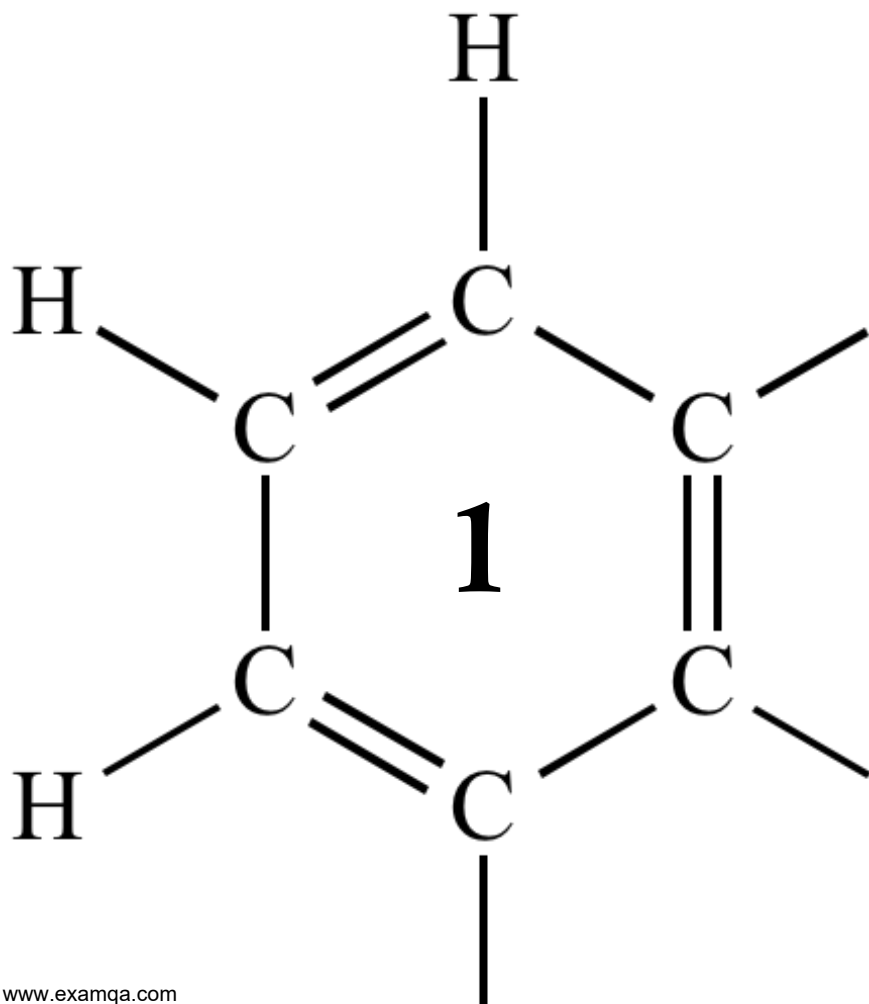


OCR A2 CHEMISTRY

MODULE 5.1

REACTION RATES



1

The initial rate of the reaction between gases **D** and **E** was measured in a series of experiments at a constant temperature. The results are shown in the table.

Expt	Initial [D] / mol dm ⁻³	Initial [E] / mol dm ⁻³	Initial rate / mol dm ⁻³ s ⁻¹
1	1.25 × 10 ⁻²	5.81 × 10 ⁻¹	1.16 × 10 ⁻²
2	1.88 × 10 ⁻²	8.73 × 10 ⁻¹	3.92 × 10 ⁻²
3	1.88 × 10 ⁻²	1.75	1.57 × 10 ⁻¹

(a) Deduce the order of reaction with respect to **D** and the order with respect to **E**.

Order with respect to **D**

Order with respect to **E**

Space for working

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(2)

(b) Suggest why initial rates of reaction are used to determine these orders rather than rates of reaction at other times during the experiments.

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(1)

(c) State how the initial rate is obtained from a graph of the concentration of the product against time.

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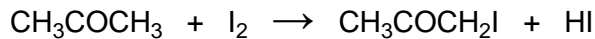
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(2)

(Total 5 marks)

2

The reaction between propanone and iodine in the presence of hydrochloric acid was studied at a constant temperature.



The following rate equation was deduced.

$$\text{rate} = k [\text{CH}_3\text{COCH}_3][\text{H}^+]$$

(a) Suggest why the order with respect to iodine is zero.

