# Lab 4 - A Freezing Point Depression Study

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# 1 Pre-Lab

## 1.1 Purpose

1. Calculate molal concentration.
2. Calculate the freezing point depression of a solution.
3. Calculate the molecular weight of a substance from its freezing point depression.

## 1.2 Definitions

**Colligative Properties** - Properties of solutions that depend on the amount of solute particles in the solution (concentration) and are independent of the nature of the solute.

**Freezing Point Depression** - Drop in temperature at which a substance freezes, caused when a smaller amount of another, non-volatile substance is added. Colligative property.

**Boiling Point Elevation** - The difference in temperature between the boiling point of the pure solvent and that of the solution. Colligative property.

**Vapor Pressure Lowering** - When a solute is added to a solvent, the vapor pressure of the solvent (above the resulting solution) is lower than the vapor pressure above the pure solvent. Colligative property.

**Osmotic Pressure** - The pressure that would have to be applied to a pure solvent to prevent it from passing into a given solution by osmosis, often used to express the concentration of the solution. Colligative property.

**Solvent** - A solvent is a substance that dissolves a solute, resulting in a solution.

**Solubility** - The maximum amount of a substance that will dissolve in a given amount of a solvent at a specified temperature.

**Van't Hoff Factor i** - Measure of the effect of a solute on colligative properties. The ratio between the actual concentration of particles produced when the substance is dissolved and the concentration of a substance as calculated from its mass.

**Molality** - The number of moles of solute in a solution corresponding to 1 kg or 1000 g of solvent. (mol/kg)

### 1.3 Equations

Freezing Point Depression

$$\Delta T_f = k_f * m * i \tag{1}$$

$$\Delta T_f = \text{The change in the freezing point:}$$
$$T_{solution} - T_{puresolvent} = negativenumber \tag{2}$$

$k_f$ = molal freezing-point-depression constant (depends on solvent)

m = molal concentration of the solute particles.

i = van't Hoff factor, which, for dilute solutions, is the number of dissolved ions per mole of solute. (pg 496 in textbook)

density of water = 1.0 $\frac{g}{mL}$

$K_f$ for water = -1.86 $\frac{^{\circ}C}{m}$ (m = molal)

