AP Chemistry By Satellite
Laboratory Manual
Instructor's Edition

Experiment 10:
PROPERTIES OF WATER

Contents | Page

Reagents

CHEMICAL	**Amount**
sodium chloride	2.0 g
sucrose	7.0 g
rock salt	Several chunks
ice	
distilled water	100 mL

Equipment

EQUIPMENT	Amount
laboratory balance	1
beaker, 100 mL	1
beaker,600 mL	1
glass rod	1
graduated cylinder, 25 or 50 mL	1
paper towels	several
ring stand	1
rubber stopper, 2-hole with slit	1
thermometer, ˚C	1
test tube, 18x150 mm	2
utility clamp	1
wire stirrer	1

Useful Experimental Comments

Experiment Scheduling

This experiment will require approximately 2-3 hours to complete.

Parts I and **II** can usually be completed during the first one hour laboratory period and **Part III** during the second. If necessary, students can stop work after recording the freezing point of any solution in any part of the experiment.

College Board Recommendations: This experiment is a variation of the type of freezing point depression experiment recommended for AP Chemistry students by the College Board. It emphasizes such recommended skills as observation, recording data, and calculating and interpreting results based on quantitative data obtained. The use of a sensitive balance is recommended.

Remind your students to put the thermometer through the *split* side of the 2–hole rubber stopper.

Students should read thermometers to the greatest accuracy permitted by the calibration. Be sure that students understand how the thermometer is calibrated before they begin the experiment. Some students become very confused when reading temperatures below zero.

Students should be sure the level of the liquid in the test tube is at or slightly below the level of the liquid in the ice–water bath when determining the freezing points.

If the students are not careful they will completely freeze the solution in the test tube around the thermometer. If this happens, it is likely the student will break the thermometer if they are not careful. The student should remove the test tube from the ice–water bath and warm it carefully using their hands. If they are observant they may be able to obtain a good freezing point without returning the test tube to the ice–water bath.

The sucrose may not dissolve quickly, so remind your students to check that all the sucrose dissolves before measuring the freezing point. Use a glass rod to promote dissolution, however be careful not to poke out the bottom of the test tube.

Students should not use the rock salt intended for depressing the freezing point of the water baths as the salt in their test solutions in **Part III** of the experiment.

Approximate experimental values to expect for the freezing point depression are given below. Have the students repeat experiments if their data is not close to these values.

The temperature change for 2.2 g of $C_{12}H_{22}O_{11}$ should be something less than 1˚C (0.8 ˚C).
The temperature change for 4.4 g of $C_{12}H_{22}O_{11}$ should be close to –1.5 ˚C.
The temperature change for 0.5 g of NaCl should be close to –2.1 ˚C.
The temperature change for 1.0 g of NaCl should be close to –4.2 ˚C.

Students should be reminded of the definition of molality (m) and molarity (M).

$$m = \frac{\text{moles solute}}{\text{kg solvent}} \qquad M = \frac{\text{moles solute}}{\text{L solution}}$$

Some students will also need guidance in the questions dealing with proportionality. For example;

$a \propto b$ states that 'a' is directly proportional to 'b'. As 'a' values increase 'b' values increase by a like (or multiple) amount.

$a \propto \frac{1}{b}$ states that 'a' is inversely proportional to 'b'. As 'a' values increase 'b' values decrease by a like (or multiple) amount.

EXPERIMENT 10: PROPERTIES OF WATER

Pre-lab Questions:

The following preparatory questions should be answered before coming to class. They are intended to introduce you to several ideas important to aspects of the experiment. You must turn-in your work to your instructor before you will be allowed to begin the experiment. Be sure to bring a calculator and paper to laboratory.

1. Calculate the molality of glucose in a solution prepared by mixing 1.258 g $C_6H_{12}O_6$ in 14.0 mL of water.

$$\mathbf{1.258\ g \left(\frac{1\ mol}{180\ g}\right)\left(\frac{1}{0.0140\ kg}\right) = 0.499\ molal}$$

2. Briefly discuss the term *freezing point depression*.

Freezing point depression is the lowering of the temperature at which a solution freezes relative to the freezing point of the pure solvent. The depression is a result of the lower solvent vapor pressure of the solution compared to the pure solvent.

3. (a) Complete the table.

