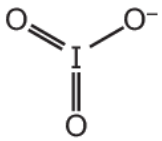
Reaction Kinetics - 2021

1. March/2021/Paper_42/No.3

Iodates are compounds that contain the IO_3^- anion.

(a) The IO_3^- anion is shown.



Explain, with reference to the qualitative model of electron-pair repulsion, why the IO_3^- anion has a pyramidal shape.

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..... [1]

(b) The reaction of iodine and hot aqueous sodium hydroxide is similar to that of chlorine and hot aqueous sodium hydroxide. Sodium iodate, $NaIO_3$, is formed as one of the products.

Suggest an equation for the reaction of iodine and hot aqueous sodium hydroxide.

..... [1]

(c) The decomposition of hydrogen peroxide, H_2O_2, is catalysed by acidified IO_3^-.

H_2O_2 reduces acidified IO_3^- as shown.

$$5H_2O_2 + 2H^+ + 2IO_3^- \rightarrow I_2 + 5O_2 + 6H_2O$$

This reaction is followed by the oxidation of I_2 by H_2O_2.

half-equation	$E^\ominus$/V
$H_2O_2 + 2H^+ + 2e^- \rightleftharpoons 2H_2O$	+1.77
$IO_3^- + 6H^+ + 5e^- \rightleftharpoons \frac{1}{2}I_2 + 3H_2O$	+1.19
$O_2 + 2H^+ + 2e^- \rightleftharpoons H_2O_2$	+0.68

(i) Use the data to show that the separate reactions of H_2O_2 with IO_3^- and with I_2 are both feasible under standard conditions.

In your answer, give the equation for the reaction of H_2O_2 with I_2.

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..... [3]

(ii) Write the overall equation for the decomposition of H_2O_2 catalysed by acidified IO_3^-.

..... [1]

(d) A student collects some data for the reaction of H_2O_2 with acidified IO_3^-, as shown in the table.

experiment	$[H_2O_2]$ / mol dm^{-3}	$[IO_3^-]$ / mol dm^{-3}	$[H^+]$ / mol dm^{-3}	initial rate of reaction / mol dm^{-3} s^{-1}
1	0.0500	0.0700	0.025	1.47×10^{-5}
2	0.100	0.0700	0.050	2.94×10^{-5}
3	0.100	0.140	0.025	5.88×10^{-5}
4	0.150	0.140	0.025	8.82×10^{-5}

(i) Use the data to determine the order of reaction with respect to $[H_2O_2]$, $[IO_3^-]$ and $[H^+]$.

Show your reasoning.

order with respect to $[H_2O_2]$ = ...

.....

.....

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order with respect to $[IO_3^-]$ = ...

.....

.....

.....

order with respect to $[H^+]$ = ...

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.....

.....

[3]

(ii) Use your answer to **(d)(i)** to write the rate equation for this reaction.

rate = ... [1]

(iii) Calculate the value of the rate constant, k, using data from experiment 4 and your answer to **(d)(ii)**.

Give the units of k.

k = ...

units = ...
[2]

(e) $Pb(IO_3)_2$ is only sparingly soluble in water at 25 °C.

The solubility product, K_{sp}, of $Pb(IO_3)_2$ is 3.69×10^{-13} mol^3 dm^{-9} at 25 °C.

(i) Write an expression for the solubility product of $Pb(IO_3)_2$.

K_{sp} =

[1]

(ii) Calculate the solubility, in mol dm^{-3}, of $Pb(IO_3)_2$ at 25 °C.

solubility = ... mol dm^{-3} [2]

(f) NH_4IO_3 is an unstable compound that readily decomposes when warmed. The decomposition reaction is shown.

$$NH_4IO_3(s) \rightarrow \tfrac{1}{2}N_2(g) + \tfrac{1}{2}O_2(g) + \tfrac{1}{2}I_2(g) + 2H_2O(l) \qquad \Delta H = -154.6\,kJ\,mol^{-1}$$

(i) Use the data in the table to calculate the entropy change of reaction, ΔS, of the decomposition of $NH_4IO_3(s)$.

compound	S / $J\,K^{-1}\,mol^{-1}$
$NH_4IO_3(s)$	42
$N_2(g)$	192
$O_2(g)$	205
$I_2(g)$	261
$H_2O(l)$	70

ΔS = ... $J\,K^{-1}\,mol^{-1}$ [2]

(ii) This reaction is feasible at all temperatures.

Explain why, using the data in **(f)** and your answer to **(f)(i)**.

