Questions

Q1.

The progress of the reaction between iodine and propanone with an acid catalyst can be followed in an experiment using a titrimetric method.

Procedure

Step **1** Mix 25 cm^3 of 1 mol dm^{-3} aqueous propanone with 25 cm^3 of 1 mol dm^{-3} sulfuric acid in a beaker. Both these reactants are in excess.

Step **2** Start the stop clock as 50 cm^3 of 0.02 mol dm^{-3} iodine solution is added to the beaker. Mix the reactants thoroughly.

Step **3** Withdraw a 10.0 cm^3 sample of the reaction mixture, using a pipette, and transfer it to a conical flask.

Step **4** Add a spatula measure of sodium hydrogencarbonate, noting the exact time.

Step **5** Titrate the iodine present in the 10.0 cm^3 sample with 0.01 mol dm^{-3} sodium thiosulfate solution, using starch indicator.

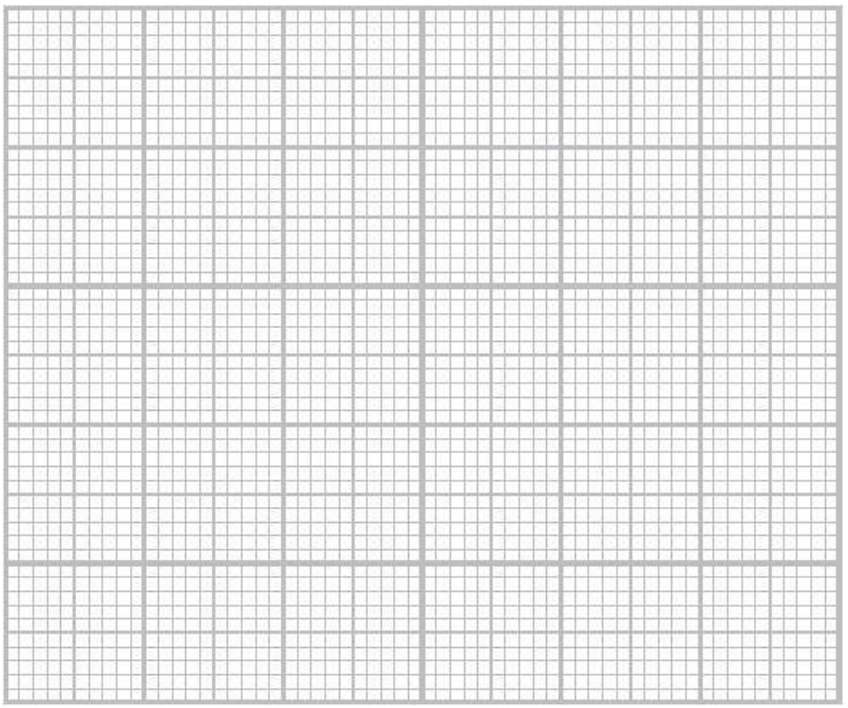
Step **6** Continue to withdraw 10.0 cm^3 samples about every two minutes, repeating Steps **4** and **5** with each sample.

Some data from the experiment are shown.

Time sodium hydrogencarbonate is added / min	2.0	5.0	6.5	8.0	10.5	12.0
Volume of sodium thiosulfate / cm^3	19.2	15.5	14.0	12.1	9.5	7.2

(i) Plot a graph of the volume of sodium thiosulfate against the time the sodium hydrogencarbonate is added.

(1)



(ii) Explain how the graph of volume of thiosulfate against time confirms the reaction is zero order with respect to iodine, I_2.

(3)

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(Total for question = 4 marks)

Q2.

The progress of the reaction between iodine and propanone with an acid catalyst can be followed in an experiment using a titrimetric method.

Procedure

Step **1** Mix 25 cm^3 of 1 mol dm^{-3} aqueous propanone with 25 cm^3 of 1 mol dm^{-3} sulfuric acid in a beaker. Both these reactants are in excess.

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Some data from the experiment are shown.

