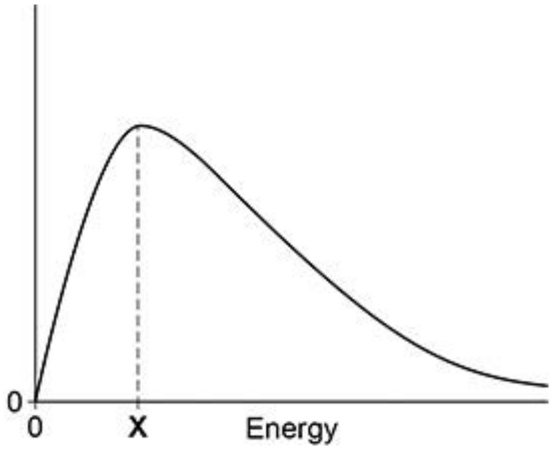
Q1.

The figure below shows the Maxwell–Boltzmann distribution of molecular energies in a sample of gas.



(a) Label the *y*-axis on the figure above.

(1)

(b) State why the curve starts at the origin.

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(1)

(c) State what **X** indicates on the figure above.

X indicates ___

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(1)

(d) Half of the gas molecules in the sample are removed.
The remaining gas molecules are kept at the same temperature.

Draw the new distribution of molecular energies for the remaining gas on the figure above.

(2)

(Total 5 marks)

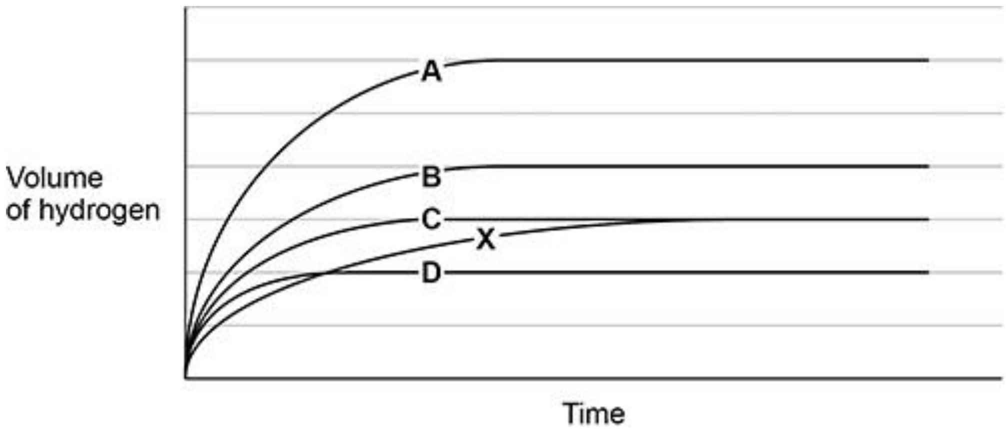
Q2.

An excess of magnesium reacts with hydrochloric acid to form hydrogen gas.

Line **X** on the graph shows how the volume of hydrogen produced changes with time as magnesium reacts with 30 cm^3 of 1.0 mol dm^{-3} hydrochloric acid.

The reaction is repeated using 20 cm^3 of 2.0 mol dm^{-3} hydrochloric acid, with all other conditions the same.

Which line shows how the volume of hydrogen produced changes with time?



A ⬭

B ⬭

C ⬭

D ⬭

(Total 1 mark)

Q3.

A mixture of 2 dm^3 of hydrogen and 1 dm^3 of oxygen is at room temperature.

Which statement is correct?

A There is no reaction to form water because the molecules do not collide with sufficient energy. ⬭

B There is no reaction to form water because the molecules do not collide with sufficient frequency. ⬭

C The mean velocity of the hydrogen molecules is less than that of the oxygen molecules. ⬭

D The partial pressure of each gas is the same. ⬭

(Total 1 mark)

Q4.

A student investigates the effect of temperature on the rate of reaction between sodium thiosulfate solution and dilute hydrochloric acid.

$$Na_2S_2O_3(aq) + 2\ HCl(aq) \rightarrow 2\ NaCl(aq) + SO_2(g) + S(s) + H_2O(l)$$

The student mixes the solutions together in a flask and places the flask on a piece of paper marked with a cross.

The student records the time for the cross to disappear. The cross disappears because the mixture becomes cloudy.

The table shows the student's results.

