# Experiment 18: Determination of Molecular Mass by Melting Point Depression

## Abstract

In this experiment, you will explore the colligative properties of the melting point depression of water. An aqueous sodium chloride solution of known concentration will be used to calibrate your PASCO temperature sensor. The molecular mass of an unknown compound will be determined by measuring the melting point depression when added to water.

## Background

Impure mixtures or solutions will exhibit lower melting points than their pure solvents. The degree of depression is dependent on the amount, rather than identity, of the impurity introduced, and is called a *colligative property*. In everyday life, these colligative properties are used to melt ice on sidewalks and streets during freezing weather by sprinkling salt on top. Salt is often added to the ice in a home ice cream freezer to melt the ice and depress the temperature of the ice-salt mixture in the outer bucket of the freezer to about -10°C.

The freezing point depression ($\Delta T_f$), or the number of degrees the freezing point of the solution deviates from the freezing point of the pure solvent, for a solution can be calculated from Eqn. 1:

$$\Delta T_f = K_f \cdot m \qquad (1)$$

Here, $K_f$ is the molal freezing point depression constant that is characteristic of the solvent. For water, $K_f$ = 1.86 °C/m. The molality of the solution (m) is calculated from Eqn. 2:

$$\text{molality (m)} = \frac{\text{moles of solute}}{\text{kg of solvent}} \qquad (2)$$

Remember that the molar mass for a compound is simply grams compound/moles compound.

Thus the molar mass of an unknown compound can be determined by measuring the freezing point depression when in solution with a known solvent. Using Eqn. 1, the molality of the solution can be calculated. If the kg of solvent is known, Eqn. 2 will allow calculation of moles of solute present. The molar mass of the solute is then the grams of solute divided by the moles.

**EXAMPLE:** When 2.35 grams of unknown was dissolved into 11.8 grams of water, the observed freezing point of the solution was -2.15°C. Determine the molar mass of the unknown.

From Eqn. 1: $\Delta T_f = K_f \cdot m$
Rearranging: $m = \Delta T_f/K_f$ = (2.15°C)/(1.86 °C/m) = 1.16 m
Converting: 11.8 g $H_2O$ · (1 kg/1000 g) = 0.0118 kg $H_2O$
From Eqn. 2: 1.16 m = (mols solute)/(kg solvent) = (mols solute)/(0.0118 kg)
Rearranging: mols solute = 0.0137 mols
Molar Mass: (2.35 g solute)/(0.0137 mols solute) = <u>172 g/mol</u>

**Procedure**

COMPUTER SETUP

1. Plug the temperature sensor probe into Analog A on the science workshop. Turn on the Science Workshop Interface and computer.
2. Open the necessary file on the computer:
   - If you have a temperature sensor with a box attached that says RTD, open the folder RTD on the desktop and then the file RTDVC14.ds

   - If you have a temperature sensor that has a blue end (model CI-6605A), open the folder BLUE on the desktop and then the file BLUEVC14.ds

   - If you have an all black temperature sensor (model CI-6505B), open the folder BLACK on the desktop and then the file BLACKVC14.ds

3. Delete any data left on the screen by left clicking on the data to select it and then pressing the delete key on your keyboard.
4. The "Sampling Options" for this lab have been set to 2 per second (2 Hz).

CALIBRATION (You may do this part of the experiment with a lab partner)

1. Add 10 mL of deionized water to a 6 inch test tube and stand the tube in a test tube rack or beaker. Insert the temperature sensor into the test tube with the water.
2. Nest two 8 oz. Styrofoam cups together. Fill the inner cup about ¾ full of ice. Add about 50 mL of saturated salt (NaCl(aq), brine) solution to the ice already present in the cup. Sprinkle approximately 10 to 15 grams of solid rock salt (NaCl) on top of the ice in the cup.
3. On the computer, click START.
4. Insert the test tube with the probe into the ice-salt mixture in the cups. Stir the mixture while gently moving the probe up and down inside the tube to stir its contents and continue to record.
5. As the tube cools, the curve on the graph will continue to decrease until it eventually shows a definite change of slope. Continue collecting data until the straight line created after the slope change is at least half as long as the previous line or until the slope of the cooling curve changes again and the sample has entirely frozen. Click STOP.
6. Adjust the graph axes so that your data fills the graph. Do this by clicking on the axes and adjusting the minimum and maximum values to fit the range of your data.
7. Select FIT at the top of the graph, then "Linear". Select the first portion of the curve (it should be a fairly straight line) by dragging a box around this area. The equation for this straight line should appear in the right of the window. Record it in your lab book.
8. Print a copy of your graph for your report by first clicking on the graph then clicking on file, then "Print."
   - Do not print your data table; it is too long.
9. Locate a straight-line portion of your cooling curve present after the slope changed. Select it by clicking and dragging a box around this area. Record this equation in your lab book.

10. The point where the two lines intercept is the freezing point and, at this point, will have the same values of x and y. Therefore, setting the two equations equal to one another will allow the value of x to be calculated through algebra. Once the value of x is known, it can be substituted into either equation to solve for the value of y. This is the melting point for water measured by your sensor.

    This melting point may or may not correspond to water's actual melting point of 0°C. Its variation from the actual melting point serves as your sensor's calibration.
11. If time permits, repeat the calibration again and use the average of multiple trials.

