Marks 3

- Peroxydisulfate and iodide ions react according to the following equation.

$$S_2O_8^{2-}(aq) + 3I^-(aq) \rightarrow 2SO_4^{2-}(aq) + I_3^-(aq)$$

The following rate data were collected at room temperature.

Experiment	$[S_2O_8^{2-}(aq)]_0$ (M)	$[I^-(aq)]_0$ (M)	Initial rate (mol L^{-1} s^{-1})
1	0.080	0.034	2.2×10^{-4}
2	0.080	0.017	1.1×10^{-4}
3	0.160	0.017	2.2×10^{-4}

Determine the rate law for the reaction.

Between experiments (1) and (2), $[S_2O_8^{2-}(aq)]_0$ is constant and $[I^-(aq)]_0$ is halved from 0.034 M to 0.017 M. This halves the initial rate. The rate is directly proportional to $[I^-(aq)]$: the reaction is first order with respect to $[I^-(aq)]$.

Between experiments (2) and (3), $[I^-(aq)]_0$ is constant and $[S_2O_8^{2-}(aq)]_0$ is doubled from 0.080 M to 0.160 M. This doubles the initial rate. The rate is directly proportional to $[S_2O_8^{2-}(aq)]$: the reaction is first order with respect to $[S_2O_8^{2-}(aq)]$.

Hence:

$$\textbf{rate} = k[S_2O_8^{2-}(aq)][I^-(aq)]$$

Calculate the value of the rate constant at room temperature.

From experiment (1), the rate = 2.2×10^{-4} M s^{-1} when $[S_2O_8^{2-}(aq)]$ = 0.080 M and $[I^-(aq)]$ = 0.034 M. Hence:

$$2.2 \times 10^{-4}\ \text{M s}^{-1} = k(0.080\ \text{M})(0.034\ \text{M})$$

$$k = 0.081\ \text{M}^{-1}\ \text{s}^{-1}$$

(Note that the units are worked out by requiring that the units on the two sides of the rate law are the same.)

Answer: **$0.081\ \text{M}^{-1}\ \text{s}^{-1}$**