Name

IB 6 HL EQ 16w to 99s P2A&B 447 marks

Q# 1/ IB Chem/2016/w/TZ0/Paper 2 Section A/Higher Level/

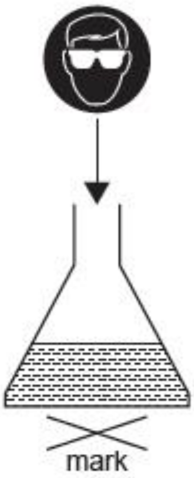
3. Sodium thiosulfate solution reacts with dilute hydrochloric acid to form a precipitate of sulfur at room temperature.

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

(a) Identify the formula and state symbol of X. [1]

(b) Suggest why the experiment should be carried out in a fume hood or in a well-ventilated laboratory. [1]

(c) The precipitate of sulfur makes the mixture cloudy, so a mark underneath the reaction mixture becomes invisible with time.



$10.0\,cm^3$ of $2.00\,mol\,dm^{-3}$ hydrochloric acid was added to a $50.0\,cm^3$ solution of sodium thiosulfate at temperature, T_1. Students measured the time taken for the mark to be no longer visible to the naked eye. The experiment was repeated at different concentrations of sodium thiosulfate.

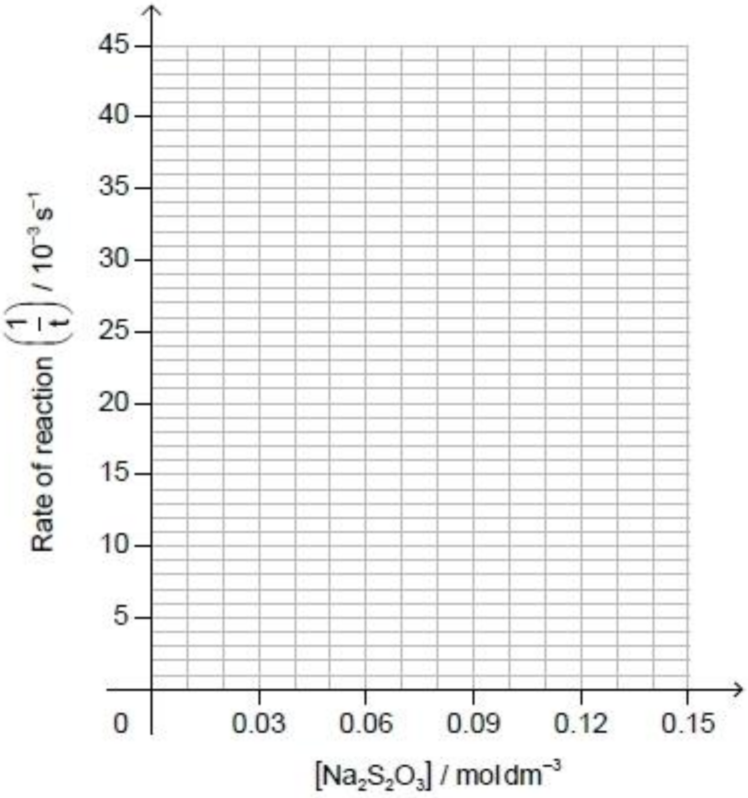
Experiment	**[$Na_2S_2O_3$(aq)] / mol dm^{-3}**	**Time, t, for mark to disappear / s ± 1 s**	**$\frac{1^*}{t}$ / $10^{-3}\,s^{-1}$**
1	0.150	23	43.5
2	0.120	27	37.0
3	0.090	36	27.8
4	0.060	60	16.7
5	0.030	111	9.0

* The reciprocal of the time in seconds can be used as a measure of the rate of reaction.

[Source: Adapted from http://www.flinnsci.com/]

Show that the hydrochloric acid added to the flask in experiment 1 is in excess. [2]

(d) Draw the best fit line of $\frac{1}{t}$ against concentration of sodium thiosulfate on the axes provided. [2]



(e) (i) Using the graph, explain the order of reaction with respect to sodium thiosulfate. [2]

(ii) In a different experiment, this reaction was found to be first order with respect to hydrochloric acid. Deduce the overall rate expression for the reaction. [1]

(f) A student decided to carry out another experiment using 0.075 mol dm^{-3} solution of sodium thiosulfate under the same conditions. Determine the time taken for the mark to be no longer visible. [2]

(g) An additional experiment was carried out at a higher temperature, T_2.

(i) On the same axes, sketch Maxwell–Boltzmann energy distribution curves at the two temperatures T_1 and T_2, where $T_2 > T_1$. [2]

(ii) Explain why a higher temperature causes the rate of reaction to increase. [2]

(h) Suggest one reason why the values of rates of reactions obtained at higher temperatures may be less accurate. [1]

Q# 2/ IB Chem/2016/s/TZoSP/Paper 2 Section A/Higher Level/Q4

(e) Describe what you understand by the term *frequency (pre-exponential) factor, A*. *[1]*

(f) The activation energy, E_a, for the reaction of 1-iodoethane with sodium hydroxide is $87.0\,\text{kJ}\,\text{mol}^{-1}$, and the frequency (pre-exponential) factor, A, is $2.10 \times 10^{11}\,\text{mol}^{-1}\,\text{dm}^3\,\text{s}^{-1}$.

Calculate the rate constant, k, of the reaction at 25 °C, indicating the units of k, and giving a reason for your choice. *[2]*

Q# 3/ IB Chem/2016/s/TZ1/Paper 2 Section A/Higher Level/

3. The reaction between hydrogen and nitrogen monoxide is thought to proceed by the mechanism shown below.

$2NO(g) \rightleftharpoons N_2O_2(g)$	fast equilibrium
$N_2O_2(g) + H_2(g) \rightarrow N_2O(g) + H_2O(g)$	slow reaction
$N_2O(g) + H_2(g) \rightarrow N_2(g) + H_2O(g)$	fast reaction

(a) (i) State the equation for the overall reaction. [1]

(ii) Deduce the rate expression consistent with this mechanism. [1]

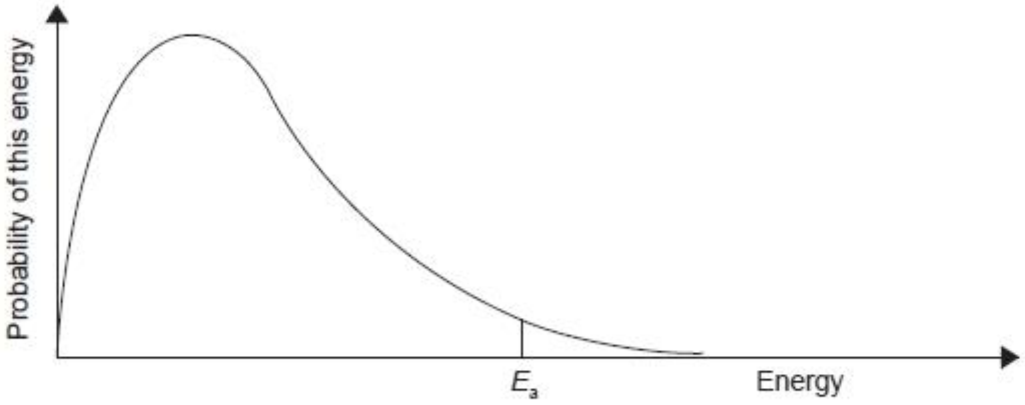
(iii) Explain how you would attempt to confirm this rate expression, giving the results you would expect. [3]

(iv) State, giving your reason, whether confirmation of the rate expression would prove that the mechanism given is correct. [1]

(v) Suggest how the rate of this reaction could be measured experimentally. [1]

(b)

(iii) Sketch and label a second Maxwell–Boltzmann energy distribution curve representing the same system but at a higher temperature, T_{higher}. [1]



(iv) Explain why an increase in temperature increases the rate of this reaction. [2]

Q# 4/ IB CHEM/2015/w/TZ0/Paper 2 Section B/Higher Level/Q7

(c) The oxidation of nitrogen monoxide takes place as follows:

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

The following experimental data was obtained at 101.3 kPa and 298 K.

Experiment	Initial [NO] / mol dm^{-3}	Initial [O_2] / mol dm^{-3}	Initial rate / mol dm^{-3} s^{-1}
1	2.30×10^{-2}	1.15×10^{-2}	1.05×10^{-3}
2	2.30×10^{-2}	2.30×10^{-2}	2.09×10^{-3}
3	4.60×10^{-2}	4.60×10^{-2}	1.68×10^{-2}

(i) Deduce the orders of reaction with respect to O_2 and NO. [2]

Order with respect to O_2:

Order with respect to NO:

(ii) State the rate expression for the reaction. [1]

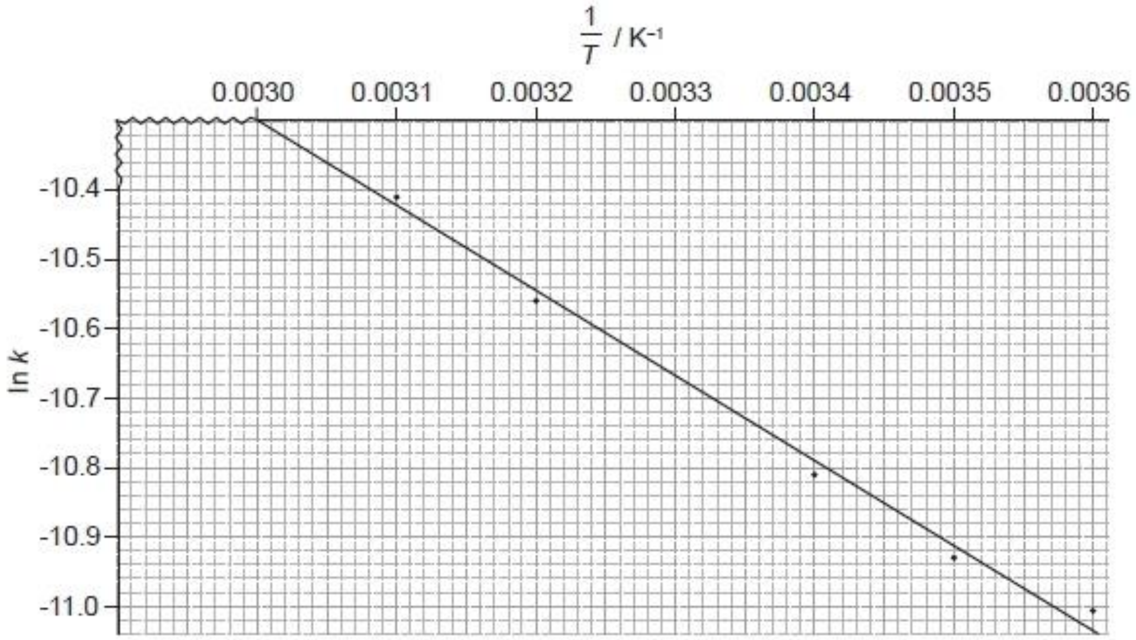
(iii) Calculate the value of the rate constant, k, and include its units. [2]

(iv) Suggest a mechanism that is consistent with the rate expression, indicating the rate-determining step. [3]

Q# 5/ IB Chem/2015/w/TZ0/Paper 2 Section A/Higher Level/Q1

(f) The student found by further experimentation that oxidation of ascorbic acid follows first-order kinetics. The graph of ln *k* against $\frac{1}{T}$ is shown below.

Determine the activation energy to **three** significant figures, including units. [3]



Q# 6/ IB Chem/2015/s/TZ2/Paper 2 Section A/Higher Level/

2. (a) Define the term *rate of reaction*. [1]

(b) Explain why increasing the particle size of a solid reactant decreases the rate of reaction. [2]

(c) Nitrogen(II) oxide reacts with hydrogen according to the equation below.

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

A suggested mechanism for this reaction is:

Step 1:	$NO + H_2 \rightleftharpoons X$	fast
Step 2:	$X + NO \rightarrow Y + H_2O$	slow
Step 3:	$Y + H_2 \rightarrow N_2 + H_2O$	fast

(i) Identify the rate-determining step. [1]

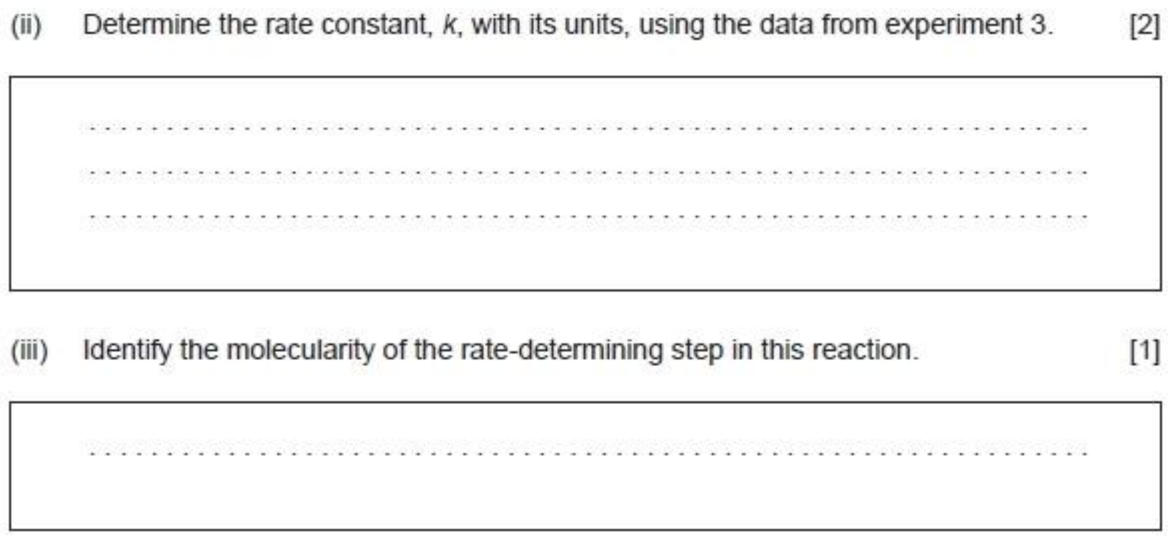
(ii) A student hypothesized that the order of reaction with respect to H_2 is 2. Evaluate this hypothesis. [2]

Q# 7/ IB CHEM/2015/s/TZ1/Paper 2 Section B/Higher Level/Q7

(c) (i) Determine the orders of reaction of the reactants and the overall rate expression for the reaction between 2-bromobutane and aqueous sodium hydroxide using the data in the table.

Experiment	[NaOH] / mol dm^{-3}	[C_4H_9Br] / mol dm^{-3}	Rate / mol dm^{-3} s^{-1}
1	1.00	1.00	1.66×10^{-3}
2	0.50	1.00	8.31×10^{-4}
3	0.25	0.25	1.02×10^{-4}
4	1.00	0.50	8.29×10^{-4}

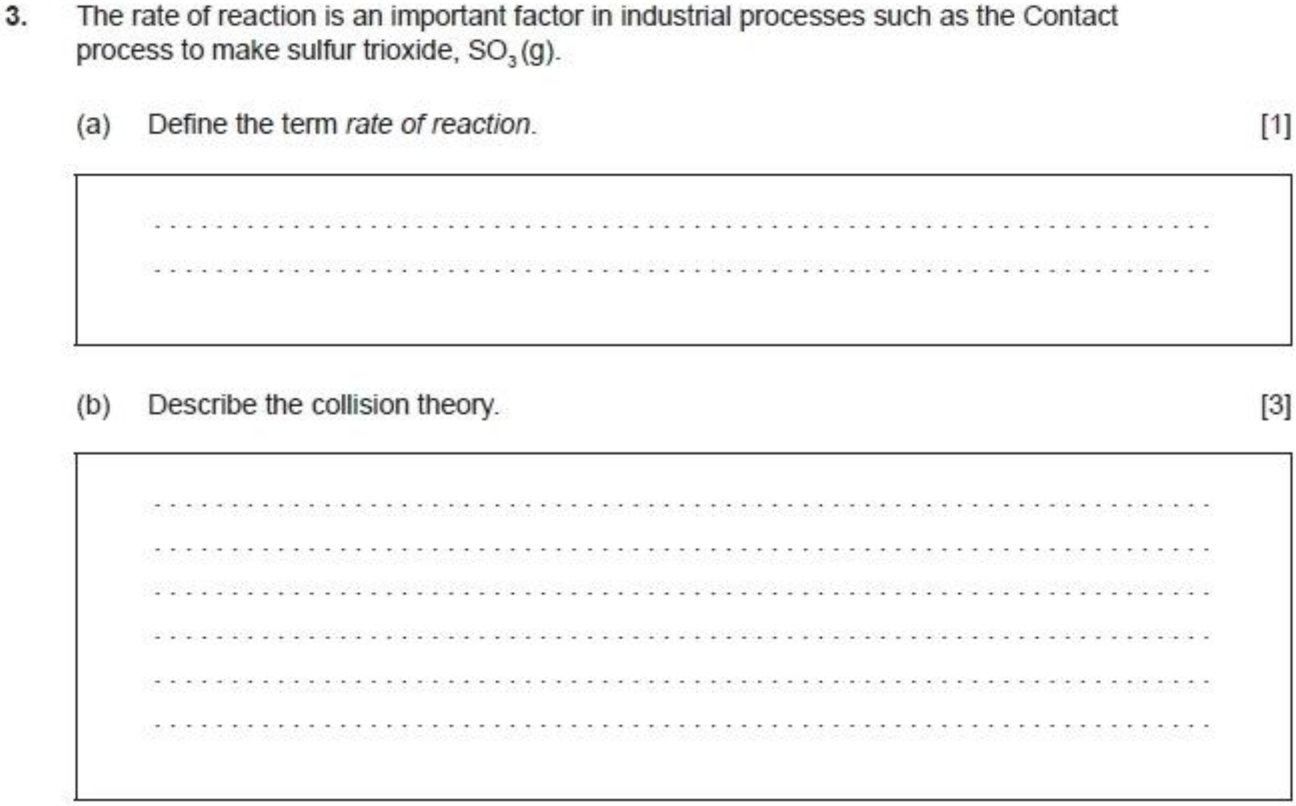
[2]



(ii) Determine the rate constant, k, with its units, using the data from experiment 3. [2]

(iii) Identify the molecularity of the rate-determining step in this reaction. [1]

Q# 8/ IB Chem/2015/s/TZ1/Paper 2 Section A/Higher Level/



3. The rate of reaction is an important factor in industrial processes such as the Contact process to make sulfur trioxide, $SO_3(g)$.

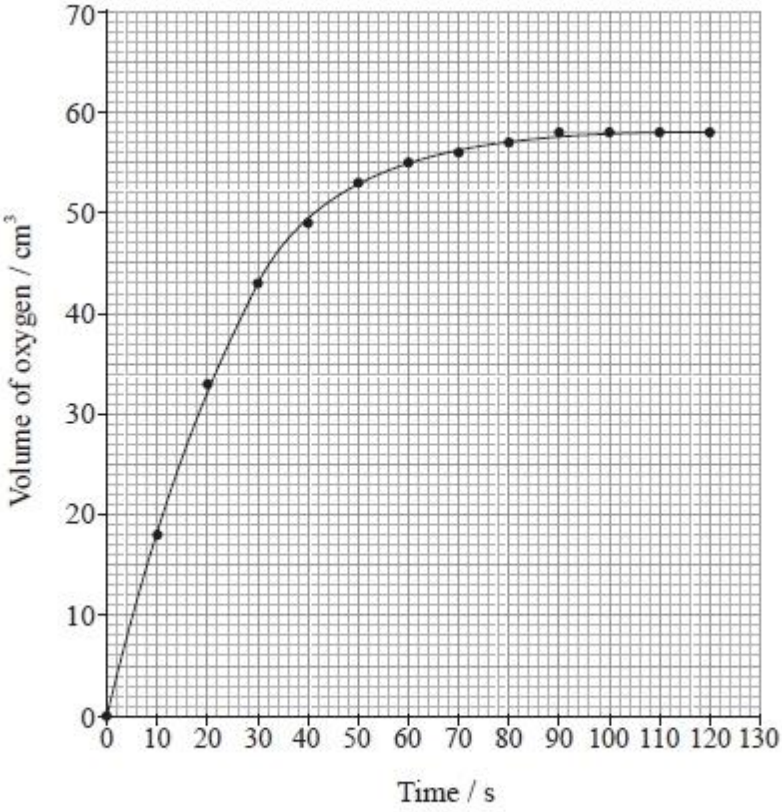
(a) Define the term *rate of reaction*. [1]

(b) Describe the collision theory. [3]

11. Hydrogen peroxide decomposes according to the equation below.

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

The rate of the decomposition can be monitored by measuring the volume of oxygen gas released. The graph shows the results obtained when a solution of hydrogen peroxide decomposed in the presence of a CuO catalyst.



(a) (i) Outline how the initial rate of reaction can be found from the graph. *[2]*

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(ii) Explain how and why the rate of reaction changes with time. *[3]*

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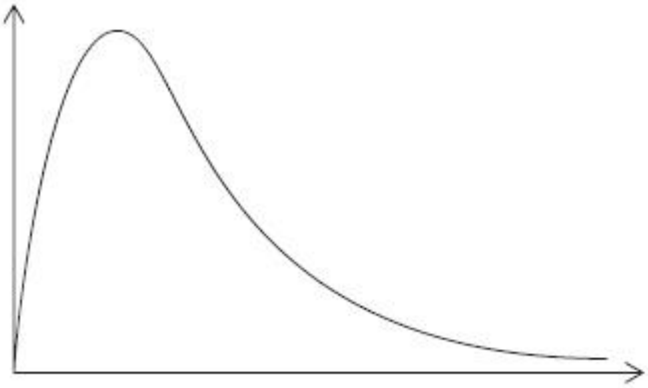
(b) A Maxwell–Boltzmann energy distribution curve is drawn below. Label both axes and explain, by annotating the graph, how catalysts increase the rate of reaction. *[3]*



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(c) (i) In some reactions, increasing the concentration of a reactant does not increase the rate of reaction. Describe how this may occur. *[1]*

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(ii) Consider the reaction

$$2\,A + B \rightarrow C + D$$

The reaction is first order with respect to **A**, and zero order with respect to **B**. Deduce the rate expression for this reaction. *[1]*

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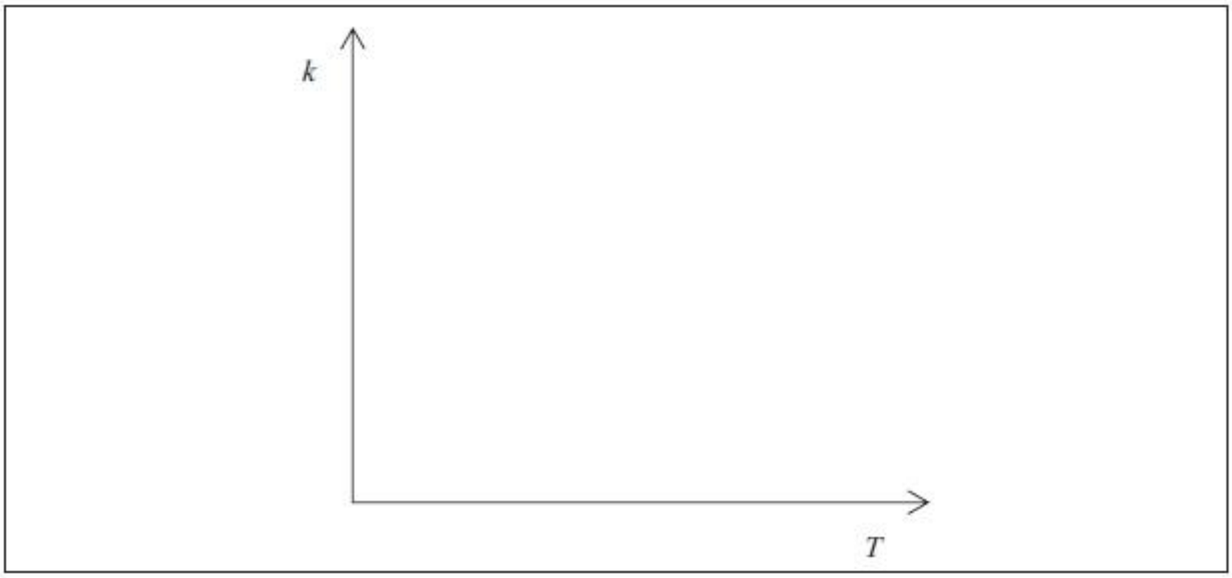
(d) Sketch a graph of rate constant (k) versus temperature. *[1]*

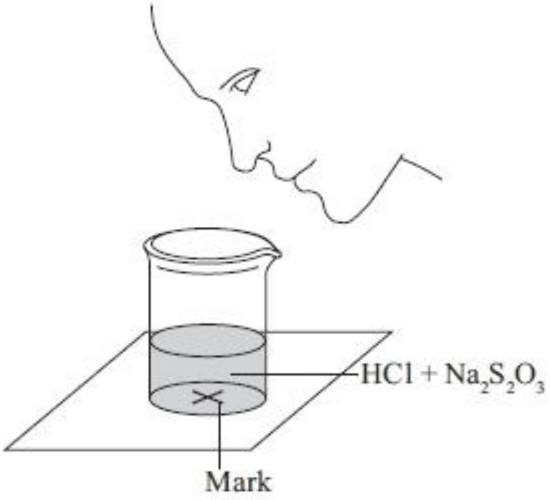


Q# 10/ IB CHEM/2014/s/TZ2/Paper 2 Section B/Higher Level/

6. A group of students investigated the rate of the reaction between aqueous sodium thiosulfate and hydrochloric acid according to the equation below.

