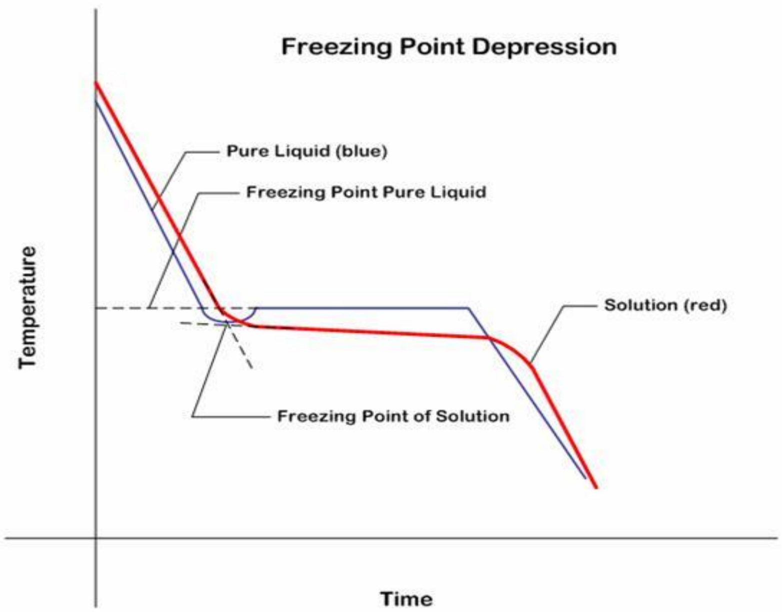
Freezing Point of Solutions

Experiment
12

A solution freezes at a lower temperature than does the pure solvent. This phenomenon is called *freezing point depression*. The freezing point depression of a solution is a colligative property of the solution which is dependent upon the number of dissolved particles in the solution. The higher the solute concentration, the greater the freezing point depression of the solution. The freezing point plot of a pure solvent and a solution are shown below:



The freezing point of the pure solvent is at constant temperature but the freezing point of the solution slowly decreases. The decrease caused by the increase in solute concentration as the solvent freezes. The dissolved solutes can be non-electrolytes or electrolytes. *Non-electrolytes* are molecules that remain intact when they dissolve in water. *Electrolytes* are solutes that dissociate into ions when dissolved in solution to give an electrically conducting solution. The equation describing the change in freezing point from pure solvent to solution is:

$$\Delta T_f = K_f\, m \quad \text{(non-electrolytes)}$$

$$\Delta T_f = i\, K_f\, m \quad \text{(electrolytes)}$$

K_f is the molal freezing point depression constant of the solvent (1.86 °C/m for water).

m = molality = moles of solute per kilogram of solvent.

i = the number of dissolved particles (Van't Hoff Factor).

In this experiment, the freezing points of aqueous solutions of methanol, a non-electrolyte, and sodium chloride, an electrolyte, will be measured and the molality of each solution will be calculated.

Equipment and Reagents

8 inch test tube	rubber stopper with hole	methanol in water
Stirring apparatus	distilled water	sodium chloride in water
Thermometer	rock salt	

Procedure

1. Place 10 mL of distilled water in an 8 inch test tube. Insert the stirring apparatus as shown in the Figure below. Make sure that the thermometer bulb is immersed in the liquid.

2. Place the entire apparatus in a salt ice-bath. Allow the water to cool with continuous stirring until there is an ice-water slush in the tube.

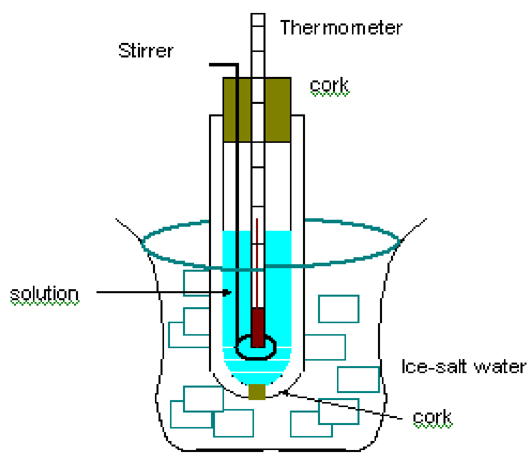


Figure 1 – Freezing point apparatus in ice-salt water bath.

3. Remove the test tube from the bath, stir vigorously and note the constant temperature during the time that ice and water ate both present. This is the freezing point of pure water.

4. In a similar manner, find the freezing point of each of the following solutions by replacing the water in the test tube with 10 mL of the solution.

a) Methanol, CH_3OH, in water.

b) Sodium Chloride, NaCl, in water.

(Each solution contains 1.00 g of solute dissolved in 10.0 g of water.)

5. Remove the test tube from the bath, stir vigorously and record the temperature at which the last ice crystal melts. This is the freezing point of the solution. Repeat steps 4 and 5 with the other solution.

6. The solutions may be washed down the drain when done.

Calculations

Freezing point depression:

$$\Delta T_f = T_f^\circ - T_f$$

T_f° = freezing point of pure solvent

T_f = freezing point of solution

Non-Electrolyte Solution

1. Calculate the theoretical molality of the methanol solution from the concentration reported on the bottle (1.00 g methanol per 10.0 g of water).

2. Calculate the experimental molality of the methanol solution from the freezing point of the solution.

$$\Delta T_f = K_f\, m$$

3. Calculate the percent difference between the experimental and the theoretical molality.

Electrolyte Solution

1. Calculate the theoretical molality of the sodium chloride solution from the concentration reported on the bottle (1.00 g sodium chloride per 10.0 g of water).

2. Calculate the experimental molality of the sodium chloride solution from the freezing point of the solution.

$$\Delta T_f = i\, K_f\, m$$

3. Calculate the percent difference between the experimental and the theoretical molality.