Chem 1403
Arrhenius Equation Key

1. **Derive the logarithmic form of Arrhenius Equation from its exponential form**

$$k = Ae^{\frac{-E_a}{RT}}$$

$$lnk = \ln(Ae^{\frac{-E_a}{RT}})$$

$$lnk = \ln A + lne^{\frac{-E_a}{RT}}$$

$$lnk = \ln A + \frac{-E_a}{RT}$$

2. **The rate constant of a reaction at 34.9°C is 1.24 x 10^{-2} Ms^{-1}. Calculate the rate constant of this reaction at 56.2°C with E_a equal to 41.7 kJ/mol .**

$$\frac{k_1}{k_2} = \frac{e^{\frac{-E_a}{RT_1}}}{e^{\frac{-E_a}{RT_2}}}$$

$$\frac{k_1}{k_2} = e^{\frac{-E_a}{RT_1} - \frac{-E_a}{RT_2}}$$

$$\frac{k_1}{k_2} = e^{\frac{E_a}{RT_2} - \frac{E_a}{RT_1}}$$

$$\frac{k_1}{k_2} = e^{\frac{E_a}{R}(\frac{1}{T_2} - \frac{1}{T_1})}$$

$$k_1 = k_2e^{\frac{E_a}{R}(\frac{1}{T_2} - \frac{1}{T_1})}$$

$$k_1 = 1.24 \times 10^{-2}e^{\frac{41.7}{.008314}(\frac{1}{308.05} - \frac{1}{329.35})} = .0355\, Ms^{-1}$$

3. **A particular reaction has a concentration of 9.6 x 10^{-4} $M^{-1}s^{-1}$ at 16.0°C and 1.45 x 10^{-3} $M^{-1}s^{-1}$ at 27.0°C. With these given information, calculate the activation energy of this reaction.**

$$k_1 = k_2e^{\frac{E_a}{R}(\frac{1}{T_2} - \frac{1}{T_1})}$$

$$lnk_1 = \ln(k_2 e^{\frac{E_a}{R}\left(\frac{1}{T_2} - \frac{1}{T_1}\right)})$$

$$lnk_1 = \ln k_2 + lne^{\frac{E_a}{R}\left(\frac{1}{T_2} - \frac{1}{T_1}\right)}$$

$$lnk_1 = \ln k_2 + \frac{E_a}{R}\left(\frac{1}{T_2} - \frac{1}{T_1}\right)$$

$$lnk_1 = \ln k_2 + \frac{E_a}{R}\left(\frac{1}{T_2} - \frac{1}{T_1}\right)$$

$$lnk_1 - \ln k_2 = +\frac{E_a}{R}\left(\frac{1}{T_2} - \frac{1}{T_1}\right) \rightarrow E_a = R\frac{\ln\left(\frac{k_1}{k}\right)}{\left(\frac{1}{T_2} - \frac{1}{T_1}\right)}$$

$$E_a = R\frac{\ln\left(\frac{k_1}{k_2}\right)}{\left(\frac{1}{T_2} - \frac{1}{T_1}\right)} = 8.314\frac{\ln\left(\frac{9.6 \times 10^{-4}}{1.45 \times 10^{-3}}\right)}{\left(\frac{1}{300.15} - \frac{1}{289.15}\right)} = \frac{27050J}{mol} = 27.050kJ/mol$$

4. **At 30.0°C, decomposition of NH_3 has rate constant of 3.18 x 10^{-3} $M^{-2}s^{-1}$. Given activation energy of this reaction is 135.0kJ/mol. What is the rate constant at 28.0°C?**

Same method as question 2

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

$$k_1 = 3.18 \times 10^{-3} e^{\frac{135.0}{.008314}\left(\frac{1}{303.15} - \frac{1}{301.15}\right)} = 2.23 \times 10^{-3} M^{-2}s^{-1}$$

5. **A reaction has rate constant of 19.0 s^{-1} at 27.6°C and 15.0 s^{-1} at 19.9°C. Calculate the activation energy of this compound.**

Same method as question 3

$$lnk_1 - \ln k_2 = +\frac{E_a}{R}\left(\frac{1}{T_2} - \frac{1}{T_1}\right) \rightarrow E_a = R\frac{\ln\left(\frac{k_1}{k_2}\right)}{\left(\frac{1}{T_2} - \frac{1}{T_1}\right)}$$

$$E_a = R\frac{\ln\left(\frac{19}{15}\right)}{\left(\frac{1}{293.05} - \frac{1}{300.75}\right)} = 8.314\frac{\ln\left(\frac{19}{15}\right)}{\left(\frac{1}{293.05} - \frac{1}{300.75}\right)} = \frac{22495J}{mol} = 22.495kJ/mol$$

6. **At 29.0°C, the production of benzaldehyde has rate constant of 8.9 x 10^{-4} Ms^{-1}. Given the activation energy of this reaction is 445.0kJ/mol. What is the temperature when the rate constant is 1.3 x 10^{-3} Ms^{-1}?**

$$lnk_1 = \ln k_2 + \frac{E_a}{R}\left(\frac{1}{T_2} - \frac{1}{T_1}\right)$$

$$lnk_1 = \ln k_2 + \frac{E_a}{R}\left(\frac{1}{T_2} - \frac{1}{T_1}\right)$$

$$lnk_1 - \ln k_2 = \frac{E_a}{R}\left(\frac{1}{T_2} - \frac{1}{T_1}\right)$$

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

$$\ln\left(\frac{8.9 \times 10^{-4}}{1.3 \times 10^{-3}}\right)\frac{.008314}{445.0} = \left(\frac{1}{T_2} - \frac{1}{302.15}\right)$$

$$-7.079 \times 10^{-6} = \left(\frac{1}{T_2} - \frac{1}{302.15}\right)$$

$$\frac{1}{T_2} = (-7.079 \times 10^{-6}) + \frac{1}{303.15}$$

$$T_2 = \frac{1}{\left[(-7.079 \times 10^{-6}) + \frac{1}{303.15}\right]} = 303.80°K = 30.65℃$$