# How Fast

**1.** A student investigates the reaction between ethanoic acid, $CH_3COOH(l)$ and methanol, $CH_3OH(l)$, in the presence of an acid catalyst. The equation is shown below.

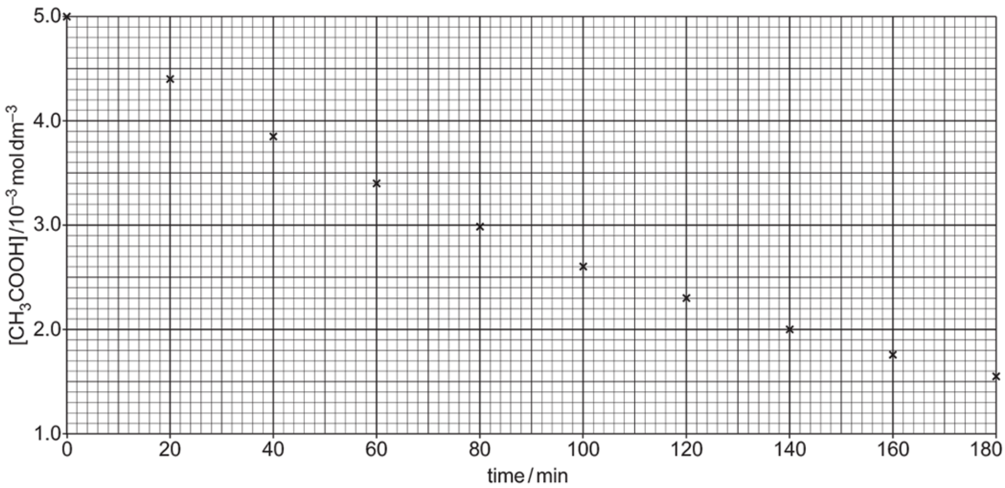
$CH_3COOH(l) + CH_3OH(l) \rightleftharpoons CH_3COOCH_3(l) + H_2O(l)$

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 = ................................ mol $dm^{-3}$ $min^{-1}$ **[2]**

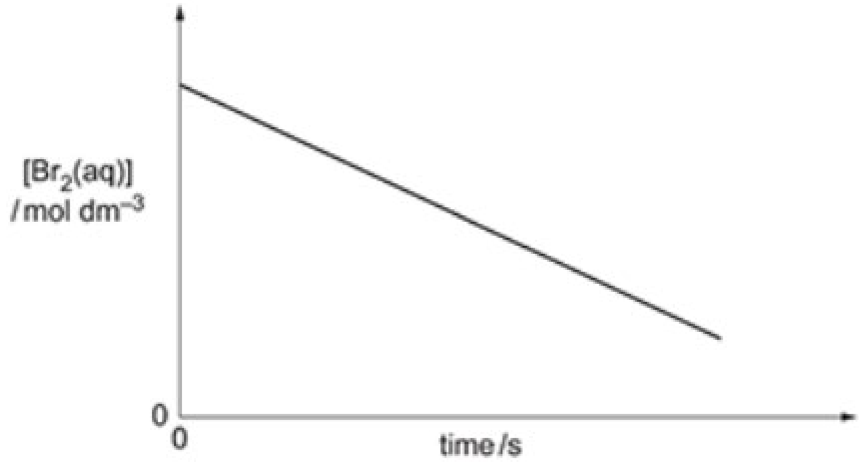
**2.** Three students carry out a rates investigation on the reaction between bromine and propanone in the presence of hydrochloric acid.

$CH_3COCH_3(aq) + Br_2(aq) \rightarrow CH_3COCH_2Br(aq) + HBr(aq)$

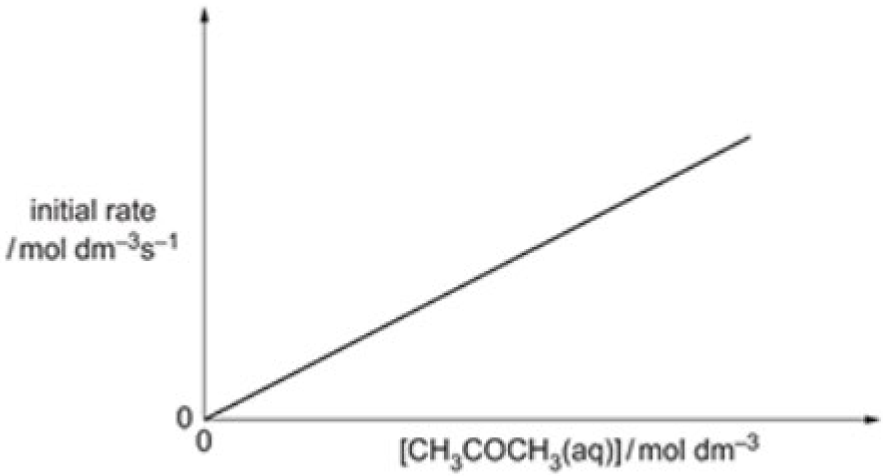
Each student investigates the effect of changing the concentration of one of the reactants whilst keeping the other concentrations constant.

