

Experiment 2

SOLUTIONS AND THEIR PROPERTIES

INTRODUCTION

I. Factors Affecting Solubility

Solutions are familiar to us. The water we drink is a solution of various minerals dissolved in water. Plants derive their food, in part, from the water-soluble constituents of the soil. In the digestion of food, a chemical change takes place whereby constituents of food are converted into soluble substances that can be absorbed by the blood and carried to various tissues in the body. The atmosphere we breathe is a solution of nitrogen and oxygen gases. Fish and marine life derive the oxygen necessary for their life processes from oxygen dissolved in water. Many chemical changes are brought about in solution since the rate of reaction in a solution is faster than between the substances in the pure state. This is probably due to the better contact of the reacting substances.

A solution is a homogeneous mixture. It is a mixture because it contains two or more substances and because their proportions can be varied. It is homogeneous because its composition and properties are uniform throughout the mixture. A solution is composed of two components: the solute and the solvent. The solute is the component that is present in lesser amount and is said to be dissolved in the solvent. The solvent is the component that is present in the greatest amount or that determines the state of matter in which the solution exists.

Solutions can be classified according to the amount of solute dissolved in a given amount of solvent. A saturated solution contains the maximum amount of solute that can be dissolved in a given amount of solvent at a certain temperature. An unsaturated solution contains less amount of solute than the maximum amount that the solvent can dissolve at a certain temperature. This solution can still dissolve more solute until it becomes saturated. A supersaturated solution contains more solute than the maximum amount that the solvent can dissolve at a certain temperature. This solution can be prepared by cooling a saturated solution. Ordinarily when a saturated solution is cooled, excess solute will crystallize out at the lower temperature. In some instances, it is possible to cool a saturated solution without the separation of the excess solute. The cooled solution then contains more of the solute than is present in a saturated solution at that particular temperature. This solution is very unstable. Addition of a tiny crystal of the solute (known as seeding) or agitation causes immediate crystallization.

The solubility of a substance is the weight of a substance dissolved by a given weight or volume of solvent to form a saturated solution at a given temperature. For a solution of a solid in a liquid, solubility is commonly expressed in terms of grams of solute per 100grams of solvent. Solubility varies greatly with the nature of solute and solvent, temperature, and pressure.

A. Effect of Nature of Solute and Solvent

Based on molecular structure, solvents are classed as polar and nonpolar. A polar solvent consists of dipoles — molecules in which constituent atoms are unsymmetrically arranged so that the center of positive and negative electrical charge is not located at the same point in the molecule. A nonpolar solvent consists of symmetrical molecules with the center of positive and negative electrical charge at the same point within the molecule. Cohesive forces between molecules of polar solvents are large compared with forces tending to hold molecules of nonpolar solvents together since positive ends of the dipoles will tend to attract negative ends of other dipoles and hold the molecules together. Polar solvent acts as a good solvent for electrovalent solute since there is a definite attraction between the ions present and the dipoles of the polar solvent. The electrovalent solute is insoluble in nonpolar solvent since there is no electrical attraction between the ions in the solute and the molecules of the solvent. Polar solvents can dissolve polar and electrovalent solutes and nonpolar solvents can dissolve nonpolar and covalent solutes. In other words, like dissolves like.

B. Effect of Temperature

The effect of temperature to the solubility of a solid in a liquid depends on the heat of solution. The heat of solution of a substance is the thermal change that accompanies the dissolving of a substance in a solvent. If the heat of solution is positive (endothermic), increasing the temperature will increase the solubility. If the heat of solution is negative (exothermic), increasing the temperature will decrease the solubility. In most cases, the solubility of a solid in a liquid increases as the temperature is raised. The effect of temperature to the solubility of a solid in a liquid can be shown in a graph called the solubility curve. Solubility curve is a graph of solubility expressed in grams of solute per 100grams of solvent versus the saturation temperature. The solubility of a gas in a liquid is also dependent on the temperature. It decreases as the temperature is raised.

C. Effect of Pressure

Pressure generally affects the solubility of a gas in a liquid. The effect is governed by Henry's law, which states, " The amount of a gas dissolved by a

given amount of liquid at constant temperature is directly proportional to the pressure of the gas over the liquid."

