Q1. Gases **A** and **B** react as shown in the following equation.

$$2\mathbf{A}(g) + \mathbf{B}(g) \longrightarrow \mathbf{C}(g) + \mathbf{D}(g)$$

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

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

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)

Q2. (a) In the presence of the catalyst rhodium, the reaction between NO and H_2 occurs according to the following equation.

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

The kinetics of the reaction were investigated and the rate equation was found to be

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

The initial rate of reaction was 6.2×10^{-6} mol dm^{-3} s^{-1} when the initial concentration of NO was 2.9×10^{-2} mol dm^{-3} and the initial concentration of H_2 was 2.3×10^{-2} mol dm^{-3}.

(i) Calculate the value of the rate constant under these conditions and give its units.

Calculation .....

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Units .....

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(3)

(ii) Calculate the initial rate of reaction if the experiment is repeated under the same conditions but with the concentrations of NO and of H_2 both doubled from their original values.

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(1)

(b) Using the rate equation and the overall equation, the following three-step mechanism for the reaction was suggested. X and Y are intermediate species.

Step **1** $NO + NO \longrightarrow X$

Step **2** $X + H_2 \longrightarrow Y$

Step **3** $Y + H_2 \longrightarrow N_2 + 2H_2O$

Suggest which **one** of the three steps is the rate-determining step.

Explain your answer.

Rate-determining step.....

Explanation .....

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(Extra space) .....

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(2)
(Total 6 marks)

Q3. 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	$3.78 \times 10_{-3}$

(a) Show how these data can be used to deduce the rate expression for the reaction

between **A** and **B**.

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(3)

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

Table 2

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

The rate equation for this reaction is

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

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

k = ... Units = ...

(3)

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

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

(1)

(d) The rate equation for a reaction is

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

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

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(3)

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

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

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

Activation energy = ...

(2)
(Total 12 marks)

Q4. The initial rate of the reaction between two gases **P** and **Q** was measured in a series of experiments at a constant temperature. The following rate equation was determined.

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

(a) Complete the table of data below 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	0.20	0.30	$1.8 = 10^{-3}$
2	0.40	0.60	
3	0.60		$5.4 = 10^{-3}$
4		0.90	$12.2 = 10^{-3}$

(Space for working) .....

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(3)

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

Calculation .....

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Units .....

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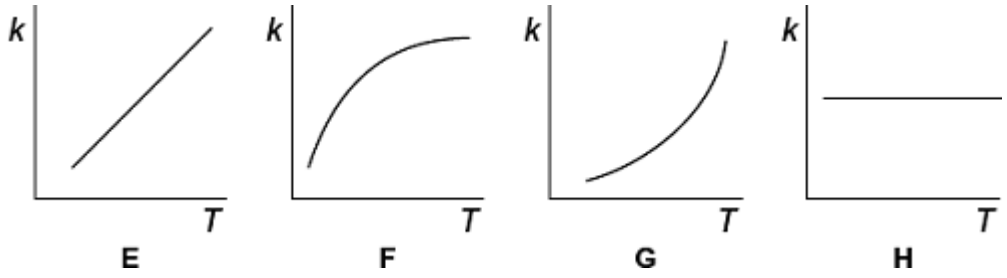
(3)

(c) Consider the graphs **E**, **F**, **G** and **H** below.



Write in the box below the letter of the graph that shows how the rate constant *k* varies with temperature.

(1)
(Total 7 marks)

Q5. Gases **P** and **Q** react as shown in the following equation.

$$2P(g) + 2Q(g) \longrightarrow R(g) + S(g)$$

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

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

(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}

(Space for working) .....

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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)

Q6.(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

rate = k[**B**]^2[**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 $(CH_3)_3CBr$ reacts with aqueous sodium hydroxide as shown in the folfollowing equation.

$$(CH_3)_3CBr + OH^- \longrightarrow (CH_3)_3COH + Br^-$$

This reaction was found to be first order with respect to $(CH_3)_3CBr$ but zero order with respect to hydroxide ions.

The following two-step process was suggested.

Step **1** $(CH_3)_3CBr \longrightarrow (CH_3)_3C^+ + Br^-$

Step **2** $(CH_3)_3C^+ + OH^- \longrightarrow (CH_3)_3COH$

(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)

Q7.(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 4.55×10^{-2} mol dm^{-3} and the initial concentration of **D** is 1.70×10^{-2} mol dm^{-3}, the initial rate of reaction is 6.64×10^{-5} mol dm^{-3} 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)

Q8. 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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(Space for working) .....

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(1)

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

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(Space for working) .....

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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

$$rate = \mathbf{k}[\mathbf{G}]^2[\mathbf{H}]$$

(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)

Q9. The rate equation for the hydrogenation of ethene

$$C_2H_4(g) + H_2(g) \longrightarrow C_2H_6(g)$$

is $Rate = k[C_2H_4][H_2]$

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)