

Quiz (Arrhenius Equation)

1. Determine graphically the activation energy for the decomposition of phosphine (PH_3) on a quartz surface from the following data:

Rate constant (s^{-1})	1.995×10^{-8}	5.012×10^{-8}	3.162×10^{-7}	1.000×10^{-6}
Temperature (K)	635	690	800	870

(Given: $R = 8.31 \text{ J K}^{-1} \text{ mol}^{-1}$)

Decomposition of phosphine: $2\text{PH}_3(\text{g}) \longrightarrow 2\text{P}(\text{s}) + 3\text{H}_2(\text{g})$

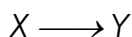
2. The rate constants of a reaction at different temperatures were found by experiments and are shown below:

Temperature (K)	200	300	400	500	600	700	800
Rate constant (s^{-1})	0.236	0.301	0.340	0.366	0.385	0.399	0.409

What is the activation energy of this reaction?

(Given: $R = 8.31 \text{ J K}^{-1} \text{ mol}^{-1}$)

3. Consider the following reaction:



At 300°C , the rate constant for the reaction is $2.41 \times 10^{-10} \text{ s}^{-1}$. At 400°C , the rate constant is $1.16 \times 10^{-6} \text{ s}^{-1}$. Determine the activation energy, E_a of this reaction.

4. Using the following data, determine the activation energy for the decomposition of HI. (Given: $R = 8.31 \text{ J K}^{-1} \text{ mol}^{-1}$)

Temperature (K)	Rate constant ($\text{mol dm}^{-3} \text{ s}^{-1}$)
673	8.38×10^{-4}
773	7.65×10^{-2}

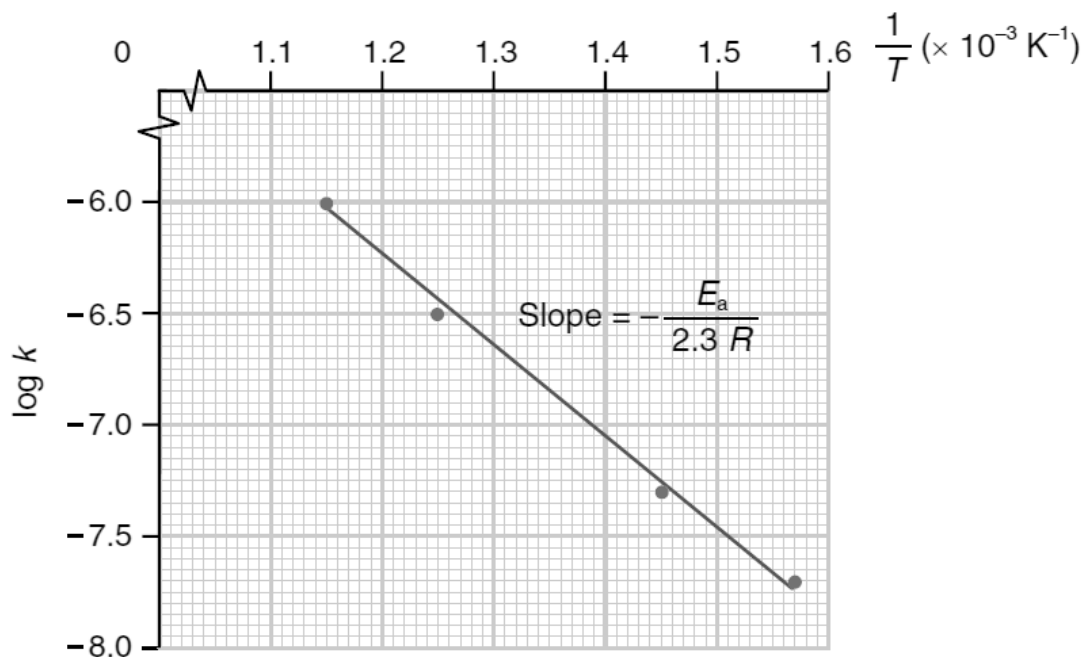
5. Consider a reaction with $E_a = 52 \text{ kJ mol}^{-1}$. Comment on the effect of increasing the reaction temperature from 298 K to 308 K on the rate of reaction.

Suggested Answer

1. Arrhenius equation: $\log k = \text{constant} - E_a / 2.3 RT$

By plotting a graph of $\log k$ against $1/T$, the slope of the line obtained should be equal to $-E_a / 2.3 R$.

$1/T (\times 10^{-3} \text{ K}^{-1})$	1.57	1.45	1.25	1.15
$\log k$	-7.7	-7.3	-6.5	-6.0



