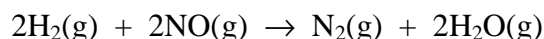


- Hydrogenation of NO to N₂ and water is a potential means of reducing smog-forming NO_x gases:



The initial rates of this reaction at constant temperature were determined at the following combination of initial pressures (P_0).

Experiment	$P_0 \text{ H}_2$ (kPa)	$P_0 \text{ NO}$ (kPa)	Rate (kPa s ⁻¹)
1	53.3	40.0	0.137
2	53.3	20.3	0.033
3	38.5	53.3	0.213
4	19.6	53.3	0.105

What is the order of the reaction? Show all working.

The rate law is in the form:

$$\text{rate} = k(P(\text{H}_2))^n(P(\text{NO}))^m$$

The order of the reaction is equal to $n + m$.

Between experiments (1) and (2), $P(\text{H}_2)$ is constant. $P(\text{NO})$ is decreased from 40.0 kPa to 20.3 kPa. It is almost halved. This causes the rate to drop from 0.137 kPa s⁻¹ to 0.033 kPa s⁻¹. It decreases by $(0.033 / 0.137) \% = 24.1 \%$

As halving the amount of NO reduces the rate by a factor of 4, $m = 2$.

Between experiments (3) and (4), $P(\text{NO})$ is constant. $P(\text{H}_2)$ is decreased from 38.5 kPa to 19.6 kPa. It is almost halved. This causes the rate to drop from 0.213 kPa s⁻¹ to 0.105 kPa s⁻¹. It decreases by $(0.105 / 0.213) \% = 49.2 \%$

As halving the amount of NO reduces the rate by a factor of 2, $n = 1$.

The order of the reaction is equal to $n + m$ so is 3.

Answer: 3

What is the value of the rate constant?

From above, the rate law is:

$$\text{rate} = k(P(\text{H}_2))(P(\text{NO}))^2$$

From experiment (1), rate = 0.137 kPa s⁻¹ when $P(\text{H}_2) = 53.3$ kPa and $P(\text{NO}) = 40.0$ kPa. Hence:

$$\begin{aligned} k &= \text{rate} / [(P(\text{H}_2))(P(\text{NO}))^2] \\ &= (0.137 \text{ kPa s}^{-1}) / (53.3 \text{ kPa})(40.0 \text{ kPa})^2 \\ &= 1.61 \times 10^{-6} \text{ kPa}^{-2} \text{ s}^{-1} \end{aligned}$$

(Note that the units can be derived from the equation and would be required to gain the marks in this question.)

Answer: $1.61 \times 10^{-6} \text{ kPa}^{-2} \text{ s}^{-1}$