## 2 Flowcharts

*Please see the following page for flowchart.*

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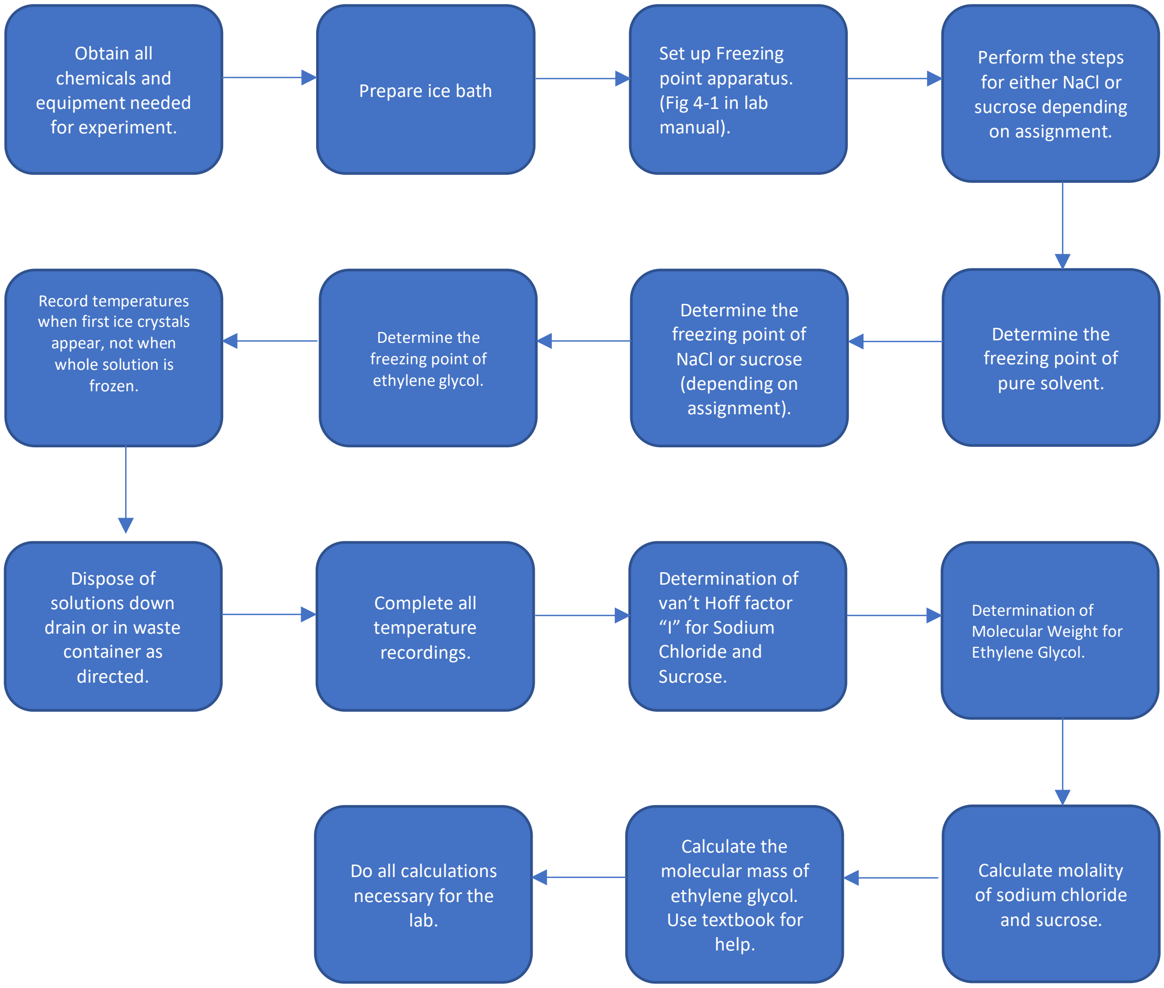

# 3 Waste Collection

## 3.1 NaCl and/or Sucrose

These solutions can be poured down the drain when finished.

## 3.2 Ethylene Glycol

Dispose in the properly labeled waste container in the fume hood.

# 4 Declaration of Academic Integrity

I certify that this is my own work and I understand that if I am found to be in violation of the honor code, I will be subject to the highest penalty.

Morgan Benavidez

# G. C. 2-4: LABORATORY REPORT

**Name:** Morgan Benavidez

**Lab partner's names:** Felicia, Rachelle

**Units, calculations, and significant figures will be counted toward your grade!**

Freezing point of DI water

Water

Sucrose - 34.0g

| Water | Time, min. | Temperature, °C |
|---|---|---|
| 18.0 | 0.5 | 20.1 |
| 16.2 | 1.0 | 16.5 |
| 11.2 | 1.5 | 12.9 |
| 9.2 | 2.0 | 9.9 |
| 7.5 | 2.5 | 7.5 |
| 4.6 | 3.0 | 5.4 |
| 3.2 | 3.5 | 3.9 |
| 1.7 | 4.0 | 2.8 |
| 1.1 | 4.5 | 1.7 |
| 0.0 | 5.0 | 0.7 |
| | 5.5 | -0.1 |
| | 6.0 | -0.6 |
| | 6.5 | -1.0 |
| | 7.0 | -1.4 |
| | 7.5 | -1.6 |
| | 8.0 | -1.8 |
| | 8.5 | -2.0 |
| | 9.0 | -1.7 |
| | 9.5 | -2.0 |
| | 10.0 | |
| | 10.5 | |
| | 11.0 | |
| | 11.5 | |
| | 12.0 | |

