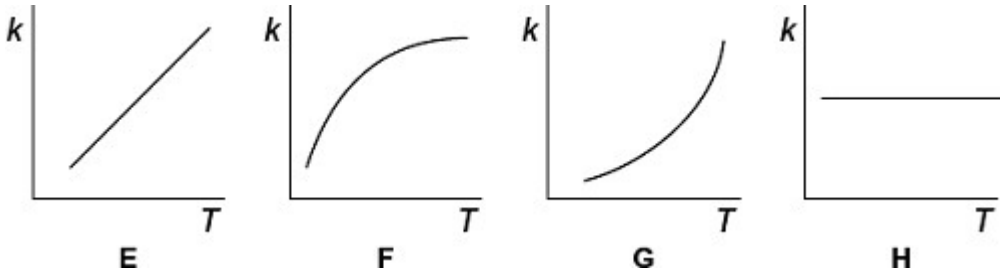
_____

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)

Q13.

(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.

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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**.

_____

_____

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_____

(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)

Q14.

(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

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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)

Q15.

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)

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(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

$$A + 3B \rightarrow AB_3$$

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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)

Q16.

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

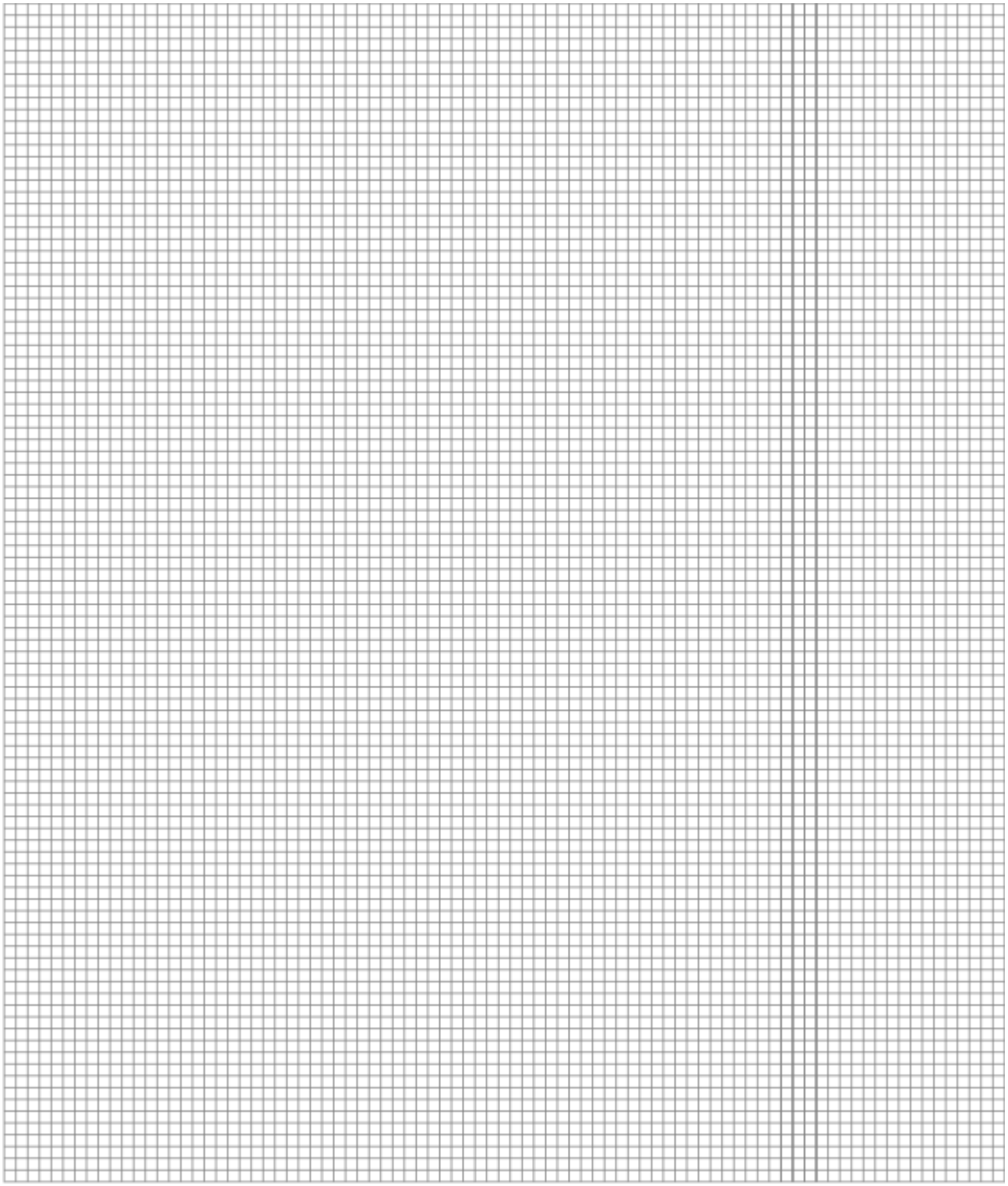
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3	10	1.00	37	−1.57
4	15	1.18	25	−1.40
5	20	1.30	19	−1.28
6	25	1.40	14	−1.15

(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.

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(5)

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

_____

_____

_____

(3)

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(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)

Q17.

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 _____

_____

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_____

_____

Units _____

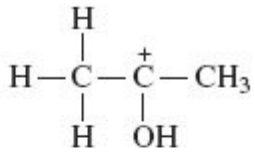
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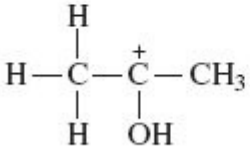
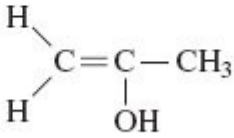
(c) Deduce how the initial rate of reaction changes when the concentration of iodine is doubled but the concentrations of propanone and of hydrochloric acid are unchanged.

