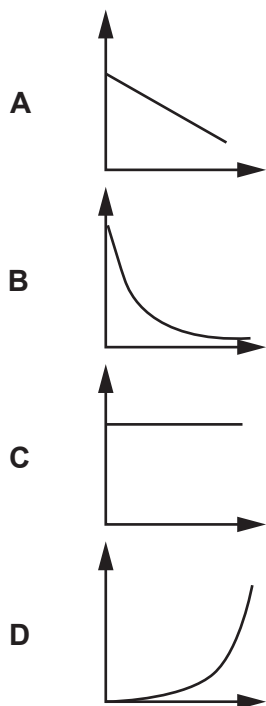


- 1 A reaction is zero order with respect to a reactant **A**.

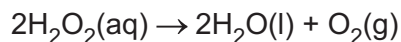
Which concentration–time graph for reactant **A** is the correct shape?



Your answer

[1]

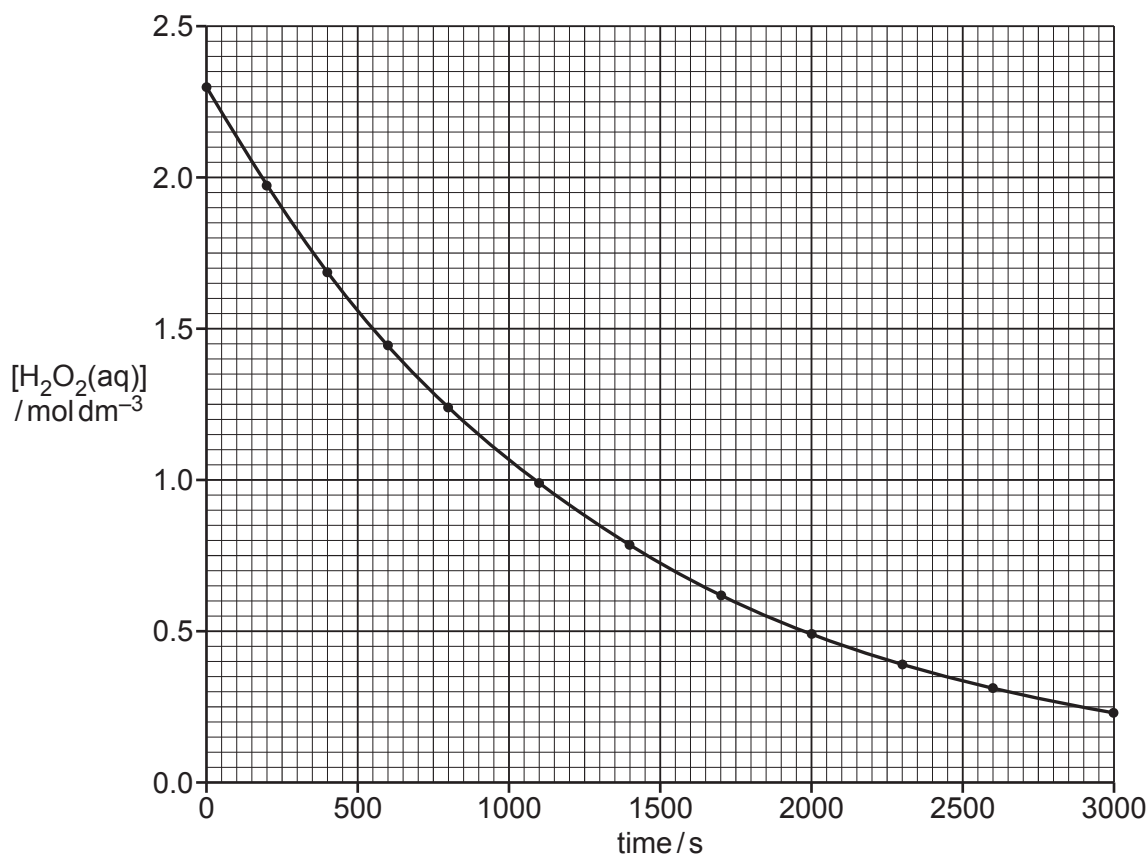
- 2 Aqueous solutions of hydrogen peroxide, $\text{H}_2\text{O}_2(\text{aq})$, decompose as in the equation below.



A student investigates the decomposition of $\text{H}_2\text{O}_2(\text{aq})$ by measuring the volume of oxygen gas produced over time. All gas volumes are measured at room temperature and pressure.

The student uses 25.0 cm^3 of 2.30 mol dm^{-3} H_2O_2 .

From the results, the student determines the concentration of $\text{H}_2\text{O}_2(\text{aq})$ at each time. The student then plots a concentration–time graph.



- (a) Determine the total volume of oxygen, measured at room temperature and pressure, that the student should be prepared to collect in this investigation.

Suggest apparatus that would allow this gas volume to be collected, indicating clearly the scale of working.

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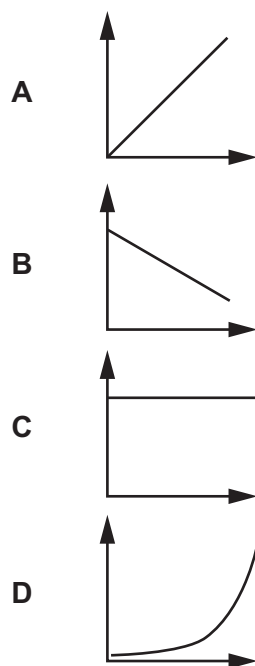
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Source: OCR H432/01 June 2017, original question 19(b)-(c)*

- 3 A reaction is first order with respect to a reactant **X**.

Which rate–concentration graph for reactant **X** is the correct shape?



Your answer

[1]

4 Which statement(s) explain(s) why reaction rates increase as temperature increases?

- 1 The activation energy is less.
- 2 Collisions between molecules are more frequent.
- 3 A greater proportion of molecules have energy greater than the activation energy.

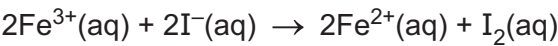
- A** 1, 2 and 3
- B** Only 1 and 2
- C** Only 2 and 3
- D** Only 1

Your answer

[1]

5 This question is about reaction rates.

Aqueous iron(III) ions, $\text{Fe}^{3+}(\text{aq})$, react with aqueous iodide ions, $\text{I}^{-}(\text{aq})$, as shown below.



A student carries out three experiments to investigate how different concentrations of $\text{Fe}^{3+}(\text{aq})$ and $\text{I}^{-}(\text{aq})$ affect the initial rate of this reaction. The results are shown below.

