Freezing Point Depression, Molar Mass, Percentage Composition

Problem 13-70

Nicotinamide is a water-soluble vitamin important in metabolism. A deficiency in this vitamin results in the debilitating condition known as pellagra. Nicotinamide is 59.0% C, 5.0% H, 22.9% N, 13.1% O, by mass. Addition of 3.88 g of nicotinamide to 30.0 mL nitrobenzene, $C_6H_5NO_2$ (d = 1.204 g/mL), lowers the freezing point from 5.7 to -1.4 °C. What is the molecular formula of this compound?

Solution:

This problem combines the determination of the empirical and molecular formula with the colligative property of freezing point depression.

Start with calculating the molar mass of Nicotinamide:

$$K_f^{\text{nitrobenzene}} = 8.1\,{}^{\circ}\text{C}/m \text{ (Table 13.2)}$$

$$T_{\text{fp}}^{\text{o}} = 5.7^{\circ}\text{C} \qquad T_{\text{fp}}^{\text{soln}} = -1.4^{\circ}\text{C} \qquad \Delta T = -7.1^{\circ}\text{C}$$

$$\Delta T = -iK_f c \qquad c = \frac{\Delta T}{-iK_f} = -\frac{7.1^{\circ}\text{C}}{(1)(8.1\,{}^{\circ}\text{C}/m)} = 0.8765\ m$$

The molar quantity of nicotinamide can be calculated from the molal concentration and the mass of nitrobenzene (solvent) used to make the solution.

$$m_{\text{solvent}} = 30.0\text{ mL} \times 1.204\,\tfrac{\text{g}}{\text{mL}} = 36.12\text{ g} = 0.03612\text{ kg}$$

$$n_{\text{nicotinamide}} = 0.8765\ m \times 0.03612\text{ kg} = 0.03166\text{ mol nicotinamide}$$

This quantity is represented by 3.88 g of nicotinamide so the molar mass is

$$M = \frac{3.88\text{ g}}{0.03166\text{ mol}} = 122.5\,\tfrac{\text{g}}{\text{mol}}$$

Now work on the empirical formula from the microanalysis:

Assume 100 g of nicotinamide

$$n_{\text{C}} = {}^{59.0\text{ g C}}\!\big/_{12.01\,\frac{\text{g}}{\text{mol}}} = 4.913\text{ mol C} \qquad 6.0\text{ mol C}$$

$$n_{\text{H}} = {}^{5.0\text{ g H}}\!\big/_{1.008\,\frac{\text{g}}{\text{mol}}} = 4.960\text{ mol H} \qquad \text{normalize to 1 O} \qquad 6.06\text{ mol H}$$

$$n_{\text{N}} = {}^{22.9\text{ g N}}\!\big/_{14.007\,\frac{\text{g}}{\text{mol}}} = 1.635\text{ mol N} \qquad 2.00\text{ mol N}$$

$$n_{\text{O}} = {}^{13.1\text{ g O}}\!\big/_{16.00\,\frac{\text{g}}{\text{mol}}} = 0.8188\text{ mol O} \qquad 1.0\text{ mol O}$$

Empirical Formula = $C_6H_6N_2O$ (122.1 g/mol)

which is the same as the molecular formula