Percentage	
Grade	

TUTOR ZONE

YOUR STEP TO SUCCESS

A Level Chemistry

Rate Equations

Duration: 50 min

Total Marks:

Information for Candidates:

- Use black or blue ink. HB pencil may be used for graphs and diagrams only.
- Answer **all** the questions.
- Read each question carefully. Make sure you know what you have to do before starting your answer.
- Write your answer to each question in the space provided. If additional paper is used, the question number(s) must be clearly shown
- The number of marks is given in brackets **[ ]** at the end of each question or part question.
- You may use an electronic calculator.
- You are advised to show all the steps in any calculations.

Do not write in this table	
Question	Mark
TOTAL	

2 In the presence of acid, $H^+(aq)$, aqueous bromate(V) ions, $BrO_3^-(aq)$, react with aqueous bromide ions, $Br^-(aq)$, to produce bromine, $Br_2(aq)$.

A student carried out an investigation into the kinetics of this reaction.

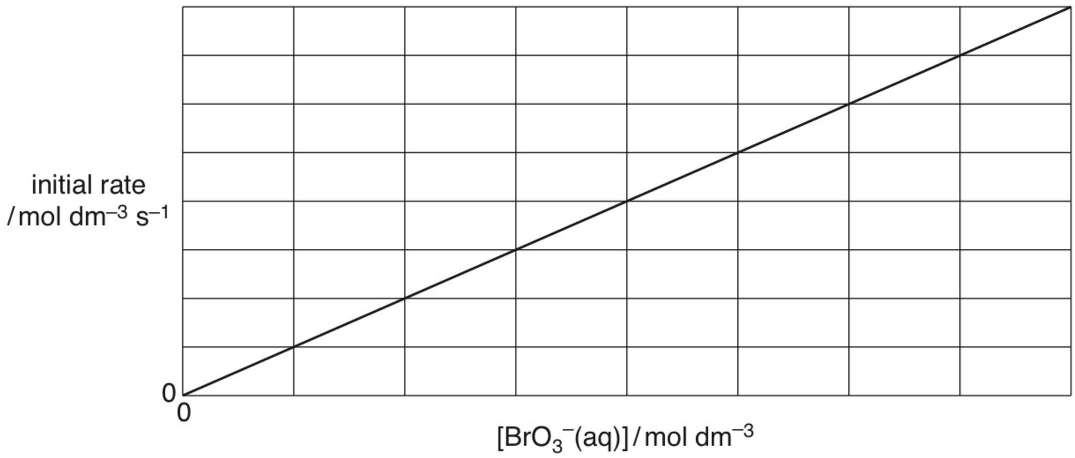
(a) Balance the ionic equation for this reaction.

$$BrO_3^- + \text{.........}Br^- + \text{.........}H^+ \longrightarrow \text{.........}Br_2 + \text{.........}H_2O$$

[1]

(b) The student investigated how different concentrations of $BrO_3^-(aq)$ affect the initial rate of the reaction.

A graph of initial rate against $[BrO_3^-(aq)]$ is shown below.



The student then investigated how different concentrations of $Br^-(aq)$ and $H^+(aq)$ affect the initial rate of the reaction.

The results are shown below.

$[BrO_3^-(aq)]$ / $mol\,dm^{-3}$	$[Br^-(aq)]$ / $mol\,dm^{-3}$	$[H^+(aq)]$ / $mol\,dm^{-3}$	initial rate / $mol\,dm^{-3}\,s^{-1}$
5.0×10^{-2}	1.5×10^{-1}	3.1×10^{-1}	1.19×10^{-5}
5.0×10^{-2}	3.0×10^{-1}	3.1×10^{-1}	2.38×10^{-5}
5.0×10^{-2}	1.5×10^{-1}	6.2×10^{-1}	4.76×10^{-5}

- Using the results from the student's experiments, what conclusions can be drawn about the kinetics of this reaction? Justify your reasoning.

