

# Experiment 1: Colligative Properties

## Determination of the Molar Mass of a Compound by Freezing Point Depression.

Objective: The objective of this experiment is to determine the molar mass of an unknown solute by measuring the freezing point depression of a solution of this solute in a solvent as compared to the freezing point of the pure solvent.

Background: **Colligative properties** are properties of a solvent, such as freezing point depression and boiling point elevation, which depend on the concentration of solute particles dissolved in the solvent. The decrease in freezing point,  $\Delta T_f$  (freezing point depression) for a near ideal solution can be described by the equation:

$$\Delta T_f = k_f \cdot m \quad \text{Eq 1}$$

where  $k_f$  is the **molal freezing point depression constant** of the solvent with units  $^{\circ}\text{C} \cdot \text{kg}$  solvent/mole solute.  $m$  is the molal concentration of the solute dissolved in the solvent expressed as moles of solute/kg solvent.

Since the molar mass **M** (traditionally and often, but erroneously called the molecular weight) of a compound has units g/mole, we can solve for moles and substitute the result into the molal concentration relationship, and then into Eq 1 as is shown below.

$$\mathbf{M} = \text{g/ mole} \quad \text{Eq 2}$$

Rearranging Eq 2 gives

$$\text{moles} = \text{g}/\mathbf{M} \quad \text{Eq 3}$$

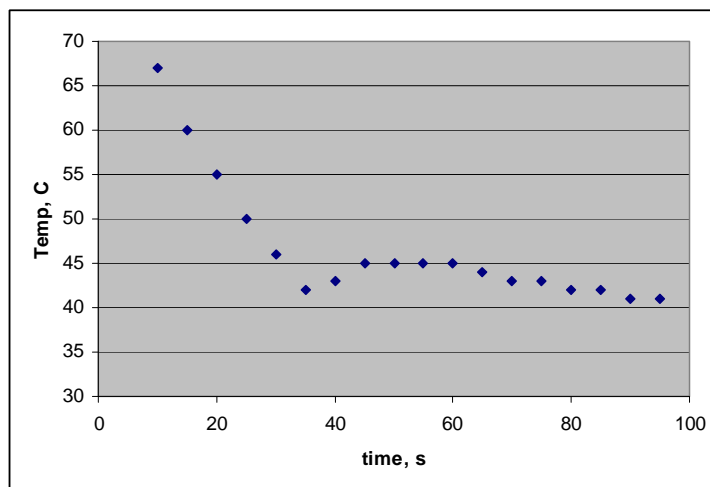
Now substituting Eq 3 into the unit definition of molality yields

$$m = \text{g}/(\mathbf{M} \cdot \text{kg solvent}) \quad \text{Eq 4}$$

And substituting Eq 4 into Eq 1 gives

$$\Delta T_f = (k$$

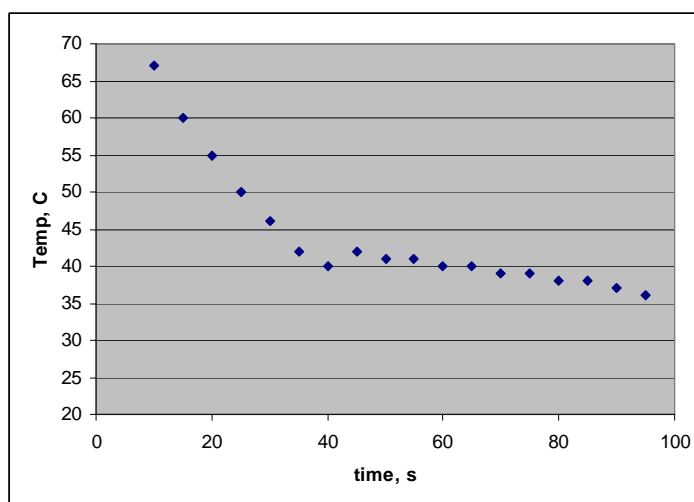
The temperature at which this equilibrium exists is the freezing point of the substance. Sometimes this temperature is difficult to determine, so the use of **cooling curves** is required. To construct a cooling curve one would warm their sample, pure solvent or solution, to well above its melting point, then allow it to cool. As the sample cools the temperature of the sample is monitored as a function of time. As the sample begins to solidify the change in temperature will slow, and at the equilibrium shown by Eq 7 the temperature will be constant until all of the sample has solidified. A graph is made by plotting the temperature vs. time. An example of a cooling curve is shown below in Figure 1.