Temperature / °C	22	31	36	42	49	54
Time, t, for cross to disappear / s	87	48	36	26	44	12
$\frac{1}{t}$ / s^{-1}	0.0115	0.0208	0.0278	0.0385	0.0227	

(a) The student uses a stopwatch to measure the time. The stopwatch shows each time to the nearest 0.01 s

Suggest why the student records the times to the nearest second and not to the nearest 0.01 s

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(1)

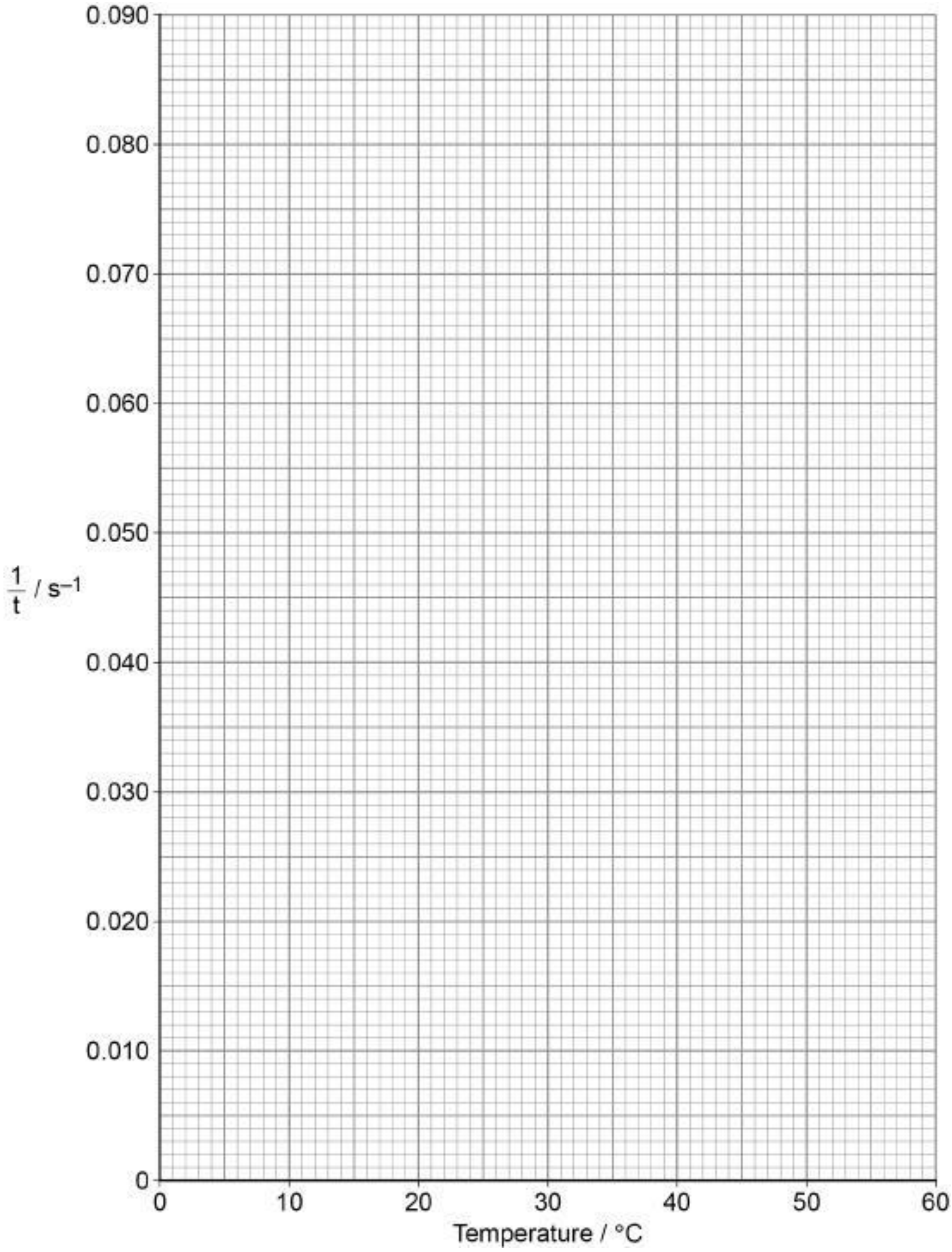
(b) The rate of reaction is proportional to $\frac{1}{t}$

Complete the table above.

(1)

(c) Plot the values of $\frac{1}{t}$ against temperature on the graph below.

Draw a line of best fit.



(2)

(d) Use your line of best fit to estimate the time for the cross to disappear at 40 °C
Show your working.

Time _______________ s

(1)

(e) Suggest, by considering the products of this reaction, why small amounts of reactants are used in this experiment.

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(1)

(f) The student could do the experiment at lower temperatures using an ice bath.

Suggest why the student chose **not** to carry out experiments at temperatures in the range 1–10 °C

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(1)

(Total 7 marks)

Q5.

Which statement about the molecules in a sample of a gas is correct?

A At a given temperature they all move at the same speed. ○

B At a given temperature their average kinetic energy is constant. ○

C As temperature increases, there are more molecules with the most probable energy. ○

D As temperature decreases, there are fewer molecules with the mean energy. ○

(Total 1 mark)

Q6.

