

$$\Delta S [\text{H}_2\text{O}_2 (g)] = 232.9 \text{ J K}^{-1} \text{ mol}^{-1} \quad \Delta H_f [\text{H}_2\text{O}_2 (g)] = -136.1 \text{ kJ mol}^{-1}$$

$$\Delta S_{\text{universe}} = \Delta S_{\text{sys}} + \Delta S_{\text{surr}} = 5.3 \text{ J K}^{-1} \text{ mol}^{-1} \text{ spontaneous}$$

7. The reaction $\text{NH}_3(g) + \text{HCl}(g) \rightarrow \text{NH}_4\text{Cl}(s)$ is spontaneous at room temperature. Determine the temperature at which it is no longer spontaneous. Given information

$$S^\circ [\text{NH}_3(g)] = 193.0 \text{ J K}^{-1} \text{ mol}^{-1}$$

$$S^\circ [\text{HCl}(g)] = 187.0 \text{ J K}^{-1} \text{ mol}^{-1}$$

$$S^\circ [\text{NH}_4\text{Cl}(s)] = 94.56 \text{ J K}^{-1} \text{ mol}^{-1}$$

$$\Delta H_f^\circ [\text{NH}_3(g)] = -46.3 \text{ kJ mol}^{-1}$$

$$\Delta H_f^\circ [\text{HCl}(g)] = -92.3 \text{ kJ mol}^{-1}$$

$$\Delta H_f^\circ [\text{NH}_4\text{Cl}(s)] = -315.39 \text{ kJ mol}^{-1}$$

$$T < 619 \text{ K}$$

8. Calculate the change in the entropies of the system and the surroundings, and the total change in entropy, when the volume of a sample of argon gas of mass 21 g at 298 K and 1.50 bar increases from 1.20 dm³ to 4.60 dm³ in (a) an isothermal reversible expansion, (b) an isothermal irreversible expansion against $p_{\text{ext}} = 0$, and (c) an adiabatic reversible expansion

$$\text{a) } \Delta S_{\text{system}} = +3.0 \text{ J/K}, \Delta S_{\text{surrounding}} = -3.0 \text{ J/K}, \Delta S_{\text{total}} = 0 \text{ J/K}$$

$$\text{b) } \Delta S_{\text{system}} = +3.0 \text{ J/K}, \Delta S_{\text{surrounding}} = -3.0 \text{ J/K}, \Delta S_{\text{total}} = 0 \text{ J/K}$$

$$\text{c) } \Delta S_{\text{system}} = +3.0 \text{ J/K}, \Delta S_{\text{surrounding}} = +3.0 \text{ J/K}, \Delta S_{\text{total}} = +6.0 \text{ J/K}$$

9. The standard molar enthalpy of formation of $\text{H}_2\text{O}(g)$ is $-241.8 \text{ kJ mol}^{-1}$ and its standard molar entropy is $188.8 \text{ J K}^{-1} \text{ mol}^{-1}$. Calculate the standard molar Gibbs energy of formation of $\text{H}_2\text{O}(g)$ at 298 K.

$$\Delta G_f^\circ = \Delta H_f^\circ - T \Delta S_f^\circ = -241.8 \text{ kJ mol}^{-1} - (298 \text{ K})(188.8 \text{ J K}^{-1} \text{ mol}^{-1}) = -297.0 \text{ kJ mol}^{-1}$$

$$\Delta G_f^\circ = -297.0 \text{ kJ mol}^{-1}$$

$$\Delta G_f^\circ = -297.0 \text{ kJ mol}^{-1}$$

10. The standard molar Gibbs energy of formation of water vapor at room temperature, 298 K, is $-228.57 \text{ kJ mol}^{-1}$. The standard molar enthalpy of formation of water vapor at 298 K is $-241.82 \text{ kJ mol}^{-1}$. Calculate the standard molar entropy of formation of water vapor at 298 K.

$$\Delta S_f^\circ = 131.25 \text{ J K}^{-1} \text{ mol}^{-1}$$

11. Suppose that the ΔG_f° for the reaction of iron(II) sulfide is more negative at 1000 K than at 298 K. The standard molar enthalpy of formation of $\text{FeS}(s)$ is $-100.0 \text{ kJ mol}^{-1}$.

Exercises For Thermodynamics Answers

Karin Nielsen-Saines



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