IB · HL · Chemistry

6 hours 41 questions

Structured Questions

How Fast? The Rate of Chemical Change

Rate of Reaction / Measuring Rates of Reaction / Collision Theory / Factors Affecting Rates of Reaction / Activation Energy / Energy Profiles With & Without Catalysts / Maxwell-Boltzmann Distribution Curves / Rate Equation (HL) / Reaction Orders (HL) / The Rate Constant (HL) / Reaction Mechanisms (HL) / Molecularity (HL) / The Arrhenius Equation (HL) / Determining Activation Energy & the Arrheni...

Easy (12 questions)	/90
Medium (17 questions)	/162
Hard (12 questions)	/102
Total Marks	**/354**

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Easy Questions

1 (a) Describe kinetic theory in relation to energy and temperature.

(2 marks)

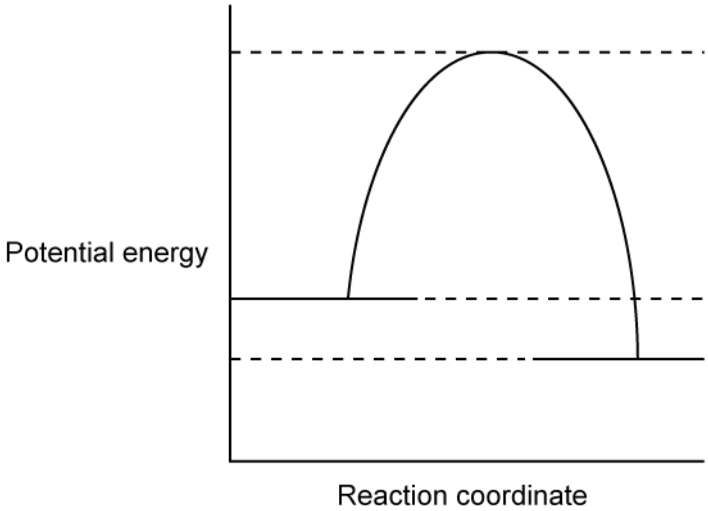
(b) State what is required for a collision to result in a reaction.

(2 marks)

(c) State the meaning of activation energy (E_a).

(2 marks)

(d) Label the activation energy on the energy profile diagram below.



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(1 mark)

State three ways of monitoring concentration changes in a reaction.

2 (a) **(3 marks)**

A reaction is monitored by measuring the volume of a gas produced every 10 seconds. State an appropriate unit to use.

(b) **(1 mark)**

Sketch a graph to show the volume of gas produced during the course of an experiment against the time taken.

(c) **(4 marks)**

State the effect that increasing concentration has on the rate of a reaction.

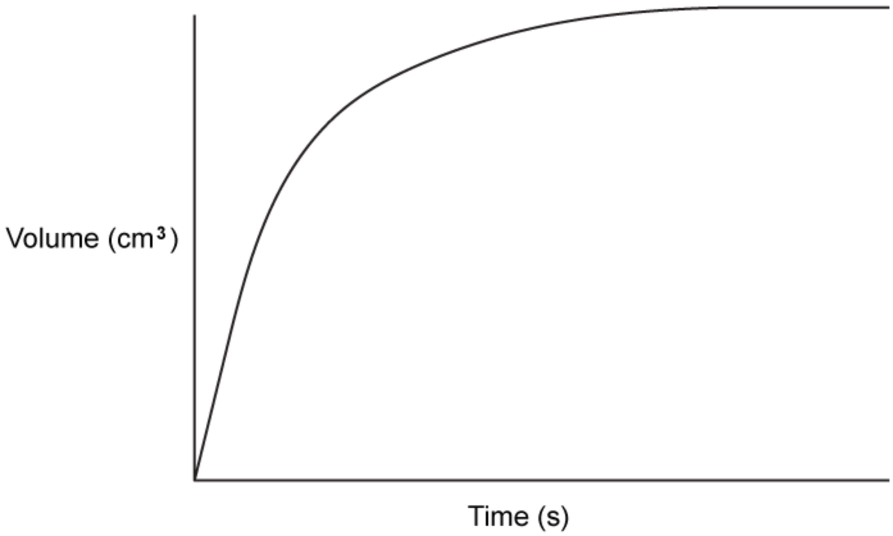
(d) **(1 mark)**

3 (a) State the effect that increasing temperature has on the rate of a reaction.

....

(1 mark)

(b) Sketch a line on the graph to show the same reaction occurring at a higher temperature.



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(3 marks)

(c) State two variables that need to be controlled when investigating the effect of temperature on rate in the following reaction:

$$2HCl\ (aq) + Mg\ (s) \rightarrow MgCl_2\ (aq) + H_2\ (g)$$

....

....

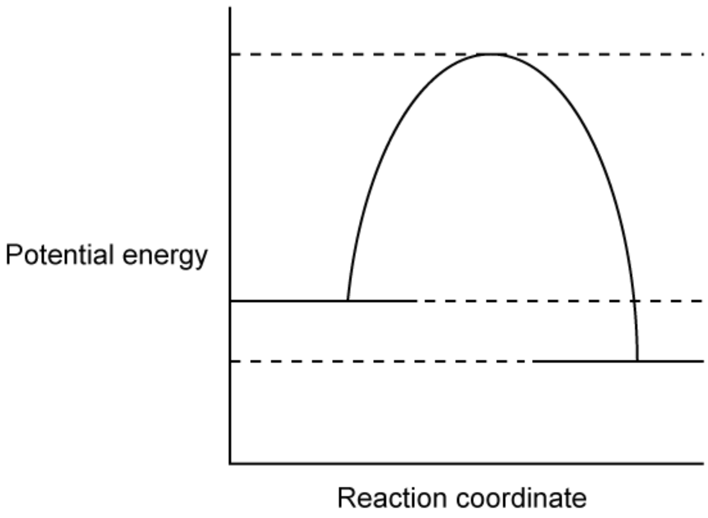
(2 marks)

(d) Suggest an appropriate piece of equipment to use to measure the volume of H_2 gas produced in the reaction between HCl and Mg.

....

(1 mark)

4 (a) Sketch a line on the potential energy profile diagram to show the pathway for the same reaction, but with a catalyst.

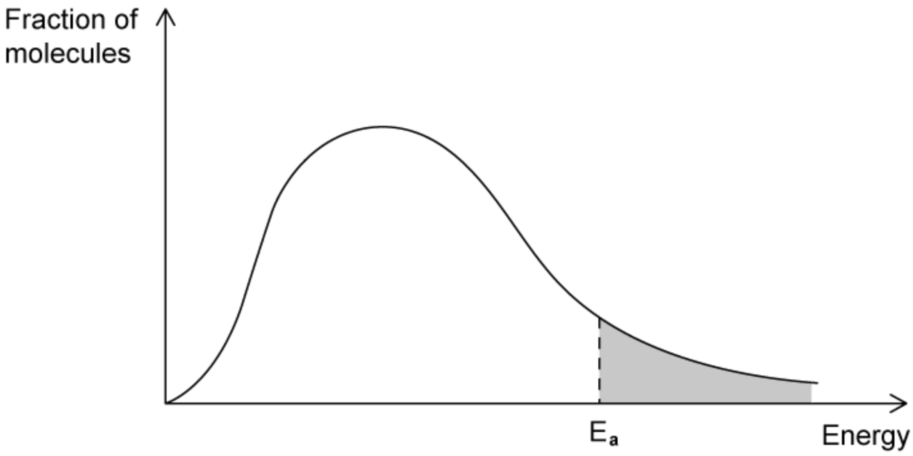


(2 marks)

(b) Explain how catalysts work.

(3 marks)

(c) Maxwell-Boltzmann distribution is shown below:



i) Draw a line on the Maxwell-Boltzmann curve below to show the effect of adding a catalyst.

[1]

ii) Shade in the area representing the number of particles that can react with the catalyst present.

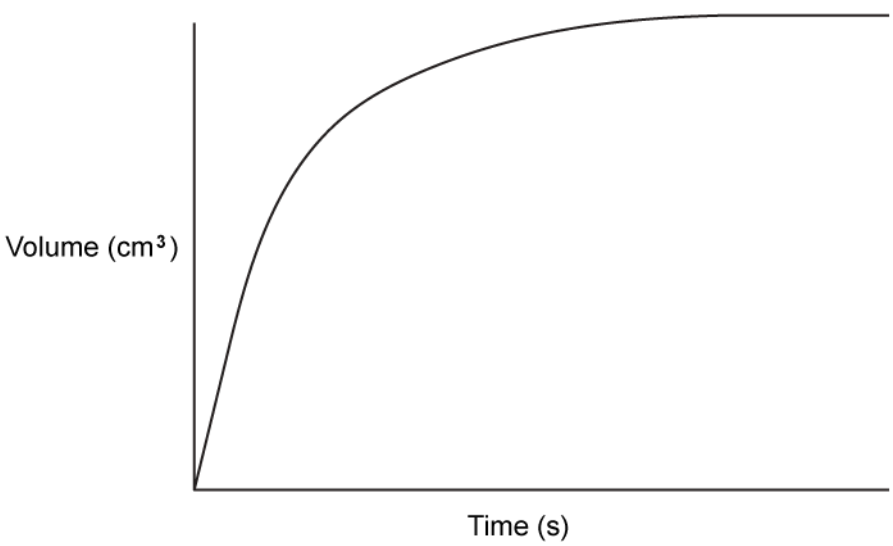
[1]

....

....

(2 marks)

(d) Sketch a line on the graph to show the same reaction occurring with a catalyst.



(3 marks)

5 (a) Outline two ways a rate of a reaction can be expressed and state the units for rate of reaction.

(2 marks)

(b) Explain what is meant by the *order* of a reaction and how it may be determined.

(2 marks)

(c) Carbon monoxide and chlorine react together to make phosgene, $COCl_2$. The equation for the reaction is given below:

$$CO\ (g) + Cl_2\ (g) \rightarrow COCl_2\ (g)$$

A possible rate equation for the reaction is:

$$\text{rate} = k[CO\ (g)]^2[Cl_2\ (g)]^{½}$$

What is the overall reaction order?

(1 mark)

(d) Determine the units of the rate constant, *k*, for the following rate equation:

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

(1 mark)

6 (a) The rate of hydrolysis of sucrose under acidic conditions can be determined experimentally. The following data was obtained:

Experiment	Initial [HCl] / mol dm^{-3}	Initial [sucrose] / mol dm^{-3}	Rate of reaction / mol dm^{-3} s^{-1}
1	0.10	0.10	0.024
2	0.10	0.15	0.036
3	0.20	0.10	0.048

Determine the order of reaction with respect to HCl.

(1 mark)

(b) Determine the order of reaction with respect to sucrose.

(1 mark)

(c) Determine the overall order of reaction, write the rate expression and state the units of the rate constant, *k*.

(3 marks)

(d) Determine the following:

i) The value of k, using Experiment 1

[1]

ii) The rate of reaction if the concentration of HCl and sucrose are both 0.20 mol dm^{-3}

[1]

(2 marks)

7 **(a)** Sketch graphs of a first order and second order reaction of concentration against time.

(2 marks)

(b) Draw sketch graphs for a first and second order reaction of rate against concentration.

(2 marks)

(c) Deduce the units of the rate constant, k, for a first order reaction.

(1 mark)

(d) State, with a reason, how the value of the rate constant, *k*, varies with increased temperature for a reaction.

(4 marks)

8 (a) State what is meant by the terms *rate determining step* and *molecularity* in a chemical reaction.

(2 marks)

(b) The following reaction mechanism has been proposed for the formation of nitrosyl bromide, NOBr, from nitrogen monoxide and bromine:

Step 1: $NO + NO \rightarrow N_2O_2$

Step 2: $N_2O_2 + Br_2 \rightarrow 2NOBr$

Deduce the overall reaction equation and comment on the molecularity of Step 1 and 2.

(2 marks)

(c) A student proposes an alternative one step mechanism for the formation of nitrosyl bromide.

