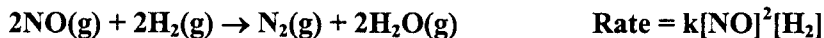


$$\text{rate} = k \quad \text{rate} = k[A] \quad \text{rate} = k[A]^2 \quad [A]_t = -kt + [A]_0 \quad \ln[A]_t = -kt + \ln[A]_0 \quad R = 8.314 \text{ J}/(\text{mol}\cdot\text{K})$$

$$1/[A]_t = kt + 1/[A]_0 \quad t_{1/2} = [A]_0/2k \quad t_{1/2} = 0.693/k \quad t_{1/2} = 1/k[A]_0 \quad \ln \frac{k_1}{k_2} = \frac{E_a}{R} \left( \frac{1}{T_2} - \frac{1}{T_1} \right) \quad e=mc^2$$

1. The equation and the rate law for the reaction between NO(g) and H<sub>2</sub>(g) are respectively:



a. (3 Pts) Which integrated rate law equation would apply to [NO] and what would one plot on each axis of a graph to obtain the rate constant and order?

Equation:  $\frac{1}{[A]_t} = kt + \frac{1}{[A]_0}$  (2<sup>nd</sup> order) Plot (x) time vs (y)  $\frac{1}{[A]_t}$

b. (3 Pts) Which integrated rate law equation would apply to [H<sub>2</sub>] and what would one plot on each axis of a graph to obtain the rate constant and order?

Equation:  $\ln[A]_t = -kt + \ln[A]_0$  (1<sup>st</sup> order) Plot (x) time vs (y)  $\ln[A]_t$

c. (2 Pts) In the rate expression below, one term is incorrect, CIRCLE THE INCORRECT TERM.

$$\text{rate} = -\frac{\Delta[\text{NO}]}{2\Delta t} = -\frac{\Delta[\text{H}_2]}{2\Delta t} = \frac{\Delta[\text{N}_2]}{\Delta t} = \frac{\Delta[\text{H}_2\text{O}]}{\Delta t} = k[\text{NO}]^2[\text{H}_2] \quad (\text{applies to above RXN})$$

2. (5 pts) A chemical reaction that is first order in X is observed to have a rate constant of  $1.2 \times 10^{-2} \text{ s}^{-1}$ . If the initial concentration of X is 2.0 M, what is the concentration of X after 200 s?

$$\begin{aligned} \ln[A]_t &= -kt - \ln[A]_0 \\ &= -1.2 \times 10^{-2} (200) + \ln[2.0] \\ &= -1.7068... \\ [A]_t &= e^{-1.7068...} = \underline{\underline{0.181 \text{ M}}} \end{aligned}$$

3. (4 Pts) In a first-order reaction the half-life is 20.0 minutes. Determine the rate constant, k, in min<sup>-1</sup>.

$$t_{1/2} = \frac{0.693}{k} \quad k = \frac{0.693}{t_{1/2}} = \frac{0.693}{20} = \underline{\underline{0.03466 \text{ min}^{-1}}}$$

4. (8 Pts) A first order reaction was found to have a rate constant of  $2.52 \times 10^{-5} \text{ s}^{-1}$  at 189.7°C. If the activation energy is  $1.6 \times 10^2 \text{ kJ/mol}$ , determine the value of the rate constant at 251.2°C.

$$\begin{aligned} \ln \frac{k_1}{k_2} &= \frac{E_a}{R} \left( \frac{1}{T_2} - \frac{1}{T_1} \right) \\ \ln \frac{k_1}{2.52 \times 10^{-5}} &= \frac{1.6 \times 10^2 \times 10^3 \text{ J/mol}}{8.314 \text{ J/mol K}} \left( \frac{1}{273+189.7} - \frac{1}{273+251.2} \right) = 4.8796... \\ \frac{k_1}{2.52 \times 10^{-5}} &= e^{4.8796...} = 131.58 \\ k_1 &= \underline{\underline{0.003315}} = \underline{\underline{3.32 \times 10^{-3} \text{ s}^{-1}}} \end{aligned}$$