Exp. #	Mass of $C_6H_{12}O_6$	Moles of $C_6H_{12}O_6$	Mass of H_2O	Volume of solution	Molality of solution	Molarity of solution	Density of solution
1	1.600 g	**0.008889**	15.50 g	16.52 mL	**0.5735**	**0.5381**	**1.035 g/mL**
2	1.160 g	**0.006444**	15.50 g	16.21 mL	**0.4158**	**0.3976**	**1.028 g/mL**
3	0.798 g	**0.00443**	15.50 g	15.98 mL	**0.286**	**0.277**	**1.02 g/mL**
4	0.465 g	**0.00258**	15.50 g	15.78 mL	**0.166**	**0.163**	**1.01 g/mL**
5	0.310 g	**0.00172**	15.50 g	15.68 mL	**0.111**	**0.110**	**1.01 g/mL**

3. (b) Why are molarity and molality more nearly equal for dilute solutions than for concentrated solutions?

As the mass of solute decreases, while the mass of the solvent remains constant, the molality of the solution and the molarity of the solution approach each other in magnitude. This can be further realized by the observed density of the solution. It is approaching the density of pure water. Therefore, for very dilute solutions the molality of the solution and the molarity of the solution are equivalent.

3. (c) Under what conditions will molarity and molality of a solution be equal? Explain.

$$\textbf{molality} = \frac{\textbf{moles solute}}{\textbf{kg solvent}} \qquad \textbf{molarity} = \frac{\textbf{moles solute}}{\textbf{L solution}}$$

$$\textbf{kg solvent} = \textbf{L solution}\left(\frac{\textbf{kg solution}}{\textbf{L solution}}\right) - \textbf{kg solute}$$

for dilute aqueous solutions: kg solute ≅ 0

$$\left(\frac{\textbf{kg solution}}{\textbf{L solution}}\right) = 1$$

therefore;

kg solvent = L solution

and

molality = molarity

4. Why is it desirable to express concentration as molality rather than molarity when performing experiments in which the temperature changes?

Molality is calculated using only mass measurements. Mass is constant with temperature variations, so the concentration of the solution expressed in molality does not change with a change in temperature. Molarity, however, is calculated based upon the volume of the solution. Most substances change in volume (expand or contract) with changes in temperature. Thus the concentration of a solution expressed as molarity would change during the experiment.

EXPERIMENT 10: PROPERTIES OF WATER

EQUIPMENT:

laboratory balance......................1
beaker, 100 mL 1
beaker, 600 mL 1
glass rod.................................1
graduated cylinder, 25 or 50 mL 1
paper towels......................... several
ring stand................................1
rubber stopper, 2-hole with slit.......1
thermometer, ˚C........................1
test tube, 18x150 mm.................2
utility clamp............................1
wire stirrer...............................1

It is suggested that students work in groups of two while collecting the experimental data.

The apparatus to be used in the following experiments is illustrated in Figure I (next page). The apparatus consists of an 18 x 150 mm test tube which will hold the solution whose freezing point is to be determined. The test tube is fitted with the two-holed rubber stopper. A thermometer is carefully inserted into the split hole of the rubber stopper and the wire stirrer is inserted through the other hole. A ring clamp is used to hold the test tube assembly, which is immersed in a ice/water/salt bath in a 600 mL beaker. A glass stirring rod is used to stir the ice/water/salt solution, while the wire stirrer is used mix the solution in the test tube. If you have any questions about the apparatus consult with your instructor.

Part I: The Freezing Point of a Pure Liquid

Add approximately 15 mL of distilled water to the 18 x 150 mm (3/4-inch) test tube. Assemble the apparatus, with the test tube containing the water immersed in the ice-rock salt mixture. (For best results use mostly ice & salt with just a little water.) The test tube should be inserted deeply enough that the entire water sample is below the level of the ice-rock salt level in the beaker. Stir both the water and the ice-rock salt solution continuously. Watch the mercury level in the thermometer drop. The lowest constant temperature observed is the freezing point of the water. When the temperature reaches a constant value, read the thermometer to the nearest 0.5 ˚C. NOTE: The solution may cool below the true equilibrium freezing point because there are no nucleation sites for solidification to occur. This is called supercooling. Stir the solution rapidly and the temperature will increase to the equilibrium freezing point and remain constant. (See Figure II - next page).

Obs. #1 Freezing Point of Pure Water 0.0 ˚C

Remove the test tube from the beaker. Dispose of the water sample and dry the test tube for later use.