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(1)

(b) In an experiment the initial concentrations of propanone, iodine and hydrochloric acid were as shown in the table. The initial rate of reaction in this experiment was $8.64 \times 10^{-7} \text{ mol dm}^{-3} \text{ s}^{-1}$.

	Initial concentration / mol dm ⁻³
CH ₃ COCH ₃	5.82×10^{-2}
I ₂	1.78×10^{-3}
H ⁺	4.76×10^{-1}

Use the data in the table and the rate equation to calculate a value for the rate constant at this temperature.

Give units with your answer.

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(2)

- (c) A series of experiments was carried out using concentrations of propanone approximately 100 times the concentrations of iodine and hydrochloric acid.

Suggest the rate equation under these conditions.
Explain your answer.

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(2)
(Total 5 marks)

3

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

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

Table 1

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

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

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(3)

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

Table 2

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

The rate equation for this reaction is

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

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

$$k = \dots\dots\dots \text{Units} = \dots\dots\dots$$

(3)

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

$$\text{Initial rate} = \dots\dots\dots \text{mol dm}^{-3} \text{s}^{-1}$$

(1)

- (d) The rate equation for a reaction is

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

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

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(3)

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

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

The constant $A = 2.57 \times 10^{10} \text{ mol}^{-1} \text{ dm}^3$.

The gas constant $R = 8.31 \text{ J K}^{-1} \text{ mol}^{-1}$.

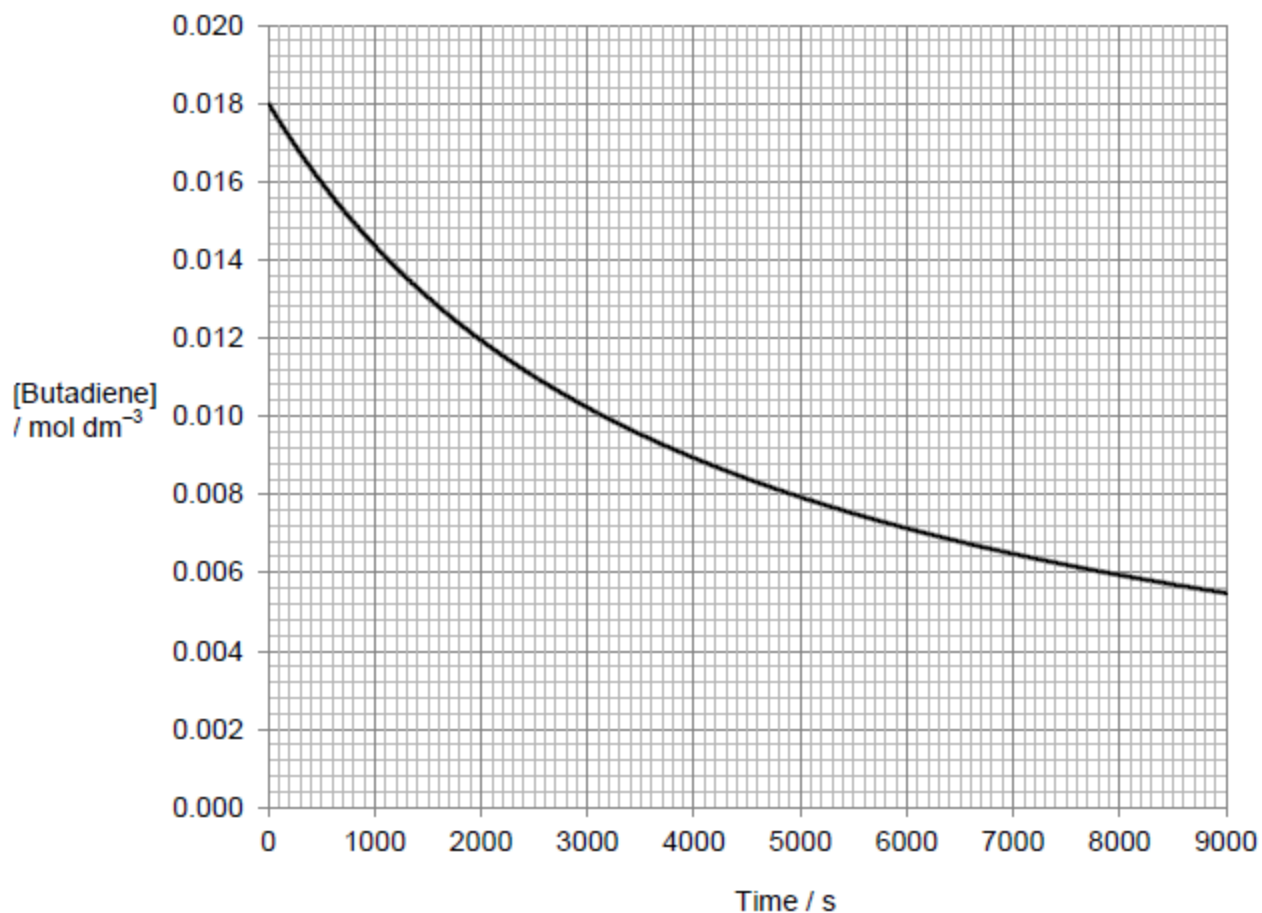
Activation energy =

(2)
(Total 12 marks)