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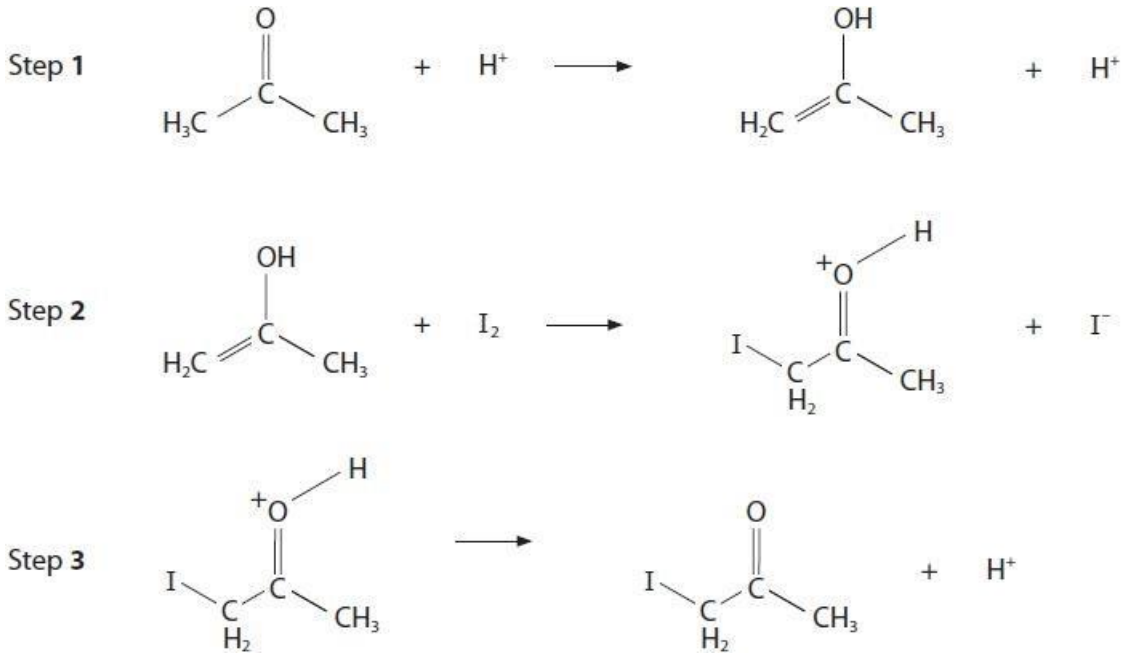
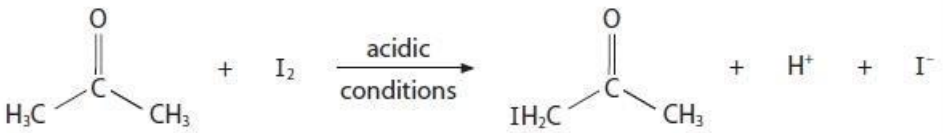
Volume of sodium thiosulfate / cm^3	19.2	15.5	14.0	12.1	9.5	7.2

The overall rate equation for the reaction is rate = $k[H^+(aq)][CH_3COCH_3(aq)]$.

A student researching the mechanism for the reaction found this example.



Overall reaction



(i) Predict which of the three steps is the rate-determining step. Justify your answer.

(2)

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(ii) The student stated that
'The hydrogen ions cannot be acting as a catalyst.
One hydrogen ion is a reactant in Step **1** but two hydrogen ions are formed as products in Steps **1** and **3**.'
Explain whether or not this statement is valid.

(2)

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(Total for question = 4 marks)

Q3.

At high temperatures, ethanal decomposes to form methane and carbon monoxide. The reaction is second order with respect to ethanal and second order overall.

$$CH_3CHO \rightarrow CH_4 + CO$$

(a) Write the rate equation for this reaction.

(1)

(b) Deduce the units of the rate constant given that the units of rate are mol dm^{-3} s–1.

(1)

(Total for question = 2 marks)

Q4.

At high temperatures, ethanal decomposes to form methane and carbon monoxide. The reaction is second order with respect to ethanal and second order overall.

$$CH_3CHO \rightarrow CH_4 + CO$$

(c) The table shows the concentration of ethanal in a sample at different times.

Time / s	Concentration of ethanal / mol dm^{-3}
0	0.72
420	0.36
1260	0.18

Calculate average values for the rate of reaction between 0 and 420 seconds and between 420 and 1260 seconds.
Give your answers to an appropriate number of significant figures.

(2)

0 s – 420 s .....

420 s – 1260 s .....

(d) Explain why the data given and your answers in (c) show that the reaction is **neither** zero order **nor** first order.

(2)

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(Total for question = 4 marks)

Q5.