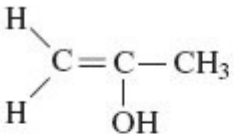
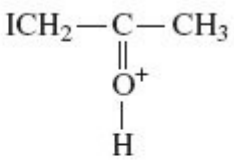
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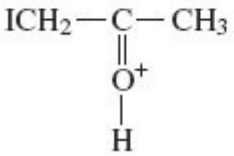
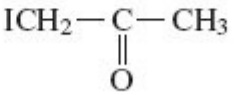
(1)

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

Step 1 $CH_3COCH_3 + H^+ \longrightarrow$ 

Step 2  $\longrightarrow$  $+ H^+$

Step 3  $+ I_2 \longrightarrow$  $+ I^-$

Step 4  $\longrightarrow$  $+ H^+$

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

Rate-determining step _____

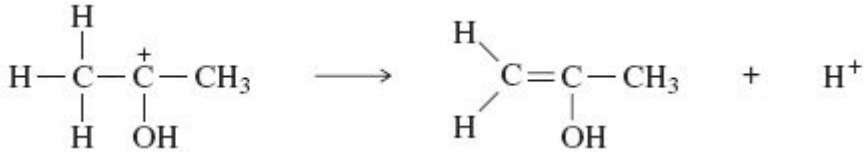
Explanation _____

_____

(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.

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Step 2 

(1)

(Total 8 marks)

Q18.

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.

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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)

Q19.

(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}.

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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)

Q20.

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 _____

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_____

(4)

(Total 6 marks)

Q21.

(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.

_____

_____

_____

_____

_____

(6)

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

_____

(1)

(Total 7 marks)

Q22.

(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}

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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**.

_____

_____

(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	9.30×10^{-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)

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Q1.

Bromate(V) ions and bromide ions react in acid conditions according to the equation

$$BrO_3^-(aq) + 5Br^-(aq) + 6H^+(aq) \rightarrow 3Br_2(aq) + 3H_2O(l)$$

(a) A series of experiments was carried out at a given temperature. The results were used to deduce the rate equation for the reaction.

$$rate = k\,[BrO_3^-][Br^-][H^+]^2$$

Table 1 shows an incomplete set of results.

Table 1

Experiment	Initial $[BrO_3^-]$ / mol dm^{-3}	Initial $[Br^-]$ / mol dm^{-3}	Initial $[H^+]$ / mol dm^{-3}	Initial rate of reaction / mol dm^{-3} s^{-1}
1	0.10	0.20	0.30	2.4×10^{-2}
2		0.20	0.30	3.6×10^{-2}
3	0.20	0.40	0.50	
4	0.10	0.10		2.7×10^{-2}

Use the data from Experiment 1 to calculate a value for the rate constant, k, at this temperature and give its units.

Give your answer to an appropriate number of significant figures.

k _____ Units _____

(3)

(b) Complete **Table 1**.

Space for working

(3)

(c) A second series of experiments was carried out to investigate how the rate of the reaction varies with temperature.
The results were used to obtain a value for the activation energy of the reaction, E_a

Identical amounts of reagents were mixed at different temperatures.

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The time taken, *t*, for a fixed amount of bromine to be formed was measured at different temperatures.

The results are shown in **Table 2**.

Table 2

Temperature, *T* / K	$\frac{1}{T}$ / K^{-1}	**Time, t / s**	$\frac{1}{t}$ / s^{-1}	$\ln \frac{1}{t}$
286	3.50×10^{-3}	54	1.85×10^{-2}	−3.99
295	3.39×10^{-3}	27	3.70×10^{-2}	
302		15	6.67×10^{-2}	−2.71
312	3.21×10^{-3}	8	1.25×10^{-1}	−2.08

Complete **Table 2**

(2)

(d) The Arrhenius equation can be written as

$$\ln k = -\frac{E_a}{R}\left(\frac{1}{T}\right) + C_1$$

In this experiment, the rate constant, *k*, is directly proportional to $\frac{1}{t}$

