Q. Consider the following reversible reaction, $A(g) + B(g) \rightleftharpoons AB(g)$ The activition energy of the backward reaction exceeds that of the forward reaction by 2RT $\left(\text{in Jmol}^{-1}\right)$. 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) [JEE Advance 2018]

Sol. 8500

$$A_{(g)} + B_{(g)} \rightleftharpoons AB_{(g)}$$

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

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

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

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