Experiment	$[\text{Fe}^{3+}(\text{aq})]$ / mol dm^{-3}	$[\text{I}^{-}(\text{aq})]$ / mol dm^{-3}	Initial rate / $\text{mol dm}^{-3} \text{s}^{-1}$
1	4.00×10^{-2}	3.00×10^{-2}	8.10×10^{-4}
2	8.00×10^{-2}	3.00×10^{-2}	1.62×10^{-3}
3	4.00×10^{-2}	6.00×10^{-2}	3.24×10^{-3}

(a)* Determine the rate constant and a possible two-step mechanism for this reaction that are consistent with these results. [6]

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Additional answer space if required

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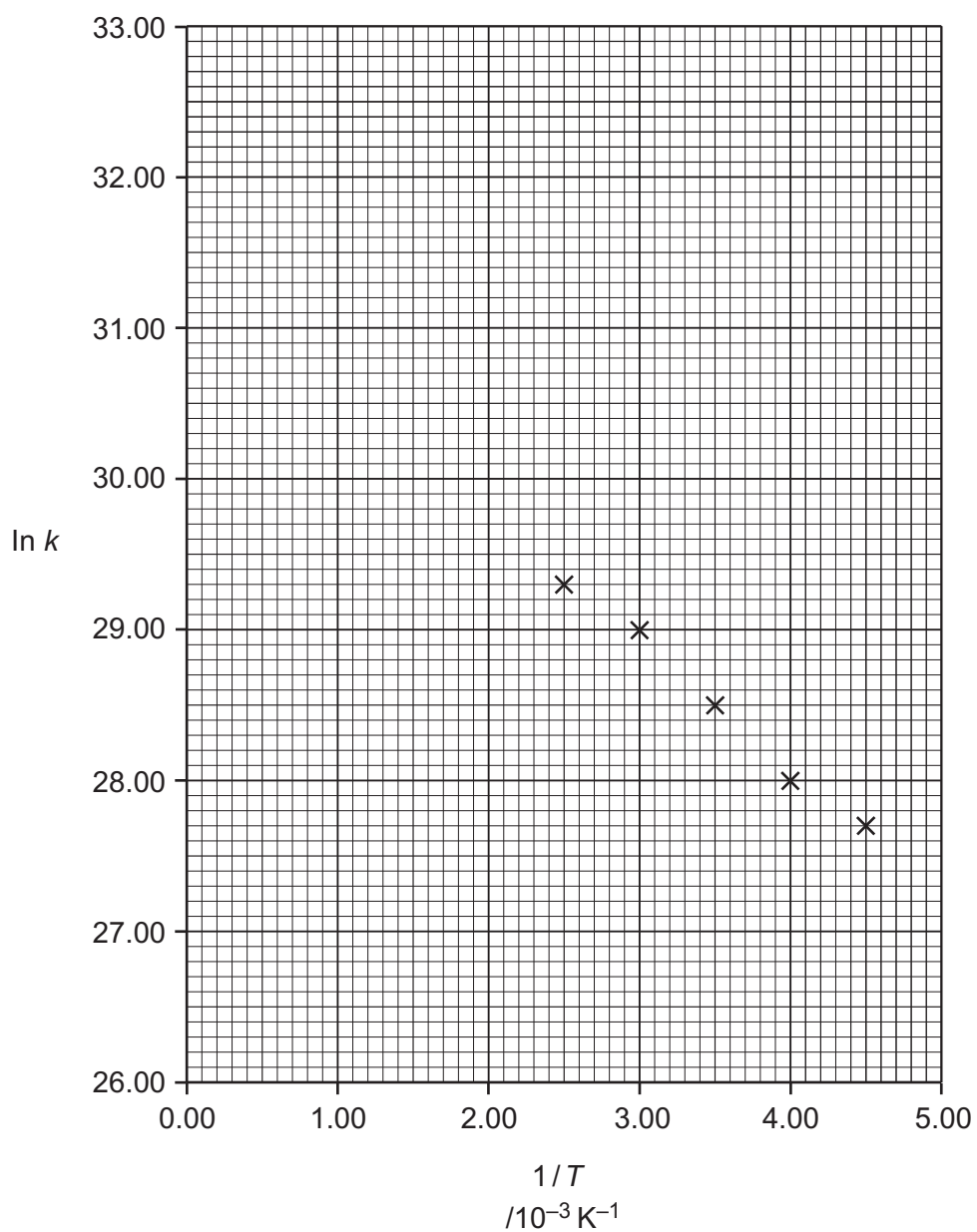
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- (b) A student carries out an investigation to find the activation energy, E_a , and the pre-exponential factor, A , of a reaction.

The student determines the rate constant, k , at different temperatures, T .
The student then plots a graph of $\ln k$ against $1/T$ as shown below.



- (i) Draw a best-fit straight line and calculate the activation energy, in J mol^{-1} .
Give your answer to **three** significant figures.

Show your working.

activation energy, $E_a = + \dots\dots\dots \text{J mol}^{-1}$ [3]

- (ii) Use the graph to calculate the value of the pre-exponential factor, A .

Show your working.

pre-exponential factor, $A = \dots\dots\dots$ [2]

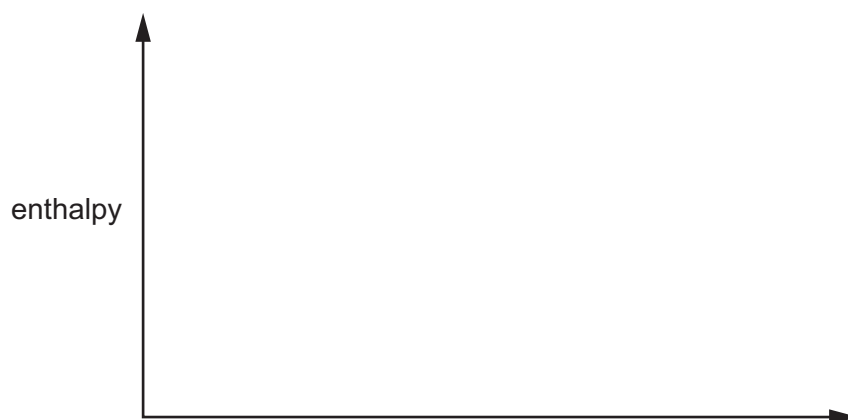
- 6(a) Hydrogen iodide, HI, is decomposed by heat into its elements:



The decomposition is much faster in the presence of a platinum catalyst.

Complete the enthalpy profile diagram for this reaction using formulae for the reactants and products.

- Use E_a to label the activation energy **without** a catalyst.
- Use E_c to label the activation energy **with** a catalyst.
- Use ΔH to label the enthalpy change of reaction.