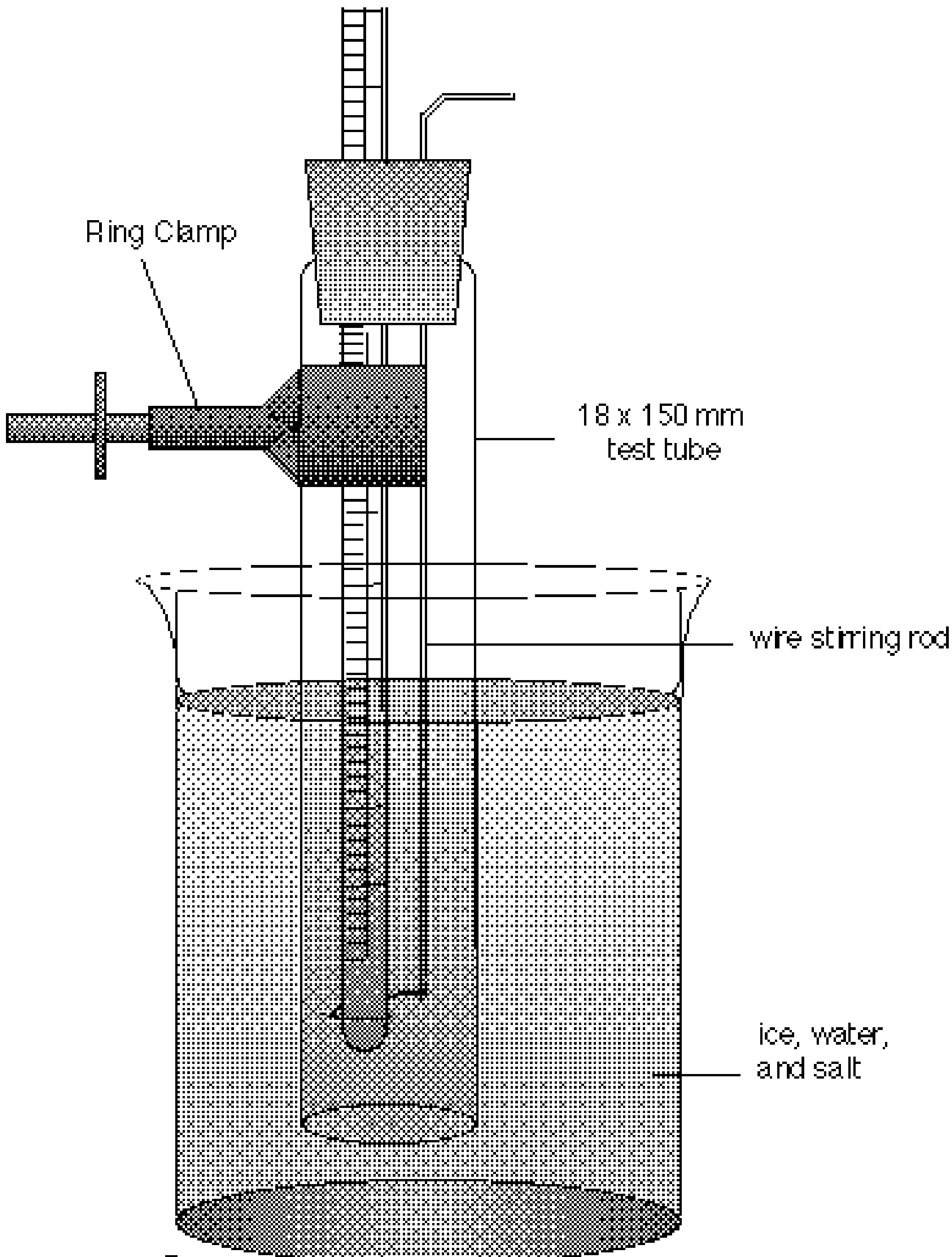

Figure I.