Therefore

$$\ln \frac{1}{t} = -\frac{E_a}{R}\left(\frac{1}{T}\right) + C_2$$

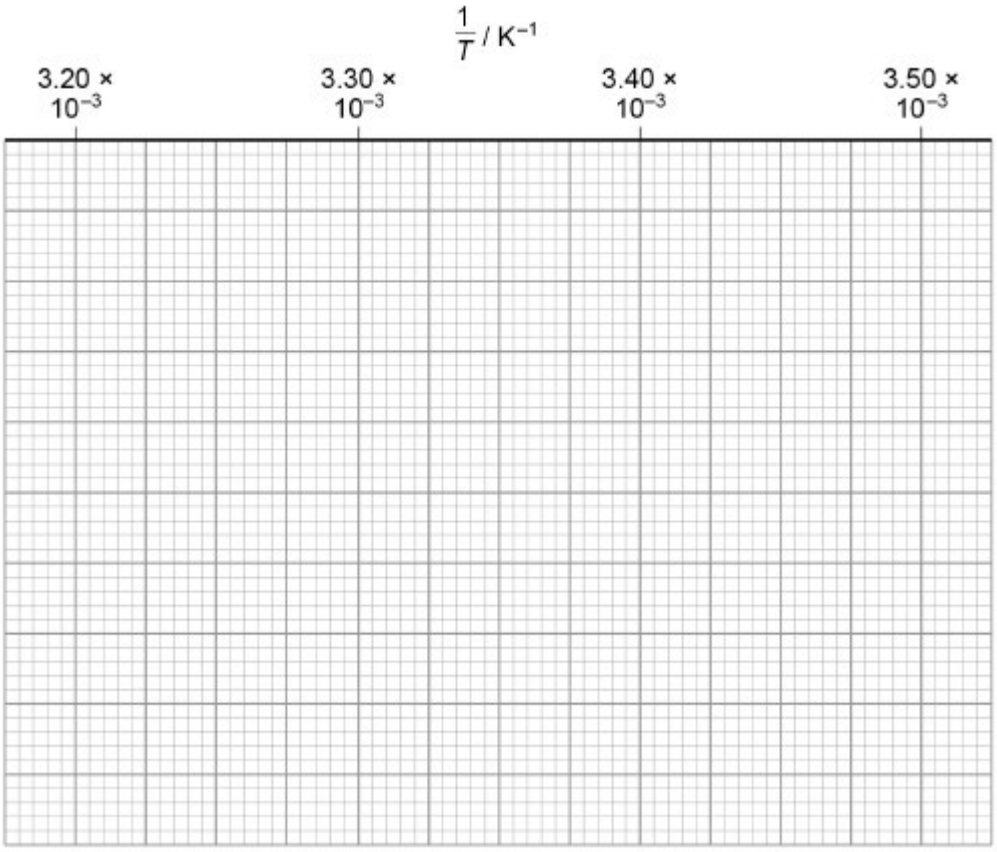
where C_1 and C_2 are constants.

Use values from **Table 2** to plot a graph of $\ln \frac{1}{t}$ (y axis) against $\frac{1}{T}$ on the grid.

Use your graph to calculate a value for the activation energy, in kJ mol^{-1}, for this reaction.

The value of the gas constant, $R = 8.31$ J K^{-1} mol^{-1}

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Activation energy _____ kJ mol^{-1}

(6)
(Total 14 marks)

Q2.

Iodide ions are oxidised to iodine by hydrogen peroxide in acidic conditions.

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

The rate equation for this reaction can be written as

$$rate = k\,[H_2O_2]^a\,[I^-]^b\,[H^+]^c$$

In an experiment to determine the order with respect to $H^+(aq)$, a reaction mixture is made containing $H^+(aq)$ with a concentration of 0.500 mol dm^{-3}

A large excess of both H_2O_2 and I^- is used in this reaction mixture so that the rate equation can be simplified to

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$$rate = k_1 [H^+]^c$$

(a) Explain why the use of a large excess of H_2O_2 and I^- means that the rate of reaction at a fixed temperature depends only on the concentration of $H^+(aq)$.

(2)

(b) Samples of the reaction mixture are removed at timed intervals and titrated with alkali to determine the concentration of $H^+(aq)$.

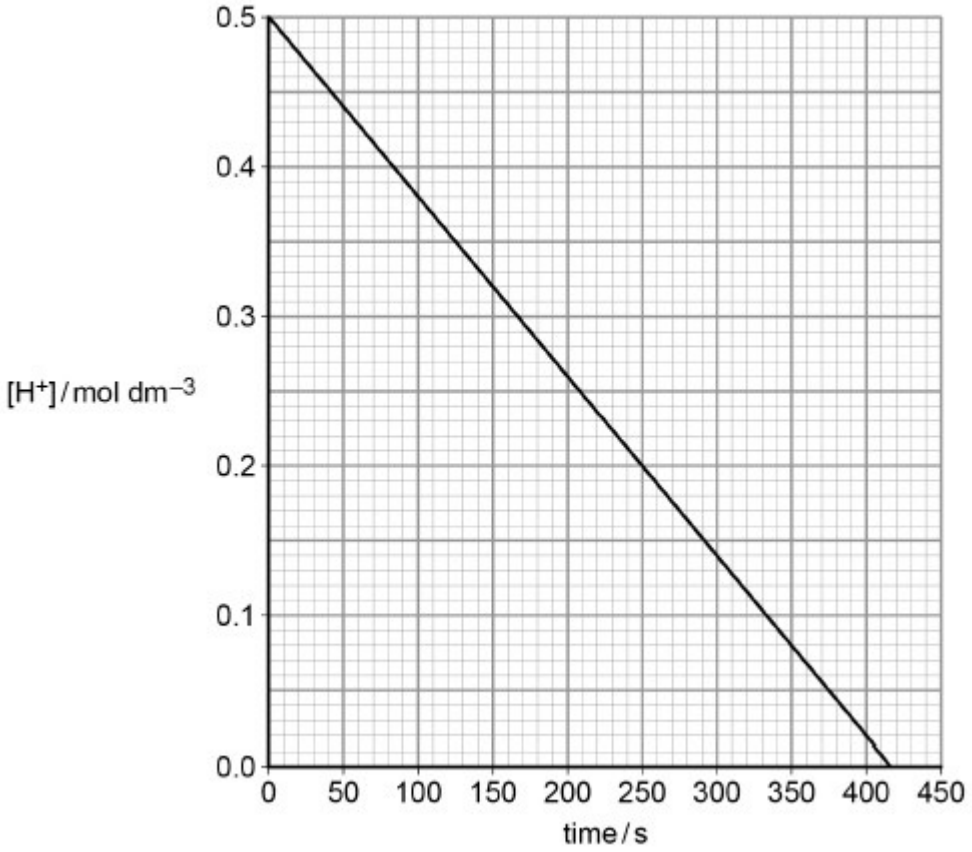
State and explain what must be done to each sample before it is titrated with alkali.

(2)

(c) A graph of the results is shown in **Figure 1**.