Their results are shown below.

**Results of student 1**




**Results of student 2**




**Results of student 3**

| Experiment | $[Br_2(aq)]$ / mol $dm^{-3}$ | $[CH_3COCH_3(aq)]$ / mol $dm^{-3}$ | $[H^+(aq)]$ / mol $dm^{-3}$ | Initial rate / $10^{-5}$ mol $dm^{-3}$ $s^{-1}$ |
|---|---|---|---|---|
| 1 | 0.004 | 1.60 | 0.20 | 1.25 |
| 2 | 0.004 | 1.60 | 0.40 | 2.50 |

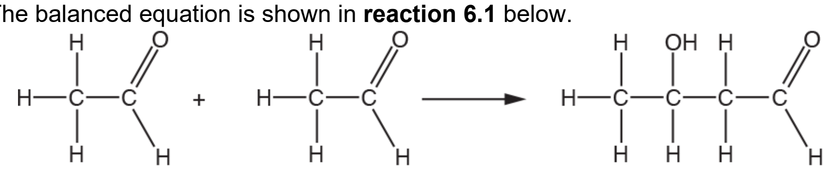
Explain how the reaction orders can be determined from the students' results, and determine the rate equation and rate constant.

**[6]**

**3.** This question is about organic reactions.

Compound **A** is formed when ethanal is mixed with $OH^-(aq)$ ions, which act as a catalyst.

The balanced equation is shown in **reaction 6.1** below.



**Reaction 6.1**

Compound **A**

i. Give the systematic name for compound **A**.

**[1]**

ii. What type of reaction has taken place?

**[1]**

iii. **Reaction 6.1** takes place in two steps. $OH^-$ ions act as a catalyst.

In **step 1**, ethanal reacts with $OH^-$ ions to set up an acid–base equilibrium.
In **step 2**, compound **A** is formed.

- Complete the equilibrium for **step 1** and label the conjugate acid–base pairs as: **A1**, **B1** and **A2**, **B2**.

$CH^3CHO$ + $OH^-$ ⇌ ________ + ________

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- Suggest the equation for **step 2**.

**[3]**

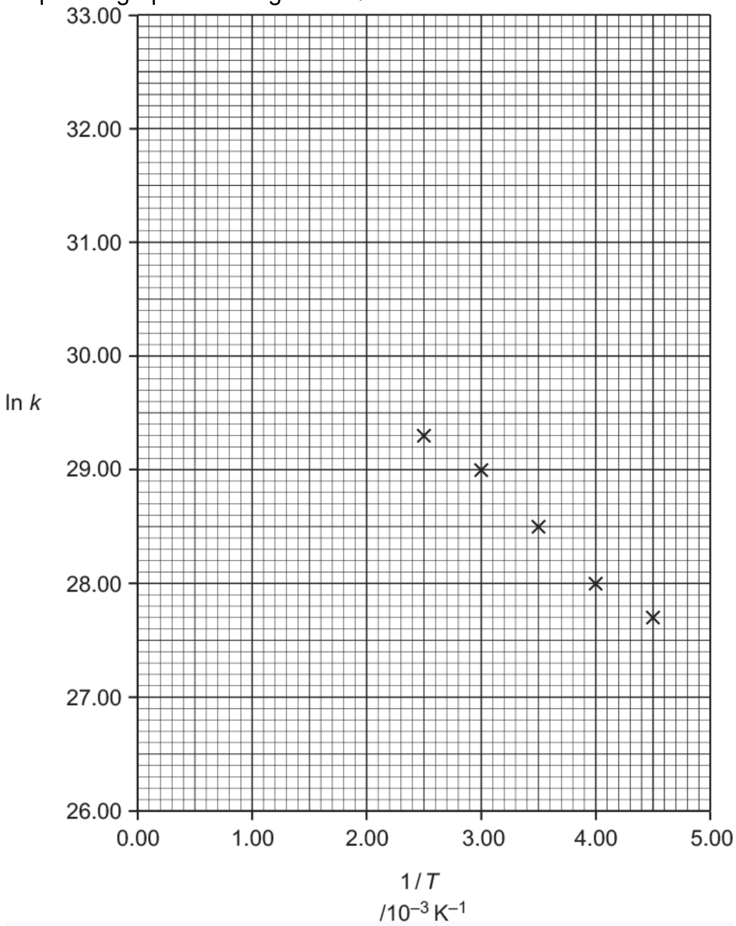
iv. A similar reaction takes place when propanone, $(CH_3)_2CO$, is mixed with $OH^-(aq)$ ions.

Draw the structure of the organic product of this reaction.

**[1]**

**4(a).** 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, $E_a$ = + ............................................ 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* = ........................................................ **[2]**

**(b).** This question is about reaction rates.

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

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

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

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

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

**[6]**

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**5(a).** Aqueous solutions of hydrogen peroxide, $H_2O_2(aq)$, decompose as in the equation below.

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

A student investigates the decomposition of $H_2O_2(aq)$ 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 H2O2(aq) at each time. The student then plots a concentration–time graph.

