Marks 3

- Hydrogenation of nitric oxide to nitrogen and water is a potential means of reducing smog-forming NO_x gases:

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

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

Run	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

Derive an expression for the rate law for this reaction.

Between Run 1 and 2, P_0 (H_2) is constant and P_0 (NO) is halved. This causes the rate to be reduced by a factor of four. The rate is second order with respect to NO.

Between Run 3 and 4, P_0 (H_2) is halved and P_0 (NO) is constant. This causes the rate to be reduced by a factor of two. The rate is first order with respect to H_2.

Overall,

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

Answer: **rate = $k \times P(H_2) \times P^2(NO)$**

Calculate the value of the rate constant.

Using Run 1, rate = 0.137 kPa s^{-1} when $P(H_2)$ = 53.3 kPa and $P(NO)$ = 40.0 kPa:

0.137 kPa s^{-1} = k × 53.3 kPa × (40.0 kPa)^2

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

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

What is the order of the reaction? **1 (H_2) + 2 (NO) = 3 (third order)**