Purpose

In this experiment, we will determine the concentration of a solution from freezing point depression data.

Introduction

Physical chemistry is the study of the relationships between physical properties and chemical constitution. In this experiment, we will be concerned with the freezing points and boiling points of solutions. Neither temperature is itself a colligative property. The relevant colligative properties—ones whose values depend only on the number of solute particles present, not on their chemical identity—are the lowering of the freezing point and the raising of the boiling point of the solutions as compared to that of the pure solvent.

In both cases, the difference is proportional to the molality, according to the equations:

$$\Delta T_\mathrm{f} = iK_\mathrm{f} \cdot m \qquad \text{(Equation 3.1)}$$

$$\Delta T_\mathrm{b} = iK_\mathrm{b} \cdot m \qquad \text{(Equation 3.2)}$$

...where K_f is the molal-freezing point depression constant for the pure solvent, K_b is the molal-boiling point elevation constant for the pure solvent, i is the van 't Hoff factor related to the number of particles created by each solute molecule, and m is the molality of the solution. The freezing point lowering is defined such that it is positive, and is calculated as the freezing point of the pure liquid minus that of the solution. Similarly, the boiling point elevation is defined as positive, and is calculated as the boiling point of the solution minus the boiling point of the pure liquid. Molality is the number of moles of solute per kilogram of solvent:

$$m = \frac{\text{mol solute}}{\text{1 kg solvent}} \qquad \text{(Equation 3.3)}$$

The van 't Hoff factor can be demonstrated by measuring the boiling point elevation of an aqueous salt solution. The relationship between molality and percent composition can be demonstrated by measuring the freezing point depression of alcohol to estimate the percent water in a wet *t*-butyl alcohol sample.

Before Coming to Lab

1. Show the dissolution—how a compound breaks into ions—of sodium nitrate in water.

2. If $\Delta T_f = 3.5\,°C$ for a saltwater solution, what is the freezing point of the solution?

3. Calculate the molality of a solution of 5.50 grams of sodium nitrate in 30.2 grams of water. If you had exactly 1 kg of water, how many grams of sodium nitrate would give the same molality?

4. Calculate the percent water in each of the above solutions. As a reminder, you can calculate the percent water from the formula:

$$\%H_2O = \frac{\text{mass } H_2O}{\text{mass solution}} \times 100\% \qquad \text{(Equation 3.4)}$$

Procedure

Part 1. Percent Water in *t*-Butyl Alcohol by Freezing Point Depression

Calibrating the Thermometer at the Ice Point

1. Fill a 150 mL beaker two-thirds full of crushed ice.

2. Add the minimum amount of distilled water to just cover the ice.

3. Clamp your thermometer with its bulb in the slush and stir with a stirring rod until the temperature becomes constant. Record this temperature to the nearest 0.1 °C.

Measuring the Freezing Point of Wet *t*-Butyl Alcohol

4. Create a warm water bath by placing some water from the hot water tap in a 250 mL beaker.

5. Use your ice-slush as a cold bath. Place 2–3 mL of the wet *t*-butyl alcohol in a medium test tube. Place the thermometer in the test tube, and clamp the test tube in the cold bath.

6. Stir gently with the thermometer, recording the temperature every 15 seconds. Mark the temperature at which you see crystals form. This is the temperature when the liquid suddenly turns cloudy. Continue to record temperatures for 2 more minutes.

7. Re-melt the wet *t*-butyl alcohol solution in the warm water bath. Stir until the liquid is clear. Repeat your determination of the freezing point two more times, and use the average of the three values as the freezing point. Correct the average using the calibration data obtained earlier.

Analysis

The melting point of pure *t*-butyl alcohol is 25.82 °C, and its freezing point depression constant is 8.37 °C per molal. Use your corrected average temperature to find the molality of the water in the alcohol. Find the percent water in the alcohol.

Discussion

1. How did you calibrate your thermometers? Why is it important to calibrate your boiling and freezing points in these experiments?

2. If this same approach were used to determine the percent by mass of a solution with a different solvent, that had a smaller K_f value, what would you expect to be true of the measured ΔT? Would this change make it more or less difficult to accurately determine the concentration of the solution? Explain your reasoning.

3. In this experiment, you are never asked to record the mass of the *t*-butyl alcohol solution. Why don't you need to record this to calculate the mass percent of water in the solution?

Temperature of ice bath (°C) ______________

Report Table 3.1: Freezing Point of Wet *t*-Butyl Alcohol

	Trial 1	**Trial 2**	**Trial 3**
Freezing point (°C)	______________	______________	______________

Average freezing point (°C) ______________

Corrected average freezing point (°C) ______________