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

The two reagents were rapidly mixed together in a beaker and placed over a mark on a piece of paper. The time taken for the precipitate of sulfur to obscure the mark when viewed through the reaction mixture was recorded.



Initially they measured out 10.0 cm^3 of 0.500 mol dm^{-3} hydrochloric acid and then added 40.0 cm^3 of 0.0200 mol dm^{-3} aqueous sodium thiosulfate. The mark on the paper was obscured 47 seconds after the solutions were mixed.

(a) The teacher asked the students to measure the effect of halving the concentration of sodium thiosulfate on the rate of reaction.

(i) State the volumes of the liquids that should be mixed. *[1]*

Liquid	0.500 mol dm^{-3} HCl	0.0200 mol dm^{-3} $Na_2S_2O_3$	Water
Volume / cm^3			

(ii) State why it is important that the students use a similar beaker for both reactions. *[1]*

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(iii) If the reaction were first order with respect to the thiosulfate ion, predict the time it would take for the mark on the paper to be obscured when the concentration of sodium thiosulfate solution is halved. *[1]*

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(b) One proposed mechanism for this reaction is:

$$S_2O_3^{2-}(aq) + H^+(aq) \rightleftharpoons HS_2O_3^-(aq) \qquad \textit{Fast}$$

$$HS_2O_3^-(aq) + H^+(aq) \rightarrow SO_2(g) + S(s) + H_2O(l) \qquad \textit{Slow}$$

(i) Deduce the rate expression of this mechanism. *[1]*

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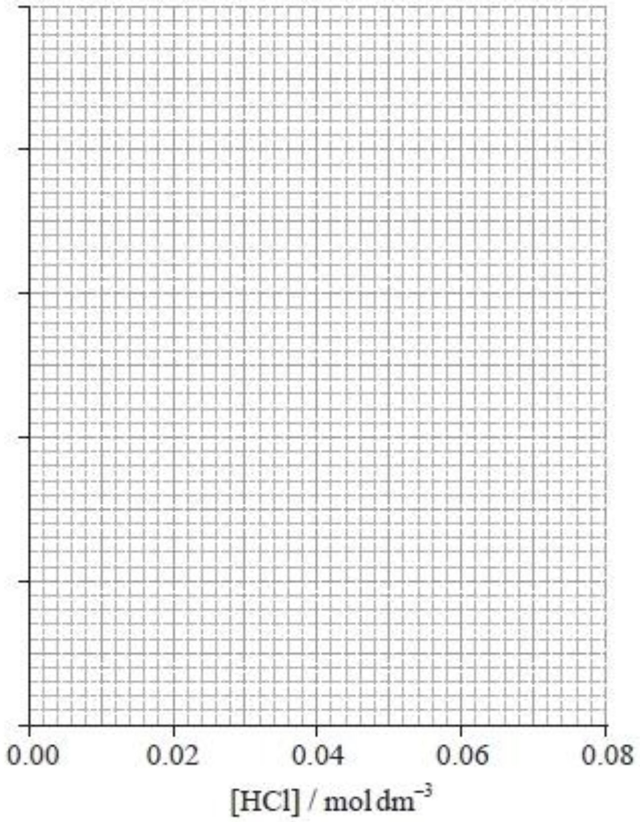
(ii) The results of an experiment investigating the effect of the concentration of hydrochloric acid on the rate, while keeping the concentration of thiosulfate at the original value, are given in the table below.

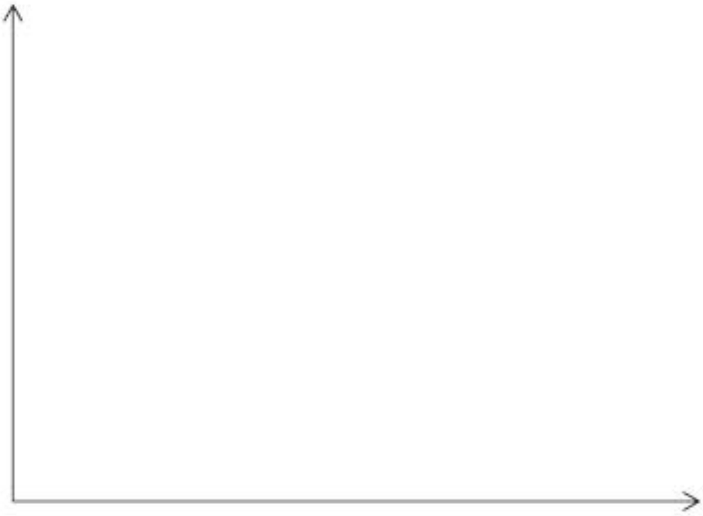
[HCl] / mol dm^{-3}	0.020	0.040	0.060	0.080
Time / s	89.1	72.8	62.4	54.2

On the axes provided, draw an appropriate graph to investigate the order of the reaction with respect to hydrochloric acid. *[3]*



(iii) Identify **two** ways in which these data **do not** support the rate expression deduced in part (i). *[2]*

(c) (i) Sketch and label, indicating an approximate activation energy, the Maxwell–Boltzmann energy distribution curves for two temperatures, T_1 and T_2 ($T_2 > T_1$), at which the rate of reaction would be significantly different. *[3]*



(ii) Explain why increasing the temperature of the reaction mixture would significantly increase the rate of the reaction. *[3]*

Q# 11/ IB Chem/2014/s/TZ1/Paper 2 Section A/Higher Level/

2. Sodium thiosulfate solution, $Na_2S_2O_3(aq)$, and hydrochloric acid, $HCl(aq)$, react to produce solid sulfur as in the equation below.

$$S_2O_3^{2-}(aq) + 2H^+(aq) \rightarrow S(s) + SO_2(g) + H_2O(l)$$

The following results to determine the initial rate were obtained:

Experiment	**$[S_2O_3^{2-}(aq)]$ / mol dm^{-3}**	**$[H^+(aq)]$ / mol dm^{-3}**	**Initial rate / mol dm^{-3} s^{-1}**
1	0.200	2.00	0.036
2	0.200	1.00	0.036
3	0.100	1.00	0.018

(a) Deduce, with a reason, the order of reaction with respect to each reactant. *[2]*

(b) State the rate expression for this reaction. *[1]*

(c) Determine the value of the rate constant, k, and state its units. *[2]*

(d) State an equation for a possible rate-determining step for the reaction. *[1]*

(e) Suggest how the activation energy, E_a, for this reaction may be determined. *[3]*

Q# 12/ IB Chem/2013/w/TZ0/Paper 2 Section A/Higher Level/

1. Reaction kinetics can be investigated using the iodine clock reaction. The equations for two reactions that occur are given below.

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

Reaction B: $I_2(aq) + 2S_2O_3^{2-}(aq) \rightarrow 2I^-(aq) + S_4O_6^{2-}(aq)$

Reaction B is much faster than reaction A, so the iodine, I_2, formed in reaction A immediately reacts with thiosulfate ions, $S_2O_3^{2-}$, in reaction B, before it can react with starch to form the familiar blue-black, starch-iodine complex.

In one experiment the reaction mixture contained:

$5.0 \pm 0.1\,cm^3$ of $2.00\,mol\,dm^{-3}$ hydrogen peroxide (H_2O_2)
$5.0 \pm 0.1\,cm^3$ of 1% aqueous starch
$20.0 \pm 0.1\,cm^3$ of $1.00\,mol\,dm^{-3}$ sulfuric acid (H_2SO_4)
$20.0 \pm 0.1\,cm^3$ of $0.0100\,mol\,dm^{-3}$ sodium thiosulfate ($Na_2S_2O_3$)
$50.0 \pm 0.1\,cm^3$ of water with $0.0200 \pm 0.0001\,g$ of potassium iodide (KI) dissolved in it.

After 45 seconds this mixture suddenly changed from colourless to blue-black.

(a) The concentration of iodide ions, I^-, is assumed to be constant. Outline why this is a valid assumption. *[1]*

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(b) For this mixture the concentration of hydrogen peroxide, H_2O_2, can also be assumed to be constant. Explain why this is a valid assumption. *[2]*

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(c) Explain why the solution suddenly changes colour. *[2]*

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(e) The colour change occurs when 1.00×10^{-4} mol of iodine has been formed. Use the total volume of the solution and the time taken, to calculate the rate of the reaction, including appropriate units. *[4]*

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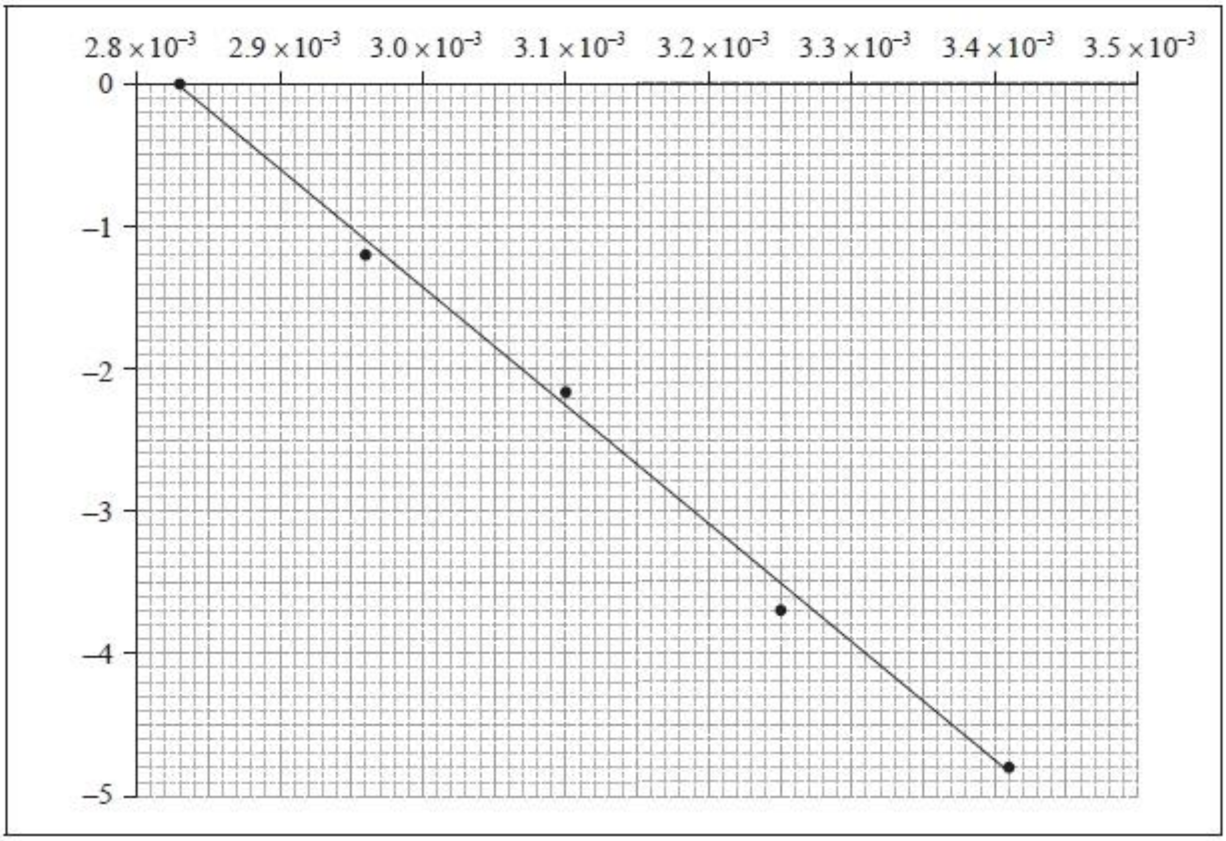
(f) The activation energy can be determined using the Arrhenius equation, which is given in Table 1 of the Data Booklet. The experiment was carried out at five different temperatures. An incomplete graph to determine the activation energy of the reaction, based on these results, is shown below.



(i) State the labels for each axis. *[2]*

x-axis:

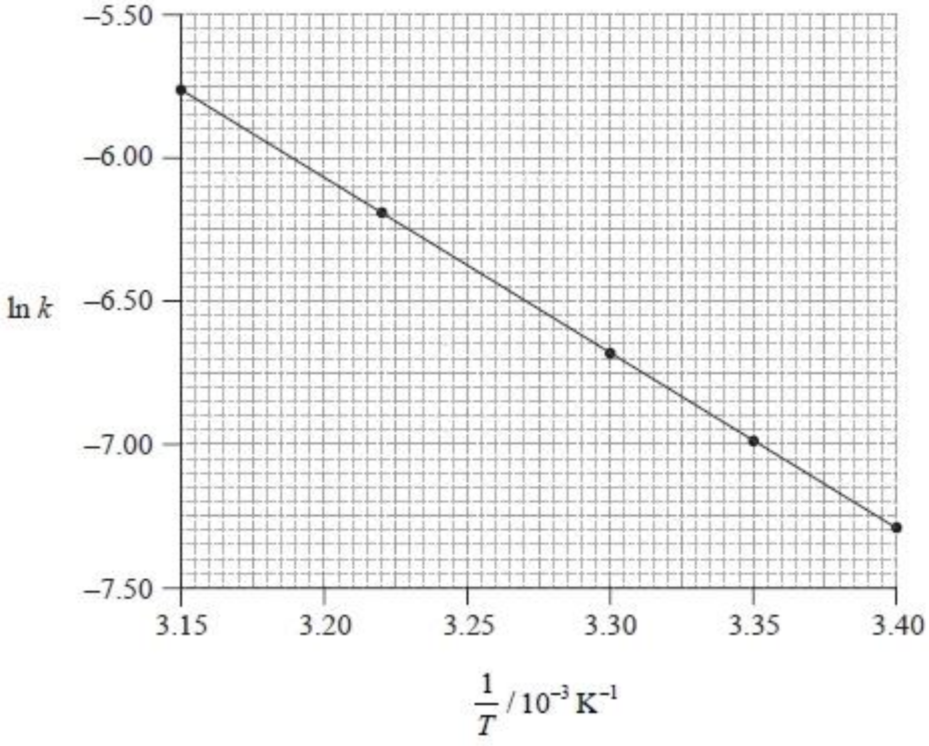
y-axis:

(ii) Use the graph to determine the activation energy of the reaction, in kJ mol^{-1}, correct to **three** significant figures. *[3]*

(g) In another experiment, 0.100 g of a black powder was also added while all other concentrations and volumes remained unchanged. The time taken for the solution to change colour was now 20 seconds. Outline why you think the colour change occurred more rapidly and how you could confirm your hypothesis. *[2]*

Q# 13/ IB Chem/2013/s/TZ2/Paper 2 Section A/Higher Level/

2. Consider the following graph of ln k against $\frac{1}{T}$.



(a) A catalyst provides an alternative pathway for a reaction, lowering the activation energy, E_a. Define the term *activation energy*, E_a. *[1]*

(b) State how the rate constant, k, varies with temperature, T. *[1]*

(c) Determine the activation energy, E_a, correct to **three** significant figures and state its units. *[3]*

Q# 14/ IB Chem/2013/s/TZ1/Paper 2 Section B/Higher Level/Q9

(c) The time taken to produce a certain amount of product using different initial concentrations of C_4H_9Br and NaOH is measured. The results are shown in the following table.

Reaction	**[C_4H_9Br] / 10^{-2} mol dm^{-3}**	**[NaOH] / 10^{-3} mol dm^{-3}**	***t* / s**
A	1.0	2.0	46
B	2.0	2.0	23
C	2.0	4.0	23

(i) Deduce the order of reaction with respect to C_4H_9Br and NaOH, using the data above. *[3]*

C_4H_9Br:

NaOH:

(ii) Deduce the rate expression. *[1]*

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(iii) Based on the rate expression obtained in (c) (ii) state the units of the rate constant, *k*. *[1]*

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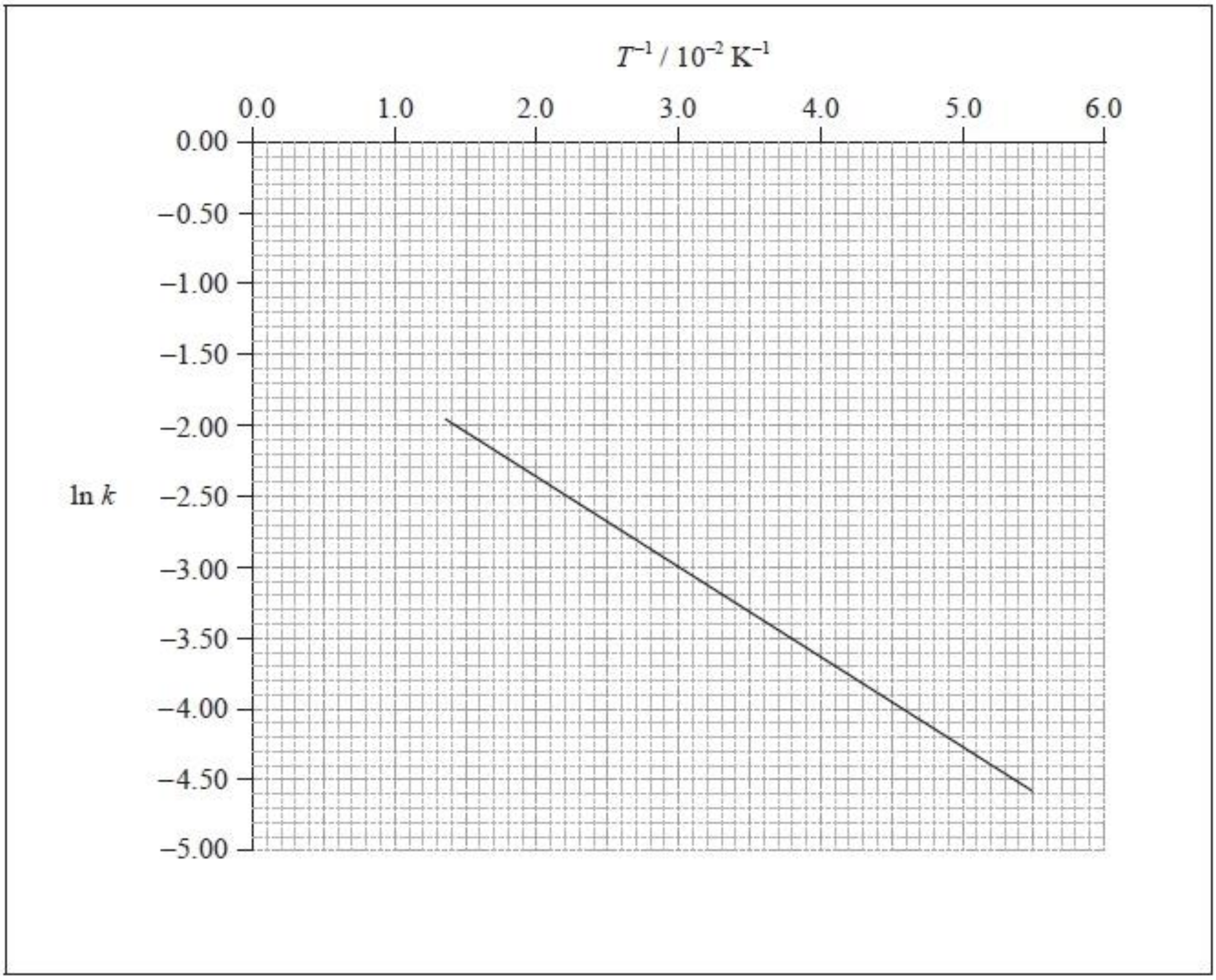
Q# 15/ IB Chem/2013/s/TZ1/Paper 2 Section A/Higher Level/

3. To determine the activation energy of a reaction, the rate of reaction was measured at different temperatures. The rate constant, k, was determined and $\ln k$ was plotted against the inverse of the temperature in Kelvin, T^{-1}. The following graph was obtained.



(a) Define the term *activation energy*, E_a. [1]

(b) Use the graph on page 8 to determine the value of the activation energy, E_a, in $kJ\,mol^{-1}$. [2]

(c) On the graph on page 8, sketch the line you would expect if a catalyst is added to the reactants. [1]

Q# 16/ IB Chem/2012/w/TZ0/Paper 2 Section B/Higher Level/

6. Chemical kinetics involves an understanding of how the molecular world changes with time.

(a) (i) Define the term *rate of reaction*. [1]

(ii) Temperature and the addition of a catalyst are two factors that can affect the rate of a reaction. State **two** other factors. [2]

(iii) In the reaction represented below, state **one** method that can be used to measure the rate of the reaction.

$$ClO_3^-(aq) + 5Cl^-(aq) + 6H^+(aq) \rightarrow 3Cl_2(aq) + 3H_2O(l)$$ *[1]*

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(b) A catalyst provides an alternative pathway for a reaction, lowering the activation energy, E_a.

