Chem 112 – Experiment 1 – Simulation – Molecular Mass by Freezing Point Depression

Background

Freezing Point Depression

The freezing point of a pure solvent is lowered or depressed when a solute is added to form a solution. This can be seen in winter, when salting roads and sidewalks lowers the freezing point of melted snow such that it will not freeze unless the temperature falls more than a few degrees below zero. The number of degrees by which the melting point is depressed is proportional to the concentration of the solute.

The properties of a solution differ from the properties of the pure solvent due to interactions between the solute and solvent molecules. These are called colligative properties and include freezing point depression, boiling point elevation, vapor pressure decrease, and osmotic pressure increase. These properties depend only on the total concentration of the solute particles and not on the identity of the particles.

For example, a solution of glucose only contains glucose as solute particles, whereas a solution of sodium chloride contains both sodium and chloride ions as solute particles. Therefore, a solution of sodium chloride will cause a freezing point depression twice as large as a solution of glucose with the same concentration.

A non-volatile solute is a solute that does not readily evaporate. When the solute is dissolved in a solvent, the vapor pressure of the solvent is lowered, which leads to both an increase in the solvent's boiling point and a decrease in its freezing point. The amount by which the freezing point is depressed depends on the solute concentration as shown in the formula below.

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

Where:

ΔT_f : is the change in freezing point.
i : is the van't Hoff factor.
K_f : is the molal freezing point depression constant.
*K_f depends only on the solvent, and for water **K_f is 1.86 °C/m**.*
m : is the molality of the solution.

The change in temperature is the difference between the freezing point temperature of the pure solvent and the freezing point temperature of the solution. To calculate the change in freezing point temperature, use the formula below.

$$\Delta T_f = T^0_f - T_f$$

Where:

T_{f0} : is the freezing temperature of the pure solvent and
T_f : is the freezing temperature of the solution.

The van't Hoff factor is a measure of the number of particles per unit of solute. It is equal to the ratio between the actual concentration of solute particles in solution and the concentration of the solute as calculated from its mass. In general, i is 1 for compounds that do not dissociate and equal to the number of ions formed in solution for most ionic substances. For example, for glucose i is 1, whereas for NaCl i is 2. The van't Hoff factor is also about 1 for weak electrolytes such as weak acids which only slightly dissociate in water.

This rule is an approximation and mostly true for ideal solutions. Deviations occur when the solutions are concentrated and solute particles can associate in solution. For example, carboxylic acids such as acetic acid dimerize in benzene, decreasing the number of solute particles in half. As another example, a 0.001 m NaCl solution has a van't Hoff factor of 1.97, whereas a 0.1 m NaCl solution has a van't Hoff factor of 1.87.

The molality of a solution is defined as the number of moles of solute per kilogram solvent. Molality is used instead of molarity because molarity can change based on temperature. Large temperature changes change the solution volume, therefore changing molarity. Molality is a more accurate measure of concentration because the mass of solvent in kilograms does not vary with temperature. To calculate molality, use the formula below.

$$m = \frac{n_{solute}}{m_{solvent(kg)}}$$

Where

n_{solute} : is the number of moles of solute.

$m_{solvent}(kg)$: is the mass of the solvent in kg.

Molar Mass Determination

The molar mass of an unknown compound can be determined by measuring the freezing point of a solution of that compound. Using water as the solvent, the molar mass of an unknown water soluble substance can be calculated.

To find the molar mass, consider the definition of molar mass below.

$$MM = \frac{m}{n}$$

Where

MM: is the molar mass in g/mol

m: is the mass of solute in g.

n: is the number of moles of solute.

Combining this with the molality equation gives the equation below for molality.

$$m = \frac{m_{solute(g)}}{MM_{solute} \times m_{solvent(kg)}}$$

Rearranging this equation, gives the following formula for molar mass.

$$MM_{solute} = \frac{m_{solute(g)}}{m \times m_{solvent(kg)}}$$

Where

m: is the molality of the solution in mol/kg.

$m_{solute(g)}$: is the mass of solute in g.

$m_{solvent(kg)}$: is the mass of solvent in kg.

Measuring the Freezing Point

The temperature of a pure substance remains constant at the freezing point because heat equal to the enthalpy of fusion, the heat required for the liquid to change states, is released as the liquid solidifies. The temperature will only start to decrease again once all the liquid has frozen.