**Figure 1: Cooling curve for a pure solvent.**

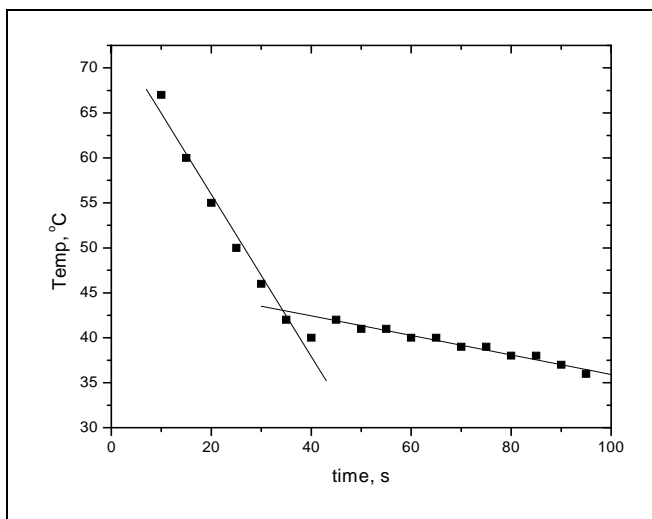
In this cooling curve you see a steady decrease in temperature followed by a dip which is followed by a slight rise in the temperature. This dip is not unusual and results from **supercooling** during the early stages of the freezing process. In this example the dip is followed by a short plateau in the temperature. This plateau is at the freezing point of the pure solvent as shown in Figure 1.

When solute is added to the solvent the shape of the cooling curve sometimes changes so that we don't see a clear horizontal plateau as the example shown in Figure 1.



**Figure 2: Cooling curve for a solution.**

In Figure 2 we don't see a clear horizontal plateau. In this case we must draw a trend line through the data points corresponding to the cooling of the liquid and a trend line through the data points corresponding to the freezing of the liquid. The temperature at the point where those two lines intersect is the freezing point of the solution.



**Figure 3: Solution cooling curve**  
showing best fit straight lines through the two portions  
of the curve discussed in the text.

Figure 3 shows an example with the trend lines drawn in and the intersection of the lines. In the example shown in Figure 3 the freezing point would be measured as about 43 °C. If we continue to record and plot the temperature of the solid, the data points may start to deviate from the trend line

**direct flame!)** Insert the stirring loop into the test tube and then insert the thermometer so that the loop of the stirrer surrounds the thermometer. Periodically stir the *t*-butanol with the stirring loop by an up and down motion while warming it.

Place about 250-300 ml of ice in the other 400 ml beaker and enough cold tap water to just cover the ice. Once the temperature of the *t*-butanol has warmed to about 40 °C, transfer the test tube to the ice-water bath making sure that most of the *t*-butanol is below the surface of the ice-water bath, add more ice if needed. Immediately begin to take temperature readings and record them in the table every 15

**Data:**

- 1) Mass of test tube: \_\_\_\_\_
- 2) Mass of test tube and *t*-butanol: \_\_\_\_\_
- 3)     a) Mass of *t*-butanol (line 2 – line 1): \_\_\_\_\_  
       b) Mass of *t*-butanol in kgs: \_\_\_\_\_
- 4) Mass of first sample of unknown:
- 5) Mass of second sampp

**Data Table:**

<i>t</i> -butanol, pure solvent		<i>t</i> -butanol plus first sample portion, solution 1		<i>t</i> -butanol plus second sample portion, solution 2	
time	temperature	time	temperature	time	temperature

### **Data Handling, Calculations and Questions:**

Use graph paper and plot temperature vs. time for the pure *t*-butanol and for each solution analyzed; make 3 different graphs. Use the discussion in the background section above as a guide and determine the freezing point,  $T_f$ , of *t*-butanol and of each solution. Clearly mark on each graph all your data points and the best fit lines you used to determine each freezing point. Determine the  $\Delta T_f$  of solution 1 by subtracting the  $T_f$  of solution 1 from the  $T_f$  of pure *t*-butanol. Determine the  $\Delta T_f$  of solution 2 by subtracting the  $T_f$  of solution 2 from the  $T_f$  of pure *t*-butanol.

Use the  $\Delta T_f$  for solution 1 along with the mass of unknown in solution 1, line 4 of the data table, the mass of solvent, *t*-butanol, line 3b of the data table, and the  $k_f$  of *t*-butanol,  $9.10\text{ }^\circ\text{C}\cdot\text{kg solvent/mol solute}$ , in Eq 6 to determine the molar mass, **M**, of your unknown compound.

Repeat the calculation above for solution 2 remembering to use the total mass of solute, line 6 of the data table, and the  $\Delta T_f$  for solution 2.

### **Report:**

In your lab report briefly discuss the theory behind why the freezing point of a solution is typically lower than the freezing point of pure solvent.

Reproduce the data from the Data page and the Data Table, pages 5 and 6, and turn in with your report, along with all three graphs.

Make a data and calculations page to report the  $T_f$  of pure *t*-butanol and for each solution. Show all calculations and report the molar mass of unknown as determined in each of the two solutions.

Average the two molar masses you determined and calculate the percent difference for each determination vs. the average. This should give you about a 6 or 7 page lab report, not all pages will be full of text.

### **Reference:**

Some information at the following lab experiment was used for this experiment.

<http://infohost.nmt.edu/~jaltig/FreezingPtDep.pdf>