

rate = k rate = k[A] rate = k[A]² [A]_t = -kt + [A]₀ ln[A]_t = -kt + ln[A]₀ R = 8.314 J/(mol·K)

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

SHOW ALL WORK TO RECEIVE CREDIT

1. (8 Pts) Determine the value of the rate constant at 430°C for the reaction: $2\text{NO}_2(\text{g}) \rightarrow 2\text{NO}(\text{g}) + \text{O}_2(\text{g})$. The reaction has an activation energy of 111 kJ/mol and at 400°C the rate constant has a value of $7.8 \text{ L}\cdot\text{mol}^{-1}\cdot\text{s}^{-1}$.

$$\ln \frac{k_2}{k_1} = \frac{E_a}{R} \left(\frac{1}{T_1} - \frac{1}{T_2} \right)$$

$$\ln \frac{k_2}{7.8} = \frac{111 \times 10^3 \frac{\text{J}}{\text{mol}}}{8.314 \frac{\text{J}}{\text{mol}\cdot\text{K}}} \left(\frac{1}{400+273} - \frac{1}{430+273} \right)$$

$$\ln \frac{k_2}{7.8} = 0.84657\dots$$

$$\frac{k_2}{7.8} = e^{0.84657}$$

$$k_2 = 18.1 \text{ L}\cdot\text{mol}^{-1}\cdot\text{s}^{-1}$$

2. (8 Pts) The reaction $2\text{NOCl}(\text{g}) \rightarrow 2\text{NO}(\text{g}) + \text{Cl}_2(\text{g})$ has $k = 9.3 \times 10^{-5} \text{ L}\cdot\text{mol}^{-1}\cdot\text{s}^{-1}$ at 100°C and $k = 1.3 \times 10^{-3} \text{ L}\cdot\text{mol}^{-1}\cdot\text{s}^{-1}$ at 130°C. Determine E_a for the reaction in kJ/mol.

$$\ln \frac{9.3 \times 10^{-5}}{1.3 \times 10^{-3}} = \frac{E_a}{8.314 \frac{\text{J}}{\text{mol}\cdot\text{K}}} \left(\frac{1}{130+273} - \frac{1}{100+273} \right)$$

$$E_a = 109874 \text{ J/mol} \Rightarrow 110 \text{ kJ/mol}$$

3. (5 Pts) The reaction, $\text{SO}_2\text{Cl}_2(\text{g}) \rightarrow \text{SO}_2(\text{g}) + \text{Cl}_2(\text{g})$ has a first order rate constant of $2.2 \times 10^{-5} \text{ s}^{-1}$ at 593K. If the initial concentration of SO_2Cl_2 is 0.0040 M, what will its concentration be after 5.00 hours?

$$\ln [A]_t = -kt + \ln [A]_0$$

$$\ln [A]_t = -2.2 \times 10^{-5} \text{ s}^{-1} \cdot 18000 \text{ s} + \ln [0.0040]$$

$$\ln [A]_t = -5.967\dots \quad [A]_t = e^{-5.967\dots} = 0.00269 \text{ M}$$

4. (4 Pts) A second order reaction has a rate constant of $1.6 \times 10^{-3} \text{ L}\cdot\text{mol}^{-1}\cdot\text{s}^{-1}$ at 700°C. If the initial concentration of the reactant is $3.4 \times 10^{-2} \text{ M}$, how many minutes will it take for the concentration to be reduced to $7.0 \times 10^{-4} \text{ M}$?

$$\frac{1}{[A]_t} = kt + \frac{1}{[A]_0}$$

$$\frac{1}{7.0 \times 10^{-4}} = 1.6 \times 10^{-3} t + \frac{1}{3.4 \times 10^{-2}}$$

$$t = 874474 \text{ sec} \Rightarrow 14600 \text{ min} \quad 1.5 \times 10^4 \text{ min}$$