## Calculated Values for Electrolyte Solutions

Remember that colligative properties depend on the total concentration of solute particles regardless of their identity. The changes in colligative properties caused by electrolytes will be proportional to the total molality of all dissolved particles, not to formula units. For the same molal concentrations of sucrose and sodium chloride, you would expect the effect on colligative properties to be twice as large for sodium chloride as for sucrose. What about barium nitrate, $Ba(NO_3)_2$? Each mole of barium nitrate yields 3 mol of ions in solution.

$$Ba(NO_3)_2(s) \longrightarrow Ba^{2+}(aq) + 2NO_3^-(aq)$$

You would expect a $Ba(NO_3)_2$ solution of a given molality to lower the freezing point of its solvent three times as much as a nonelectrolytic solution of the same molality.

**SAMPLE PROBLEM F** *For more help, go to the* ***Math Tutor*** *at the end of this chapter.*

**What is the expected change in the freezing point of water in a solution of 62.5 g of barium nitrate, $Ba(NO_3)_2$, in 1.00 kg of water?**

**SOLUTION**

**1** ***ANALYZE***

**Given:** solute mass and formula = 62.5 g $Ba(NO_3)_2$
solvent mass and identity = 1.00 kg water
$\Delta t_f = K_f m$

**Unknown:** expected freezing-point depression

**2** ***PLAN***

The molality can be calculated by converting the solute mass to moles and then dividing by the number of kilograms of solvent. That molality is in terms of formula units of $Ba(NO_3)_2$ and must be converted to molality in terms of dissociated ions in solution. It must be multiplied by the number of moles of ions produced per mole of formula unit. This adjusted molality can then be used to calculate the freezing-point depression.

$$\frac{\text{mass of solute (g)}}{\text{mass of solvent (kg)}} \times \frac{\text{1 mol solute}}{\text{molar mass of solute (g)}} = \text{molality of solution}\left(\frac{\text{mol}}{\text{kg}}\right)$$

$$\text{molality of solution}\left(\frac{\text{mol}}{\text{kg}}\right) \times \text{molality conversion}\left(\frac{\text{mol ions}}{\text{mol}}\right) \times K_f\left(\frac{{}^\circ\text{C} \bullet \text{kg } H_2O}{\text{mol ions}}\right)$$
$$= \text{expected freezing-point depression } ({}^\circ\text{C})$$

This problem is similar to Sample Problem E, except that the solute is ionic rather than a nonionizing molecular solute. The number of particles in solution will therefore equal the number of ions of the solute.

**3** ***COMPUTE***

$$\frac{62.5\ \cancel{\text{g } Ba(NO_3)_2}}{1.00 \text{ kg } H_2O} \times \frac{\text{mol } Ba(NO_3)_2}{261.35\ \cancel{\text{g } Ba(NO_3)_2}} = \frac{0.239 \text{ mol } Ba(NO_3)_2}{\text{kg } H_2O}$$