Figure 1

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Explain how the graph shows that the order with respect to $H^+(aq)$ is zero.

_____

_____

_____

_____

(2)

(d) Use the graph in **Figure 1** to calculate the value of k_1

Give the units of k_1

k_1 _____

Units _____

(3)

(e) A second reaction mixture is made at the same temperature. The initial concentrations of $H^+(aq)$ and $I^-(aq)$ in this mixture are both 0.500 mol dm^{-3}

There is a large excess of H_2O_2

In this reaction mixture, the rate depends only on the concentration of $I^-(aq)$.

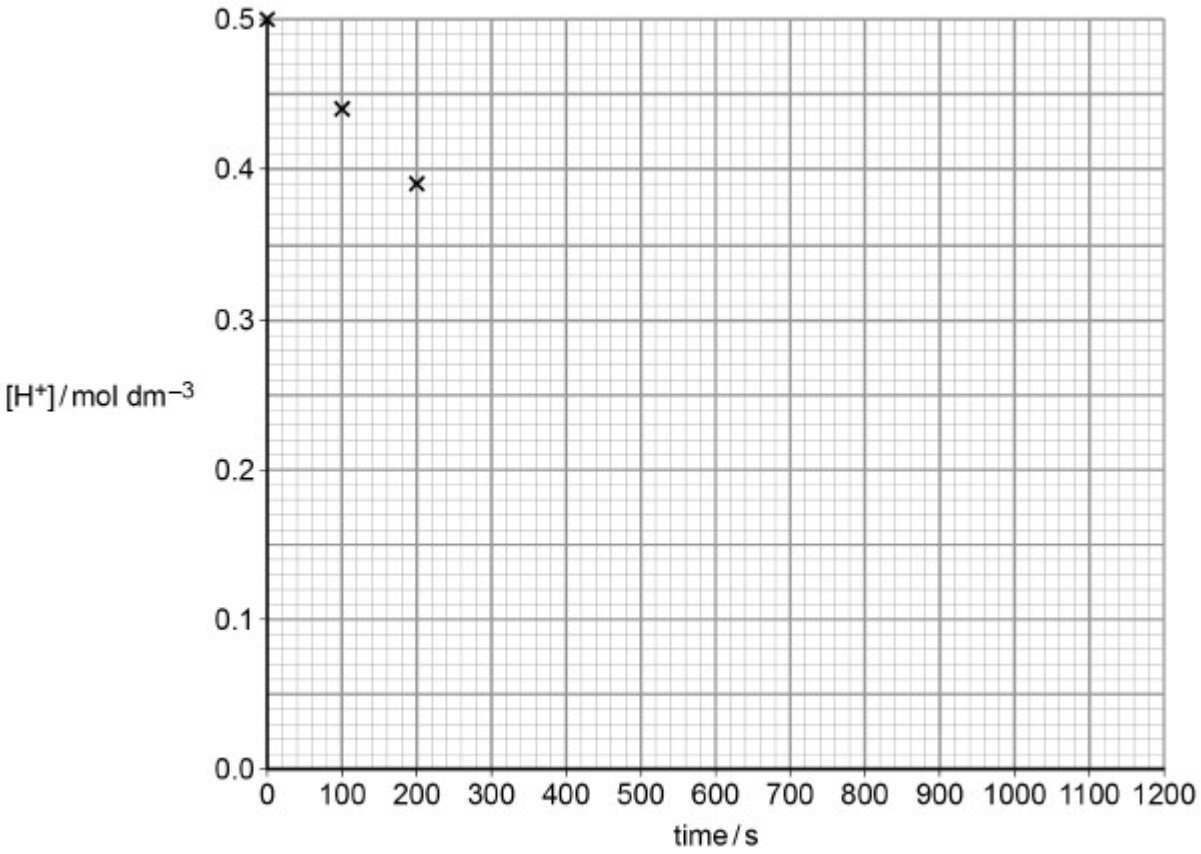
The results are shown in the table.

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Time / s	0	100	200	400	600	800	1000	1200
[H^+] / mol dm^{-3}	0.50	0.44	0.39	0.31	0.24	0.19	0.15	0.12

Plot these results on the grid in **Figure 2**. The first three points have been plotted.

Figure 2



(1)

(f) Draw a line of best fit on the grid in **Figure 2**.

(1)

(g) Calculate the rate of reaction when [H^+] = 0.35 mol dm^{-3}

Show your working using a suitable construction on the graph in **Figure 2**.

Rate _____ mol dm^{-3} s^{-1}

(2)

(h) A general equation for a reaction is shown.

A(aq) + **B**(aq) + **C**(aq) → **D**(aq) + **E**(aq)

In aqueous solution, **A**, **B**, **C** and **D** are all colourless but **E** is dark blue.

A reagent (**X**) is available that reacts rapidly with **E**. This means that, if a small

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amount of **X** is included in the initial reaction mixture, it will react with any **E** produced until all of the **X** has been used up.

Explain, giving brief experimental details, how you could use a series of experiments to determine the order of this reaction with respect to **A**. In each experiment you should obtain a measure of the initial rate of reaction.

_____

_____

_____

_____

_____

_____

_____

_____

_____

_____

_____

(6)

(Total 19 marks)

Q3.

Solutions of two compounds, **W** and **X**, react together in the presence of a soluble catalyst, **Y**, as shown in the equation

$$2\mathbf{W} + \mathbf{X} \rightarrow \mathbf{Z}$$

When the concentrations of **W**, **X** and **Y** are all doubled, the rate of reaction increases by a factor of four.

Which is a possible rate equation for this reaction?