At high temperatures, ethanal decomposes to form methane and carbon monoxide. The reaction is second order with respect to ethanal and second order overall.

$$CH_3CHO \rightarrow CH_4 + CO$$

The rate constant for the reaction was determined at five temperatures.
The results are given in the table.

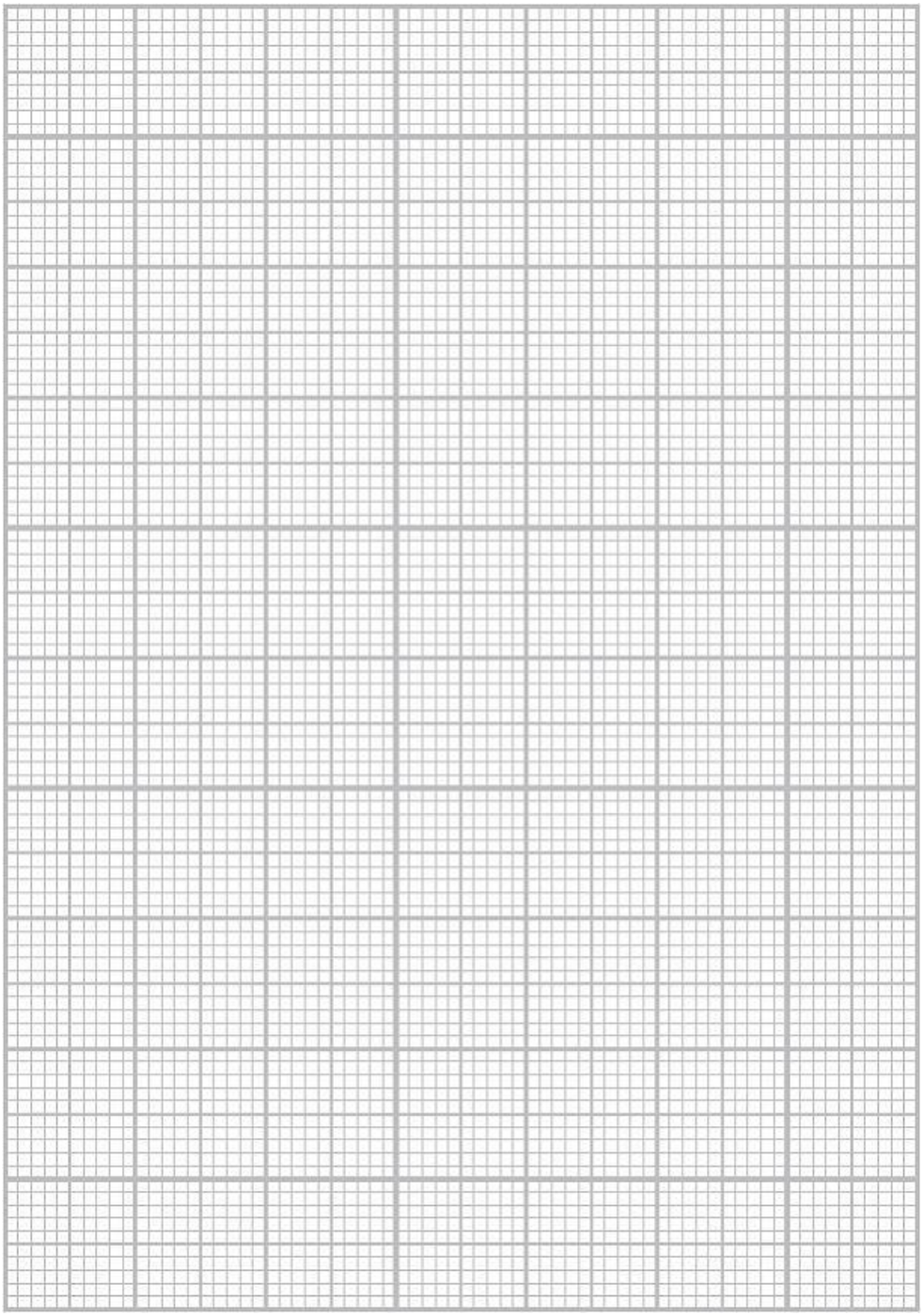
Temperature (T) / K	1/ Temperature ($1/T$) / K^{-1}	Rate constant (k) / units in (b)	ln k
700	1.43×10^{-3}	0.011	–4.51
730	1.37×10^{-3}	0.035	–3.35
760	1.32×10^{-3}	0.105	–2.25
790		0.343	
810	1.23×10^{-3}	0.787	–0.24

Determine the activation energy, E_a, in kJ mol^{-1}, by completing the data in the table and plotting a graph of ln k against $1/T$.