NaCl - 5.78g

| Time, min. | Temperature, °C |
|---|---|
| 0.5 | 19.1 |
| 1.0 | 16.3 |
| 1.5 | 12.9 |
| 2.0 | 9.3 |
| 2.5 | 6.2 |
| 3.0 | 3.9 |
| 3.5 | 2.2 |
| 4.0 | 0.7 |
| 4.5 | -0.4 |
| 5.0 | -1.3 |
| 5.5 | -1.8 |
| 6.0 | -2.4 |
| 6.5 | -2.9 |
| 7.0 | -3.4 |
| 7.5 | -3.9 |
| 8.0 | |
| 8.5 | |
| 9.0 | |
| 9.5 | |
| 10.0 | |
| 10.5 | |
| 11.0 | |
| 11.5 | |
| 12.0 | |

EG - 6.27g

| Time, min. | Temperature, °C |
|---|---|
| 0.5 | 23.1 |
| 1.0 | 19.0 |
| 1.5 | 15.3 |
| 2.0 | 12.9 |
| 2.5 | 10.9 |
| 3.0 | 5.3 |
| 3.5 | 4.4 |
| 4.0 | 1.9 |
| 4.5 | 0.3 |
| 5.0 | -1.0 |
| 5.5 | -0.8 |
| 6.0 | -1.5 |
| 6.5 | -1.8 |
| 7.0 | |
| 7.5 | |
| 8.0 | |
| 8.5 | |
| 9.0 | |
| 9.5 | |
| 10.0 | |
| 10.5 | |
| 11.0 | |
| 11.5 | |
| 12.0 | |

Freezing point of ______

Freezing point of ethylene glycol

Part 1

$\frac{34\cancel{g}}{342\ \cancel{g}/mol} = 0.1\ mols / 0.1\ Kg = 1\ \frac{mol}{Kg}$

Sucrose

$C_{12}\ (12) = 144g$

$H_{22}\ (1) = 22\ g$

$O_{11}\ (16) = \underline{176\ g}$

342 g/mol ← Molecular weight

$\Delta T_f = -2.0^{\circ C} - 0^{\circ C} = -2.0^{\circ}C$

$-2.0^{\circ}C = -1.86 \frac{^{\circ}C}{\cancel{mols}/\cancel{Kg}} \left(\frac{0.1\ \cancel{mols}}{0.1\ \cancel{Kg}}\right) i$

$\frac{-2.0^{\circ}\cancel{C}}{-1.86^{\circ}/C} = -1.86^{\circ C}(1)\ i$

$1.1 \approx i$

mols of Solute ↑ ↓ Mass of Solvent ↑ ↓ Molality ↑ ↓

Sodium Chloride

$\frac{5.78\cancel{g}}{58.45\ \cancel{g}/mol} = 0.1\ mols / 0.1\ Kg = 1\ mol/Kg$

$N_a\ (23) = 23\ g$

$Cl\ (35.45) = \underline{35.45\ g}$

58.45 g/mol ← molecular weight

$\Delta T_f = -3.9^{\circ C} - 0^{\circ C} = -3.9^{\circ}C$

$-3.9^{\circ}C = -1.86 \frac{^{\circ}C}{\cancel{mols}/\cancel{Kg}} \left(\frac{0.1\ \cancel{mols}}{0.1\ \cancel{Kg}}\right) i$

$\frac{-3.9^{\circ}\cancel{C}}{-1.86^{\circ}/C} = -1.86^{\circ C}(1)\ i$

$2.1 \approx i$

Part 2

$\Delta T_f = -1.8^{\circ C} - 0^{\circ C} = -1.8^{\circ}C$

$\Delta T_f = K \cdot \frac{m}{n} \cdot i = \frac{\Delta T_f}{K i} = \frac{m}{n} = \boxed{\frac{\Delta T_f (n)}{K \cdot i} = m}$

$$\frac{(-1.8^{\circ}\cancel{C})(0.1\,\cancel{kg})}{-1.86\,\frac{^{\circ}\cancel{C}}{\text{mols}/\cancel{kg}}\,(1)} = \frac{-0.18}{-1.86\,\text{mols}} = 0.1 \text{ mols} = m$$

$$\frac{6.27\ g}{0.1\ \text{mols}} = 62.7\ g/\text{mol}$$

## Part I: Determination of van't Hoff factor "i" for Sodium Chloride and Sucrose

Please perform calculations in your laboratory notebook.