- 7 A graph of $\ln k$ against $\frac{1}{T}$ (T in K) for a reaction has a gradient with the numerical value of -4420 .

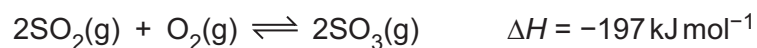
What is the activation energy, in kJ mol^{-1} , for this reaction?

- A -532
- B -36.7
- C $+36.7$
- D $+5.32 \times 10^5$

Your answer

[1]

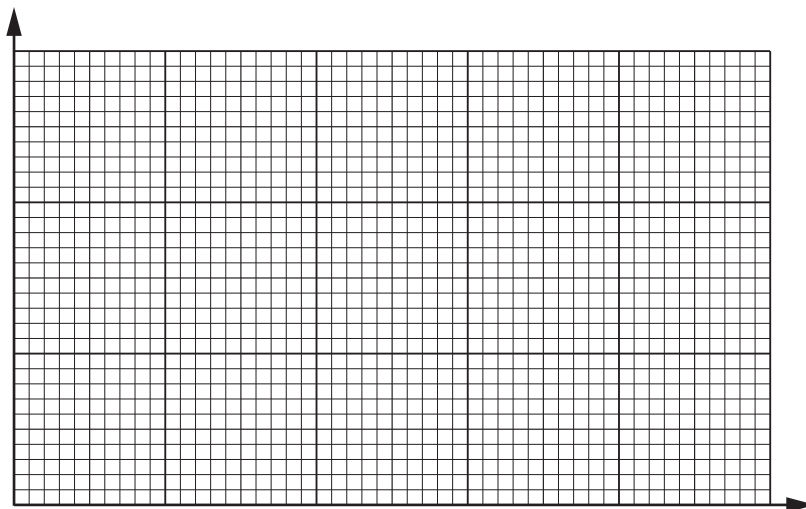
8(a) Vanadium(V) oxide, $V_2O_5(s)$, is used as a catalyst in



(i) Explain how the presence of $V_2O_5(s)$ increases the rate of reaction.

Include a labelled sketch of the Boltzmann distribution, on the grid below.

Label the axes.



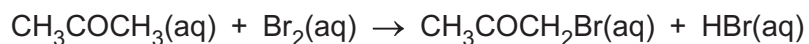
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 [4]

(ii) Explain whether vanadium(V) oxide is acting as a homogeneous or heterogeneous catalyst.

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 [1]

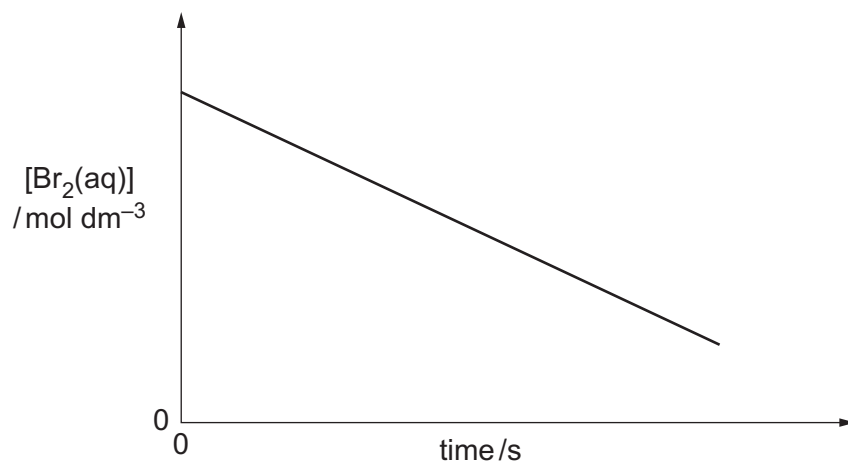
- 9 Three students carry out a rates investigation on the reaction between bromine and propanone in the presence of hydrochloric acid.



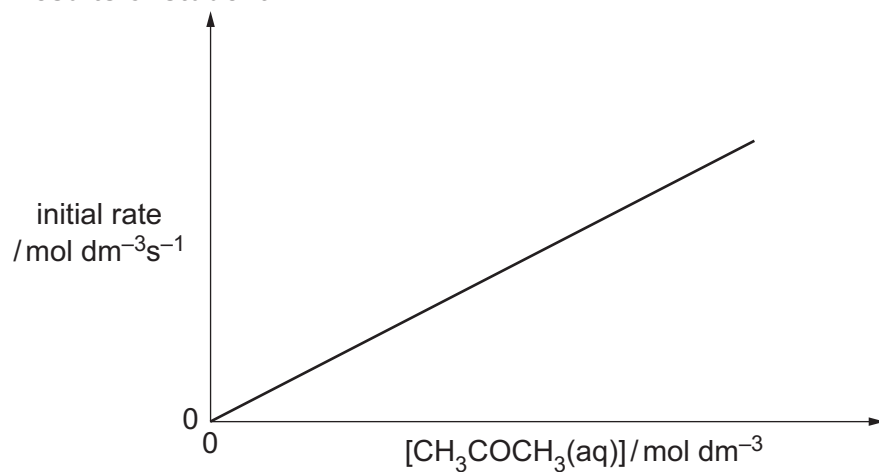
Each student investigates the effect of changing the concentration of one of the reactants whilst keeping the other concentrations constant.

Their results are shown below.

Results of student 1



Results of student 2



Results of student 3

Experiment	$[\text{Br}_2(\text{aq})]$ $/\text{mol dm}^{-3}$	$[\text{CH}_3\text{COCH}_3(\text{aq})]$ $/\text{mol dm}^{-3}$	$[\text{H}^+(\text{aq})]$ $/\text{mol dm}^{-3}$	Initial rate $/10^{-5}\text{mol dm}^{-3}\text{s}^{-1}$
1	0.004	1.60	0.20	1.25
2	0.004	1.60	0.40	2.50

Explain how the reaction orders can be determined from the students' results, and determine the rate equation and rate constant. **[6]**

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Additional answer space if required

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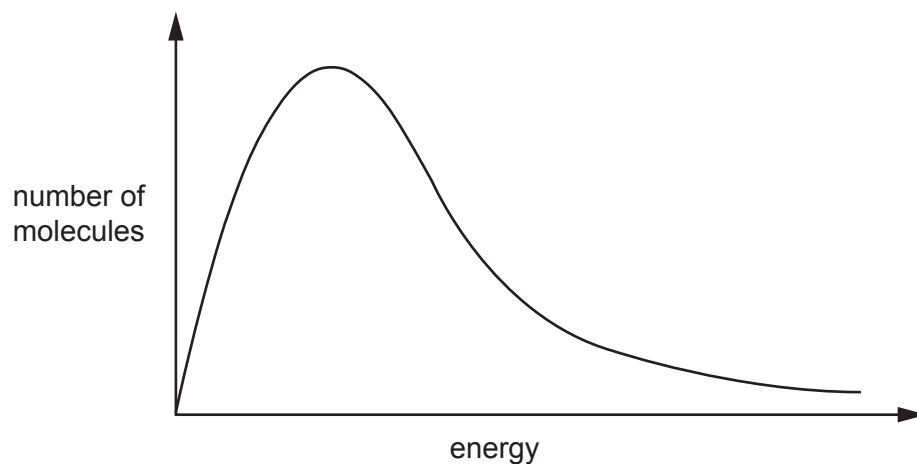
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- 10 The diagram represents a Boltzmann distribution curve of molecules at a given temperature.