(i) Define the term *activation energy*, E_a. *[1]*

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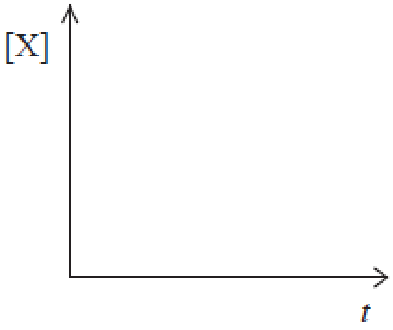
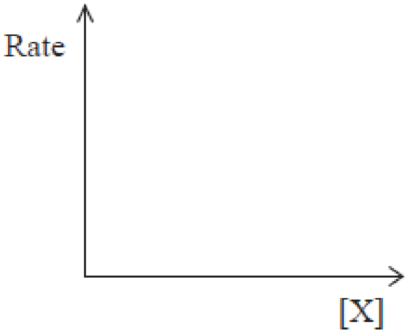
(ii) Sketch the **two** Maxwell–Boltzmann energy distribution curves for a fixed amount of gas at two different temperatures, T_1 and T_2 ($T_2 > T_1$). Label **both** axes. *[3]*

(c) Sketch graphical representations of the following reactions, for X → products.

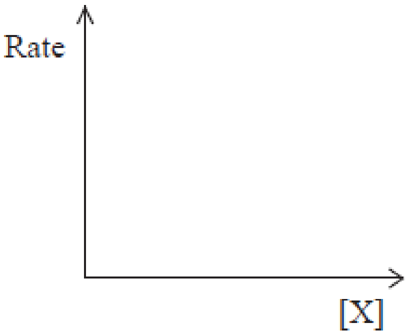
(i) Concentration of reactant X against time for a **zero-order** reaction. *[1]*



(ii) Rate of reaction against concentration of reactant X for a **zero-order** reaction. *[1]*



(iii) Rate of reaction against concentration of reactant X for a **first-order** reaction. *[1]*



(d) 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 $[ClO_2(aq)]$ / $mol\,dm^{-3}$**	**Initial $[OH^-(aq)]$ / $mol\,dm^{-3}$**	**Initial rate / $mol\,dm^{-3}\,s^{-1}$**
1	1.00×10^{-1}	1.00×10^{-1}	2.30×10^{-1}
2	5.00×10^{-2}	1.00×10^{-1}	5.75×10^{-2}
3	5.00×10^{-2}	3.00×10^{-2}	1.73×10^{-2}

(i) Deduce the rate expression. *[2]*

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(ii) Determine the rate constant, *k*, and state its units, using the data from Experiment 2. *[2]*

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(iii) Calculate the rate, in $mol\,dm^{-3}\,s^{-1}$, when $[ClO_2(aq)] = 1.50\times10^{-2}\,mol\,dm^{-3}$ and $[OH^-(aq)] = 2.35\times10^{-2}\,mol\,dm^{-3}$. *[1]*

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(e)

(ii) Describe **qualitatively** the relationship between the rate constant, *k*, and temperature, *T*. *[1]*

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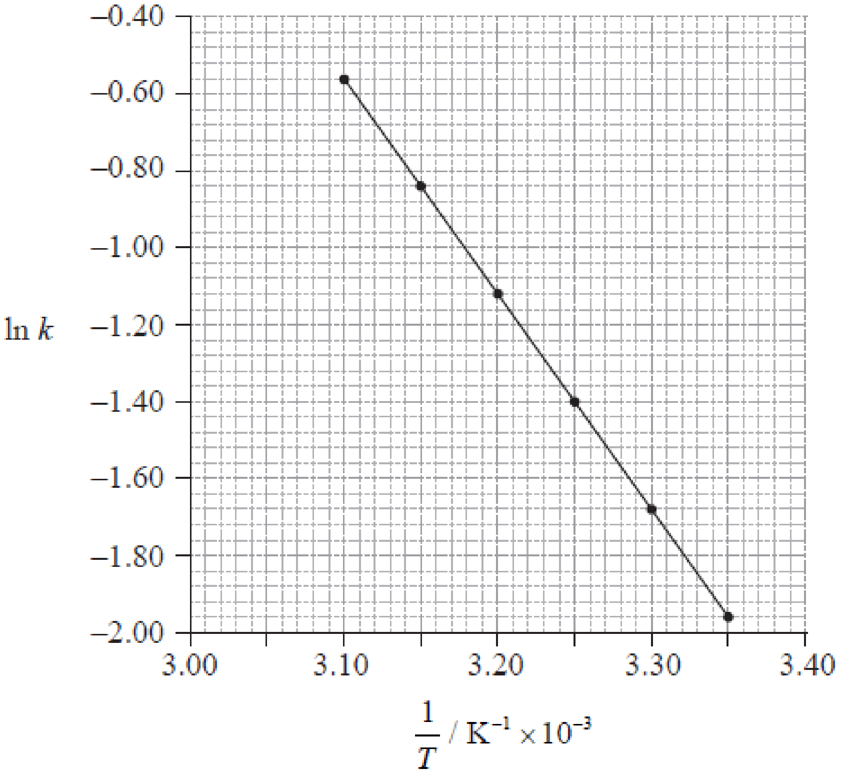
(iii) The rate of this reaction was measured at different temperatures and the following data were recorded.



Using data from the graph, determine the activation energy, E_a, correct to **three** significant figures and **state its units**. *[4]*

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(f) A two-step mechanism has been proposed for the following reaction.

Step 1: $ClO^-(aq) + ClO^-(aq) \rightarrow ClO_2^-(aq) + Cl^-(aq)$

Step 2: $ClO_2^-(aq) + ClO^-(aq) \rightarrow ClO_3^-(aq) + Cl^-(aq)$

(i) Deduce the overall equation for the reaction. *[1]*

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(ii) Deduce the rate expression for each step. *[2]*

Step 1:

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Step 2:

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Q# 17/ IB Chem/2012/w/TZ0/Paper 2 Section A/Higher Level/Q1a

(iv) Using a digital thermometer, the students discovered that the reaction was exothermic. State the sign of the enthalpy change of the reaction, ΔH. *[1]*

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Q# 18/ IB Chem/2011/w/TZ0/Paper 2 Section B/Higher Level/Q8/Data Booklet

(e) $BF_3(g)$ reacts with $NH_3(g)$ to form $F_3BNH_3(g)$ according to the equation below.

$$BF_3(g) + NH_3(g) \rightarrow F_3BNH_3(g)$$

(i) Identify the type of bond present between BF_3 and NH_3 in $F_3BNH_3(g)$ and state another example of a compound with this type of bonding. *[2]*

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(ii) The table below shows initial rates of reaction for different concentrations of each reactant for this reaction at temperature, T.

Experiment	**$[BF_3(g)]$ / mol dm^{-3}**	**$[NH_3(g)]$ / mol dm^{-3}**	**Initial rate / mol dm^{-3} s^{-1}**
1	1.00×10^{-1}	6.67×10^{-2}	2.27×10^{-2}
2	1.00×10^{-1}	3.75×10^{-2}	1.28×10^{-2}
3	2.50×10^{-1}	2.50×10^{-1}	2.13×10^{-1}
4	3.00×10^{-1}	1.00×10^{-1}	1.02×10^{-1}

Deduce the rate expression, the overall order of the reaction and determine the value of k, the rate constant, with its units, using the data from Experiment 4. *[3]*

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(f) The following is a proposed mechanism for the reaction of $NO(g)$ with $H_2(g)$.

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

Step 2: $N_2O_2(g) + H_2(g) \rightarrow N_2O(g) + H_2O(g)$

(i) Identify the intermediate in the reaction. *[1]*

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(ii) The observed rate expression is rate $= k[NO]^2[H_2]$. Assuming that the proposed mechanism is correct, comment on the relative speeds of the two steps. *[1]*

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(g) The following two-step mechanism has been suggested for the reaction of $NO_2(g)$ with $CO(g)$, where $k_2 >> k_1$.

Step 1: $NO_2(g) + NO_2(g) \xrightarrow{k_1} NO(g) + NO_3(g)$

Step 2: $NO_3(g) + CO(g) \xrightarrow{k_2} NO_2(g) + CO_2(g)$

Overall: $NO_2(g) + CO(g) \longrightarrow NO(g) + CO_2(g)$

The experimental rate expression is rate $= k_1[NO_2]^2$. Explain why this mechanism produces a rate expression consistent with the experimentally observed one. *[2]*

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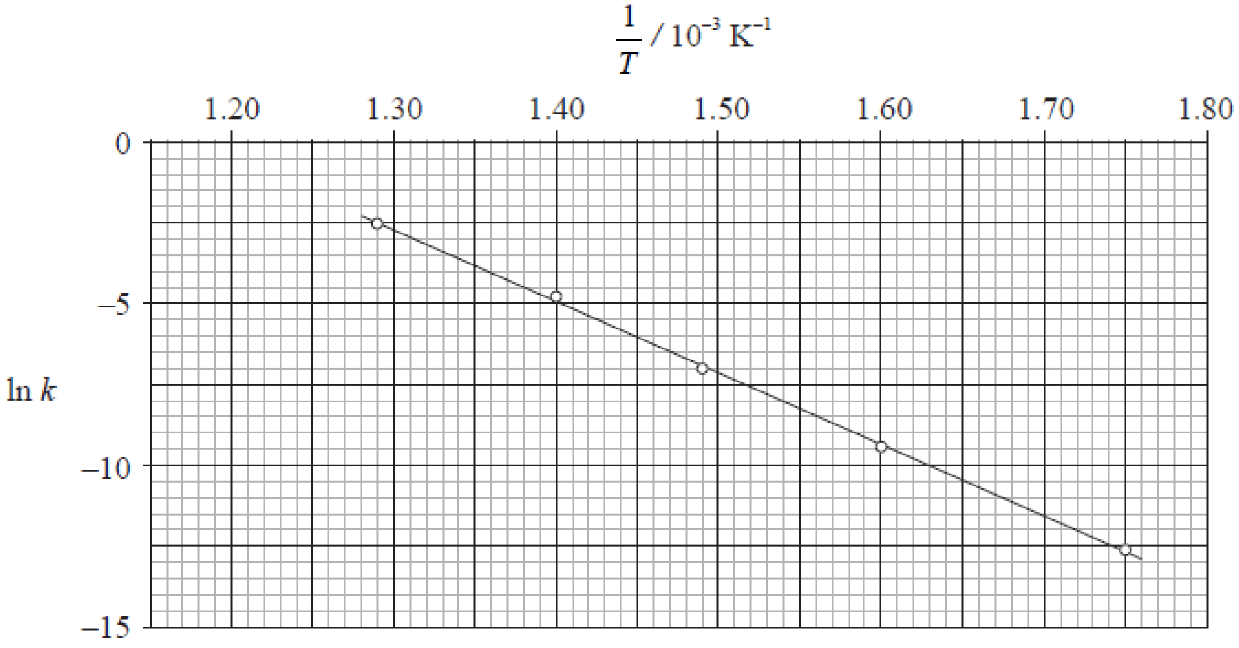
(h) HI (g) decomposes into H_2 (g) and I_2 (g) according to the reaction below.

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

The reaction was carried out at different temperatures and a value of the rate constant, k, was obtained for each temperature. A graph of ln k against $\frac{1}{T}$ is shown below.



Calculate the activation energy, E_a, for the reaction using these data and Table 1 of the Data Booklet showing your working. *[4]*

Q# 19/ IB Chem/2011/w/TZ0/Paper 2 Section B/Higher Level/

8. (a) Define the term *activation energy*, E_a. *[1]*

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(b) Nitrogen monoxide, NO, is involved in the decomposition of ozone according to the following mechanism.

$$O_3 \rightarrow O_2 + O\bullet$$

$$O_3 + NO \rightarrow NO_2 + O_2$$

$$NO_2 + O\bullet \rightarrow NO + O_2$$

$$\text{Overall: } 2O_3 \rightarrow 3O_2$$

State and explain whether or not NO is acting as a catalyst. *[2]*

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(c) (i) Define the term *endothermic reaction*. *[1]*

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(ii) Sketch the Maxwell-Boltzmann energy distribution curve for a reaction with and without a catalyst, and label both axes. *[3]*

(d) Nitrogen reacts with hydrogen to form ammonia in the Haber process, according to the following equilibrium.

$$N_2(g) + 3H_2(g) \rightleftharpoons 2NH_3(g) \qquad \Delta H^\ominus = -92.6\ \text{kJ}$$

(i) Define the term *rate of reaction*. *[1]*

Q# 20/ IB Chem/2011/w/TZ0/Paper 2 Section A/Higher Level/ NOT with 1b

Total amount of gas produced = $34.1 dm^3$

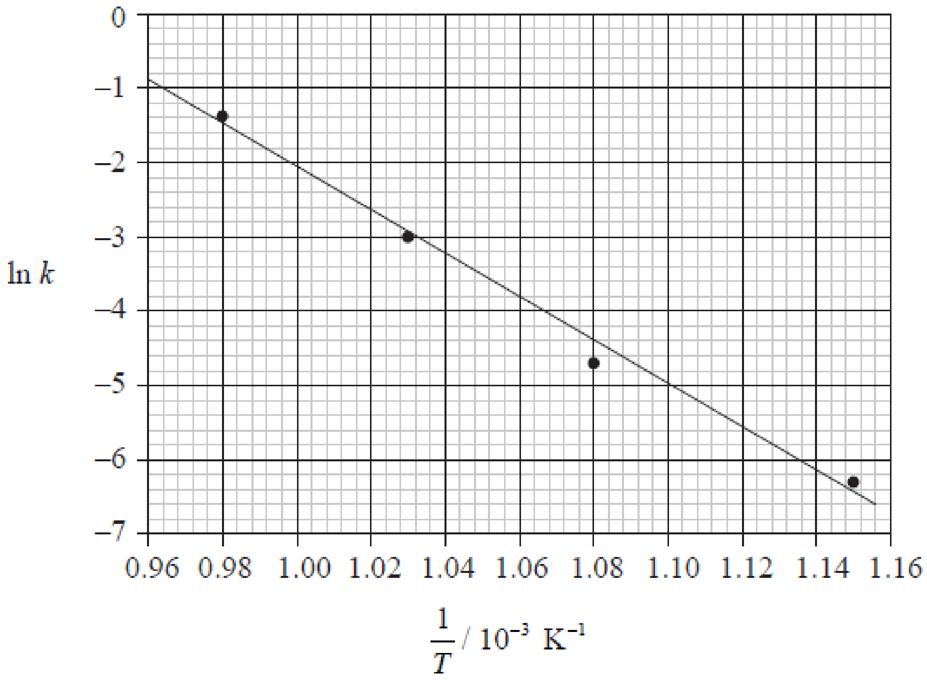
(d) An airbag inflates very quickly.

(i) It takes just 0.0400 seconds to produce nitrogen gas in the simulation. Calculate the average rate of formation of nitrogen in (b) (iii) and state its units. *[1]*

(ii) The students also discovered that a small increase in temperature (*e.g.* 10 °C) causes a large increase (*e.g.* doubling) in the rate of this reaction. State **one** reason for this. *[1]*

2. Consider the following graph of ln k against $\frac{1}{T}$ (temperature in Kelvin) for the second order decomposition of N_2O into N_2 and O.

$$N_2O \rightarrow N_2 + O$$



(a) State how the rate constant, k varies with temperature, T. *[1]*

(b) Determine the activation energy, E_a, for this reaction. *[3]*

(c) The rate expression for this reaction is rate $= k[N_2O]^2$ and the rate constant is $0.244\ \text{dm}^3\ \text{mol}^{-1}\ \text{s}^{-1}$ at $750\ ^\circ\text{C}$.

A sample of N_2O of concentration $0.200\ \text{mol dm}^{-3}$ is allowed to decompose. Calculate the rate when 10 % of the N_2O has reacted. *[2]*

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Q# 22/ IB Chem/2010/s/TZ1/Paper 2 Section A/Higher Level/

2. Alex and Hannah were asked to investigate the kinetics involved in the iodination of propanone. They were given the following equation by their teacher.

$$CH_3COCH_3(aq) + I_2(aq) \xrightarrow{H^+(aq)} CH_2ICOCH_3(aq) + HI(aq)$$

Alex's hypothesis was that the rate will be affected by changing the concentrations of the propanone and the iodine, as the reaction can happen without a catalyst. Hannah's hypothesis was that as the catalyst is involved in the reaction, the concentrations of the propanone, iodine and the hydrogen ions will all affect the rate.

They carried out several experiments varying the concentration of one of the reactants or the catalyst whilst keeping other concentrations and conditions the same, and obtained the results below.

Experiment	**Composition by volume of mixture / cm^3**				**Initial rate / mol dm^{-3} s^{-1}**
	1.00 mol dm^{-3} $CH_3COCH_3(aq)$	**Water**	**1.00 mol dm^{-3} $H^+(aq)$**	**5.00×10^{-3} mol dm^{-3} I_2 in KI**	
1	10.0	60.0	10.0	20.0	4.96×10^{-6}
2	10.0	50.0	10.0	30.0	5.04×10^{-6}
3	5.0	65.0	10.0	20.0	2.47×10^{-6}
4	10.0	65.0	5.0	20.0	2.51×10^{-6}

(a) Explain why they added water to the mixtures. *[1]*

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(b) (i) Deduce the order of reaction for each substance and the rate expression from the results. *[2]*

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(ii) Comment on whether Alex's or Hannah's hypothesis is correct. *[1]*

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(c) Using the data from Experiment 1, determine the concentration of the substances used and the rate constant for the reaction including its units. *[3]*

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(d) (i) This reaction uses a catalyst. Sketch and annotate the Maxwell-Boltzmann energy distribution curve for a reaction with and without a catalyst on labelled axes below. *[3]*

(ii) Describe how a catalyst works. *[1]*

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6. (a) Consider the following reaction studied at 263 K.

$$2NO(g) + Cl_2(g) \rightleftharpoons 2NOCl(g)$$

It was found that the forward reaction is first order with respect to Cl_2 and second order with respect to NO. The reverse reaction is second order with respect to NOCl.

(i) State the rate expression for the forward reaction. *[1]*

(ii) Predict the effect on the rate of the forward reaction and on the rate constant if the concentration of NO is halved. *[2]*

(b) Consider the following reaction.

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

Possible reaction mechanisms are:

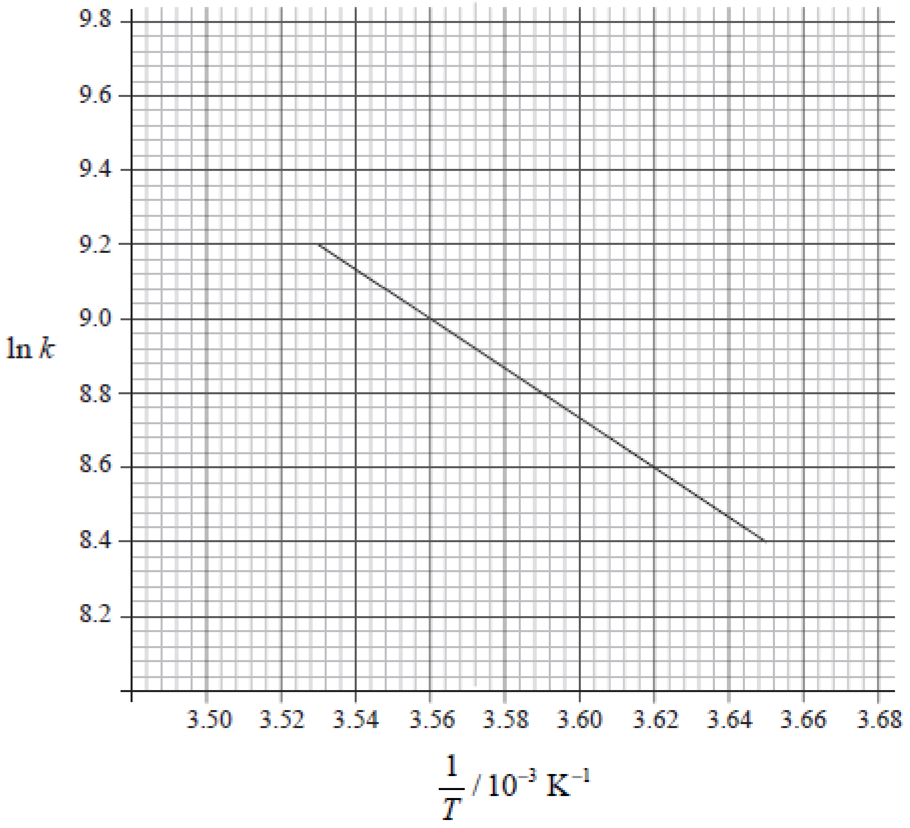
Above 775 K: $NO_2 + CO \rightarrow NO + CO_2$ *slow*

Below 775 K: $2NO_2 \rightarrow NO + NO_3$ *slow*
$NO_3 + CO \rightarrow NO_2 + CO_2$ *fast*

Based on the mechanisms, deduce the rate expressions above and below 775 K. *[2]*

(c) State two situations when the rate of a chemical reaction is equal to the rate constant. *[2]*

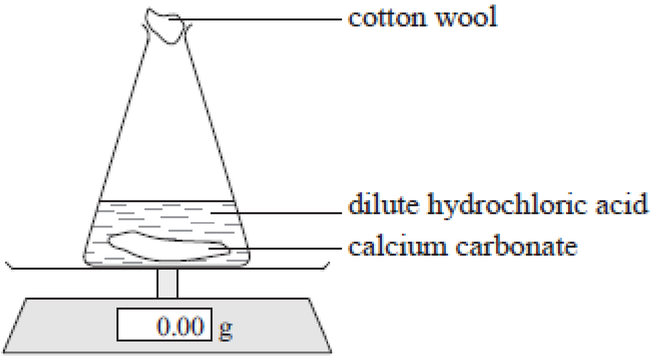
(d) Consider the following graph of ln *k* against $\frac{1}{T}$ for the first order decomposition of N_2O_4 into NO_2. Determine the activation energy in kJ mol^{-1} for this reaction. *[2]*



Q# 24/ IB Chem/2008/w/TZ0/Paper 2 Section A/Higher Level/

1. (a) The diagram shows the apparatus used to study the rate of reaction between calci carbonate and hydrochloric acid.

$$CaCO_3(s) + 2HCl(aq) \rightarrow CaCl_2(aq) + H_2O(l) + CO_2(g)$$



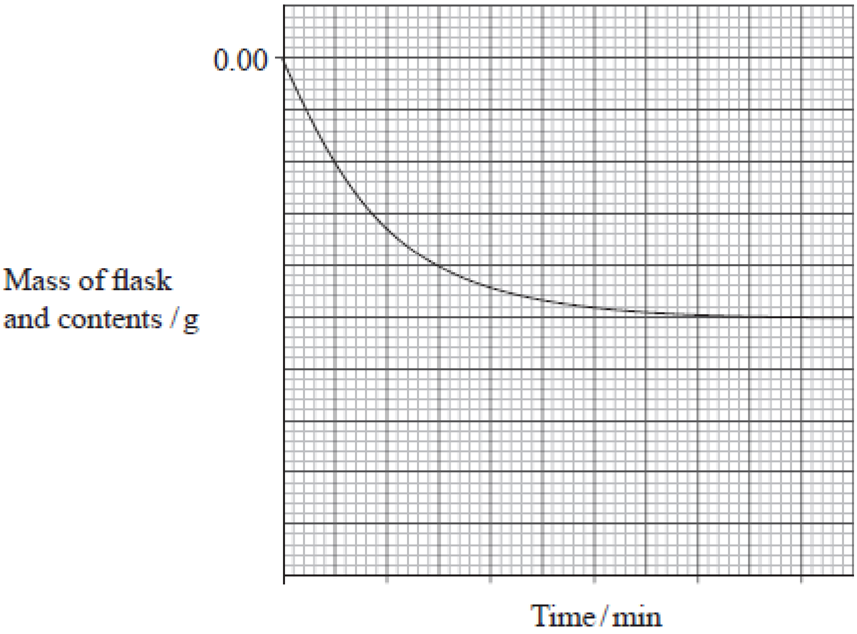
The quantities of reactants added to the flask in one experiment carried out at ro temperature were:

mass of single piece of $CaCO_3(s)$ = 5.00 g

volume of 1.00 mol dm^{-3} $HCl(aq)$ = 50.0 cm^3

The balance was set to zero at the start of the experiment.

