# 16.2 Activation Energy

## Question Paper

| | |
|---|---|
| Course | DP IB Chemistry |
| Section | 16. Chemical Kinetics (HL only) |
| Topic | 16.2 Activation Energy |
| Difficulty | Hard |

**Time allowed:** 40

**Score:** /26

**Percentage:** /100

## Question 1a

a)
A series of experiments were carried out to investigate how the rate of the reaction of bromate and bromide in acidic conditions varies with temperature.

The time taken, *t*, was measured for a fixed amount of bromine to form at different temperatures. The results are shown below.

| Temperature (*T*) / K | $\frac{1}{T}$ x $10^{-3}$ / $K^{-1}$ | Time (t) / s | $\frac{1}{t}$ / $s^{-1}$ | ln $\frac{1}{t}$ |
|---|---|---|---|---|
| 408 | 2.451 | 21.14 | 0.0473 | -3.051 |
| 428 | 2.336 | 10.57 | | |
| 448 | | 5.54 | 0.1805 | -1.712 |
| 468 | 2.137 | 3.02 | 0.3311 | -1.106 |
| 488 | 2.049 | | | -0.536 |

Complete the table above.

[3]

**[3 marks]**

## Question 1b

b)

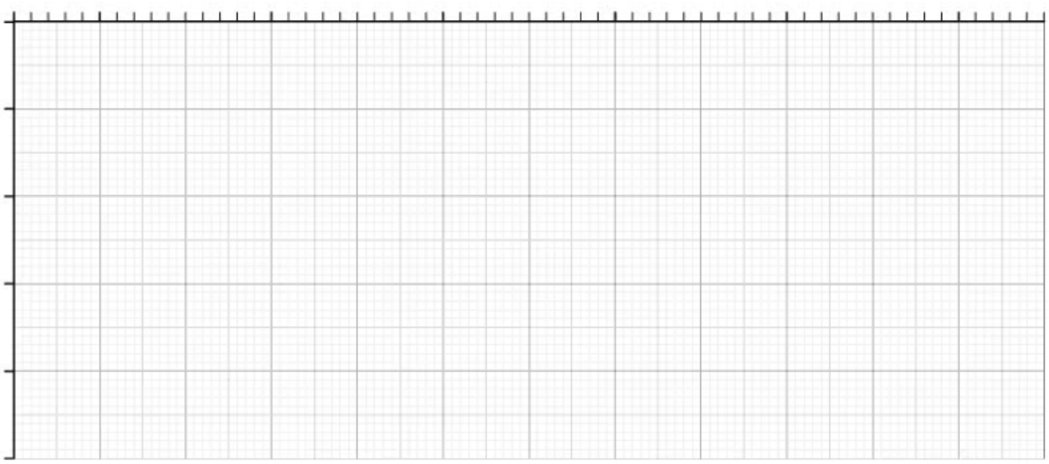
The Arrhenius equation relates the rate constant, $k$, to the activation energy, $E_a$, and temperature, $T$.

$$\ln k = \ln A + \frac{-E_a}{RT}$$

In this experiment, the rate constant, $k$, is directly proportional to $\frac{1}{t}$. Therefore,

$$\ln \frac{1}{t} = \ln A + \frac{-E_a}{RT}$$

Use your answers from part (a) to plot a graph of $\ln \frac{1}{t}$ against $\frac{1}{T} \times 10^{-3}$ on the graph below.



[4]

**[4 marks]**

## Question 1c

c)
Use section 2 of the data booklet along with your graph and information from part (b) to calculate a value for the activation energy, in kJ mol$^{-1}$, for this reaction.

To gain full marks you must show all of your working.

[4]

**[4 marks]**

## Question 2a

a)
Three experiments were carried out at a temperature, $T_1$, to investigate the rate of the reaction between compounds **F** and **G**. The results are shown in the table below:

| | **Experiment 1** | **Experiment 2** | **Experiment 3** |
|---|---|---|---|
| Initial concentration of **F** / mol dm$^{-3}$ | $1.71 \times 10^{-2}$ | $5.34 \times 10^{-2}$ | $7.62 \times 10^{-2}$ |
| Initial concentration of **G** / mol dm$^{-3}$ | $3.95 \times 10^{-2}$ | $6.24 \times 10^{-2}$ | $3.95 \times 10^{-2}$ |
| Initial rate / mol dm$^{-3}$ s$^{-1}$ | $3.76 \times 10^{-3}$ | $1.85 \times 10^{-2}$ | $1.68 \times 10^{-2}$ |

Use the data in the table to deduce the rate equation for the reaction between compounds **F** and **G**.

[3]

**[3 marks]**

## Question 2b

b)
Use the information in the table in part (a) to calculate a value for the rate constant, *k*, for this reaction between 0.0534 mol $dm^{-3}$ **F** and 0.0624 mol $dm^{-3}$ **G**.

Give your answer to the appropriate number of significant figures.

State the units for *k*.

(If you did not get an answer for (a), you may assume that *rate* = *k* [**F**]$^2$ [**G**]$^2$. This is **not** the correct answer)

[2]

**[2 marks]**

## Question 2c

c)
The Arrhenius equation shows how the rate constant, *k*, for a reaction varies with temperature, *T*.

$$k = Ae^{\frac{-E_a}{RT}}$$

For the reaction between 0.0534 mol $dm^{-3}$ **F** and 0.0624 mol $dm^{-3}$ **G** at 25 °C, the activation energy, $E_a$, is 16.7 kJ $mol^{-1}$.

Use section 2 of the data booklet and your answer to part (b) to calculate a value for the Arrhenius constant, *A*, for this reaction.

Give your answer to the appropriate number of significant figures.

(If you did not get an answer for (b), you may assume that *k* has a value of 4.97. This is **not** the correct answer)

[2]

**[2 marks]**

## Question 2d

d)
The temperature of the reaction is increased to twice the original temperature, $T_1$.
The value of k increases to 0.28 $mol^{-1}\ dm^3\ s^{-1}$ at this new temperature.

Using sections 1 and 2 of the data booklet and your answer to part (b), determine the original temperature, $T_1$.
(If you did not get an answer for (b), you may assume that k = 16700 $mol^{-1}\ dm^3\ s^{-1}$ This is **not the correct answer)**

[2]

**[2 marks]**

## Question 3a

a)
The rate constant for a reaction doubles when the temperature is increased from 25.0 °C to 35 °C.

Calculate the activation energy, $E_a$, in kJ $mol^{-1}$ for the reaction using section 1 and 2 of the data booklet.

[2]

**[2 marks]**

## Question 3b

b)
The rate constant is $6.2 \times 10^3\ s^{-1}$ when the temperature is reduced by a factor of a fifth from the original starting temperature, 25 °C.
Calculate the rate constant, in $min^{-1}$, using sections 1 and 2 of the data booklet.

[2]

**[2 marks]**

## Question 3c

c)

A different reaction route is used which reduces the activation energy of the reaction.

Explain how the rate constant calculated in part(b) would differ.

[2]

**[2 marks]**