Which statement for this Boltzmann distribution curve is correct at a higher temperature?

- A** The peak increases in height and moves to the left.
- B** The peak increases in height and moves to the right.
- C** The peak decreases in height and moves to the left.
- D** The peak decreases in height and moves to the right.

Your answer

[1]

- 11** A graph is plotted of $\ln(k)$ against $1/T$.
(k = rate constant, T = temperature in K)

The gradient has the numerical value of $-55\,000$.

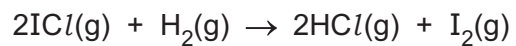
What is the activation energy, in kJ mol^{-1} ?

- A** $+1.5 \times 10^{-7}$
B $+2.22 \times 10^{-6}$
C $+6.62$
D $+457$

Your answer

[1]

- 12 The equation for the reaction of ICl and H_2 is shown below.



The rate constant k for this reaction is $1.63 \times 10^{-6} \text{ dm}^3 \text{ mol}^{-1} \text{ s}^{-1}$.

What is the overall order of the reaction?

A 0

B 1

C 2

D 3

Your answer

[1]

13(a) A catalyst is used in the production of methanol

State **two** ways that the use of catalysts helps chemical companies to make their processes more sustainable and less harmful to the environment.

1

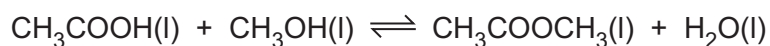
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2

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[2]

- 14 A student investigates the reaction between ethanoic acid, $\text{CH}_3\text{COOH(l)}$ and methanol, $\text{CH}_3\text{OH(l)}$, in the presence of an acid catalyst. The equation is shown below.

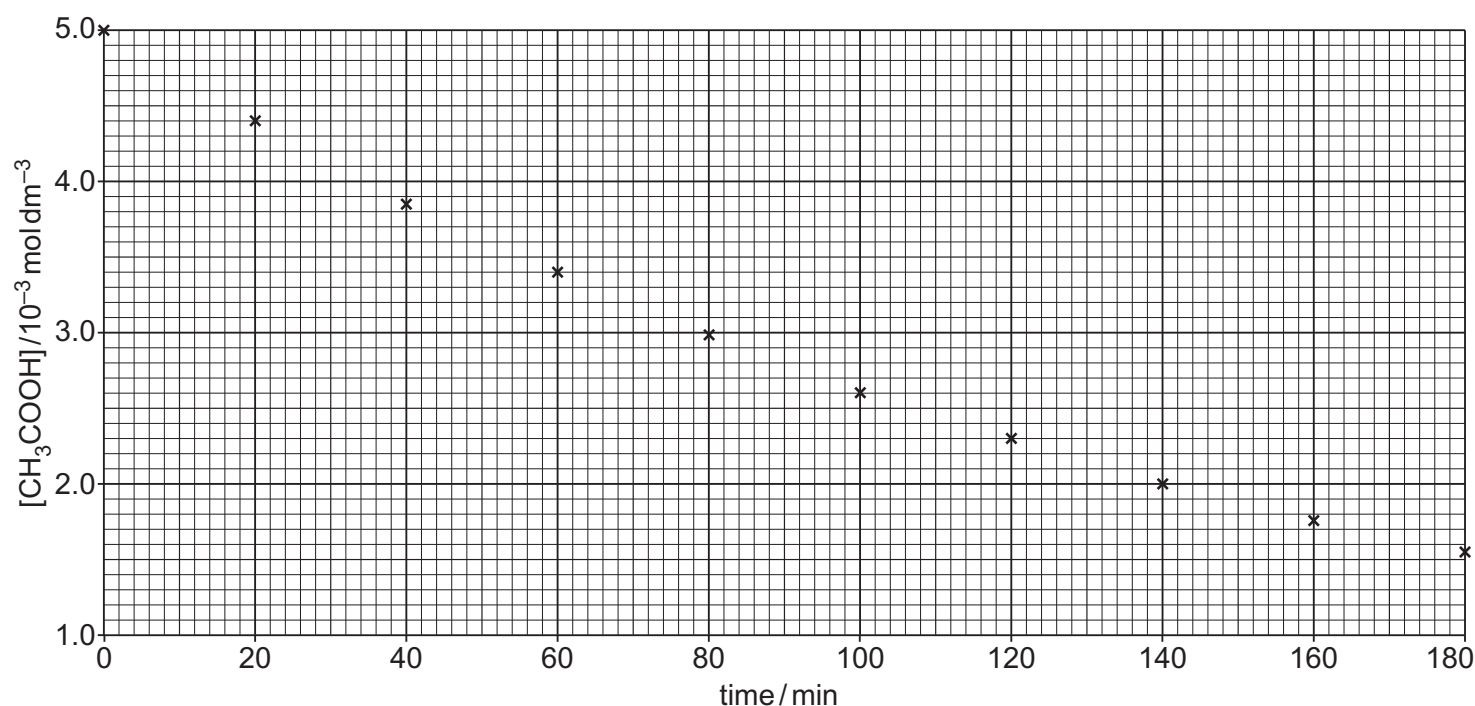


- (a) The student carries out an experiment to determine the order of reaction with respect to CH_3COOH .

The student uses a large excess of CH_3OH . The temperature is kept constant throughout the experiment.

The student takes a sample from the mixture every 20 minutes, and then determines the concentration of the ethanoic acid in each sample.

From the experimental results, the student plots the graph below.



- (i) Explain why the student uses a large excess of methanol in this experiment.

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..... [1]

- (ii) Use the half-life of this reaction to show that the reaction is first order with respect to CH_3COOH .

Show your working on the graph and below.

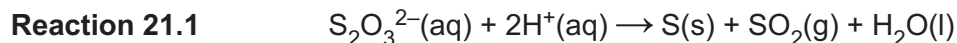
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..... [2]

- (iii) Determine the initial rate of reaction.

initial rate = ... $\text{mol dm}^{-3} \text{min}^{-1}$ [2]

- 15 This question is about how the rate of reaction is affected by changes in conditions.

