Freezing Point Depression

Introduction

At a given pressure, the temperature at which a specific pure substance undergoes a particular phase change (freezing, boiling, sublimation, etc.) is always the same. For example, pure water at 1 bar pressure will always freeze at 0°C and boil at 100°C. However, if impurities (solutes) are added to the pure solvent in the liquid phase, the temperatures at which phase changes from the liquid occur will be different than those for the pure solvent. In particular the freezing point of a solution formed by mixing a pure liquid with a small amount of solute will decrease and the boiling point of the solution will increase. These effects are referred to as freezing point depression and boiling point elevation and both are part of a larger class of related phenomena called colligative properties - other examples include the lowering of the vapor pressure of a pure liquid by the addition of solutes and the phenomena of osmotic pressure, which is important in keeping blood cells from collapsing. In our everyday lives, the phenomenon of freezing point depression explains why road ice will melt when salt is poured on it (the freezing point is lowered by the addition of the impurity) and why we put ethylene glycol (antifreeze) in our automobile radiators to protect our cooling systems from freezing in extreme cold. Physically, freezing point depression and boiling point elevation occur because a solution consisting of a solvent (the formerly pure liquid) with a small amount of solute added is thermodynamically more stable than the pure liquid itself—primarily due to the fact that the addition of the solute increases the entropy of the solution. This additional stability increases the range of temperatures at which the liquid is stable - lowering the freezing point and increasing the boiling point.

In this lab you will be exploring how the freezing point depression of water depends upon the amount and nature of the added solute.

Figure 1: A New Yorker salts the sidewalk in front of their store before a blizzard[2].
Pre-lab

**Safety:** Goggles must be worn at all times.

Solid CO$_2$ (dry ice) can cause severe frostbite. Do not allow dry ice to touch bare skin. Do not ingest any of the chemicals used in this laboratory. In addition, make sure to wash your hands carefully before leaving the laboratory.

**Pre-lab Assignment:** Please write out the following in your lab notebook. This assignment must be completed before the beginning of lab. You will not be allowed to start the experiment until this assignment has been completed and accepted by your TA.

1) List all of the chemicals you will use for this week's experiment. For each chemical, list specific safety precaution(s) that must be followed. In order to find specific safety information, please obtain a Materials Safety Data Sheet (MSDS) on the chemical of interest. MSDSs can be found through an internet search (e.g., google) or from the following website: [www.hazard.com](http://www.hazard.com). Read the MSDS and find specific safety concerns for each chemical. List the chemicals you will use for the hands-on portion of this week's lab experiment. For each chemical, list specific safety precaution(s) that must be followed.

2) An automobile mechanic mixes ethylene glycol with water to make a solution of automobile antifreeze. This mixture contains 1104 g of ethylene glycol and 1.50 kg of pure water. What is the freezing point of this solution in degrees Celsius? (Assume 1 atm pressure.)

3) A zoologist combines NaCl with pure water for use in an outdoor aquatic facility. She makes the salt/water mixture by dissolving 15.0 kg of NaCl and 400.0 kg of pure water. What is the freezing point of this solution in degrees Celsius? (Assume 1 atm pressure.)

**Hint:** See pp. 600-602 in your Tro Chemistry textbook for help with Prelab questions #2 and #3.

Procedure

**Part 1 – Freezing of alcohol solutions**

For this investigation you will use the Vernier LabPro box and LoggerPro software to record the cooling curve and to determine the freezing point of a water/alcohol mixture. Each group will be assigned a different alcohol to use in the investigation; your group will be assigned one of the following:

<table>
<thead>
<tr>
<th>Alcohol</th>
<th>Formula</th>
<th>Molecular Weight (g/mol)</th>
<th>Freezing Point (°C)</th>
</tr>
</thead>
<tbody>
<tr>
<td>Methanol</td>
<td>CH$_3$OH</td>
<td>32.04</td>
<td>-93.9</td>
</tr>
<tr>
<td>Ethanol</td>
<td>CH$_3$CH$_2$OH</td>
<td>46.07</td>
<td>-117.3</td>
</tr>
<tr>
<td>1-Propanol (<strong>n</strong>-propyl alcohol)</td>
<td>CH$_3$CH$_2$CH$_2$OH</td>
<td>60.11</td>
<td>-126.5</td>
</tr>
<tr>
<td>2-Propanol (<strong>isopropyl</strong> alcohol)</td>
<td>(CH$_3$)$_2$CHOH</td>
<td>60.11</td>
<td>-89.5</td>
</tr>
<tr>
<td>1-Butanol (<strong>n</strong>-butyl alcohol)</td>
<td>CH$_3$CH$_2$CH$_2$CH$_2$OH</td>
<td>74.12</td>
<td>-89.5</td>
</tr>
</tbody>
</table>
Determine the freezing point of the alcohol solutions you are assigned. Make sure that you make replicate measurements of the freezing point. Follow your TA's instructions for sharing your data with the other groups.