Suggest a different experimental method that would allow the rate of this reaction to be followed over time.

**[1]**

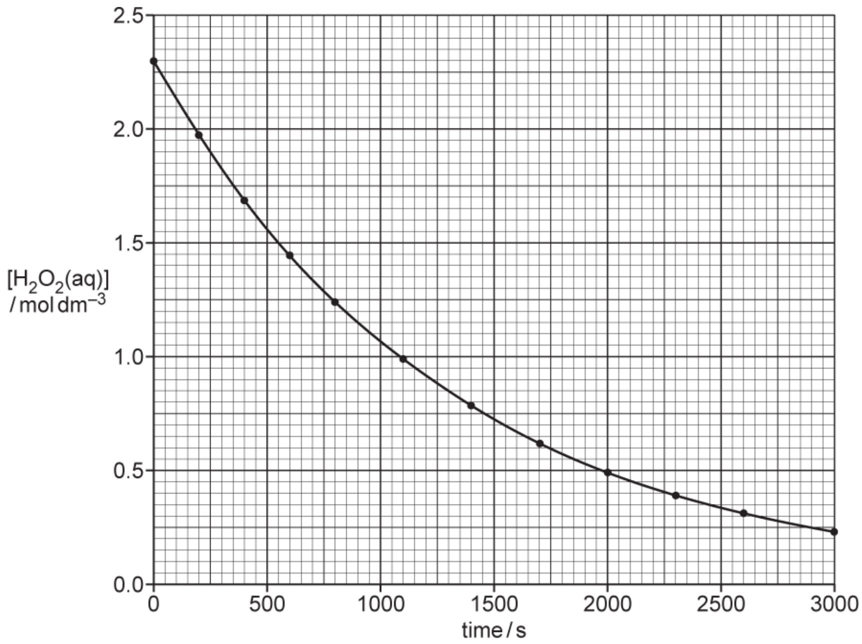
**(b).** Aqueous solutions of hydrogen peroxide, $H_2O_2(aq)$, decompose as in the equation below.

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

A student investigates the decomposition of $H_2O_2(aq)$ 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 $H_2O_2(aq)$ at each time. The student then plots a concentration–time graph.




Determine the initial rate of reaction, the order with respect to $H_2O_2$, and the rate constant.

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

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**(c).** 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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**6.** This question is about reactions of hydrogen peroxide, $H_2O_2$.

Hydrogen peroxide, $H_2O_2$, iodide ions, $I^-$, and acid, $H^+$, react as shown in the equation below.

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

A student carries out several experiments at the same temperature, using the initial rates method, to determine the rate constant, *k*, for this reaction.

The results are shown below.

| | **Initial concentrations** | | | |
|---|---|---|---|---|
| **Experiment** | **$[H_2O_2(aq)]$ / mol dm$^{-3}$** | **$[I^-(aq)]$ / mol dm$^{-3}$** | **$[H^+(aq)]$ / mol dm$^{-3}$** | **Rate / $10^{-6}$ mol dm$^{-3}$ s$^{-1}$** |
| **1** | 0.0100 | 0.0100 | 0.100 | 2.00 |
| **2** | 0.0100 | 0.0200 | 0.100 | 4.00 |
| **3** | 0.0200 | 0.0100 | 0.100 | 4.00 |
| **4** | 0.0200 | 0.0100 | 0.200 | 4.00 |

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

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*k* = ____________________ units **[3]**

ii. The rate constant, *k*, for this reaction is determined at different temperatures, *T*.

Explain how the student could determine the activation energy, $E_a$, for the reaction graphically using values of *k* and *T*.

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---[3]

**7(a).** A student investigates the rate of reaction between iodine, $I_2$, and propanone, $CH_3COCH_3$, in the presence of $H^+$ ions. The student uses $HCl(aq)$ to supply $H^+$ ions.

$$I_2(aq) + CH_3COCH_3(aq) \rightarrow CH_3COCH_2I(aq) + HI(aq)$$

The student follows the method outlined below.

**1.** The student starts the reaction by mixing the following solutions.

1.00 $cm^3$ of 1.00 mol $dm^{-3}$ $I_2(aq)$
49.5 $cm^3$ of 1.00 mol $dm^{-3}$ $CH_3COCH_3(aq)$
49.5 $cm^3$ of 1.00 mol $dm^{-3}$ $HCl(aq)$

**2.** The student places a sample of the reaction mixture in a colorimeter, immediately starts a stopwatch, and records the absorbance.

**3.** The student records the absorbance every 100 s. The results are shown below.

| Time/s | Absorbance |
|---|---|
| 0 | 0.80 |
| 100 | 0.67 |
| 200 | 0.51 |
| 300 | 0.44 |
| 400 | 0.28 |
| 500 | 0.18 |
| 600 | 0.05 |

Explain why absorbance decreases during the experiment.

