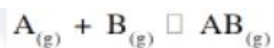


Q. Consider the following reversible reaction, $A(g) + B(g) \rightleftharpoons AB(g)$ The activation energy of the backward reaction exceeds that of the forward reaction by $2RT$ (in J mol^{-1}). If the pre-exponential factor of the forward reaction is 4 times that of the reverse reaction, the absolute value of ΔG^θ (in J mol^{-1}) for the reaction at 300 K is _____. (Given ; $\ln(2) = 0.7$, $RT = 2500 \text{ J mol}^{-1}$ at 300 K and G is the Gibbs energy)

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$$E_{ab} - E_{af} = 2RT \quad \Rightarrow \quad \Delta H = -2RT \quad \text{and} \quad \frac{A_f}{A_b} = 4$$

Sol. 8500

$$K_{eq} = \left(\frac{K_f}{K_b} \right) = \frac{A_f e^{-E_{af}/RT}}{A_b e^{-E_{ab}/RT}} = 4(e^2)$$

$$\Delta G^\circ = -RT \ln K = -2500 \times \ln(4 \times e^2) = -8500 \text{ J/mol}$$

$$\therefore \text{Absolute value of } \Delta G^\circ = 8500 \text{ J/mol}$$