The freezing point of the solution is determined in the same way as the freezing point of the solvent. The cooling curve, however, does not become horizontal at the freezing point. Rather, the freezing point may decrease as the solution freezes due to changes in the composition of the liquid.

At the freezing point, the solution enters a solid phase as a pure solvent. This is because it is energetically unfavorable to incorporate solute particles into the crystal structure of the solid solvent. Therefore, the solute is kept out of the solid phase as the solvent freezes which takes extra energy. As freezing occurs the concentration of the solute in the liquid phase increases until it is saturated, thereby decreasing the freezing point even further.

Freezing point depression can also be explained using entropy changes. Adding a solute to a solvent increases the entropy (disorder) of the system and therefore more energy is required for freezing. This is because freezing is a process that creates order and therefore decreases the entropy of the system.

About This Lab

In this lab, you will determine the molar mass of two unknown substances by measuring the freezing point depression of aqueous solutions of these substances. The **first unknown is an organic compound** with a van't Hoff factor of 1, whereas the **second unknown is an ionic compound** with a van't Hoff factor of 2. To determine the molar mass of each unknown substance from the freezing point depression, the freezing points of both the pure solvent and the solution will be measured.

Open the simulation by clicking on the virtual lab icon shown on the left on the Hayden-McNeil Web Site. The simulation will launch in a new window.

You may need to move or resize the window in order to view both the Procedure and the simulation at the same time.

Follow the instructions in the Procedure to complete each part of the simulation. When instructed to record your observations, record data, or complete calculations, record them for your own records in order to use them later to complete the post-lab assignment.

Procedure

Experiment 1a – Measure the Freezing Point of Pure Water.

1. Take a clean **test tube** from the **Containers shelf** and place it on the workbench.
2. Take **a balance** from the **Instruments shelf** and place it on the workbench.
3. Place the test tube on the balance. Zero the balance.
4. Take **water** from the **Materials shelf** and **add 10mL** to the **test tube**.
   NOTE: *In a classroom laboratory, you would remove the container from the balance before you add any chemicals or solutions. Filling the container separately prevents any spills, sprays, or splashes that might affect your measurements or damage the balance. Luckily virtual labs are spill-, spray-, and splash-free, so you can skip that step to save time.*
5. **Record** the **mass of the water** to reference later and move the test tube to the workbench.
6. Take a **thermometer** from the **Instruments shelf** and **attach it to the test tube**.
7. Take a **constant temperature bath** from the **Instruments shelf** and place it on the workbench. **Run the bath at -15°C.**
8. **Move the test tube** into the **constant temperature bath**.
   NOTE: *In a classroom laboratory, the test tube would be completely submerged in the temperature bath. In the virtual lab, the test tube is situated so that you can observe the process clearly.*
9. **Watch** the **temperature of the water** in the test tube **decrease**. The water should begin to freeze. You will see some solid ice form in the test tube. You may want to zoom in on the test tube for a closer view.
10. **Record** the **first temperature** when **ice just begins to appear** as the freezing point of the solvent.
11. Clear your station by dragging the thermometer back to the shelf, emptying the test tube into the waste, then placing the test tube in the sink.

Experiment 1b – Measure the Freezing Point of a Solution of an Unknown Substance.

1. Take a clean **test tube** from the **Containers shelf** and place it onto the workbench.
2. Move the **test tube** onto the **balance**. **Zero the balance**.
3. Take **FP sample 1** from the **Materials shelf** and **add about 2g** to the **test tube**. **Record** the **exact mass of FP sample 1** added.
4. Take **water** and add **10mL** to the test tube. **Record** the **total mass of the water and FP sample 1**.
5. Take a **thermometer** from the **Instruments shelf** and **attach it to the test tube**.
6. **Move the test tube** into the **constant temperature bath**. Make sure it is **still set to -15°C**.

7. **Record** the **freezing point of the solution** as the first temperature **when ice begins to appear**.

8. Clear your station by emptying the test tube in the waste, then placing in the sink.

9. **Repeat steps 1-7** but using **unknown FP sample 2**.

10. Clear the bench of all materials, containers, and instruments, then use the **Report File** that you downloaded from the **General Chemistry web site**. When completed **send it to your TA.**