A student carries out two investigations using the reaction between aqueous thiosulfate ions, $\text{S}_2\text{O}_3^{2-}(\text{aq})$, and aqueous hydrogen ions, $\text{H}^+(\text{aq})$.



- (a) In **Investigation 1**, the student determines how the rate of **Reaction 21.1** is affected by changes in concentration.

The results are shown in the table.

Experiment	$[\text{S}_2\text{O}_3^{2-}(\text{aq})]$ $/\text{mol dm}^{-3}$	$[\text{H}^+(\text{aq})]$ $/\text{mol dm}^{-3}$	initial rate $/\text{mol dm}^{-3} \text{s}^{-1}$
1	0.16	1.00	0.0120
2	0.08	1.00	0.0060
3	0.02	0.50	0.0015

From the results, the student concludes that the rate equation is
 $\text{rate} = k [\text{S}_2\text{O}_3^{2-}(\text{aq})]$

- (i) Explain how the student's results support this rate equation.

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..... [2]

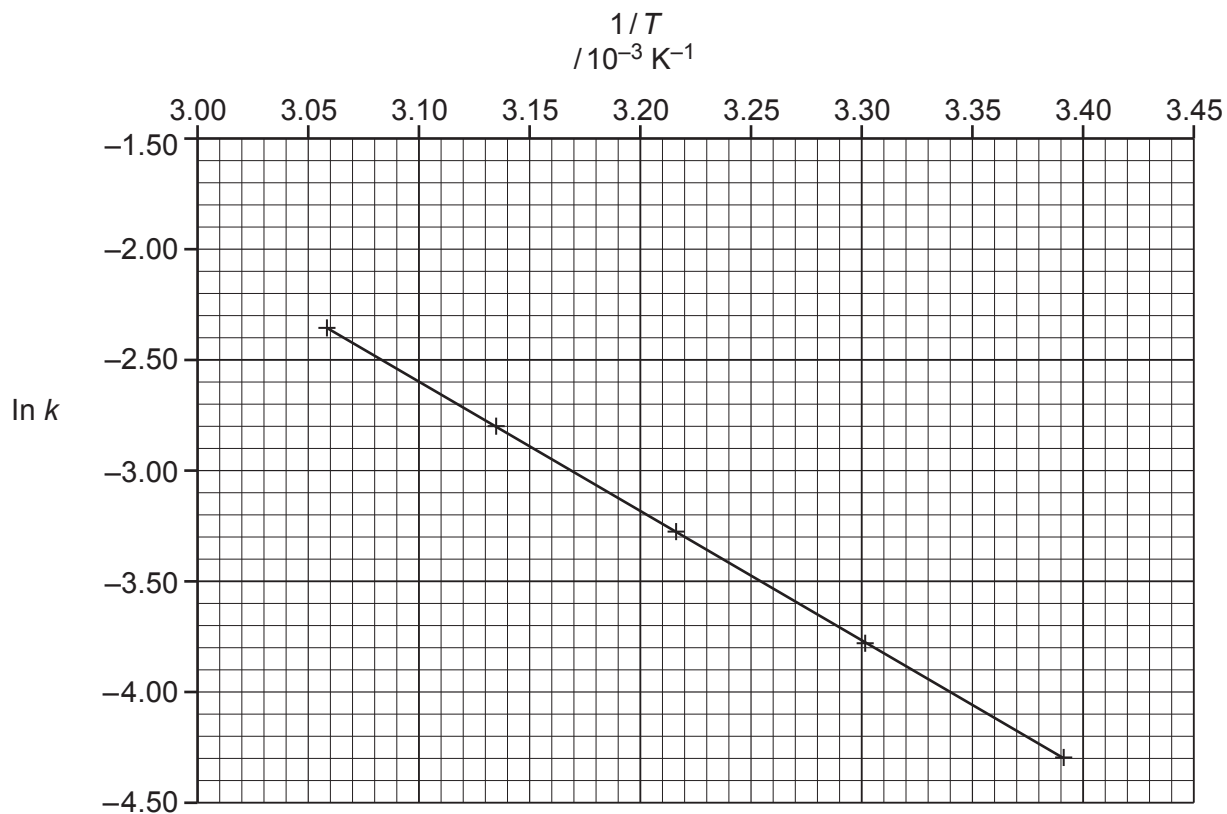
- (ii) Predict a possible two-step mechanism for **Reaction 21.1**.
The first step is the rate-determining step.

Step 1

Step 2 [2]

- (b) In **Investigation 2** the student determines the rate constant k of **Reaction 21.1** at different temperatures, T .

From the results, the student plots a graph of $\ln k$ against $1/T$ as shown below.



Graph 21.2

- (i) Calculate the activation energy, E_a , for **Reaction 21.1**, in kJ mol^{-1} .

Give your answer to **3** significant figures.

$$E_a = \dots\dots\dots \text{kJ mol}^{-1} \text{ [3]}$$

- (ii) From the graph the student estimates the value of $\ln A$ as -2.00 .
(A is the pre-exponential factor.)

Explain what mistake the student has made.

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..... [1]

- (iii) The student calculates the value of k in **Investigation 1** as 0.075s^{-1} .

Using **Graph 21.2**, determine the temperature, in $^{\circ}\text{C}$, at which **Investigation 1** was carried out.

temperature = ... $^{\circ}\text{C}$ [2]

- 16** The half-life for a first order reaction is 80 s.

What is the rate constant k , in s^{-1} , for this reaction?

A 8.66×10^{-3}

B 0.0125

C 55.5

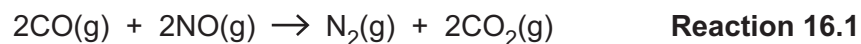
D 115

Your answer

[1]

- 17 A catalytic converter in a car removes nitrogen monoxide, NO, and carbon monoxide, CO, from the exhaust gases.

