

Answer on the question #62947, Chemistry / Physical Chemistry

Question:

The normal boiling point of ethanol is 352.6K. The enthalpy of formation in the gas phase is -235.10 kJ mol⁻¹, while at liquid phase is -277.69kJ mol⁻¹. Given that $\Delta C_p = 0$ for the vapourisation process, compute the equilibrium vapour pressure of ethanol at 298K

Solution:

To solve this, we use Clausius-Clapeyron equation. This equation relates the vapour pressure, enthalpy of vaporization and temperature:

$$\ln(P) = \text{const} - \frac{\Delta H_{vap}}{RT}$$

Then, vapour pressure change for different temperatures is:

$$\ln\left(\frac{P_2}{P_1}\right) = \frac{\Delta H_{vap}}{R} \left(\frac{1}{T_1} - \frac{1}{T_2}\right)$$

Then, $T_1 = 352.6K$ and $T_2 = 298 K$.

In its turn, $P_1 = 101325 Pa$, and P_2 is the value we need to compute (vapour pressure of ethanol at 298K).

R is the gas constant, $R = 8.314 \frac{J}{K \cdot mol}$.

If we assume, that the change in heat capacity for this temperature interval is zero ($\Delta C_p = 0$), vaporization enthalpy ΔH_{vap} can be calculated as the difference between the enthalpies of formation of ethanol in gas phase and in liquid phase:

$$\Delta H_{vap} = \Delta H_{f,gas} - \Delta H_{f,liquid}$$

$$\Delta H_{vap} = -235.10 \frac{kJ}{mol} - \left(-277.69 \frac{kJ}{mol}\right) = 42.59 \frac{kJ}{mol} = 42.59 \cdot 10^3 \frac{J}{mol}$$

Now, we should rearrange the equation to get $\ln(P_2)$, knowing that $\ln\left(\frac{P_2}{P_1}\right) = \ln(P_2) - \ln(P_1)$:

$$\begin{aligned} \ln(P_2) &= \ln(P_1) + \frac{\Delta H_{vap}}{R} \left(\frac{1}{T_1} - \frac{1}{T_2}\right) \\ \ln(P_2) &= \ln(101325 Pa) + \frac{42.59 \cdot 10^3 \frac{J}{mol}}{8.314 \frac{J}{K \cdot mol}} \left(\frac{1}{352.6K} - \frac{1}{298K}\right) \\ P_2 &= 7074.1 Pa \end{aligned}$$

Answer: 7074.1 Pa

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