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---[1]

**(b).** Absorbance is proportional to the concentration of $I_2$.

Calculate the concentration of $I_2$ at the start of the experiment and after 500 s.

| **Time/s** | **Absorbance** | **$[I_2(aq)]$/mol $dm^{-3}$** |
|---|---|---|
| 0 | 0.80 | |
| 500 | 0.18 | |

**[2]**

**(c).** i. Plot a graph of absorbance against time and draw a line of best fit.

**[3]**

ii. Use your graph to find the order of reaction with respect to iodine.

Explain your reasoning.

Order ..........

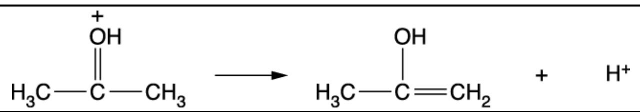
Explanation ..........

**[2]**

**(d).** A three step mechanism has been proposed for this reaction.

$$I_2(aq) + CH_3COCH_3(aq) \rightarrow CH_3COCH_2I(aq) + HI(aq)$$

Complete the mechanism by adding equations for **Step 1** and **Step 3** in the boxes below.

| | |
|---|---|
| **Step 1** (slow) | |
| **Step 2** (fast) |  |
| **Step 3** (fast) | |

**[2]**

**8.** This question is about redox reactions.

*Bromine, $Br_2$, is formed in the redox reaction shown below.

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

A student plans an investigation, using the initial rates method, to determine the rate equation and rate constant for this reaction.

The student is supplied with solutions containing the following:

- 0.300 mol dm$^{-3}$ $Br^-(aq)$
- 0.300 mol dm$^{-3}$ $BrO_3^-(aq)$
- 0.300 mol dm$^{-3}$ $H^+(aq)$.

The student is also supplied with distilled water and normal laboratory glassware.

The student uses a total volume of 30 cm$^3$ for each experiment and measures the initial rate of formation of $Br_2(aq)$.

The results of the student's experiments are shown below.

| Experiment | $[Br^-(aq)]$ / mol $dm^{-3}$ | $[BrO_3^-(aq)]$ / mol $dm^{-3}$ | $[H^+(aq)]$ / mol $dm^{-3}$ | Initial rate / $10^{-3}$ mol $dm^{-3}$ $s^{-1}$ |
|---|---|---|---|---|
| **1** | 0.100 | 0.100 | 0.100 | 1.20 |
| **2** | 0.025 | 0.100 | 0.100 | 0.30 |
| **3** | 0.100 | 0.050 | 0.100 | 0.60 |
| **4** | 0.100 | 0.050 | 0.050 | 0.15 |

Show how the student could obtain the concentrations for experiments **1–4** and determine the rate constant for this reaction.

**[6]**

**9(a).** Hydrogen peroxide reacts with iodide ions in acid conditions, as shown below.

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

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** | **$[H_2O_2(aq)]$ / mol dm$^{-3}$** | **$[I^-(aq)]$ / mol dm$^{-3}$** | **$[H^+(aq)]$ / mol dm$^{-3}$** | **Initial rate / mol dm$^{-3}$ s$^{-1}$** |
|---|---|---|---|---|
| **1** | 0.0010 | 0.20 | 0.10 | $5.70 \times 10^{-6}$ |
| **2** | 0.0020 | 0.20 | 0.10 | $1.14 \times 10^{-5}$ |
| **3** | 0.0020 | 0.20 | 0.20 | $1.14 \times 10^{-5}$ |
| **4** | 0.0040 | 0.40 | 0.10 | $4.56 \times 10^{-5}$ |

The rate equation is: *rate* = *k* $[H_2O_2(aq)]\,[I^-(aq)]$

- Show that the student's results support this rate equation.
- Calculate the rate constant, *k*, for this reaction.

Give your answer to **two** significant figures, in standard form and with units.

*In your answer you should make clear how the experimental results provide evidence for the rate equation.*

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______________________________________________________________________ **[6]**

**(b).** The student concluded that $H^+(aq)$ ions act as a catalyst.

Explain why the student's conclusion is **not** correct.

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______________________________________________________________________ **[1]**

**(c).** A four-step mechanism has been proposed for this reaction.
The rate-determining step is the first step.

i. State what is meant by the term *rate-determining step.*

______________________________________________________________________

______________________________________________________________________ **[1]**

ii. The equation for **Step 3** in the four-step mechanism is shown below.

Suggest equations for the other three steps.
State symbols are **not** required.