II. Colligative Properties of a Solution of a Nonvolatile, Nonelectrolyte Solute

Solutions are homogeneous mixtures that contain two or more substances. The major component is called the solvent and the minor component is called the solute. Since the solution is primarily composed of solvent, physical properties of a solution resemble those of the solvent. Some of these physical properties, called colligative properties, are independent of the nature of the solute and depend only upon the solute concentration. The colligative properties include vapor pressure lowering, boiling point elevation, freezing point depression, and osmotic pressure. The following discussion includes the colligative properties of a solution of a nonvolatile, nonelectrolyte solute.

A. Vapor Pressure Lowering

The addition of a nonvolatile solute lowers the vapor pressure of the solvent. This is because the molecules of the solute hold the molecules of the solvent. Hence, the rate of evaporation of the solvent from a solution is less than that from the pure solvent resulting in a decrease in vapor pressure. According to Raoult's law, the lowering of the vapor pressure of a solvent by a nonvolatile solute is directly proportional to the concentration of the solute in the solution. This is expressed mathematically as $\Delta P = P^\circ - P = P^\circ X_2$ where ΔP is the vapor pressure lowering, P° is the vapor pressure of the solvent, P is the vapor pressure of the solution, and X_2 is the mole fraction of the solute. Expressed in terms of the mole fraction of the solvent (X_1), the equation becomes $P = P^\circ X_1$. This is because $X_1 + X_2 = 1$.

B. Boiling Point Elevation

The boiling point of a liquid is the temperature at which its vapor pressure becomes equal to the atmospheric pressure. A solution of a nonvolatile solute has a boiling point that is higher than that of the pure solvent. This is the result of the decrease in the vapor pressure of the pure solvent. Hence, it is necessary to heat the solution to a higher temperature so that the vapor pressure would equal the atmospheric pressure. The boiling point elevation of a solution is directly proportional to the concentration of the solute in the solution, expressed mathematically as $\Delta T_b = K_b m$. ΔT_b is the boiling point elevation which is equal to the boiling point of the solution (BP_{solution}) minus the boiling point of the solvent (BP_{solvent}). K_b is the molal boiling-point-elevation constant which is

dependent on the kind of solvent. If water is the solvent, its K_b is $0.52\text{ }^{\circ}\text{C/molal}$. m is the molality of the solution.

C. Freezing Point Depression

The freezing point of a liquid is the temperature at which its vapor pressure is equal to that of the solid into which it freezes. Since a nonvolatile solute lowers the vapor pressure of a solvent, it follows that the freezing point of the solution would be lower than that of the pure solvent. The freezing point depression of a solution is directly proportional to the concentration of the solute in the solution, expressed mathematically as $\Delta T_f = K_f m$. ΔT_f is the freezing point depression which is equal to the freezing point of the solution (FP_{solution}) minus the freezing point of the solvent (FP_{solvent}). K_f is the molal freezing-point-depression constant which is dependent on the kind of solvent. If water is the solvent, its K_f is $1.86\text{ }^{\circ}\text{C/molal}$. M is the molality of the solution.

D. Osmotic Pressure

Osmosis is the selective passage of solvent molecules through a semipermeable membrane from a dilute solution to a more concentrated solution. The pressure that must be exerted on the more concentrated solution to stop osmosis is called the osmotic pressure. The osmotic pressure of a solution is directly proportional to the concentration of the solute in the solution. The osmotic pressure of a solution is given by $\pi = MRT$ where π is the osmotic pressure of the solution, M is the molarity of the solution, R is the gas constant, and T is the absolute temperature of the solution.

III. Colligative Properties of a Solution of a Nonvolatile, Electrolyte Solute

An electrolyte dissociates into ions in solution. One formula unit of an electrolyte separates into two or more ions when it dissolves. This increase in the number of ionic particles must be taken into consideration since colligative properties depend on the number of particles in a solution. The following equations are applicable for a solution of an electrolyte:

$\Delta P = iP^{\circ}X_2$	(Vapor Pressure Lowering)
$\Delta T_b = iK_b m$	(Boiling Point Elevation)
$\Delta T_f = iK_f m$	(Freezing Point Depression)
$\pi = iMRT$	(Osmotic Pressure)

where i is the van't Hoff factor. The van't Hoff factor is the number of ionic particles produced when a formula unit of the electrolyte is dissolved in the

solution. In reality, the van't Hoff factor is less than that expected from a formula unit because of the electrostatic force between a cation and an anion.