A rate = k [**W**]^2 [**X**] ⬭

B rate = k [**W**]^2 [**Y**] ⬭

C rate = k [**X**] [**Y**] ⬭

D rate = k [**X**] [**Z**] ⬭

(Total 1 mark)

Q4.

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A series of experiments is carried out with compounds **C** and **D**. Using the data obtained, the rate equation for the reaction between the two compounds is deduced to be

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

In one experiment at 25 °C, the initial rate of reaction is 3.1×10^{-3} mol dm^{-3} s^{-1} when the initial concentration of C is 0.48 mol dm^{-3} and the initial concentration of D is 0.23 mol dm^{-3}

(a) Calculate a value for the rate constant at this temperature and give its units.

Rate constant _____ Units _____

(3)

(b) An equation that relates the rate constant, k, to the activation energy, E_a, and the temperature, T, is

$$\ln k = \frac{-E_a}{RT} + \ln A$$

Use this equation and your answer from part (a) to calculate a value, in kJ mol^{-1}, for the activation energy of this reaction at 25 °C.

For this reaction ln A = 16.9

The gas constant R = 8.31 J K^{-1} mol^{-1}

(If you were unable to complete part (a) you should use the value of 3.2×10^{-3} for the rate constant. This is not the correct value.)

Activation energy _____ kJ mol^{-1}

(4)

(Total 7 marks)

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Q5.

The mechanisms of some reactions can be deduced from kinetic data.

The table shows the results of five experiments involving the reaction between benzaldehyde (C_6H_5CHO) and potassium hydroxide solution.

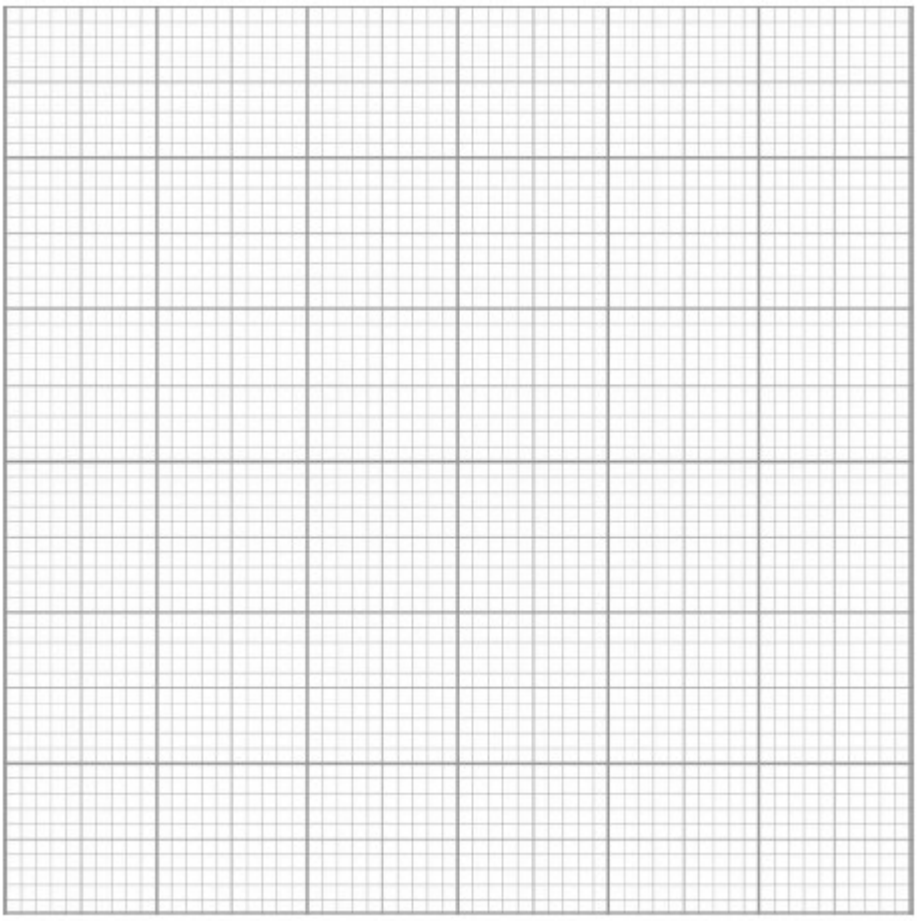
Experiment	Initial rate /mol dm^{-3} s^{-1}	[C_6H_5CHO] /mol dm^{-3}	$[C_6H_5CHO]^2$	[KOH] /mol dm^{-3}
1	0	0		0
2	2.90×10^{-6}	0.25		0.25
3	1.16×10^{-5}	0.50		0.25
4	2.61×10^{-5}	0.75		0.25
5	4.65×10^{-5}	1.00		0.25

(a) Calculate the value of $[C_6H_5CHO]^2$ for each experiment.

Write your answers in the table above.

Plot a graph of initial rate against $[C_6H_5CHO]^2$ on the grid below.

Label the axes with units.



.

(4)

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(b) Deduce the order of the reaction with respect to C_6H_5CHO from the graph.

Explain your answer.

Order _____

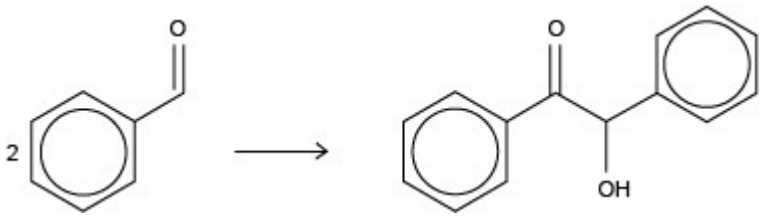
Explanation _____

_____

_____

(2)

(c) Benzaldehyde (C_6H_5CHO) undergoes a dimerisation reaction as shown:



The rate equation for the reaction is

$$rate = k\,[C_6H_5CHO]\,[CN^-]$$

Suggest the role of the cyanide ion in the reaction.

