Marks 5

- The following data were obtained for the iodide-catalysed decomposition of hydrogen peroxide, H_2O_2.

Experiment	$[I^-]$(M)	$[H_2O_2]$ (M)	Initial rate(M s^{-1})
1	0.375	0	0
2	0.375	0.235	0.000324
3	0.375	0.470	0.000657
4	0.375	0.705	0.001024
5	0.375	0.940	0.001487
6	0	0.948	0
7	0.050	0.948	0.00045
8	0.100	0.948	0.00095
9	0.150	0.948	0.00140
10	0.200	0.948	0.00193

Determine the rate law from these data.

The rate law is of the form, rate = $k[I^-]^x[H_2O_2]^y$.

Between experiments 2 and 3, $[I^-]$ is unchanged. The increase in rate is due to the increase in $[H_2O_2]$:

$$\frac{\text{rate (3)}}{\text{rate (2)}} = \frac{\cancel{k(0.375)^x}(0.470)^y}{\cancel{k(0.375)^x}(0.235)^y} = \frac{(0.470)^y}{(0.235)^y} = \frac{0.000657}{0.000324} \qquad \text{so } y = 1$$

Between experiments 7 and 8, $[H_2O_2]$ is unchanged. The increase in rate is due to the increase in $[I^-]$:

$$\frac{\text{rate (8)}}{\text{rate (7)}} = \frac{k(0.100)^x\cancel{(0.948)^y}}{k(0.050)^x\cancel{(0.948)^y}} = \frac{(0.100)^y}{(0.050)^y} = \frac{0.00095}{0.00045} \qquad \text{so } x = 1$$

Hence,

rate = $k[I^-][H_2O_2]$

Use the data from Experiment 10 to calculate the rate constant for this reaction.

In experiment 10, $[I^-]$ = 0.200 M, $[H_2O_2]$ = 0.948 M and rate = 0.00193 M s^{-1}. Hence:

k = rate / $[I^-][H_2O_2]$ = (0.00193 M s^{-1}) / (0.200 M × 0.948 M)

= 0.0102 M^{-1} s^{-1}

k = **0.0102 M^{-1} s^{-1}**

ANSWER CONTINUES ON THE NEXT PAGE

Iodide ion is used as a catalyst in this reaction. What is the role of a catalyst in a chemical reaction?

A catalyst provides a reaction pathway of lower activation energy and hence increases the rate of the reaction. It is unchanged at the end of the reaction and does not change the equilibrium position.