**Step 1**: ..........................................................................................

**Step 2**: ..........................................................................................

**Step 3**: $HIO + I^- \rightarrow I_2 + OH^-$

**Step 4**: ..........................................................................................

**[3]**

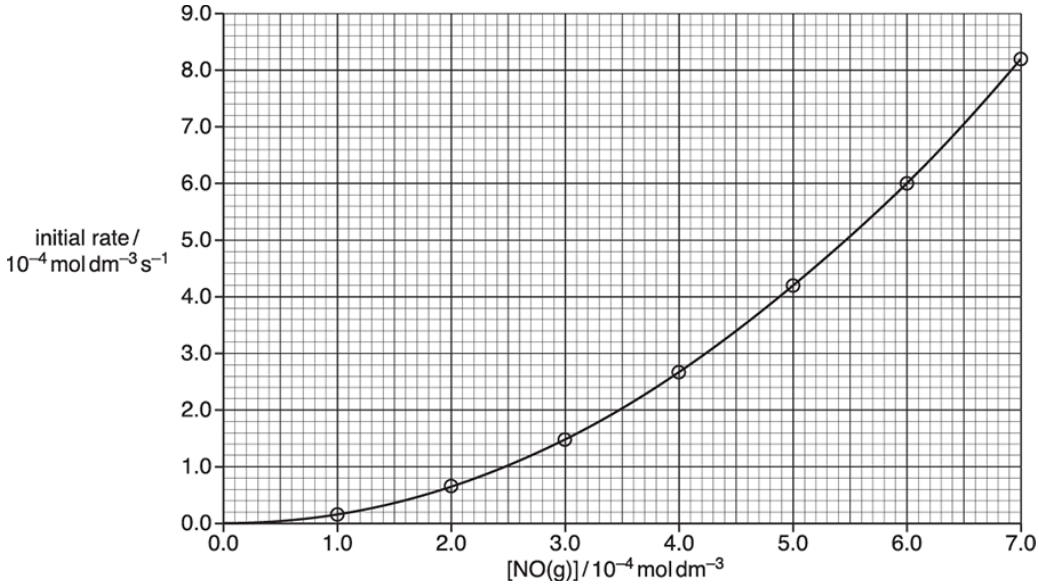
**10(a).** Hydrogen, $H_2$, reacts with nitrogen monoxide, NO, as shown below:

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

The rate equation for this reaction is:

*rate* = $k[H_2(g)][NO(g)]^2$

The concentration of NO(g) is changed and a rate–concentrationgraph is plotted.




The chemist uses $H_2$(g) of concentration $2.0 \times 10^{-2}$ mol dm$^{-3}$.

Using values from the graph, calculate the rate constant, *k*, for this reaction.

Give your answer to **two** significant figures and in **standard form**.

Show your working.

*k* = .............................................. units .............................................. **[4]**

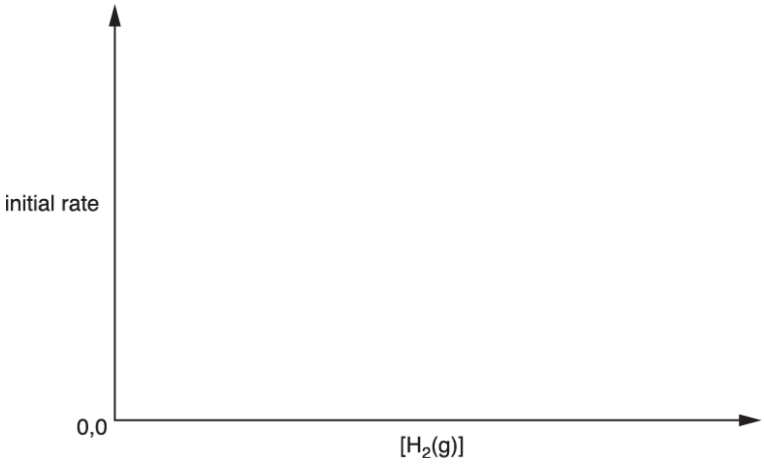
**(b).** A chemist investigates the effect of changing the concentration of $H_2$(g) on the initial reaction rate at two different temperatures.

The reaction is first order with respect to $H_2$(g).

i. Using the axes below, sketch **two** graphs of the results.

Label the graphs as follows:
- **L** for the lower temperature

- **H** for the higher temperature.



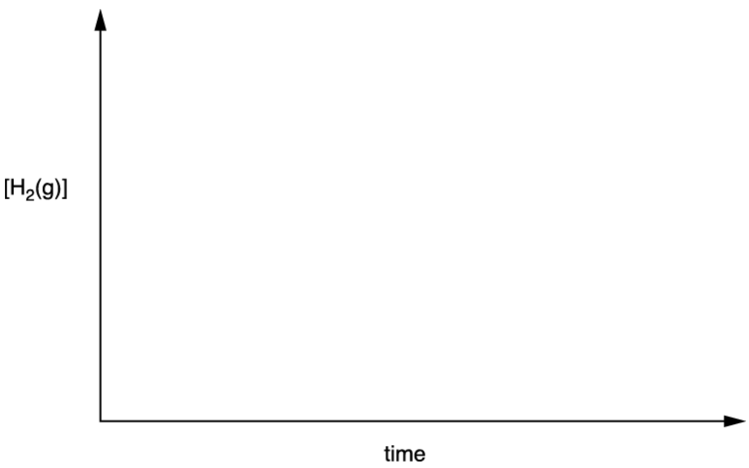

**[2]**

ii. State the effect of the higher temperature on the rate constant, *k*.

_______________________________________________________________ **[1]**

**(c).** The reaction can also be shown as being first order with respect to $H_2(g)$ by continuous monitoring of $[H_2(g)]$ during the course of the reaction.

- Using the axes below, sketch a graph to show the results.
- State how you would use the graph to show this first order relationship for $H_2(g)$.