Explain your answer.

Role of CN^- _____

Explanation _____

_____

_____

(2)

(Total 8 marks)

Q6.

The Arrhenius equation can be written as:

$$\ln k = -\frac{E_a}{RT} + \ln A$$

The table below shows the value of the rate constant at different temperatures for a reaction.

Rate constant k/s^{-1}	**ln k**	**Temperature/K**	$\frac{1}{T}$
6.13×10^{-5}		700	
2.75×10^{-4}		727	

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8.25×10^{-4}		746	
8.23×10^{-3}		793	

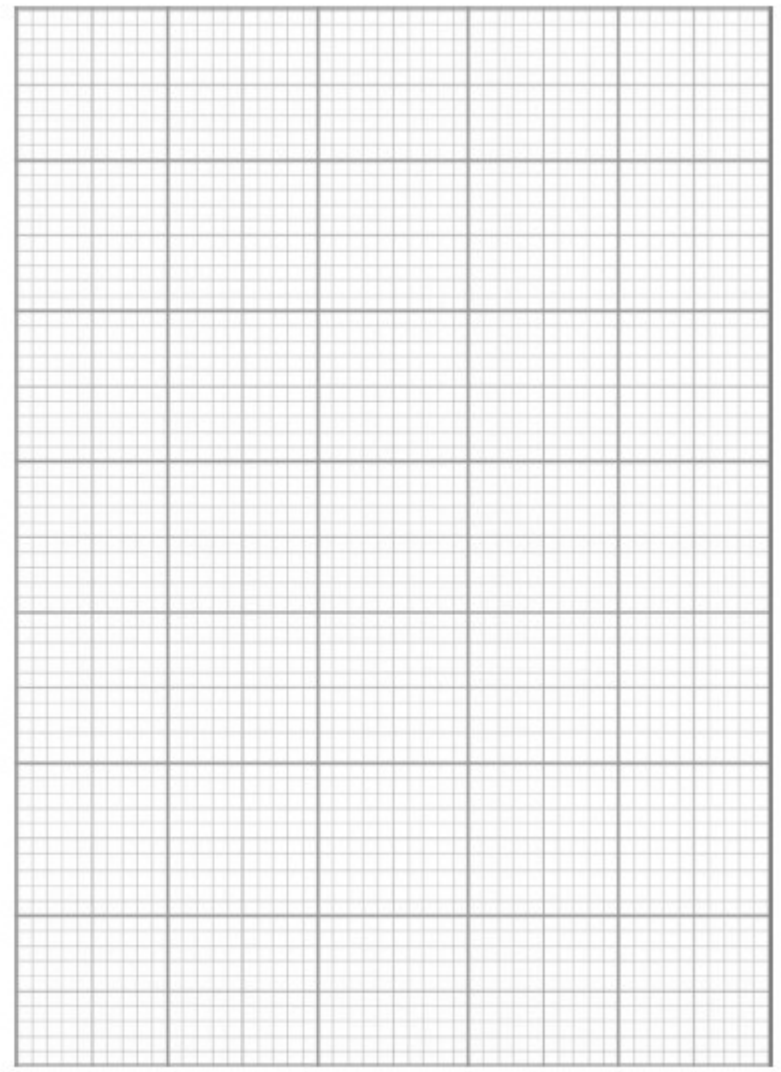
Complete the table by calculating the values of ln k and $\frac{1}{T}$ at each temperature.

Plot a graph of ln k against $\frac{1}{T}$ on the grid opposite.

Use your graph to calculate a value for the activation energy, in kJ mol^{-1}, for this reaction.

To gain full marks you must show all your working.

The gas constant R = 8.31 J K^{-1} mol^{-1}



E_a _____ kJ mol^{-1}

(Total 8 marks)

Q7.

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A rate investigation was carried out on a reaction involving three reactants, **X**, **Y** and **Z**. The concentrations of the reactants were varied and the relative rate for each mixture determined.

Experiment	[**X**]/mol dm^{-3}	[**Y**]/mol dm^{-3}	[**Z**]/mol dm^{-3}	Relative rate
1	1×10^{-3}	1×10^{-3}	2×10^{-3}	1
2	2×10^{-3}	2×10^{-3}	2×10^{-3}	4
3	5×10^{-4}	2×10^{-3}	4×10^{-3}	0.5

The reaction is zero order with respect to **Y**.

What is the overall order of reaction?

A 0 ⬭

B 1 ⬭

C 2 ⬭

D 3 ⬭

(Total 1 mark)

Q8.

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	2.10×10^{-4}
2	0.36	0.26	1.89×10^{-3}
3	0.72	0.13	3.78×10^{-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	1.9×10^{-2}	3.5×10^{-2}	7.2×10^{-4}
5	3.6×10^{-2}	5.4×10^{-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**.

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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**.

_____

_____

_____

_____

_____

_____

(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)

Q9.

The rate equation for the hydrogenation of ethene

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

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

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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)

Q10.

The initial rate of the reaction between gases **D** and **E** was measured in a series of experiments at a constant temperature. The results are shown in the table.

Expt	Initial [D] / mol dm^{-3}	Initial [E] / mol dm^{-3}	Initial rate / mol dm^{-3} s^{-1}
1	1.25×10^{-2}	5.81×10^{-1}	1.16×10^{-2}
2	1.88×10^{-2}	8.73×10^{-1}	3.92×10^{-2}
3	1.88×10^{-2}	1.75	1.57×10^{-1}

(a) Deduce the order of reaction with respect to **D** and the order with respect to **E**.