$NO + NO + Br_2 \rightarrow NOBr_2$

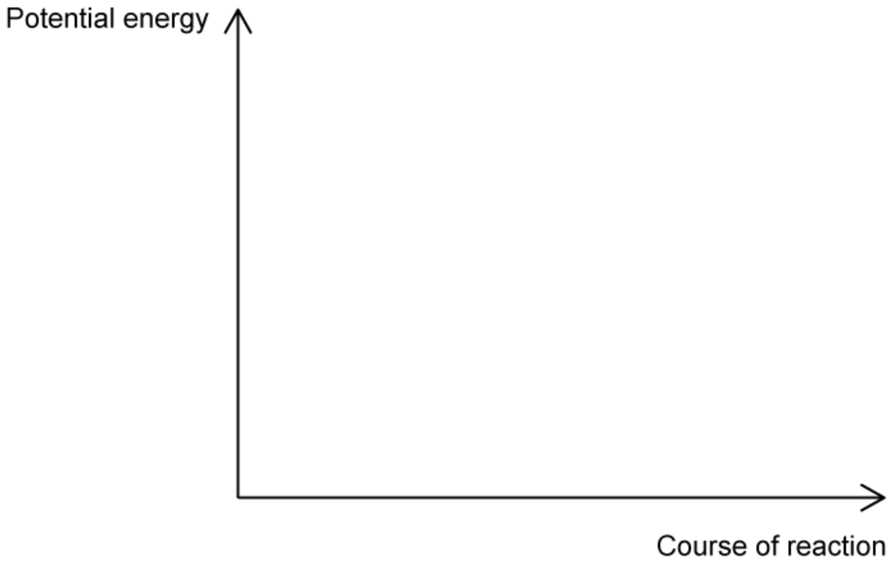
Explain why this mechanism is not likely to take place.

(2 marks)

(d) State the role of N_2O_2 in the mechanism in part b).

(1 mark)

9 (a) Draw a labelled diagram, on the follow grid, showing a potential energy profile in a two step reaction. The second step is the slow step of the reaction.



(3 marks)

(b) State which step of the mechanism in a) is affected by the addition of a catalyst.

(1 mark)

(c) A reaction mechanism is shown below.

Step 1: $NO_2 + NO_2 \rightarrow NO + NO_3$ (slow)

Step 2: $NO_3 + CO \rightarrow NO_2 + CO_2$ (fast)

Deduce the overall reaction equation and the rate equation for the reaction.

(2 marks)

(d) State the overall reaction order in part c) and state the units of the rate constant.

(2 marks)

10 (a) The Arrhenius equation can be written as:

$$k = Ae^{\frac{-E_a}{RT}}$$

State what each of the following terms represents, including units where applicable.

- A
- E_a
- R
- T

(5 marks)

(b) Rearrange the Arrhenius equation given in part (a) to make A the subject.

(1 mark)

(c) State how the rate constant, k varies with temperature, T.

(1 mark)

(d) State how the activation energy, Ea, varies with rate constant, k.

(1 mark)

11 (a) The Arrhenius equation can also be written in natural logarithmic forms.

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

A plot of ln k against $\frac{1}{T}$ gives a straight-line graph of the type y = mx + c.

Complete the table below which relates the terms from the natural logarithmic Arrhenius equation to the equation of a straight line.

Straight-line term	**Arrhenius term**
y	ln k
m	
x	
c	

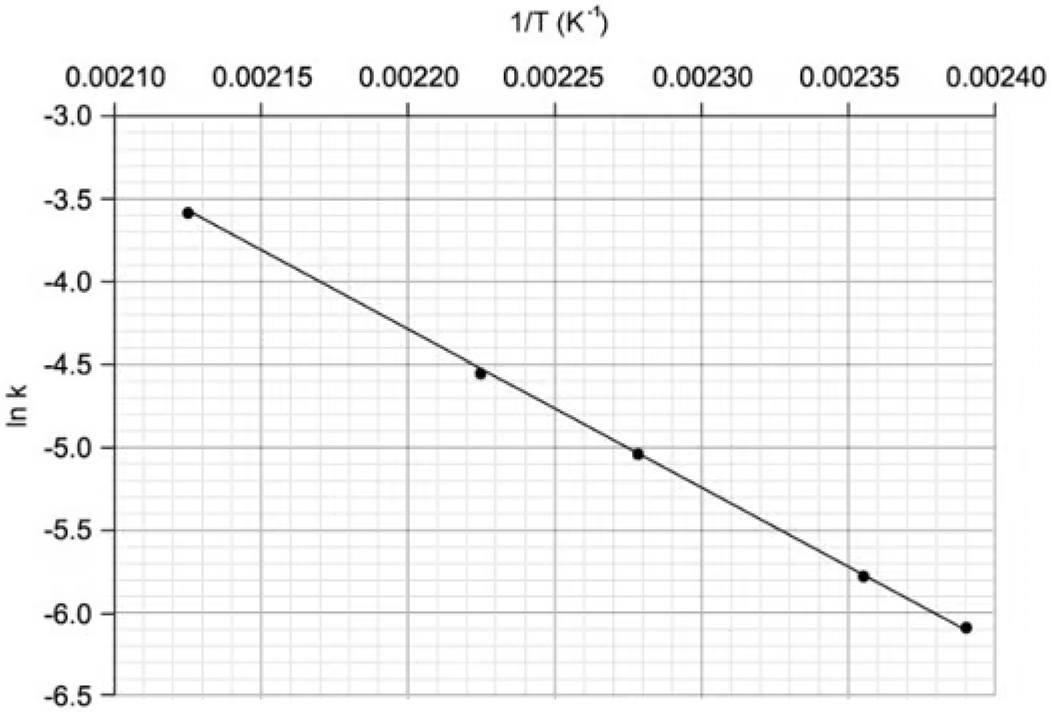
....

....

....

(3 marks)

(b) A graph of ln k against $\frac{1}{T}$ is shown below.



Calculate the gradient of the straight line.

....

....

(2 marks)

(c) Using section 2 of the data booklet, calculate the activation energy, E_a for the graph in part b).

....

(1 mark)

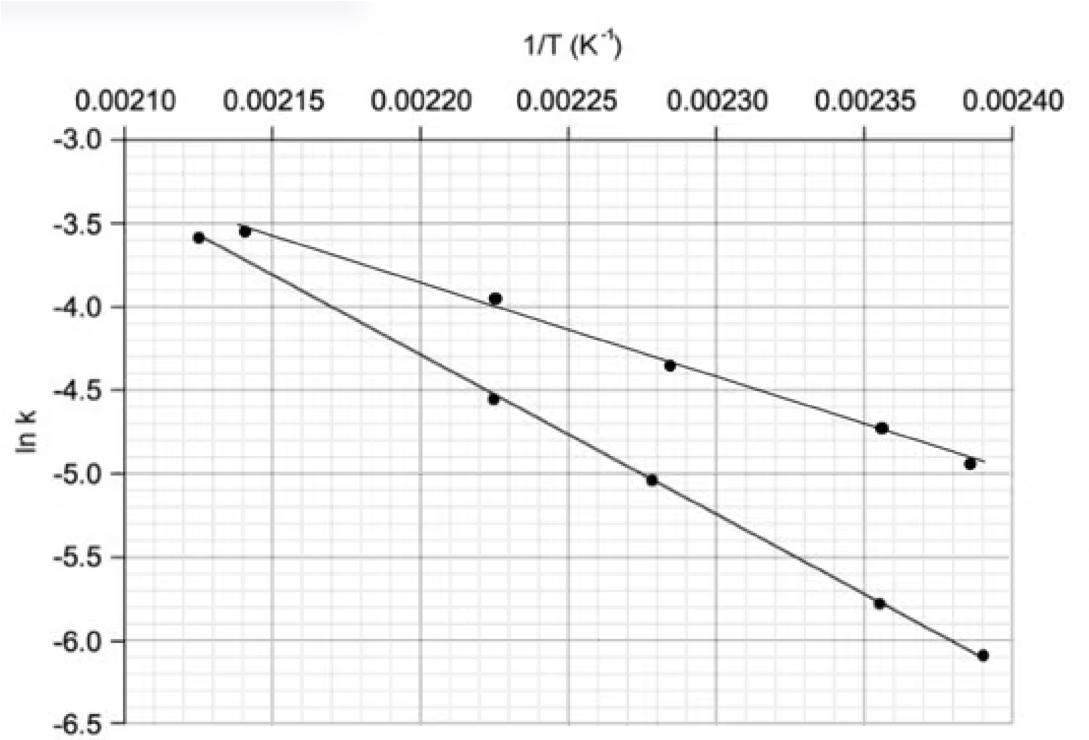
(d) Calculate the frequency factor, *A*, for the graph in part b) to 2 decimal places.

....

....

(2 marks)

12 **(a)** Arrhenius plots for two reactions with different activation energies are shown below.



State which plot shows the reaction with the greatest activation energy.

....

(1 mark)

(b) The temperature of both reactions from part a) is increased from 20° to 45°.

Using section 1 of the data booklet, determine which of the reactions will experience the largest change in the rate of reaction.

....

(1 mark)

(c) The decomposition of hydrogen peroxide into water and oxygen occurs at a slow rate with a rate constant of k = 6.42×10^{-4} mol dm^{-3} s^{-1} and at a temperature of 290 K.

When the temperature is increased to 340 K the rate constant k = 6.47×10^{-2} mol dm^{-3} s^{-1}.

Using sections 1 and 2 of the data booklet, calculate the activation energy for this reaction.

(2 marks)

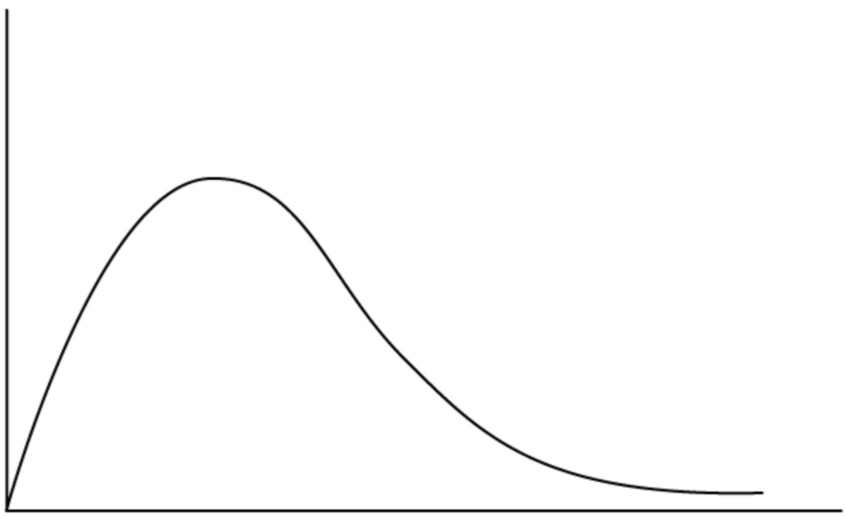
Medium Questions

1 (a) In any chemical reaction, the particles will all be moving around in different directions, at different speeds, with different amounts of energy.

A Maxwell-Boltzmann distribution is a graph which shows the distribution of energy amongst particles within a chemical reaction.

Figure 1 below shows the Maxwell-Boltzmann distribution in a sample of a gas at a fixed temperature, T_1.

Figure 1



i) Label the x and y axes of the graph.

[2]

ii) Sketch a distribution for this same sample of gas, at a higher temperature, and label it as T_2.

[2]

(4 marks)

(b) State why a Maxwell-Boltzmann distribution curve always starts at the origin and what the area under the curve represents.

(2 marks)

(c) Chemical reactions take place at different speeds. For a chemical reaction to take place, particles must collide with each other in the correct orientation and with sufficient energy.

i) Explain why most collisions between particles in the gas phase do not result in a reaction taking place.

[1]

ii) State and explain one way that the rate of reaction could be increased, other than by increasing the temperature.

[2]

(3 marks)

(d) Give one reason why a reaction may be slow at room temperature.

(1 mark)

2 (a) State the meaning of the term *rate of reaction*.

....