The graph shows how the mass of the flask and contents changed during Experiment



(i) Explain why the mass decreased. *[1]*

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(ii) Calculate the amount, in moles, of each reactant at the start of Experiment 1. *[2]*

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(iii) Use your answers to (a) (ii), and the equation for the reaction, to deduce which reactant was added in excess. *[1]*

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(iv) The experiment was repeated with **small** pieces of calcium carbonate. Draw **two** lines (labelled 2 and 3) on the graph to show how the mass of the flask and contents changes in the following experiments at the same temperature.

Experiment	**Mass of small pieces of $CaCO_3(s)$ / g**	**Volume of 1.00 mol dm^{-3} HCl(aq) / cm^3**
2	2.50	50.0
3	5.00	25.0

[4]

(b) For the reaction between compounds **A** and **B** the initial rate was measured in a series of reactions carried out at the same temperature.

Experiment	**Initial [A] / mol dm^{-3}**	**Initial [B] / mol dm^{-3}**	**Initial rate / mol dm^{-3} min^{-1}**
1	4.2×10^{-2}	7.8×10^{-2}	8.8×10^{-4}
2	4.2×10^{-2}	3.9×10^{-2}	2.2×10^{-4}
3	8.4×10^{-2}	3.9×10^{-2}	2.2×10^{-4}

(i) Deduce the order of reaction with respect to **A** and to **B**, giving a reason in each case. *[2]*

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(ii) Deduce the rate expression for the reaction. *[1]*

(iii) Use the data for Experiment 1 to determine the value, including units, of the rate constant for the reaction. *[2]*

Q# 25/ IB Chem/2008/s/TZ1/Paper 2 Section B/Higher Level/

7. (a) The oxidation of nitrogen monoxide takes place as follows:

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

The following experimental data was obtained at a constant temperature.

Experiment	Initial [NO] / mol dm^{-3}	Initial [O_2] / mol dm^{-3}	Initial rate / mol dm^{-3}s^{-1}
1	3.50×10^{-2}	1.75×10^{-2}	3.75×10^{-3}
2	3.50×10^{-2}	3.50×10^{-2}	7.50×10^{-3}
3	7.00×10^{-2}	7.00×10^{-2}	6.00×10^{-2}

(i) Deduce the order of reaction with respect to O_2 and with respect to NO. *[2]*

(ii) State the rate expression for the reaction. *[1]*

(iii) Calculate the value of the rate constant and state the units. *[2]*

(iv) Calculate the expected rate of reaction if the reactants from experiment 2 were placed in a container of twice the original volume. Explain your answer. *[2]*

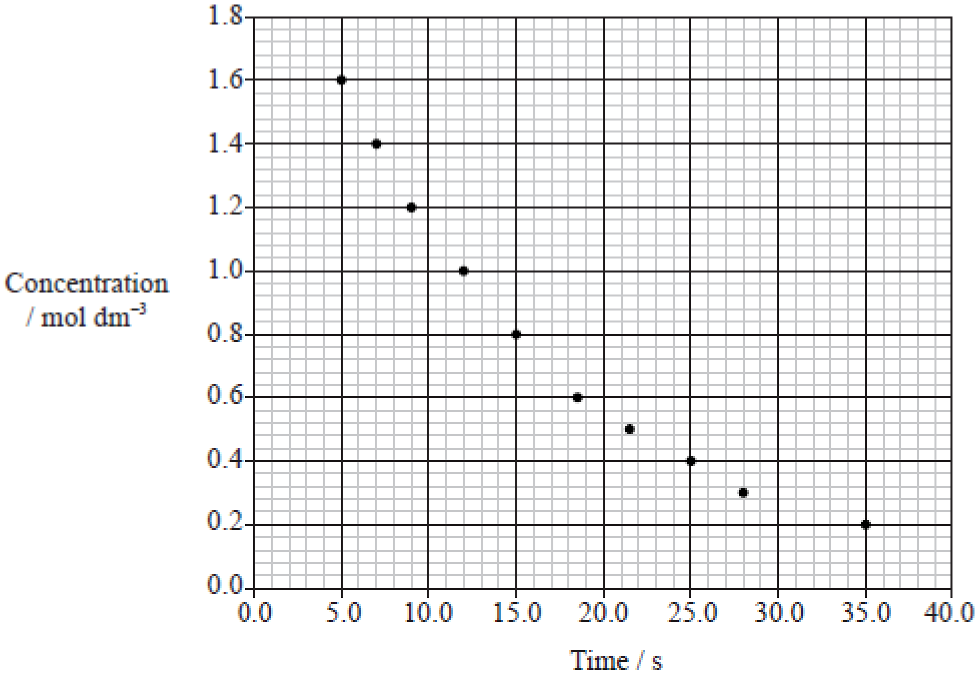
(v) Suggest a possible two step mechanism that is consistent with the rate expression. Identify the rate-determining step. *[3]*

(b) The reaction in (a) is faster at 323 K than at 298 K. Explain the increase in rate in terms of collision theory and Maxwell-Boltzmann energy distribution curves. *[5]*

8. (a) (i) The reaction between propanone, CH_3COCH_3 and bromine, Br_2 in the presence of acid, H^+, is found to be second order overall, but the rate is independent of the bromine concentration. Write **three** possible rate expressions for the reaction. *[3]*

(ii) The concentration of each of the three reactants was doubled in three separate experiments. Choose **one** of the rate expressions in (a) (i) and predict the effect on the rate of the reaction of each of these changes. *[2]*

(iii) The graph below shows how the concentration of propanone changes with time in a reaction.



Use the graph to confirm that the reaction is first order with respect to propanone showing your working. *[2]*

(iv) The overall reaction is:

$$CH_3COCH_3(aq) + Br_2(aq) \xrightarrow{H^+(aq)} CH_3COCH_2Br(aq) + HBr(aq)$$

Describe **one** observation that would allow you to follow the progress of the reaction. State and explain the role of the acid in the reaction. *[4]*

Q# 27/ IB Chem/2007/s/TZ0/Paper 2 Section A/Higher Level/

1. Nitrogen(II) oxide reacts with bromine according to the following equation.

$$2NO(g) + Br_2(g) \rightarrow 2NOBr(g) \qquad \Delta H = \text{negative}$$

The data below were obtained for the reaction between NO(g) and $Br_2(g)$ at a specified temperature and pressure.

Experiment	Initial [NO] / mol dm^{-3}	Initial [Br_2] / mol dm^{-3}	Initial rate / mol dm^{-3}s^{-1}
1	2.00×10^{-2}	5.00×10^{-3}	3.20×10^{-3}
2	2.00×10^{-2}	2.50×10^{-3}	1.60×10^{-3}
3	4.00×10^{-2}	5.00×10^{-3}	1.30×10^{-2}

(a) Determine, giving a reason, the order of reaction with respect to NO and the order of reaction with respect to Br_2. *[2]*

(b) Derive the rate expression for the reaction between NO and Br_2. *[1]*

(c) Calculate the rate constant for the rate expression using experiment 1 and state its units. *[2]*

(d) If the total volume of the reaction mixture was doubled at constant temperature, state the effect, if any, on

(i) the rate constant. *[1]*

(ii) the rate of change of the $Br_2(g)$ concentration. *[1]*

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(e) A heterogeneous catalyst may be used in this reaction. Outline how *heterogeneous catalysts* work. *[2]*

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(f) Draw a labelled enthalpy level diagram for the reaction between NO(g) and $Br_2(g)$, with and without the use of a catalyst. *[3]*

Q# 28/ IB Chem/2006/w/TZ0/Paper 2 Section B/Higher Level/

9. (a) The compound iodine chloride, ICl, reacts with hydrogen to form iodine and hydrogen chloride.

(i) Deduce the equation for this reaction. *[1]*

(ii) The kinetics of this reaction were studied at a certain temperature, when all the reactants and products were in the gas phase. The table shows the initial rate of reaction for different concentrations of reactants.

Experiment	[ICl] / mol dm^{-3}	$[H_2]$ / mol dm^{-3}	Initial rate / mol dm^{-3} s^{-1}
1	0.100	0.0500	5.00×10^{-3}
2	0.200	0.0500	1.00×10^{-2}
3	0.200	0.0250	2.50×10^{-3}

Deduce and explain the order of reaction with respect to ICl and to H_2. *[4]*

(iii) Write the rate expression for the reaction. *[1]*

(iv) Use information from Experiment 1 to determine the value, with units, of the rate constant for the reaction. *[2]*

(v) Determine the rate of reaction when the concentrations of reactants in Experiment 1 are both doubled. *[1]*

(b) The overall equation for a reaction that occurs in two steps is

$$P + 2Q \rightarrow R$$

The rate expression for the reaction is

$$\text{rate} = k[P][Q]$$

(i) Suggest a mechanism for this reaction and identify the rate-determining step. *[2]*

(ii) Deduce the molecularity of the reaction. *[1]*

(c) In a first order reaction it takes four minutes for the concentration of a reactant to decrease from $0.08\ \text{mol dm}^{-3}$ to $0.04\ \text{mol dm}^{-3}$. Deduce the time for the concentration to decrease from $0.04\ \text{mol dm}^{-3}$ to $0.02\ \text{mol dm}^{-3}$. *[1]*

(d) The variation of the rate constant, k, for a reaction with temperature is shown by the Arrhenius equation. Two versions of this equation are shown in Table 1 of the Data Booklet.

(i) Explain the significance of the Arrhenius constant, A, in this equation. *[1]*

(ii) Explain what is meant by the term *activation energy*, E_a. *[1]*

(iii) Describe how, using a graphical method, values of A and E_a can be obtained for a reaction. *[5]*

Q# 29/ IB Chem/2006/s/TZ0/Paper 2 Section B/Higher Level/

6. Consider the following reaction in the Contact process for the production of sulfuric acid for parts (a) to (f) in this question.

$$2SO_2 + O_2 \rightleftharpoons 2SO_3$$

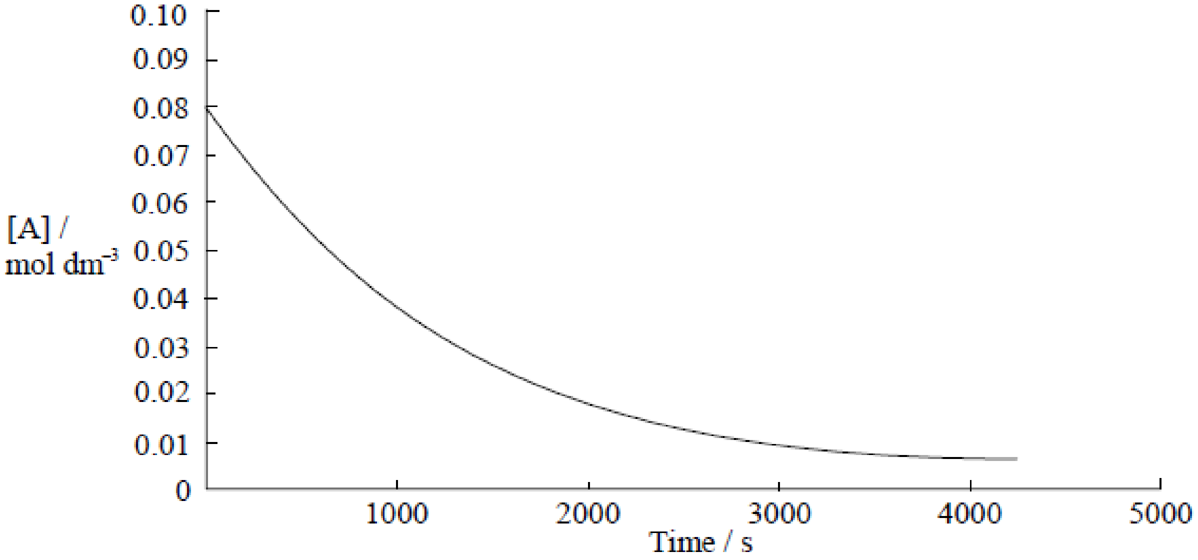
(c) Use the collision theory to explain why increasing the temperature increases the rate of the reaction between sulfur dioxide and oxygen. *[2]*

Q# 30/ IB Chem/2005/w/TZ0/Paper 2 Section A/Higher Level/

1. An equation for the decomposition of substance **A** is

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

A graph showing the change in concentration of **A** against time as the reaction proceeds at a particular temperature is shown below.



(a) Define the term *half-life* of reaction. *[1]*

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(b) Use the graph to measure values of half-life of reaction, starting from *[2]*

time = zero . . . .

time = 2000 s . . . .

(c) Deduce the order of the reaction with respect to **A**, giving a reason for your choice, and write the rate expression for the reaction. *[3]*

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(d) For a different reaction, between compounds **D** and **E**, the rate expression is

$$\text{rate} = k[\text{D}]^2[\text{E}]$$

Calculate the value of k, including units, for the reaction when the concentrations of both **D** and **E** are 1.35×10^{-2} mol dm^{-3} and the reaction rate is 3.75×10^{-5} mol dm^{-3} min^{-1}. *[3]*

Q# 31/ IB Chem/2005/s/TZ0/Paper 2 Section A/Higher Level/

3. When a small quantity of a strongly smelling gas such as ammonia is released into the air, it can be detected several metres away in a short time.

(a) Use the kinetic molecular theory to explain why this happens. *[2]*

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(b) State and explain how the time taken to detect the gas changes when the temperature is increased. *[2]*

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Q# 32/ IB Chem/2005/s/TZ0/Paper 2 Section A/Higher Level/

4. The following reaction

$$2N_2O_5(g) \rightarrow 4NO_2(g) + O_2(g)$$

is described as first order with respect to N_2O_5.

(a) Write the rate expression for the reaction. *[1]*

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(b) One possible mechanism for this reaction is given below.

$$N_2O_5(g) \rightarrow NO_2(g) + NO_3(g) \qquad \text{Step 1}$$
$$N_2O_5(g) + NO_3(g) \rightarrow 2NO_2(g) + O_2(g) \qquad \text{Step 2}$$

Describe the rate expression that would result if the rate determining step in the mechanism is

(i) Step 1. *[1]*

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(ii) Step 2. *[2]*

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Outline your reasoning.

(c) Explain what is meant by the term *half-life* for this reaction. *[1]*

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(d) State what is characteristic about the half-life of a first order reaction. *[1]*

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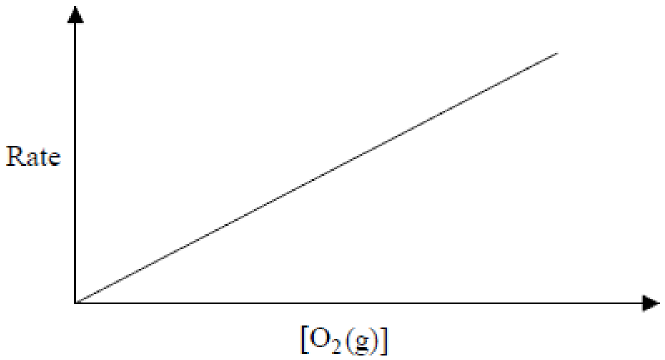
Q# 33/ IB Chem/2004/w/TZ0/Paper 2 Section B/Higher Level/

8. Oxygen and nitrogen monoxide react together to form nitrogen dioxide.

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

The graph below shows how the initial rate of reaction changed during an experiment in which the initial [NO(g)] was kept constant whilst the initial $[O_2(g)]$ was varied.



(a) Deduce, giving a reason, the order of reaction with respect to O_2. *[2]*

(b) In a series of experiments, the initial $[O_2(g)]$ was kept constant while the initial [NO(g)] was varied. The results showed that the reaction was second order with respect to NO. Sketch a graph to show how the rate of reaction would change if the initial [NO(g)] was increased. *[2]*

(c) Deduce the overall order of this reaction. *[1]*

(d) State and explain what would happen to the initial rate of reaction if the initial concentration of NO was doubled and that of O_2 was halved. *[3]*

(e) When the initial values are $[O_2(g)] = 1.0 \times 10^{-2}$ mol dm^{-3} and $[NO(g)] = 3.0 \times 10^{-2}$ mol dm^{-3}, the initial rate of reaction is 6.3×10^{-4} mol dm^{-3} s^{-1}. Write the rate expression for this reaction and calculate the rate constant, stating its units. *[4]*

(f) Nitrogen monoxide may also be converted into nitrogen dioxide at high temperature according to the equation below.

$$NO(g) + CO(g) + O_2(g) \rightarrow NO_2(g) + CO_2(g)$$

(i) Sketch a graph of concentration of $NO_2(g)$ produced against **time** for this reaction and annotate the graph to show how the initial rate of reaction could be deduced. *[3]*

(ii) The results from a series of experiments for this reaction are shown below.
Deduce, giving a reason, the order of reaction with respect to each of the reactants. *[6]*

Experiment	**[NO(g)] / mol dm^{-3}**	**[CO(g)] / mol dm^{-3}**	**[O_2(g)] / mol dm^{-3}**	**Initial rate / mol dm^{-3} s^{-1}**
1	1.00×10^{-3}	1.00×10^{-3}	1.00×10^{-1}	4.40×10^{-4}
2	2.00×10^{-3}	1.00×10^{-3}	1.00×10^{-1}	1.76×10^{-3}
3	2.00×10^{-3}	2.00×10^{-3}	1.00×10^{-1}	1.76×10^{-3}
4	4.00×10^{-3}	1.00×10^{-3}	2.00×10^{-1}	7.04×10^{-3}

(g) Explain why the order of a reaction cannot be obtained directly from the stoichiometric equation. *[1]*

(h) (i) First order reactions have a constant *half-life*. Define the term *half-life* and explain why it is constant for first order reactions. *[2]*

(ii) The value of *k* for a first order reaction is 440 s^{-1}. Calculate the half-life for this reaction. *[1]*

Q# 34/ IB Chem/2004/s/TZ0/Paper 2 Section A/Higher Level/

6. Describe in molecular terms the processes that occur when

(a) a mixture of ice and water is maintained at the melting point. *[2]*

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(b) a sample of a very volatile liquid (such as ethoxyethane) is placed on a person's skin. *[2]*

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Q# 35/ IB Chem/2004/s/TZ0/Paper 2 Section A/Higher Level/

1. Nitrogen(II) oxide reacts with hydrogen as shown by the following equation.

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

The table below shows how the rate of reaction varies as the reactant concentrations vary.

Experiment	**Initial [NO] / mol dm^{-3}**	**Initial [H$_2$] / mol dm^{-3}**	**Initial rate / mol N$_2$ dm^{-3} s^{-1}**
1	0.100	0.100	2.53×10^{-6}
2	0.100	0.200	5.05×10^{-6}
3	0.200	0.100	10.10×10^{-6}
4	0.300	0.100	22.80×10^{-6}

(a) Determine the order of reaction with respect to NO and with respect to H_2. Explain how you determined the order for NO. *[3]*

NO ....

....

....

....

H_2 ....

....

(b) Write the rate expression for the reaction. *[1]*

....

(c) Calculate the value for the rate constant, including its units. *[2]*

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(d) A suggested mechanism for this reaction is as follows.

$$H_2 + NO \rightleftharpoons X \quad \text{fast step}$$
$$X + NO \rightarrow Y + H_2O \quad \text{slow step}$$
$$Y + H_2 \rightarrow N_2 + H_2O \quad \text{fast step}$$

State and explain whether this mechanism agrees with the experimental rate expression in (b). *[4]*

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(e) Explain why a single step mechanism is unlikely for a reaction of this kind. *[2]*

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(f) Deduce how the initial rate of formation of $H_2O(g)$ compares with that of $N_2(g)$ in experiment 1. Explain your answer. *[2]*

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Q# 36/ IB Chem/2003/w/TZ0/Paper 2 Section B/Higher Level/

8. (a) The following data were obtained for the reaction of nitrogen monoxide gas, NO(g), with oxygen gas to form nitrogen dioxide gas, $NO_2(g)$, at 25 °C .

Experiment	**[NO] / mol dm^{-3}**	**[O_2] / mol dm^{-3}**	**Initial rate / mol dm^{-3} s^{-1}**
1	0.50	0.20	3.0×10^{-3}
2	0.50	0.40	6.0×10^{-3}
3	1.00	0.80	4.8×10^{-2}

(i) Calculate the order with respect to the two reactants and write the rate expression for the reaction. Show your reasoning. *[5]*

(ii) Explain why the following mechanism is **not** consistent with the rate expression. *[2]*

$$NO\ (g) + O_2\ (g) \rightarrow NO_2\ (g) + O\ (g)\text{; slow step}$$
$$NO\ (g) + O\ (g) \rightarrow NO_2\ (g)\text{; fast step}$$

(iii) Explain why the following mechanism is consistent with the rate expression, **but** is unlikely. *[2]*

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

(iv) Explain why the following mechanism is consistent with the rate expression. *[3]*

$$NO\ (g) + O_2\ (g) \rightarrow NO_3\ (g)\text{; fast}$$
$$NO_3\ (g) + NO\ (g) \rightarrow 2NO_2\ (g)\text{; slow}$$

(v) Suggest, giving a reason, **one** other mechanism that would be consistent with the rate expression. *[3]*

Q# 37/ IB Chem/2003/s/TZ0/Paper 2 Section B/Higher Level/

6. (a) (i) Draw a graph to show the distribution of energies in a sample of gas molecules. Label the axes and label your curve T_1. Using the same axes, draw a second curve to represent the distribution of energies at a higher temperature. Label this curve T_2. *[3]*

(ii) State and explain, with reference to your graph, what happens to the rate of a reaction when the temperature is increased. *[2]*

(b) (i) State and explain the effect of a catalyst on the rate of a reaction. *[3]*

(ii) Distinguish between the terms *heterogeneous* and *homogeneous* catalyst. *[2]*

(iii) For **each** type of catalyst, state an example and write an equation for a reaction it catalyses. *[4]*

(c) The data below refer to a reaction between X and Y.

