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

$$NO(g) + O_3(g) \rightarrow NO_2(g) + O_2(g)$$

The following rate data were collected at a specified temperature.

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?

Rate = $k[NO]^x[O_3]^y$

Between trials (1) and (2), $[O_3]_{initial}$ 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 $[O_3]_{initial}$ is doubled. This leads to the rate doubling: $y = 1$.

Hence:

rate = $k[NO][O_3]$

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

Using trial (1),

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

Answer: **3.6×10^{6} M^{-1} s^{-1}**