Which statement is correct for the distribution curve of molecular energies in a gas?

A	The curve is symmetrical about the maximum.	⬭
B	There are always some molecules with zero energy.	⬭
C	The position of the maximum of the curve is not dependent on the temperature.	⬭
D	The mean energy of the molecules is greater than the most probable energy of the molecules.	⬭

(Total 1 mark)

Q7.

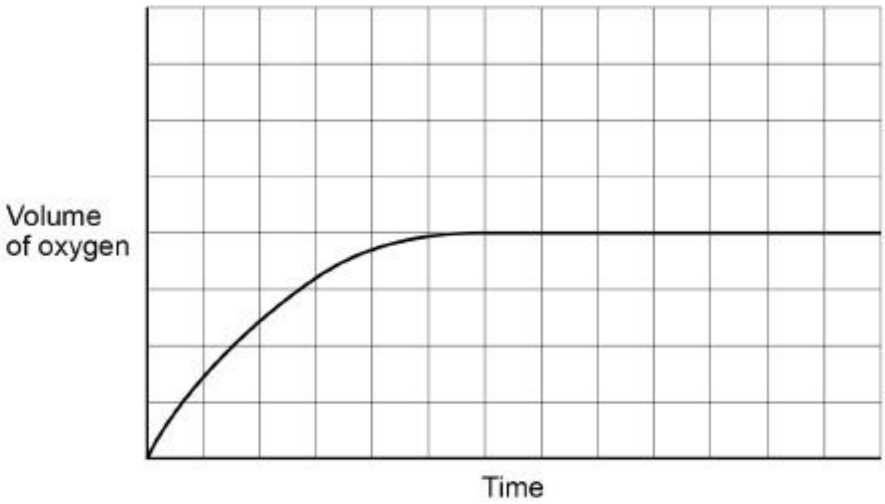
Hydrogen peroxide solution decomposes slowly to form water and oxygen.
The reaction is much faster in the presence of a manganese(IV) oxide catalyst.

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

Three experiments, shown in the table, were carried out to investigate how the volume of oxygen produced varied over time under different conditions. The same mass of catalyst was used in each experiment.

Experiment	Concentration of H_2O_2(aq) / mol dm^{-3}	Volume of H_2O_2(aq) / cm^3	Temperature / °C	Catalyst
1	1.0	50	20	lumps
2	1.0	50	20	powder
3	0.5	50	20	lumps

The graph shows how the volume of oxygen collected varied with time in Experiment **1**.



(a) Explain, in general terms, how a catalyst increases the rate of a reaction.

(2)

(b) Draw **two** lines on the graph to show how the volume of oxygen collected varied with time in Experiments **2** and **3**.
Label each line with the experiment number.

(2)

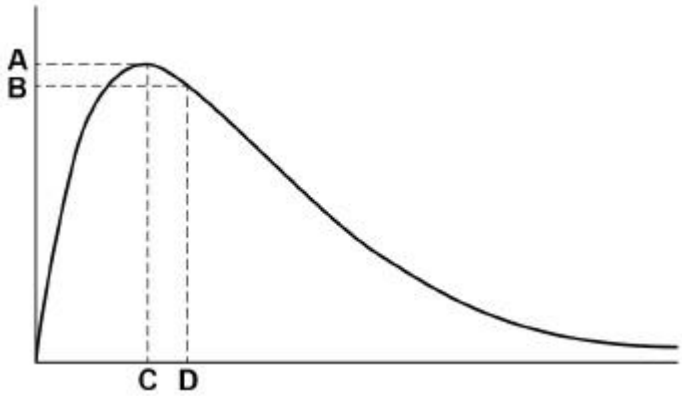
(c) Explain, in terms of collision theory, the effect of increasing the concentration of hydrogen peroxide on the rate of reaction.

(2)
(Total 6 marks)

Q8.

The Maxwell–Boltzmann distribution of molecular energies in a sample of gas at a fixed temperature is shown.

Which letter represents the mean energy of the molecules?



A ⬭

B ⬭

C ⬭

D ⬭

(Total 1 mark)

Q9.