	Initial concentration / mol dm^{-3}		**Initial rate of reaction / $\text{mol dm}^{-3}\,\text{s}^{-1}$**
Experiment	**X**	**Y**	
1	0.25	0.25	1.0×10^{-2}
2	0.50	0.25	4.0×10^{-2}
3	0.50	0.50	8.0×10^{-2}

(i) Define the term *order of reaction*. *[1]*

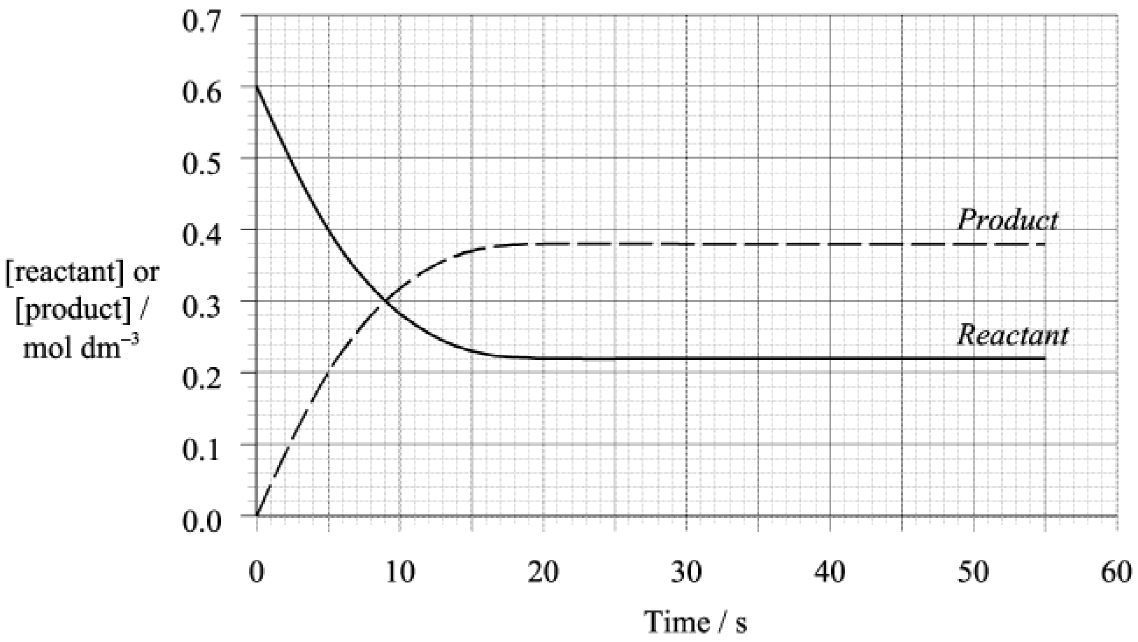
(ii) Deduce the order of reaction with respect to **both** X and Y. Explain your reasoning. *[4]*

(iii) Write the rate expression for the reaction and calculate the rate constant, including its units. *[4]*

(iv) Calculate the initial rate of reaction when the initial concentrations of X and Y are 0.40 mol dm^{-3} and 0.60 mol dm^{-3} respectively. *[2]*

Q# 38/ IB Chem/2003/s/TZ0/Paper 2 Section A/Higher Level/Q3

(b) The following graph represents the change of concentration of reactant and product during a reaction.



(i) Calculate the average rate of reaction over the first 15 s, stating the units. *[3]*

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Q# 39/ IB Chem/2002/w/TZ0/Paper 2 Section A/Higher Level/

3. A balloon, which can hold a maximum of 1100 cm^3 of air before bursting, contains 955 cm^3 of air at 5 °C.

(a) Determine whether the balloon will burst if the temperature is increased to 25 °C. Assume that the pressure of the gas in the balloon remains constant. *[3]*

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(b) Use the kinetic theory to explain what happens to the air particles inside the balloon as the temperature is increased to 25 °C . *[2]*

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Q# 40/ IB Chem/2002/s/TZ0/Paper 2 Section A/Higher Level/

1. The following data were obtained for the reaction between gases **A** and **B**:

Experiment	**Initial concentration of reactants / mol dm^{-3}**		**Initial rate of reaction / mol dm^{-3} min^{-1}**
	A	**B**	
1	1.0×10^{-3}	2.0×10^{-3}	3.0×10^{-4}
2	2.0×10^{-3}	2.0×10^{-3}	3.0×10^{-4}
3	1.0×10^{-3}	4.0×10^{-3}	1.2×10^{-3}

(a) Define the term *overall order of reaction*. *[2]*

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(b) Deduce the order of reaction with respect to **A** and the order of reaction with respect to **B**. *[2]*

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(c) Write the rate expression for the reaction between **A** and **B**. *[1]*

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(d) Use the data from Experiment 1 to calculate the value of the rate constant for the reaction and state its units. *[2]*

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(e) The reaction between gases **A** and **B** is repeated at a pressure double that of the original. Determine how many times faster the reaction will be when the pressure is doubled. *[1]*

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(f) Some solids act as *heterogeneous* catalysts in this reaction. State what is meant by the term *heterogeneous* and outline how such catalysts work. *[3]*

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(g) State and explain how the following changes affect the rate of reaction between **A** and **B**:

(i) Using a catalyst *[2]*

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(ii) Decreasing the temperature *[2]*

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Q# 41/ IB Chem/2001/w/TZ0/Paper 2 Section B/Higher Level/

8. When steam is passed over heated carbon, the following equilibrium is established:

$$C(s) + H_2O(g) \rightleftharpoons H_2(g) + CO(g)$$

(c) In the above experiment, the surface of the carbon is sprayed with a catalyst. Draw an appropriate energy level diagram for the reaction, showing the effect of the catalyst. State the way in which the catalyst affects the rate of the forward reaction, the rate of the reverse reaction and the overall position of equilibrium. *[6]*

Q# 42/ IB Chem/2001/w/TZ0/Paper 2 Section A/Higher Level/Q3

(b) As a volatile liquid in an isolated container evaporates, its temperature drops. Account for this observation in terms of the behaviour of the molecules. *[2]*

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Q# 43/ IB Chem/2001/w/TZ0/Paper 2 Section A/Higher Level/Q1

(b) The reaction between methanol and oxygen is spontaneous at all temperatures. Explain why methanol vapour is stable in air, but the mixture reacts quickly when ignited. *[2]*

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Q# 44/ IB Chem/2001/s/TZ0/Paper 2 Section B/Higher Level/

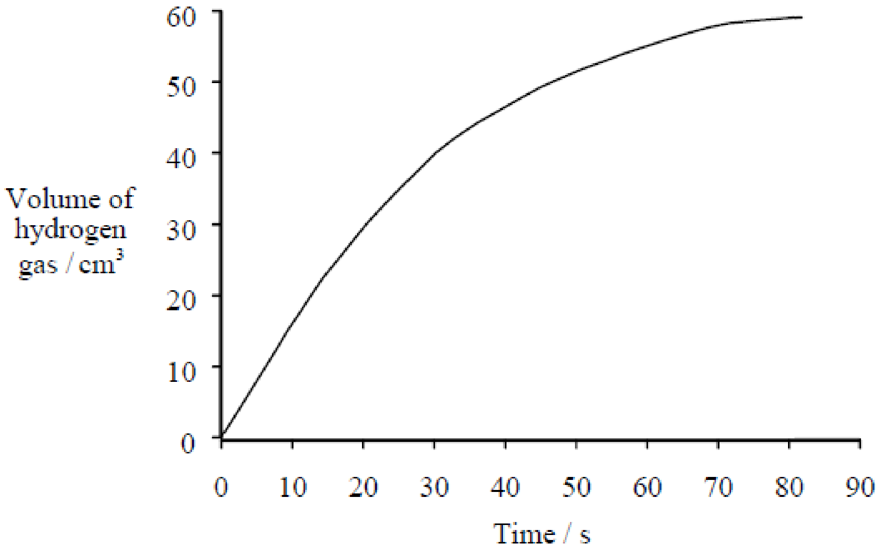
5. Magnesium reacts exothermically with dilute sulfuric acid according to the following equation:

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

(a) Outline an experimental procedure by which you could obtain a value for the rate of this reaction. *[6]*

(b) From the results of such an experiment using excess magnesium ribbon and sulfuric acid of concentration **0.6 mol dm^{-3}**, the graph shown in Figure 1 was obtained. Describe how and explain why the slope of the curve changes with time. *[2]*

Figure 1



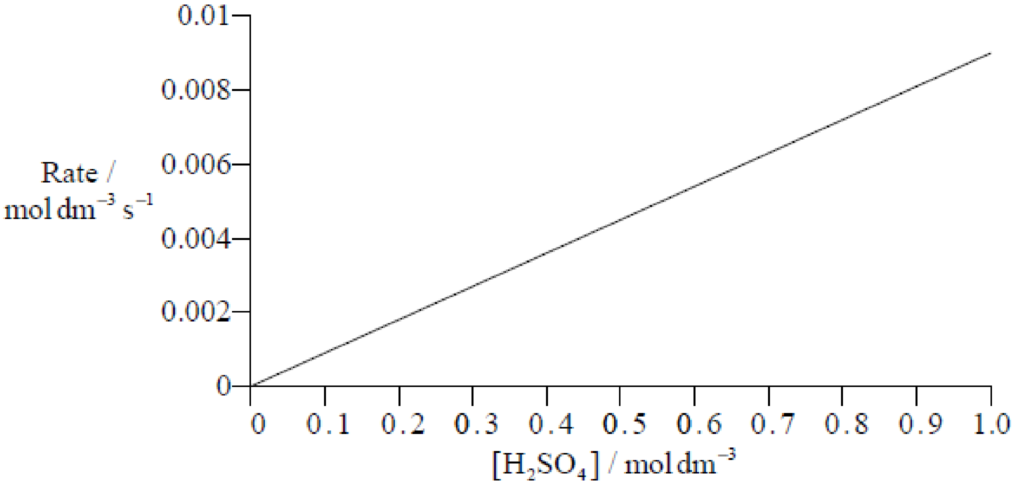
(c) Sketch the graph shown in Figure 1 and label the curve A.

(i) The experiment is repeated using the same mass of magnesium ribbon and the same volume of acid of concentration **0.3 mol dm^{-3}**. Show on the same sketch the curve you would expect to obtain and label it B. Explain your choice at a molecular level. *[3]*

(ii) The experiment is repeated using the same mass of magnesium **powder** and the same volume of acid of concentration **0.6 mol dm^{-3}**. Show on the same sketch the curve you would expect to obtain and label it C. Explain your choice at a molecular level. *[3]*

(d) From the results of a series of experiments involving magnesium and sulfuric acid, a graph of rate against acid concentration was plotted (Figure 2). Use Figure 2 to deduce the order of reaction with respect to sulfuric acid. *[2]*

Figure 2



(e) Under the conditions used, the order of reaction with respect to magnesium is zero. Give the rate expression for this reaction. Calculate the value of the rate constant and give its units. State how the value of the rate constant would change if the experiment were repeated at a higher temperature. *[4]*

Q# 45/ IB Chem/2000/w/TZ0/Paper 2 Section A/Higher Level/

3. Dinitrogen oxide decomposes to give nitrogen and oxygen according to the following equation:

$$2N_2O(g) \rightarrow 2N_2(g) + O_2(g) \qquad \Delta H^{\ominus} = -82 \text{ kJ mol}^{-1}$$

(a) The decomposition is a first order reaction in the presence of gold as a catalyst. The half-life of the catalysed reaction at 834 °C is 1.62×10^4 s.

(i) Calculate the rate constant (velocity constant), k, for the reaction at this temperature and give the units of k. *[1]*

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(ii) Calculate the activation energy of the reaction at this temperature, given the Arrhenius constant, A = 25 s^{-1}. *[2]*

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(iii) The decomposition of dinitrogen oxide without a catalyst is bimolecular. Suggest a possible mechanism for the reaction indicating the equation for each step: *[2]*

Slow step: ....

....

Fast step: ....

....

(b) Draw a labelled diagram showing the potential energy changes during the catalysed and uncatalysed reaction given above. *[2]*

Q# 46/ IB Chem/2000/s/TZ0/Paper 2 Section A/Higher Level/1

(b) The depletion of ozone, O_3, in the upper atmosphere can be caused by the reaction of automobile exhaust gases, such as NO, with the ozone. The reaction between $O_3(g)$ and $NO(g)$ has been studied and the following data were obtained at 25 °C.

Experiment	$[NO(g)]$ / mol dm^{-3}	$[O_3(g)]$ / mol dm^{-3}	Rate / mol dm^{-3} s^{-1}
1	1.00×10^{-6}	3.00×10^{-6}	0.660×10^{-4}
2	1.00×10^{-6}	6.00×10^{-6}	1.32×10^{-4}
3	3.00×10^{-6}	9.00×10^{-6}	5.94×10^{-4}
4	4.50×10^{-6}	7.20×10^{-6}	

(i) Give the rate equation for the reaction between $NO(g)$ and $O_3(g)$, showing your reasoning. *[3]*

....

....

....

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....

(ii) Calculate the value of the rate constant, *k*, stating its units. *[2]*

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....

(iii) Calculate the rate of the reaction for Experiment 4. *[1]*

....

....

Q# 47/ IB Chem/1999/s/TZ0/Paper 2 Section A/Higher Level/

4. The following data were obtained for the reaction between A and B:

$$A(aq) + 2B(aq) \rightarrow 3C(aq) + D(aq)$$

Experiment	Initial concentration of Reactants ($mol\,dm^{-3}$)		Initial Rate of Reaction ($mol\,dm^{-3}\,hr^{-1}$)
	A	B	
1	0.200	0.200	0.50
2	0.400	0.200	2.00
3	0.400	0.800	8.00

(a) Give the order with respect to A. *[1]*

. . . .

(b) Give the order with respect to B. *[1]*

. . . .

(c) Write the rate expression for this reaction. *[1]*

. . . .

(d) Using the data from the first experiment, calculate the value of the rate constant and give its units. *[1]*

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Q# 48/ IB Chem/1999/s/TZ0/Paper 2 Section A/Higher Level/

5. Evidence suggests that reaction between the gases nitrogen dioxide and fluorine is a two-step process:

$$2NO_2(g) + F_2(g) \rightarrow 2NO_2F(g)$$

Step 1 $NO_2 + F_2 \rightarrow NO_2F + F$ (slow)

Step 2 $F + NO_2 \rightarrow NO_2F$ (fast)

(a) State and explain which step is the rate determining step. *[1]*

. . . .

. . . .

(b) State and explain which of the two steps is expected to have the higher activation energy. *[2]*

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. . . .

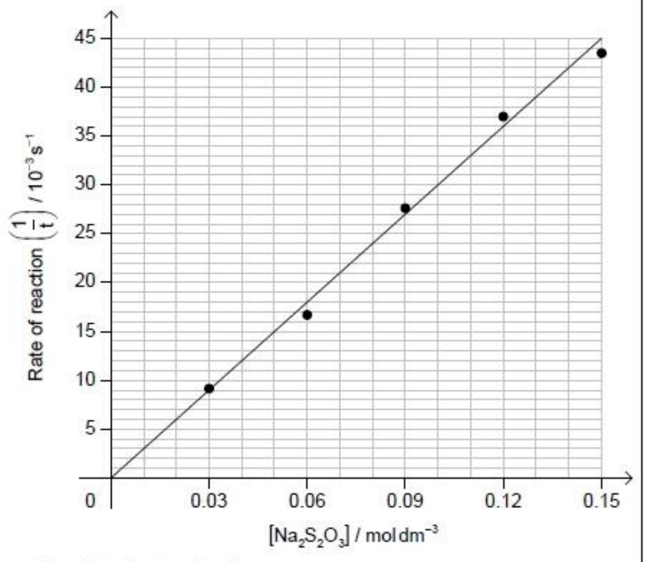
(b) State and explain which of the two steps is expected to have the higher activation energy. *[2]*

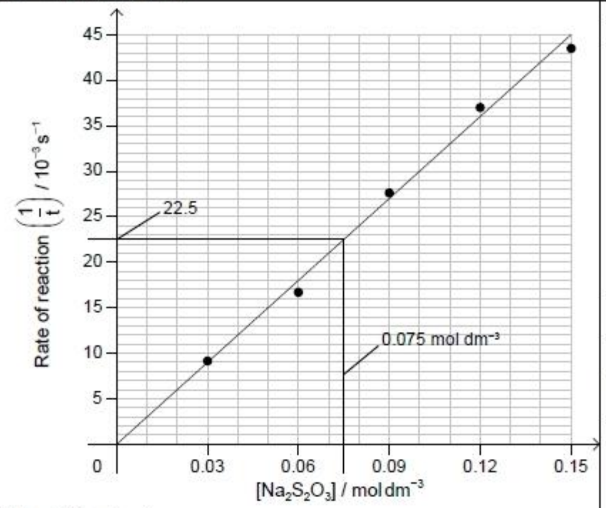
(c) Give the rate expression of the reaction based on your answer to (a). *[1]*

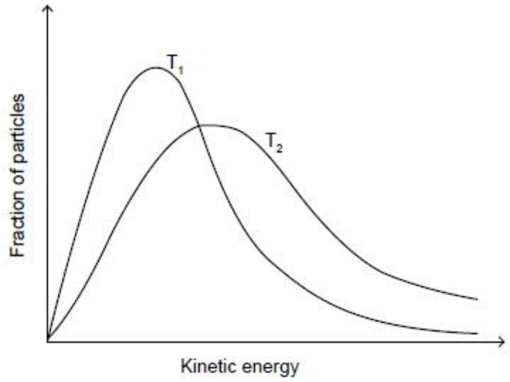
Mark Scheme IB 6 HL EQ 16w to 99s P2A&B 447 marks

Q# 1/ IB Chem/2016/w/TZ0/Paper 2 Section A/Higher Level/

Question			Answers	Notes	Total
3.	a		H_2O *AND* (l) ✓	*Do not accept* H_2O *(aq).*	1
3.	b		SO_2(g) is an irritant/causes breathing problems *OR* SO_2(g) is poisonous/toxic ✓	*Accept* SO_2*(g) is acidic but do not accept "causes acid rain".* *Accept* SO_2*(g) is harmful.* *Accept* SO_2*(g) has a foul/pungent smell.*	1
3.	c		$n(HCl) = «\frac{10.0}{1000} dm^3 \times 2.00 \, mol \, dm^{-3} =» 0.0200 / 2.00 \times 10^{-2} «mol»$ *AND* $n(Na_2S_2O_3) = «\frac{50}{1000} dm^3 \times 0.150 \, mol \times dm^{-3} =» 0.00750 / 7.50 \times 10^{-3} «mol»$ ✓ $0.0200 «mol» > 0.0150 «mol»$ *OR* $2.00 \times 10^{-2} «mol» > 2 \times 7.50 \times 10^{-3} «mol»$ *OR* $\frac{1}{2} \times 2.00 \times 10^{-2} «mol» > 7.50 \times 10^{-3} «mol»$ ✓	*Accept answers based on volume of solutions required for complete reaction.* *Award* ***[2]*** *for second marking point.* *Do not award M2 unless factor of 2 (or half) is used.*	2

Question			Answers	Notes	Total
3.	d		 five points plotted correctly ✓ best fit line drawn with ruler, going through the origin ✓		2

Question			Answers	Notes	Total
3.	e	i	first order ✓ «because» $[Na_2S_2O_3]$ is «directly» proportional to rate of reaction «$\frac{1}{t}$» ✓	*Do not accept "linear" for M2.*	2
3.	e	ii	rate = $k\,[Na_2S_2O_3][HCl]$ ✓		1
3.	f		 22.5×10^{-3} «s^{-1}» ✓ «Time = $\frac{1}{22.5 \times 10^{-3}}$ =» 44.4 «s» ✓	*Award **[2]** for correct final answer.* *Accept value based on candidate's graph.* *Award M2 as ECF from M1.* *Award **[1 max]** for methods involving taking mean of appropriate pairs of $\frac{1}{t}$ values.* *Award **[0]** for taking mean of pairs of time values.* *Award **[2]** for answers between 42.4 and 46.4 «s».*	2

Question			Answers	Notes	Total
3.	g	i	 correctly labelled axes ✓ peak of T_2 curve lower ***AND*** to the right of T_1 curve ✓	*Accept "probability «density» / number of particles / N / fraction" on y-axis.* *Accept "kinetic E/KE/E_k" but not **just** "Energy/E" on x-axis.*	2

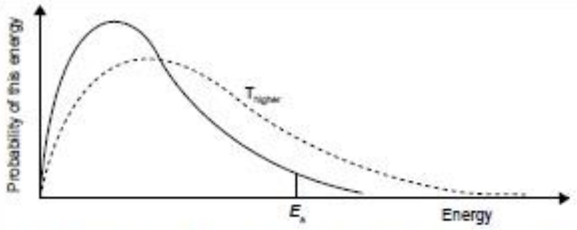
Question			Answers	Notes	Total
3.	g	ii	greater proportion of molecules have $E \geq E_a$ or $E > E_a$ ***OR*** greater area under curve to the right of the E_a ✓ greater frequency of collisions «between molecules» ***OR*** more collisions per unit time/second ✓	*Accept more molecules have energy greater than E_a.* ***Do not** accept just "particles have greater kinetic energy".* *Accept "rate/chance/probability/likelihood" instead of "frequency".* *Accept suitably shaded/annotated diagram.* ***Do not** accept just "more collisions".*	2
3.	h		shorter reaction time so larger «%» error in timing/seeing when mark disappears ✓	*Accept cooling of reaction mixture during course of reaction.*	1

Q# 2/ IB Chem/2016/s/TZoSP/Paper 2 Section A/Higher Level/Q4

	e		*A* is indicative of frequency of collisions and probability that collisions have proper orientations ✓		1
	f		$k = \left\langle \exp\left[\frac{(-87.0\times1000)}{(8.31\times298)} + \ln(2.10\times10^{11})\right] = \right\rangle 1.2\times10^{-4}$ ✓ S_N2 implies second-order so mol^{-1} dm^3 s^{-1} ✓		2

Q# 3/ IB Chem/2016/s/TZ1/Paper 2 Section A/Higher Level/

3.	a	i	$2NO(g) + 2H_2(g) \rightarrow N_2(g) + 2H_2O(g)$ ✓		1
3.	a	ii	rate = $k[NO]^2[H_2]$ ✓		1
3.	a	iii	test the effect «on the reaction rate» of varying each concentration «independently» ***OR*** test the effect of varying [NO] «on rate», whilst keeping $[H_2]$ constant ***AND*** test effect of varying $[H_2]$ «on rate», whilst keeping [NO] constant ✓ rate proportional to $[NO]^2$ ***OR*** doubling [NO] quadruples rate ✓ rate proportional to $[H_2]$ ***OR*** doubling $[H_2]$ doubles rate ✓	*Remember to refer back to a (ii) for ECF.* *If only one species in rate expression, third mark can be awarded for zero order discussion.*	3
3.	a	iv	no ***AND*** different mechanisms could give the same rate expression ***OR*** no ***AND*** mechanisms can only be disproved ***OR*** no ***AND*** just suggest it is consistent with the mechanism given ***OR*** no ***AND*** does not give information about what occurs after RDS ✓		1
3.	a	v	change of pressure «at constant volume and temperature» with time ***OR*** change of volume «at constant pressure and temperature» with time ✓	*Accept other methods where rate can be monitored with time.*	1

Question			Answers	Notes	Total
3.	b	iii	 second curve at a higher temperature is correctly drawn (maximum lower and to right of original) ✓		1
3.	b	iv	greater proportion of molecules have $E \geq E_a$ or $E > E_a$ ***OR*** greater area under curve to the right of the E_a ✓ greater frequency of collisions «between molecules» ***OR*** more collisions per unit time/second ✓	*Do **not** accept just particles have greater kinetic energy.* *Do **not** accept just "more collisions".*	2

Q# 4/ IB CHEM/2015/w/TZ0/Paper 2 Section B/Higher Level/Q7

(c) (i) *Order with respect to* O_2: first (order);
Order with respect to NO: second (order); ***[2]***

(ii) (rate =) $k\,[NO]^2\,[O_2]$; ***[1]***

(iii) 172.6 / 173;
$mol^{-2}\,dm^6\,s^{-1}$; ***[2]***
Accept $M^{-2}\,s^{-1}$/$mol^{-2}\,L^2\,s^{-1}$.