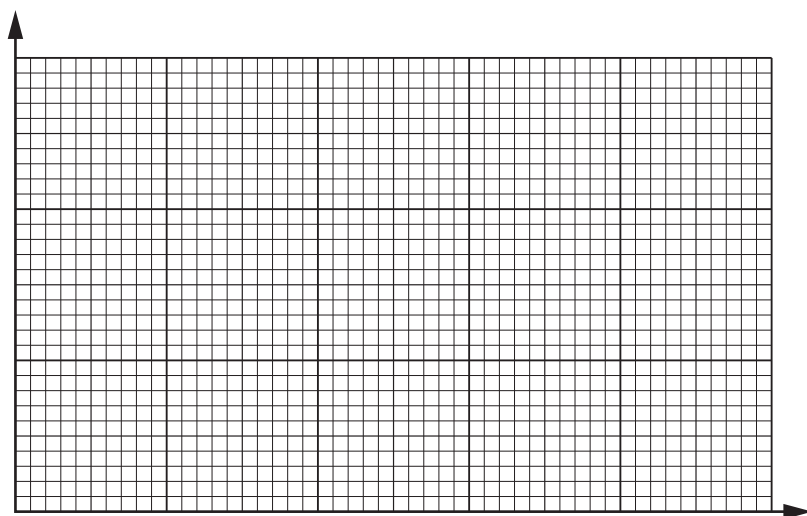
(a) One reaction that happens in a catalytic converter is shown below.



- (i) Explain how increasing the temperature increases the rate of **Reaction 16.1**.

Include a labelled sketch, using Boltzmann distributions, on the grid below.

Label the axes.



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..... [3]

- (ii) The rate of **Reaction 16.1** is investigated by carrying out three experiments at the same temperature. The results are shown below.

Experiment	[NO(g)] / mol dm ⁻³	[CO(g)] / mol dm ⁻³	Initial rate / mol dm ⁻³ s ⁻¹
1	2.75×10^{-4}	7.25×10^{-4}	1.85×10^{-4}
2	5.50×10^{-4}	7.25×10^{-4}	7.40×10^{-4}
3	1.10×10^{-3}	2.90×10^{-3}	1.18×10^{-2}

Determine the orders with respect to NO and CO, the rate equation, and the rate constant, k , including units.

Explain your reasoning.

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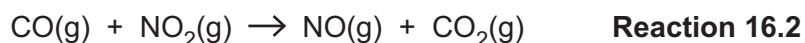
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$k =$ units [5]

- (b) Carbon monoxide also reacts with nitrogen dioxide as shown in **Reaction 16.2**.



The rate equation for **Reaction 16.2** is shown below:

$$\text{rate} = k[\text{NO}_2\text{(g)}]^2$$

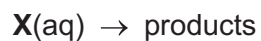
Suggest a possible two-step mechanism for **Reaction 16.2**.

The first step is much slower than the second step.

step 1

step 2

- 18** The reaction below is first order with respect to reactant **X**.



When the initial concentration of **X** is 1.0 mol dm^{-3} , the half-life is 16 minutes.

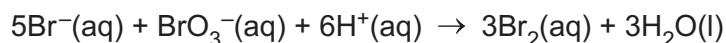
What is the half-life when the initial concentration of **X** is 2.0 mol dm^{-3} ?

- A** 2 minutes
- B** 4 minutes
- C** 8 minutes
- D** 16 minutes

Your answer

[1]

- 19 Bromine, Br₂, can be produced by the reaction:



A student investigates the rate of this reaction by carrying out four experiments at the same temperature. The student's results are shown below.

Experiment	[Br ⁻] /mol dm ⁻³	[BrO ₃ ⁻] /mol dm ⁻³	[H ⁺] /mol dm ⁻³	Initial rate /mol dm ⁻³ s ⁻¹
1	2.00 × 10 ⁻²	1.20 × 10 ⁻¹	8.00 × 10 ⁻²	2.52 × 10 ⁻⁴
2	6.00 × 10 ⁻²	1.20 × 10 ⁻¹	8.00 × 10 ⁻²	7.56 × 10 ⁻⁴
3	4.00 × 10 ⁻²	6.00 × 10 ⁻²	8.00 × 10 ⁻²	2.52 × 10 ⁻⁴
4	2.00 × 10 ⁻²	6.00 × 10 ⁻²	4.00 × 10 ⁻¹	3.15 × 10 ⁻³

Explain how the reaction orders can be determined from the student's results, and determine the rate equation and rate constant for this reaction.

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Additional answer space if required.

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- 20 Which row in the table explains how a catalyst affects the activation energy (E_a) and the proportion of molecules with energy $> E_a$?

	How the activation energy changes	Proportion of molecules with energy $> E_a$
A	decreases	decreases
B	decreases	increases
C	increases	decreases
D	increases	increases

Your answer

[1]

- 21 A graph of $\ln(k)$ is plotted against $1/T$ for a reaction.
(k = rate constant, T = temperature in K.)

The gradient has the numerical value of $-16\,000$.

What is the activation energy, in kJ mol^{-1} , for this reaction?

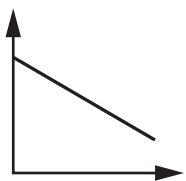
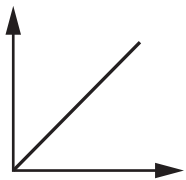
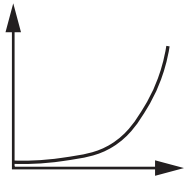
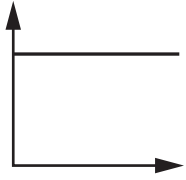
- A +1.92
- B +133
- C +1920
- D +133 000

Your answer

[1]

- 22 A reaction is zero order with respect to a reactant.

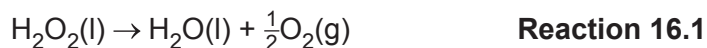
Which rate–concentration graph for the reactant is the correct shape?

A	
B	
C	
D	

Your answer

[1]

Hydrogen peroxide decomposes as shown in **Reaction 16.1**.



For Reaction 16.1, enthalpy change = -98 kJ mol⁻¹.

- 23(a)** The rate of decomposition of hydrogen peroxide shown in **Reaction 16.1** can be increased by adding a small amount of powdered manganese(IV) oxide, MnO₂.

The MnO₂ acts as a catalyst.

- (i)** Complete the enthalpy profile diagram for **Reaction 16.1** using formulae for the reactants and products.

- Use E_a to label the activation energy **without** MnO₂.
- Use E_c to label the activation energy **with** MnO₂.
- Use ΔH to label the enthalpy change of reaction.



[3]

- (ii)** Explain why MnO₂ is described as a **heterogeneous** catalyst for this reaction.

