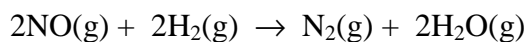


- The following data were obtained for the reaction between gaseous nitric oxide and hydrogen at 1280 °C.

Marks
4



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.

<p>RATE LAW</p> <p>Between experiments 1 and 2, [H₂] is constant and [NO] is doubled. The rate increases by a factor of four. The reaction is second order with respect to NO.</p> <p>Between experiments 2 and 3, [NO] is constant and [H₂] is doubled. The rate increases by a factor of two. The reaction is first order with respect to H₂.</p> <p>rate = $k[\text{NO}]^2[\text{H}_2]$</p>	<p>RATE CONSTANT</p> <p>Using experiment 1,</p> <p>rate = $k[\text{NO}]^2[\text{H}_2]$</p> <p>$1.3 \times 10^{-5} = k \times (5.0 \times 10^{-3})^2 \times (2.0 \times 10^{-3})$</p> <p>$k = 260 \text{ M}^{-2} \text{ min}^{-1}$</p> <p>The units of k can be deduced from balancing those of the other terms:</p> <p>$\text{M min}^{-1} = (\text{units of } k) \times (\text{M})^2 \times (\text{M})$</p>
Answer: rate = $k[\text{NO}]^2[\text{H}_2]$	Answer: $260 \text{ M}^{-2} \text{ min}^{-1}$

THE REMAINDER OF THIS PAGE IS FOR ROUGH WORKING ONLY.