

Experiment: Colligative Properties, Do it at Home Lab

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Objective: The objective of this experiment is to determine the freezing point depression and boiling point elevation of water by addition of sodium chloride, and to explore the osmosis process pass through the membrane of deshelled un-cooked eggs.

Introduction: The addition of salt to water lowers the freezing point of water. This is the mechanism by which roads are cleared of ice in the winter in snowy climates. This is also the same trick used in making home-made ice cream since ice cream solidifies below 0°C. Many salts and solutes will change the properties of the solvent in which they are dissolved. This phenomenon is called colligative properties of a solution. Most often, colligative properties take the form of lowering the freezing point of the solvent or raising the boiling point of the solvent. One of the features of colligative properties is that the change in freezing/boiling points is often independent of the exact solute but rather the moles of the solute.

In this experiment, we will determine the freezing point of water without and with different amount of salt. The objective is to demonstrate the freezing point depression of water with the addition of known amounts of salt. The second part of the experiment is to observe the osmosis process of eggs. The size of an egg depends on the solution it is placed in. If eggs are placed in a hypertonic solution, they will shrink in size. If eggs are placed in a hypotonic solution, they will expand in size.

The decrease in freezing point, ΔT_f (freezing point depression) for a near ideal solution can be described by the equation:

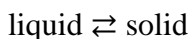
$$\Delta T_f = i K_f \cdot m$$

Where i is the van't Hoff factor, which is the number of particles a substance becomes when it dissolves. We use the predicted i , which is based on the composition of the strong electrolytes. For example, the predicted i for NaCl is 2, whereas the predicted i for CaCl₂ will be 3, because NaCl dissociates into two ions, whereas CaCl₂ dissociates into three ions. The van't Hoff factor, i , may be less than one if the solute particle aggregate to form complex as some proteins do.

K_f is the molal freezing point depression constant of the solvent with units °C · kg solvent/mole solute. m is the molal concentration of the solute dissolved in the solvent expressed as moles of solute/kg solvent.

$$T_f = \text{freezing point of pure solvent} - \Delta T_f$$

At the freezing point of any substance, an equilibrium exists in which both liquid and solid are present.



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 start with their sample, pure solvent or solution, well above its freezing 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 equilibrium, the temperature will be ideally constant until all the sample has solidified. A graph is made by plotting the temperature vs. time. In this experiment, 5 consecutive readings indicate the freezing point for the solution.

Similarly, the increase in boiling point, ΔT_b (boiling point elevation) for a near ideal solution can be described by the equation:

$$\Delta T_b = i K_b \cdot m$$

Where i is the van't Hoff factor, K_b is the molal boiling point elevation 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.

$$\text{The boiling point of the solution } T_b = \text{boiling point of pure solvent} + \Delta T_b$$

Some substances form **semipermeable membranes**, allowing some smaller particles to pass through but blocking larger particles. The net movement of solvent molecules from solution of low to high concentration of solute across a semipermeable membrane is **osmosis**. The applied pressure to stop it is **osmotic pressure, Π**

$$\Pi = iMRT$$

Where i is the van't Hoff factor, M is the molarity of the solute substance. R is the gas law constant ($\frac{0.08206 \text{ atm}\cdot\text{L}}{\text{mol}\cdot\text{K}}$), T is temperature in Kelvin. Osmotic pressure is a colligative property.

If two solutions separated by a semipermeable membrane have the same osmotic pressure, no osmosis will occur. Isotonic solutions have same osmotic pressure; solvent passes the membrane at the same rate both ways. Hypotonic solutions have lower osmotic pressure; solvent will leave this solution at a higher rate than it enters with. Solute concentration is low in hypotonic solutions. Hypertonic solutions have higher osmotic pressure; solvent will enter this solution at a higher rate than it leaves with. Solute concentration is high in hypertonic solutions.

Red blood cells have semipermeable membranes. If stored in a hypertonic solution, they will shrivel as water leaves the cell; this is called **crenation**. If stored in a hypotonic solution, they will grow until they burst; this is called **hemolysis**. Intravenous (IV) solutions **MUST** be isotonic to blood.

Procedures:

Part-A Freezing point depression:

1. Place about 1 cup of ice in a cup (coffee cup or some other plastic cup).
2. Put ~200 mL of tap water in the cup with the ice.
3. Start your timer.
4. Record the temperature for every 30 seconds in **Data Table 1**.
5. Stir the solution continuously to keep the temperature even throughout the cup.
6. When measuring the temperature, keep the thermometer **near the ice** in the cup, remembering that ice floats in water.
7. Record the temperature at each time point (every 30 seconds) in **Data Table 1**.
8. After recording is completed (when the readings remain consistent for at least 5 consecutive readings), empty the water and ice into a sink.
9. Dry the cup completely. Repeat step 1 and 2.
10. Pour 1.0 teaspoon (tsp) of table salt into the cup over the ice and stir.
11. Repeat step 3 to 9.
12. Rinse the cup with tap water three times, pouring out the rinse water each time.
13. Dry the cup completely. Repeat step 1 and 2.
14. Pour 2.0 teaspoon (tsp) of table salt into the cup over the ice and stir.
15. Repeat step 3 to 8.
16. Make a graph of your data for all three columns of **Data Table 1**. Submit your graph with your lab report.
 - a. Plot temperature on the y-axis and time on the x-axis.
 - b. The 5 consecutive readings indicate the freezing point for the solution