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_______________________________________________________________ **[2]**

**(d).** The chemist proposes a three-step mechanism for the reaction:

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

i. On the dotted line below, write the equation for step 3.

| | | |
|---|---|---|
| step 1: | $2NO \rightarrow N_2O_2$ | fast |
| step 2: | $H_2 + N_2O_2 \rightarrow N_2O + H_2O$ | slow |
| step 3: | .................................................. | fast |

**[1]**

ii. Explain why this mechanism is consistent with the rate equation *rate* = $k[H_2(g)][NO(g)]^2$.

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---------------------------------------------------------------------------------------------------- **[1]**

**11(a).** A student carries out an initial rates investigation on the reaction below.

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

From the results, the student determines the rate equation for this reaction:

$$\text{rate} = k\,[I^-(aq)]^2\,[IO_3^-(aq)]\,[H^+(aq)]^2$$

i. What is the overall order of reaction?

---------------------------------------------------------------------------------------------------- **[1]**

ii. A proposed mechanism for this reaction takes place in several steps.

Suggest **two** reasons why it is unlikely that this reaction could take place in one step.

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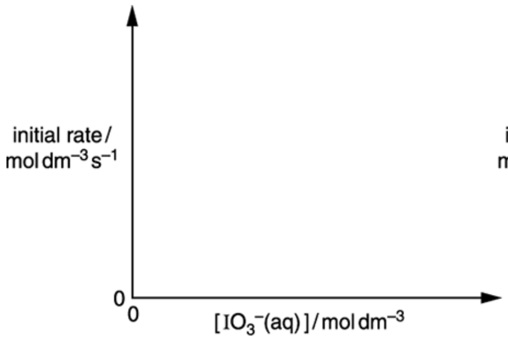
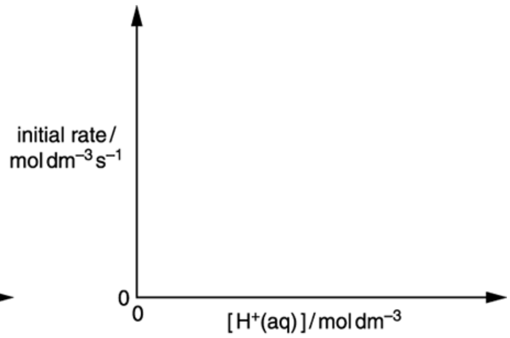
---------------------------------------------------------------------------------------------------- **[2]**

**(b).** On the rate—concentration graphs below, sketch lines to show the relationship between initial rate and concentration for $IO_3^-$(aq) and $H^+$(aq).




**[2]**

**(c).** The table below shows some of the student's results.

i. Complete the table by adding the missing initial rates in the boxes.

| | **[$I^-$(aq)] / mol dm$^{-3}$** | **[$IO_3^-$(aq)] / mol dm$^{-3}$** | **[$H^+$(aq)] / mol dm$^{-3}$** | **Initial rate / mo dm$^{-3}$ s$^{-1}$** |
|---|---|---|---|---|
| **Experiment 1** | 0.015 | 0.010 | 0.020 | 0.60 |
| **Experiment 2** | 0.045 | 0.010 | 0.020 | |
| **Experiment 3** | 0.060 | 0.040 | 0.080 | |

**[2]**

ii. Calculate the rate constant, *k*, for this reaction. Include units.

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

*k* = ........................................... units ...........................................**[3]**

iii. The student repeats Experiment 1 using 0.020 mol dm$^{-3}$ methanoic acid, HCOOH(aq) ($pK_a$ = 3.75), instead of 0.020 mol dm$^{-3}$ HC*l*(aq) as a source of $H^+$(aq).

Determine the initial rate in this experiment. Show your working.

initial rate = .............................................. mol dm$^{-3}$ s$^{-1}$ **[3]**

**12.** This question is about numbers and patterns in chemistry.

This question looks at number relationships. For calculations, show your working.

i. What is the oxidation number of nitrogen in each species?

$N_2O_4$ ............. $NO_3^-$ ............. $NH_4^+$ .............

**ii. [1]**

iii. What mass of $KMnO_4$ is needed to prepare a 250.0 $cm^3$ solution with a concentration of 0.200 mol $dm^{-3}$ $KMnO_4$?

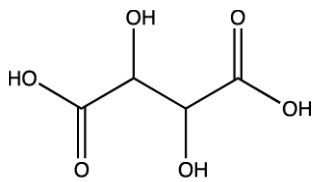
mass = ............................ g **[2]**

iv. What are the units of the rate constant for a reaction with an overall order of 3?

units = ............................ **[1]**

v. How many molecules are in 38.25 g of tartaric acid?

Give your answer to an **appropriate** number of significant figures and in standard form.