<u>UNKNOWN DETERMINATION</u> (This part of the experiment must be done individually)

1. Weigh a clean, dry 6 inch test tube in a 150 mL beaker on an analytical balance. Record the mass of the empty tube.
2. Add about 10 mL of deionized water to the tube and reweigh. Record the mass.

**CAUTION:** Exercise care when handling your unknown in order to ensure the highest accuracy.

3. Use a dropper to transfer 0.8-1.0 grams of the unknown into the water inside the test tube.
    - It is important to use a clean, dry dropper to dispense your unknown to eliminate possible contamination, which will ruin your results.
    - Some unknowns evaporate readily. You must drop the unknown down into the water in the bottom of the tube and not let it drizzle down the sides of the tube in order to prevent your unknown from evaporating away. Hold the tube and dropper vertical while dispensing.
4. Reweigh the test tube, beaker, water, and unknown. Record the mass.
5. Determine the freezing point of the solution of unknown in the same manner you used above to find the freezing point of water.
6. Print a copy of the graph for your report.
7. If time permits, repeat the determination again and use the average of multiple trials.

**NAME: ____________________________________**

# REPORT SHEET
## Experiment 18

CALIBRATION

**Trial 1**

Equation for the first part of the cooling curve:

Equation for the second part of the cooling curve:

Freezing point of water observed with the sensor (show calculation set-up): ______________°C

**Trial 2**

Equation for the first part of the cooling curve:

Equation for the second part of the cooling curve:

Freezing point of water observed with the sensor (show calculation set-up): ______________°C

**Average Freezing Point of Water: ________________°C**

UNKNOWN DETERMINATION

**Data Table** **UNKNOWN NUMBER: _____________**

| | TRIAL 1 | TRIAL 2 |
|---|---|---|
| **Mass of empty test tube** | | |
| **Mass of tube + water** | | |
| **Mass of tube + water + sample** | | |
| **Mass of water** | | |
| **Mass of sample** | | |

**Trial 1**

Equation for the first part of the cooling curve:

Equation for the second part of the cooling curve:

Freezing point of solution observed with the sensor (show calculation setup): ____________°C

$\Delta T_f$ of solution observed (observed FP of $H_2O$ – observed FP solution): _________________°C

Molality of the solution (show calculation setup): ______________m

Moles of solute present in solution (show calculation setup): _________mols

Molar mass of the unknown solute (show calculation setup): ____________g/mol

**Trial 2**

Equation for the first part of the cooling curve:

Equation for the second part of the cooling curve:

Freezing point of solution observed with the sensor (show calculation setup): ____________°C

$\Delta T_f$ of solution observed (observed FP of $H_2O$ – observed FP solution): _________________°C

Molality of the solution (show calculation setup): __________________________________m

Moles of solute present in solution (show calculation setup): ________________________mols

Molar mass of the unknown solute (show calculation setup): _______________________g/mol

**Average Molar Mass for Trials 1 & 2:**____________________ g/mol

**Unknown Number:** _________________________

**NAME: ___________________________________**

**Post-Lab Assignment – Experiment 18**

1. Is the density of a compound a colligative property? Why or why not?

2. While calibrating the temperature sensor, a student carelessly wrote down the wrong observed freezing point of pure water as -2.5°C rather than the true reading of -5°C. The student performed the rest of the experiment accurately. Describe this single error's effect (high, low, or the same) on the following values and justify your answers.

a. $K_f$ of water

b. $\Delta T_f$ of the unknown solution

c. Molality of the unknown solution

d. Mass of the unknown compound

e. Molar mass of the unknown compound

A corollary colligative property to freezing point depressing is boiling point elevation. In this case, a solution will begin to boil at a higher temperature than its corresponding pure solvent from the following equation:

$$\Delta T_b = K_b \cdot m$$

Here, $\Delta T_b$ is the change in boiling point, $K_b$ is the molal boiling point elevation constant individual to the solvent, and m is the solution's concentration in units of molality.

Answer the following questions about boiling point elevation.

3. An aqueous solution is prepared with 11.2 g of NaCl dissolved into 42.7 g $H_2O$. Calculate the boiling point elevation of this solution, if $K_b$ for water is 0.512 °C/m.

4. Benzene has a $K_b$ of 2.53 °C/m and a normal boiling point of 80.2°C. Calculate the molality of a solution of benzene and an unknown compound that boils at a temperature of 95.6°C.

5. If the above solution was created by dissolving 15.62 g of unknown into 42.3 g benzene, calculate the molecular mass of the unknown compound in g/mol.

**NAME: ___________________________________**

**Pre-Lab Assignment – Experiment 18**

1. Calculate the molality of a solution prepared from 0.9813 grams glucose ($C_6H_{12}O_6$) dissolved into 8.9547 g water.

2. Find the new freezing point for the above solution ($K_f$ for water is 1.86 °C/m).

3. Why is it important to drop your liquid unknown directly into the water rather than allowing it to run down the inside surface of the test tube?

4. A solution is prepared by dissolving 0.8763 g unknown into 10.4832 g of water.  The freezing point of the solution was found to be -3.12 °C.  Find the molar mass of the unknown.