OBJECTIVES

1. To discuss the effects of the various factors that influence the solubility of one substance into another
2. To prepare the solubility curve of potassium chloride in water
3. To compute the molar mass of urea from boiling point elevation
4. To compute the van't Hoff factor of sodium chloride
5. To demonstrate osmotic pressure using raisin and water

LIST OF CHEMICALS

10 drops	ethyl alcohol	0.15 g	naphthalene
3 mL	kerosene	0.15 g	sodium chloride
10 drops	coconut oil	5.40 g	potassium chloride
1 can	carbonated drink	25 g	urea
1 piece	raisin		distilled water
50 mL	4 molal aqueous sodium chloride		

LIST OF APPARATUS

stirring rod, 4 micro test tubes, 600-mL beaker, wire gauze, iron ring, iron stand, iron clamp, Bunsen burner, macro test tube, test tube holder, thermometer, 10-mL measuring pipette, two 150-mL beaker, 50-mL graduated cylinder, round bottom flask, iron clamp, test tube rack, electronic balance, cork, test tube

TO THE PROFESSOR: Tell the students to perform part G of the experiment first. It will require a minimum of 3 hours to see good result.

PROCEDURE

A. Effect of Nature of Solute and Solvent

1. Solubility of a solid in a liquid

SAFETY PRECAUTION: Naphthalene vapor is harmful. Work in a well-ventilated area.

Weigh 0.15 g each of sodium chloride and naphthalene. Place each solid in separate micro test tubes. Add 20 drops of distilled water to each micro test tube. Stir and observe if the solids are soluble in water. Repeat this procedure using kerosene as the solvent.

Waste Disposal: Dispose the mixtures to the sink with plenty of water.

2. Solubility of a liquid in another liquid

SAFETY PRECAUTION: Ethyl alcohol is flammable. Keep it away from the flame.

Place 10 drops each of ethyl alcohol and coconut oil in separate micro test tubes. Add 10 drops of distilled water to each micro test tube. Stir and observe if the liquids are miscible in water. Repeat this procedure using kerosene as the solvent.

Waste Disposal: Dispose the mixtures to the sink with plenty of water.

B. Effect of Temperature on the Solubility of a Solid in a Liquid

1. Place 400 mL of tap water in a 600-mL beaker. Boil the water.
2. Using an electronic balance, weigh accurately 5.40 grams of potassium chloride. Transfer it carefully in a macro test tube.
3. Measure 11 mL of distilled water using a 10-mL measuring pipette. Add this slowly to the macro test tube, washing the potassium chloride adhering to the sides.

Note: Any potassium chloride particle that adheres to the sides of the macro test tube will decrease the amount of the potassium chloride used and therefore, will cause an error.

4. Using an iron clamp to hold the macro test tube, immerse the macro test tube and its contents into the beaker of boiling water. See Figure 1.

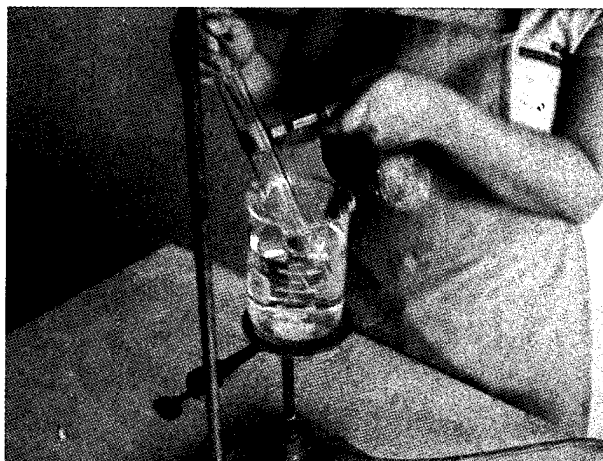


Figure 1. Dissolving the potassium chloride

5. Stir gently the mixture until the solids have been dissolved.

Note: Do not heat the mixture more than necessary because excessive vaporization of water will yield erroneous results. Excessive vaporization will greatly reduce the amount of water used in preparing the solution.

6. After all the solids have been dissolved, remove the macro test tube from the boiling water and allow it to cool in the air while stirring the contents constantly and very gently with a thermometer, observing the solution carefully. When crystals begin to form, note the temperature.

Note 1: Do not cool the solution by immersing the macro test tube in a beaker of water. It will give erroneous results.