Data from your lab partner must be shared and collected during class time and signed off by your TA.

Freezing point of DI water, °C: 0 °C

| Collected Data | Sodium Chloride | Sucrose |
|---|---|---|
| Freezing point (°C) | -3.9 °C | -2.0 °C |
| $\Delta T_f$ | 0 - 3.9 - 0 °C = -3.9 °C | -2.0 °C - 0 °C = -2.0 °C |
| Mass of solute (g) | 5.78 g | 34.0 g |
| Number of moles of solute | .1 mols | .1 mols |
| Mass of solvent (kg) | .1 Kg | .1 Kg |
| Molality (mol/kg) | 1 mol/Kg | 1 mol/Kg |
| Calculation of van't Hoff factor "i" | 2.1 | 1.1 |

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## Part II: Determination of Molecular Weight for Ethylene Glycol

| Freezing Point, °C | $\Delta T_f$ | Mass of Solute, g | Mass of Solvent, kg | Calculated Molecular Weight, g/mol |
|---|---|---|---|---|
| -1.8 °C | -1.8 °C - 0 °C = -1.8 °C | 6.27 g | .1 Kg | $\frac{6.27 g}{0.1 mol}$ = 62.7 g/mol |

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## Calculations, Conversions, and Values

1. Calculate the molality of the sodium chloride and the sucrose solutions.
2. For the mass of solvent, assume that the **density of water is 1.0 g/mL**.
3. Determine the value of "i," where **$K_f$ for water = −1.86°C/m (m = molal)**. Hint: Use Equation 1.
4. Calculate the molecular mass of ethylene glycol based on the freezing point depression. You may want to refer to your textbook for this calculation.

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Lab 4 – Freezing Point Depression Study

**Discussion and Conclusion**

When I prepared the ice bath, I added salt to lower the freezing point of the water. This allowed the bath to have a temperature lower than the freezing point of water without the water freezing. "If a solute is dissolved in a solvent, when the solution is cooled, it does not begin to freeze until the temperature has gone beyond the freezing temperature of the pure solvent." This is stated in our lab manual, and this applies to the principles/concepts behind this experiment, because the bath itself and all the solutions tested in this experiment have a freezing point that is less than that of the pure solvent (water). This means that nothing in this lab freezes until the temperature dips below 0 degrees Celsius (the freezing point of water).

The effect of electrolytes on colligative properties is greater than that of nonelectrolytes. The Van't Hoff factor for most non-electrolytes dissolved in water is 1. This was demonstrated in our experiment because I calculated the Van't Hoff factors of sodium chloride (electrolyte) and sucrose (nonelectrolyte) to be 2.1 and 1.1, respectively. These results vary from the ideal Van't Hoff factors for these compounds, but the experiment is not free of error (measurements and rounding).

The theoretical molecular mass of ethylene glycol is 62.07 g/mol. The molecular mass calculated through experimental data was 62.7 g/mol. This turned out to be surprisingly accurate, but I could see how this number could have a wide variance. It comes down to the precision and accuracy of measurements. This includes the initial measurement of the substance (ethylene glycol) and the freezing point. Changes in these numbers can lead to the molecular weight appearing to be much less or much greater than the theoretical molecular weight. I assumed the value of “i” to be 1 in my calculation because ethylene glycol is a nonelectrolyte and does not dissociate to form ions in aqueous solution.