- Calculate the rate constant for this reaction, including the units.

In your answer you should make clear how your conclusions fit with the experimental results.

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.... **[9]**

[Total: 10]

Answer **all** the questions.

1 Hydrogen, H_2, reacts with nitrogen monoxide, NO, as shown in the equation below.

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

A chemist carries out a series of experiments and determines the rate equation for this reaction:

$$\text{rate} = k[H_2(g)][NO(g)]^2$$

(a) In one of the experiments, the chemist reacts together:

- $1.2 \times 10^{-2}\,\text{mol dm}^{-3}\,H_2(g)$
- $6.0 \times 10^{-3}\,\text{mol dm}^{-3}\,NO(g)$

The initial rate of this reaction is $3.6 \times 10^{-2}\,\text{mol dm}^{-3}\,\text{s}^{-1}$.

Calculate the rate constant, k, for this reaction. State the units, if any.

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

(b) Predict what would happen to the initial rate of reaction for the following changes in concentrations.

(i) The concentration of $H_2(g)$ is doubled.

.... **[1]**

(ii) The concentration of NO(g) is halved.

.... **[1]**

(iii) The concentrations of $H_2(g)$ and NO(g) are **both** increased by four times.

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.... **[1]**

(c) The chemist carries out the reaction between hydrogen and nitrogen monoxide at a higher pressure.

(i) Explain, with a reason, what happens to the initial rate of reaction.

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(ii) State what happens to the rate constant.

.... **[1]**

(d) This overall reaction between hydrogen and nitrogen monoxide takes place by a two-step mechanism. The first step is much slower than the second step.

Suggest a possible two-step mechanism for the overall reaction.

step 1: ....

step 2: .... **[2]**

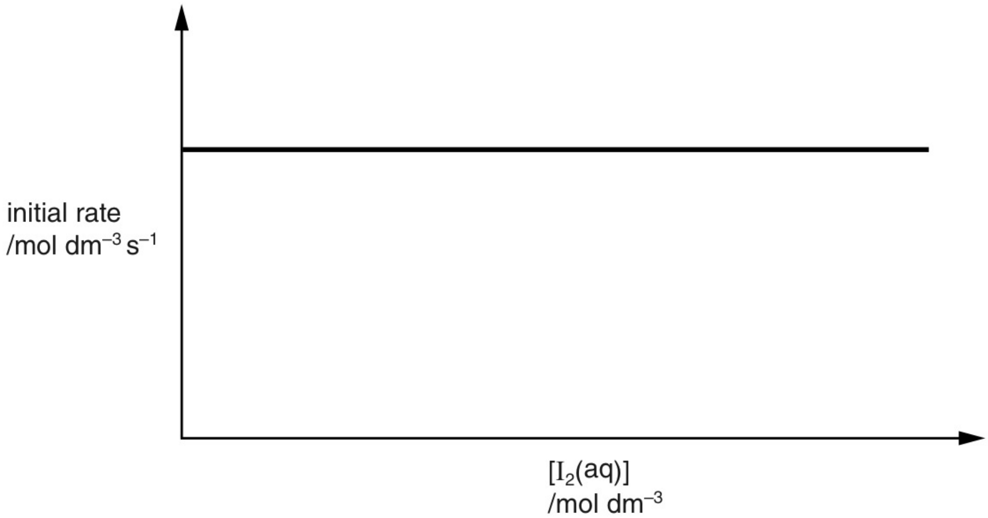
[Total: 10]

Answer **all** the questions.

1 A student investigates the reaction between iodine, I_2, and propanone, $(CH_3)_2CO$, in the presence of aqueous hydrochloric acid, $HCl(aq)$.

The results of the investigation are shown below.