Order with respect to **D** _____

Order with respect to **E** _____

Space for working _____

_____

_____

_____

(2)

(b) Suggest why initial rates of reaction are used to determine these orders rather than rates of reaction at other times during the experiments.

_____

_____

_____

(1)

(c) State how the initial rate is obtained from a graph of the concentration of the product against time.

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(2)
(Total 5 marks)

Q11.

The reaction between propanone and iodine in the presence of hydrochloric acid was studied at a constant temperature.

$$CH_3COCH_3 + I_2 \longrightarrow CH_3COCH_2I + HI$$

The following rate equation was deduced.

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

(a) Suggest why the order with respect to iodine is zero.

(1)

(b) In an experiment the initial concentrations of propanone, iodine and hydrochloric acid were as shown in the table. The initial rate of reaction in this experiment was 8.64×10^{-7} mol dm^{-3} s^{-1}.

	Initial concentration / mol dm^{-3}
CH_3COCH_3	5.82×10^{-2}
I_2	1.78×10^{-3}
H^+	4.76×10^{-1}

Use the data in the table and the rate equation to calculate a value for the rate constant at this temperature.
Give units with your answer.

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(2)

(c) A series of experiments was carried out using concentrations of propanone approximately 100 times the concentrations of iodine and hydrochloric acid.

Suggest the rate equation under these conditions.
Explain your answer.

_____

_____

_____

_____

_____

_____

(2)
(Total 5 marks)

Q12.

(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}	1.60×10^{-2}	2.40×10^{-2}	3.60×10^{-2}
Initial concentration of B / mol dm^{-3}	4.20×10^{-2}	6.30×10^{-2}	6.30×10^{-2}
Initial rate /mol dm^{-3} s^{-1}	8.00×10^{-5}	1.80×10^{-4}	4.05×10^{-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.

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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 4.55×10^{-2} mol dm^{-3} and the initial concentration of **D** is 1.70×10^{-2} mol dm^{-3}, the initial rate of reaction is 6.64×10^{-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)

Q13.

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.

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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	4.2×10^{-3}	2.8×10^{-3}	3.3×10^{-5}
2	7.9×10^{-3}	2.8×10^{-3}	
3		5.6×10^{-3}	1.8×10^{-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 _____

_____

_____

_____

_____

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.

_____

_____

_____

(1)

(Total 6 marks)

Q14.

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.

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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	0.42×10^{-3}
2	0.45	0.24	3.78×10^{-3}
3	0.90	0.12	7.56×10^{-3}

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

_____

(1)

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

_____

(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	3.8×10^{-2}	2.6×10^{-2}	8.6×10^{-4}
5	6.3×10^{-2}	7.5×10^{-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 _____

_____

_____

_____

Units _____

_____

(3)

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

_____

_____

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(1)
(Total 6 marks)

Q15.

(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}

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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)

Q16.

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

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

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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	2.5×10^{-2}	1.8×10^{-2}	5.0×10^{-5}
2	7.5×10^{-2}	1.8×10^{-2}	
3	5.0×10^{-2}		5.0×10^{-5}
4		5.4×10^{-2}	4.5×10^{-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 _____

_____

_____

_____

Units _____

_____

(3)

(Total 6 marks)

Q17.

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}

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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 _____

_____

_____

_____

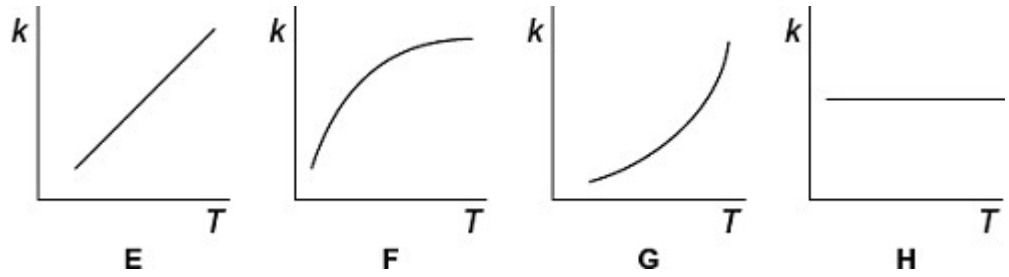
_____

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)

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Q18.

(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 _____

_____

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_____

(2)

(Total 6 marks)

Q19.

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.

_____

_____

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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)

Q20.

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**.

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_____

(1)

(b) The initial rate of the reaction at 298K was found to be 2.2×10^{-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.

_____

_____

_____

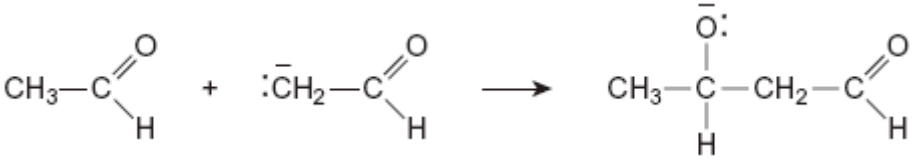
(2)

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

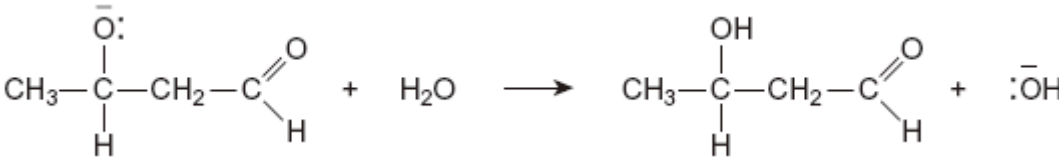
Step **1**



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)

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(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)

Q21.

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 _____

_____

_____

_____

Units _____

(3)

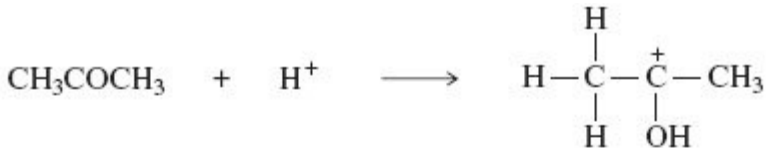
(c) Deduce how the initial rate of reaction changes when the concentration of iodine is doubled but the concentrations of propanone and of hydrochloric acid are unchanged.

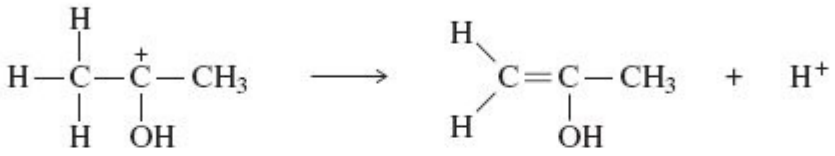
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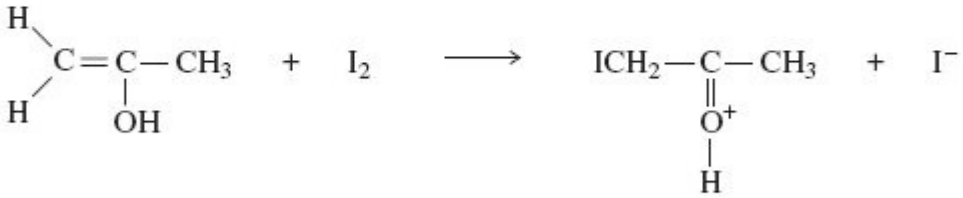
(1)

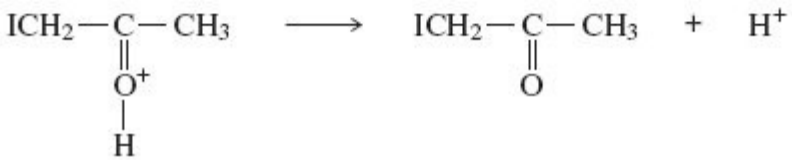
(d) The following mechanism for the overall reaction has been proposed.

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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.

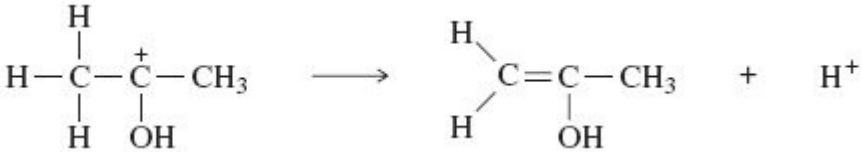
Rate-determining step _____

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)

Q22.

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	Initial concentration	Initial rate/

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	of **A**/mol dm^{-3}	of **B**/mol dm^{-3}	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.

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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)

Q23.

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}
1.15×10^{-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}
4.56×10^{-5}	8.94×10^{-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)

Q24.

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.

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

(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	3.0×10^{-3}	1.0×10^{-3}	1.8×10^{-5}
2	3.0×10^{-3}		7.2×10^{-5}
3	1.5×10^{-3}	1.0×10^{-3}	
4		0.50×10^{-3}	8.1×10^{-5}

(3)

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

_____

_____

_____

_____

_____

_____

(3)

(Total 6 marks)

Q25.

(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 _____

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(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	4.0×10^{-4}
2	0.040	0.045	6.0×10^{-4}
3	0.060	0.045	9.0×10^{-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 9.0×10^{-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.

_____

(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.

_____

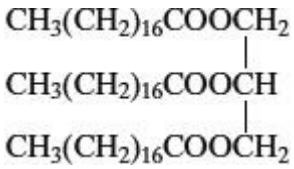
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_____

_____

(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)

Q26.

(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**.

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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)

Q27.

(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	4.80×10^{-2}	6.60×10^{-2}	10.4×10^{-3}
2	4.80×10^{-2}	3.30×10^{-2}	
3		13.2×10^{-2}	5.20×10^{-3}
4	1.60×10^{-2}		10.4×10^{-3}

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

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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.

_____

(1)

(Total 7 marks)

Q28.

(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

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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)

Q29.

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.

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Rate equation _____

Units for the rate constant _____

_____

(4)

(Total 6 marks)

Q30.

(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.

_____

_____

_____

_____

_____

(6)

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

_____

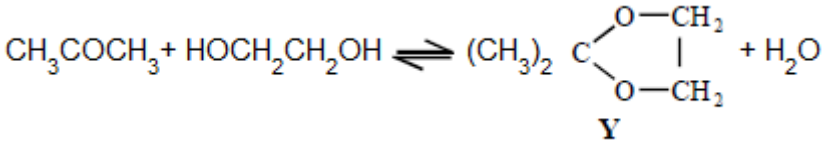
(1)

(Total 7 marks)

Q31.

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

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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)

Q32.

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

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

(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 2.00×10^{-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	2.00×10^{-2}
H^+	3.00×10^{-2}

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

Rate constant _____

_____

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_____

Units _____

(3)

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

_____

_____

_____

(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

_____

_____

_____

_____

_____

_____

(3)

(Total 10 marks)

Q33.

(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**.

_____

_____

(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[**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.

_____

_____

_____

(4)

(Total 6 marks)

Q34.

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)

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Q35.

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)

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