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..... [1]

[Total: 18]

2. June/2021/Paper_41/No.5

Dinitrogen pentoxide, N_2O_5, is dissolved in an inert solvent (solv) and the rate of decomposition of N_2O_5 is investigated. This reaction produces nitrogen dioxide, which remains in solution, and oxygen gas.

$$N_2O_5(solv) \rightarrow 2NO_2(solv) + \tfrac{1}{2}O_2(g)$$

(a) Suggest what measurements could be used to follow the rate of this reaction from the given information.

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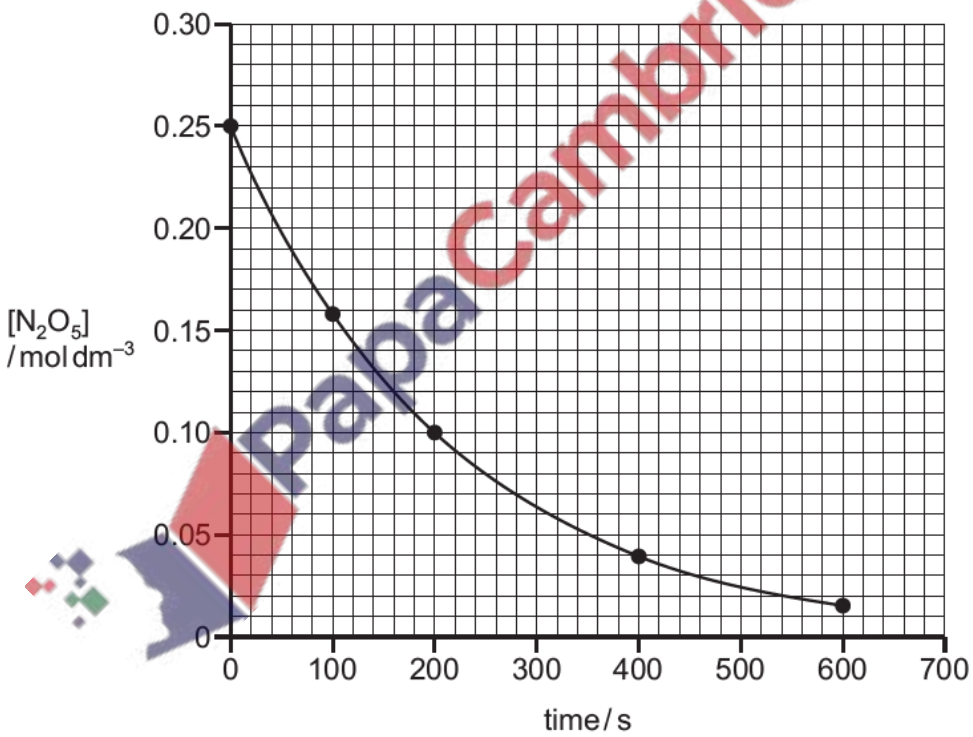
..... [1]

(b) In a separate experiment, the rate of the decomposition of $N_2O_5(g)$ is investigated.

$$N_2O_5(g) \rightarrow 2NO_2(g) + \tfrac{1}{2}O_2(g)$$

The graph shows the results obtained.



The reaction is first order with respect to N_2O_5. This can be confirmed from the graph using half-lives.

(i) Explain the term *half-life of a reaction*.

.....

..... [1]

(ii) Determine the half-life of this reaction. Show your working on the graph.

half-life = ... s [1]

(iii) Suggest the effect on the half-life of this reaction if the initial concentration of N_2O_5 is halved.

..... [1]

(c) (i) Use the graph in **5(b)** to determine the rate of reaction at 200 s. Show your working.

rate = ...

units = ...
[2]

The rate equation for this reaction is shown.

$$\text{rate} = k[N_2O_5]$$

(ii) Use your answer to **(c)(i)** to calculate the value of the rate constant, k, for this reaction and state its units.

k = ... units ..... [1]

(d) Nitrogen dioxide reacts with ozone, O_3, as shown.

$$2NO_2 + O_3 \rightarrow N_2O_5 + O_2$$

The rate equation for this reaction is rate = $k[NO_2][O_3]$.

Suggest a possible two-step mechanism for this reaction.

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..... [2]

[Total: 9]

3. June/2021/Paper_42/No.7

(a) In aqueous solution, chlorine dioxide, ClO_2, reacts with hydroxide ions as shown.

$$2ClO_2 + 2OH^- \rightarrow ClO_3^- + ClO_2^- + H_2O$$

A series of experiments is carried out using different concentrations of ClO_2 and OH^-. The table shows the results obtained.

experiment	$[ClO_2]$ / mol dm^{-3}	$[OH^-]$ / mol dm^{-3}	initial rate / mol dm^{-3} min^{-1}
1	0.020	0.030	7.20×10^{-4}
2	0.020	0.120	2.88×10^{-3}
3	0.050	0.030	4.50×10^{-3}

(i) Explain the term *order of reaction*.

