1. The normal boiling point of ethanol (C₂H₅OH) is 78.3 °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 C₂H₅OH (g) at 1 atm condenses to a liquid at the normal boiling point?

\[ \Delta S = -\frac{\Delta H}{T} = \frac{1.485 \text{ mol C}_2\text{H}_5\text{OH}}{68.3 \text{ g}} \times \frac{38.56 \text{ kJ/mol}}{351 \text{ K}} \]

\[ = 0.163 \frac{\text{kJ}}{\text{K}} = 163 \frac{\text{J}}{\text{K}} \]

2. The normal freezing point of 1-propanol (C₃H₇O) is −127 °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  
(b) < −127 °C  
(c) > −127 °C  
(d) Equilibrium occurs at −127 °C and 1 atm.

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

\[ \text{BaCO}_3(s) \rightarrow \text{BaO}(s) + \text{CO}_2(g) \]

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

\[ \Delta H = [-553.5 \text{ kJ/mol}] - [-1216.3 \text{ kJ/mol}] = 26.93 \text{ kJ} \]

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

\[ \Delta S = [70.42 \text{ J/mol.K}] + [213.6 \text{ J/mol.K}] - [112.1 \text{ J/mol.K}] = 171.32 \text{ J/K} \]

\[ \Delta G = \Delta H - T \Delta S = 26.93 \text{ kJ} - 298 \text{ K} (0.17132 \text{ kJ/K}) = 218.2 \text{ kJ} \]

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

\[ \Delta G = [-525.1 \text{ kJ/mol}] - [-394.4 \text{ kJ/mol}] - [-1137.6 \text{ kJ/mol}] = 218.2 \text{ kJ} \]

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

$$C_2H_5OH (g) \rightleftharpoons C_2H_4 (g) + H_2O (g)$$

$$\Delta G = -168.5 \frac{kJ}{mol} - 68.11 \frac{kJ}{mol} \leq -228.57 \frac{kJ}{mol}$$

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

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

$$= [68.11 \frac{kJ}{mol} + (-228.57 \frac{kJ}{mol})] - (-168.5 \frac{kJ}{mol})$$

$$= 8.04 \frac{kJ}{mol}$$

at equilibrium

$$\Delta G = 0 = \Delta G^\circ + RT \ln K$$

$$\Rightarrow \Delta G^\circ = -RT \ln K$$

$$R = 8.314 \frac{J}{mol \cdot K}$$

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

$$= 0.0390$$