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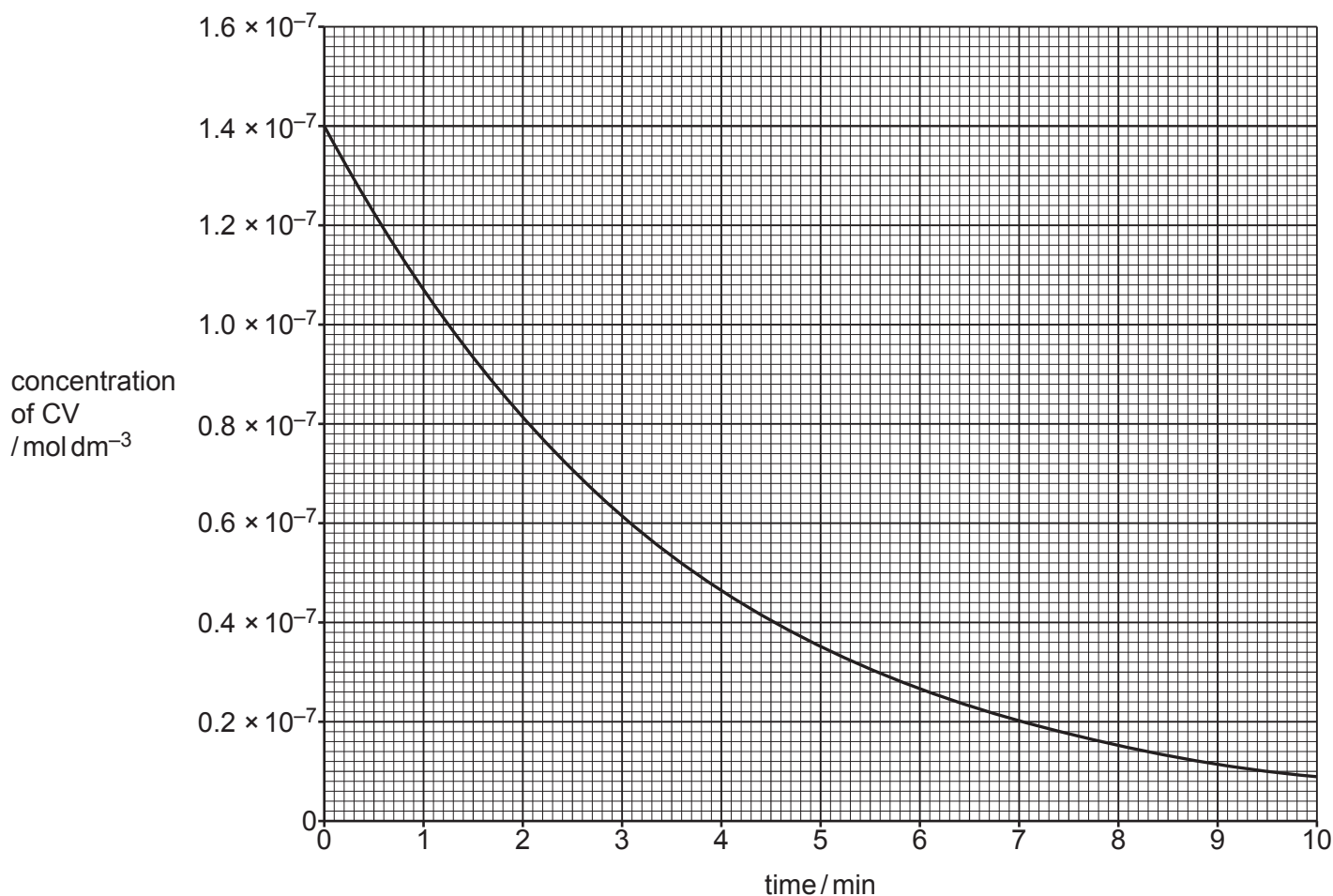
..... **[1]**

- 24 Crystal violet (CV) is a purple dye. In the presence of an alkali, CV reacts to form a colourless product.

A student uses a colorimeter to investigate the rate of the reaction between CV and sodium hydroxide, NaOH.

- The student mixes 10.0 cm^3 of $2.8 \times 10^{-7}\text{ mol dm}^{-3}$ CV with 10.0 cm^3 of 0.016 mol dm^{-3} NaOH.
- A large excess of NaOH is used, so that the reaction is effectively zero-order with respect to OH^- ions.
- The student places a sample of the reaction mixture in a colorimeter and measures the absorbance over time.

The student uses the absorbance readings to calculate the concentration of CV and plots a graph of concentration of CV against time, as shown below.



- (a) Using collision theory, explain why the gradient decreases over time.

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(b)* Use the graph to determine the order of reaction with respect to CV, the rate of the reaction at three minutes and the rate constant, k .

Your answer must show full working on the graph and on the lines below.

[6]

[illegible]

Extra answer space if required.

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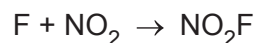
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25(a) Fluorine, F_2 , reacts with nitrogen dioxide, NO_2 , to form nitryl fluoride, NO_2F , as shown in **Reaction 20.1** below.



The mechanism for this reaction involves two steps. **Step 1** is the 'slow' step and **Step 2** is the 'fast' step.

The equation for **Step 2** is shown below:



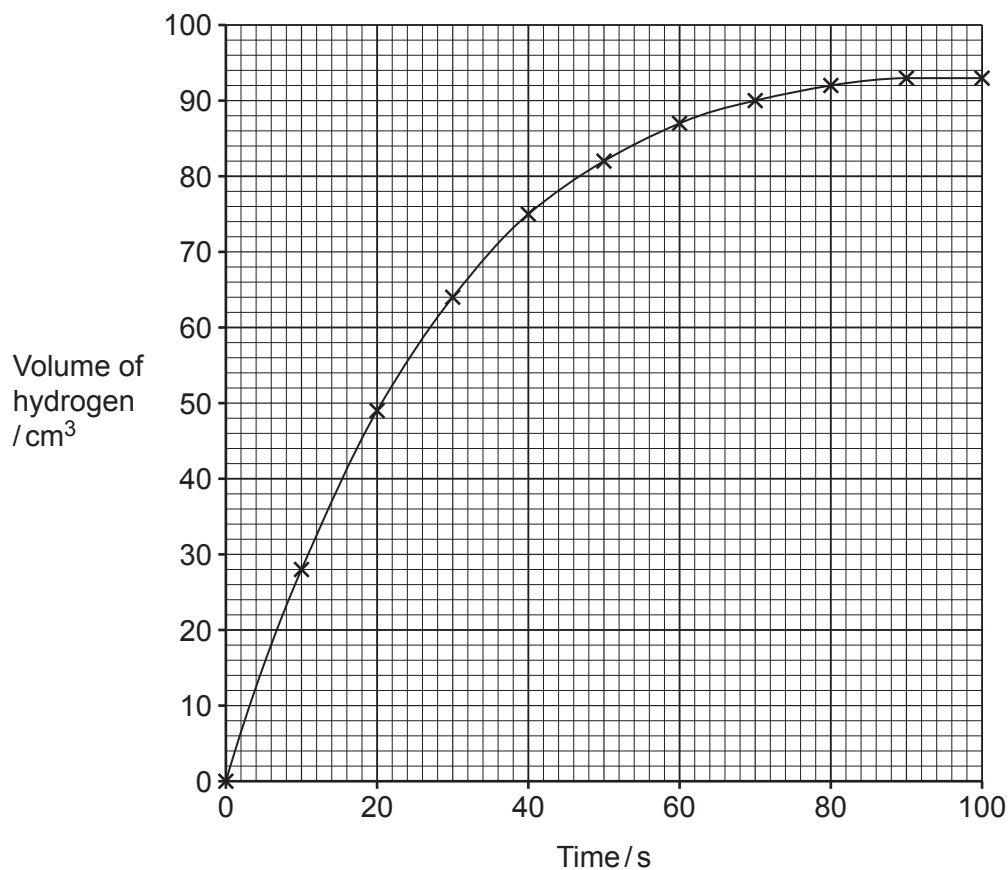
Suggest the equation for **Step 1** and the rate equation for **Reaction 20.1**.

