Marks
3

- Hydrogenation of NO to N_2 and water is a potential means of reducing smog-forming NO_x gases:

$$2H_2(g) + 2NO(g) \rightarrow N_2(g) + 2H_2O(g)$$

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

Experiment	P_0 H_2 (kPa)	P_0 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:

rate = $k(P(H_2))^n(P(NO))^m$

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

Between experiments (1) and (2), $P(H_2)$ is constant. P(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 is 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(NO) is constant. $P(H_2)$ 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 is 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:

rate = $k(P(H_2))(P(NO))^2$

From experiment (1), rate = 0.137 kPa s^{-1} when $P(H_2)$ = 53.3 kPa and P(NO) = 40.0 kPa. Hence:

$$k = \text{rate} / [(P(H_2))(P(NO))^2]$$

$$= (0.137 \text{ kPa s}^{-1}) / (53.3 \text{ kPa})(40.0 \text{ kPa})^2$$

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

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

Answer: **1.61×10^{-6} kPa^{-2} s^{-1}**