You should include the value of the gradient of the line and its units.

The Arrhenius equation can be expressed as $\ln k = -\frac{E_a}{R} \times \frac{1}{T} + \text{constant}$

(7)



(Total for question = 7 marks)

Q6.

This question is about reaction kinetics.

Compound **A** decomposes in a first order reaction.

Calculate the time it takes for the mass of **A** to decrease from 600 g to 37.5 g if the decomposition has a constant half-life of 14 minutes.

(1)

(Total for question = 1 mark)

Q7.

This question is about transition metals and their ions.

Tungsten wire catalyses the decomposition of ammonia.

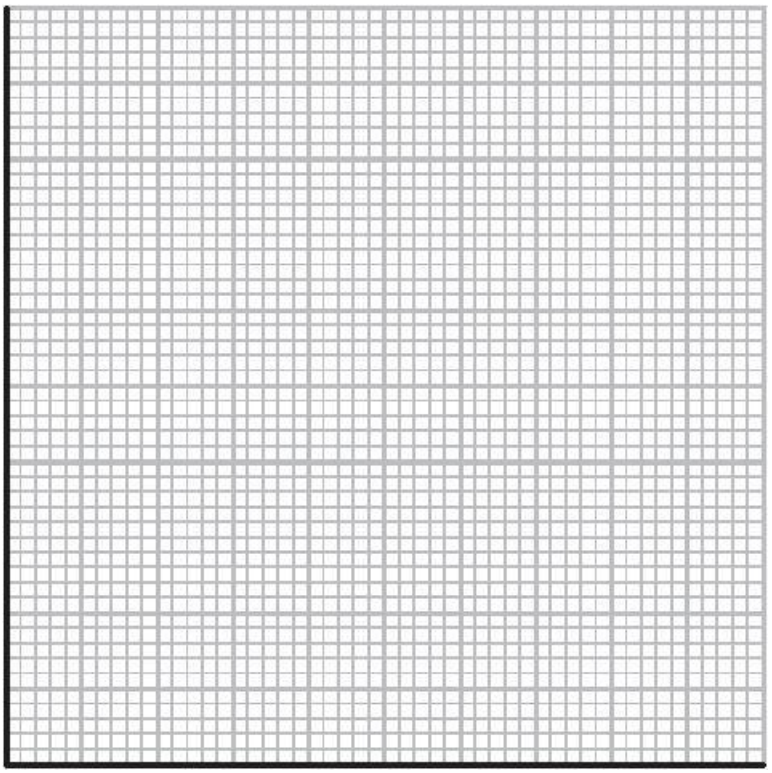
$$2NH_3(g) \rightarrow N_2(g) + 3H_2(g)$$

In an experiment, the following results were obtained.

Time /s	Partial pressure of ammonia / kPa
0	0.350
100	0.335
200	0.319
300	0.303
400	0.287
500	0.271

(i) Plot a graph of partial pressure of ammonia against time.

(2)



(ii) Deduce the rate equation for this reaction by using your graph in (i).
Justify your answer.

(2)

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(iii) Use the graph to calculate the rate constant. Include units in your answer.

(2)

(iv) Describe the stages in the catalytic decomposition of ammonia by tungsten.

(3)

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(Total for question = 9 marks)

Q8.

