Marks 4

- The following data were obtained for the reaction between gaseous nitric oxide and hydrogen at 1280 °C.

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

Experiment number	INITIAL [NO] (M)	INITIAL [H_2] (M)	INITIAL REACTION RATE (M min^{-1})
1	5.0×10^{-3}	2.0×10^{-3}	1.3×10^{-5}
2	1.0×10^{-2}	2.0×10^{-3}	5.0×10^{-5}
3	1.0×10^{-2}	4.0×10^{-3}	1.0×10^{-4}

Deduce the rate law for this reaction and calculate the value of the rate constant.

RATE LAW	RATE CONSTANT
Between experiments 1 and 2, [H_2] is constant and [NO] is doubled. The rate increases by a factor of four. The reaction is second order with respect to NO. **Between experiments 2 and 3, [NO] is constant and [H_2] is doubled. The rate increases by a factor of two. The reaction is first order with respect to H_2.** **rate = $k[NO]^2[H_2]$**	**Using experiment 1,** **rate = $k[NO]^2[H_2]$** **$1.3\times10^{-5} = k \times (5.0\times10^{-3})^2 \times (2.0\times10^{-3})$** **$k$ = 260 M^{-2} min^{-1}** **The units of k can be deduced from balancing those of the other terms:** **M min^{-1} = (units of k) × $(M)^2$ × (M)**
Answer: **rate = $k[NO]^2[H_2]$**	Answer: **260 M^{-2} min^{-1}**

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