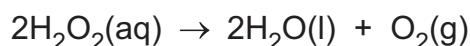


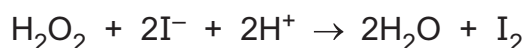
- 3 (a)** Solid manganese(IV) oxide, MnO_2 , catalyses the decomposition of hydrogen peroxide.



State the type of catalysis for this reaction. Explain your answer.

.....
 [1]

- (b)** Hydrogen peroxide reacts with iodide ions in acidic conditions as shown.



The initial rate of this reaction is investigated with different concentrations of H_2O_2 , I^- and H^+ . The results obtained are shown in Table 3.1.

Table 3.1

experiment	$[\text{H}_2\text{O}_2]/\text{mol dm}^{-3}$	$[\text{I}^-]/\text{mol dm}^{-3}$	$[\text{H}^+]/\text{mol dm}^{-3}$	initial rate/ $\text{mol dm}^{-3}\text{s}^{-1}$
1	0.0450	0.0300	0.0125	2.42×10^{-3}
2	0.0225	0.0600	0.0125	2.42×10^{-3}
3	0.0225	0.120	0.0125	4.84×10^{-3}
4	0.0450	0.120	0.0500	9.68×10^{-3}

- (i)** Use the information in Table 3.1 to deduce the rate equation for this reaction.

Explain your reasoning.

.....

 [4]

- (ii)** Use your rate equation from **(b)(i)** and the data from Experiment 1 to calculate the rate constant, k , for this reaction. Include the units of k .

$k = \frac{1}{\text{1}}$ units [2]

- (c) The rate of the thermal decomposition of azomethane, $\text{CH}_3\text{N}=\text{NCH}_3$, is investigated.

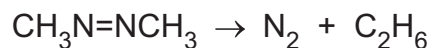


Fig. 3.1 shows the results obtained. The reaction is first order with respect to $\text{CH}_3\text{N}=\text{NCH}_3$.

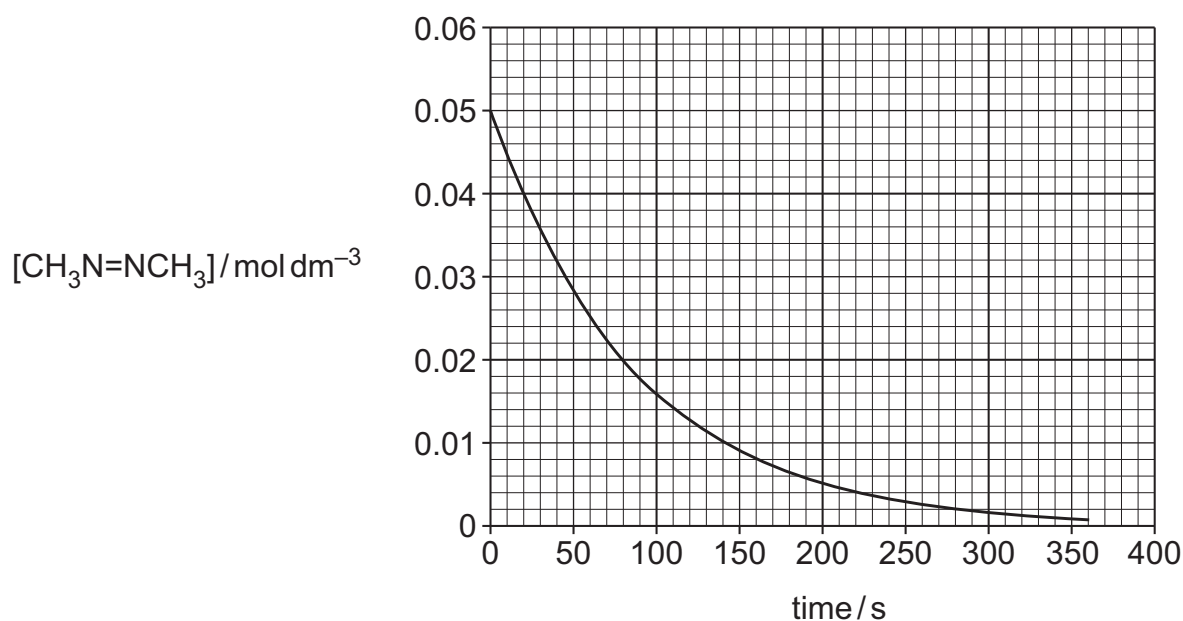


Fig. 3.1

- (i) Use Fig. 3.1 to calculate **two** half-lives, $t_{\frac{1}{2}}$, to show that the reaction is first order.