1. Weigh out 10.0 grams of the alcohol that you were assigned. Pour the sample into a clean, dry 250-mL Erlenmeyer flask.

2. Weigh out 100.0 grams of distilled water and pour it into the Erlenmeyer flask containing the alcohol. Mix the liquids.

3. Prepare a dry-ice bath by filling a 600-mL beaker 2/3-full with dry-ice chunks. **CAUTION:** Be careful handling the dry ice as it can cause severe burns when it comes in contact with unprotected skin.

4. Set up the Temperature Probe and the LoggerPro software for collection of temperature data. You may find it helpful to open “Exp 02 Freeze Melt Water” from the Chemistry with Vernier Experiment files of Logger Pro. Each run should take less than 10 minutes.

5. Pour approximately 30 mL of distilled water into a 50-mL Erlenmeyer flask. (Running the experiment first with pure water allows you to determine a baseline freezing temperature.)

6. Place the temperature probe into the mouth of the flask; the probe should be submerged in the liquid.

7. Place the flask in the dry ice bath. **CAUTION:** Do not touch the dry ice with bare hands.

8. Begin collecting data. Make certain that you continually swirl the water in the flask to maintain a uniform temperature in your solution.

9. After the freezing point has been reached and is relatively stable, you may stop the data collection.

10. Repeat steps 6-10 using approximately 30 mL of your alcohol/water solution instead of pure water.

How does the freezing point you measured compare with the freezing point of the pure alcohol or the pure water? Explain your observations. Do you observe trends in the data for the different alcohols? Explain any trends that you observe. How might you use these results to determine the molecular weight of an unknown solute?
Part 2 – Determination of unknown from freezing point depression

Use your data from Part 1 to determine the change in freezing point expected for the alcohols used in this experiment at different concentrations. From this information you will determine the molecular weight of the unknown alcohol you investigate.

Following the procedure from Part 1, determine the freezing point of your unknown alcohol solutions. Using this data, and the freezing point depression data from the known alcohols above, determine the molecular weight of your unknown alcohol.

A trend you may have observed in Part 1 is that as the molecular weight of the alcohol increases, the magnitude of the freezing point depression ($\Delta T$) decreases. This is because the magnitude of a colligative property is, by definition, based on the number of solute particles in a solution. Each group prepares their sample by adding 10.0 grams of their assigned alcohol to 100.0 grams of water. Thus, the total number of alcohol molecules in a particular sample is inversely related to the molecular weight of the constituent alcohol. For example, 10.0 grams of methanol (molecular weight 32.04 g/mole) consists of a greater number of molecules than 10.0 grams of 1-butanol (molecular weight 74.12 g/mole). The equation relating the freezing point depression and concentration of a solution is

$$\Delta T_f = K_f m$$

where $\Delta T_f$ is the change in freezing point (the difference in the freezing point of a pure solvent and the freezing point of a solution using that solvent), $K_f$ is a constant for the solvent (the same for each of the solutions since each have water as the solvent), and $m$ is the concentration of the solution in molal units (moles of solute per kilogram of solvent). The $K_f$ of water is 1.86 ($^\circ$C/m). From your measurements of the amounts of solute and solvent used, determine the molality and also the number of moles of solute in your solution. Using the number of moles of solute and its mass, determine the molecular weight of the solute in g/mole.

Reference(s)

Glossary

**Boiling point elevation**

the effect of a solute that causes a solution to have a boiling point higher than that of the pure solvent

**Colligative property**

a property that depends on the amount but not type or identity

**Freezing point depression**

The effect of a solute that causes a solution to have a lower freezing point than that of the pure solvent