ESSENTIAL EXPERIMENTS
for
CHEMISTRY

Morrison

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Sample Experiment

Freezing Point Depression

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Experiment16B Freezing Point Depression

The temperature at which a substance changes state (melting, freezing, boiling) at standard conditions is a definite fixed quantity, characteristic of that substance. However, if a second substance is mixed with the original substance, the change of state temperature of the original gets altered. This is particularly important when dealing with aqueous solutions. For example, dissolving a solute in water causes the freezing point to be lowered, as the molecules of water have more difficulty aligning with each other to form a pure crystal. On the other hand, the boiling point of an aqueous solution is higher than the boiling point of pure water. In this experiment you will measure the depression of the freezing point of three different solutions.

When a pure substance like water freezes, the temperature stays constant throughout the phase change. However with a solution, the temperature continues to drop as more water freezes and is removed from the solution, making the remaining solution more concentrated. (In a similar manner, the boiling point of a solution continues to rise as more water evaporates.) In addition, when a solution freezes, *supercooling* often occurs. This is a situation in which the temperature briefly drops below the freezing point, then rises slightly when the crystals actually form. The temperature you need to record is not the lowest temperature reached, but the temperature at which the ice crystals first start forming. (Refer to the sketch graphs in Figure 16B-1.)

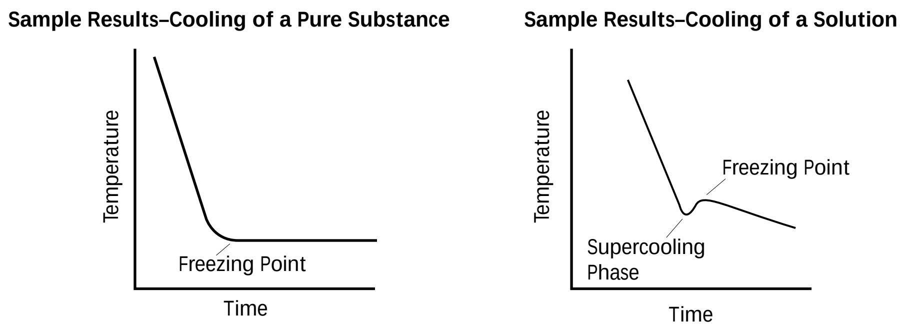

Figure 16B-1 *Sample results for cooling of a pure substance and a solution*

The concentration of a solution used in freezing point depression experiments must be expressed in molality, not molarity. The molality of a solution is the number of moles of solute added to 1 kg of solvent. The symbol used for molality is a small case m, italicized: *m*. The freezing point depression of a solute (ΔT) is given by the equation:

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

where K_f is the freezing point depression constant (°C/*m*),
m is the molality of the solution,
and *i* is the number of particles (molecules or ions) produced by one molecule of the solute.

After measuring the freezing point of water, the three solutions for which you will measure the freezing point depression are 1 *m* sucrose, 1 *m* NaCl and an unknown. You will first determine the K_f for water from the freezing point depression observed with a 1 *m* solution of sucrose. Then you will use the value of K_f to determine *i*, the number of particles produced by a molecule of sodium chloride, NaCl. Finally, you will determine the molality of a solution of an unknown molecular solute, and from that, the molar mass of the solute.

OBJECTIVES

1. to determine K_f (the freezing point depression constant) for water using a 1 *m* sucrose solution
2. to determine the number of particles produced by one molecule of NaCl from the freezing point depression of a 1 *m* solution of NaCl
3. to determine the molar mass of an unknown molecular solid

SUPPLIES

Equipment
large styrofoam cup
18 mm × 150 mm test tubes (4)
thermometer (preferably –30°C to + 50°C) or temperature probe
centigram balance
100 mL beaker
stirring rods (2)
test tube rack
crushed ice
lab apron
safety goggles

Chemical Reagents
1.00 *m* sucrose (sugar, $C_{12}H_{22}O_{11}$) solution
1.00 *m* sodium chloride (table salt, NaCl) solution
unknown molecular solid
sodium chloride (NaCl) crystals
distilled water

PROCEDURE

1. Put on your lab apron and safety goggles.
2. Obtain approximately 20 mL of NaCl crystals in a 100 mL beaker, and add it to a large styrofoam cup half-filled with crushed ice. Stir to mix the contents.
3. Measure the temperature of the ice/salt mixture, and proceed with the next step when the temperature is –10°C or lower.
4. Obtain 4 test tubes, label them 1-4 and place them in a test tube rack.
5. Add distilled water to test tube 1 until it is about one-third full, and place it in the ice/salt mixture. Stir constantly, and periodically lift the test tube out enough to observe the contents. When ice crystals are seen beginning to form, insert the thermometer and measure the temperature of the test tube contents to obtain the freezing point of water. Record this value in your copy of Table 1 in your notebook.