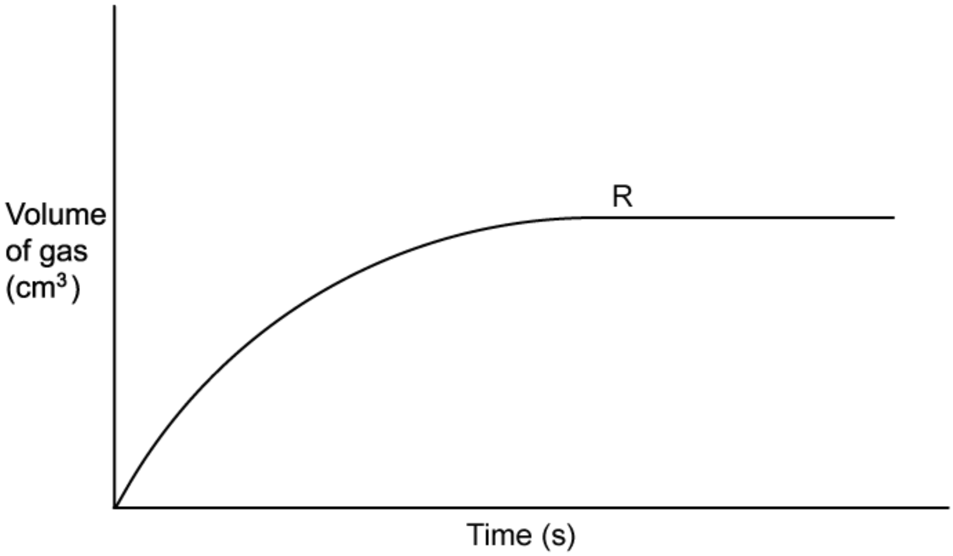
(1 mark)

(b) A group of students were completing a practical, investigating the factors which affect the rate of the chemical reaction shown below.

$$\mathbf{A\ (s) + B\ (aq) \rightarrow C\ (g)}$$

The students collected the gas produced and plotted the graph shown in **Figure 1**.

Figure 1



i) State and explain what the letter R represents on the students graph in **Figure 1.**

[1]

ii) In the original reaction above, the students used 0.5 g of **A** and 50 cm^3 of 1.0 mol dm^{-3} **B**.

Sketch a curve on the graph to show how the total volume of gas collected would change if the students still used 0.5 g of **A,** but used 50 cm^3 of 2.0 mol dm^{-3} of **B**.

[2]

....

(3 marks)

(c) Explain why the gradient of the curve in part (b) decreases as the time of the reaction progresses.

(2 marks)

(d) Another way to increase the rate of reaction is to increase the temperature.

Explain why a small increase in temperature has a large effect on the initial rate of a chemical reaction.

(2 marks)

3 **(a)** The decomposition of hydrogen peroxide into water and oxygen is a very slow chemical reaction.

Write the equation for the decomposition of hydrogen peroxide.

....

(1 mark)

(b) The rate of decomposition of hydrogen peroxide can be found by collecting and measuring the volume of gas formed at specific time intervals.

i) Draw a labelled diagram to show the apparatus that you would use to collect and measure the volume of gas formed during this reaction.

[2]

ii) Explain how you would use the results to determine the initial rate of the reaction.

[3]

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....

(5 marks)

(c) The decomposition of hydrogen peroxide is a slow reaction, so a catalyst is often added to speed up the rate of the reaction. Catalysts are used in many chemical reactions to increase the rate.

The following shows a two-step reaction mechanism of a chemical reaction, where a catalyst, **X** is used.

STEP 1: $W + X \rightarrow Y + Z$

STEP 2: $Y + W \rightarrow Z + A + X$

OVERALL REACTION: $2W \rightarrow 2Z + A$

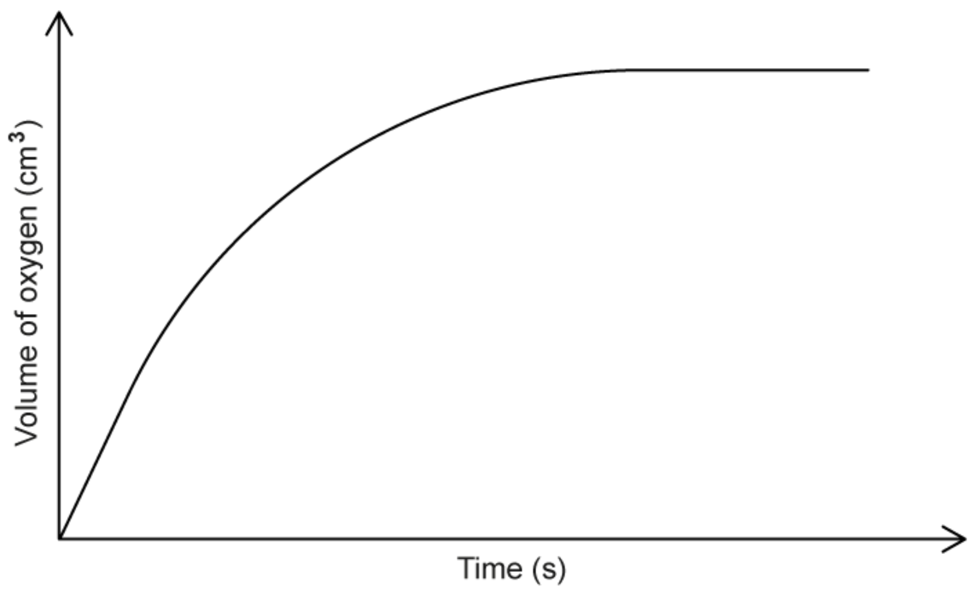
Give a reason, other than the rate of reaction increasing, why it can be deduced from the three equations above that X is a catalyst.

....

(1 mark)

(d) The graph shown below represents the decomposition of hydrogen peroxide.

Figure 1



The graph starts to level out as the reaction slows down.

State why the rate of the reaction slows down over time.

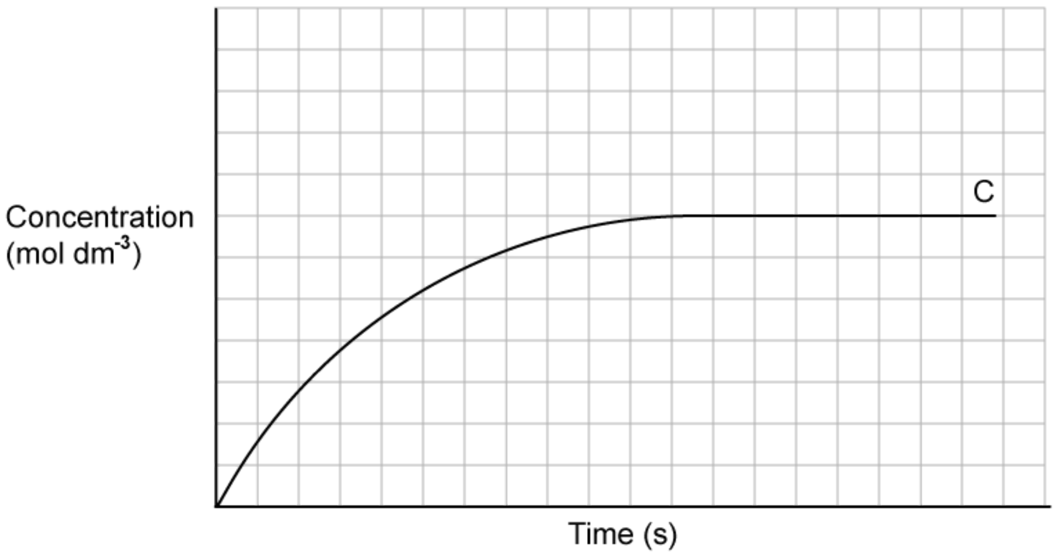
(1 mark)

4 (a) During the following reaction, **A** and **B** react together to produce **C**.

$$\mathbf{A} + 2\mathbf{B} \rightleftharpoons \mathbf{C}$$

Figure 1 shows the production of C over time.

Figure 1



i) Sketch a graph to show what happens to **A** and **B** during the progress of the reaction.

[1]

ii) On your graph, write the letter **E** at the point at which an equilibrium is first established.

[1]

(2 marks)

(b) In the reaction in part (a), large pieces of **A** were used.

Use collision theory to explain what would happen to the rate of the reaction if powdered **A** was used instead of large pieces.

(3 marks)

(c) In a different reaction, gaseous W and X were added together to produce Y and Z as shown in the equation below:

$$\mathbf{2W\,(g) \; + \; X\,(g) \; \rightarrow \; Y\,(g) \; + 2Z\,(g)}$$

A catalyst was added to speed up the rate of reaction.

i) Sketch a Maxwell-Boltzmann distribution on the axes below in **Figure 2** to show the distribution of molecular energies at a constant temperature with **and** without a catalyst.
Use E_a to label the activation energy without a catalyst and E_c to label the activation energy with a catalyst.

[3]

ii) Explain what your distribution shows.

[3]

Figure 2

(6 marks)

(d) Some changes were made individually to the experiment completed in part (c).

Consider your Maxwell-Boltzmann distribution curve from part (c). For each of the changes in parts (i), (ii) and (iii) below, state and explain the effect that the change would have on:

- The area under the curve
- The value of the most probable energy of the molecules (E_{mp})
- The proportion of molecules with energy greater than or equal to E_a

i) The temperature of the original reaction is increased, but no other changes are made.

[2]

ii) The number of molecules in the original reaction mixture is increased, but no other changes are made.

[2]

iii) A catalyst is added to the original reaction mixture, but no other changes are made.

[2]

(6 marks)

5 (a) Iodine reacts with propanone in an acid catalyzed reaction, according to reaction equation below.

$$CH_3COCH_3\ (aq) + I_2\ (aq) \rightarrow CH_3COCH_2I\ (aq) + H^+\ (aq) + I^-\ (aq)$$

Suggest how the change in concentration of iodine could be used to determine the rate of the above reaction.

....

(1 mark)

(b) A group of students completed the iodination of propanone reaction using the same acid catalyst, but with different concentrations. The results achieved are shown in the table below:

Table 1

Concentration of acid, $[H^+]$ / moldm^{-3}	**Relative Rate of Iodination Reaction**
0.100	0.0046
0.200	0.0092
0.300	0.0138

Use the table to state and explain the relationship between the concentration of acid used in the reaction and the rate.

....

....

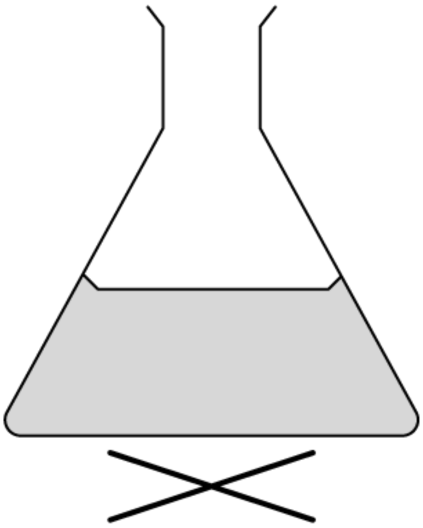
(2 marks)

(c) Sodium thiosulfate and hydrochloric acid will react together readily, as shown by the equation below:

$$Na_2S_2O_3 + 2HCl \rightarrow 2NaCl + S + SO_2 + H_2O$$

This reaction is often referred to as the 'disappearing cross' experiment. The cross disappears when viewed from above because the solution turns cloudy as a sulfur precipitate is formed, covering the cross.

Figure 1



The speed of the reaction can be increased, by raising the temperature of the sodium thiosulfate solution in the reaction. The thiosulfate solution is heated to different temperatures before the acid is added, and the time it takes for the cross to disappear is recorded. The times can then be compared.

Suggest one reason why the value for the rate of reaction when a higher temperature was used may be less accurate than at a lower temperature.

....

(1 mark)

(d) Collision theory can be used to explain why different factors affect the rate of a chemical reaction.

Describe collision theory.

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(3 marks)

6 (a) A student carried out a metal displacement reaction between zinc powder and copper(II) sulfate solution. The equation for the reaction is

$$Zn\ (s) + CuSO_4\ (aq) \rightarrow ZnSO_4\ (aq) + Cu\ (s)$$

3.78 g of zinc powder was added to 50.0 cm^3 of 0.250 $moldm^{-3}$ copper(II) sulfate solution.

Determine the limiting reagent showing your working.

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(3 marks)

(b) The reaction between the zinc and copper sulfate was carried out in a polystyrene cup and the temperature change was measured using a temperature probe. The maximum temperature rise the student recorded was 8.5 ^{o}C.