.....
.....
.....
..... [2]

- (ii) Use your answer to (c)(i) to calculate the rate constant, k , for the decomposition of azomethane.

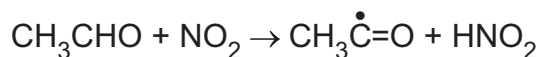
$k = \dots\dots\dots \text{s}^{-1}$ [1]

- (d) Describe the effect of increasing temperature on the rate constant and on the rate of a reaction.

.....
.....
..... [1]

[Total: 11]

- 2 Ethanal, CH_3CHO , reacts with nitrogen dioxide, NO_2 . The products of the first step of this reaction are a $\text{CH}_3\dot{\text{C}}=\text{O}$ radical and a molecule of nitrous acid, HNO_2 .



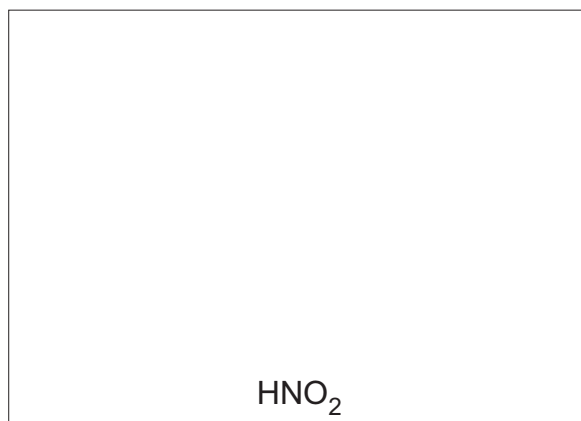
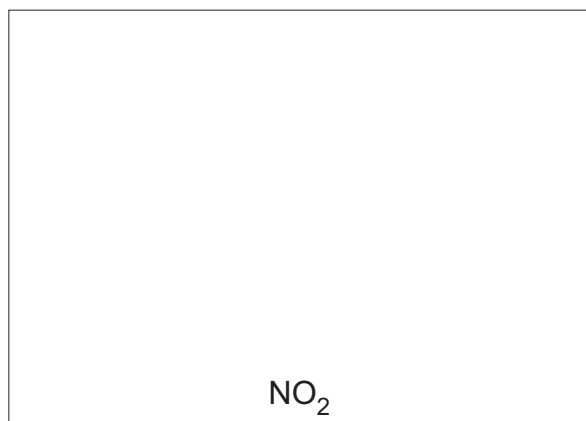
- (a) (i) Use **two** words to complete the sentence.

This reaction involves of the single covalent bond between a hydrogen atom and a carbon atom in CH_3CHO .

[1]

- (ii) The hydrogen atom mentioned in (a)(i) forms a covalent bond with one of the oxygen atoms of an NO_2 molecule. An NO_2 molecule has a single, unpaired electron on the nitrogen atom. All electrons are paired in an HNO_2 molecule.

Draw dot-and-cross diagrams of NO_2 and HNO_2 in the boxes. Show outer shell electrons only.



[1]

- (iii) Use VSEPR theory to predict the bond angle at the nitrogen atom in an HNO_2 molecule.

bond angle = [1]

- (b) The rate equation for the reaction between CH_3CHO and NO_2 is shown.

$$\text{rate} = k[\text{CH}_3\text{CHO}][\text{NO}_2]$$

Under certain conditions, when the concentrations of both CH_3CHO and NO_2 are $0.200 \text{ mol dm}^{-3}$, the rate of the reaction is $1.53 \times 10^{-4} \text{ mol dm}^{-3} \text{ s}^{-1}$.

Calculate the value of the rate constant, k , under these conditions. Give the units of k .

$k = \dots\dots\dots$ units $\dots\dots\dots$ [2]

- (c) The reaction mixture described in (b) is monitored over a period of time.

Predict whether the graph of $[\text{NO}_2]$ against time shows a constant half-life.

Explain your answer.

prediction

explanation

[1]

- (d) NO_2 also reacts with ozone, O_3 .

The rate equation is shown.

$$\text{rate} = k_1[\text{NO}_2]$$

Under certain conditions, the value of k_1 is 0.0848 s^{-1} .

The reaction has a constant half-life under these conditions.