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

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

The rate equation for this reaction can be written as

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

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

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

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

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

(2)

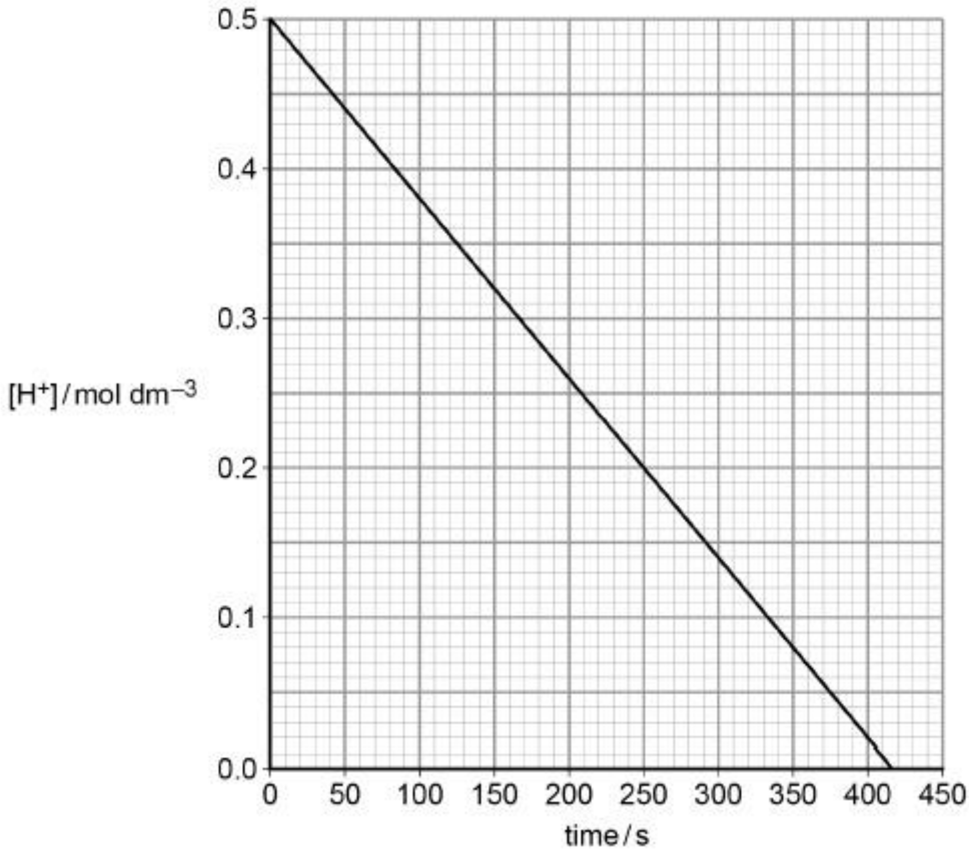
(b) Samples of the reaction mixture are removed at timed intervals and titrated with alkali to determine the concentration of H^+(aq).

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

(2)

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

Figure 1



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

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(2)

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

Give the units of k_1

k_1 ______________________

Units ______________________

(3)

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

There is a large excess of H_2O_2

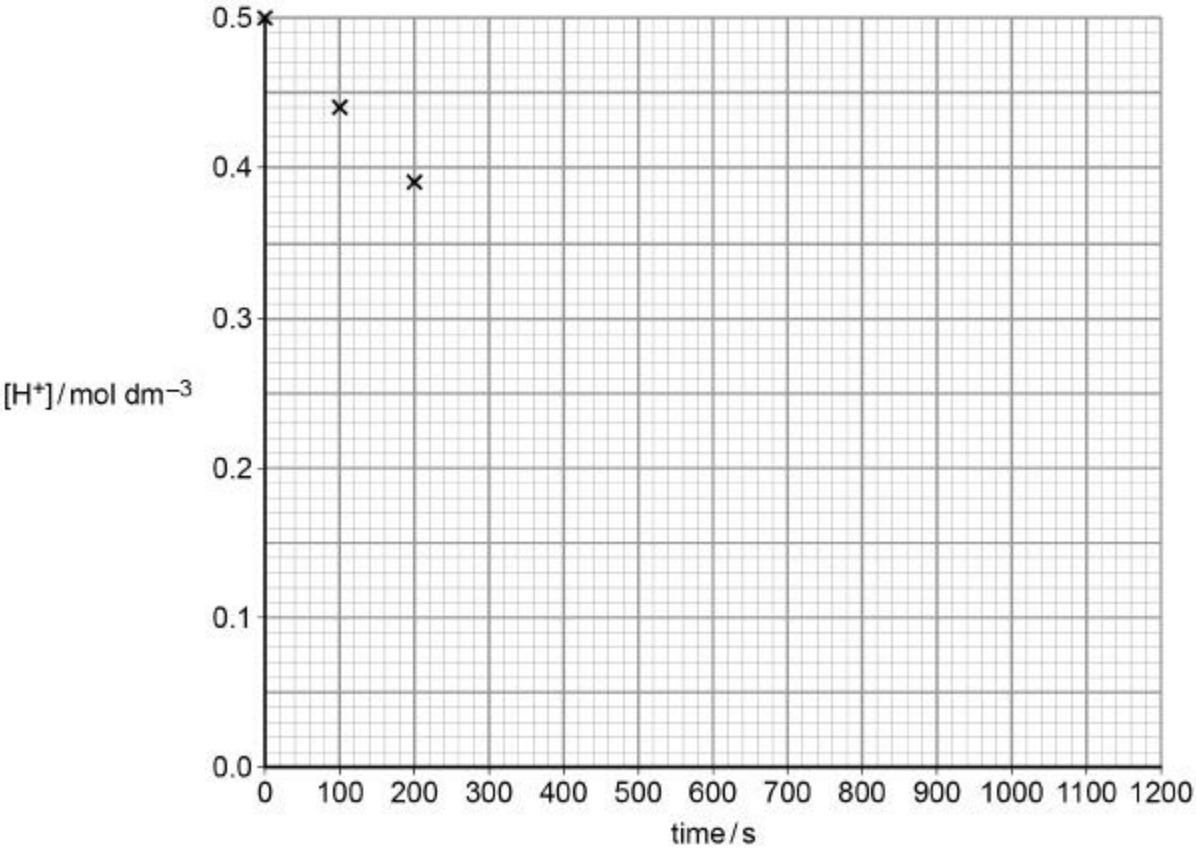
In this reaction mixture, the rate depends only on the concentration of I^-(aq).

