1.

(c) Hydrogen reacts with iodine monochloride, ICl, to produce iodine and hydrogen chloride.

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

In an experiment, the following value for the initial rate was obtained.

Iodine monochloride concentration / mol dm^{-3}	*Hydrogen concentration* / mol dm^{-3}	*Initial rate of reaction* / mol dm^{-3} s^{-1}
$1{\cdot}60 \times 10^{-3}$	$1{\cdot}60 \times 10^{-3}$	$6{\cdot}40 \times 10^{-7}$

(i) The reaction is first order for both hydrogen and iodine monochloride. Calculate the value of the rate constant, k, and state its units. [3]

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(ii) Write a chemical equation showing a possible rate-determining step for this reaction. [1]

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(iii) Write a further equation to show how the products shown in the overall equation are obtained from (ii). [1]

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2.

(a) The thermal decomposition of ozone is shown in the equation below.

$$2O_3(g) \longrightarrow 3O_2(g)$$

Kinetic studies have shown that the reaction is second order with respect to ozone.

(i) Write the rate equation for the reaction and use it to explain the term ***order of reaction***. [2]

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(ii) The value of the rate constant at 298 K is $3{\cdot}4 \times 10^{-5}$ $dm^3\ mol^{-1}s^{-1}$. If the concentration of ozone is 0·023 mol dm^{-3}, calculate the rate of reaction at 298 K and state its units. [2]

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(iii) In the stratosphere, chlorine radicals act as catalysts and speed up the decomposition of ozone.

Explain how catalysts increase the rate of reaction. [2]

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(iv) Give **one** problem caused as a consequence of ozone depletion. [1]

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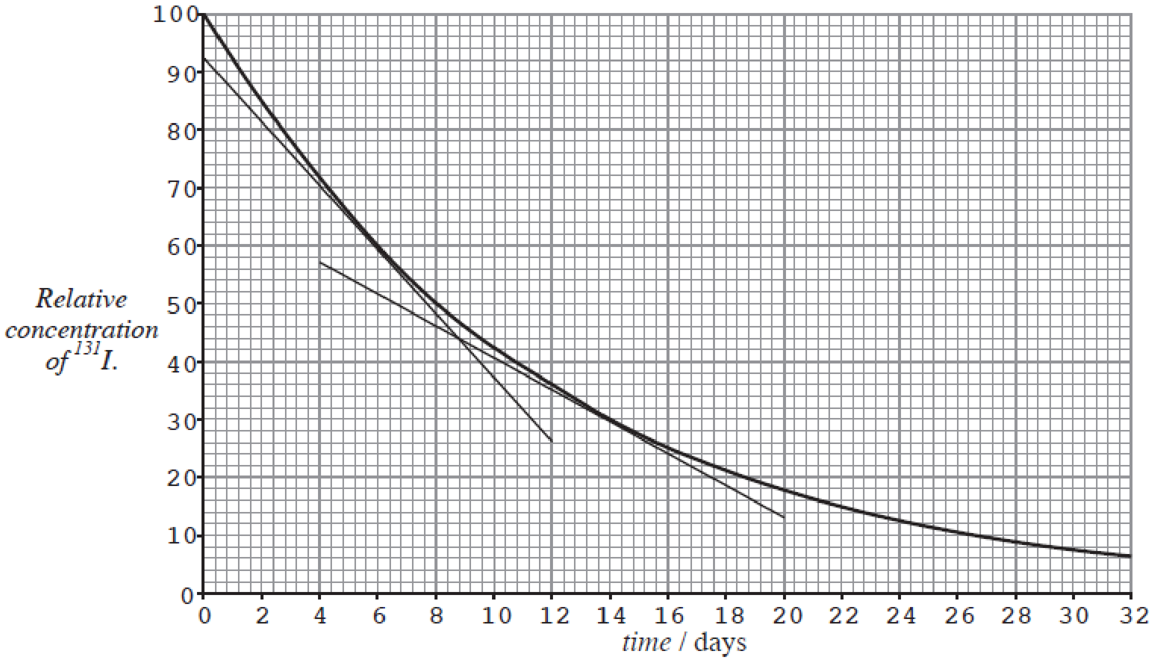
(*b*) Radioactive decay shows first-order kinetics.

(i) The graph below shows the decay of a sample of radioactive iodine ^{131}I.



I. Use the graph to find the half-life of ^{131}I. [1]

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II. Using the tangents drawn at relative concentrations of ^{131}I of 60 and 30, or otherwise, explain how the graph shows that the radioactive decay is first-order. [2]

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3.

The gaseous reaction between carbon monoxide, CO, and nitrogen dioxide, NO_2, is shown by the equation below.

$$CO(g) \quad + \quad NO_2(g) \longrightarrow CO_2(g) \quad + \quad NO(g)$$

(a) At temperatures below 500 K, experiments have shown the rate equation to be given by

$$\text{Rate} = k\,[NO_2]^2$$

The mechanism can be represented as two stages, one of which is the ***rate determining step***:

Stage 1	$NO_2(g) + NO_2(g)$	$\longrightarrow$	Intermediates
Stage 2	Intermediates + $CO(g)$	$\longrightarrow$	$NO(g) + CO_2(g) + NO_2(g)$

Explain the term ***rate determining step*** and, giving your reasoning, state which of the above two stages is the rate determining step in this reaction. [2]

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(b) At higher temperatures, the products are the same but the reaction proceeds by a different mechanism. Experiments at a temperature of 600 K produced the following results.

Initial Rate of Reaction / $mol\ dm^{-3}\ s^{-1}$	Initial Concentration NO_2 $[NO_2]$ / $mol\ dm^{-3}$	Initial Concentration CO [CO] / $mol\ dm^{-3}$
$1{\cdot}12 \times 10^{-4}$	0·050	0·050
$2{\cdot}24 \times 10^{-4}$	0·050	0·100
$4{\cdot}48 \times 10^{-4}$	0·100	0·100
$8{\cdot}96 \times 10^{-4}$	0·100	0·200

(i) Determine the order of reaction with respect to NO_2 and the order with respect to CO and state the rate equation for this reaction at 600 K. [3]

NO_2 order .... *CO order* ....

Rate equation ....

(ii) Calculate the value of the rate constant, k, at 600 K, and give its units. [2]

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(iii) Suggest a rate-determining step which fits the rate equation at 600 K. [1]

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