Marks
5

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

$$2NO(g) + H_2(g) \rightarrow N_2O(g) + H_2O(g)$$

The following rate data were collected at 225 °C.

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), $[H_2]_0$ 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 ~ 4, the rate is second order with respect to $[NO]_0$.

Between experiments (1) and (3), $[H_2]_0$ 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,

rate = $k[NO]^2[H_2]$

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

Using experiment (1),

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

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

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

Calculate the rate of appearance of N_2O when [NO] = $[H_2]$ = 6.6×10^{-3} M.

rate of disappearance of NO = $k[NO]^2[H_2]$
= $(2.9 \times 10^2\ M^2\ s^{-1}) \times (6.6 \times 10^{-3}\ M)^2 \times (6.6 \times 10^{-3}\ M)$
= 8.3×10^{-5} M s^{-1}

The rate of appearance of N_2O is *half* this value as, from the chemical equation, NO is disappearing at twice the rate than N_2O is appearing.

Answer: **4.1×10^{-5} 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.

Step 1: $2NO(g) \rightleftharpoons N_2O_2(g)$

This is a fast equilibrium and so $K = \frac{[N_2O_2(g)]}{[NO(g)]^2}$ **or** $[N_2O_2(g)] = K[NO(g)]^2$

Step 2: $N_2O_2(g) + H_2(g) \rightarrow N_2O(g) + H_2O(g)$ **slow (*i.e.* rate determining)**

As this is rate determining,

$$\text{rate} = k_2[N_2O_2][H_2]$$

As $[N_2O_2(g)] = K[NO(g)]^2$**, this can be rewritten as,**

$$\text{rate} = kK[NO]^2[H_2]$$

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