Using sections 1 and 2 of the data booklet, calculate the enthalpy change, ΔH, for the reaction, in kJ.

Assume that all the heat evolved was absorbed by the solution, and that the density and specific heat capacity of the copper(II) sulfate solution are the same as pure water.

....

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(2 marks)

(c) State **two** further assumptions made in the calculation of ΔH.

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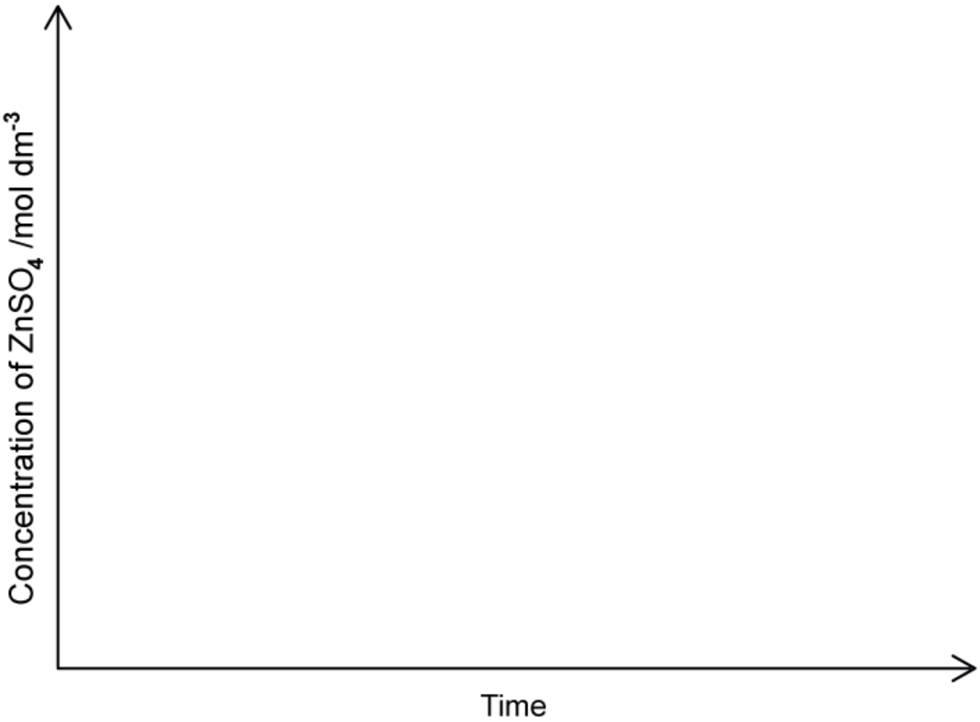
....

(2 marks)

(d) Using **Figure 1**, sketch a graph of the concentration of zinc sulfate, $ZnSO_4$ (aq), versus time and show how the graph may be used to find the initial rate of reaction.

Figure 1



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(4 marks)

7 (a) A student investigated the rate of decomposition of hydrogen peroxide, H_2O_2, at a temperature of 45 ^{o} The decomposition reaction occurs in the presence of a catalyst, MnO_2.

$$2H_2O_2\ (aq) \xrightarrow{MnO_2} O_2\ (g) + 2H_2O\ (l)$$

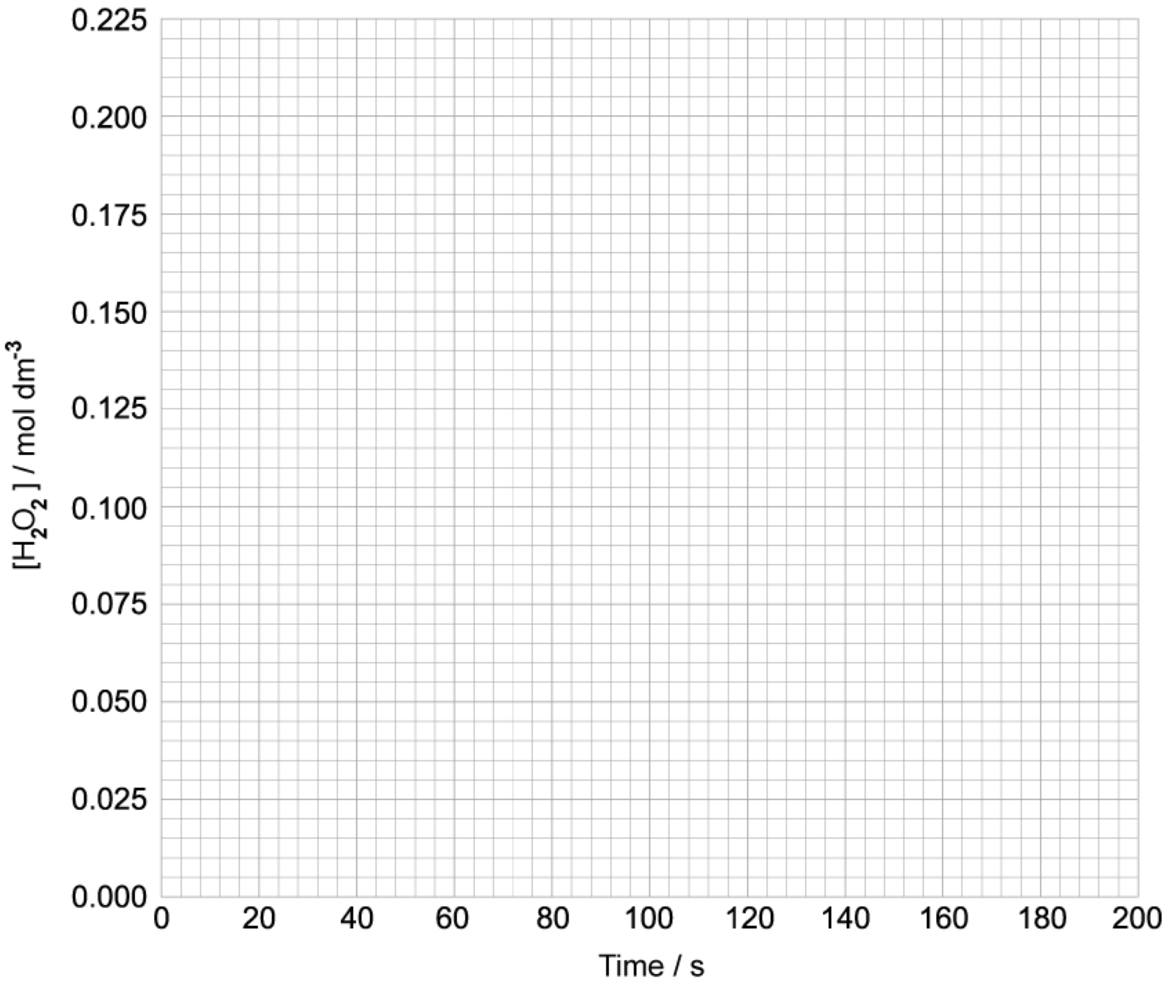
The results she obtained are shown in **Table 1** below.

Table 1

Time / s	**Concentration of H_2O_2 / moldm^{-3}**	**Time / s**	**Concentration of H_2O_2 / moldm^{-3}**
0	0.200	120	0.068
20	0.155	140	0.063
40	0.124	160	0.058
60	0.102	180	0.055
80	0.085	200	0.052
100	0.075		

Plot a graph on the axes below in **Figure 1** and from it determine the rate of reaction after 60 s.

Figure 1



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(4 marks)

(b) On the same graph sketch the shape obtained if the student had carried out the same reaction at 60 ^{o}C. Explain the shape of the graph at 60 ^{o}C.

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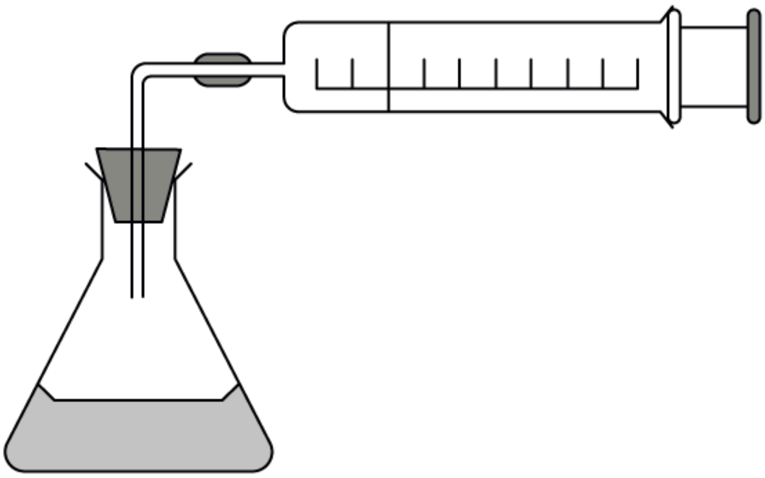
(3 marks)

(c) The decomposition of hydrogen peroxide can be investigated by measuring the volume of oxygen given off using the apparatus shown in **Figure 2**.

Figure 2



i) Explain why the volume of oxygen given off can be used as a measure of the concentration of hydrogen peroxide.

[1]

ii) Suggest one limitation of using the apparatus used in **Figure 2**.

[1]

iii) Suggest an alternative method of measuring the rate of reaction.

[1]

(3 marks)

(d) Two students decide to measure the rate of decomposition for H_2O_2 using the change in mass as oxygen escapes from the reaction container.

One student says that they should use a three decimal place rather than two decimal place balance because it will make their results more accurate. The second student disagrees and says it will make their results more precise, but not more accurate.

Which student is correct?

....

....

(2 marks)

8 (a) For the reaction below, consider the following experimental data.

$$X\ (aq) + Y\ (aq) \rightarrow Z\ (aq)$$

Experiment	Initial [X] / mol dm^{-3}	Initial [Y] / mol dm^{-3}	Initial rate / mol dm^{-3} s^{-1}
1	0.030	0.040	4.0×10^{-4}
2	0.045	0.040	6.0×10^{-4}
3	0.060	0.120	2.4×10^{-3}

Deduce the order of reaction with respect to X.

....

....

(2 marks)

(b) Deduce the order of the reaction with respect to Y.

....

....

(2 marks)

(c) Write the rate expression for the reaction between X and Y.

....

(1 mark)

(d) Determine the rate constant, *k*, correct to three significant figures and state its units, using data from Experiment 2.

Experiment	Initial [X] / mol dm^{-3}	Initial [Y] / mol dm^{-3}	Initial rate / mol dm^{-3} s^{-1}
1	0.030	0.040	4.0×10^{-4}
2	0.045	0.040	6.0×10^{-4}
3	0.060	0.120	2.4×10^{-3}

....

....

....

(3 marks)

9 (a) Explain why the reaction represented below is a redox reaction.

$$2ClO_2\ (aq) + 2NaOH\ (aq) \rightarrow NaClO_3\ (aq) + NaClO_2\ (aq) + H_2O\ (l)$$

....

....

(2 marks)

(b) For the reaction below, consider the following experimental data.

$$2ClO_2\ (aq) + 2OH^-\ (aq) \rightarrow ClO_3^-\ (aq) + ClO_2^-\ (aq) + H_2O\ (l)$$

Experiment	**Initial [ClO_2] / mol dm^{-3}**	**Initial [OH^-] / mol dm^{-3}**	**Initial rate / mol dm^{-3} s^{-1}**
1	0.85	1.70	9.30×10^{-5}
2	1.70	1.70	3.72×10^{-4}
3	1.70	0.85	1.86×10^{-4}

Deduce the rate expression.

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....

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(3 marks)

(c) Determine the rate constant, *k*, and state its units, using data from Experiment 3.

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(3 marks)

(d) Calculate the rate when [ClO_2 (aq)] = 3.10 x 10^{-2} mol dm^{-3} and [OH^- (aq)] = 1.50 x 10^{-2} mol dm^{-3}.

....

....

(2 marks)

10 (a) Sketch a graph to show how the rate constant, *k*, varies with temperature.

....