**tartaric acid**

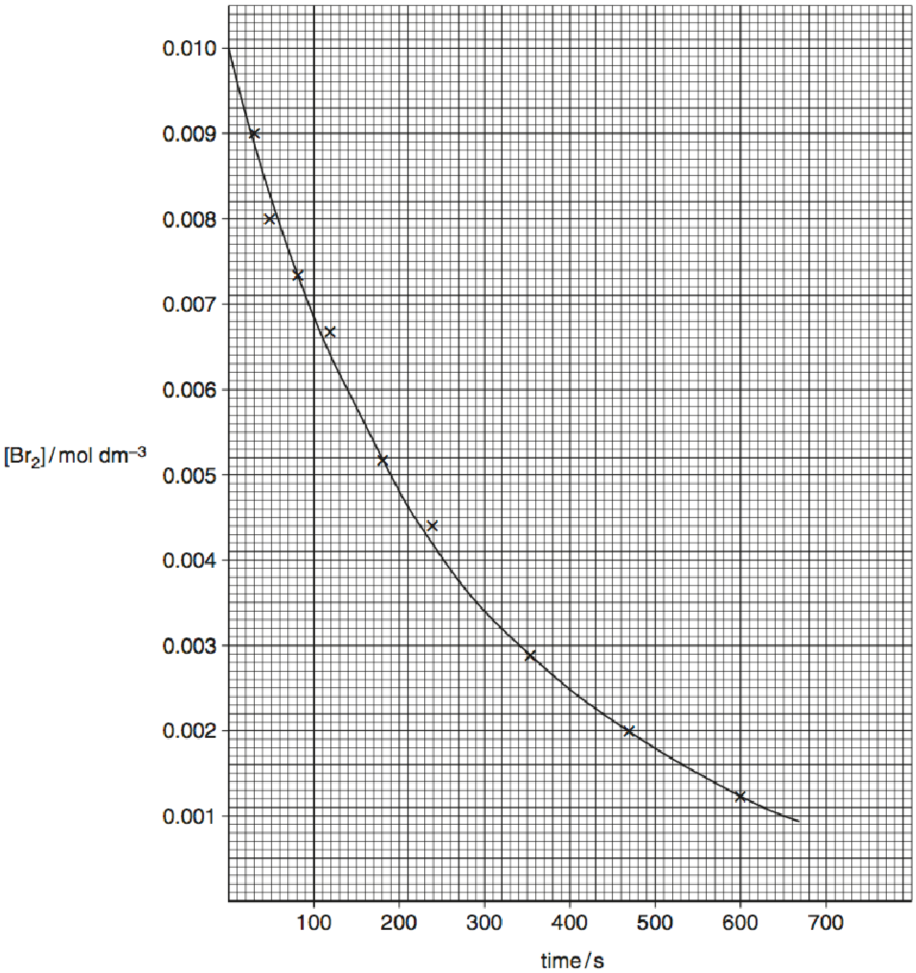
number of molecules = ............................ **[2]**

**13.** Methanoic acid and bromine react as in the equation below.

$Br_2(aq) + HCOOH(aq) \rightarrow 2H^+(aq) + 2Br^-(aq) + CO_2(g)$

A student investigates the rate of this reaction by monitoring the concentration of bromine over time.
The student uses a large excess of HCOOH to ensure that the order with respect to HCOOH will be effectively zero.

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




a. Suggest how the concentration of the bromine could have been monitored.

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_______________________________________________________________________________ **[1]**

b. Suggest a different experimental method that would allow the rate of this reaction to be followed over time.

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_______________________________________________________________________________ **[1]**

c. Why would use of excess HCOOH ensure that the order with respect to HCOOH is effectively zero?

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------------------------------------------------------------------------------------------------ **[1]**

d. * Using the graph, determine
   - the initial rate of reaction
   - the rate constant.

Your answer must show full working using the graph and the lines below as appropriate.

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------------------------------------------------------------------------------------------------ **[6]**

**14(a).** This question is about reaction rates.

Ozone, $O_3$, reacts with nitrogen dioxide, $NO_2$, as shown below.

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

A student carries out three experiments to investigate how different concentrations of $O_3(g)$ and $NO_2(g)$ affect the initial rate of this reaction.

The results of the three experiments are shown below

| **Experiment** | $[O_3(g)]$ / mol dm$^{-3}$ | $[NO_2(g)]$ / mol dm$^{-3}$ | Initial rate, / mol dm$^{-3}$ s$^{-1}$ |
|---|---|---|---|
| **1** | $1.00 \times 10^{-3}$ | $2.50 \times 10^{-3}$ | $3.20 \times 10^{-8}$ |
| **2** | $3.00 \times 10^{-3}$ | $2.50 \times 10^{-3}$ | $9.60 \times 10^{-8}$ |
| **3** | $3.00 \times 10^{-3}$ | $5.00 \times 10^{-3}$ | $1.92 \times 10^{-7}$ |

* Determine the rate constant and a possible two-step mechanism for this reaction that is consistent with these results.

Your response should clearly show how your calculations and explanations are linked to the experimental results.

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______________________________________________________________________ **[6]**

**(b).** A student carries out an investigation to find the activation energy, $E_a$, of a reaction.

From the results, the student determines the rate constant, *k*, at different temperatures, *T*.

The student then calculates 1 / *T* and ln *k*, as shown in **Table 19.1**.

