

Seymafurat -Kinetics (9) - Activation Energy (1)

- * ① Rate constant for 1st order decomposition of a reaction: $A \rightarrow B$ are $k = 4.75 \times 10^{-4} /s$ at 293 K, and $k = 1.63 \times 10^{-3} /s$ at 303 K. What is the activation energy of the reaction?

Ans $\ln \frac{k_2}{k_1} = \frac{E_a}{R} \left[\frac{1}{T_1} - \frac{1}{T_2} \right]$ $T = K$

$$\ln \frac{1.63 \times 10^{-3} /s}{4.75 \times 10^{-4} /s} = \frac{E_a}{\frac{8.3145 J/mol \cdot K}{1000}} \left(\frac{1}{293} - \frac{1}{303} \right) K$$

$$1.233 = E_a \times (120.3) \times (1.13 \times 10^{-4})$$

$$\rightarrow E_a = \frac{1.233}{120.3 \times 1.13 \times 10^{-4}} = \boxed{91 \text{ kJ/mol}}$$

- * ② Decomposition of DTBP is 1st order with half life of 320 mins at 135°C and 100 mins at 145°C. Calculate the activation energy for this reaction: $k = ?$ $t_{1/2} = \frac{0.693}{k}$

Ans $\left\{ \begin{array}{l} R_1 = \frac{0.693}{320 \text{ min}} = 2.17 \times 10^{-3} /min \\ R_2 = \frac{0.693}{100 \text{ min}} = 6.93 \times 10^{-3} /min \end{array} \right.$

$T_1 = 135 + 273 = 408 \text{ K}, T_2 = 145 + 273 = 418 \text{ K}$

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

$$\ln \frac{6.93 \times 10^{-3} /min}{2.17 \times 10^{-3} /min} = \frac{E_a}{\frac{8.314 J/mol \cdot K}{1000}} \left[\frac{1}{408} - \frac{1}{418} \right] K$$

$$1.16 = E_a \times 120.3 \text{ mol/K} \times 5.86 \times 10^{-5}$$

$$\boxed{E_a = 165 \text{ kJ/mol}}$$

Sepnafyote -

Kinetics (7) - Activation Energy (1)

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Ans

$$\ln \frac{k_2}{k_1} = \frac{E_a}{R} \left[\frac{1}{T_1} - \frac{1}{T_2} \right] \quad [T = K]$$
$$\ln \frac{1.63 \times 10^{-3} /s}{4.75 \times 10^{-4} /s} = \frac{E_a}{\frac{8.3145 \text{ J/mol} \cdot K}{1000}} \left(\frac{1}{293} - \frac{1}{303} \right) K$$
$$1.233 = E_a \times (120.3) \times (1.13 \times 10^{-4})$$
$$\rightarrow E_a = \frac{1.233}{120.3 \times 1.13 \times 10^{-4}} = \boxed{91 \text{ kJ/mol}}$$

- * ② Decomposition of DTBP is 1st order with half life of 320 mins at 135°C and 100 mins at 145°C. Calculate the activation energy for this reaction: $k = ?$ $t_{1/2} = \frac{0.693}{k}$

Ans

$$\begin{cases} k_1 = \frac{0.693}{320 \text{ mins}} = 2.17 \times 10^{-3} / \text{min} & k_2 = \frac{0.693}{100 \text{ mins}} = 6.93 \times 10^{-3} / \text{min} \\ T_1 = 135 + 273 = 408 \text{ K} & T_2 = 145 + 273 = 418 \text{ K} \end{cases}$$

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

$$\ln \frac{6.93 \times 10^{-3} / \text{min}}{2.17 \times 10^{-3} / \text{min}} = \frac{E_a}{\frac{8.314 \text{ J/mol} \cdot K}{1000}} \left[\frac{1}{408} - \frac{1}{418} \right] K$$
$$1.16 = E_a \times 120.3 \text{ mol/KJ} \times 5.86 \times 10^{-5}$$
$$\boxed{E_a = 165 \text{ kJ/mol}}$$