(1 mark)

(b) The following mechanism is proposed for the reaction where ethanal dimerises in dilute alkaline solution to form 3-hydroxybutanal:

Step 1: $CH_3CHO + :OH^- \rightarrow :CH_2CHO + H_2O$
Step 2: $CH_3CHO + :CH_2CHO \rightarrow CH_3CH(O:^-)CH_2CHO$
Step 3: $CH_3CH(O:^-)CH_2CHO + H_2O \rightarrow CH_3CH(OH)CH_2CHO + :OH^-$

Classify OH^-, CH_2CHO and $CH_3CH(O:^-)CH_2CHO$ as reactant, product, catalyst or intermediate, based on the proposed mechanism.

....

....

....

(3 marks)

(c) Using the following information about the proposed mechanism, deduce the rate expression.

Step 1: $CH_3CHO + :OH^- \rightarrow :CH_2CHO + H_2O$ slow step
Step 2: $CH_3CHO + :CH_2CHO \rightarrow CH_3CH(O:^-)CH_2CHO$ fast step
Step 3: $CH_3CH(O:^-)CH_2CHO + H_2O \rightarrow CH_3CH(OH)CH_2CHO + :OH^-$ fast step

....

(1 mark)

(d) Calculate the initial rate of reaction for experiment 2, if measured under the same conditions.

Experiment	Initial [CH_3CHO] / mol dm^{-3}	Initial [OH^-] / mol dm^{-3}	Initial rate / mol dm^{-3} s^{-1}
1	0.25	0.20	4.2×10^{-2}
2	0.25	0.30	

....

(1 mark)

(e) State the effect, if any, increasing the concentration of a reactant would have on the value of the rate constant, *k*.

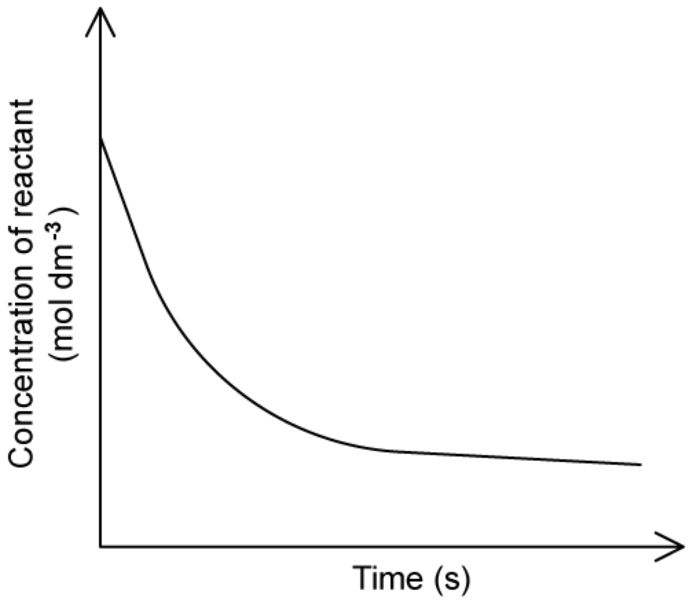
....

(1 mark)

11 (a) Nitrogen dioxide and carbon monoxide react according to the following equation.

$$NO_2\ (g) + CO\ (g) \rightarrow NO\ (g) + CO_2\ (g)$$

Using the following graph, what is the order with respect to NO_2?



....

(1 mark)

(b) The rate expression for the reaction of nitrogen dioxide and carbon monoxide at T < 227 ^{O}C is:

$$\text{Rate} = k\ [NO_2]^2$$

Sketch a labelled graph of concentration against time for carbon monoxide.

....

....

(2 marks)

(c) A student proposed the following single step mechanism for the reaction of nitrogen dioxide and carbon monoxide.

$$NO_2 + CO \rightarrow NO + CO_2 \qquad \text{slow}$$

$$\text{Rate} = k\,[NO_2]^2$$

Justify whether the student's proposed mechanism is correct.

....

....

(2 marks)

(d) Another student proposed the following mechanism for the reaction of nitrogen dioxide and carbon monoxide.

Step 1: $NO_2 + NO_2 \rightarrow NO + NO_3$

Step 2: $NO_3 + 2CO \rightarrow NO + 2CO_2$

$$\text{Rate} = k\,[NO_2]^2$$

Explain which of the student's proposed mechanism steps is the rate determining step.

....

(1 mark)

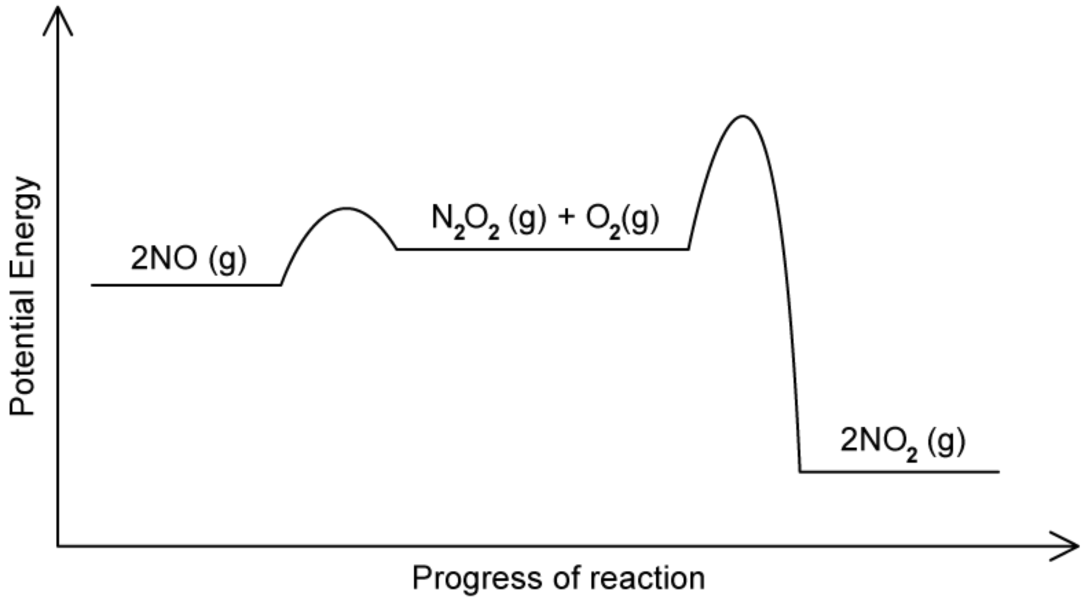
12 (a) Nitrogen(II) oxide is oxidised according to the following equation.

$$2NO\ (g) + O_2\ (g) \rightarrow 2NO_2\ (g)$$

The following mechanism is proposed for the two-step oxidation of nitrogen(II) oxide.

Step 1: $NO\ (g) + NO\ (g) \rightarrow N_2O_2\ (g)$
Step 2: $N_2O_2\ (g) + O_2\ (g) \rightarrow 2NO_2\ (g)$

The potential energy profile for this two-step reaction is shown.



Explain which step is the rate determining step.

....

(1 mark)

(b) Energy profile diagrams give evidence for or against a proposed mechanism or proposed rate expression.

i) Explain why the rate expression for the oxidation of nitrogen(II) oxide is not rate = $k\ [N_2O_2]\ [O_2]$.

[2]

ii) Deduce the rate expression for the oxidation of nitrogen(II) oxide.

[1]

....

(3 marks)

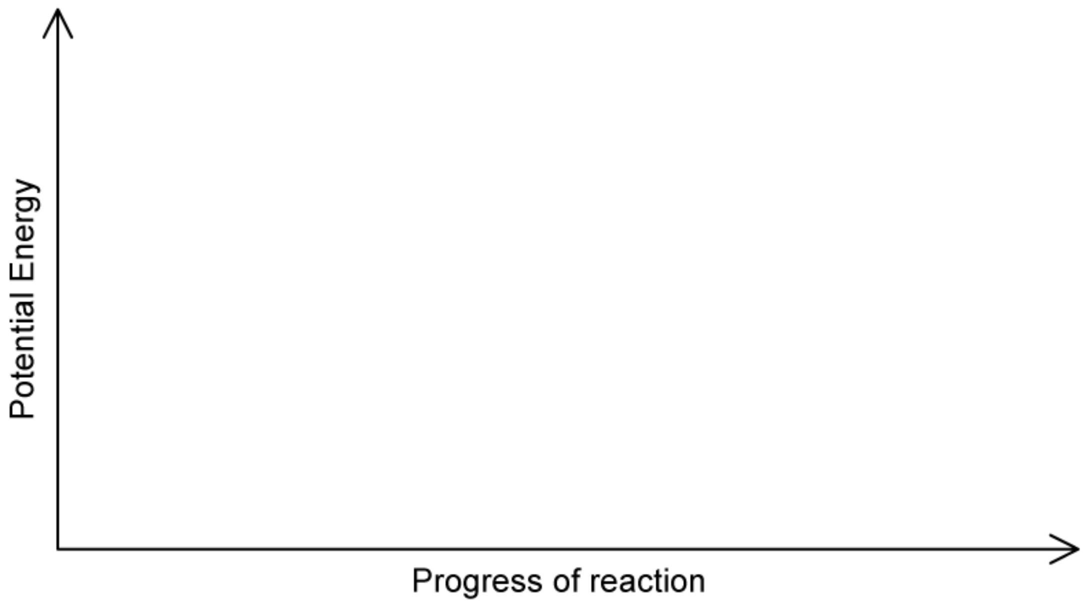
(c) Explain why the following reaction between iodide ions and peroxodisulfate ions has a high activation energy.

$$S_2O_8^{2-} (aq) + 2I^- (aq) \rightarrow 2SO_4^{2-} (aq) + I_2 (aq)$$

(2 marks)

(d) Sketch the potential energy diagram for the reaction of iodide ions with peroxodisulfate ions catalysed by iron(II) ions according to the following mechanism.

$2Fe^{2+} (aq) + S_2O_8^{2-} (aq) \rightarrow 2Fe^{3+} (aq) + 2SO_4^{2-} (aq)$	slow
$2Fe^{3+} (aq) + 2I^- (aq) \rightarrow 2Fe^{2+} (aq) + I_2 (aq)$	fast



(2 marks)

(e) Deduce the rate expression for the reaction of iodide ions with peroxodisulfate ions catalysed by iron(II) ions according to the following mechanism.

$$2Fe^{2+}(aq) + S_2O_8^{2-}(aq) \rightarrow 2Fe^{3+}(aq) + 2SO_4^{2-}(aq) \qquad \text{slow}$$
$$2Fe^{3+}(aq) + 2I^{-}(aq) \rightarrow 2Fe^{2+}(aq) + I_2(aq) \qquad \text{fast}$$

....

(1 mark)

13 (a) The decomposition of hydrogen peroxide into water and oxygen occurs at a slow rate with a rate constant of k = 6.62 x 10^{-3} mol dm^{-3} s^{-1} and at a temperature of 290 K.

Using Sections 1 and 2 of the Data Booklet, calculate the activation energy, E_a, correct to three significant figures and state its units.

The constant, A = 3.18 × 10^{11} mol^{-1} dm^3.

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(3 marks)

(b) Hydrogen peroxide decomposes to form water and oxygen as shown in the equation below.

$$2H_2O_2\ (aq) \rightarrow 2H_2O\ (l) + O_2\ (g)$$

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**	$\frac{1}{T}$
0.000493		295	
0.000656		298	
0.001400		305	
0.002360		310	
0.006120		320	

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

(2 marks)

(c) The results of the experiment can be used to calculate the activation energy, E_a. Use the results table to plot a graph of *ln k* against $\frac{1}{T}$.

(4 marks)

(d) Using Sections 1 and 2 of the Data Booklet and your graph, calculate a value for the activation energy, E_a, for this reaction. To gain full marks you must show all of your working.

(4 marks)

14 (a) The Arrhenius equation can be represented as $k = Ae^{-Ea/RT}$ in its exponential form.

State the effect on *k* of an increase in;

i) The constant, *A*, (frequency factor)

[1]

ii) Activation energy, E_a

[1]

iii) Temperature, *T*

[1]

....

....

....

(3 marks)

(b) Using Sections 1 and 2 of the Data Booklet, calculate the activation energy, E_a, of a reaction at 57 °C and a rate constant of 1.30 x 10^{-4} mol dm^{-3} s^{-1}.

The constant $A = 4.55 \times 10^{13}$.

....

