Name ______________________ Date __________ Class __________

Boiling Points of Solutions — *Lab* 44

Introduction

You may know that the boiling and freezing points of a solution are different from the boiling and freezing points of a pure solvent. Adding a solute to a solvent elevates the boiling point and depresses the freezing point. There are many everyday examples of this chemical principle. Antifreeze is added to the water in automobile radiators to prevent both boiling over in hot weather and freezing in cold weather. Saltwater harbors may freeze during prolonged cold spells, but only when the temperature of the salt water is much lower than 0°C. Adding table salt to cooking water is probably the most familiar example of a solute elevating the boiling point of a solution.

When antifreeze is added to a car radiator there are well-defined mixing directions to be followed. As with all other solutions, the molality of the antifreeze-water mixture determines the temperature at which the solution will boil. When one mole of particles is dissolved in one kilogram of solvent (a one molal solution), the boiling point is elevated by a precise amount depending on the solvent. If the solvent is water, the increase will be $K_b = 0.52°C$ for a one molal solution. (This figure is a constant.) For nondissociating solutes, there is no change in the number of particles as the solute dissolves, so the temperature change for the solution is determined by multiplying the molality by the solvent constant, according to the following formula:

$$\Delta T_b = K_b m$$

As ionic solids dissolve, they dissociate, resulting in at least two moles of particles in solution for each mole of crystalline solute. The molal concentration of the ions in solution increases—doubling, tripling, and so on. Accordingly, the change in temperature for a solvent's boiling point is two, three, or more times the value of the solvent constant.

In this investigation, you will find the boiling points of distilled water, a urea-water solution, and a salt-water solution. Temperature readings will be taken for the samples as each is heated from room temperature to the boiling point. The time and temperature data will be plotted on a single graph. From the graph you can compare the boiling points of the different samples. Where the boiling points of the samples differ, you can suggest reasons for the differences based on an examination of the chemical formulas of the solutes.

Pre-Lab Discussion

Read the entire laboratory investigation and the relevant pages of your textbook. Then answer the questions that follow.

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1. What is the molality of a solution if 0.0300 moles of nonionizing solute dissolve in 15.0 mL of water?

2. What is the concentration of the particles in solution in Question 1?

3. Calculate the concentration of particles in solution if 0.0300 moles of an ionic solid such as $CaCl_2$ dissolve in 15.0 mL of water.

4. How many grams of urea (NH_2CONH_2) equal 0.0300 moles of urea?

5. How many grams of sodium chloride (NaCl) equal 0.0300 moles of sodium chloride?

6. How should the test tube be heated so that the temperature increases continuously but not rapidly?

7. What precautions should you take when using the burner?

Problem

What effect do solutes have on the boiling point of a solvent?

Materials

chemical splash goggles
laboratory apron
graduated cylinder, 100-mL
distilled water
large test tube
boiling chips
thermometer
#2 cork stopper, cutaway
ring stand
ring
test-tube clamp
wire gauze square
lab burner
watch or clock with a second hand
urea (NH_2CONH_2)
laboratory balance
stirring rod
sodium chloride (NaCl)

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Safety

Wear your goggles and lab apron at all times during the investigation. Tie back loose hair and clothing to avoid any fire hazard. Always point the open end of a test tube away from yourself and others when heating a substance. Since glass retains heat without looking hot, heated glassware should be given ample time to cool before it is handled.

Note the caution alert symbols here and with certain steps of the Procedure. Refer to page *xi* for the specific precautions associated with each symbol.

Procedure

1. Put on your goggles and lab apron.
2. Precisely measure out 15.0 mL of distilled water and pour it into the test tube. Put two or three boiling chips into the test tube.
3. Carefully insert the thermometer into the cutaway cork stopper, positioning it so the calibrations are fully visible. Place the thermometer in the test tube so that the bulb of the thermometer is about 1 cm above the bottom of the test tube.
4. Set up the ring stand and attach the ring and test-tube clamp. Place the wire gauze on the ring. Adjust the height of the ring so that the burner will fit underneath it. Put the test tube into the test-tube clamp and position it so that the bottom of the test tube just rests on the wire gauze. Rotate the clamp and test tube away from the ring.
5. Light the burner and adjust the flame to a medium setting. **CAUTION:** *Tie back loose hair and clothing to avoid any fire hazard.* Move the burner underneath the wire gauze and adjust the height of the ring so the inner cone of the burner flame just hits the gauze. Readjust the position of the test tube if necessary. Preheat the wire gauze for a minute or two without the test tube in place.
6. Find the initial temperature of the water and record it in the Data Table. Turn the clamp back so the test tube again rests on the wire gauze. The test tube should be slightly off to the side rather than directly over the flame. Record time and temperature readings at 30-second intervals. Continue until the temperature remains constant for three or four readings. If necessary, change the position of the burner so the temperature of the water increases continuously but not rapidly. **CAUTION:** *Always point the open end of a test tube away from yourself and others when heating a substance.*
7. Turn off the burner. Remove the thermometer, rinse it, and set it aside for the next trial. Keeping the hot test tube in the clamp, carefully remove the clamp and test tube from the ring stand. Pour the water in the sink and rinse out the test tube. Discard the boiling chips. **CAUTION:** *Since glass retains heat without looking hot, heated glassware should be given ample time to cool before it is handled.*