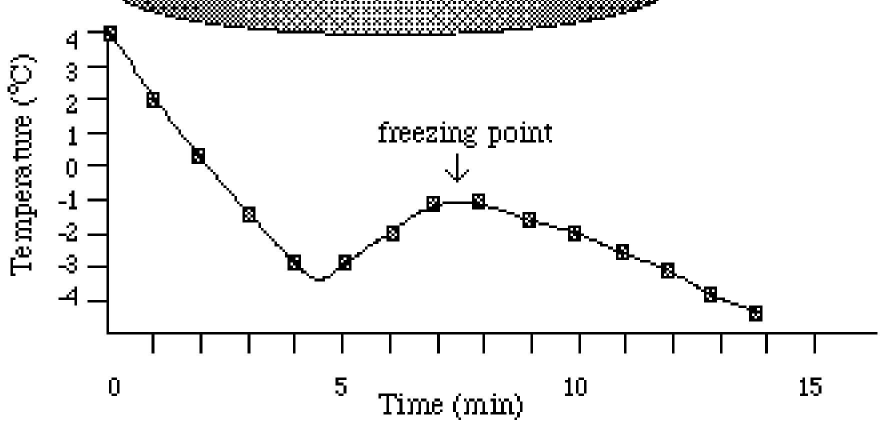

Figure II.

Part II: The Freezing Point of Sucrose Solutions

Weigh a clean, dry 18 x 150 mm test tube, supported in a 100 mL beaker, using the laboratory balance and record the mass below. Add between 15 and 16 mL of distilled water to the test tube. Record the mass of the water+test tube+beaker below (Note: This mass is entered twice for ease of calculations). Carefully add about 2.2 g of sucrose ($C_{12}H_{22}O_{11}$) to the test tube and record the mass of sucrose+water+test tube+beaker. After all of the sucrose has dissolved, assemble the apparatus with the test tube containing the solution immersed in the ice-rock salt mixture. The test tube should be inserted deeply enough that the entire solution sample is below the level of the ice-rock salt level in the beaker. Stir both the sucrose solution in the test tube and the ice-rock salt solution continuously. Measure the freezing point of this solution and record the temperature. Watch for supercooling!

Mass: sucrose + H_2O + test tube + beaker	**92.108 g**	||	Mass: H_2O + test tube + beaker	**89.695 g**
Mass: H_2O + test tube + beaker	**89.695 g**	||	Mass: test tube + beaker	**74.320 g**
Mass of sucrose	**2.413 g**	||	Mass of water	**15.375 g**

Calculate the molality of the sucrose solution.

Calc. #1

$$\mathbf{2.413\ g\left(\frac{1\ mol}{342\ g}\right)\left(\frac{1}{0.015375\ kg}\right) = 0.4589\ molal}$$

Obs. #2 Freezing Point of the Solution –0.8 ˚C

Obs. #3 Change in Freezing Point of the Solution 0.8 ˚C (Obs #1 - Obs #2)

Repeat the procedure above using about 4.4 g of sucrose. Record your data and observed freezing point below.

Mass: sucrose + H_2O + test tube + beaker	**87.200 g**	||	Mass: H_2O + test tube + beaker	**82.684 g**
Mass: H_2O + test tube + beaker	**82.684 g**	||	Mass: test tube + beaker	**68.024 g**
Mass of sucrose	**4.516 g**	||	Mass of water	**14.660 g**

Calculate the molality of the sucrose solution.

Calc. #2

$$\mathbf{4.516\ g\left(\frac{1\ mol}{342\ g}\right)\left(\frac{1}{0.014660\ kg}\right) = 0.9007\ molal}$$

Obs. #4 Freezing Point of the Solution –1.5 ˚C

Obs. #5 Change in Freezing Point of the Solution 1.5 ˚C (Obs #1 - Obs #4)

Complete the following table.

Table I.

	Molality of sucrose solution (m)	Change in freezing point (ΔT)
Experiment #1	**0.4589**	**0.8**
Experiment #2	**0.9007**	**1.5**

How did the freezing point change with change in the molality of the sucrose solution?

Expl. #1

The most obvious observation is, as the molality of the solution increased the freezing point increased. However, some of your students may go a step further and set–up a ratio;

$$\frac{0.9007}{0.4589} = 1.936 \cong 1.9 \qquad \frac{1.5}{0.8} = 1.9$$

Write a mathematical proportionality statement based on your data in Table I.

$$\textbf{Molality} \ \alpha \ \Delta\textbf{T}$$

In order to make a proportionality into an equality, we need to incorporate a constant into the equation. (For example, if a variable *a* is directly proportional to another variable *b*, we can write *a* α *b*. To convert the proportionality statement to an equation we write $a = kb$, where k is a proportionality constant.) We need an equation relating the molality of the solution to its change in freezing point. We will call the required constant, K_f, the freezing point constant. Write the equation obtained by including K_f in your proportionality statement. So that all answers in the class will be consistent, be sure to place the constant on the "molality side" of the equation. Calculate the magnitude of the freezing point depression constant based on your data. (<u>Be sure to include the correct units for the constant</u>.)