The results are shown in the table.

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

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

Figure 2



(1)

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

(1)

(g) Calculate the rate of reaction when $[H^+] = 0.35$ mol dm^{-3}
Show your working using a suitable construction on the graph in **Figure 2**.

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

(2)

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

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

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

A reagent (**X**) is available that reacts rapidly with **E**. This means that, if a small amount of **X** is included in the initial reaction mixture, it will react with any **E** produced until all of the **X** has been used up.

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

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(6)
(Total 19 marks)

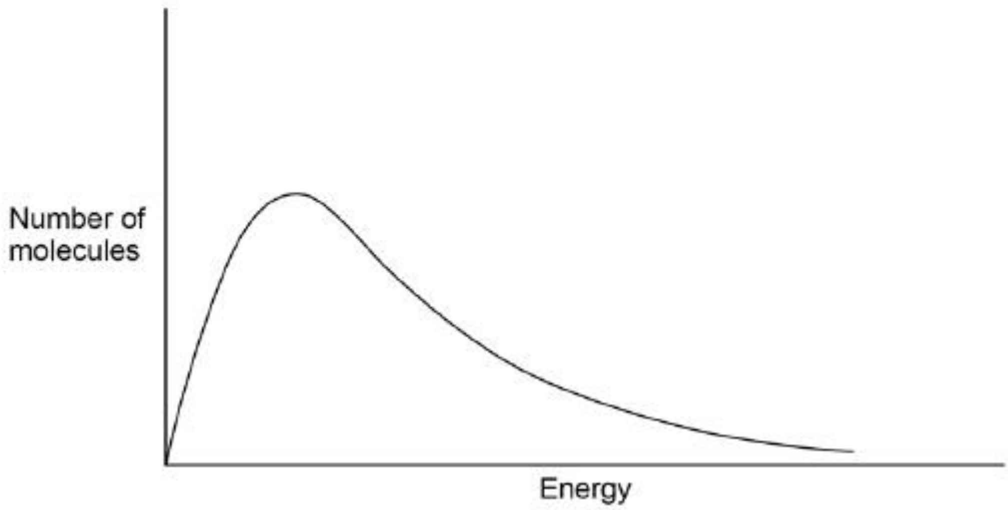
Q10.

Nitryl chloride reacts with nitrogen monoxide according to the equation:

$$ClNO_2(g) + NO(g) \longrightarrow NO_2(g) + ClNO(g)$$

The Maxwell–Boltzmann distribution curve in **Figure 1** shows the distribution of molecular energies in 1 mol of this gaseous reaction mixture (sample **1**) at 320 K.

Figure 1



(a) On the same axes, draw a curve for sample **1** at a lower temperature.

(2)

(b) Explain the effect that lowering the temperature would have on the rate of reaction.

