Marks 4

- The following initial rate data have been obtained for the gas phase reaction of nitrogen dioxide, $NO_2(g)$, and ozone, $O_3(g)$, at 300 K.

$$2NO_2(g) + O_3(g) \rightarrow N_2O_5(g) + O_2(g)$$

$[NO_2(g)]$ M	$[O_3(g)]$ M	Rate M s^{-1}
0.65	0.80	2.61×10^4
1.10	0.80	4.40×10^4
1.10	1.60	8.80×10^4

What is the order of this reaction with respect to each reagent?

Between the first and second experiments, $[O_3(g)]$ is kept constant and $[NO_2(g)]$ increases by about 1.7. This leads to the rate also increasing by about 1.7. The reaction is therefore first order with respect to $[NO_2(g)]$.

Between the second and third experiments, $[NO_2(g)]$ is kept constant and $[O_3(g)]$ is doubled. This also leads to a doubling of the rate. The reaction is therefore also first order with respect to $[O_3(g)]$.

$$\textbf{rate} = k[NO_2(g)][O_3(g)]$$

What is the rate constant of the reaction?

Using the rate law rate = $k[NO_2(g)][O_3(g)]$, k can be worked out using any of the three experiments. For example, using the first experiment:

$$\text{rate} = 2.61 \times 10^4 = k[NO_2(g)][O_3(g)] = k \times (0.65) \times (0.80)$$

$$\text{Hence } k = \frac{2.61\times10^4}{(0.65)(0.80)} = 5.0\times10^4$$

The units of k are obtained by balancing those in the rate law: the rate has units of M s^{-1} and the concentrations both have units of M. Hence:

$$\text{Units of } k = \frac{M\ s^{-1}}{(M)(M)} = M^{-1}\ s^{-1}$$

Answer: **5.0×10^4 M^{-1} s^{-1}**