Calc #3

$$\Delta T = K_f \text{ Molality} \qquad \frac{\Delta T}{\text{molality}} = K_f$$

For Experiment #1

$$\frac{0.8}{0.4589} = 1.74 \ \frac{^\circ C}{\text{molal}}$$

For Experiment #2

$$\frac{1.5}{0.9007} = 1.67 \ \frac{^\circ C}{\text{molal}}$$

$$\text{Average} = 1.71 \ \frac{^\circ C}{\text{molal}}$$

Part III: Freezing Point of Sodium Chloride Solutions

Weigh a clean, dry 18 x 150 mm test tube, supported in a 100 mL beaker, using the laboratory balance and record the mass in the data sheet. Add between 15 and 16 mL of distilled water to the test tube. Record the mass of the water+test tube+beaker on the data sheet. Carefully add about 0.5 g of sodium chloride to the test tube and record the mass of sodium chloride+water+test tube+beaker. After all of the sodium chloride has dissolved, assemble the apparatus with the test tube containing the solution immersed in the ice-rock salt mixture. The test tube should be inserted deeply enough that the entire solution sample is below the level of the ice-rock salt level in the beaker. Stir both the salt solution in the test tube and the ice-rock salt solution continuously. Measure the freezing point of this solution and record the temperature. Watch for supercooling!

Mass: NaCl + H_2O + test tube + beaker	**89.723 g**	Mass: H_2O + test tube + beaker	**89.226 g**
Mass: H_2O + test tube + beaker	**89.226 g**	Mass: test tube + beaker	**74.335 g**
Mass of NaCl	**0.497 g**	Mass of water	**14.891 g**

Calculate the molality of the sodium chloride solution.

Calc. #4

$$0.497\ \text{g} \left(\frac{1\ \text{mol}}{58.45\ \text{g}}\right)\left(\frac{1}{0.014891\ \text{kg}}\right) = 0.571\ \text{molal}$$

Obs. #6 Freezing Point of the Solution –2.1 ˚C

Obs. #7 Change in Freezing Point of the Solution 2.1 ˚C

Repeat the procedure above using about 1.0 g of sodium chloride . Record your data and observed freezing point below.

Mass: NaCl + H_2O + test tube + beaker	**90.642 g**	Mass: H_2O + test tube + beaker	**89.655 g**
Mass: H_2O + test tube + beaker	**89.655 g**	Mass: test tube + beaker	**74.674 g**
Mass of NaCl	**0.987 g**	Mass of water	**14.981 g**

Calculate the molality of the sodium chloride solution.

Calc. #5

$$0.987\ \text{g} \left(\frac{1\ \text{mol}}{58.45\ \text{g}}\right)\left(\frac{1}{0.014891\ \text{kg}}\right) = 1.127\ \text{molal}$$

Obs. #8 Freezing Point of the Solution –4.2 ˚C

Obs. #9 Change in Freezing Point of the Solution 4.2 ˚C

Complete the data table.

Table II.

	Molality of sucrose solution (m)	Change in freezing point (ΔT)
Experiment #1	**0.571**	**2.1**
Experiment #2	**1.127**	**4.2**

How did the freezing point change with change in the molality of the sodium chloride solution?

Expl. #2

The most obvious observation is, as the molality of the solution increased the freezing point change increased. However, some of your students may go a step further and set–up a ratio;

$$\frac{1.127}{0.571} = 1.974 \cong 2.0 \qquad \frac{4.2}{2.1} = 2.0$$

Write a mathematical proportionality statement based on your data in Table I.

Molality α ΔT

In order to make a proportionality into an equality, we need to incorporate a constant into the equation. (For example, if a variable *a* is directly proportional to another variable *b*, we can write $a \alpha b$. To convert the proportionality statement to an equation we write $a = kb$, where k is a proportionality constant.) We need an equation relating the molality of the solution to its change in freezing point. We will call the required constant, K_f, the freezing point constant. Write the equation obtained by including K_f in your proportionality statement. So that all answers in the class will be consistent, be sure to place the constant on the "molality side" of the equation. Calculate the magnitude of the freezing point depression constant based on your data. (Be sure to include the correct units for the constant.)