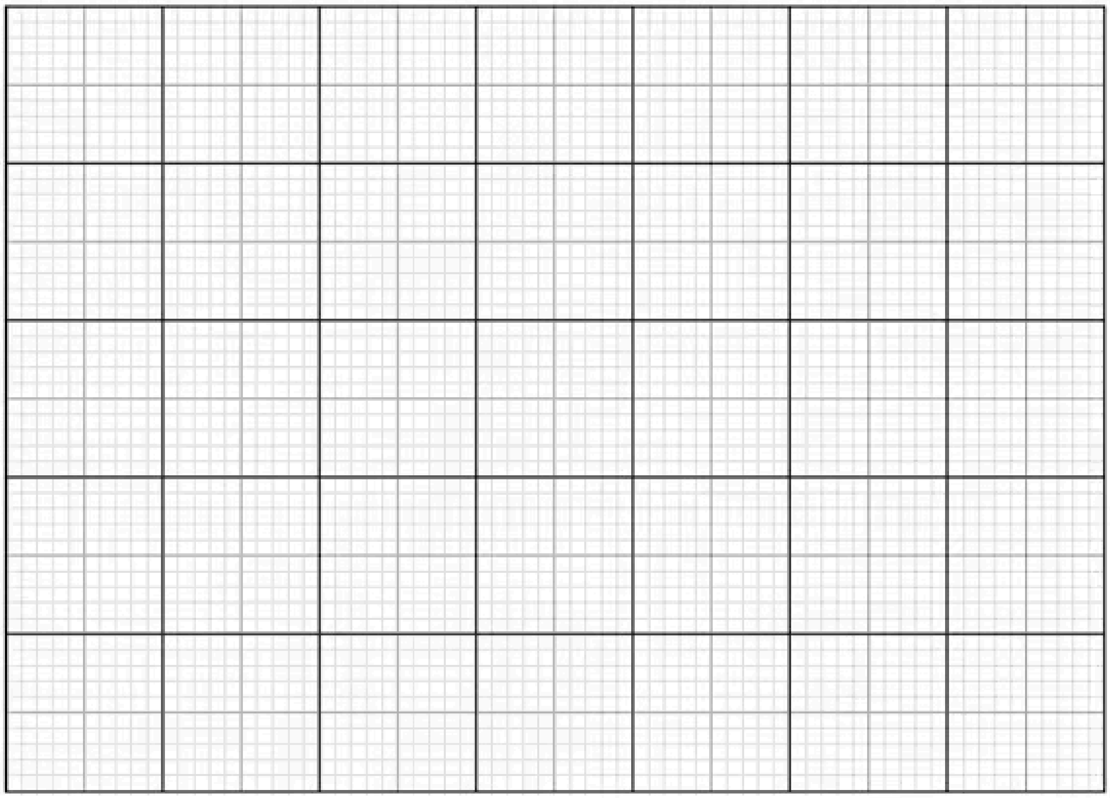
....

(2 marks)

(c) The table below shows how temperature affects the rate of reaction.

Rate constant k/s^{-1}	***ln k***	**Temperature / K**	$\frac{1}{T}$
2.0 x 10^{-5}	-10.8	278	0.00360
4.7 x 10^{-4}	-7.7	298	0.00336
1.7 x 10^{-3}	-6.4	308	0.00325
5.2 x 10^{-3}	-5.3	318	0.00314

Use the results to plot a labelled graph of *ln k* against $\frac{1}{T}$.



(3 marks)

(d) Using Sections 1 and 2 of the Data Booklet and your graph, calculate a value for the activation energy, E_a, for this reaction.

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(4 marks)

15 (a) Nitrogen dioxide and ozone react according to the following equation.

$$2NO_2\,(g) + O_3(g) \rightarrow N_2O_5\,(g) + O_2\,(g)$$

Experimental data shows the reaction is first order with respect to NO_2 and first order with respect to O_3.

State the rate expression for the reaction.

....

(1 mark)

(b) At 30°C, the initial rate of reaction is 3.46×10^{-3} mol dm^{-3} s^{-1} when the initial concentration of NO_2 is 0.50 mol dm^{-3} and the initial concentration of O_3 is 0.21 mol dm^{-3}.

Calculate a value for the rate constant *k* at this temperature and state its units.

....

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(3 marks)

(c) Using Sections 1 and 2 of the Data Booklet and your answer from part (b), calculate a value for the activation energy of this reaction at 30 °C.

For this reaction *ln A* = 15.8 mol^{-1} dm^{3}.

....

....

....

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(4 marks)

(d) The relationship between the rate constant and temperature is given by the Arrhenius equation, $k = Ae^{-\frac{Ea}{RT}}$

State how temperature affects activation energy.

....

(1 mark)

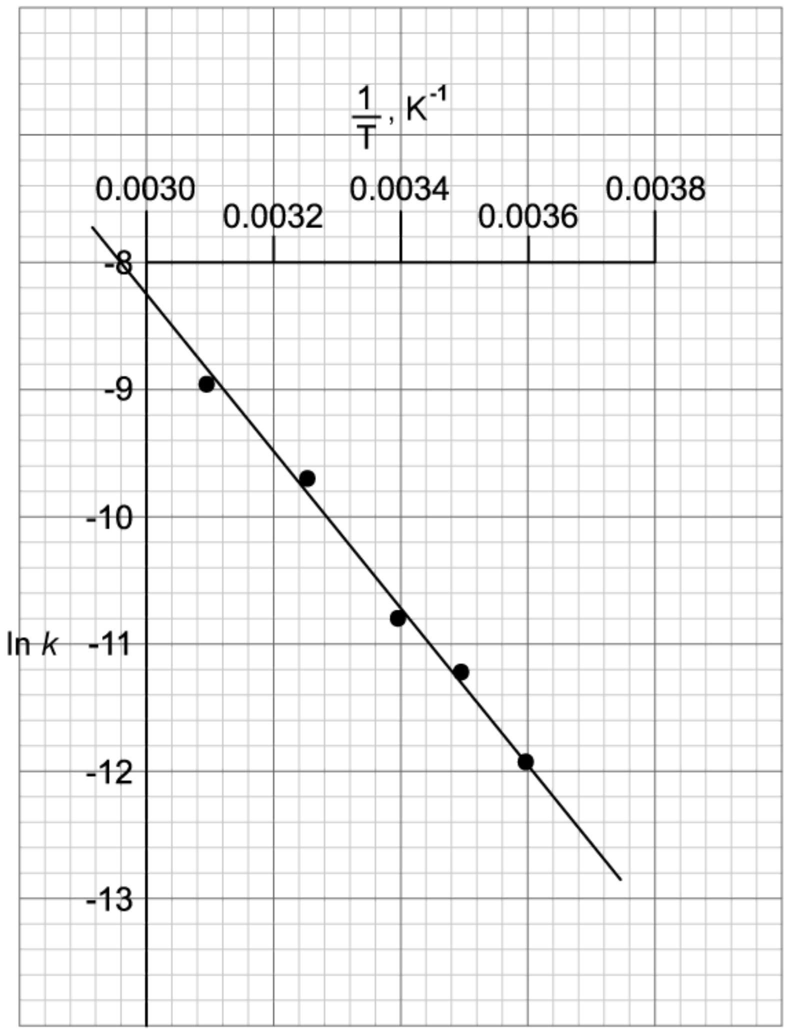
16 (a) A common relationship exists between temperature and rate.

What temperature change is associated with a doubling of rate?

....

(1 mark)

(b) An Arrhenius plot of ln k against $\frac{1}{T}$ for the reaction between A (g) and B (g) at different temperatures is shown in **Figure 1** below.



The equation of the line of best fit was found to be:

$$\ln k = -6154\left(\frac{1}{T}\right) - 8.2$$

Calculate the activation energy, E_a, for the reaction between A (g) and B (g).

(2 marks)

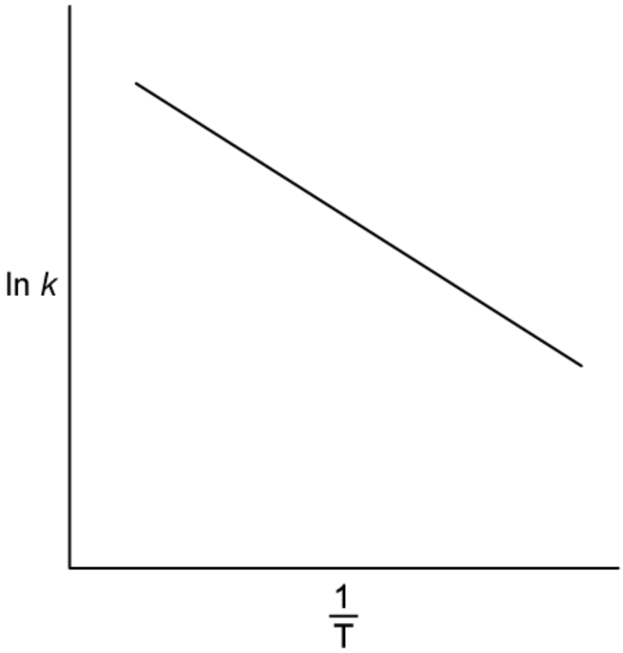
(c) Define the Arrhenius constant, *A*.

(2 marks)

(d) Using the Arrhenius plot, calculate an approximate value for the constant, *A*.

(2 marks)

17 (a) The graph of ln *k* against $\frac{1}{T}$ for a general reaction is shown.

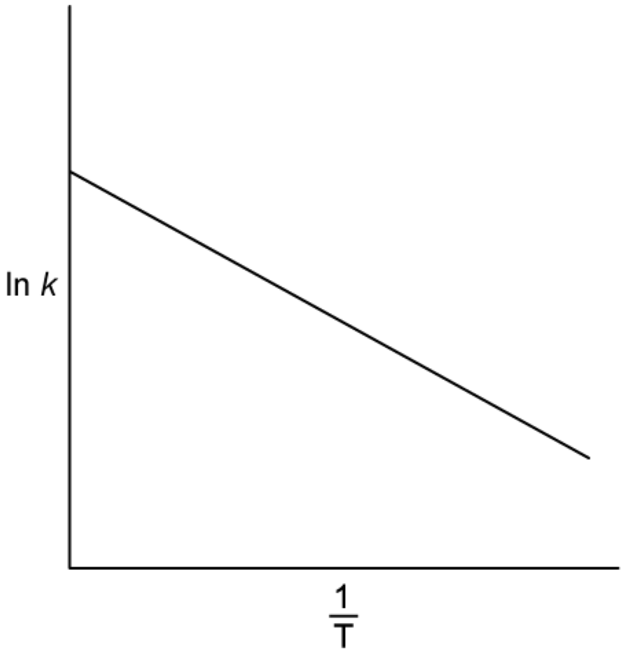


Sketch the expected line for a **different** reaction with a higher activation energy.

....

(1 mark)

(b) A graph of ln *k* against $\frac{1}{T}$ for another general reaction is shown.



Sketch the expected line for the **same** reaction with an added catalyst.

(2 marks)

(c) Rate constant data for the reaction of hydrogen and iodine at two different temperatures is shown in the table below.

$$H_2 (g) + I_2 (g) \rightarrow 2HI (g)$$

Table 1

Experiment	**Temperature / K**	**Rate constant, *k* / mol dm^{-3} s^{-1}**
1	599	5.40×10^{-4}
2	683	2.80×10^{-2}

Using Sections 1 and 2 of the Data Booklet, calculate the activation energy, in kJ mol^{-1}, for the reaction.

(3 marks)

(d) Using the data from experiment 1 and Sections 1 and 2 in the Data Booklet, calculate a value for the constant, *A*.

Table 2

Experiment	**Temperature / K**	**Rate constant, *k* / mol dm^{-3} s^{-1}**
1	599	5.40×10^{-4}
2	683	2.80×10^{-2}

(2 marks)

Hard Questions

1 (a) A group of students planned how to investigate the effect of changing the concentration of H_2SO_4 on the initial rate of reaction with magnesium:

$$Mg\ (s) + H_2SO_4\ (aq) \rightarrow MgSO_4\ (aq) + H_2\ (g)$$

They decided to measure how long the reaction took to complete when similar masses of magnesium were added to acid.

Two methods were suggested:

Method 1 - Use small pieces of magnesium ribbon, an excess of acid and record the time taken for the magnesium ribbon to disappear

Method 2 - Use large strips of magnesium ribbon, an excess of magnesium and record the time taken for bubbles to stop forming

Deduce, giving a reason, which of method 1 and method 2 would be the least affected if the masses of magnesium ribbon used varied slightly between each experiment.

