Answer on Question #73551, Chemistry / General Chemistry :

The enthalpy change (ΔH) and entropy change(ΔS) for the reaction of 30.54 kj mol^{-1} and 0.06 kj mol^{-1} respectively. Calculate the temperature at equilibrium. Also predict the spontaneity below the temperature at which Gibb's free energy is zero. Justify your answer with a valid reason.

Solution.

$$\Delta H = 30.54 kj / mol$$
$$\Delta S = 0.06 kj / mol \cdot K$$
$$\Delta G = 0$$

$$T - ?$$

Gibb's free energy is:

$$\Delta G = \Delta H - T \cdot \Delta S$$

When $\Delta G = 0$, and:

$$0 = \Delta H - T \cdot \Delta S$$
$$T \cdot \Delta S = \Delta H$$
$$T = \frac{\Delta H}{\Delta S} = \frac{30.54 kj / mol}{0.06 kj / mol \cdot K}$$
$$T = 509 K$$

If $T > 509K$:

$$\Delta H - T \cdot \Delta S < 0$$
$$\Delta G < 0$$

Reaction with a positive Gibbs free energy will not proceed spontaneously.

When $\Delta G < 0$, the process is exergonic and will proceed spontaneously in the forward direction to form more products.

Answer: $T = 509K$.