.....

..... [1]

(ii) Use the data in the table to determine the order of reaction with respect to each reactant, ClO_2 and OH^-.

Explain your reasoning.

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..... [2]

(iii) Use your answer to **(a)(ii)** to construct the rate equation for this reaction.

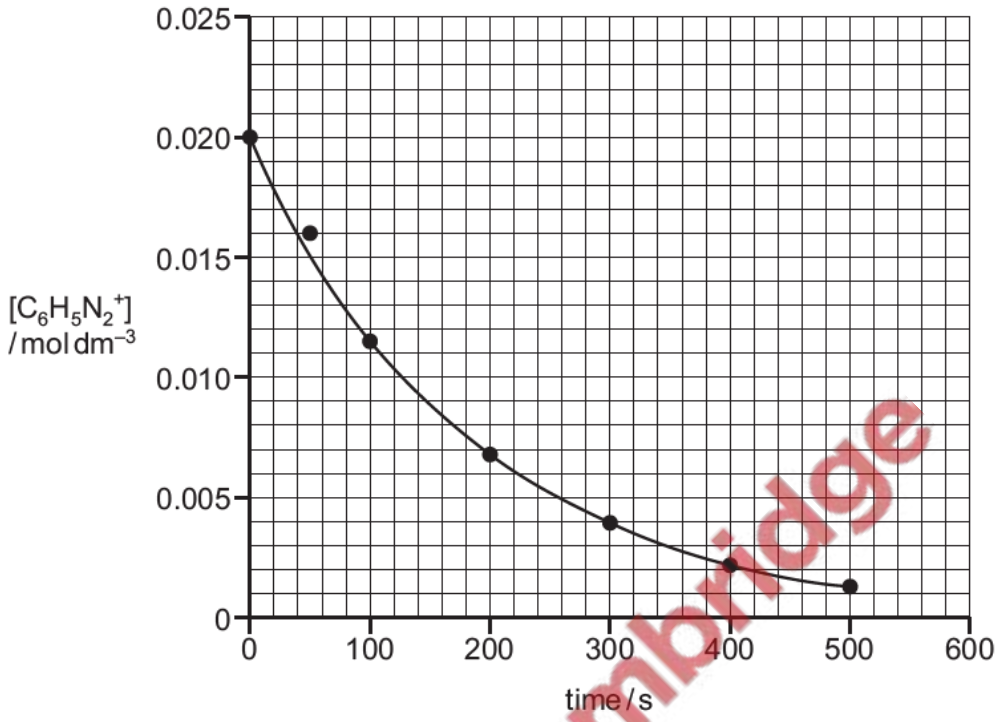
rate = ... [1]

(iv) Use your rate equation and the data from experiment 1 to calculate the rate constant, k, for this reaction.
Include the units of k.

k = ... units ..... [2]

(b) The decomposition of benzenediazonium ions, $C_6H_5N_2^+$, using a large excess of water, is a first-order reaction.

The graph shows the results obtained.



(i) Draw the structure of the organic product formed in this reaction.

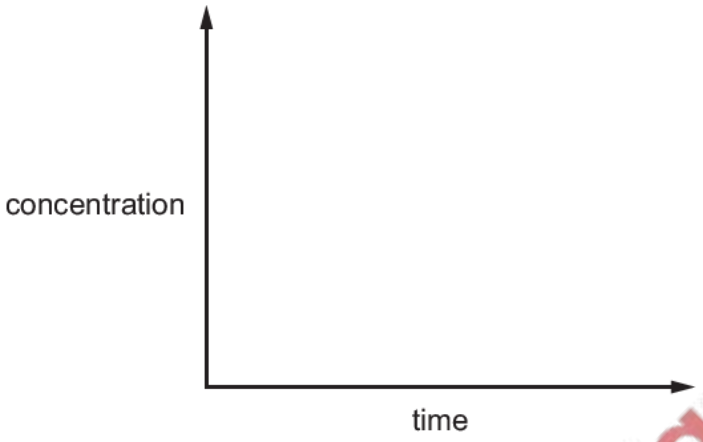
[1]

(ii) Use the graph to determine the rate of reaction at 100 s. Show your working.

rate = ... $mol\,dm^{-3}\,s^{-1}$ [1]

(c) Sketch a concentration–time graph for a **zero-order** reaction.

Use your graph to suggest how successive half-lives for a zero-order reaction vary as the concentration of a reactant decreases. Indicate this by placing a tick (✓) in the appropriate box in the table.



successive half-lives decrease	no change in successive half-lives	successive half-lives increase

[1]

[Total: 9]