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8. Using the calculations from Pre-Lab Discussion Question 4, obtain 0.0300 moles of urea and record its mass to the nearest 0.01 g. Place the urea and 15.0 mL of distilled water into the test tube. Stir gently with the stirring rod until the urea dissolves. Put two or three boiling chips into the solution.
9. Repeat Steps 3–7 for the urea-water solution.
10. Using the calculations from Pre-Lab Discussion Question 5, obtain 0.0300 moles of sodium chloride and record its mass. Place the sodium chloride and 15.0 mL of distilled water into the test tube. Stir gently with the stirring rod until the sodium chloride dissolves. Put two or three boiling chips into the solution.
11. Repeat Steps 3–7 for the sodium chloride–water solution.
12. Clean up your work area and wash your hands before leaving the laboratory.

Observations

DATA TABLE

	Temperature (°C)		
Time (min)	**Distilled Water**	**Urea-Water**	**Salt-Water**
0.0			
0.5			
1.0			
1.5			
2.0			
2.5			
3.0			
3.5			
4.0			
4.5			
5.0			
5.5			
6.0			
6.5			

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Calculations

1. Predict the new boiling point for the urea-water solution.

2. Predict the new boiling point for the sodium chloride–water solution.

Critical Thinking: Analysis and Conclusions

1. Construct a boiling-point graph using only the data collected for temperatures above 80°C, plotting temperature as a function of time. Plot the data for all three samples—the distilled water, the urea-water solution, and the sodium chloride–water solution—on the same graph. Use different symbols or colors to differentiate among the data points for each of the three samples. Draw best-fit curves for all three sets of data points. *(Interpreting data)*

2. Examine the graph. Do you see any pattern in the boiling points of the two solutions compared to the boiling point of the distilled water? *(Making comparisons)*

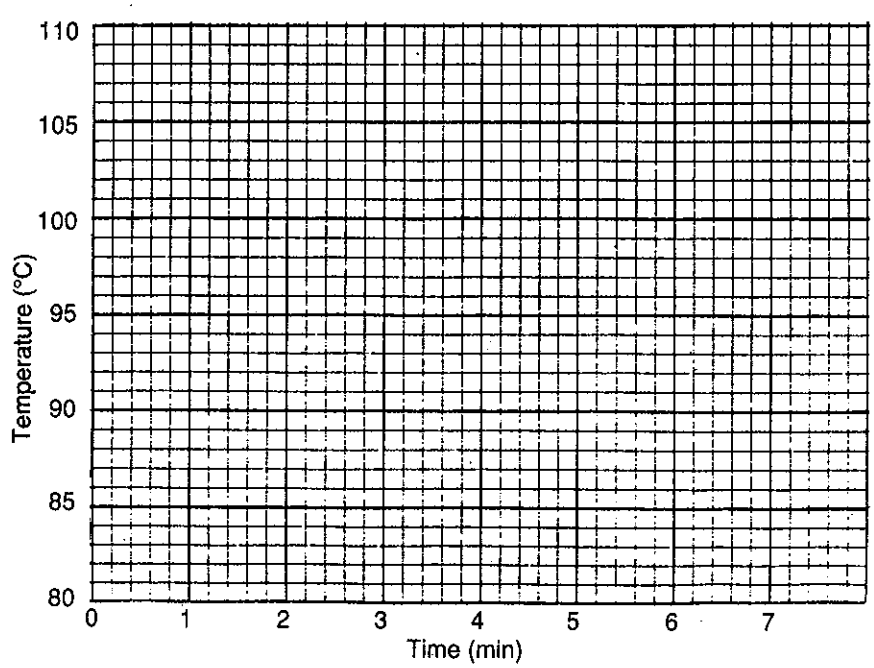

Figure 44–1

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3. Examine the formulas for the two solutes, urea (NH_2CONH_2) and sodium chloride ($NaCl$). What inferences can you make about the effects of each of these two solutes on the boiling point of water? *(Making inferences)*

4. Compare the boiling points of the two solutions from your graph to the predicted values from Calculations 1 and 2. Do the experimental and predicted values match? If they differ, how might you account for the differences? *(Making comparisons)*

Critical Thinking: Applications

1. What do you think would happen to the temperature of either solution if you kept heating it? Explain. *(Making predictions)*

2. In terms of particles, explain why the boiling points of the solutions are higher than the boiling point of distilled water. *(Developing models)*