By using the points (0.00119, -6.2) and (0.00141, -7.1),

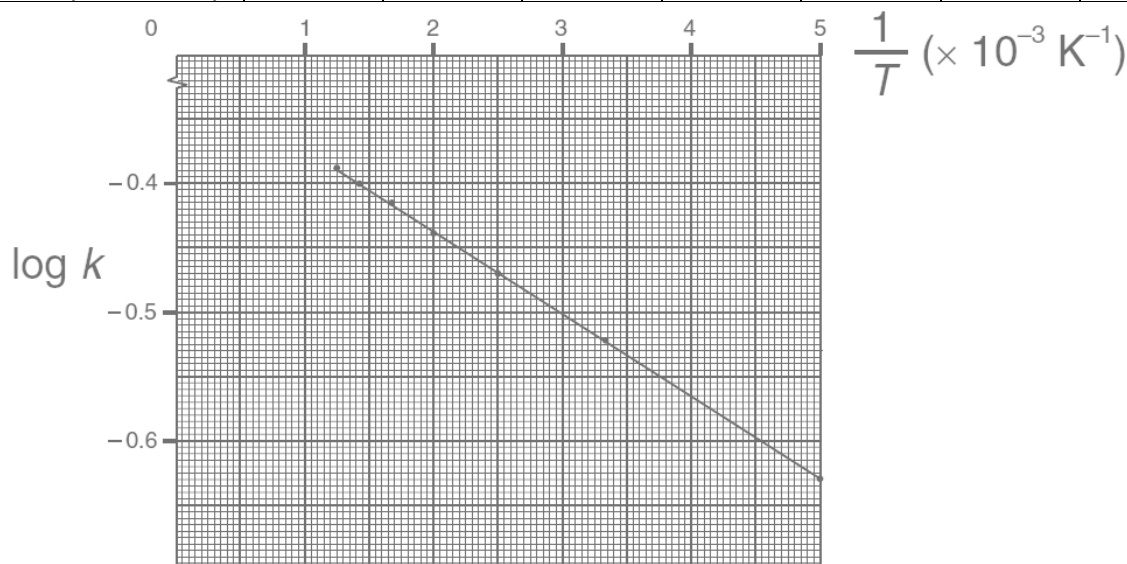
the slope = $[-7.1 - (-6.2)] / (0.00141 - 0.00119) = -4091$

$\therefore -E_a / 2.3 R = -4091$

$$\begin{aligned} E_a &= 4091 \times 2.3 \times 8.31 \\ &= 78\,191 \text{ J mol}^{-1} \\ &= 78.2 \text{ kJ mol}^{-1} \end{aligned}$$

2. Arrhenius equation states that $\log k = \text{constant} - E_a / 2.3 RT$
 By plotting a graph of $\log k$ again $1/T$, the slope of the line obtained should be equal to $- E_a / 2.3 R$

$\log k$	-0.627	-0.521	-0.469	-0.437	-0.415	-0.399	-0.388
$1 / T (\times 10^{-3} \text{ K}^{-1})$	5.00	3.33	2.50	2.00	1.67	1.43	1.25



By using the points (0.00333, -0.521) and (0.005, -0.627),

the slope = $[-0.627 - (-0.521)] / (0.005 - 0.00333) = -63.47$

$$- E_a / 2.3 R = -63.47$$

$$\begin{aligned} E_a &= 63.47 \times 2.3 \times 8.31 \\ &= 1213 \text{ J mol}^{-1} \\ &= 1.213 \text{ kJ mol}^{-1} \end{aligned}$$

3. By substituting values of k_1 , k_2 , T_1 and T_2 into the Arrhenius equation, E_a can be solved.

$$\begin{aligned} k_1 &= 2.41 \times 10^{-10} & T_1 &= 300 + 273 \text{ K} = 573 \text{ K} \\ k_2 &= 1.16 \times 10^{-6} & T_2 &= 400 + 273 \text{ K} = 673 \text{ K} \end{aligned}$$

$$\log k_1 - \log k_2 = E_a / 2.3 R (1 / T_2 - 1 / T_1)$$

$$\log 2.41 \times 10^{-10} - \log 1.16 \times 10^{-6} = E_a / 2.3 \times 8.31 (1/673 - 1/573)$$

$$\text{Solving for } E_a = 271 \text{ kJ mol}^{-1}$$

4. By substituting values of k_1 , k_2 , T_1 and T_2 into the Arrhenius equation, E_a can be solved.

$$\begin{array}{ll} k_1 = 8.38 \times 10^{-4} & T_1 = 673 \text{ K} \\ k_2 = 7.65 \times 10^{-2} & T_2 = 773 \text{ K} \end{array}$$

$$\log k_1 - \log k_2 = E_a / 2.3 R (1 / T_2 - 1 / T_1)$$

$$\log 8.38 \times 10^{-4} - \log 7.65 \times 10^{-2} = E_a / 2.3 \times 8.31 (1/773 - 1/673)$$

$$\text{Solving for } E_a = 195 \text{ kJ mol}^{-1}$$

5. According to the Arrhenius equation,

$$\begin{array}{l} \text{At 298 K, } k_{298} = A \exp (- 52 \times 10^3 / 8.31 \times 298) \\ \quad \quad \quad = A (7.59 \times 10^{-10}) \end{array}$$

$$\begin{array}{l} \text{At 308 K, } k_{308} = A \exp (- 52 \times 10^3 / 8.31 \times 308) \\ \quad \quad \quad = A (15.02 \times 10^{-10}) \end{array}$$

$$k_{308} / k_{298} = A (15.02 \times 10^{-10}) / A (7.59 \times 10^{-10}) = 1.98 \approx 2$$

$$k_{308} = 2 \times k_{298}$$

The reaction rate nearly doubles when the temperature increases by 10 K. Small temperature changes lead to an obvious influence on the rates of reactions.