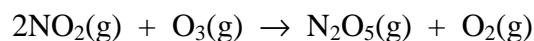


- The following initial rate data have been obtained for the gas phase reaction of nitrogen dioxide, $\text{NO}_2(\text{g})$, and ozone, $\text{O}_3(\text{g})$, at 300 K.



$[\text{NO}_2(\text{g})]$ M	$[\text{O}_3(\text{g})]$ 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, $[\text{O}_3(\text{g})]$ is kept constant and $[\text{NO}_2(\text{g})]$ 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 $[\text{NO}_2(\text{g})]$.

Between the second and third experiments, $[\text{NO}_2(\text{g})]$ is kept constant and $[\text{O}_3(\text{g})]$ is doubled. This also leads to a doubling of the rate. The reaction is therefore also first order with respect to $[\text{O}_3(\text{g})]$.

$$\text{rate} = k[\text{NO}_2(\text{g})][\text{O}_3(\text{g})]$$

What is the rate constant of the reaction?

Using the rate law $\text{rate} = k[\text{NO}_2(\text{g})][\text{O}_3(\text{g})]$, k can be worked out using any of the three experiments. For example, using the first experiment:

$$\text{rate} = 2.61 \times 10^4 = k[\text{NO}_2(\text{g})][\text{O}_3(\text{g})] = k \times (0.65) \times (0.80)$$

$$\text{Hence } k = \frac{2.61 \times 10^4}{(0.65)(0.80)} = 5.0 \times 10^4$$

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:

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

Answer: $5.0 \times 10^4 \text{ M}^{-1} \text{ s}^{-1}$

Marks
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