Calc #6

$$\frac{\Delta T}{\text{molality}} = K_f$$

For Experiment #1

$$\frac{2.1}{0.571} = 3.7 \frac{{}^\circ C}{\text{molal}}$$

For Experiment #2

$$\frac{4.2}{1.127} = 3.7 \frac{{}^\circ C}{\text{molal}}$$

Is there any difference between the constant obtained from Part II (sucrose solution) and the constant obtained in Part III (sodium chloride solution)? If so, how do they differ?

Expl. #3

Yes, there is a difference. The constant in Part III is 3.7 ˚C/molal while in Part II it is about 1.71 ˚C/molal. The constant in Part III is twice as large as the constant determined in Part II.

What would you have to do to the constant from Part III to get close to the constant from Part II?

Expl. #4

If the constant in Part III is halved it would almost equal the constant in Part II.

Write the chemical equation for what occurs when sucrose is dissolved in water.

Equ. #1

$$C_{12}H_{22}O_{11(s)} \xrightarrow{H_2O} C_{12}H_{12}O_{11(aq)}$$

Write the chemical equation for what occurs when sodium chloride is dissolved in water.

Equ. #2

$$NaCl_{(s)} \xrightarrow{H_2O} Na^+_{(aq)} + Cl^-_{(aq)}$$

For every mole of sucrose that dissolves, how many moles of 'particles' (sucrose molecules) are in solution?

Expl. #5

For every mole of sucrose that dissolves there is one mole of sucrose molecules in solution. The sucrose molecule does not dissociate when it dissolves. In aqueous solution sucrose is a molecular species, just as it is in the solid phase.

For every mole of sodium chloride that dissolves, how many moles of 'particles' (ions) are in solution?

Expl. #6

For every mole of sodium chloride that dissolves there are two moles of ions formed in solution. There are twice as many particles produced per mole of solute for sodium chloride compared to sucrose.

The equation "$\Delta T = K_f m$" describes the freezing point depression of a solvent due to the addition of a nonvolatile solute. The freezing point constant, K_f, for a given solvent is constant and independent of the identity of the nonvolatile solute particles. In other words, the freezing point constant has to be the same in Part II and Part III of the experiment because water is the solvent in both parts. Explain the reason for the difference in the values obtained in Calc. #3 and Calc. #6.

Expl. #7

The difference is due to the number of particles in solution. If the constants must be the same in Part II and Part III, then an additional factor is needed. This factor must be related to the number of particles produced when a solute dissolves in water. In the relationship;

$$\Delta T = K_f m$$

'm' must be interpreted as the molality of *particles* in solution. For sucrose the molality of the solution is the same as the molality of particles. But for sodium chloride the molality of the solution is half the molality of particles in solution. Given equal concentrations of sucrose and sodium chloride the freezing point change is twice as much for the sodium chloride solution compared to the sucrose solution because the sodium chloride solution contains twice as many particles.

If the freezing point constant, K_f, is constant, on what does the freezing point of an aqueous solution of a nonvolatile solute depend?

Expl. #8

The freezing point of a particular solution will depend on the number of particles present in the solution. Solutions of equal concentrations which have different numbers of particles in solution will have different freezing point depressions.

Post-Laboratory Questions

These problems are to be turned in with the laboratory write-up. SHOW ALL CALCULATIONS.

1. Write a chemical equation which describes what happens when the following substances are added to water.

   (a) $C_2H_4(OH)_{2(l)}$ **—H_2O→ $C_2H_4(OH)_{2(aq)}$**

   (b) $Ca(NO_3)_{2(s)}$ **—H_2O→ $Ca^{2+}(aq)$ + $2NO_3^{-}(aq)$**

2. Complete the following table.

Substance	Molar Mass	Mass of Solute	Mass of Water	Ideal Freezing Point ˚C	Molality of Solution
Sugar	342	122.0 g	**1.1×10^3 g**	**–0.60**	0.32
Potassium Iodide	166	**2.1 g**	140 g	**–0.33**	0.090
Ethylene Glycol	**62**	1.5 g	25.0 g	–1.80	**0.968**
Calcium nitrate	164	**126**	1000 g	-4.30	**0.771†**
Urea	60	**68.1 g**	610 g	-3.46	**1.86**

