**Question● Consider the following reversible reaction, $A(g) + B(g) \rightleftarrows AB(g)$. 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 ΔGθ (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)

$A(g) + B(g) \rightleftharpoons AB(g)$

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

$A_f / A_b = 4$

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

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

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

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

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

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

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

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

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

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

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

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