

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)

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

1. The rate constant for the formation of hydrogen for the reaction: $H_2(g) + I_2(g) \rightarrow 2HI(g)$ is $2.7 \times 10^{-4} \text{ L}/(\text{mol}\cdot\text{s})$ at 600 K and $3.5 \times 10^{-3} \text{ L}/(\text{mol}\cdot\text{s})$ at 650 K .

a. (5 Pts) Determine the activation energy.

$$\ln \frac{k_1}{k_2} = \frac{E_a}{R} \left(\frac{1}{T_2} - \frac{1}{T_1} \right) \Rightarrow \ln \left(\frac{2.7 \times 10^{-4}}{3.5 \times 10^{-3}} \right) = \frac{E_a}{8.314 \frac{\text{J}}{\text{mol}\cdot\text{K}}} \left(\frac{1}{650 \text{ K}} - \frac{1}{600 \text{ K}} \right)$$

$$E_a = 166000 \frac{\text{J}}{\text{mol}} = 166 \frac{\text{kJ}}{\text{mol}}$$

b. (6 Pts) Determine the rate constant at 750 K

→ Same formula use either k_1 or k_2

$$\ln \left(\frac{k_1}{2.7 \times 10^{-4}} \right) = \frac{166150 \frac{\text{J}}{\text{mol}}}{8.314 \frac{\text{J}}{\text{mol}\cdot\text{K}}} \left(\frac{1}{600 \text{ K}} - \frac{1}{750 \text{ K}} \right) = 6.6615 \dots$$

$$\frac{k_1}{2.7 \times 10^{-4}} = e^{6.6615 \dots}$$

$$k_1 = 2.1 \times 10^{-1} \frac{\text{L}}{\text{mol}\cdot\text{s}}$$

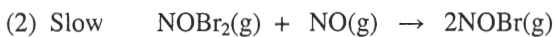
c. (6 Pts) Determine the temperature in K at which the value of the rate constant is $3.2 \times 10^{-2} \text{ L}/(\text{mol}\cdot\text{s})$

$$\ln \frac{2.7 \times 10^{-4}}{3.2 \times 10^{-2}} = \frac{166150}{8.314} \left(\frac{1}{T_2} - \frac{1}{600} \right)$$

$$0.0014277 = \frac{1}{T_2}$$

$$T = 700 \text{ K}$$

2. The following **STEPS** are proposed for a reaction mechanism.



a. (4 Pts) Identify any intermediates and show the overall reaction.

intermediate: $\text{NOBr}_2(g)$



b. (4 Pts) Determine the rate law. Show all work

From slow step: rate = $k_2 [\text{NOBr}_2^\ddagger] [\text{NO}]$

Since NOBr_2 is an intermediate, we must use (1) to solve and substitute for it.

From (1) $\frac{k_1 [\text{NO}] [\text{Br}_2]}{k_{-1}} = \frac{k_{-1} [\text{NOBr}_2^\ddagger]}{k_1}$

therefore: $\text{rate} = \frac{k_2 k_1}{k_{-1}} [\text{NO}]^2 [\text{Br}_2] = k [\text{NO}]^2 [\text{Br}_2]$