Note 2: Stir the solution constantly and gently to avoid formation of a supersaturated solution.

Note 3: Do not remove the thermometer from the solution while reading the temperature.

Note 4: If you have any doubt regarding the appearance of the first crystals, let the solution cool further for verification but record the first temperature noted. If a considerable mass of crystals has been allowed to form before the temperature is noted, the concentration of the solution has changed to such an extent that the observed temperature is not the point of saturation of the original solution. Consequently, you must

observe, as nearly as possible, the temperature at which crystallization just begins.

7. Immerse the macro test tube and its contents in the boiling water once more and again stir the contents gently until the crystallized solid has redissolved. Then remove the macro test tube and repeat the procedure for determining the temperature at which crystals first form. If there is a large difference from the first result (greater than $\pm 2^{\circ}\text{C}$), repeat the procedure once more.
8. After a satisfactory value has been obtained, dilute the original solution by adding exactly 1 mL of distilled water using the measuring pipette and observe, in the same manner as before, the temperature at which the new solution is saturated. Repeat this procedure until the saturation temperatures of four different concentrations have been observed.
9. **Waste Disposal:** Dispose the solution to the sink.
10. Calculate for each solution the solubility of potassium chloride in water expressed in grams of KCl per 100 grams of water. The density of water is 1 g/mL.
11. Plot a graph with average saturation temperature on the horizontal axis and with solubility on the vertical axis.

C. Effect of Temperature on the Solubility of a Gas in a Liquid

Place 100 mL of tap water in a 150-mL beaker. Warm gently the water. Observe the formation of bubbles.

Waste Disposal: Dispose the water to the sink.

D. Effect of Pressure on the Solubility of a Gas in a Liquid

Open a can of carbonated drink. Observe the formation of bubbles.

E. Molar Mass Determination from Boiling Point Elevation

1. Set-up the apparatus as shown in Figure 2.

SAFETY PRECAUTION: Use boiling chips to avoid sudden boiling of the water in the round bottom flask. There must be a notch in the cork to release the generated steam.

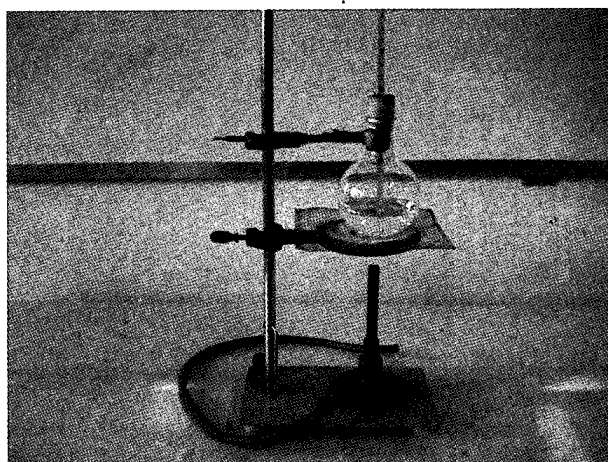


Figure 2. Determining the boiling point of water

2. Using a 50-mL graduated cylinder, measure 50 mL of distilled water.
3. Pour the distilled water in the round bottom flask.
4. Make sure that the thermometer bulb is immersed at the center of the distilled water.
5. Boil the distilled water.
6. While boiling, determine the temperature of the distilled water. This temperature is the boiling point of water.
7. **Waste Disposal:** Dispose the hot distilled water to the sink.
8. Using an electronic balance, weigh accurately 25 grams of urea. Place the urea in a 150-mL beaker.
9. Measure 50 mL of distilled water using a 50-mL graduated cylinder and pour the distilled water in the beaker containing the urea.
10. Stir the mixture to dissolve the urea.
11. Set-up again the apparatus as shown in Figure 2.
12. Transfer the urea-distilled water solution into the round bottom flask. The thermometer bulb must be immersed at the center of the solution.
13. Heat the solution to boiling and get the boiling point of the solution.
14. **Waste Disposal:** Dispose the solution to the sink.