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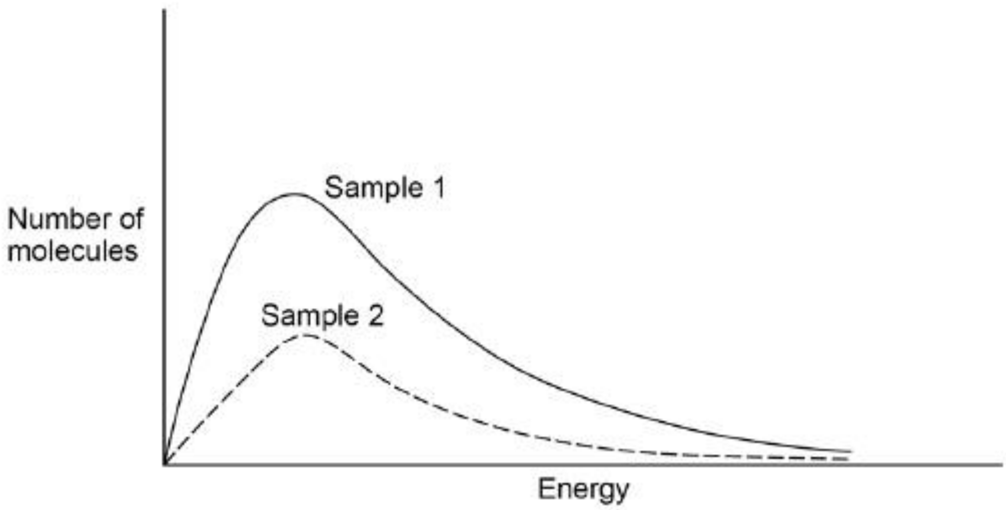
(2)

(c) A Maxwell–Boltzmann distribution curve was drawn for a second sample of the reaction mixture in the same reaction vessel. **Figure 2** shows the results.

Figure 2



Deduce the change that was made to the reaction conditions.

Explain the effect that this change has on the rate of reaction.

Change

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Explanation

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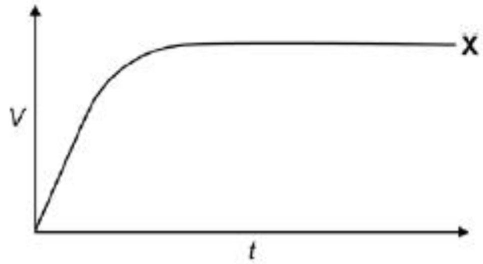
(3)
(Total 7 marks)

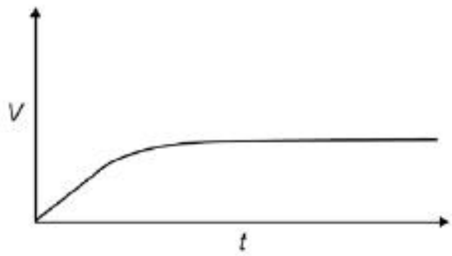
Q11.

Line **X** in the diagram represents the volume (V) of gas formed with time (t) in a reaction between an excess of magnesium and aqueous sulfuric acid.

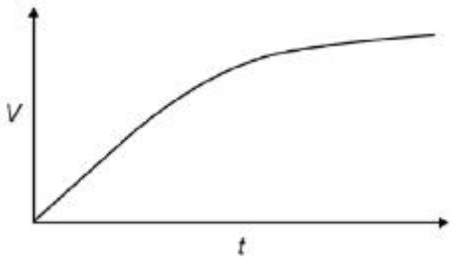


Which line represents the volume of hydrogen formed, at the same temperature and pressure, when the concentration of sulfuric acid has been halved?