4 Butadiene dimerises according to the equation



The kinetics of the dimerisation are studied and the graph of the concentration of a sample of butadiene is plotted against time. The graph is shown below.



(a) Draw a tangent to the curve when the concentration of butadiene is $0.0120 \text{ mol dm}^{-3}$.

(1)

(b) The initial rate of reaction in this experiment has the value $4.57 \times 10^{-6} \text{ mol dm}^{-3} \text{ s}^{-1}$.

Use this value, together with a rate obtained from your tangent, to justify that the order of the reaction is 2 with respect to butadiene.

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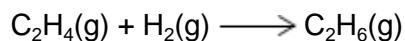
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(5)
(Total 6 marks)

5

The rate equation for the hydrogenation of ethene



is $\text{Rate} = k[\text{C}_2\text{H}_4][\text{H}_2]$

At a fixed temperature, the reaction mixture is compressed to triple the original pressure.

What is the factor by which the rate of reaction changes?

- A 6
- B 9
- C 12
- D 27

(Total 1 mark)

6

Gases **A** and **B** react as shown in the following equation.



The initial rate of the reaction was measured in a series of experiments at a constant temperature. The following rate equation was determined.

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

An incomplete table of data for the reaction between **A** and **B** is shown in the table.

Experiment	Initial [A] / mol dm ⁻³	Initial [B] / mol dm ⁻³	Initial rate / mol dm ⁻³ s ⁻¹
1	4.2 × 10 ⁻³	2.8 × 10 ⁻³	3.3 × 10 ⁻⁵
2	7.9 × 10 ⁻³	2.8 × 10 ⁻³	
3		5.6 × 10 ⁻³	1.8 × 10 ⁻⁴

- (a) Use the data from Experiment **1** to calculate a value for the rate constant, *k*, at this temperature.
Deduce the units of *k*.

Calculation

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Units

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(3)

- (b) Use your value of *k* from (a) to complete the table for the reaction between **A** and **B**.
(If you have been unable to calculate an answer for (a), you may assume a value of 2.3.
This is **not** the correct answer.)

(2)

- (c) The reaction is zero order with respect to **B**.

State the significance of this zero order for the mechanism of the reaction.

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(1)
(Total 6 marks)

7

- (a) The table shows the results of three experiments to investigate the rate of reaction between compounds **A** and **B** dissolved in a given solvent. All three experiments were carried out at the same temperature.

	Experiment 1	Experiment 2	Experiment 3
Initial concentration of A / mol dm ⁻³	1.60×10^{-2}	2.40×10^{-2}	3.60×10^{-2}
Initial concentration of B / mol dm ⁻³	4.20×10^{-2}	6.30×10^{-2}	6.30×10^{-2}
Initial rate / mol dm ⁻³ s ⁻¹	8.00×10^{-5}	1.80×10^{-4}	4.05×10^{-4}

- (i) Deduce the order of reaction with respect to **A**. Tick (✓) **one** box.

Order of reaction with respect to A	Tick (✓)
0	
1	
2	

(1)

- (ii) Deduce the order of reaction with respect to **B**.
Tick (✓) **one** box.

Order of reaction with respect to B	Tick (✓)
0	
1	
2	

(1)

- (b) The reaction between two different compounds, **C** and **D**, is studied at a given temperature. The rate equation for the reaction is found to be

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

- (i) When the initial concentration of **C** is $4.55 \times 10^{-2} \text{ mol dm}^{-3}$ and the initial concentration of **D** is $1.70 \times 10^{-2} \text{ mol dm}^{-3}$, the initial rate of reaction is $6.64 \times 10^{-5} \text{ mol dm}^{-3} \text{ s}^{-1}$.

Calculate the value of the rate constant at this temperature and deduce its units.

Calculation

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Units of rate constant

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(3)

- (ii) The experiment in part (i) is repeated at the same temperature but after the addition of extra solvent so that the total volume of the mixture is doubled.