15. From the results obtained, compute the molar mass of urea as follows:

- a. Calculate the molality of the solution by using $BP_{\text{solution}} - BP_{\text{solvent}} = K_b m$.
 K_b of water is 0.52°C/molal .
- b. Determine the number of moles of urea. The density of water is 1 g/mL.
 $\text{No. of moles of urea} = (\text{mass of water in kg}) \times (\text{molality of the solution})$
- c. Compute the molar mass of urea.

$$\text{Molar mass of urea} = (\text{mass of urea used}) / (\text{no. of moles of urea})$$

16. Calculate the percent error. This is to be done by the instructor.

F. Determination of the van't Hoff Factor for Sodium Chloride

1. Set-up the apparatus as shown in Figure 2.
2. Place 50 mL of a 4 molal aqueous sodium chloride in the round bottom flask. The thermometer bulb must be at the center of the solution.
3. Boil the solution and determine its boiling point.
4. **Waste Disposal:** Discard the solution to the sink.
5. Calculate the van't Hoff factor (i) for sodium chloride using the equation $BP_{\text{solution}} - BP_{\text{water}} = iK_b m$. Use the boiling point of water obtained in part E.

G. Osmotic Pressure

1. Place 10 mL of distilled water in a test tube.
2. Get a piece of raisin and observe its size.
3. Drop the raisin into the distilled water.
4. Set aside for a minimum time of 3 hours.
5. Observe the change in size of the raisin.
6. **Waste Disposal:** Dispose the water to the sink. Place the raisin in the trash can.

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Group No. _____

Name:	Date Performed:
Course & Section:	Date Submitted:
Program & Year:	Professor:

Experiment 2 SOLUTIONS AND THEIR PROPERTIES

A. Effect of Nature of Solute and Solvent

1. Solubility of a solid in a liquid

Write in the table provided whether the solids are soluble or insoluble in the given solvents.

Solute	Water	Kerosene
Sodium Chloride		
Naphthalene		

Tell whether each solid solute is polar or nonpolar.

2. Solubility of a liquid in another liquid

Write in the table provided whether the liquids are miscible or immiscible in the given solvents.

Solute	Water	Kerosene
Ethyl Alcohol		
Coconut Oil		

Tell whether each liquid solute is polar or nonpolar.

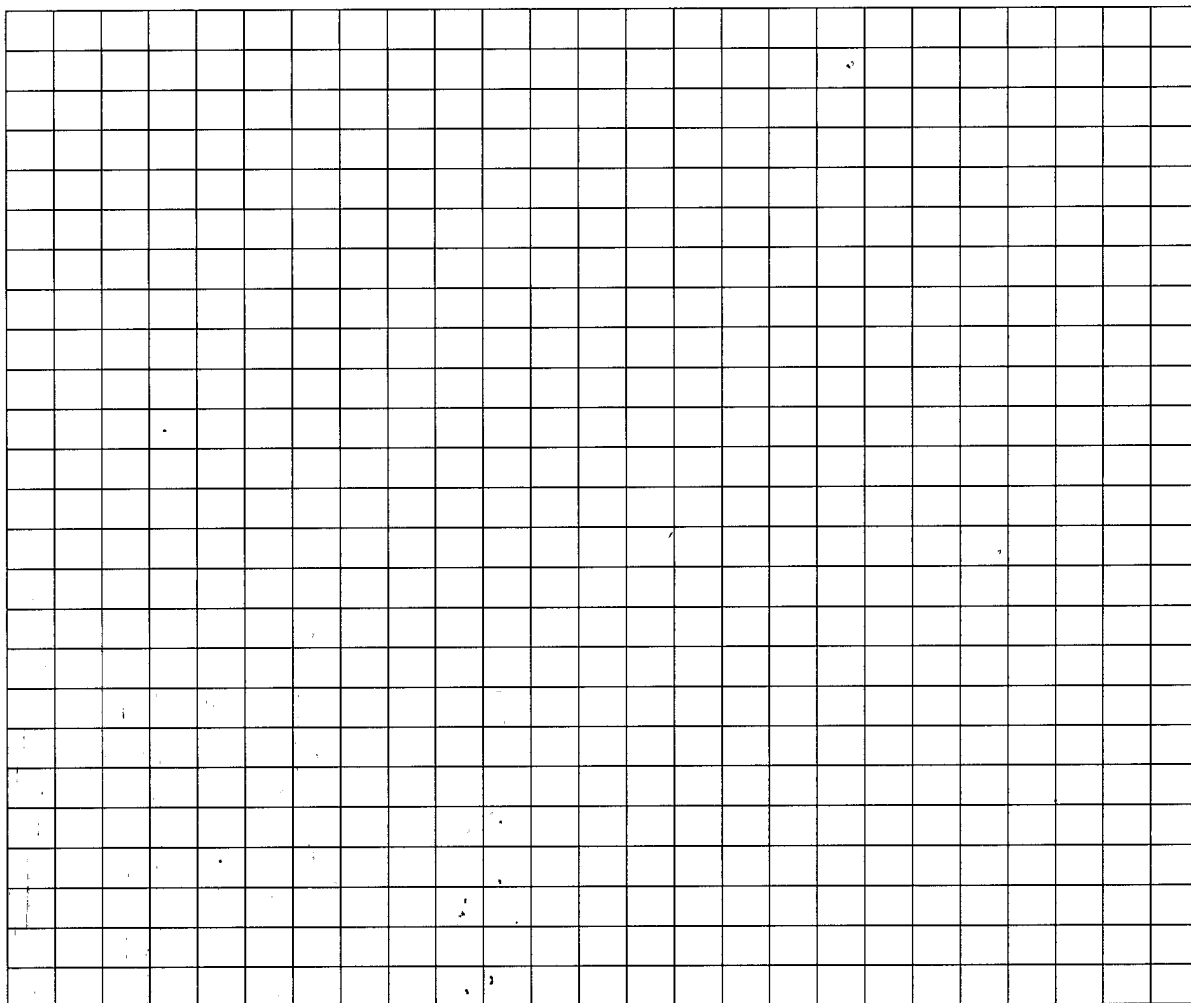
Based on the results, make a general correlation between solubility and nature of solute and solvent.