† This is the molality of $Ca(NO_3)_2$ not the molality of particles. To get this number the calculation is;

$$\frac{4.30}{1.86} = 2.31 \text{ molal of particles} \left(\frac{1 \text{ mol } Ca(NO_3)_2}{3 \text{ mol particles}}\right) = 0.771 \text{ molal } Ca(NO_3)_2 \text{ soln.}$$

3. Calculate the boiling point of each of the solutions in Problem #2.

Substance	**Boiling Point**
Sugar	**100.16 ˚C**
Potassium Iodide	**100.092 ˚C**
Ethylene Glycol	**100.49 ˚C**
Calcium nitrate	**101.18 ˚C**
Urea	**100.80 ˚C**

4. Freezing point depression is often used to experimentally determine the molecular mass of a solute, but boiling point elevation is rarely used. Considering the calculations you have just performed in questions #2 and #3, explain why this is so.

   The magnitude of the freezing point depression is much greater than the boiling point elevation. Also, the experimental conditions and equipment requirements for determining a freezing point elevation are less rigorous compared to those required for boiling point elevations.

5. A solution of 0.684 g of chlorous acid, $HClO_2$, in 100 g of water freezes at –0.24 ˚C. Would you classify $HClO_2$ as a nonelectrolyte, a weak electrolyte, or a strong electrolyte?

   $$\mathbf{0.684\ g\ HClO_2 \left(\frac{1\ mol}{68.45\ g}\right)\left(\frac{1}{0.1\ kg}\right) = 0.0999\ molal}$$

   The concentration of particles can be calculated from $\Delta T/K_f$;

   $$\mathbf{\frac{0.24}{1.86} = 0.129\ molal\ of\ particles}$$

   If $HClO_2$ behaved as a nonelectrolyte, the concentration should be 0.0999 molal; however the concentration is greater. So more than 1 mol of particles are produced per mol of $HClO_2$. If $HClO_2$ were a strong electrolyte,

	$HClO_2$ $\xrightarrow{H_2O}$	$H^+(aq)$ +	$ClO_2^-(aq)$
Initial	**.0999**	**0**	**0**
Final	**0**	**.0999**	**.0999**

   Total moles of particles = 0.200 molal. The experimental freezing point suggests 0.129 molal of particles, so $HClO_2$ must be a weak electrolyte.

INSTRUCTOR EVALUATION NAME: ____________________
EXPERIMENT 10: PROPERTIES OF WATER SCHOOL: ____________________

Please complete the form as soon as possible after your students have completed the laboratory. Include any comments you have on each section of the experiment. If the answer to any question is "no" please note the specific problems or difficulties encountered. Attach extra sheets if necessary. At the end of the semester, return all forms to **Dr. John Gelder, Department of Chemistry, Oklahoma State University, Stillwater, OK 74078**. Your comments and suggestions are very important in helping to correct errors and improve the overall quality of this manual.

1. How much time was required to complete the experiment? _______ hours
   Briefly describe those section of the experiment which were completed during each laboratory period. (Note: You may include Part numbers or page numbers for simplicity.)

	NO	YES
2. Was the pre-lab exercise ...		
A. ...completed by the the students prior to the laboratory session?	_____	_____
B. ... adequate introduction to the ideas introduced in the experiment?	_____	_____

Comments:

	NO	YES
3. Were the laboratory instructions ...		
A. ...understood by the students with little or no assistance from you?	_____	_____
B. ...leading to the collection necessary data?	_____	_____
C. ...resulting in data with acceptable experimental error?	_____	_____

Comments:

	NO	YES
4. Were the questions and calculations included in the experiment ...		
A. ...completed by most students?	_____	_____
B. ...relevant to the experiment?	_____	_____

Comments:

	NO	YES
5. Were the post-lab problems ...		
A. ...completed by most students?	_____	_____
B. ...relevant to the experiment?	_____	_____
C. ...sufficient to illustrate the overall goals of the experiment?	_____	_____

Comments:

	NO	YES
6. Was the experiment as a whole ...		
A. ...interesting to the students?	_____	_____
B. ...relevant to the course work?	_____	_____
C. ...written at an appropriate level of difficulty?	_____	_____

Comments: