Marks 3

- Given the following experimental data, find the rate law and the rate constant for the following reaction:

$$NO(g) + NO_2(g) + O_2(g) \rightarrow N_2O_5(g)$$

Run	$[NO(g)]$ / M	$[NO_2(g)]$ / M	$[O_2(g)]$ / M	Rate / M s^{-1}
1	0.10	0.10	0.10	2.1×10^{-2}
2	0.20	0.10	0.10	4.2×10^{-2}
3	0.20	0.30	0.20	1.26×10^{-1}
4	0.10	0.10	0.20	2.1×10^{-2}

Between run (1) and (2), $[NO(g)]$ is doubled but $[NO_2(g)]$ and $[O_2(g)]$ are kept constant. The doubling in $[NO(g)]$ causes a doubling in the rate: the rate is proportional to $[NO(g)]^1$.

Between run (1) and (4), $[O_2(g)]$ is doubled but $[NO(g)]$ and $[NO_2(g)]$ are kept constant. The doubling in $[O_2(g)]$ causes no change in the rate: the rate is independent of $[O_2(g)]$.

Between run (2) and (3), $[NO_2(g)]$ is trebled but $[NO(g)]$ is kept constant. Although $[O_2(g)]$ is doubled, this has no effect on the rate (see directly above). The trebling in $[NO_2(g)]$ causes the rate to treble: the rate is proportional to $[NO_2(g)]^1$

Overall:

$$\text{rate} = k[NO(g)][NO_2(g)]$$

Using run (1), rate = 2.1×10^{-2} M and $[NO(g)] = [NO_2(g)] = 0.10$ M. Hence:

$$\begin{aligned} k &= \text{rate} / [NO(g)][NO_2(g)] \\ &= 2.1 \times 10^{-2} \text{ M s}^{-1} / (0.10 \text{ M})^2 \\ &= 2.1 \text{ M}^{-1} \text{ s}^{-1} \end{aligned}$$

Rate = $k[NO(g)][NO_2(g)]$	$k = 2.1$ M^{-1} s^{-1}

ANSWER CONTINUES ON THE NEXT PAGE

- The rate constant for a reaction is 5.0×10^{-3} s^{-1} at 215 °C and 1.2×10^{-1} s^{-1} at 452 °C. What is the activation energy of the reaction in kJ mol^{-1}? **3**

The rate constant varies with temperature according to the Arrhenius equation:

$$\ln\left(\frac{k_2}{k_1}\right) = \frac{E_a}{R}\left(\frac{1}{T_1} - \frac{1}{T_2}\right)$$

At $T_1 = (215 + 273)$ K = 488 K and $k_1 = 5.0 \times 10^{-3}$ s^{-1}. At $T_2 = (452 + 273)$ K = 725 K and $k_2 = 1.2 \times 10^{-1}$ s^{-1}. Hence:

$$\ln\left(\frac{1.2 \times 10^{-1}}{5.0 \times 10^{-3}}\right) = \frac{E_a}{8.314}\left(\frac{1}{488} - \frac{1}{725}\right)$$

E_a = 39 kJ mol^{-1}

Answer: **39 kJ mol^{-1}**

What is the rate constant for this reaction at 100 °C?

Using T_1 = 488 K and $k_1 = 5.0 \times 10^{-3}$ s^{-1}, when $T_2 = (100 + 273)$ K = 373 K:

$$\ln\left(\frac{k_2}{5.0 \times 10^{-3}}\right) = \frac{39 \times 10^3}{8.314}\left(\frac{1}{488} - \frac{1}{373}\right)$$

$k_2 = 2.5 \times 10^{-4}$ s^{-1}

Answer: **2.5×10^{-4} s^{-1}**