Calculate the half-life in seconds.

half-life = s [1]

- (e) NO_2 is present in the exhaust gases of cars. It can react with carbon monoxide, CO , on the surface of a heterogeneous catalyst in the car's catalytic converter.

Describe the mode of action of this heterogeneous catalyst.

.....
.....
.....
.....
.....
..... [2]

[Total: 9]

4 (a) Nitrogen monoxide, NO, reacts with hydrogen, as shown in reaction 3.



(i) The rate equation for reaction 3 is shown.

$$\text{rate} = k[\text{H}_2][\text{NO}]^2$$

Complete Table 4.1.

Table 4.1

the order of reaction with respect to $[\text{H}_2]$	
the order of reaction with respect to $[\text{NO}]$	
the overall order of the reaction	

[1]

(ii) Predict how the initial rate for reaction 3 changes when the concentration of NO is halved.

..... [1]

(iii) Predict how the initial rate for reaction 3 changes when the concentrations of NO and H_2 are both increased **three** times.

..... [1]

(iv) Suggest why reaction 3 is unlikely to proceed by a mechanism involving only a single step.

.....

..... [1]

(v) Suggest equations for the **three** steps of the reaction mechanism for reaction 3.

Each step involves a reaction between **two** molecules.

step 1 $\rightarrow$

step 2 + $\rightarrow \text{N}_2\text{O} +$

step 3 $\text{N}_2\text{O} +$ $\rightarrow$ +

[2]

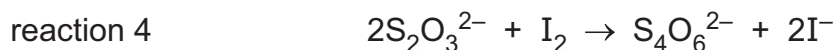
(vi) Suggest the role of N_2O in this mechanism.

Explain your reasoning.

.....

..... [1]

- (b) Iodine, I_2 , reacts with thiosulfate ions, $S_2O_3^{2-}$, as shown in reaction 4.



Reaction 4 is carried out in the presence of a large excess of I_2 . Under these conditions, the reaction is first order with respect to $[S_2O_3^{2-}]$ and zero order with respect to $[I_2]$.

The half-life, $t_{\frac{1}{2}}$, for reaction 4 is 720 s under certain conditions.

Calculate the value of the rate constant, k , for reaction 4. Include the units of k .

$k =$ units [1]

- (c) The reaction between iodide ions, $I^-(aq)$, and peroxydisulfate ions, $S_2O_8^{2-}(aq)$, is catalysed by $Co^{3+}(aq)$. The mechanism is similar to the mechanism of this reaction when $Fe^{3+}(aq)$ is used as the catalyst.

- (i) State the type of catalysis that occurs in this reaction.

Explain your reasoning.

.....
..... [1]

- (ii) Write **two** equations to show how $Co^{3+}(aq)$ catalyses this reaction.

equation 1

equation 2 [2]

- (iii) Suggest why this reaction is slow in the absence of $Co^{3+}(aq)$.

.....
..... [1]

[Total: 12]

- 2 A student uses the following method to investigate the kinetics of the reaction between iodine and tin to produce tin(IV) iodide, SnI_4 .

- step 1** Rinse a block of tin with distilled water and then rinse it with propanone.
- step 2** Place 50 cm^3 of a 0.400 mol dm^{-3} solution of iodine dissolved in methylbenzene in a 100 cm^3 beaker.
- step 3** Suspend the block of tin from a three decimal place balance as shown in Fig. 2.1. Start a timer.
- step 4** Record the balance reading every 100 seconds.