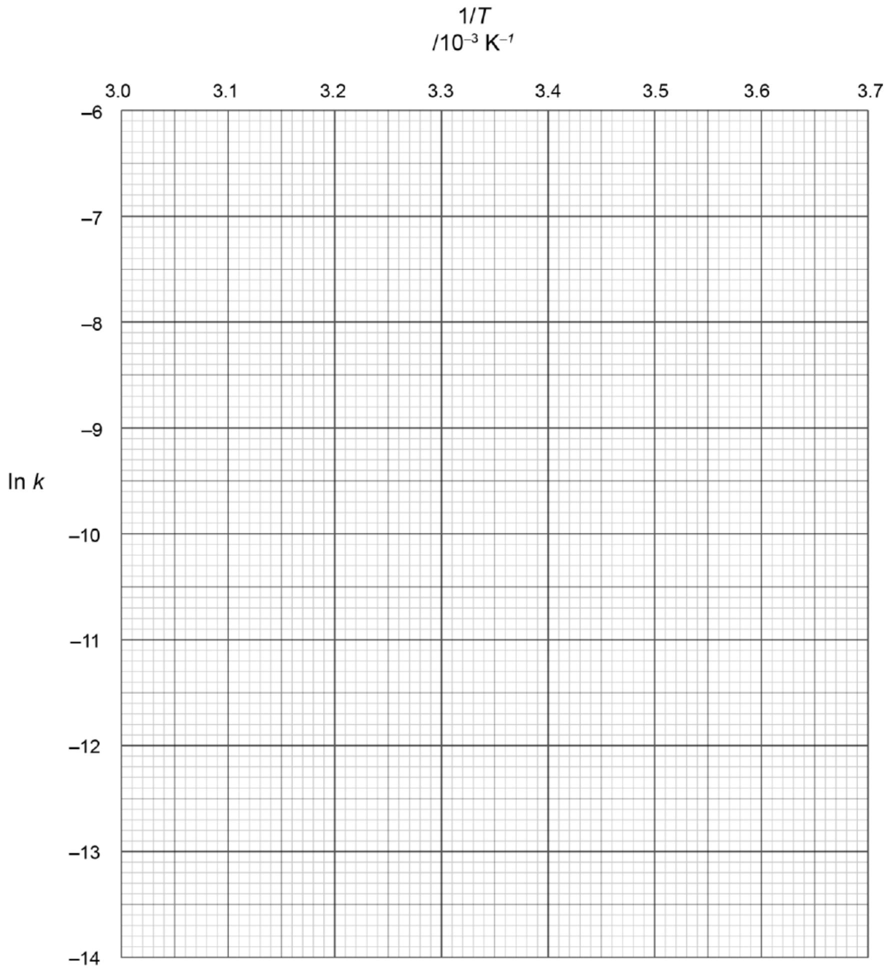
| **Temperature, *T* / K** | **Rate constant, *k* / s$^{-1}$** | **1 / *T* / K$^{-1}$** | **ln *k*** |
|---|---|---|---|
| 278 | $1.50 \times 10^{-6}$ | $3.60 \times 10^{-3}$ | −13.41 |
| 290 | $1.51 \times 10^{-5}$ | | −11.10 |
| 298 | $4.11 \times 10^{-5}$ | $3.34 \times 10^{-3}$ | −10.10 |
| 308 | $1.99 \times 10^{-4}$ | $3.23 \times 10^{-3}$ | |
| 323 | $1.40 \times 10^{-3}$ | $3.10 \times 10^{-3}$ | −6.57 |

**Table 19.1**

Add the missing values to **Table 19.1** and plot a graph of ln *k* against 1 / *T* on the graph paper opposite.

Using your graph, calculate the activation energy of the reaction.

Show your working.




activation energy, $E_a$ = .......................................................... kJ mol$^{-1}$ **[4]**

**15.** Iodine monochloride, ICl, can react with hydrogen to form iodine.

$2ICl + H_2 \rightarrow 2HCl + I_2$

This reaction was carried out several times using different concentrations of ICl or $H_2$. The initial rate of each experiment was calculated and the results are shown below. Initial concentrations are shown for each experiment.

| | **[ICl] / mol dm$^{-3}$** | **[$H_2$] / mol dm$^{-3}$** | **Rate / mol dm$^{-3}$ s$^{-1}$** |
|---|---|---|---|
| **Experiment 1** | 0.250 | 0.500 | $2.04 \times 10^{-2}$ |
| **Experiment 2** | 0.500 | 0.500 | $4.08 \times 10^{-2}$ |
| **Experiment 3** | 0.125 | 0.250 | $5.10 \times 10^{-3}$ |

i. Calculate the rate constant, *k*, for this reaction. Include units in your answer.

Show **all** your working.

*k* = .......................................... units .......................................... **[4]**

ii. Calculate the rate of reaction when ICl has a concentration of $3.00 \times 10^{-3}$ mol dm$^{-3}$ and $H_2$ has a concentration of $2.00 \times 10^{-3}$ mol dm$^{-3}$.

Show **all** your working.

rate = .......................................... mol dm$^{-3}$ s$^{-1}$ **[1]**

**END OF QUESTION PAPER**