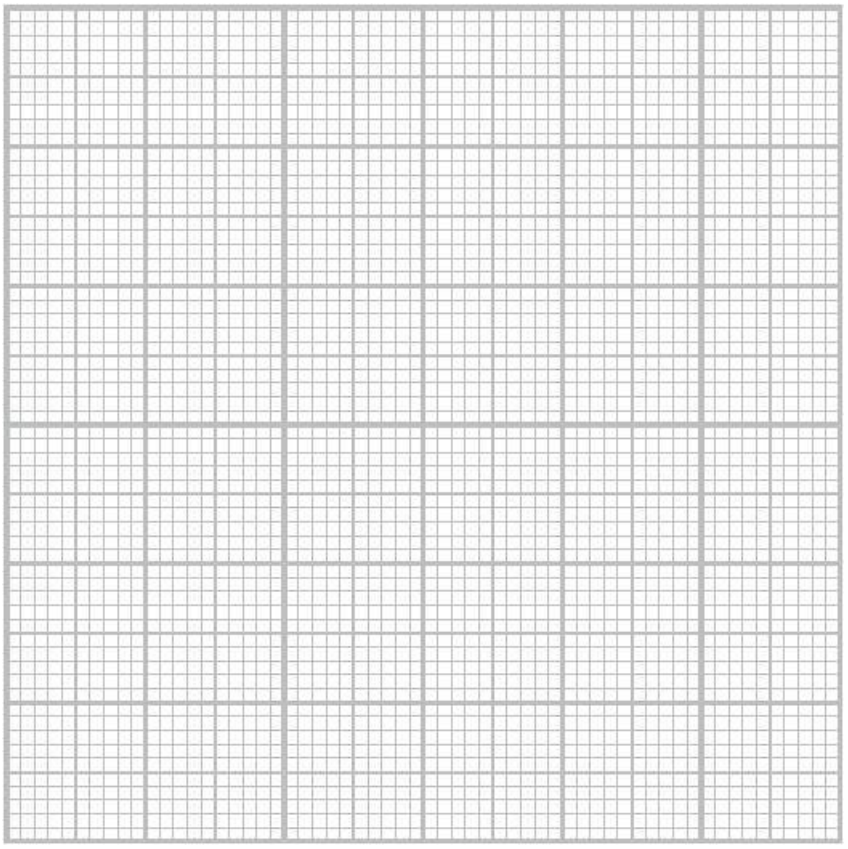
A series of experiments was carried out to determine the kinetics of the reaction between a chloroalkane, RCl, and potassium hydroxide in aqueous solution.
A large excess of the chloroalkane was used.

The data obtained are shown.

$[OH^-]$ / mol dm^{-3}	Time / s
0.00100	39
0.00200	31
0.00300	23
0.00400	16
0.00500	8

(a) Plot a graph of the concentration of the hydroxide ions against time.

(2)



(b) State the order with respect to hydroxide ions.
Justify your answer by reference to your graph in (a).

(2)

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(c) Deduce the type of mechanism occurring.
Justify your answer.

(2)

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(d) Give the classification of the chloroalkane in this reaction.

(1)

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(Total for question = 7 marks)

Q9.

This question is about the reaction kinetics of an 'iodine clock' reaction.

One example of an 'iodine clock' reaction that involves the iodate(V) ions and iodide ions in acidic solution is

$$IO_3^-(aq) + 5I^-(aq) + 6H^+(aq) \rightarrow 3I_2(aq) + 3H_2O(l)$$

A series of experiments was carried out by a student to determine the order of reaction with respect to iodate(V) ions. The concentrations of the iodide ions and the acid were in large excess and the volume of the iodate(V) solution was varied. The total volume of the reaction mixture was kept constant by the addition of suitable volumes of deionised water.
The following results were obtained:

Experiment Number	1	2	3	4	5	6
Volume of iodate(V) solution / cm^3	10.0	7.0	5.0	3.0	2.0	1.0
Time (t) / s	180	260	357	606	900	800
(1000/t) / s^{-1}	5.56				1.11	1.25

(i) In experiment 6, the student forgot to add deionised water to keep the total volume the same for each experiment.

State why the total volume should be kept the same.

(1)