B. Effect of Temperature on the Solubility of a Solid in a Liquid

Solution No.	Weight of KCl (grams)	Volume of Water (mL)	Solubility in g KCl per 100 g Water	Saturation Temperature		
				Trial 1	Trial 2	Average
1						
2						
3						
4						

Show the computation for expressing the solubility in grams KCl per 100grams of water for solution1.

Plot the solubility in grams of potassium chloride per 100 grams of water versus the average saturation temperature in degrees Celsius.



Based on the trend of the solubility curve, what is the effect of temperature to the solubility of potassium chloride in water?

If you heat the solutions excessively, what error might you encounter in determining the solubility at different temperatures?

Why is it necessary to stir the solution constantly while cooling?

C. Effect of Temperature on the Solubility of a Gas in a Liquid

What were formed when the water was gently heated?

The appearance of bubbles indicates the evolution of a gas. What are expected to be the gases in the tap water?

What is the effect of temperature to the solubility of a gas in a liquid?

D. Effect of Pressure to the Solubility of a Gas in a Liquid

What were produced when a can of carbonated drink was opened?

Explain why bubbles appeared after opening a can of carbonated drink.

What is the effect of pressure to the solubility of a gas in a liquid?

E. Molar Mass Determination from Boiling Point Elevation

Mass of Urea (grams)	
Volume of Water (mL)	
Density of Water (g/mL)	
Mass of Water (grams)	
Boiling Point of Water ($^{\circ}\text{C}$)	
Boiling Point of Solution ($^{\circ}\text{C}$)	
K_b of Water ($^{\circ}\text{C/molal}$)	
Molality of Solution (moles urea / Kg water)	
No. of Moles of Urea	
Experimental Value of Molar Mass of Urea (g/mole)	
True Value of Molar Mass of Urea (g/mole)	
Percent error	

Show the computations for the following:

1. Molality of Solution
2. Number of Moles of Urea
3. Molar Mass of Urea

What are the possible sources of errors in this experiment?

F. Determination of the van't Hoff Factor for Sodium Chloride

Volume of Aqueous Sodium Chloride (mL)	
Molality of Aqueous Sodium Chloride (moles NaCl/Kg water)	
Boiling Point of Water ($^{\circ}\text{C}$)	
Boiling Point of Aqueous Sodium Chloride ($^{\circ}\text{C}$)	
K_b of Water ($^{\circ}\text{C/molal}$)	
Experimental van't Hoff Factor	
Theoretical van't Hoff Factor (from general chemistry book)	

Compute the experimental van't Hoff factor of sodium chloride.

Compare the experimental van't Hoff factor with the theoretical value. Account for any discrepancy in values.

G. Osmotic Pressure

What happened to the size of the raisin after being soaked in distilled water?

Explain the result of the experiment.