(iv) $NO + NO \rightleftharpoons N_2O_2$ (fast);
$N_2O_2 + O_2 \rightarrow 2NO_2$ (slow);
second step is rate determining step;

OR

$NO + O_2 \rightleftharpoons O + NO_2$ (fast);
$O + NO \rightarrow NO_2$ (slow);
second step is rate determining step;

OR

$NO + O_2 \rightleftharpoons NO_3$ (fast);
$NO_3 + NO \rightarrow 2NO_2$ (slow);
second step is rate determining step;
Accept single arrow instead of reversible sign.
Accept fast/slow next to the equations.

*Do **not** accept the following mechanisms:*
$NO + NO \rightleftharpoons NO_2 + N$ (fast)
$N + O_2 \rightarrow NO_2$ (slow)

OR

$NO + NO + O_2 \rightleftharpoons N_2O_4$ (slow)
$N_2O_4 \rightarrow 2NO_2$ (fast) *[3]*

Q# 5/ IB Chem/2015/w/TZ0/Paper 2 Section A/Higher Level/

(f) gradient $= -\frac{E_a}{R}$ / -1233 (K);
Accept value from –1200 to –1260 (K).

$E_a = (1233 \times 8.31 =)\ 1.02 \times 10^4\ \text{J mol}^{-1}$ / $10.2\ \text{kJ mol}^{-1}$
correct E_a value;
correct units; *[3]*
Allow value in range 9.97 – 10.5 kJ mol^{-1}.
Answer must be given to three significant figures.
*Award [3] for correct final answer **and** units.*
Accept J or kJ instead of J mol^{-1} or kJ mol^{-1}.

Q# 6/ IB Chem/2015/s/TZ2/Paper 2 Section A/Higher Level/

2. (a) change in concentration of reactant/product with time / rate of change of concentration; **[1]**
Accept "increase" instead of "change" for product and "decrease" instead of "change" for reactant.
Accept "mass/amount/volume" instead of "concentration".
Do not accept substance.

(b) surface area decreases;
frequency/probability of collisions decreases; **[2]**
Accept number of collisions per unit time decreases.

(c) (i) step 2 / $X + NO \rightarrow Y + H_2O$ / slow; **[1]**

(ii) invalid / unlikely as order most likely one (with respect to hydrogen);
rate $= k[NO]^2[H_2]$ / H_2 only involved once in the formation of the intermediate before the slow step / *OWTTE*; **[2]**
Award M2 only if M1 is correct.

Q# 7/ IB CHEM/2015/s/TZ1/Paper 2 Section B/Higher Level/Q7

(c) (i) [NaOH] / [OH^-] is 1/first order **and** [C_4H_9Br] is 1/first order;
rate = $k[OH^-][C_4H_9Br]$ / rate = $k[NaOH][C_4H_9Br]$; **[2]**
Square brackets must be used for M2.

(ii) $\left(\frac{1.02\times10^{-4}}{0.25\times0.25}=\right)0.0016 / 1.6\times10^{-3}$;
$mol^{-1}\,dm^3\,s^{-1}$; **[2]**
Accept $M^{-1}\,s^{-1}$.
Ignore order of units.
Must use experiment 3 data.

(iii) bimolecular/2; **[1]**
Accept dimolecular.

Q# 8/ IB Chem/2015/s/TZ1/Paper 2 Section A/Higher Level/

3. (a) change in concentration of reactant/product with time / rate of change of concentration; **[1]**
Accept "increase" instead of "change" for product and "decrease" instead of "change" for reactant.
Accept "mass/amount/volume" instead of "concentration".
Do not accept substance.

(b) collision frequency;
two particles must collide;
particles must have sufficient energy to overcome the activation energy/ $E \geq E_a$;
Concept of activation energy must be mentioned.

appropriate collision geometry/orientation; **[3 max]**

(c) (i) increases yield;
(equilibrium shifts to the right/products as) more gaseous moles in reactants/on left / fewer gaseous moles in products/on right; **[2]**

(ii) Eqm$[O_2] = 2.6(mol\,dm^{-3})$;
Eqm$[SO_2] = 1.2(mol\,dm^{-3})$;
$K_c = \frac{[SO_3]^2}{[SO_2]^2[O_2]}$;
$K_c = 0.17$; **[4]**
Award ***[4]*** *for correct final answer.*
Ignore units.

(iii) (K_c) decreases; **[1]**

(d) catalyst increases rate of reaction / equilibrium reached faster / increases yield of product per unit time;
reduces costs / reduces energy needed; **[2]**
Do not accept just "increases the yield".

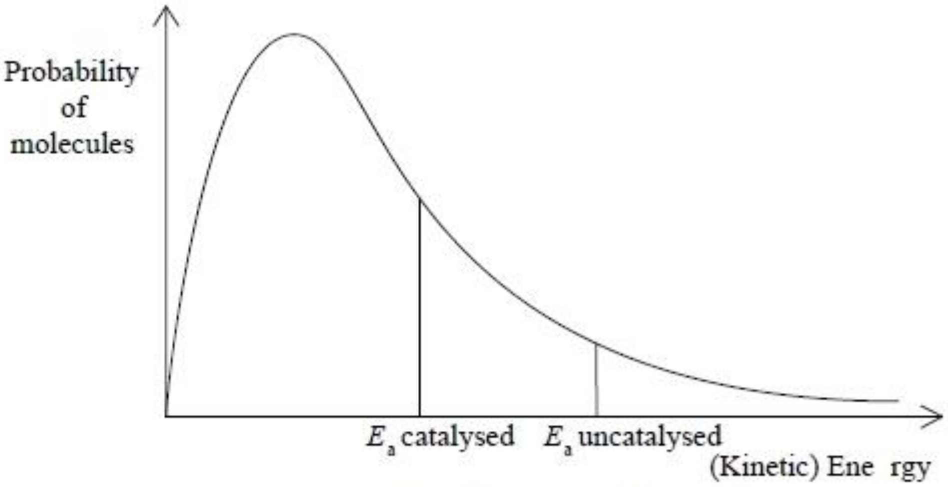
Q# 9/ IB CHEM/2014/w/TZ0/Paper 2 Section B/Higher Level/

11. (a) (i) (draw a) tangent to the curve at origin/time = 0/start of reaction;
(calculate) the gradient/slope (of the tangent); *[2]*

(ii) rate decreases (with time);
concentration/number of (reactant) molecules per unit volume decreases (with time);
Do not accept "number of molecules decreases" or "amount of reactant decreases".

collisions (between reactant molecules/reactant and catalyst) become less frequent; *[3]*
Do not accept "fewer collisions" without reference to frequency (eg, no. collisions per second).

(b) *y-axis*: probability / fraction of molecules/particles / probability density
Allow "number of particles/molecules" on y-axis.
and
x-axis: (kinetic) energy;
Accept "speed/velocity" on x-axis.



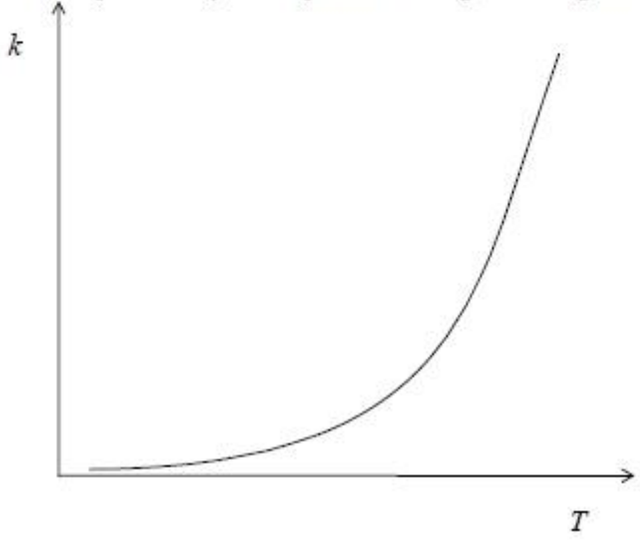
correct relative position of E_a catalysed and E_a uncatalysed;
more/greater proportion of molecules/collisions have the lower/required/catalysed E_a (and can react upon collision); *[3]*
M3 can be scored by stating ***or*** *shading and annotating the graph.*
Accept "a greater number/proportion of successful collisions as catalyst reduces E_a".

(c) (i) reactant not involved in (or before) the slowest/rate-determining step/RDS;
reactant is in (large) excess; *[1 max]*

(ii) (rate =)$k[\text{A}]$; *[1]*
Accept rate $= k[A]^1[B]^0$.

(d) curve with a positive slope curving upwards; [1]

Do not penalize if curve passes through the origin.



Q# 10/ IB CHEM/2014/s/TZ1/Paper 2 Section B/Higher Level/

6. (a) (i)

Liquid	0.500 mol dm^{-3} HCl	0.0200 mol dm^{-3} $Na_2S_2O_3$	Water
Volume / cm^3	10(.0)	20(.0)	20(.0)

; [1]

Accept other volumes in a 1:2:2 ratio.

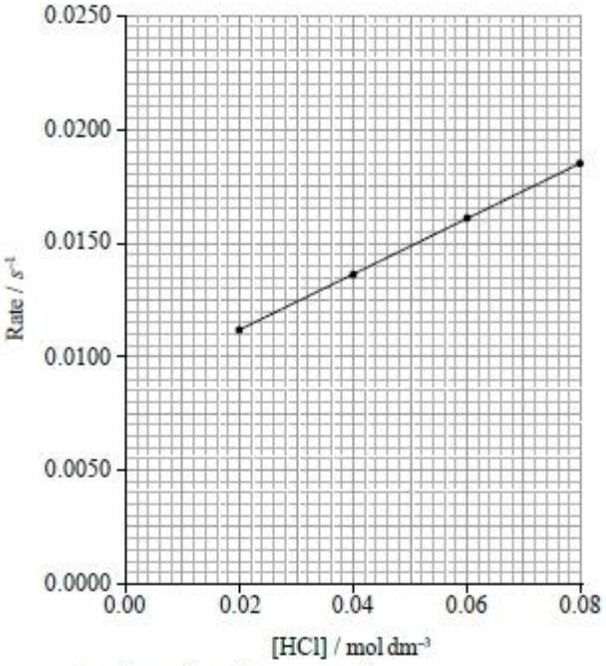
(ii) depth of liquid in the beaker must remain constant / *OWTTE*; [1]

Accept "same thickness of glass" and any other valid point, such as answers framed around minimizing uncontrolled variables / making it a "fair test".

(iii) 94 (s) / 1 min 34 s; [1]

(b) (i) rate $= k[S_2O_3^{2-}][H^+]^2$ / rate $= k[Na_2S_2O_3][HCl]^2$; [1]

(ii)



correct scale and units on *y*-axis;

Accept other suitable scales (such as 1/t) and units (such as ms^{-1}).

Axes do not have to show origin/start at zero.

correct calculation of rate in s^{-1};

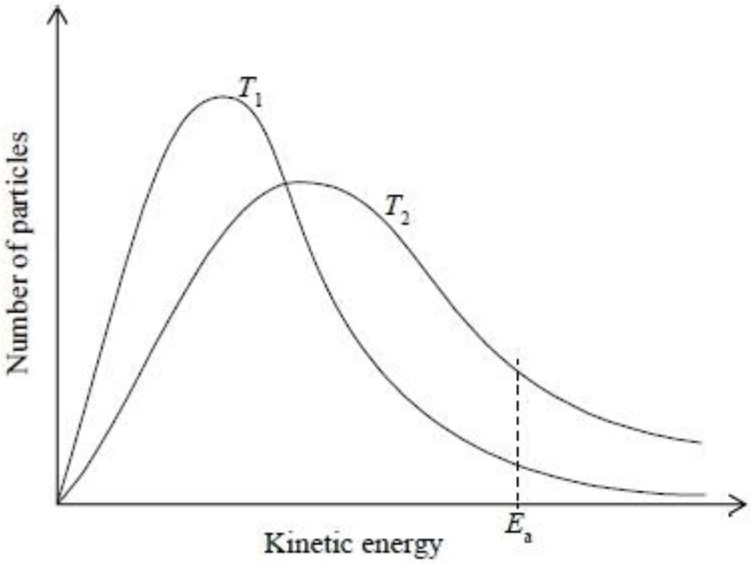
[HCl] / mol dm^{-3}	0.02	0.04	0.06	0.08
Time / s	89.1	72.8	62.4	54.2
Rate / s^{-1}	0.0112	0.0137	0.0161	0.0185

If graph correct, assume this has been done on calculator and not written down.

correct plotting of points that the student decides to use **and** a connecting line; *[3]*
Award final mark if 3 or more points are correct, irrespective of what is plotted on y-axis.
If line goes through the correct values at given concentrations of HCl, assume that points are marked there.

(iii) linear dependence on [HCl] (so not second order in [H$^+$]);
Accept that doubling of concentration does not result in quadrupling of rate / OWTTE .
does not go through origin; *[2]*
Remember to allow ECF from (b) (i).

(c) (i)



labelled y-axis: number of particles / probability of particles (with that kinetic energy) **and** *labelled x-axis:* (kinetic) energy;
Allow fraction/proportion/amount of particles (with kinetic energy) for y-axis label.
Allow speed/velocity for x-axis label.

T_2 curve broader **and** with maximum lower **and** to right of T_1 curve;
Do not award this mark if both curves not asymmetric.
Curves must pass through the origin and be asymptotic to x axis.
Do not award this mark if curves not labelled.

E_a marked on graph; *[3]*

(ii) kinetic energy of molecules increases;
This may be answered implicitly in the final marking point.

frequency of collision/number of collisions per unit time increases;
*Do **not** accept "number of collisions increases".*

greater proportion of molecules have energy greater than/equal to activation energy / rate related to temperature by the Arrhenius equation; *[3]*
Award [1 max] for statements such as "there will be more successful collisions" if neither of last two marking points awarded.

Q# 11/ IB Chem/2014/s/TZ1/Paper 2 Section A/Higher Level/

2. (a) experiments 1 and 2 ([$S_2O_3^{2-}$] remains constant) change in [H^+] does not affect the rate so zero order with respect to H^+(aq) / *OWTTE*;
experiment 1/2 and 3 ([H^+] has no effect) [$S_2O_3^{2-}$] is halved and rate is also halved so first order with respect to [$S_2O_3^{2-}$] / *OWTTE*; *[2]*
Accept explanation given in mathematical terms.
*Award [1 max] if both [$S_2O_3^{2-}$] is first order, **and** [H^+] is zero order are stated without reason.*

(b) rate $= k[S_2O_3^{2-}]$; *[1]*

(c) 0.18;
s^{-1}; *[2]*

(d) $S_2O_3^{2-} \rightarrow S + SO_3^{2-}$; *[1]*
Accept any balanced equation that starts with only one $S_2O_3^{2-}$.
Equations must be balanced in terms of number of atoms and charge.

(e) determine rate at a range of temperatures (while keeping concentrations constant);
calculate k for each temperature;
plot graph of ln k against T^{-1};
gradient is $\frac{-E_a}{R}$ /*OWTTE*; *[3 max]*

Q# 12/ IB Chem/2013/w/TZ0/Paper 2 Section A/Higher Level/

1. (a) KI/I^-/potassium iodide/iodide (ion) (rapidly) reformed (in second stage of reaction); *[1]*

(b) amount (in mol) of H_2O_2/hydrogen peroxide >> amount (in mol) $Na_2S_2O_3$/$S_2O_3^{2-}$ /sodium thiosulfate/ thiosulfate (ion);
Accept amount (in mol) of H_2O_2/hydrogen peroxide >> amount (in mol) KI/I^- /potassium iodide/iodide (ion).
Accept "H_2O_2/hydrogen peroxide is in (large) excess/high concentration".

(at end of reaction) [H_2O_2] is only slightly decreased/virtually unchanged; *[2]*

(c) all $Na_2S_2O_3$/sodium thiosulfate/$S_2O_3^{2-}$/thiosulfate consumed/used up;
Accept "iodine no longer converted to iodide".

(free) iodine is formed / iodine reacts with starch / forms iodine-starch complex; *[2]*

(d) (i) $(5 \times 0.1) = (\pm)0.5\ (\text{cm}^3)$; *[1]*

(ii) $(\pm)0.7(\%)$; *[1]*

Comprises both mass of KI = $\pm 0.5\%$ *and volume of KI =* $\pm 0.2\%$.

(iii) $0.5 + 0.7 = (\pm)1.2\%$; *[1]*

Sum of (i) and (ii) (percentage uncertainty of total volume = absolute uncertainty as 100 cm^3*).*

(e) total volume $= 0.100\,(\text{dm}^3) / 100\,(\text{cm}^3)$;

$$\left(\text{change in concentration} = \frac{1.00 \times 10^{-4}}{0.100} = \right) 1.00 \times 10^{-3}\ (\text{mol dm}^{-3});$$

$$\left(\text{rate} = \frac{1.00 \times 10^{-3}}{45} = \right) 2.2 \times 10^{-5};$$

Award [3] for the correct final answer.

$\text{mol dm}^{-3}\ \text{s}^{-1}$; *[4]*

(f) (i) *x-axis:* $\dfrac{1}{\text{Temperature}} / \dfrac{1}{T} / \text{T}^{-1}$;

Ignore units.

y-axis: ln rate/$\log_e$ rate / ln rate constant/$\log_e$ rate constant / ln k/$\log_e k$; *[2]*

(ii) gradient $= \dfrac{-E_a}{R}$;

$$\text{gradient} = \frac{-4.00}{(3.31 \times 10^{-3} - 2.83 \times 10^{-3})} = -8333 \;/\; = \frac{-4.80}{(3.41 \times 10^{-3} - 2.83 \times 10^{-3})} = -8276;$$

$$E_a = \left(\frac{8.31 \times 8333}{1000}\right) = 69.3\,(\text{kJ mol}^{-1}) \;/\; = \left(\frac{8.31 \times 8276}{1000}\right) = 68.8\,(\text{kJ mol}^{-1});$$ *[3]*

Award [3] for correct final answer.
Accept values from 65.0 to 73.0 kJ mol^{-1}.
Deduct [1] for final answer in J mol^{-1}.
Deduct [1] for final answer not to 3 significant figures.

(g) acting as a catalyst / black powder reacts with thiosulfate ions / solid dissolves to give blue-black solution;

Accept any other valid suggestion which will make colour change more rapid.

For catalyst: amount/mass of black powder remains constant / no new/different products formed / activation energy decreased;
For other suggestions: any appropriate way to test the hypothesis; *[2 max]*
Award [1] for valid hypothesis, [1] for appropriate method of testing the stated hypothesis.

Q# 13/ IB Chem/2013/s/TZ2/Paper 2 Section A/Higher Level/

2. (a) minimum energy needed (by reactants/colliding particles) to react/start/initiate a reaction / for a successful collision; *[1]*
Allow energy difference between reactants and transition state.

(b) *k* increases with *T*; *[1]*
Do not accept k proportional to T or statement of Arrhenius equation from Data booklet.

(c) slope/gradient/$m = \frac{-E_a}{R}$ / -6.20×10^3;
Allow range of m from -5.96×10^3 *to* -6.44×10^3.
Award M1 for $m = \frac{-E_a}{R}$ *even if gradient is out of range.*

$E_a = (6.20\times10^3 \times 8.31) = 51.5\,\text{kJ mol}^{-1}$ / $5.15\times10^4\,\text{J mol}^{-1}$
E_a value correct;
units correct; *[3]*
Award [3] for correct final answer.
Allow range of E_a *from 49.5 to 53.5* kJ mol^{-1}/4.95×10^4 *to* $5.35\times10^4\,\text{J mol}^{-1}$.
Answer must be given correct to three significant figures.
M3 can be scored independently.

Q# 14/ IB Chem/2013/s/TZ1/Paper 2 Section B/Higher Level/Q9

(c) (i) *C_4H_9Br:*
[C_4H_9Br] doubles **and** time halves/rate doubles/rate proportional to [C_4H_9Br];
Do not accept rate increases when [C_4H_9Br] increases.

NaOH:
[NaOH] doubles **and** time/rate does not change/rate independent of [NaOH];

C_4H_9Br: first order **and** *NaOH:* zero order; *[3]*

(ii) rate $= k[C_4H_9Br]$; *[1]*
Accept ECF.

(iii) s^{-1}; *[1]*
Accept ECF.

Q# 15/ IB Chem/2013/s/TZ1/Paper 2 Section A/Higher Level/

3. (a) minimum energy needed to react/start a reaction / energy difference between reactants and transition state; *[1]*

(b) *gradient of the line: –63;*
Accept –60 to –65.

$E_a (= -R\times \text{gradient}) = 0.52\,(\text{kJ mol}^{-1})$; *[2]*
Accept 0.50 to 0.54.

(c) gradient of the line less steep (less negative); *[1]*
Accept any position as long as gradient less steep.

Q# 16/ IB Chem/2012/w/TZ0/Paper 2 Section B/Higher Level/

6. (a) (i) change in concentration of reactant/product with time / rate of change of concentration; *[1]*

Increase can be used instead of change for product or decrease can be used instead of change for reactant.
Allow mass/amount/volume instead of concentration.
Do not accept substance.

(ii) concentration;
particle size / surface area;
light;
pressure; *[2 max]*
Allow pH.

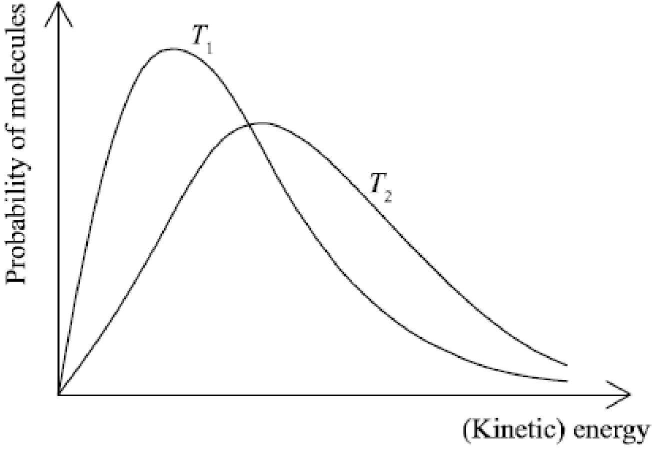
(iii) (measuring electrical) conductivity / (measuring) pH; *[1]*
Accept other suitable method.

(b) (i) minimum/least/smallest energy needed (by reactants/colliding particles) to react/start/initiate a reaction; *[1]*
Allow energy difference between reactants and transition state.
Minimum/least/smallest required for the mark.

(ii) *x-axis label*: (kinetic) energy/(K)E **and** *y-axis label*: probability/fraction of molecules/particles / probability density;
Allow number of molecules/particles for y-axis.