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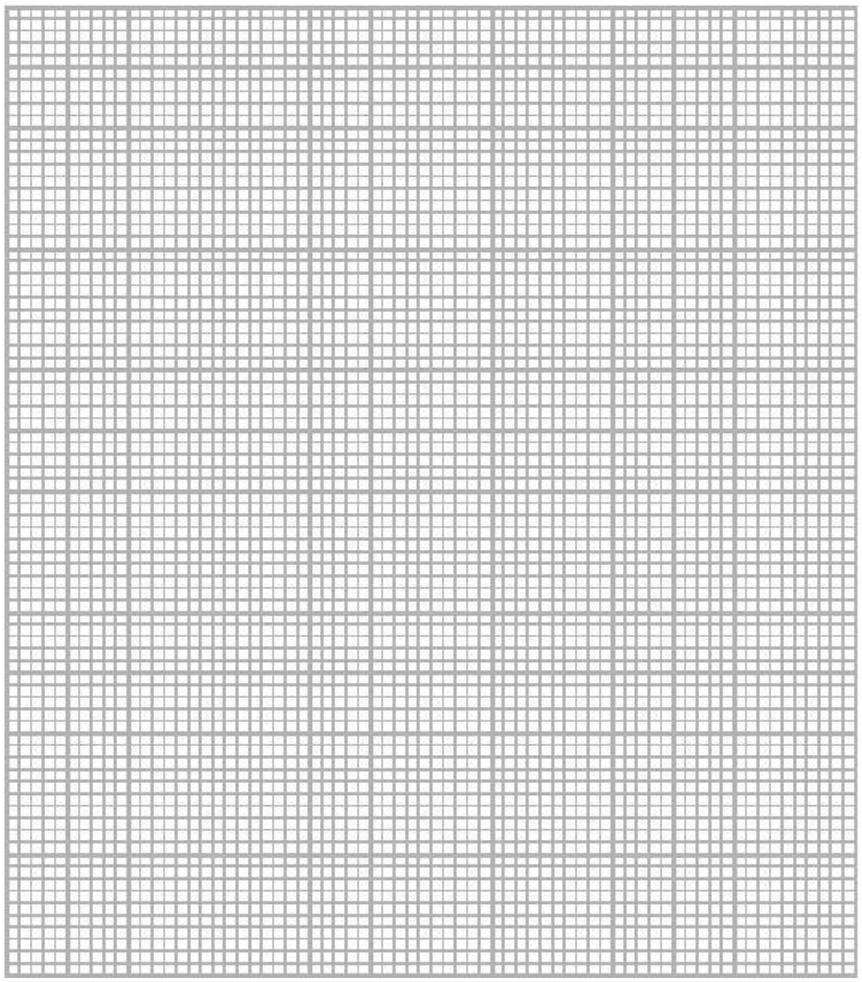
(ii) Complete the table and use the results from experiments 1, 2, 3, 4 and 5 to plot a graph of 1000/t against volume of iodate(V) ions.

(4)



(iii) Deduce the order of reaction with respect to the iodate(V) ions. Justify your answer.

(2)

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(Total for question = 7 marks)

Q10.

This question is about reaction kinetics.

The 'initial rates' method was used to investigate the orders of reaction with respect to reactants **X**, **Y** and **Z**. The table shows the results obtained.

Run	Initial concentration / mol dm^{-3}			Initial rate / mol dm^{-3} s^{-1}
	X	Y	Z	
1	0.00100	0.00300	0.00600	2.17×10^{-6}
2	0.00100	0.00600	0.00600	8.68×10^{-6}
3	0.00050	0.00600	0.00600	4.34×10^{-6}
4	0.00300	0.00300	0.00300	6.51×10^{-6}

(i) Calculate the orders with respect to **X**, **Y** and **Z**.

X .....
Y .....
Z .....

(ii) Give the rate equation for the reaction and hence calculate the rate constant, *k*, to an appropriate number of significant figures. Include units in your answer.

(4)

(Total for question = 4 marks)

Q11.

This question is about transition metals.

Manganate(VII) ions, MnO_4^-, react with ethanedioate ions in acid solution.

$$2MnO_4^- + 5C_2O_4^{2-} + 16H^+ \rightarrow 2Mn^{2+} + 10CO_2 + 8H_2O$$

The reaction starts slowly, the rate of reaction then increases, before it decreases again.

Explain this sequence.

(3)

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(Total for question = 3 marks)

Q12.

This question is about the reaction kinetics of an 'iodine clock' reaction.

One example of an 'iodine clock' reaction that involves the iodate(V) ions and iodide ions in acidic solution is

$$IO_3^-(aq) + 5I^-(aq) + 6H^+(aq) \rightarrow 3I_2(aq) + 3H_2O(l)$$

A different version of the 'iodine clock' reaction involves mixing hydrogen peroxide with aqueous solutions of potassium iodide, sodium thiosulfate and starch.

The main reaction is

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

The reaction is first order with respect to hydrogen peroxide and iodide ions but zero order with respect to hydrogen ions.

(i) In one experiment, the following data were obtained:

Reactants	Initial concentration / mol dm^{-3}
$H_2O_2(aq)$	1.50×10^{-3}
$I^-(aq)$	2.10×10^{-3}
$H^+(aq)$	2.10×10^{-3}

Initial rate = 1.24×10^{-3} mol dm^{-3} s^{-1}

Write the rate equation and hence deduce the value of the rate constant, *k*, from these data. Include units and give your answer to an appropriate number of significant figures.

(2)

(ii) Explain the purpose of the starch present in the reaction mixture when starch is neither in the rate equation, nor in the reaction equation.

(2)

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(Total for question = 4 marks)

Q13.

Nitrogen monoxide and chlorine react together to form nitrosyl chloride.

$$2NO(g) + Cl_2(g) \rightarrow 2NOCl(g)$$

The rate equation for the formation of nitrosyl chloride is

Rate = $k[NO]^2[Cl_2]$

(i) Complete the table by adding the missing values.

Experiment	[NO] / mol dm^{-3}	[Cl_2] / mol dm^{-3}	Rate / mol dm^{-3} s^{-1}
1	0.122	0.241	1.09×10^{-2}
2		0.482	8.72×10^{-2}
3	0.366		4.91×10^{-2}

(2)

(ii) Calculate the rate constant, *k*, using data from Experiment 1.

Include units with your answer.

(3)

(iii) Explain how using a catalyst increases the rate constant, *k*.

(2)

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(iv) The heterogeneous catalyst palladium was suggested for use in this reaction.

Explain how impurities in the gaseous reactants could make the catalyst less effective.

(3)

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(Total for question = 10 marks)

Q14.

This question is about carbon monoxide, CO, which is a toxic and colourless gas used widely in the chemical industry.

Haemoglobin (Hb) found in red blood cells reacts almost irreversibly with carbon monoxide. Initial rate experiments were carried out to investigate the effect of the concentrations of Hb and CO on the rate of this reaction.

Experiment	[Hb] / mol dm^{-3}	[CO] / mol dm^{-3}	Rate / mol dm^{-3} s^{-1}
1	2.09×10^{-6}	1.40×10^{-6}	8.20×10^{-7}
2	4.18×10^{-6}	1.40×10^{-6}	1.64×10^{-6}
3	3.26×10^{-6}	2.80×10^{-6}	2.56×10^{-6}

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

(1)

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(ii) Determine the order with respect to carbon monoxide using your answer to (i) and the data in the table.
Justify your answer.

(2)

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(iii) Write the rate equation for this reaction using your answers to (i) and (ii).

(1)

(iv) Calculate the rate constant, *k*, for the reaction, using the data from Experiment 1 and the rate equation from (iii).
Include units in your answer.

(3)

(Total for question = 7 marks)

Q15.

This question is about the reaction kinetics of an 'iodine clock' reaction.

One example of an 'iodine clock' reaction that involves the iodate(V) ions and iodide ions in acidic solution is

$$IO_3^-(aq) + 5I^-(aq) + 6H^+(aq) \rightarrow 3I_2(aq) + 3H_2O(l)$$

State why the order of reaction with respect to iodide ions cannot be five, even though 5 mol of iodide ions are shown in the equation.

(1)

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(Total for question = 1 mark)

Q16.

This question is about reaction kinetics.

The kinetics of the 'bromine clock' were investigated and the rate equation was found to be

$$\text{Rate} = k[BrO_3^-][Br^-][H^+]^2$$

(i) What is the overall reaction order?

(1)

- ☒ **A** First
- ☒ **B** Second
- ☒ **C** Third
- ☒ **D** Fourth

(ii) Calculate the concentration of bromide ions required to produce a reaction rate of 4.08×10^{-3} mol dm^{-3} s^{-1} at 298 K given that

$k = 8.00\ dm^9\ mol^{-3}\ s^{-1}$

$[BrO_3^-] = 0.200$ mol dm^{-3}

$[H^+] = 0.100$ mol dm^{-3}

(2)

(Total for question = 3 marks)

Q17.

A bromoalkane, RBr, reacts with aqueous hydroxide ions in a nucleophilic substitution reaction.

$$RBr + OH^- \rightarrow R—OH + Br^-$$

This reaction is first order with respect to the bromoalkane and the rate equation is

$$\text{rate} = k[RBr]^1[OH^-]^x$$

where x is the order of the reaction with respect to hydroxide ions.