Part-B Boiling Point Elevation:

1. Pour 1000 mL of water into a small cooking pot.
2. Turn on you stove.
3. When the water is boiling, measure the temperature of the boiling water by hanging the thermometer in the middle of the water for ~2 minutes., or until the reading of the thermometer is steady. Make sure the thermometer does not touch the bottom or side of the pot. You may want to wear an oven glove if you hold the thermometer in your hand when measuring the temperature.
4. Record the boiling point in data **table 2**.
5. Let the water cool to none boiling.
6. Add 5.0 tsp of table salt into the pot and stir to make sure the salt dissolves.
7. Repeat step 2, 3, and 4.
8. Let the solution cool to none boiling.
9. Add 5.0 more tsp of table salt into the pot and stir to make sure the salt dissolves. Together with the 5.0 tsp salt from step-6, there are 10.0 tsp salt in the solution now.

10. Turn on you stove.
11. Repeat step 2, 3, and 4.
12. Discard the salt-water in the sink. Be careful with the hot boiling water!

Part-C Osmosis:

1. Place two similar size of eggs into two cups, one egg in one cup.
2. Pour enough white vinegar with ~5.0% of acetic acid into the cup until the eggs are submerged in the vinegar.
3. Cover the cup with plastic films and let it stand for ~48 hours.
4. After ~ 48 hours, wash off the remaining egg shell with hand in the running tap water.
5. Place both de-shelled eggs side-by-side on a plate and measure the length and width of them. Record the data in **data table 3**. Take a picture of both eggs together and submit with your lab report.
6. Place one egg in a cup with corn syrup, make sure there is enough syrup so the egg is fully submerged, may be with some pushing on the egg. If the egg float, it will sink as time goes on.
7. Place the other egg in tap water, make sure there is enough water so the egg is fully submerged, may be with pushing into it.
8. Cover the cups with plastic films and let them stand for ~72 hours.
9. Take the eggs out of the cup and measure the length and width of the eggs. Record the data in **data table 3**. Take a picture of both eggs together and submit with your lab report.

Data Table-1 Temperature and time. (6 points)

Time (second)	Temp (°C) of Tap water	Temp (°C) of Solution (1.0 tsp NaCl)	Temp (°C) of Solution (2.0 tsp NaCl)
30			
60			
90			
120			
150			
180			
210			
240			
270			
300			
330			
360			
390			
420			
450			
480			

Table-2 Boling point (3 points)

	Tap Water	Solution with 5.0 tsp of NaCl	Solution with 10.0 tsp of NaCl
Boling Point (°C))			

Post Lab Questions

1. Suppose your thermometer consistently read a temperature 1.2° lower than the correct temperature throughout the experiment. How would this have affected the molar mass you found? (1 pt)
2. If the freezing point of the solution had been incorrectly read 0.3° lower than the true freezing point, would the calculated molar mass of the solute be too high or too low? Explain your answer. (1 pt)
3. Arrange the aqueous solutions in order of increasing freezing points (lowest to highest temperature): 0.10M glucose, 0.10M BaCl_2 , 0.20M NaCl, and 0.20M Na_2SO_4 . (1 pt)
4. What mass of NaCl is dissolved in 150g of water in a 0.050*m* solution? *m* (molality units are moles of solute per Kg solvent) (1 pt)

5. A molar mass determination experiment was conducted. The unknown salt dissociates to form two ions. The data was (1 pt)

Mass of unknown salt = 5.021 g

Initial ice + water temperature = 0.20°C

Final melted ice + water + salt temperature = -0.40°C

The initial weight of the empty beaker = 108.5 g

The final weight of the beaker + water + salt + melted ice = 298.0 g

Assume for water, $K_f = 1.86 \text{ }^\circ\text{C/m}$, and all the ice melted.

What is the molar mass of the unknown salt? Show your calculations to receive credit.

6. Proteins frequently form complexes in which 2, 3, 4, or even more individual proteins (“monomers”) interact specifically with each other via hydrogen bonds or electrostatic interactions. The entire assembly of proteins can act as one unit in solution, and this assembly is called the “quaternary structure” of the protein. Suppose you discover a new protein whose monomer molar mass is 25,000 g/mol. You measure an osmotic pressure of 0.183 atm at 37 °C for 7.20 g of the protein in 10.00 mL of an aqueous solution. How many protein monomers form the quaternary protein structure in solution? Treat the protein as a nonelectrolyte. (2 point)

_____ of 7 pts

Total for lab _____ of 25 pt