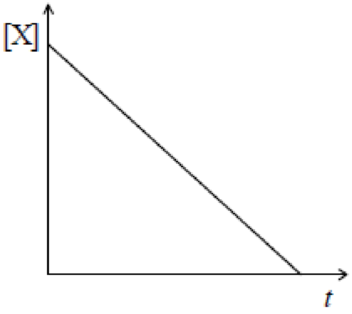
correct shape of a typical Maxwell–Boltzmann energy distribution curve;
Do not award mark if curve is symmetric, does not start at zero or if it crosses x-axis.

two curves represented with second curve for $T_2 > T_1$ to right of first curve, peak maximum lower than first curve and after the curves cross going to the right, T_2 curve needs to be above T_1 curve as illustrated;
M2 and M3 can be scored independently.



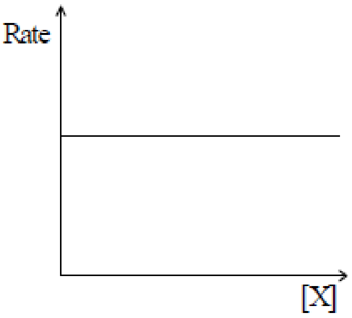
[3]

(c) (i)



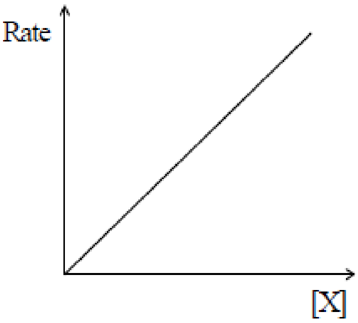
; **[1]**

(ii)



; **[1]**

(iii)



; **[1]**

(d) (i) second order in ClO_2 **and** first order in OH^-;

rate $= k[ClO_2]^2[OH^-]$; **[2]**

*Award **[2]** for correct final answer.*

(ii) $k = 2.30 \times 10^2 / 230$;

$mol^{-2}\ dm^6\ s^{-1}$; **[2]**

(iii) $1.22 \times 10^{-3} / 0.00122\ (mol\ dm^{-3}\ s^{-1})$; **[1]**

(e) (i) ethyl ethanoate; **[1]**

Do not allow ethyl acetate.

(ii) as temperature/T increases, (value of) rate constant/k increases (exponentially); **[1]**

Do not allow answers involving ln k from the Arrhenius equation.

Do not allow T directly proportional to k.

(iii) slope $= -5.6 \times 10^3 / -5600$ (K);
$E_a = -\text{slope} \times R$ / slope $= -E_a / R$;
$E_a (= 5.60 \times 10^3 K \times 8.31 J K^{-1} mol^{-1}) = 4.65 \times 10^4 (J mol^{-1})$ / 46.5 (kJ mol^{-1});
Accept answers in range $4.60 \times 10^4 J mol^{-1}$ *to* $4.67 \times 10^4 J mol^{-1}$.

J mol^{-1}/ kJ mol^{-1}; *[4]*
Accept J or kJ.
Unit mark can be scored independently but correct E_a *values with incorrect units scores only* ***[3 max]*** *(for example 46.5 J* mol^{-1}*).*

Award [4] for correct final answer.

(f) (i) $3ClO^-(aq) \rightarrow ClO_3^-(aq) + 2Cl^-(aq)$; *[1]*
Ignore state symbols.

(ii) *Step 1*: rate $= k[ClO^-]^2$;
Step 2: rate $= k[ClO_2^-][ClO^-]$; *[2]*
Penalize missing k once only.

Q# 17/ IB Chem/2012/w/TZ0/Paper 2 Section A/Higher Level/Q1a

(iv) negative/–/minus/ < 0; *[1]*

Q# 18/ IB Chem/2011/w/TZ0/Paper 2 Section B/Higher Level/Q8

(e) (i) dative (covalent)/coordinate;
carbon monoxide/CO / hydronium (ion)/ H_3O^+ / ammonium (ion)/ NH_4^+ / aluminium chloride/Al_2Cl_6 / any relevant transition metal complex (*e.g.* $[Ni(NH_3)_6]^{2+}$); *[2]*
Accept $AlCl_3$*.*

(ii) rate $= k[BF_3][NH_3]$;
second (order)/2°;
$k = 3.40$ dm^3 mol^{-1} s^{-1}; *[3]*
Allow units of L mol^{-1} s^{-1} *or* M^{-1} s^{-1}.
Units required for mark.

(f) (i) N_2O_2; *[1]*

(ii) ($[H_2]$ appears in rate expression so) step 2 rate-determining/rds/slow step; *[1]*
Allow "since step 1 involves 2NO and step 2 involves H_2 *and as all 3 molecules are involved in rate expression, then two steps must have approximately same rate" / OWTTE.*

(g) ($k_2 >> k_1$ so) step 1 rate-determining/rds/slow step;
two molecules of NO_2 involved in step 1 consistent with rate expression / rate of overall reaction must equal rate of step 1 which is rate $= k_1 [NO_2]^2$ / *OWTTE;* *[2]*

(h) $E_a = -R \times m$;

measurement of gradient from **two** points on line;

Accept a gradient in range -2.14×10^4 *K to* -2.27×10^4 (*K*).

correct answer for E_a;

correct units kJ mol^{-1}/J mol^{-1} corresponding to answer; *[4]*

Allow kJ or J.

A typical answer for $E_a = 1.85 \times 10^2$ *kJ* mol^{-1}.

Allow answers for E_a *in range* 1.75×10^2 *kJ* mol^{-1} *to* 1.91×10^2 *kJ* mol^{-1}.

Award [4] for correct final answer with some working shown.

*Award **[2 max]** for correct final answer without any working shown.*

Q# 19/ IB Chem/2011/w/TZ0/Paper 2 Section B/Higher Level/

8. (a) minimum energy needed (by reactants/colliding particles) to react/start/initiate a reaction; *[1]*

Allow energy difference between reactants and transition state.

(b) catalyst;

regenerated at end of reaction / *OWTTE*; *[2]*

(c) (i) (system) absorbs/takes in heat from surroundings / *OWTTE*; *[1]*

Allow standard enthalpy change/ $\Delta H^{\ominus}$ *positive.*

Allow bond breaking more energetic then bond formation / OWTTE.

Absorbs/takes in heat alone not sufficient for mark.

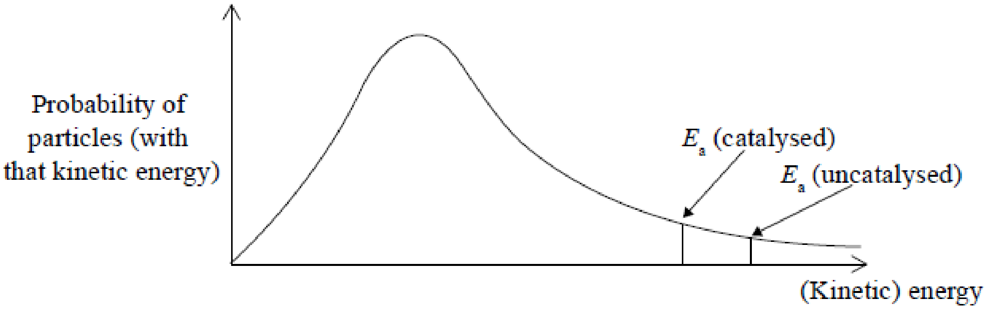
(ii) *Curve showing:*

general shape of Maxwell-Boltzmann energy distribution curve;

correct position of E_a (catalysed) **and** E_a (uncatalysed);

labelled y-axis: probability of particles (with that kinetic energy) **and** *labelled x-axis:* (kinetic) energy; *[3]*

Allow number/fraction/proportion of particles (with kinetic energy) for y-axis label, but do not allow amount or particles.



*Award **[2 max]** if a second curve is drawn, but at a higher temperature, M2 will not be scored here.*

(d) (i) change in concentration of reactant/product with time / rate of change of concentration; *[1]*

Increase can be used instead of change for product or decrease can be used instead of change for reactant.

Allow mass/amount/volume instead of concentration.

Do not accept substance.

Q# 20/ IB Chem/2011/w/TZ0/Paper 2 Section A/Higher Level/Q1

(d) (i) $\left(\frac{34.1}{0.0400}\right) = 853\ \text{dm}^3\ \text{s}^{-1} / \left(\frac{1.50}{0.0400}\right) = 37.5\ \text{mol s}^{-1}$; **[1]**

Accept 851 $dm^3\ s^{-1}$.

Units required for mark.

(ii) more energetic collisions / more species have energy $\geq E_a$; **[1]**

Allow more frequent collisions / species collide more often.

Q# 21/ IB Chem/2011/s/TZ1/Paper 2 Section A/Higher Level/

2. (a) k increases with increase in T / k decreases with decrease in T; **[1]**

Do not allow answers giving just the Arrhenius equation or involving lnk relationships.

(b) gradient $= -E_a/R$;

$-30000\ (\text{K}) = -E_a/R$;

Allow value in range –28800–31300 (K).

$E_a = (30000 \times 8.31 =)\ 2.49 \times 10^5\ \text{J mol}^{-1} / 249\ \text{kJ mol}^{-1}$; **[3]**

Allow value in range 240–260 kJ mol^{-1}.

Allow [3] for correct final answer.

(c) $0.9 \times 0.200 = 0.180\ (\text{mol dm}^{-3})$;

rate $= (0.244 \times (0.180)^2 =) 7.91 \times 10^{-3}\ \text{mol dm}^{-3}\ \text{s}^{-1}$; **[2]**

*Award **[2]** for correct final answer.*

*Award **[1 max]** for either 9.76×10^{-3} mol dm^{-3} s^{-1} or 9.76×10^{-5} mol dm^{-3} s^{-1}.*

Q# 22/ IB Chem/2010/s/TZ1/Paper 2 Section A/Higher Level/

2. (a) to maintain a constant volume / *OWTTE*; **[1]**

(b) (i) $[H^+]$ order 1, $[CH_3COCH_3]$ order 1, $[I_2]$ order 0;

(rate $=$) $k\ [H^+]\ [CH_3COCH_3]$; **[2]**

*Award **[2]** for correct rate expression.*

Allow expressions including $[I_2]^0$.

(ii) neither were correct / Alex was right about propanone and wrong about iodine / Hannah was right about propanone and hydrogen ions but wrong about iodine / *OWTTE*; **[1]**

(c) $[CH_3COCH_3] = 0.100\ \text{mol dm}^{-3}$ **and** $[H^+] = 0.100\ \text{mol dm}^{-3}$;

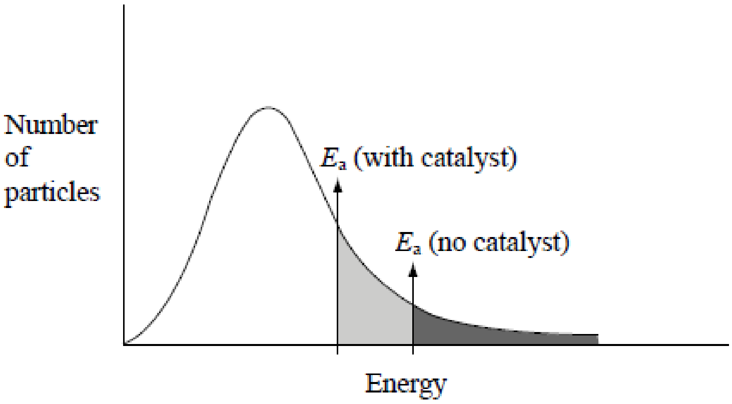
$k = \frac{4.96 \times 10^{-6}}{(0.100 \times 0.100)} = 4.96 \times 10^{-4}$;

$\text{mol}^{-1}\ \text{dm}^3\ \text{s}^{-1}$; **[3]**

Ignore calculation of $[I_2]$.

No ECF here for incorrect units.

(d) (i)



axes correctly labelled x = energy/velocity/speed, y = number/% of molecules/particles/probability;
graph showing correct curve for Maxwell-Boltzmann distribution;
If two curves are drawn, first and second marks can still be scored, but not third.
Curve(s) must begin at origin and not go up at high energy.

two activation energies shown with E_{cat} shown lower; *[3]*
Award the mark for the final point if shown on an enthalpy level diagram.

(ii) catalyst provides an alternative pathway of lower energy / *OWTTE*; *[1]*
Accept catalyst lowers activation energy (of reaction).

Q# 23/ IB Chem/2009/w/TZ0/Paper 2 Section B/Higher Level/

6. (a) (i) rate $= k[\text{NO}]^2[\text{Cl}_2]$; *[1]*

(ii) rate of reaction will decrease by a factor of 4;
no effect on the rate constant; *[2]*

(b) Above 775 K: rate $= k[\text{NO}_2][\text{CO}]$;
Below 775 K: rate $= k[\text{NO}_2]^2$; *[2]*

(c) zero order reaction;
all concentrations are $1.0\ \text{mol dm}^{-3}$; *[2]*

(d) slope $= \dfrac{9.2 - 8.4}{(3.53 - 3.65) \times 10^{-3}} = -6.67 \times 10^3$;
$(E_a = 6.67 \times 10^3 \times 8.31)$
$55.4\ (\text{kJ mol}^{-1})$; *[2]*
Accept in range 55.0 – 56.0
Award [1] if 55454 (J) stated
Award [2] for the correct final answer

Q# 24/ IB Chem/2008/w/TZ0/Paper 2 Section A/Higher Level/

1. (a) (i) gas/carbon dioxide/CO_2 is given off/evolves/escapes/formed; *[1]*

(ii) $\text{n(CaCO}_3)\left(= \dfrac{5.00}{100.09}\right) = 0.0500$;

$\text{n(HCl)}\left(= \dfrac{1.00 \times 50.0}{1000}\right) = 0.0500$; *[2]*

Do not penalize significant figures.
No penalty for using 100 instead of 100.09.

(iii) $CaCO_3$ because twice as much acid needed / *OWTTE*; *[1]*
Allow HCl is limiting.
Allow HCl in excess as correct deduction from wrong values in (a)(ii).

(iv) (line 2) steper;
levels off at same mass as Experiment 1;
Apply ECF from part (ii).

(line 3) steeper;
levels off halfway between zero and Experiment 1; *[4]*

Allow ECF for lines on graph from HCl in excess in (a)(iii). This means that the lines are the same but numbered the other way round.
If lines are unlabelled allow [2 max].

(b) (i) order wrt A is zero because increasing/doubling [A] has no effect on rate / *OWTTE*;
order wrt B is 2 because doubling [B] quadruples rate/halving [B] quarters rate / *OWTTE*; *[2]*
If explanations missing or incorrect but both orders correct, award [1] mark.
Accept other mathematical explanations.

(ii) rate $= k[\mathrm{B}]^2$ / rate $= k[\mathrm{A}]^0[\mathrm{B}]^2$; *[1]*
ECF from (b)(i)

(iii) $k = \dfrac{8.8\times10^{-4}}{(7.8\times10^{-2})^2} = 0.14$;

$\mathrm{mol^{-1}\ dm^3\ min^{-1}}$; *[2]*
ECF from (b)(ii)

Q# 25/ IB Chem/2008/s/TZ1/Paper 2 Section B/Higher Level/

7. (a) (i) (exp. 1 and 2, [NO] constant, [O_2] doubled rate doubles) first order with respect to O_2;
(exp. 2 and 3, [O_2] doubled, [NO] doubled, rate increases by a factor of 8) second order with respect to NO; *[2]*
Accept alternative mathematical method

(ii) rate $= k[\mathrm{NO}]^2[\mathrm{O_2}]$; *[1]*
Allow ECF from (i).

(iii) $\left(k = \dfrac{3.75\times10^{-3}}{(3.50\times10^{-2})^2(1.75\times10^{-2})}\right)$;
$=1.75\times10^2$;
$\mathrm{dm^6\ mol^{-2}s^{-1}}$; *[2 max]*
Allow ECF from (ii).

(iv) 9.38×10^{-4} ($\mathrm{mol\ dm^{-3}\ s^{-1}}$);
(since the volume is doubled), concentration is halved; *[2]*
Allow ECF from (ii).

(v) $NO + NO \rightleftharpoons N_2O_2$ (fast);
$N_2O_2 + O_2 \rightarrow 2NO_2$ (slow);
second step is rate-determining step/rds;

OR

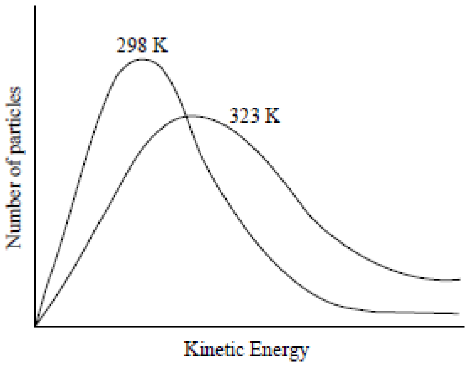
$NO + O_2 \rightleftharpoons NO_3$ (fast);
$NO_3 + NO \rightarrow 2NO_2$ (slow);
second step is rate determining step/rds; *[3 max]*

Do not allow the following 2 mechanisms.

$NO + NO \rightleftharpoons NO_2 + N$ slow
$N + O_2 \rightarrow NO_2$ fast

$NO + NO + O_2 \rightarrow N_2O_4$ slow
$N_2O_4 \rightarrow 2NO_2$ fast

(b) kinetic energy of molecules increases;
frequency of collision increases;
greater proportion of molecules have energy greater than/equal to activation energy;
correct Boltzmann-energy distribution curves showing curve at higher temperature on the right side;
broadening of the curve; *[5]*



Maxwell-Boltzmann distribution curve

Q# 26/ IB Chem/2007/w/TZ0/Paper 2 Section B/Higher Level/

8. (a) (i) $\text{rate} = k[CH_3COCH_3][H^+]$;
$\text{rate} = k[CH_3COCH_3]^2$;
$\text{rate} = k[H^+]^2$; *[3]*

(ii) $[CH_3COCH_3]$ doubles, rate doubles **and** $[H^+]$ doubles, rate doubles;
$[Br_2]$ double, no effect on rate;

OR

$[CH_3COCH_3]$ doubles, rate quadruples;
$[Br_2]$ doubles / $[H^+]$ doubles, no effect on rate;

OR

$[H^+]$ doubles, rate quadruples;
$[Br_2]$ doubles / $[CH_3COCH_3]$ doubles, no effect on rate; *[2]*

The answer given must correspond to the selected expression in 8(a)(i).

(iii) constant half-life;
at least two sets of data to justify statement; *[2]*
e.g. [ ] from 1.6 to 0.8 mol dm^{-3}, 10 s; 0.8 to 0.4, 10 s; 0.4 to 0.2, 10 s.

(iv) decrease in the colour of the bromine / *OWTTE*;
catalyst;
increases rate / speeds up reaction;
by lowering E_a / activation energy (by providing an alternate pathway); *[4]*

Q# 27/ IB Chem/2007/s/TZ0/Paper 2 Section A/Higher Level/

1. (a) *order of NO*: second/2 - [NO] doubled, rate ×4 / *OWTTE*;
order of Br_2: first/1 - as $[Br_2]$ doubled, rate of reaction doubled / *OWTTE*; *[2]*
Reason needed for each mark.

(b) rate = $k[\text{NO}]^2[\text{Br}_2]$; *[1]*
Allow ECF from (a).

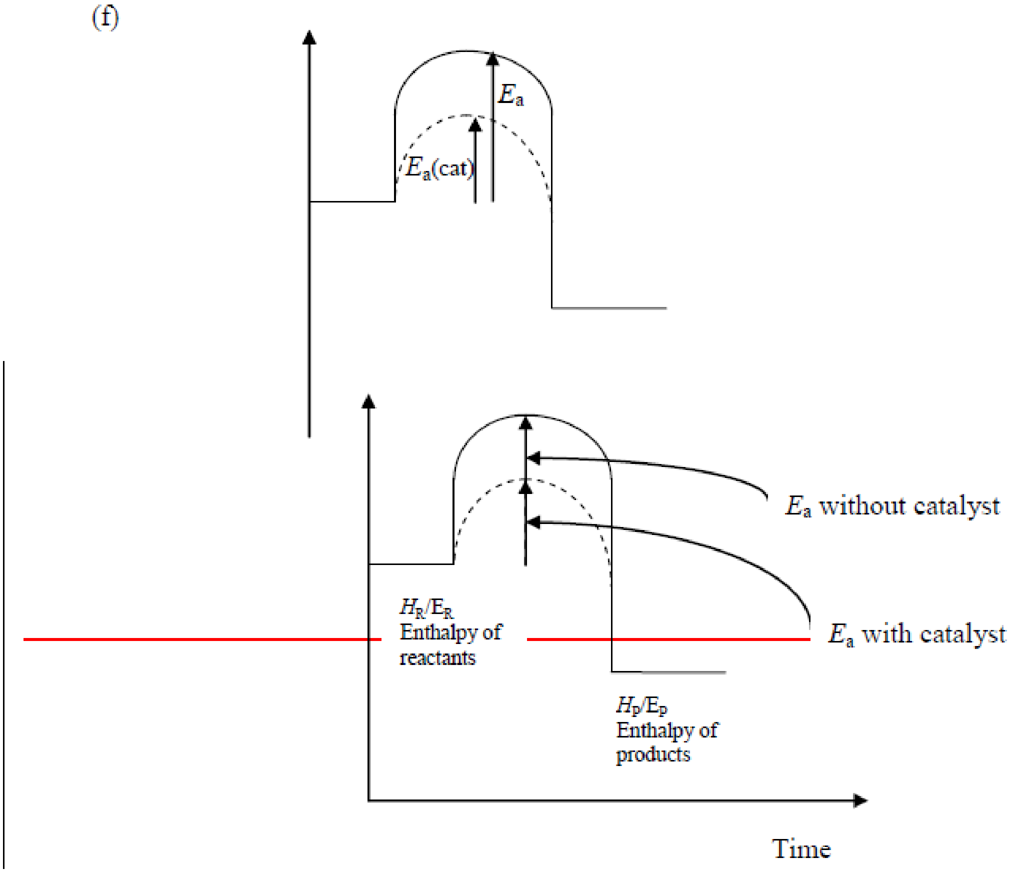
(c) $3.20\times10^{-3} = k\ (2.00\times10^{-2})^2\times5.00\times10^{-3}$
$k = 1.60\times10^{3}$;
$\text{dm}^6\ \text{mol}^{-2}\ \text{s}^{-1}$; *[2]*
Allow ECF from (b).

(d) (i) no effect / K changes only with temperature / OWTTE; *[1]*

(ii) decrease (by a factor of 2); *[1]*

(e) gases adsorb on surface of catalyst / provides surface for the reaction / OWTTE;
lowers activation energy / provides alternative pathway for reaction; *[2]*

(f)



curve clearly showing E_a without catalyst (E_a);
curve clearly showing E_a with catalyst (E_a(cat));

labelling for x axis; *[3]*
Accept time / progress of reaction / course of reaction / OWTTE.

*Award **[2 max]** if an enthalpy level diagram for an endothermic reaction has been correctly drawn.*

Q# 28/ IB Chem/2006/w/TZ0/Paper 2 Section B/Higher Level/

9. (a) (i) $2\text{ICl} + \text{H}_2 \rightarrow \text{I}_2 + 2\text{HCl}$; *[1]*

(ii) *ICl order*
1;
because doubling [ICl] doubles rate (when $[\text{H}_2]$ constant);

H_2 order
2;
because halving $[\text{H}_2]$ quarters rate (when [ICl] constant);
or doubling $[\text{H}_2]$ quadruples rate (when [ICl] constant); *[4]*

(iii) rate $= k[\text{ICl}][\text{H}_2]^2$; *[1]*
ECF from (ii).