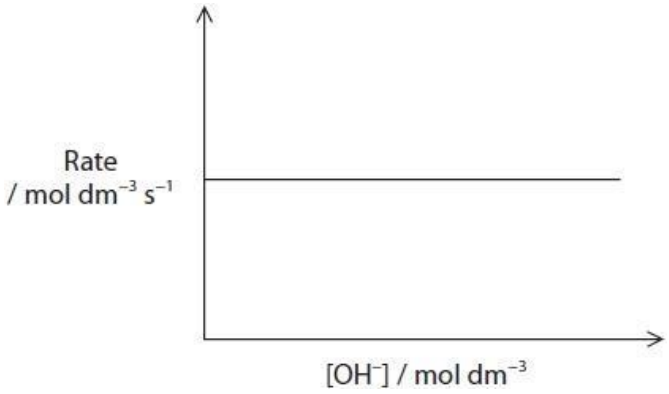
In an experiment, a sample of the bromoalkane was added to a large excess of aqueous sodium hydroxide and the concentration of the bromoalkane was determined at regular time intervals.

Results

Time / s	[RBr] / mol dm^{-3}
0	0.100
30	0.065
60	0.042
90	0.028
120	0.019
150	0.014

The experiment was repeated using equal concentrations of RBr and varying the concentration of hydroxide ions.

A graph was plotted of the results.



(i) Deduce the value of x in the rate equation.

$$\text{rate} = k[\text{RBr}]^1[\text{OH}^-]^x$$

(1)

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(ii) Give the mechanism for the reaction that is consistent with the orders of reaction with respect to R—Br and hydroxide ions.
Include curly arrows and relevant lone pairs.

(3)

(Total for question = 4 marks)

Q18.

A bromoalkane, RBr, reacts with aqueous hydroxide ions in a nucleophilic substitution reaction.

$$RBr + OH^- \rightarrow R—OH + Br^-$$

This reaction is first order with respect to the bromoalkane and the rate equation is

$$\text{rate} = k[RBr]^1[OH^-]^x$$

where x is the order of the reaction with respect to hydroxide ions.

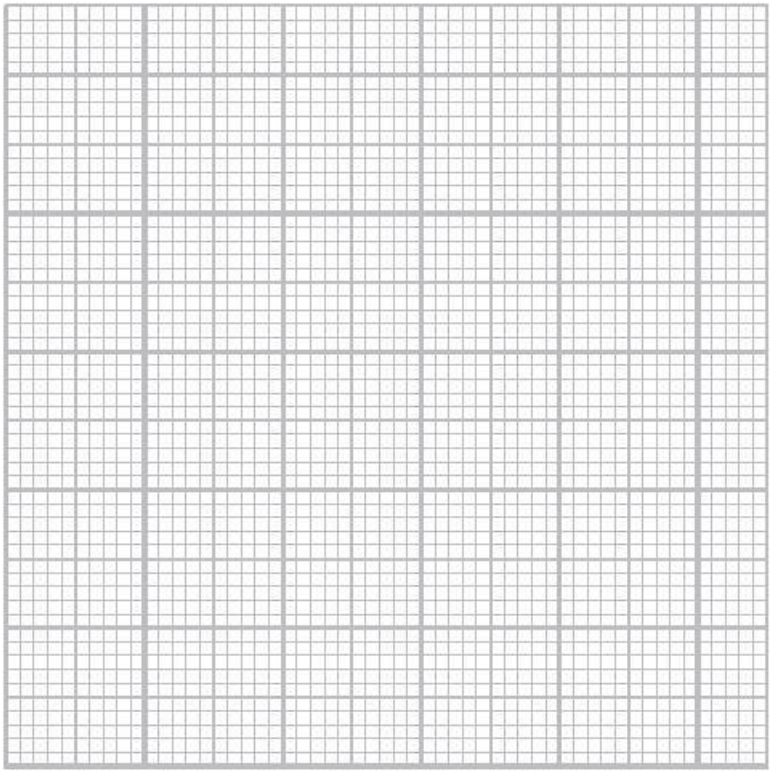
In an experiment, a sample of the bromoalkane was added to a large excess of aqueous sodium hydroxide and the concentration of the bromoalkane was determined at regular time intervals.

Results

Time / s	[RBr] / mol dm^{-3}
0	0.100
30	0.065
60	0.042
90	0.028
120	0.019
150	0.014

(i) Plot a graph of [RBr] against time.

(3)



(ii) Explain how the graph shows that the reaction is first order with respect to RBr.
Include the values of two consecutive half-lives.

You **must** show your working for the half-lives on the graph.

(2)

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(Total for question = 5 marks)