Marks 4

- The following reaction is run from 4 different starting positions.

$$H_2SeO_3 + 6I^- + 4H^+ \rightarrow Se + 2I_3^- + 3H_2O$$

Experiment	Initial $[H_2SeO_3]$ (mol L^{-1})	Initial $[I^-]$ (mol L^{-1})	Initial $[H^+]$ (mol L^{-1})	Initial rate of increase of $[I_3^-]$ (mol L^{-1} s^{-1})
1	0.100	0.100	0.100	1.000
2	0.100	0.075	0.100	0.422
3	0.075	0.100	0.100	0.750
4	0.100	0.075	0.075	0.237

Determine the rate law for the reaction.

The rate law is of the form:

$$rate = k[H_2SeO_3]^x[I^-]^y[H^+]^z$$

Between experiments (1) and (3), $[I^-]$ and $[H^+]$ are both constant. The change in rate is due to the change in $[H_2SeO_3]$:

$$\frac{\text{rate (3)}}{\text{rate (1)}} = \frac{\cancel{k}(0.075)^x\cancel{(0.100)^y}\cancel{(0.100)^z}}{\cancel{k}(0.100)^x\cancel{(0.100)^y}\cancel{(0.100)^z}} = \frac{(0.075)^x}{(0.100)^x} = \frac{0.750}{1.000} \quad \text{so } x = 1$$

Between experiments (1) and (2), $[H_2SeO_3]$ and $[H^+]$ are both constant. The change in rate is due to the change in $[I^-]$:

$$\frac{\text{rate (2)}}{\text{rate (1)}} = \frac{\cancel{k(0.100)^x}(0.075)^y\cancel{(0.100)^z}}{\cancel{k(0.100)^x}(0.100)^y\cancel{(0.100)^z}} = \frac{(0.075)^y}{(0.100)^y} = \frac{0.422}{1.000} \quad \text{so } y = 3$$

Between experiments (2) and (4), $[H_2SeO_3]$ and $[I^-]$ are both constant. The change in rate is due to the change in $[H^+]$:

$$\frac{\text{rate (4)}}{\text{rate (2)}} = \frac{\cancel{k(0.100)^x(0.075)^y}(0.075)^z}{\cancel{k(0.100)^x(0.075)^y}(0.100)^z} = \frac{(0.075)^z}{(0.100)^z} = \frac{0.237}{0.422} \quad \text{so } z = 2$$

Overall:

$$rate = k[H_2SeO_3][I^-]^3[H^+]^2$$

What is the value of the rate constant?

Using, for example, experiment (1), the initial rate of increase of $[I_3^-]$ = 1.000 mol L^{-1} s^{-1}. As $2I_3^-$ are produced in the reaction:

rate of reaction = ½ × rate of increase of I_3^- = 0.5000 mol L^{-1} s^{-1}

ANSWER CONTINUES ON THE NEXT PAGE

$$\textbf{\textit{rate}} = \boldsymbol{k[H_2SeO_3][I^-]^3[H^+]^2}$$
$$= \boldsymbol{k\ (0.100\ mol\ L^{-1})(0.100\ mol\ L^{-1})^3(0.100\ mol^{-1})^2 = 0.5000\ mol\ L^{-1}\ s^{-1}}$$

Hence:

$$\boldsymbol{k = (0.5000\ mol\ L^{-1}\ s^{-1})\ /\ (1.00 \times 10^{-6}\ mol^6\ L^{-6}) = 5.00 \times 10^5\ L^5\ mol^{-5}\ s^{-1}}$$

Answer: $\boldsymbol{5.00 \times 10^5\ L^5\ mol^{-5}\ s^{-1}}$