Equation for **Step 1**:

Rate equation for **Reaction 20.1**:

[2]

- 26 The graph shows the volume of hydrogen released over time in a chemical reaction.



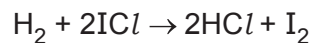
What is the rate of reaction, in cm^3s^{-1} , at 30 seconds?

- A 0.8
- B 1.2
- C 2.1
- D 2.4

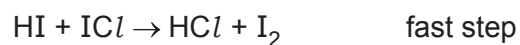
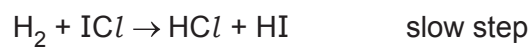
Your answer

[1]

- 27 Hydrogen reacts with iodine monochloride, ICl , as shown in the equation:



The mechanism for the reaction is:



What is the rate equation for the reaction between H_2 and ICl ?

- A** rate = $k[\text{HI}][\text{ICl}]$
B rate = $k[\text{H}_2][\text{ICl}]$
C rate = $k[\text{H}_2][\text{ICl}]^2$
D rate = $k[\text{H}_2][\text{HI}][\text{ICl}]^2$

Your answer

[1]

- 28** The rate constant, k , for a first order reaction is $6.19 \times 10^{-3} \text{ s}^{-1}$.

What is the half-life, in s, for the reaction?

A 4.29×10^{-3}

B 49

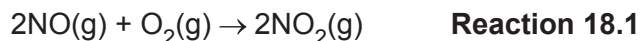
C 112

D 323

Your answer

[1]

- 29 Nitrogen monoxide, NO, reacts with oxygen to form nitrogen dioxide, NO₂.



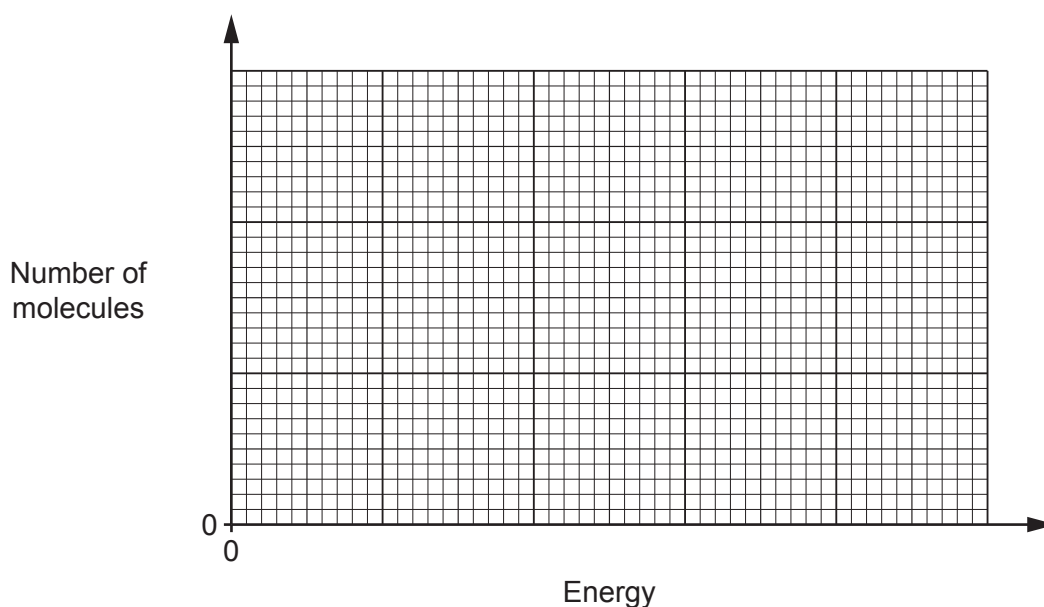
- (a) The rate of reaction increases in the presence of a heterogeneous platinum catalyst.

- (i) Explain why platinum is an example of a *heterogeneous* catalyst for **Reaction 18.1**.

.....
 [1]

- (ii) Explain how the platinum catalyst increases the rate of reaction.

Include a labelled sketch of the Boltzmann distribution on the grid.



.....

 [2]

- (iii) In industry, catalysts have great economic importance and benefits for increased sustainability.

Explain how the use of catalysts is beneficial to the environment.

.....

- Reaction 18.1** at a constant temperature.

Experiment	[NO]/mol dm ⁻³	[O ₂]/mol dm ⁻³	rate/mol dm ⁻³ s ⁻¹
1	0.040	0.035	7.45 × 10 ⁻⁴
2	0.040	0.070	1.49 × 10 ⁻³
3	0.060	0.035	1.68 × 10 ⁻³

Determine the orders with respect to NO and O₂. Explain your reasoning.

Determine the rate equation and calculate the rate constant, k , including units.

[illegible]

$k =$ units

[5]

- (c) The rate constant varies with temperature, according to the Arrhenius equation.

A technician determines the activation energy for **Reaction 18.1** by determining the rate constant, k , at different temperatures.

The table shows the technician's results.

Temperature /K	$\frac{1}{T}/\text{K}^{-1}$	k	$\ln k$
1500	6.7×10^{-4}	1.81×10^3	7.50
1750	5.7×10^{-4}	8.51×10^4	
2000	5.0×10^{-4}	2.79×10^5	12.5
2250	4.4×10^{-4}	1.47×10^6	14.2
2500	4.0×10^{-4}	5.70×10^6	15.6

- (i) Complete the table by adding the missing value of $\ln k$.

[1]

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

Use your graph to calculate the value of the activation energy, E_a , in kJ mol^{-1} .

Give your answer to **2** significant figures.

$$E_a = \text{..... kJ mol}^{-1} \text{ [4]}$$

