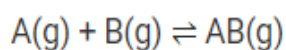


Question ● Consider the following reversible reaction, $A(g) + B(g) \rightleftharpoons AB(g)$. The activation energy of the backward reaction exceeds that of the forward reaction by $2RT$ (in $J mol^{-1}$). If the pre-exponential factor of the forward reaction is 4 times that of the reverse reaction,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

Given; $\ln(2) = 0.7$, $RT = 2500 J mol^{-1}$ at 300 K and **G** is the Gibbs energy)

Solution: (8500)



$$(E_a)_b - (E_a)_f = 2RT$$

$$A_f / A_b = 4$$

$$\Delta G^\circ = -RT \ln K_{eq}$$

$$K_f = A_f e^{-(E_a)_f / RT}$$

$$K_b = A_b e^{-(E_a)_b / RT}$$

$$K_{eq} = K_f / K_b = A_f / A_b \times e^{-(E_a)_f / RT} \times e^{+(E_a)_b / RT} = 4 \times e^{(E_a)_b - (E_a)_f / RT}$$

$$K_{eq} = 4 \times e^2$$

$$\Delta G^\circ = -RT \times \ln(4 \times e^2)$$

$$\Delta G^\circ = -RT(\ln 4 + 2 \ln e)$$

$$\Delta G^\circ = -RT(2 \times 0.7 + 2)$$

$$\Delta G^\circ = -RT(1.40 + 2)$$

$$\Delta G^\circ = -RT(3.40)$$

$$\Delta G^\circ = -2500 \times 3.40$$

$$\Delta G^\circ = -8500J.$$