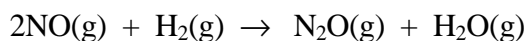


- Nitric oxide, a noxious pollutant, and hydrogen react to give nitrous oxide and water according to the following equation.



The following rate data were collected at 225 °C.

Marks
5

Experiment	[NO] ₀ (M)	[H ₂] ₀ (M)	Initial rate (d[NO]/dt, M s ⁻¹)
1	6.4 × 10 ⁻³	2.2 × 10 ⁻³	2.6 × 10 ⁻⁵
2	1.3 × 10 ⁻²	2.2 × 10 ⁻³	1.0 × 10 ⁻⁴
3	6.4 × 10 ⁻³	4.4 × 10 ⁻³	5.1 × 10 ⁻⁵

Determine the rate law for the reaction.

Between experiments (1) and (2), [H₂]₀ is constant and [NO]₀ doubles. As the rate increases by a factor of $\frac{1.0 \times 10^{-4}}{2.6 \times 10^{-5}} = 3.8 \sim 4$, the rate is second order with respect to [NO]₀.

Between experiments (1) and (3), [H₂]₀ doubles and [NO]₀ is constant. As the rate increases by a factor of $\frac{5.1 \times 10^{-5}}{2.6 \times 10^{-5}} = 2.0$, the rate is first order with respect to [NO]₀.

Overall,

$$\text{rate} = k[\text{NO}]^2[\text{H}_2]$$

Calculate the value of the rate constant at 225 °C.

Using experiment (1),

$$2.6 \times 10^{-5} \text{ M s}^{-1} = k \times (6.4 \times 10^{-3} \text{ M})^2 \times (2.2 \times 10^{-3} \text{ M})$$

$$k = 2.9 \times 10^2 \text{ M}^{-2} \text{ s}^{-1}$$

$$\text{Answer: } k = 2.9 \times 10^2 \text{ M}^{-2} \text{ s}^{-1}$$

Calculate the rate of appearance of N₂O when [NO] = [H₂] = 6.6 × 10⁻³ M.

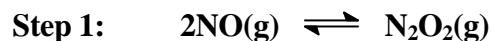
$$\begin{aligned} \text{rate of disappearance of NO} &= k[\text{NO}]^2[\text{H}_2] \\ &= (2.9 \times 10^2 \text{ M}^{-2} \text{ s}^{-1}) \times (6.6 \times 10^{-3} \text{ M})^2 \times (6.6 \times 10^{-3} \text{ M}) \\ &= 8.3 \times 10^{-5} \text{ M s}^{-1} \end{aligned}$$

The rate of appearance of N₂O is *half* this value as, from the chemical equation, NO is disappearing at twice the rate than N₂O is appearing.

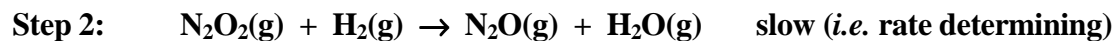
$$\text{Answer: } 4.1 \times 10^{-5} \text{ M s}^{-1}$$

ANSWER CONTINUES ON THE NEXT PAGE

Suggest a possible mechanism for the reaction based on the form of the rate law.
Explain your answer.



This is a fast equilibrium and so $K = \frac{[\text{N}_2\text{O}_2(\text{g})]}{[\text{NO}(\text{g})]^2}$ **or** $[\text{N}_2\text{O}_2(\text{g})] = K[\text{NO}(\text{g})]^2$



As this is rate determining,

$$\text{rate} = k_2[\text{N}_2\text{O}_2][\text{H}_2]$$

As $[\text{N}_2\text{O}_2(\text{g})] = K[\text{NO}(\text{g})]^2$, **this can be rewritten as,**

$$\text{rate} = kK[\text{NO}]^2[\text{H}_2]$$

This is consistent with the experimentally determined rate law with $k_{\text{exp}} = k_2K$.