

12.3: Phase Changes - the Fusion of Ice

Let us consider processes in which transfer of heat from the surroundings melts one mole of ice. We suppose that the ice is initially at 0 °C and one bar. At these conditions, the enthalpy change for melting a mole of ice is 6010 J. If the ice melts reversibly at these conditions, the temperature of the surroundings is also 0 °C. As it melts, the ice takes up 6010 J of heat, which is given up by the surroundings. For this process, we have $q_P^{rev} = \Delta_{fus}H = 6010$ J. The temperature is constant, and the entropy change for the system is

$$\Delta S = \frac{q_P^{rev}}{T} = \frac{\Delta_{fus}H}{T}$$

Since $\hat{q} = -q_P^{rev}$, we have

$$\Delta\hat{S} = \frac{-q_P^{rev}}{\hat{T}} = \frac{-\Delta_{fus}H}{\hat{T}}$$

so that

$$\Delta S = 6010 \text{ J} / 273.15 \text{ K} = 22.00 \text{ J K}^{-1}$$

and

$$\Delta\hat{S} = -6010 \text{ J} / 273.15 \text{ K} = -22.00 \text{ J K}^{-1}$$

As required for a reversible process, we have $\Delta S + \Delta\hat{S} = 0$. The Gibbs free energy change is

$$(\Delta G)_{PT} = (\Delta H)_{PT} - T(\Delta S)_{PT} = q_P^{rev} - T(q_P^{rev}/T) = 0$$

which is also as required for a reversible process.

Now let us consider a spontaneous process, in which the ice melts while in thermal contact with surroundings at 10 °C. To reach equilibrium, the system must reach the temperature of the surroundings, which we assume to be constant. In this process, the ice melts and the melt water warms to 10 °C. To find the entropy change, we must find a reversible process that effects the same change. A two-step process effects this conveniently. The first step is the one we have just considered: Surroundings at 0 °C transfer 6010 J of heat to the system, reversibly melting the ice to water at 0 °C. We have $q_1 = 6010$ J and

$$\Delta S_1 = 22.00 \text{ J K}^{-1}$$

In the second reversible step, surroundings that are always at the same temperature as the system transfer heat to the system as the temperature increases from 273.15 K to 283.15 K. The heat capacity of liquid water is 75.3 J mol^{-1} . For this step,

$$q_2 = \int_{273.15}^{283.15} C_P dT = (75.3 \text{ J K}^{-1})(10 \text{ K}) = 753 \text{ J}$$

and

$$\Delta S_2 = \int_{273.15}^{283.15} \frac{C_P}{T} dT = (75.3 \text{ J K}^{-1}) \ln \frac{283.15}{273.15} = 2.71 \text{ J K}^{-1}$$

For these reversible changes in the system, we have $\Delta S = \Delta S_1 + \Delta S_2 = 24.71 \text{ J K}^{-1}$. This is also the value of ΔS for the spontaneous process. The heat taken up by the system in the two-step reversible process is $q = q_1 + q_2 = 6763$ J. This heat is surrendered by the surroundings, and we could effect identically the same change in the surroundings by exchanging this quantity of heat reversibly. For the spontaneous process, therefore, we have $\hat{q} = -6763$ J and

$$\Delta\hat{S} = -6763 \text{ J} / 283.15 \text{ K} = -23.88 \text{ J K}^{-1}$$

For the universe, we have

$$\Delta S_{universe} = \Delta S + \Delta\hat{S} = +0.83 \text{ J K}^{-1}$$

which is greater than zero, as required for a spontaneous process.

Because this reversible two-step process does not occur at a constant temperature, its Gibbs free energy change is not zero. However, we can use the Gibbs-Helmholtz equation to estimate the Gibbs free energy change for the related process in which ice at 10 °C and 1 bar (a hypothetical substance) melts to form liquid water at the same temperature and pressure. For this process, we estimate

$$\Delta G = -220 \text{ J mol}^{-1} \text{ K}^{-1}$$

(See problem 10-21.) Since we have $\Delta G < 0$ for the process, our change criterion asserts that, in agreement with our experience, superheated ice melts spontaneously at 10 C.

This page titled [12.3: Phase Changes - the Fusion of Ice](#) is shared under a [CC BY-SA 4.0](#) license and was authored, remixed, and/or curated by [Paul Ellgen](#) via [source content](#) that was edited to the style and standards of the LibreTexts platform.