....

....

(2 marks)

(b) Neither method in part a) actually allows the initial rate to be calculated. Outline a method that would allow the calculation of initial rate.

....

....

(2 marks)

(c) The reaction is to be conducted across a few weeks.

State a factor that has a significant effect on reaction rate, which could vary between experiments across the weeks and therefore needs to be controlled.

(1 mark)

(d) One group collected the following data using 1.50 mol dm^{-3} acid:

Trial	Time/ s (± 0.01 s)
1	91.56
2	98.33
3	72.08
4	89.41

i) Comment on the use of uncertainty when calculating the mean.

[2]

ii) Calculate the mean time for the set of results.

[2]

(4 marks)

2 (a) When investigating the reaction between sulfuric acid and calcium carbonate, it was observed that a small increase of temperature of around 10 ^{o}C caused a doubling in the rate of the reaction.

Sketch and label Maxwell-Boltzmann curve for the two temperatures T and T+10, and use this diagram to help to explain this effect of temperature on rate.

(5 marks)

Why do some collisions at high temperatures still not result in the formation of the product?

(b) **(2 marks)**

(c) Identify and explain another factor that affects the number of particles present in a solution with sufficient energy to react.

(3 marks)

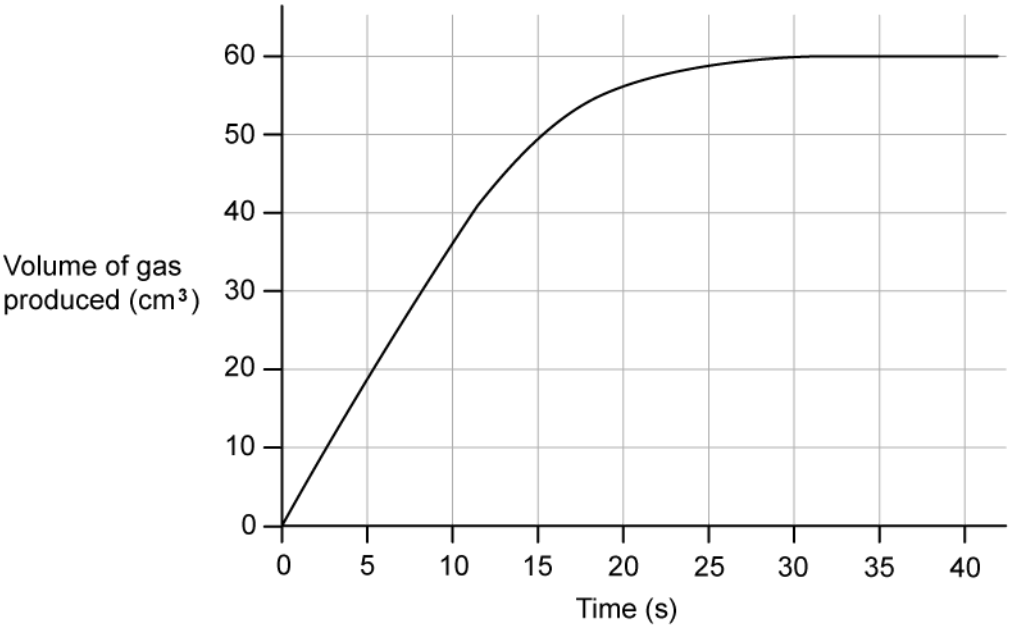
(d) Some groups investigating the effect of temperature on rate stirred their reactions, some did not.

Explain the effect of stirring upon the rate of the reaction.

(2 marks)

3 (a) 0.5 g of magnesium reacts with 50 cm^3 of 0.01 $moldm^{-3}$ nitric acid. Magnesium is in excess.

A graph monitoring the volume of hydrogen gas produced is shown below:



i) Calculate the mean rate of reaction over the first 15 seconds of the reaction

[1]

ii) Calculate the actual rate of reaction at 15 seconds

[3]

iii) Explain the difference in values for rate

[1]

....

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....

....

(5 marks)

(b) Compare the expected rate and progress of the reaction if 25 cm^3 of 0.2 mol dm^{-3} nitric acid was used instead of 50 cm^3 of 0.1 mol dm^{-3} nitric acid.

(3 marks)

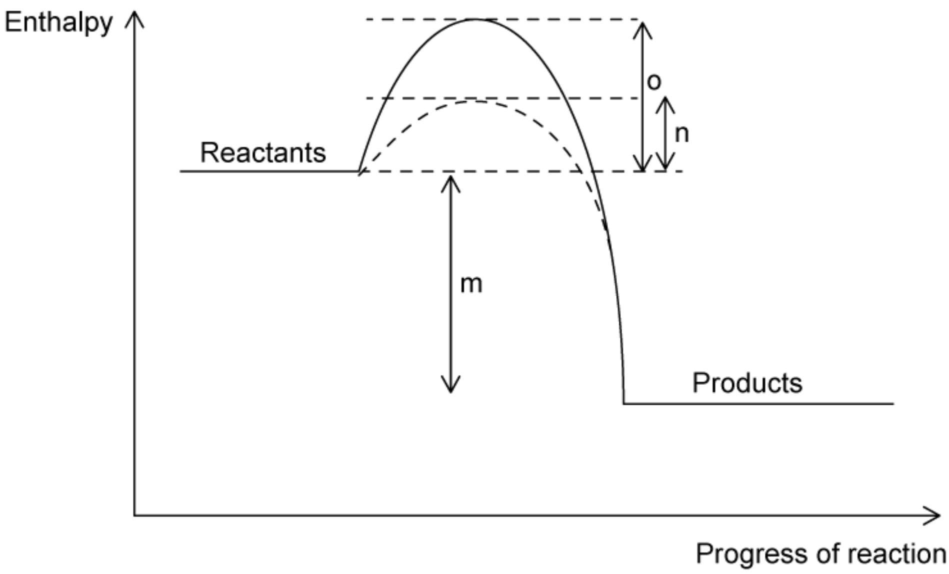
(c) Suggest one change to the reaction that could be made to produce more hydrogen gas in total and explain your choice.

(2 marks)

(d) Suggest why it is often better to study a slower reaction instead of a faster one.

(2 marks)

4 (a) The following energy profile diagram shows the pathways for both a catalysed and uncatalysed reversible reaction:



Identify the letter(s) representing the activation energy for the catalysed reverse reaction.

....

(1 mark)

(b) State and explain the effect that this catalyst will have on the equilibrium yield.

....

....

(2 marks)

(c) Vehicles with combustion engines usually have catalytic convertors added to catalyse the oxidation of carbon monoxide into carbon dioxide and to catalyse the reduction of nitrogen oxides to nitrogen. These catalysts are usually rhodium or platinum.

Leaded fuels were phased out as they were found to poison these catalysts, binding irreversibly to the metal surface.

Explain the problems for drivers of the catalysts being poisoned.

....

(2 marks)

(d) Suggest a situation in which using a catalyst would not be appropriate.

(1 mark)

5 (a) The conversion of hydrogen and iodine into hydrogen iodide proceeds via a three step reaction mechanism:

1. $I_2\ (g) \rightleftharpoons 2I\ (g)$ fast
2. $H_2\ (g) + I\ (g) \rightleftharpoons H_2I\ (g)$ fast
3. $H_2I\ (g) + I\ (g) \rightarrow 2HI\ (g)$ slow

Write the rate equation for this reaction and show how the mechanism is consistent with the stoichiometric equation.

....

....

(2 marks)

(b) An investigation into the rate of reaction between hydrogen and iodine was carried out at 298 K and the data obtained is shown below.

Experiment	$[H_2]$ / mol dm^{-3}	$[I_2]$ / mol dm^{-3}	Initial rate/ mol dm^{-3} s^{-1}
1	0.0258	0.0137	6.43×10^{-22}
2	0.0258	0.0274	1.29×10^{-21}
3	0.0516	0.0137	1.29×10^{-21}

Determine the rate equation for the reaction and justify your answer.

....

....

....

(3 marks)

(c) Calculate the rate constant using Expt 2 data, including its units.

....

(1 mark)

(d) Using section 12 of the Data booklet, determine whether the forward reaction is favoured by an increase in temperature.

....

(1 mark)

6 (a) The reaction between iodide ions and persulfate ions is a 'clock' reaction and often used to study reaction kinetics.

$$2I^- (aq) + S_2O_8^{2-} (aq) \rightarrow I_2 (aq) + 2SO_4^{2-} (aq)$$

Deduce the redox changes taking place in the reaction.

(2 marks)

(b) A persulfate-iodide clock reaction was studied and the following rate data obtained.

Experiment	$[S_2O_8^{2-}]$ / mol dm^{-3}	$[I^-]$ / mol dm^{-3}	Initial rate/ mol dm^{-3} s^{-1}
1	0.25	0.10	8.0×10^{-3}
2	0.10	0.10	3.2×10^{-3}
3	0.20	0.30	1.92×10^{-2}

Deduce the order with respect to persulfate ions and iodide ions.

(2 marks)

(c) Determine the rate equation for the reaction and calculate rate constant, including the units.

(2 marks)

(d) Four mechanisms are proposed for the persulfate-iodide reaction. Deduce which mechanism(s) is/are consistent with the rate equation in part c) and justify your answer.

Mechanism 1:

1. I^- (aq) + I^- (aq) → I_2^{2-} (aq) *slow*
2. I_2^{2-} (aq) + $S_2O_8^{2-}$ (aq) → I_2 (aq) + $2SO_4^{2-}$ (aq) *fast*

Mechanism 2:

1. I^- (aq) + $S_2O_8^{2-}$ (aq) → $S_2O_8I^{3-}$ (aq) *slow*
2. $S_2O_8I^{3-}$ (aq) + I^- (aq) → I_2 (aq) + $2SO_4^{2-}$ (aq) *fast*

Mechanism 3:

1. I^- (aq) + $S_2O_8^{2-}$ (aq) → $S_2O_8I^{3-}$ (aq) *fast*
2. $S_2O_8I^{3-}$ (aq) + I^- (aq) → I_2 (aq) + $2SO_4^{2-}$ (aq) *slow*

Mechanism 4:

1. $2I^-$ (aq) + $S_2O_8^{2-}$ (aq) → I_2 (aq) + $2SO_4^{2-}$ (aq) *slow*

....

....

....

(3 marks)

7 (a) The reaction between nitrogen monoxide and hydrogen produces nitrogen and water:

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

Rate data for this reaction is shown below.

Experiment	[NO] / mol dm^{-3}	$[H_2]$ / mol dm^{-3}	Initial rate/ mol dm^{-3} s^{-1}
1	0.001	0.004	0.002
2	0.002	0.004	0.008
3	0.004	0.001	0.016

What is the *molecularity* of the reaction?

....

(1 mark)

(b) Draw a sketch graphs of:

i) Rate against concentration of NO.

[1]

ii) Rate against concentration of H_2

[1]

....

....

(2 marks)

(c) Suggest a possible mechanism for the reaction.

....

....

(2 marks)

(d) Suggest a Lewis structure for N_2O_2 and draw the shape of the molecule.

....

....