(iv) $k = 5.00\times10^{-3} \div 0.100\times0.0500^2 = 20$;
$\text{mol}^{-2}\ \text{dm}^6\ \text{s}^{-1}$; *[2]*
ECF from (iii).

(v) rate $= 20\times0.200\times0.100^2 = 4.00\times10^{-2}$ ($\text{mol dm}^{-3}\ \text{s}^{-1}$); *[1]*
ECF from (iii).

(b) (i) $\text{P} + \text{Q} \rightarrow \text{PQ}$ and $\text{PQ} + \text{Q} \rightarrow \text{R}$;
1st step identified as rate-determining; *[2]*
Accept other symbol such as X instead of PQ.

(ii) 2 / bimolecular; *[1]*

(c) 4 minutes; *[1]*

(d) (i) it relates to the geometric requirements of the reaction / orientation of reactants on collision / *OWTTE*; *[1]*

(ii) <u>minimum</u> energy needed for reactants to react (on collision) / *OWTTE*; *[1]*

(iii) k measured at different values of temperature;
graph plotted of ln k against $1/T$;
intercept on y-axis is ln A;
$A = e^{\text{intercept}}$;
measured slope of graph $= -E_a/\text{R}$;
$E_a = -\text{R}\times$gradient; *[5 max]*
Award [1] each for any five.

Q# 29/ IB Chem/2006/s/TZ0/Paper 2 Section B/Higher Level/Q6

(c) more energetic collisions / more molecules have energy greater than activation energy;
more frequent collisions; *[2]*
Do not accept more collisions without reference to time.

Q# 30/ IB Chem/2005/w/TZ0/Paper 2 Section A/Higher Level/

1. (a) time for reactant concentration to halve / *OWTTE*; *[1]*

Accept "time for mass to halve".

(b) 1000 s;

1000 s;

Accept 900-1100 s. *[2]*

(c) first order;

constant half-life;

rate = k[A]; *[3]*

Allow ECF for rate expression from stated order.

(d) $k = \frac{\text{rate}}{[\text{D}]^2[\text{E}]} \,/\, \frac{3.75\times10^{-5}}{(1.35\times10^{-2})^3}$;

= 15.2;

Accept answer in range 15.2 to 15.3.

$\text{mol}^{-2}\ \text{dm}^6\ \text{min}^{-1}$; *[3]*

Q# 31/ IB Chem/2005/s/TZ0/Paper 2 Section A/Higher Level/

3. (a) the <u>particles/molecules</u> of ammonia gas are in rapid/random/constant motion;

and will diffuse/spread out / *OWTTE*; *[2]*

(b) less time;

(the particles/molecules of ammonia gas will have) greater velocity/greater kinetic energy/greater rate of diffusion/move faster; *[2]*

Do not accept "greater energy". Answer must indicate increased movement.

Q# 32/ IB Chem/2005/s/TZ0/Paper 2 Section A/Higher Level/

4. (a) rate $= k[\text{N}_2\text{O}_5]$ / $k[\text{N}_2\text{O}_5]^1$; *[1]*

(b) (i) Rate $\times[\text{N}_2\text{O}_5]$

One molecule of $[\text{N}_2\text{O}_5]$; *[1]*

(ii) Rate $\times[\text{N}_2\text{O}_5]\,[\text{NO}_3]$

But NO_3 comes from N_2O_5 in first step

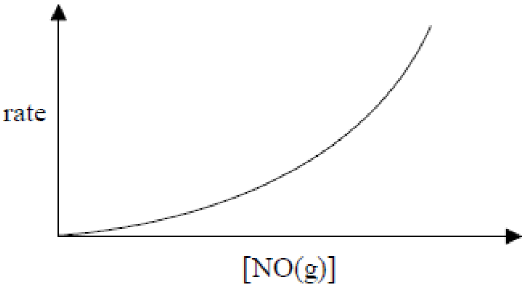
So:Rate $\times[\text{N}_2\text{O}_5]^2$ *[2]*

(c) the time taken for the concentration/amount (of N_2O_5) to decrease to half its original value / *OWTTE*; *[1]*

(d) it is constant / does not depend on concentration (of reactant); *[1]*

8. (a) 1 / first order;
rate is (directly) proportional to concentration of oxygen / *OWTTE*; *[2]*

(b)



correct axes;
correct shape curve; *[2]*

(c) 3 / third order; *[1]*
Allow ECF from (a) and (b).

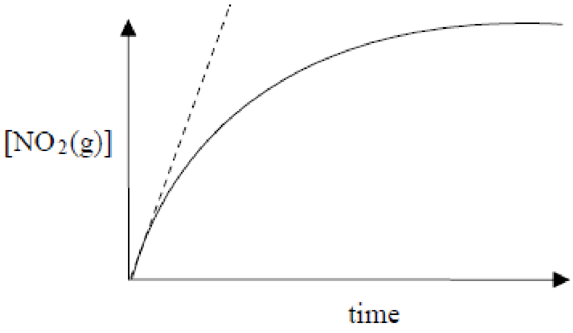
(d) overall effect on rate $= 4 \times \frac{1}{2}$ / doubled / $\times 2$;
$[NO(g)]$ doubled, rate $= \times 4$ / quadrupled;
$[O_2(g)]$ halved, rate $= \times 1$ / halved; *[3]*
Allow ECF from (a) and (b).

(e) $\text{rate} = k[NO(g)]^2\ [O_2(g)]$;

$$k = \frac{\text{rate}}{[NO(g)]^2\ [O_2(g)]} = \frac{6.3 \times 10^{-4}}{(3.0 \times 10^{-2})^2 (1.0 \times 10^{-2})};$$

$= 70$;
$\text{mol}^{-2}\ \text{dm}^6\ \text{s}^{-1}$; *[4]*
Allow ECF.
State symbols not needed.

(f) (i)



correct axes;
curve starts at origin and levels off;
tangent to curve at origin; *[3]*

(ii) second order with respect to NO;
(Expt 1+2) double [NO(g)] only and rate quadruples;
zero order with respect to CO;
(Expt 2+3) double [CO(g)] only and rate is unchanged;
zero order with respect to O_2;
(Expt 2+4) double both [NO(g)] and $[O_2(g)]$ and rate quadruples; *[6]*

(g) stoichiometric equation gives no indication of the reaction mechanism / *OWTTE*; *[1]*

(h) (i) time taken for the concentration of a reactant to fall to half of its initial value;
half-life of a first order reaction is independent of the original concentration; *[2]*

(ii) $t_{\frac{1}{2}} = \frac{0.693}{k} = \frac{0.693}{440} = 1.58 \times 10^{-3}\ (\text{s})$; *[1]*

Q# 34/ IB Chem/2004/s/TZ0/Paper 2 Section A/Higher Level/

6. (a) molecules move from ice to water and water to ice / $H_2O(s) \rightleftharpoons H_2O(l)$ / *OWTTE*
mentioning particles / molecules;

at the same rate; *[2]*

(b) molecules leave skin surface / evaporate / intermolecular forces are overcome on evaporation;
causing cooling effect / heat taken from skin / endothermic process; *[2]*

Q# 35/ IB Chem/2004/s/TZ0/Paper 2 Section A/Higher Level/

1. (a) (order with respect to) NO = 2;
(order with respect to) $H_2 = 1$;
rate increases ×4 when [NO] doubles / *OWTTE*; *[3]*

(b) $\text{rate} = k[\text{NO}]^2[\text{H}_2]$; *[1]*
ECF from (a).

(c) $(2.53 \times 10^{-6}\ \text{mol dm}^{-3}\ \text{s}^{-1} = k\ (0.100\ \text{mol dm}^{-3})^2 (0.100\ \text{mol dm}^{-3}))$
$k = 2.53 \times 10^{-3}$; *[1]*
$\text{mol}^{-2}\ \text{dm}^{6}\ \text{s}^{-1}$; *[1]*
ECF from (b).

(d) agrees / yes;

slow step depends on X and NO;
X depends on H_2 and NO;
(so) NO is involved twice and H_2 once;
Overall equation matches the stoichiometric equation;
Award [1] each for any three of the four above.

OWTTE
ECF for "no", depending on answer for 1 (b). *[4 max]*

Or agrees / yes;

and $\frac{[X]}{[H_2][NO]}$ = constant;

rate of slow step $= k[X][NO]$

$= k[H_2][NO]^2$;

ECF for "no", depending on answer for 1 (b). *[4]*

(e) reaction involves <u>four</u> molecules;
statistically / geometrically unlikely. *[2]*

(f) the rate of formation of $H_2O(g) = 2 \times$ rate for $N_2(g)$;
because 2 moles H_2O formed with 1 mole N_2 / *OWTTE;* *[2]*

Q# 36/ IB Chem/2003/w/TZ0/Paper 2 Section B/Higher Level/

8. (a) (i) [NO] constant, $[O_2]$ doubles, rate doubles;
rate $\propto [O_2]$ / first order;
[NO] doubles, $[O_2]$ doubles, rate increases 8 times;
rate $\propto [NO]^2$ / second order;
rate = k $[NO]^2 [O_2]$ / rate $\propto [NO_2]^2[O_2]$; *[5]*

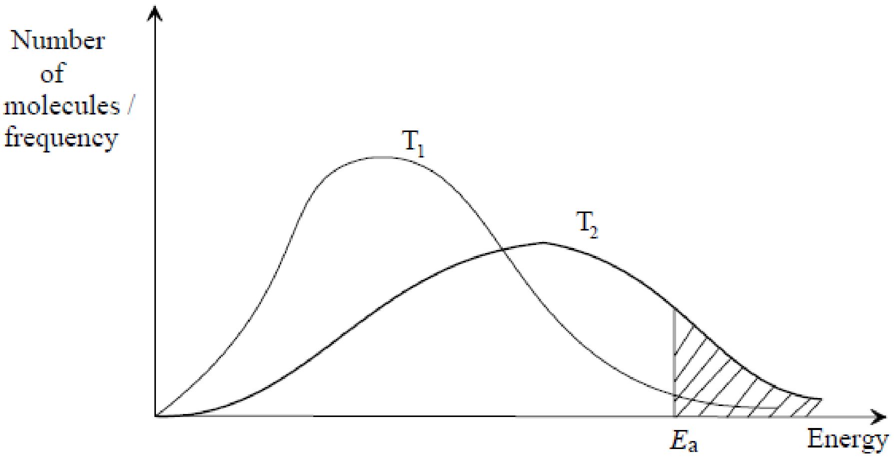
(ii) slow step / rate determining step involves only one NO and one O_2;
not two NO and one O_2 (as required by rate-expression); *[2]*

(iii) since two NO and one O_2 involved in the (one step) mechanism, correct rate expression possible;
but unlikely that three particles will collide at the same time; *[2]*

(iv) from fast step $[NO_3]$ depends on [NO] and $[O_2]$;
rate depends on $[NO_3]$ and [NO];
thus rate must depend on $[NO]^2[O_2]$ (consistent with rate expression); *[3]*

(v) $NO(g) + NO(g) \rightarrow N_2O_2(g)$ fast;
$N_2O_2(g) + O_2(g) \rightarrow 2NO_2(g)$ slow;
similar reasoning to above / rate depends on $[N_2O_2]$ and $[O_2]$, and $[N_2O_2]$ depends on [NO] and [NO] (thus consistent with the overall third order reaction); *[3]*
Accept ECF from (ii) to (v).

6. (a) (i)



both axes correctly labelled;
T_2 peak / lower;
and to right of T_1; *[3]*
Area under graph is not important.

(ii) rate increased / changes;
as more molecules with ≥ E_a; *[2]*
No explicit reference to graph required.

(b) (i) rate increased / changes;
activation energy / E_a lowered;
catalyst provides alternative route / more molecules have $E \geq E_a$; *[3]*

(ii) heterogeneous: different state / phase from reactants;
homogeneous: same state as reactants; *[2]*

(iii) any suitable example, *e.g.* MnO_2 for H_2O_2 for decomposition (heterogeneous);
e.g. $2H_2O_2 \rightarrow 2H_2O + O_2$;
H_2SO_4 / H^+ for ester production (homogeneous);
e.g. $CH_3COOH + CH_3CH_2OH \rightleftharpoons CH_3COOCH_2CH_3 + H_2O$; *[4]*

(c) (i) the power of a reactant's concentration in the rate equation / sum of powers of concentration / rate $= k[X]^n$, where n = order of reaction; *[1]*

Must be in terms of powers of concentration.

(ii) experiment 1—2 : [X] doubles and rate × 4;
2nd order for X;
experiment 2—3 : [Y] doubles and rate × 2;
1st order for Y; *[4]*

(iii) rate $= k[X]^2[Y]$ *(ECF from (ii))*;

for experiment 1, $1.0 \times 10^{-2} = k(0.25)^2(0.25)$;
$k = 0.64$;
$\text{mol}^{-2}\,\text{dm}^6\,\text{s}^{-1}$; *[4]*
Allow ECF from rate expression.

(iv) $\text{rate} = 0.64[0.40]^2[0.60]$;
$= 0.061$; *[2]*
Final answer to 2 sig figs only. Do not penalize if already penalized in 3 (b) (i).
Allow ECF from (iii).

Q# 38/ IB Chem/2003/s/TZ0/Paper 2 Section A/Higher Level/Q3

(b) (i) after 15s (product) $= 0.37$ (mol dm^{-3});

$\text{rate} = \frac{0.37}{15} = 0.025$;

2 sig figs – if penalized here do not penalize in Q6 (c) (iv) or Q8 (d) (ii)

$\text{mol dm}^{-3}\,\text{s}^{-1}$ / M s^{-1} / $\frac{\text{M}}{\text{sec}}$; *[3]*

(ii) at equilibrium / rates of forward and reverse reactions are equal / $\Delta G = 0$; *[1]*

Q# 39/ IB Chem/2002/w/TZ0/Paper 2 Section A/Higher Level/

3. (a) (P constant) $\frac{955}{278} = \frac{V}{298}$; *[1]*

$V = 1024\ \text{cm}^3$; *(accept 1020 cm^3)* *[1]*
no / it will not burst; *[1]*
Third mark has to follow from correctly calculated volume. Allow for ECF.

(b) increased temperature increases the energy content / kinetic energy of the particles / increases velocity / speed;
particles move about more / spread out;
collide more frequently / more forcefully; *[2 max]*
Any two correct answers.

Q# 40/ IB Chem/2002/s/TZ0/Paper 2 Section A/Higher Level/

1. (a) Sum of *[1]*
orders / powers of concentration terms (in rate expression) *[1]*. **[2]**

(b) (A) 0, with justification *[1]*;
(B) 2, with justification *[1]*. **[2]**
(If 0 and 2 are stated with no justification, award [1])

(c) rate = $k[\mathrm{B}]^2$ **[1]**

(d) $k = \dfrac{3.0\times10^{-4}}{(2\times10^{-3})^2} = 75$ *[1]* $\mathrm{mol^{-1}\,dm^3\,min^{-1}}$ *[1]* **[2]**
(Accept l instead of dm^3, min^{-1} is essential. Allow ECF from (c).)

(e) 4 *(Allow ECF. Answer must be consistent with (c).)* **[1]**

(f) in a phase/state different from that of the reagents *[1]*;
(Award [1] each for any two of the following:)
adsorption / absorption *[1]*;
on active sites / on surface of catalyst *[1]*;
reactants' bonds weakened/broken *[1]*;
more favourable orientation *[1]*;
lowers activation energy/E_a *[1]*;
provides alternative reaction pathway *[1]*; **[3]**

(g) (i) (rate) increases *[1]*;
(alternative route with) lower E_a *[1]*. **[2]**

(ii) (rate) decreases *[1]*;
fewer molecules with E_a / less frequent or energetic collisions between molecules / lower kinetic energy of molecules *[1]*. **[2]**

Q# 41/ IB Chem/2001/w/TZ0/Paper 2 Section B/Higher Level/Q8

(c) Correct diagram (reactants level, products level and activation energy) *[1]*
for an endothermic change (*no double jeopardy*) *[1]*
showing a lower E_a for the catalysed reaction. *[1]*

Rate of forward reaction:	increases	*[1]*
Rate of reverse reaction:	increases	*[1]*
Equilibrium position:	unchanged	*[1]*

Q# 42/ IB Chem/2001/w/TZ0/Paper 2 Section A/Higher Level/3

(b) More energetic molecules escape / *owtte*. *[1]*
Remaining molecules are less energetic (lower T) *[1]*
OR Thus its average energy drops (so does its temperature)

OR Bonds are broken as molecules escape
Energy to break these bonds comes from the remaining molecules which are less energetic (thus lower T)

Q# 43/ IB Chem/2001/w/TZ0/Paper 2 Section A/Higher Level/Q1

(b) The reaction of methanol and oxygen has a high activation energy / *OWTTE*. *[1]*
When $E \geq E_a$ / *OWTTE* is provided, a (highly exothermic) reaction takes place. *[1]*

5. (a)

Volume of gas method:	**or**	Mass loss method:
Chemicals in flask / beaker *[1]*		Chemicals in flask / beaker *[1]*
Method of collection *[1]*		Top pan balance *[1]*
Airtight *[1]*		Not airtight *[1]*

Measurement and processing:
Volume / mass monitoring *[1]*
Timing *[1]*
Appropriate processing of data *[1]*
e.g. convert time to rate / rate = $\frac{1}{\text{time}}$ / plot volume–time graph and measure slope *[6 max]*

(b) Slope decreases / slope becomes less steep / rate decreases *[1]*
Acid / reactant concentration falls / surface area of magnesium / reactant decreases *[1]* *[2 max]*

(c) (i) For line B:
Curve shown less steep than A *[1]*
Curve becomes horizontal at half the height of A / explain why curve goes to half height *[1]*
Acid particles more widely spaced / less frequent collisions between reactant particles *[1]* *[3 max]*

(ii) For line C:
Curve shown steeper than A *[1]*
Curve becomes horizontal at same height as A *[1]*
Magnesium surface area increased / more frequent collisions between reactant particles *[1]* *[3 max]*

(d) Order = 1 / first order *[1]*
Rate proportional to concentration / graph is straight line *[1]* *[2 max]*

(e) Rate = $k[H_2SO_4]$ *[1]*

e.g. $k = \frac{0.008}{0.9} \Big/ \frac{0.006}{0.675} \Big/ \approx 8.9\times10^{-3}$ *[1]* (*accept answer* $\pm 2\times10^{-3}$)

s^{-1} *[1]*
The value of *k* would be greater at a higher temperature *[1]* *[4 max]*

Q# 45/ IB Chem/2000/w/TZ0/Paper 2 Section A/Higher Level/

3. (a) (i) $t_{\frac{1}{2}} = \dfrac{0.693}{k}$;

$k = \dfrac{0.693}{1.62 \times 10^4\ \text{s}}$

$= 4.28 \times 10^{-5}\ \text{s}^{-1}$ *[1]*

(No mark without units.)

(ii) $\ln k = \dfrac{-E_a}{RT} + \ln A$

Therefore $E_a = (\ln A - \ln k)RT$

$= [3.219 - (-10.059)] \times 8.314\ \text{J K}^{-1} \times 1107\ \text{K}$ *[1]*

(Award [1] for correct temperature.)

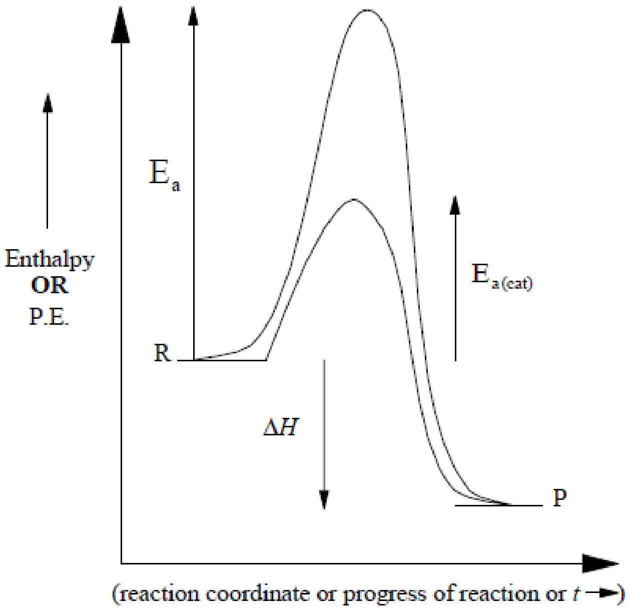
$= 122\ \text{kJ}$ *[1]*

(No mark without unit. If T taken as 834 °C, then $E_a = 92.1\ \text{kJ}$ *)*

(iii) Slow step: $N_2O + N_2O \rightarrow N_4O_2$ **OR** $N_2O + N_2O \rightarrow N_2O_2 + N_2$ *[1]*

Fast step: $N_4O_2 \rightarrow 2N_2 + O_2$ **OR** $N_2O_2 \rightarrow N_2 + O_2$ *[1]*

(b)



(Award [1] for exothermic reaction and [1] for E_a *and* $E_{a(cat)}$. *Allow [1] for showing* E_a *lower for catalysed reaction.)* *[2]*

Q# 46/ IB Chem/2000/s/TZ0/Paper 2 Section A/Higher Level/Q1

(b) (i) O_3: 1st order plus reasonable attempt to justify (*e.g.* double $[O_3]$ doubles rate). *[1]*
NO: 1st order plus reasonable attempt to justify (*e.g.* triple $[O_3]$ and triple [NO], rate is × 9.) *[1]*
*(Two correct orders but no reasoning, award **[1]**.)*

Rate = $k[O_3][NO]$ (*accept rate expression consistent with stated orders – ECF principle*) *[1]*

(ii) $0.66\times10^{-4} = k\times3\times10^{-6}\times10^{-6}$
$k = 2.2\times10^{7}$ *[1]* $dm^3\ mol^{-1}\ s^{-1}$ *[1]* *[2]*
(Apply error carried forward (ECF) from rate expression in (i). U-1 may be applied.)

(iii) Rate (experiment 4) $= 2.2\times10^{7}\times4.5\times10^{-6}\times7.2\times10^{-6} = 7.13\times10^{-4}$ *[1]*
OR Rate (experiment 4) = Rate (experiment 3) $\times\frac{3}{2}\times\frac{8}{10} = 7.13\times10^{-4}$
(Apply ECF from rate expression in (i), must be experiment 4, units not required.)

Q# 47/ IB Chem/1999/s/TZ0/Paper 2 Section A/Higher Level/

4. (a) 2 ***[1 mark]***

(b) 1 ***[1 mark]***

(c) Rate = $k[A]^2[B]$ [ ] must be used ***[1 mark]***

(d) $0.5 = k(0.2)^2(0.2)$ $k = 62.5$ error carried forward from (c) ***[1 mark]***

Q# 48/ IB Chem/1999/s/TZ0/Paper 2 Section A/Higher Level/

5. (a) Step 1 since it is the slowest. (Explanation must be given.) ***[1 mark]***

(b) Step 1 ***[1 mark]***
Slowest step, therefore has a higher activation energy relative to Step 2. ***[1 mark]***

(c) Rate = $k[NO_2][F_2]$ *mark consequentially on (a)* ***[1 mark]***