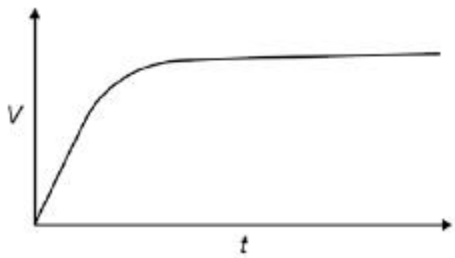
A

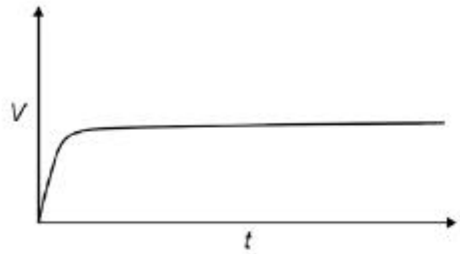


B



C

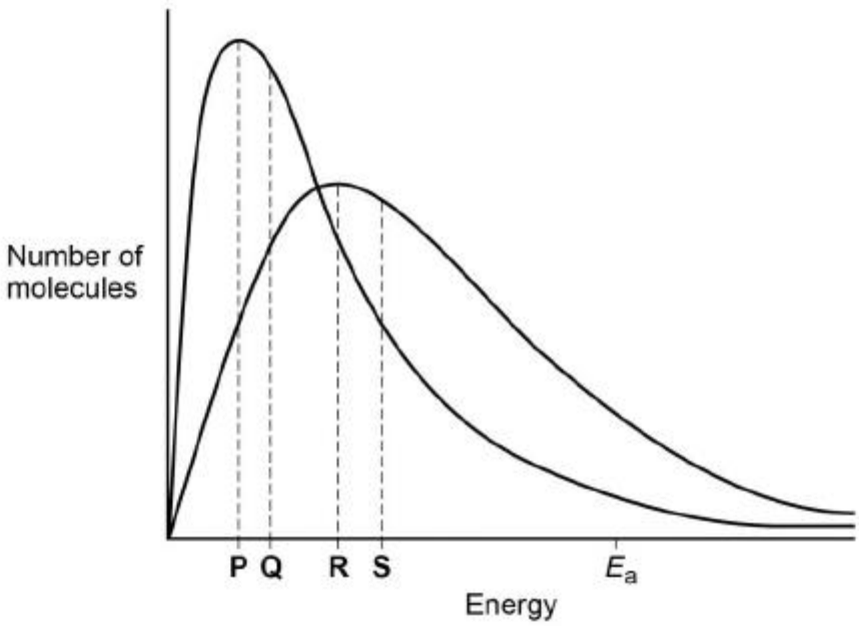


D 

(Total 1 mark)

Q12.

The question below is about the Maxwell–Boltzmann distribution shown for a sample of a gas, X, at two different temperatures.



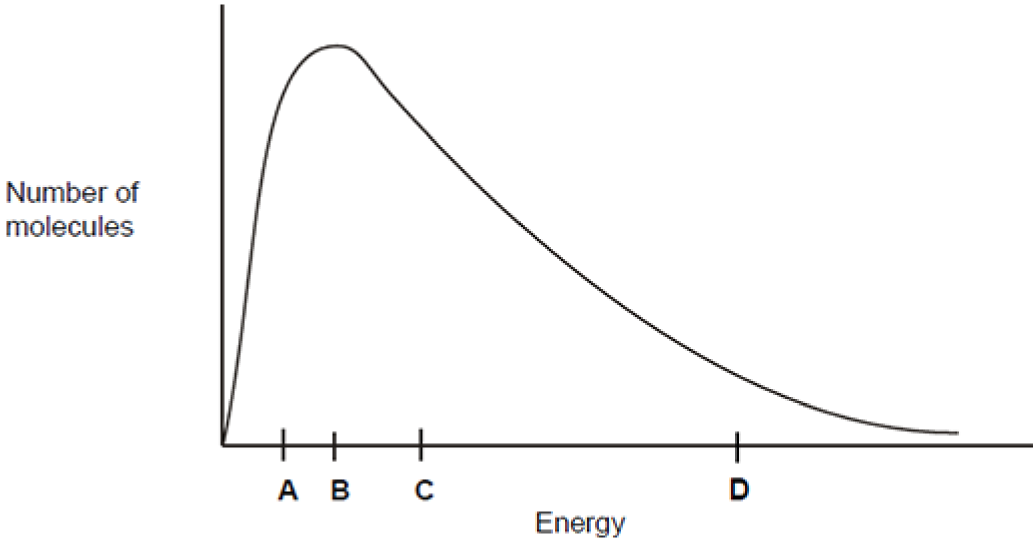
Which statement is correct for the higher temperature?

A	The area under the curve to the left of E_a decreases.	⬭
B	The total area under the curve increases.	⬭
C	The activation energy decreases.	⬭
D	More molecules have the mean energy.	⬭

(Total 1 mark)

Q13.

This question is about the Maxwell–Boltzmann distribution of molecular energies in a sample of a gas shown in the figure below.



Which letter best represents the mean energy of the molecules?

A ⬭

B ⬭

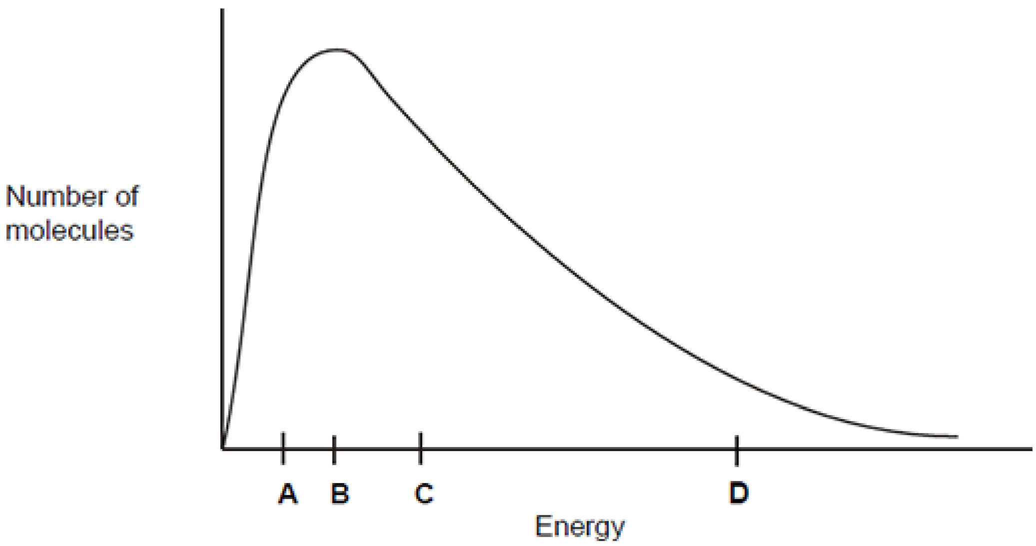
C ⬭

D ⬭

(Total 1 mark)

Q14.