(2 marks)

8 (a) The rate of reaction between manganate(VII) ions and oxalate ions, $C_2O_4^{2-}$, can be investigated by measuring how the concentration of manganate(VII) varies with time.

$$2MnO_4^-(aq) + 16H^+(aq) + 5C_2O_4^{2-} \rightarrow 2Mn^{2+}(aq) + 8H_2O(l) + 10CO_2(g)$$

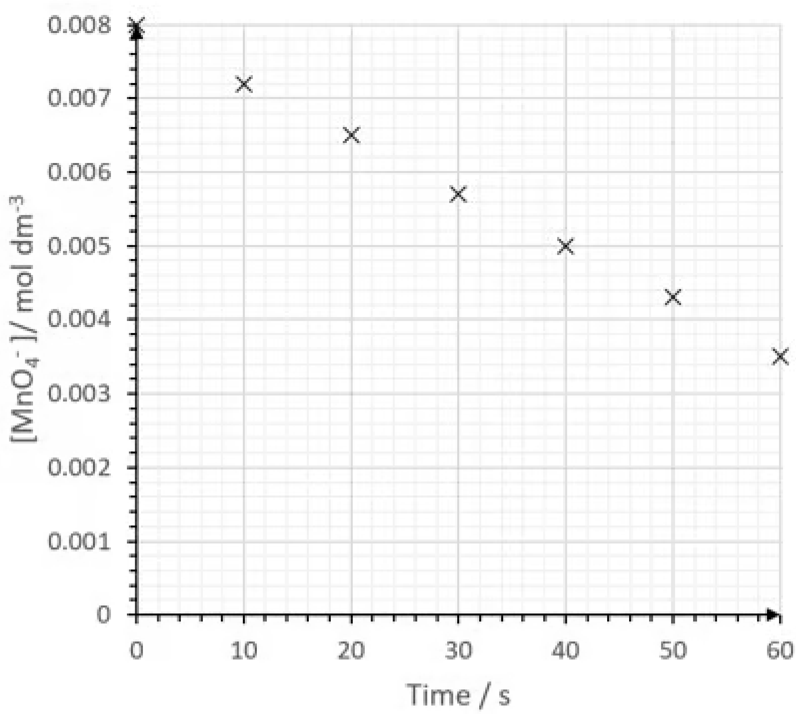
The rate is first order with respect to oxalate ions and the general rate equation for the reaction is:

$$\text{rate} = k\,[MnO_4^-]^p[C_2O_4^{2-}]^q[H^+]^r$$

i) Suggest how the change in manganate(VII) concentration can be measured.

[1]

ii) A student investigated how the concentration of manganate(VII) affected the rate of reaction and produced the following results. The oxalate ions and acid were in excess.



Determine the rate of reaction.

[2]

(3 marks)

(b) The student used an acid concentration of 1.0 mol dm^{-3}. She then varied it, keeping the other concentrations constant. She measured the rate of reaction and found the following results:

$[H^+]$/ mol dm^{-3}	Relative rate of reaction
0.5	0.0025
0.25	0.0013
0.01	0.0005

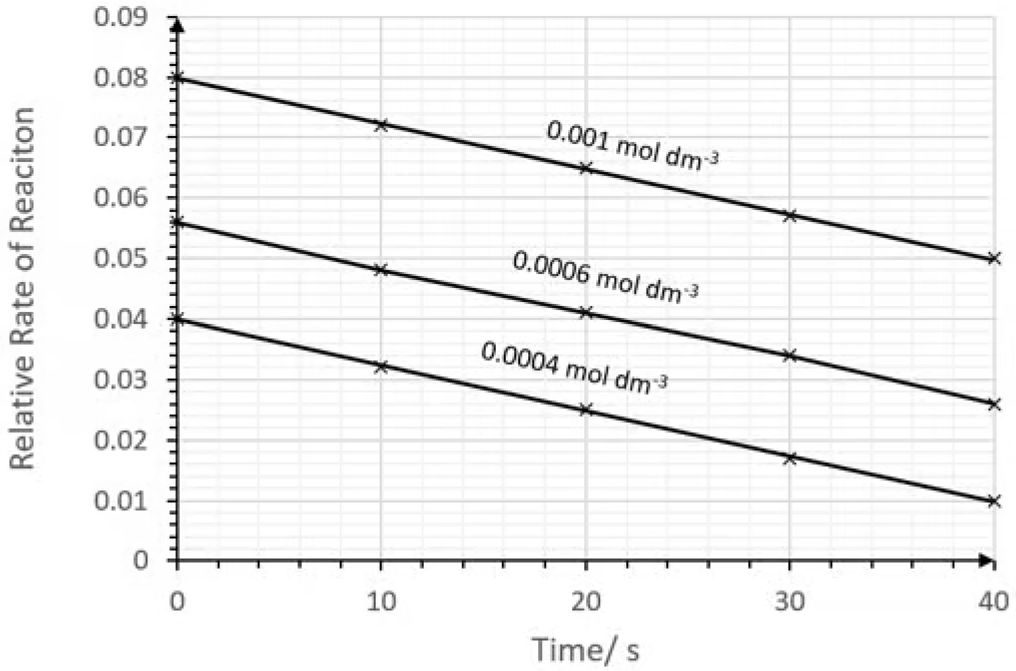
Identify the relationship between the relative rate of reaction and H^+, and hence determine the order of reaction with respect to H^+ ions.

....

....

(2 marks)

(c) The student varied the concentration of $[MnO_4^-]$ and plotted the rate against time at three different concentrations:



i) Deduce, with a reason, the order of reaction with respect to MnO_4^-.

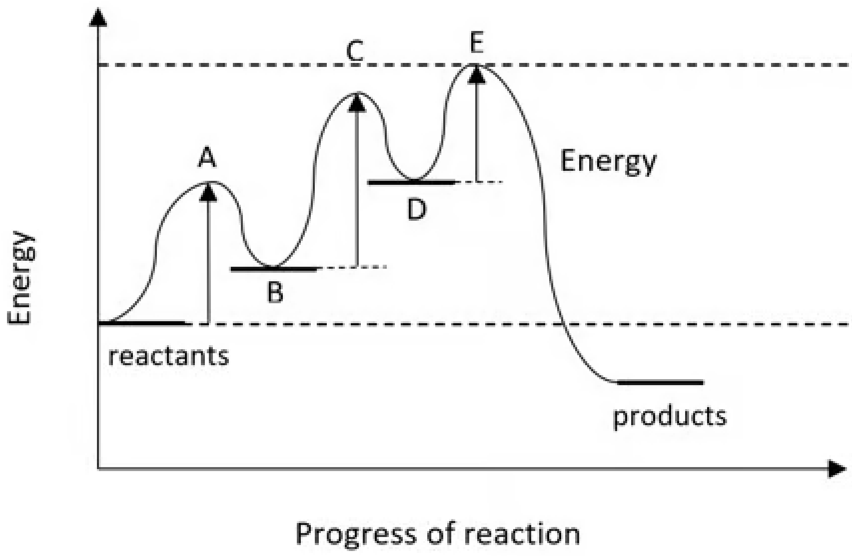
[2]

ii) Write the rate expression for the reaction.

[1]

(3 marks)

9 (a) A reaction proceeds by a three step mechanism. The energy profile for the reaction is shown below:



Explain the difference between points A, C, E and B, D on the profile.

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....

(4 marks)

(b) Deduce which step is the rate determining step of the reaction, giving a reason.

....

....

(2 marks)

10 (a) A series of experiments were carried out to investigate how the rate of the reaction of bromate and bromide in acidic conditions varies with temperature.

The time taken, *t*, was measured for a fixed amount of bromine to form at different temperatures. The results are shown below.

Temperature (T) / K	$\frac{1}{T}$ **x 10^{-3} / K^{-1}**	**Time (t) / s**	$\frac{1}{t}$ **/ s^{-1}**	**ln** $\frac{1}{t}$
408	2.451	21.14	0.0473	-3.051
428	2.336	10.57		
448		5.54	0.1805	-1.712
468	2.137	3.02	0.3311	-1.106
488	2.049			-0.536

Complete the table above.

....

....

....

(3 marks)

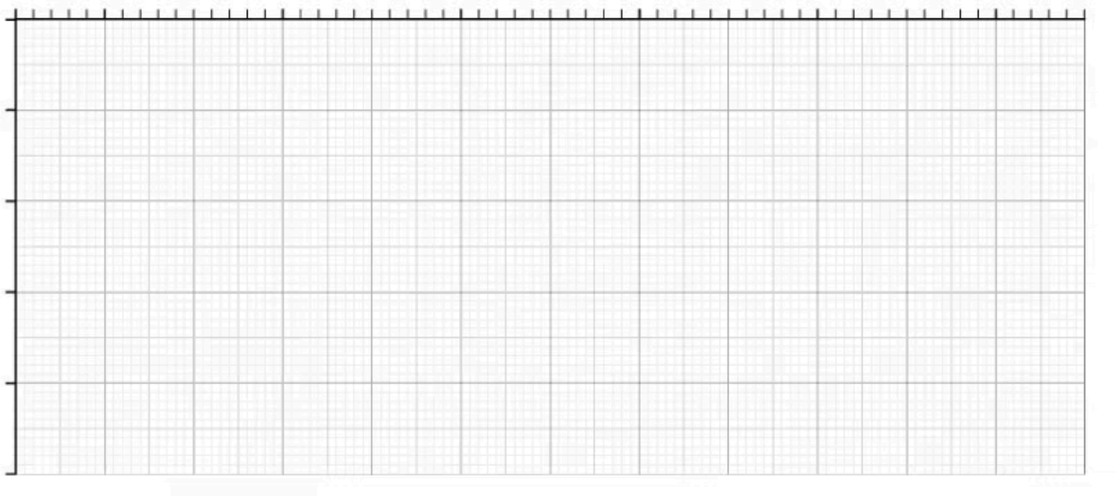
(b) The Arrhenius equation relates the rate constant, *k*, to the activation energy, E_a, and temperature, *T*.

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

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

$$\ln \frac{1}{t} = \ln A + \frac{-E_a}{RT}$$

Use your answers from part (a) to plot a graph of $\ln \frac{1}{t}$ against $\frac{1}{T} \times 10^{-3}$ on the graph below.



(4 marks)

(c) Use section 2 of the data booklet along with your graph and information from part (b) to calculate a value for the activation energy, in kJ mol^{-1}, for this reaction.

To gain full marks you must show all of your working.

(4 marks)

11 (a) Three experiments were carried out at a temperature, T_1, to investigate the rate of the reaction between compounds **F** and **G**. The results are shown in the table below:

	Experiment 1	**Experiment 2**	**Experiment 3**
Initial concentration of **F** / mol dm^{-3}	1.71×10^{-2}	5.34×10^{-2}	7.62×10^{-2}
Initial concentration of **G** / mol dm^{-3}	3.95×10^{-2}	6.24×10^{-2}	3.95×10^{-2}
Initial rate / mol dm^{-3} s^{-1}	3.76×10^{-3}	1.85×10^{-2}	1.68×10^{-2}

Use the data in the table to deduce the rate equation for the reaction between compounds **F** and **G**.

(3 marks)

(b) Use the information in the table in part (a) to calculate a value for the rate constant, *k*, for this reaction between 0.0534 mol dm^{-3} **F** and 0.0624 mol dm^{-3} **G**.

Give your answer to the appropriate number of significant figures.

State the units for *k*.

(If you did not get an answer for (a), you may assume that *rate* = *k* [**F**]^2 [**G**]^2. This is **not** the correct answer)

(2 marks)

(c) The Arrhenius equation shows how the rate constant, *k*, for a reaction varies with temperature, *T*.

$$k = Ae^{\frac{-E_a}{RT}}$$

For the reaction between 0.0534 mol dm^{-3} **F** and 0.0624 mol dm^{-3} **G** at 25 °C, the activation energy, E_a, is 16.7 kJ mol^{-1}.

Use section 2 of the data booklet and your answer to part (b) to calculate a value for the Arrhenius constant, *A*, for this reaction.

Give your answer to the appropriate number of significant figures.

(If you did not get an answer for (b), you may assume that *k* has a value of 4.97. This is **not** the correct answer)

....

....

(2 marks)

(d) The temperature of the reaction is increased to twice the original temperature, T_1. The value of *k* increases to 0.28 mol^{-1} dm^3 s^{-1} at this new temperature.

Using sections 1 and 2 of the data booklet and your answer to part (b), determine the original temperature, T_1.

(If you did not get an answer for (b), you may assume that k = 16700 mol^{-1} dm^3 s^{-1} This is **not the correct answer)**

....

....

(2 marks)

12 (a) The rate constant for a reaction doubles when the temperature is increased from 25.0 °C to 35 °C.

Calculate the activation energy, E_a, in kJ mol^{-1} for the reaction using section 1 and 2 of the data booklet.

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(2 marks)

(b) The rate constant is 6.2 x 10^3 s^{-1} when the temperature is reduced by a factor of a fifth from the original starting temperature, 25 °C.

Calculate the rate constant, in min^{-1}, using sections 1 and 2 of the data booklet.

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(2 marks)

(c) A different reaction route is used which reduces the activation energy of the reaction.

Explain how the rate constant calculated in part(b) would differ.

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(2 marks)