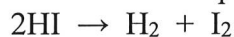


1) At a certain temperature the following data were collected for the decomposition of HI.



Determine the rate law for the reaction.

$$\text{Rate} = k[\text{HI}]^2$$

Experiment	Initial [HI] (mol L ⁻¹)	Initial rate of reaction (mol L ⁻¹ s ⁻¹)
1	1.0 × 10 ⁻²	4.0 × 10 ⁻⁶
2	2.0 × 10 ⁻²	1.6 × 10 ⁻⁵
3	3.0 × 10 ⁻²	3.6 × 10 ⁻⁵

What is the value of the rate constant for the decomposition of HI? Include units in your answer.

$$k = \frac{\text{Rate}}{[\text{HI}]^2} = \frac{4.0 \times 10^{-6} \text{ M/s}}{(1.0 \times 10^{-2} \text{ M})^2} = 0.040 \text{ M}^{-1} \text{ s}^{-1}$$

2) Nitrogen monoxide, a noxious pollutant, reacts with oxygen to produce nitrogen dioxide, another toxic



The rate data in the table was collected at 225°C.

Determine the rate law for the reaction

$$\text{Rate} = k[\text{NO}]^2[\text{O}_2]$$

Exp	[NO] ₀ (M)	[O ₂] ₀ (M)	Initial Rate, -Δ[O ₂]/Δt (M s ⁻¹)
1	1.3 × 10 ⁻²	1.1 × 10 ⁻²	1.6 × 10 ⁻³
2	1.3 × 10 ⁻²	2.2 × 10 ⁻²	3.2 × 10 ⁻³
3	2.6 × 10 ⁻²	1.1 × 10 ⁻²	6.4 × 10 ⁻³

Calculate the value of the rate constant at 225°C.

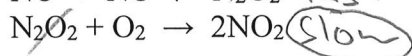
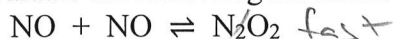
$$k = \frac{\text{Rate}}{[\text{NO}]^2[\text{O}_2]} = \frac{1.6 \times 10^{-3} \text{ M s}^{-1}}{(1.3 \times 10^{-2} \text{ M})^2 (1.1 \times 10^{-2} \text{ M})} = 860 \text{ M}^{-2} \text{ s}^{-1}$$

Calculate the rate of appearance of NO₂ when [NO] = [O₂] = 6.5 × 10⁻³ M.

$$\text{Rate} = (860 \text{ M}^{-2} \text{ s}^{-1}) (6.5 \times 10^{-3} \text{ M})^2 (6.5 \times 10^{-3} \text{ M}) = 2.4 \times 10^{-4} \text{ M s}^{-1} \text{ (for } -\text{O}_2\text{)}$$

$$\begin{array}{l} \text{O}_2 \text{ and NO}_2 \\ 1 : 2 \end{array} \quad = 2.4 \times 10^{-4} \times 2 = 4.8 \times 10^{-4} \text{ M s}^{-1}$$

Evaluate the following mechanism for the reaction. Which is the slow step? Does it match the rate law?



NO is involved in rate law + 2nd order

O₂ is involved in rate law + 1st order

N₂O₂ is an intermediate - not in rate law

3) Nitric oxide, a noxious pollutant, and hydrogen react to give nitrous oxide and water according to the reaction: $2\text{NO}(\text{g}) + \text{H}_2(\text{g}) \rightarrow \text{N}_2\text{O}(\text{g}) + \text{H}_2\text{O}(\text{g})$

The following rate data was collected at 225°C.

Determine the rate law for the reaction.

$$\text{Rate} = k[\text{NO}]^2[\text{H}_2]$$

Exp	$[\text{NO}]_0$ (M)	$[\text{H}_2]_0$ (M)	Initial Rate, $-\Delta[\text{NO}]/\Delta t$ (M s^{-1})
1	6.4×10^{-3}	2.2×10^{-3}	2.6×10^{-5}
2	1.3×10^{-2}	2.2×10^{-3}	1.0×10^{-4}
3	6.4×10^{-3}	4.4×10^{-3}	5.1×10^{-5}

Calculate the value of the rate constant at 225°C.

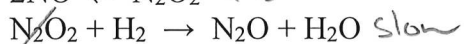
$$k = \frac{\text{Rate}}{[\text{NO}]^2[\text{H}_2]} = \frac{2.6 \times 10^{-5} \text{ M s}^{-1}}{(6.4 \times 10^{-3} \text{ M})^2 (2.2 \times 10^{-3} \text{ M})} = 290 \text{ M}^{-2} \text{ s}^{-1}$$

Calculate the rate of appearance of N_2O when $[\text{NO}] = [\text{H}_2] = 6.6 \times 10^{-3} \text{ M}$.

$$\text{Rate} = (290 \text{ M}^{-2} \text{ s}^{-1}) (6.6 \times 10^{-3} \text{ M})^2 (6.6 \times 10^{-3} \text{ M}) = 8.3 \times 10^{-5} \text{ M s}^{-1}$$

$2:1 \text{ ratio} \rightarrow 8.3 \times 10^{-5} \div 2 = 4.2 \times 10^{-5} \text{ M s}^{-1}$

Evaluate the following mechanism does it match the rate law. Which step is fast?



N_2O_2 is an intermediate, 2nd Order w/ respect to NO and 1st Order w/ respect to H_2

4) The major pollutants NO (g), CO (g), NO_2 (g) and CO_2 (g) are emitted by cars and can react according to the following equation: $\text{NO}_2(\text{g}) + \text{CO}(\text{g}) \rightarrow \text{NO}(\text{g}) + \text{CO}_2(\text{g})$ The rate data was collected at 225°C.

Determine the rate law for the reaction.

$$\text{Rate} = k[\text{NO}_2]^2$$

Exp	$[\text{NO}_2]_0$ (M)	$[\text{CO}]_0$ (M)	Initial Rate, $-\Delta[\text{CO}_2]/\Delta t$ (M s^{-1})
1	0.263	0.826	1.44×10^{-5}
2	0.263	0.413	1.44×10^{-5}
3	0.526	0.413	5.76×10^{-5}

Calculate the value of the rate constant at 225°C.

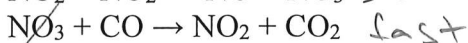
$$k = \frac{\text{Rate}}{[\text{NO}_2]^2} = \frac{1.44 \times 10^{-5} \text{ M s}^{-1}}{(0.263 \text{ M})^2} = 2.08 \times 10^{-4} \text{ M}^{-1} \text{ s}^{-1}$$

Calculate the rate of appearance of CO_2 when $[\text{NO}_2] = [\text{CO}] = 0.500 \text{ M}$.

$$\text{Rate} = (2.08 \times 10^{-4} \text{ M}^{-1} \text{ s}^{-1}) (0.500 \text{ M})^2 = 5.20 \times 10^{-5} \text{ M s}^{-1}$$

★ 1:1 ratio

Evaluate the following mechanism for the reaction based on the form of the rate law. Explain your answer.



2nd Order w/ respect to NO_2

CO is zero order (not in rate law)