Rate–concentration graph



Results of initial rates experiments

experiment	$[(CH_3)_2CO(aq)]$ / mol dm^{-3}	$[HCl(aq)]$ / mol dm^{-3}	initial rate / mol dm^{-3} s^{-1}
1	1.50×10^{-3}	2.00×10^{-2}	2.10×10^{-9}
2	3.00×10^{-3}	2.00×10^{-2}	4.20×10^{-9}
3	3.00×10^{-3}	5.00×10^{-2}	1.05×10^{-8}

(a) Determine the orders with respect to I_2, $(CH_3)_2CO$ and HCl, the rate equation and the rate constant for the reaction.

Explain all of your reasoning.

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(b) The student then investigates the reaction of hydrogen, H_2, and iodine monochloride, ICl.

The equation for this reaction is shown below.

$$H_2(g) + 2ICl(g) \rightarrow 2HCl(g) + I_2(g)$$

The rate equation for this reaction is shown below.

$$\text{rate} = k[H_2(g)]\,[ICl(g)]$$

Predict a possible two-step mechanism for this reaction. The first step should be the rate-determining step.

step 1 ....

step 2 .... [2]

[Total: 11]

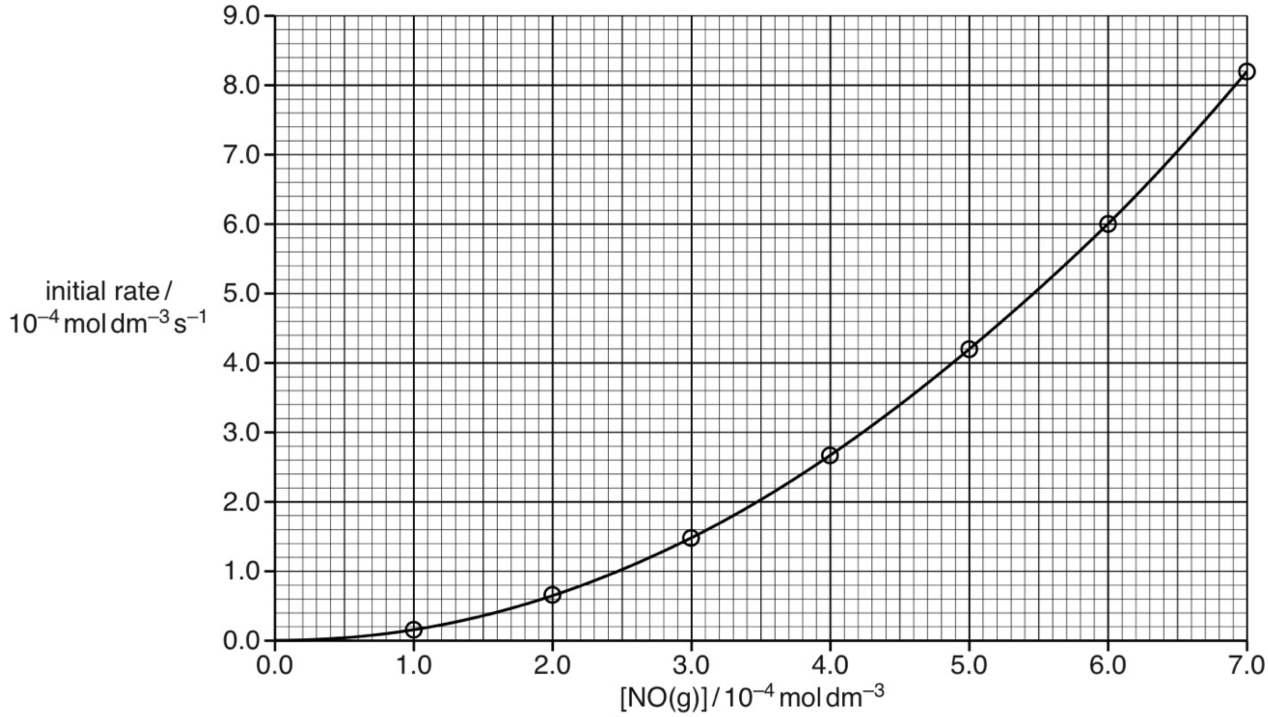
2 Hydrogen, H_2, reacts with nitrogen monoxide, NO, as shown below:

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

(a) 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–concentration graph is plotted.



