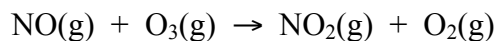


- Nitric oxide reacts with ozone according to the following equation.



The following rate data were collected at a specified temperature.

Marks
3

Trial	Initial[NO] (M)	Initial [O ₃] (M)	Initial rate of reaction (M s ⁻¹)
1	2.1×10^{-6}	2.1×10^{-6}	1.6×10^{-5}
2	6.3×10^{-6}	2.1×10^{-6}	4.8×10^{-5}
3	6.3×10^{-6}	4.2×10^{-6}	9.6×10^{-5}

What is the experimental rate law for the reaction?

$$\text{Rate} = k[\text{NO}]^x[\text{O}_3]^y$$

Between trials (1) and (2), [O₃]_{initial} is constant and [NO]_{initial} is tripled. This leads to the rate tripling: $x = 1$.

Between trials (2) and (3), [NO]_{initial} is constant and [O₃]_{initial} is doubled. This leads to the rate doubling: $y = 1$.

Hence:

$$\text{rate} = k[\text{NO}][\text{O}_3]$$

What is the value of the rate constant of this reaction?

Using trial (1),

$$k = \frac{\text{rate}}{[\text{NO}][\text{O}_3]} = \frac{1.6 \times 10^{-5} \text{ M s}^{-1}}{(2.1 \times 10^{-6} \text{ M})(2.1 \times 10^{-6} \text{ M})} = 3.6 \times 10^6 \text{ M}^{-1} \text{ s}^{-1}$$

Answer: $3.6 \times 10^6 \text{ M}^{-1} \text{ s}^{-1}$