

show work where possible

1. (6 Pts) The rate constants for the decomposition of a compound are  $6.2 \times 10^{-4} \text{ s}^{-1}$  at  $45^\circ\text{C}$  and  $2.1 \times 10^{-3} \text{ s}^{-1}$  at  $55^\circ\text{C}$ . What is the value of the activation energy for this reaction in kilojoules per mole?

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

$$\ln \frac{6.2 \times 10^{-4}}{2.1 \times 10^{-3}} = \frac{E_a}{8.314 \frac{\text{J}}{\text{mol} \cdot \text{K}}} \left( \frac{1}{328} - \frac{1}{318} \right)$$

$$E_a = \frac{105,000 \text{ J}}{\text{mol}} = 106 \frac{\text{kJ}}{\text{mol}}$$

2. (7 Pts) For the reaction  $\frac{1}{2} \text{N}_2\text{O}_4(\text{g}) \rightarrow \text{NO}_2(\text{g})$   $\Delta H = 28.6 \text{ kJ}$

at  $30^\circ\text{C}$ ,  $k = 5.1 \times 10^{-6} \text{ s}^{-1}$ , and the activation energy is  $54.0 \text{ kJ/mol}$ .

Determine the rate constant at  $55^\circ\text{C}$ .

$$\ln \frac{k_2}{5.1 \times 10^{-6}} = \frac{54.0 \times 10^3 \frac{\text{J}}{\text{mol}}}{8.314 \frac{\text{J}}{\text{mol} \cdot \text{K}}} \left( \frac{1}{303} - \frac{1}{328} \right)$$

$$\ln \frac{k_2}{5.1 \times 10^{-6}} = 1.633 \dots$$

$$\frac{k_2}{5.1 \times 10^{-6}} = e^{1.633 \dots}$$

$$k_2 = 2.6 \times 10^{-7} \text{ s}^{-1}$$

3. (6 Pts) A reaction that is second order in one reactant has a rate constant of  $1.0 \times 10^{-2} \text{ L mol}^{-1} \text{ s}^{-1}$ . If the initial concentration of the reactant is  $0.100 \text{ M}$ , how long will it take for the concentration to become  $0.0500 \text{ M}$ ?

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

$$\frac{1}{0.0500} = 1.0 \times 10^{-2} t + \frac{1}{0.100}$$

$$t = 1000 \text{ s}$$

4. (6 Pts) A first-order chemical reaction is observed to have a rate constant of  $65 \text{ min}^{-1}$ . What is the corresponding half-life for the reaction?

$$t_{1/2} = \frac{\ln 2}{k} = \frac{\ln 2}{65 \text{ min}^{-1}} = 0.0106 \text{ min}$$

$$\text{or } 0.64 \text{ sec}$$