This question is about the Maxwell–Boltzmann distribution of molecular energies in a sample of a gas shown in the following figure.



What does the area under the curve represent?

A The total energy of the particles. ☐

B The total number of particles. ☐

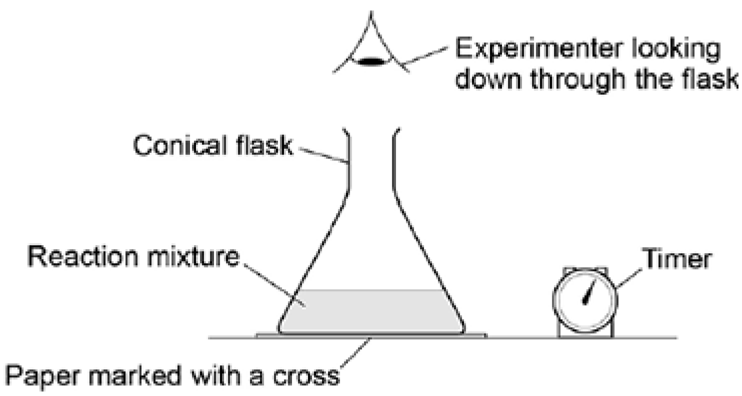
C The number of particles that can react with each other. ☐

D The total number of particles that have activation energy. ☐

(Total 1 mark)

Q15.

The apparatus in the figure below was set up to measure the time taken for 20.0 cm^3 of sodium thiosulfate solution to react with 5.0 cm^3 of hydrochloric acid in a 100 cm^3 conical flask at 20 °C. The timer was started when the sodium thiosulfate solution was added to the acid in the flask. The timer was stopped when it was no longer possible to see the cross on the paper.



What is likely to decrease the accuracy of the experiment?

A Rinsing the flask with acid before each new experiment.

B Stirring the solution throughout each experiment.

C Using the same piece of paper for each experiment.

D Using different measuring cylinders to measure the volumes of acid and sodium thiosulfate.

(Total 1 mark)

Q16.

The experiment was repeated at 20 °C using a 250 cm^3 conical flask.

Which statement is correct about the time taken for the cross to disappear when using the larger conical flask?

A The time taken will **not** be affected by using the larger conical flask.

B The time taken will be decreased by using the larger conical flask.

C The time taken will be increased by using the larger conical flask.

D It is impossible to predict how the time taken will be affected by using the larger conical flask.

(Total 1 mark)

Q17.

This question involves the use of kinetic data to deduce the order of a reaction and calculate a value for a rate constant.

The data in **Table 1** were obtained in a series of experiments on the rate of the reaction between compounds **A** and **B** at a constant temperature.

Table 1

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

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

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(3)

The data in **Table 2** were obtained in two experiments on the rate of the reaction between compounds **C** and **D** at a constant temperature.

Table 2

Experiment	**Initial concentration of C / mol dm^{-3}**	**Initial concentration of D / mol dm^{-3}**	**Initial rate / mol dm^{-3} s^{-1}**
4	1.9×10^{-2}	3.5×10^{-2}	7.2×10^{-4}
5	3.6×10^{-2}	5.4×10^{-2}	To be calculated

The rate equation for this reaction is

$$rate = k[\mathbf{C}]^2[\mathbf{D}]$$

(b) Use the data from experiment **4** to calculate a value for the rate constant, *k*, at this temperature. Deduce the units of *k*.

k = _______________ Units = _______________

(3)

(c) Calculate a value for the initial rate in experiment **5**.

Initial rate = _______________ mol dm^{-3} s^{-1}

(1)

(d) The rate equation for a reaction is

$$rate = k[\mathbf{E}]$$

Explain qualitatively why doubling the temperature has a much greater effect on the rate of the reaction than doubling the concentration of **E**.

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(3)

(e) A slow reaction has a rate constant $k = 6.51 \times 10^{-3}$ mol^{-1} dm^{3} at 300 K.

Use the equation $\ln k = \ln A - E_a / RT$ to calculate a value, in kJ mol^{-1}, for the activation energy of this reaction.

The constant $A = 2.57 \times 10^{10}$ mol^{-1} dm^{3}.
The gas constant $R = 8.31$ J K^{-1} mol^{-1}.

Activation energy = ________________

(2)