CAUTION

Your thermometer is made of glass and can easily break, leaving sharp edges that cut. Handle your thermometer gently. Do not use it to crush or stir ice. If your thermometer breaks, call your instructor. If it contains mercury, be aware that mercury liquid and vapor are very poisonous.

6. Add 1.00 *m* sucrose solution to test tube 2 until it is about one-third full, and place it in the ice/salt mixture. Proceed as in Procedure 5 to measure and record the temperature at which crystals of ice are first seen to form in the sucrose solution.
7. Repeat using the 1.00 *m* solution of sodium chloride, NaCl in test tube 3.
8. Make up your own solution of the unknown solid by weighing an empty 250 mL beaker, placing approximately 100 mL of water in it and weighing again. Add the unknown solid until you have added approximately 20 g, and again record the exact mass of the beaker and contents.
9. Stir until the solid has dissolved, then pour some into test tube 4 until it is approximately one-third full. Proceed to measure and record the freezing point of the solution as before. Make sure the temperature of the ice/salt mixture is still approximately –10°C, and add more ice if necessary.
10. Pour all left over solutions down the sink, and wash your hands before leaving the laboratory.

REAGENT DISPOSAL

All leftover solutions and the salt/ice mixture may be safely washed down the sink with plenty of water.

POST LAB CONSIDERATIONS

You may find when you calculate the result for the number of particles (ions) in the NaCl solution that it is not as high as you might predict it be, and this is to be expected. In ionic solutions there is a certain amount of interionic attraction and formation of ion pairs that leads to less dissociation of the neutral compound than there would be if there were no interaction. Chemists use the term *activity* to represent the effective concentration of an ion in a solution, but a further discussion of this topic is beyond the scope of this experiment.

The depression of the freezing point has many practical applications. The antifreeze used in a car's cooling system is a solution of ethylene glycol (1,2-ethandiol) in water. If pure water were used, the expansion that occurs when water turns to ice would cause serious damage to the engine in winter conditions. It is necessary to keep the antifreeze in the engine in summer as well, because it also increases the boiling point and prevents the engine coolant from boiling over at high operating temperatures. Windshield washing fluid contains cleaning agents but also usually has methanol added to lower the freezing point of the solution.

In areas where winter temperatures cause ice and snow to make driving conditions difficult, several different kinds of salts are sometimes spread on the roads. This lowers the freezing point and allows the snow and ice to melt. However this is only useful if the temperature is just a few degrees below 0°C. Below about –10°C salt will not help.

EXPERIMENTAL RESULTS

Table 1

Freezing point of water (test tube 1)	
Freezing point of 1.00 *m* sucrose (test tube 2)	
Freezing point of 1.00 *m* NaCl (test tube 3)	
Mass of empty beaker	
Mass of beaker + water	
Mass of beaker + water + unknown solid	
Freezing point of unknown's solution (test tube 4)	

COMPLETE IN YOUR NOTEBOOK

ANALYSIS OF RESULTS

1. Calculate the freezing point depression (ΔT, in °C) for the sucrose solution. Using the equation $\Delta T = K_f \times \mathrm{m} \times i$,calculate K_f for water. Note that the molality of the sucrose solution is 1.00 *m* and that sucrose dissolves as a complete molecule, giving *i* the value of 1.
2. Using the same equation, calculate the value of *i* for NaCl from the freezing point depression (ΔT) observed for NaCl, and the K_f value calculated in Analysis 2. Remember that the molality of the NaCl was 1.00 *m*.
3. The unknown solid dissolves as a complete molecule, giving *i* the value of 1. Using the ΔT for the unknown, and the K_f value calculated in Analysis 2, calculate the molality of the unknown's solution.
4. Using the mass of the unknown solid and the mass of the water used in making the solution, express the concentration in terms of grams of solid per 1000 g of water.
5. Calculate the molar mass of the solid from the mass /1000 g water and the molality.

FOLLOW-UP QUESTIONS

1. A sample of sea water was analyzed and found to be 0.47 *m* for the NaCl content. Disregarding other minerals that may be present, calculate the freezing point of this sample of sea water.
2. The substance mercury(I) nitrate was originally thought to have the formula $HgNO_3$, but is now known to be $Hg_2(NO_3)_2$ because the mercury(I) ion has the formula Hg_2^{2+}. Freezing point depression data provided some of the evidence for this discovery. What value for *i* would be obtained for a 1 *m* solution based on the formula $HgNO_3$?
3. Salt (NaCl) can result in corrosion problems on cars and bridges in areas where it is used to de-ice roads. Use a reference source to find out what other chemicals are now sometimes used instead, with their advantages and disadvantages.
4. Calculate the freezing point of 1 *m* acetic acid.

CONCLUSION

State the results of Objectives 1, 2, and 3.