

CH 223 - Worksheet 6

1. The normal boiling point of ethanol (C_2H_5OH) is $78.3^\circ C$, and its molar enthalpy of vaporization is 38.56 kJ/mol . What is the change in entropy in the system when 68.3 g of C_2H_5OH (g) at 1 atm condenses to a liquid at the normal boiling point?

$$68.3 \text{ g } C_2H_5OH \left(\frac{1 \text{ mol}}{46 \text{ g}} \right) = 1.485 \text{ mol } C_2H_5OH$$

$$\Delta S = \frac{-\Delta H}{T} = \frac{(1.485 \text{ mol } C_2H_5OH)(38.56 \frac{\text{kJ}}{\text{mol}})}{351 \text{ K}}$$

$$= 0.163 \frac{\text{kJ}}{\text{K}} = 163 \frac{\text{J}}{\text{K}}$$

2. The normal freezing point of 1-propanol (C_3H_8O) is $-127^\circ C$. (a) Is the freezing an endothermic or exothermic process? (b) In what temperature range is the freezing of 1-propanol a spontaneous process? (c) In what temperature range is it a nonspontaneous process? (d) Is there any temperature at which liquid and solid 1-propanol are in equilibrium? Explain.

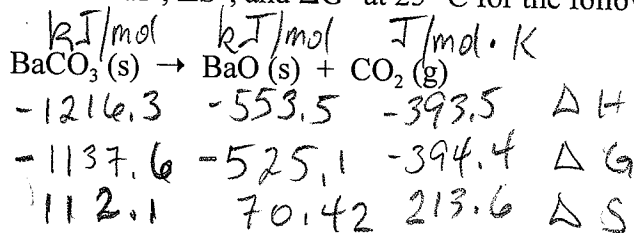
a) exothermic

d) equilibrium occurs at $-127^\circ C$ and 1 atm .

b) $< -127^\circ C$

c) $> -127^\circ C$

3. Using data from Appendix C, calculate ΔH° , ΔS° , and ΔG° at $25^\circ C$ for the following reaction:



Show that $\Delta G^\circ = \Delta H^\circ - T\Delta S^\circ$

$$\Delta H = \sum n \Delta H^\circ(\text{products}) - \sum m \Delta H^\circ(\text{reactants})$$

$$\Delta H = \left[-553.5 \frac{\text{kJ}}{\text{mol}} + -393.5 \frac{\text{kJ}}{\text{mol}} \right] - \left[-1216.3 \frac{\text{kJ}}{\text{mol}} \right] = 269.3 \text{ kJ}$$

$$\Delta S = \sum n S^\circ(\text{products}) - \sum m S^\circ(\text{reactants})$$

$$\Delta S = \left[70.42 \frac{\text{J}}{\text{mol}\cdot\text{K}} + 213.6 \frac{\text{J}}{\text{mol}\cdot\text{K}} \right] - \left[112.1 \frac{\text{J}}{\text{mol}\cdot\text{K}} \right] = 171.32 \text{ J}$$

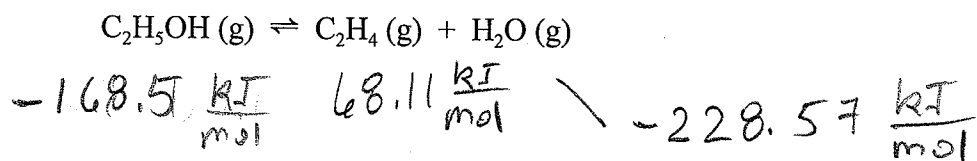
$$\Delta G = \Delta H - T\Delta S = 269.3 \text{ kJ} - 298 \text{ K} \left(0.17132 \frac{\text{kJ}}{\text{K}} \right) = 218.2 \text{ kJ}$$

or $\Delta G = \sum n \Delta G^\circ(\text{products}) - \sum m \Delta G^\circ(\text{reactants})$

$$\Delta G = \left[-525.1 \frac{\text{kJ}}{\text{mol}} + -394.4 \frac{\text{kJ}}{\text{mol}} \right] - \left[-1137.6 \frac{\text{kJ}}{\text{mol}} \right] = 218.2 \text{ kJ}$$

They are the same.

4. Use data from Appendix C to calculate the equilibrium constant, K , at 298 K for the following reaction:



$$\Delta G = \Delta G^\circ + RT \ln Q$$

$$\begin{aligned} \Delta G^\circ &= \sum n \Delta G^\circ(\text{products}) - \sum m \Delta G^\circ(\text{reactants}) \\ &= \left[68.11 \frac{\text{kJ}}{\text{mol}} + -228.57 \frac{\text{kJ}}{\text{mol}} \right] - \left[-168.5 \frac{\text{kJ}}{\text{mol}} \right] \\ &= 8.04 \frac{\text{kJ}}{\text{mol}} \end{aligned}$$

at equilibrium

$$\Delta G = 0 = \Delta G^\circ + RT \ln K \quad \text{at eq. } K = Q$$

$$\therefore \Delta G^\circ = -RT \ln K \quad R = 8.314 \frac{\text{J}}{\text{mol} \cdot \text{K}} \quad K =$$

$$-\frac{\Delta G^\circ}{RT} = \ln K$$

$$K = e^{-\Delta G^\circ / RT} = e^{-8.040 \frac{\text{J}}{\text{mol}} / 8.314 \frac{\text{J}}{\text{mol} \cdot \text{K}} \cdot 298 \text{ K}}$$

$$= 0.0390$$