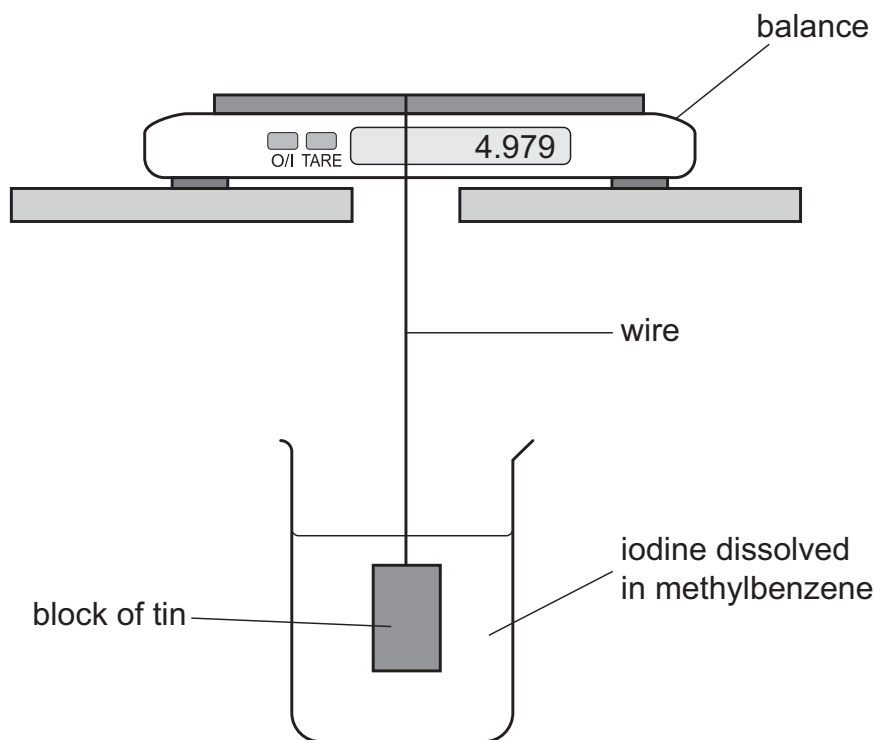


Fig. 2.1

- (a) (i) Suggest why the student rinses the block of tin with propanone after rinsing it with distilled water in step 1.

.....
..... [1]

- (ii) Suggest why water is **not** used as the solvent for iodine.

.....
..... [1]

- (iii) Suggest why a three decimal place balance is more suitable than a two decimal place balance for this experiment.

..... [1]

- (iv) Suggest a control experiment that could be used to verify that the loss in mass of tin is caused by reaction with iodine and **not** any other factor.

.....

..... [1]

- (b) The student's results are shown in Table 2.1.

Complete Table 2.1.