The chemist uses $H_2(g)$ of concentration 2.0×10^{-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]**

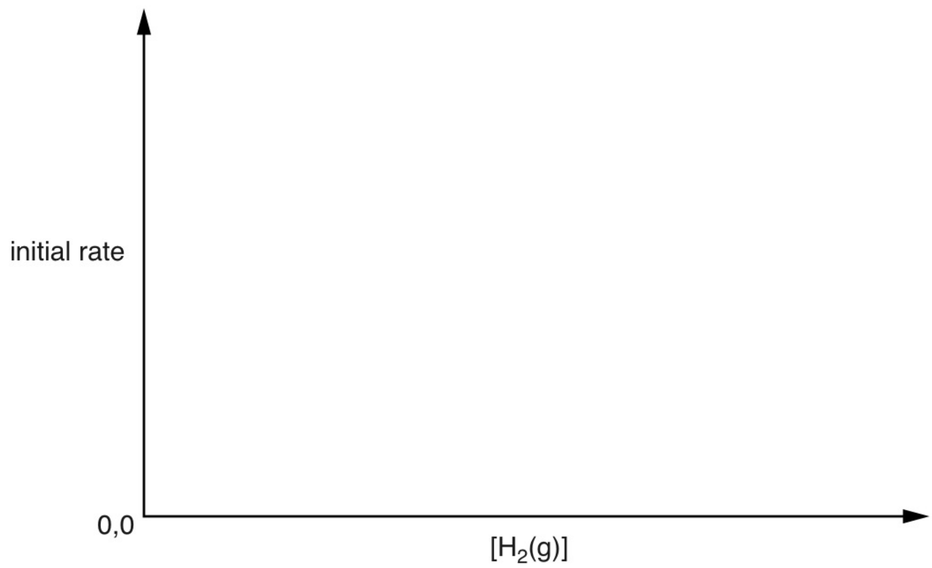
(b) A chemist investigates the effect of changing the concentration of $H_2(g)$ on the initial reaction rate at two different temperatures.

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.



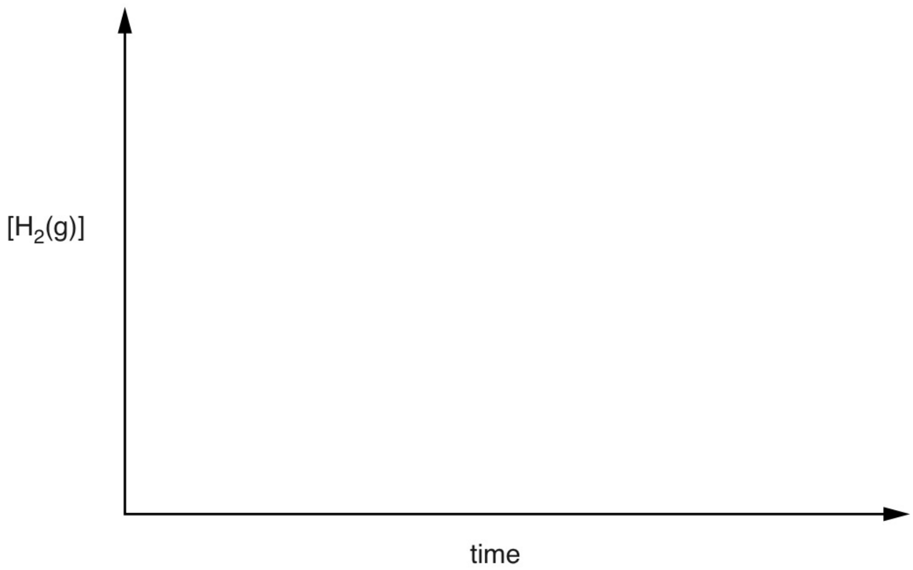
[2]

(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]**

[Total: 11]