

# MOLECULAR MASS BY FREEZING POINT DEPRESSION

## LAB FP.1

Juniata College Science Outreach

*New and Improved Introductory Manual, 2000*

## INTRODUCTION

Once a solution has been formed, the solution displays properties different from those of the solvent used to prepare it. Two common examples of solvent properties that are altered are depression of the freezing point and elevation of the boiling point. The changes in these properties are related to the number of particles present in the solution and not the type of particle present. Properties that depend only on the number of particles are referred to as colligative properties. Colligative properties such as freezing point depression can be used to calculate the molecular weight of a soluble solid. To complete this calculation, the mass of solute and solvent must be known as well as the freezing points of the pure solvent and the solution. In this experiment the molecular weight of urea will be determined.

## PURPOSE

The purpose of this laboratory activity is to determine the molecular weight of urea using the technique of freezing point depression.

## EQUIPMENT/MATERIALS

test tube or large vial	ice
urea	table salt
250 or 400 mL beaker	foam coffee cup(s)
thermometer or temperature probe and lab pro	

## SAFETY

- Always wear apron and goggles in the lab.

## **PROCEDURE**

### **Part I:** Freezing Point of Pure Water

1. Obtain a clean, dry test tube or vial. Determine the mass of the test tube or vial. Place about 10 mL of distilled water in the test tube or vial, and reweigh. Determine the mass of the water used. Record the mass of the water in the data table.
2. Prepare an ice bath in a foam cup with ice and table salt. Place the cup in a beaker to give it more stability. The ice bath should be deep enough so that it is above the level of the water in the test tube or vial but well below the top. Take care not to let any of the salt or ice get into the sample of distilled water.
3. Place a thermometer or temperature probe in the distilled water. Take time-temperature data every half-minute until ice has formed in the test tube or vial. It is not necessary to freeze the entire sample. Record the temperature at which the sample froze.
4. **Do not** discard the sample of the distilled water, because the sample will be used in Part II.

### **Part II:** Molecular Weight of the Unknown

5. Remove the test tube or vial containing the distilled water from the ice bath. Allow the ice to melt. This step can be speeded up by placing the test tube or vial in a beaker of tap water.
6. Weigh out approximately 1 gram of urea. Record the mass of the sample in the data table. Add the urea to the distilled water, and stir until it is all dissolved. Return the test tube or vial to the ice bath. Insert the thermometer or temperature probe.
7. Take time-temperature data as in Part I. Again, the sample does not have to be frozen solid in order to determine the freezing point. Record the freezing point in the data table.
8. Repeat the procedure (both Parts I and II).

Name \_\_\_\_\_

Name \_\_\_\_\_

Period \_\_\_\_\_

Date \_\_\_\_\_

## MOLECULAR WEIGHT BY FREEZING POINT DEPRESSION

### DATA TABLE

Trial 1	Trail 2
Mass, water	Mass, water
T <sub>f</sub> water	T <sub>f</sub> water
Mass urea	Mass urea
T <sub>f</sub> solution	T <sub>f</sub> solution

### CALCULATIONS

1. Using the change in freezing point, the kilograms of water used, and the freezing point constant for water, calculate the number of moles of urea used in each trial.

2. Using the mass of urea and the number of moles of urea, calculate the molecular weight for each trial.

3. Calculate the average molecular weight for urea and the percent error for the trials.



## QUESTIONS

1. What differences would be expected if an ionic compound such as sodium chloride were used instead of urea?
2. Why is it not necessary to wait for the entire sample of water to freeze in order to determine its freezing point?
3. Why is it a good idea to measure the freezing point of the water instead of assuming that its freezing point is exactly  $0^{\circ}\text{C}$ ?
4. What would have happened if a two-gram sample of urea were used in this experiment?