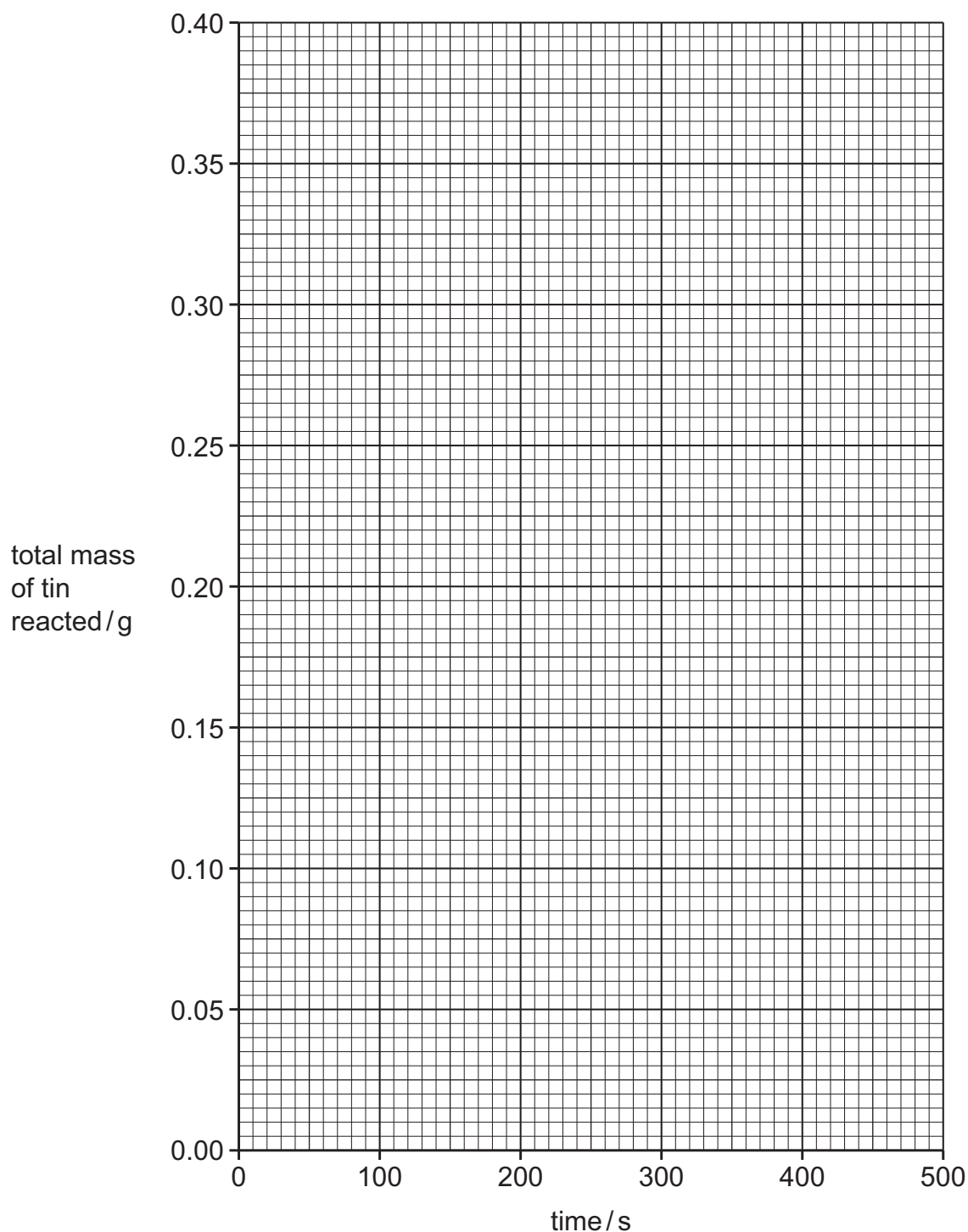
Table 2.1

time/s	balance reading/g	total mass of tin reacted/g
0	4.979	0.000
100	4.910	
200	4.859	
300	4.761	
400	4.688	
500	4.620	

[1]

- (c) (i) Use the results from Table 2.1 to plot a graph on the grid in Fig. 2.2 to show the relationship between total mass of tin reacted and time.

Use a cross (x) to plot each data point. Draw a straight line of best fit.

**Fig. 2.2**

[2]

- (ii) Circle the point on the graph in Fig. 2.2 that you consider to be most anomalous.

Suggest **one** reason why this anomaly may have occurred during this experimental procedure.

Assume all measurements of mass are accurate.

.....

..... [1]

(d) Use your graph in Fig. 2.2 to determine the gradient of the line of best fit.

State the coordinates of both points you used in your calculation. These must be selected from your line of best fit.

Give your gradient to **three** significant figures.

coordinates 1

coordinates 2

gradient =
[2]

(e) Another student makes various concentrations of solutions of iodine dissolved in methylbenzene by dilution of the $0.400 \text{ mol dm}^{-3} \text{ I}_2$ solution.

The student repeats the experiment at a different temperature using these solutions.

Table 2.2

volume of $0.400 \text{ mol dm}^{-3} \text{ I}_2$ solution used / cm^3	volume of methylbenzene used / cm^3	$[\text{I}_2]$ / mol dm^{-3}	relative rate of reaction
100.0	0.0	0.400	4.76
		0.300	3.57
		0.200	2.35
		0.100	1.15

(i) Complete Table 2.2 by adding the volumes of solutions that are mixed to make 100.0 cm^3 of a solution of iodine dissolved in methylbenzene for each required concentration.

[1]

(ii) Identify the independent variable in this experiment.

..... [1]

- (iii) The student concludes that the rate equation for the reaction between iodine and tin is as follows.

$$\text{rate} = k [\text{I}_2]^2$$

State whether the results support the student's conclusion.

Explain your answer using values from Table 2.2.

.....

.....

.....

.....

.....

..... [2]

[Total: 14]

Important values, constants and standards

molar gas constant	$R = 8.31 \text{ J K}^{-1} \text{ mol}^{-1}$
Faraday constant	$F = 9.65 \times 10^4 \text{ C mol}^{-1}$
Avogadro constant	$L = 6.02 \times 10^{23} \text{ mol}^{-1}$
electronic charge	$e = -1.60 \times 10^{-19} \text{ C}$
molar volume of gas	$V_m = 22.4 \text{ dm}^3 \text{ mol}^{-1}$ at s.t.p. (101 kPa and 273 K) $V_m = 24.0 \text{ dm}^3 \text{ mol}^{-1}$ at room conditions
ionic product of water	$K_w = 1.00 \times 10^{-14} \text{ mol}^2 \text{ dm}^{-6}$ (at 298 K (25 °C))
specific heat capacity of water	$c = 4.18 \text{ kJ kg}^{-1} \text{ K}^{-1}$ (4.18 J g ⁻¹ K ⁻¹)

The Periodic Table of Elements

Group																			
1	2													13	14	15	16	17	18