Deduce the new initial rate of reaction.

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(1)

(Total 6 marks)

8

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

(a) The data in this table were obtained in a series of experiments on the rate of the reaction between compounds **E** and **F** at a constant temperature.

Experiment	Initial concentration of E / mol dm ⁻³	Initial concentration of F / mol dm ⁻³	Initial rate of reaction / mol dm ⁻³ s ⁻¹
1	0.15	0.24	0.42 × 10 ⁻³
2	0.45	0.24	3.78 × 10 ⁻³
3	0.90	0.12	7.56 × 10 ⁻³

(i) Deduce the order of reaction with respect to **E**.

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(1)

(ii) Deduce the order of reaction with respect to **F**.

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(1)

- (b) The data in the following table were obtained in two experiments on the rate of the reaction between compounds **G** and **H** at a constant temperature.

Experiment	Initial concentration of G / mol dm ⁻³	Initial concentration of H / mol dm ⁻³	Initial rate of reaction / mol dm ⁻³ s ⁻¹
4	3.8×10^{-2}	2.6×10^{-2}	8.6×10^{-4}
5	6.3×10^{-2}	7.5×10^{-2}	To be calculated

The rate equation for this reaction is

$$rate = k[G]^2[H]$$

- (i) Use the data from Experiment **4** to calculate a value for the rate constant k at this temperature. Deduce the units of k .

Calculation

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Units

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(3)

- (ii) Calculate a value for the initial rate of reaction in Experiment **5**.

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(1)

(Total 6 marks)

9

(a) The data in the following table were obtained in two experiments about the rate of the reaction between substances **B** and **C** at a constant temperature.

Experiment	Initial concentration of B / mol dm ⁻³	Initial concentration of C / mol dm ⁻³	Initial rate / mol dm ⁻³ s ⁻¹
1	4.2×10^{-2}	2.6×10^{-2}	8.4×10^{-5}
2	6.3×10^{-2}	7.8×10^{-2}	To be calculated

The rate equation for this reaction is known to be

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

(i) Use the data from Experiment **1** to calculate a value for the rate constant *k* at this temperature and deduce its units.

Calculation

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Units

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(Extra space)

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(3)

(ii) Calculate a value for the initial rate in Experiment **2**.

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(1)

- (b) The data in the following table were obtained in a series of experiments about the rate of the reaction between substances **D** and **E** at a constant temperature.

Experiment	Initial concentration of D / mol dm ⁻³	Initial concentration of E / mol dm ⁻³	Initial rate / mol dm ⁻³ s ⁻¹
3	0.13	0.23	0.26 × 10 ⁻³
4	0.39	0.23	2.34 × 10 ⁻³
5	0.78	0.46	9.36 × 10 ⁻³

- (i) Deduce the order of reaction with respect to **D**.

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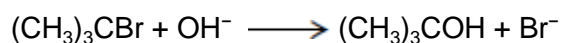
(1)

- (ii) Deduce the order of reaction with respect to **E**.

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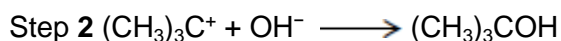
(1)

- (c) The compound (CH₃)₃CBr reacts with aqueous sodium hydroxide as shown in the following equation.



This reaction was found to be first order with respect to (CH₃)₃CBr but zero order with respect to hydroxide ions.

The following two-step process was suggested.



- (i) Deduce the rate-determining step in this two-step process.

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(1)

(ii) Outline a mechanism for this step using a curly arrow.

(1)
(Total 8 marks)

10

Gases **P** and **Q** react as shown in the following equation.



The initial rate of the reaction was measured in a series of experiments at a constant temperature. The following rate equation was determined.

$$\text{rate} = k[\text{P}]^2[\text{Q}]$$

(a) Complete the table of data for the reaction between **P** and **Q**.

Experiment	Initial [P] / mol dm ⁻³	Initial [Q] / mol dm ⁻³	Initial rate / mol dm ⁻³ s ⁻¹
1	2.5×10^{-2}	1.8×10^{-2}	5.0×10^{-5}
2	7.5×10^{-2}	1.8×10^{-2}	
3	5.0×10^{-2}		5.0×10^{-5}
4		5.4×10^{-2}	4.5×10^{-4}

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(3)

- (b) Use the data from Experiment 1 to calculate a value for the rate constant (k) at this temperature. Deduce the units of k .

Calculation

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Units

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(3)
(Total 6 marks)