

TOPIC 4:

SOLUTIONS

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Appendix 4.1: Polar and Non-polar Substances

Purpose

In this lab, you will discover which substances, when mixed, will form solutions. You will also create a general rule about solutions created from polar and non-polar substances.

Materials

- goggles
- test tubes
- test tube rack
- rubber stoppers
- graduated cylinder (10 mL)
- paper towels
- scoopulas
- Material Safety Data Sheets (MSDS)

Substances

- *Polar Substances*
 - copper(II) sulphate crystals, CuSO_4
 - water, H_2O
 - vinegar
- *Non-polar Substances*
 - solid iodine crystals
 - vegetable oil
 - kerosene

Procedure

1. Using the following chart as a guide, mix all possible combinations of the substances above. The teacher will show you how much solid to add.
2. If the substance is a liquid, then use 10 mL.
3. If the solution is a solid to be mixed with a liquid, then add the solid first to a clean, dry test tube.
4. Once the substances are added to the test tubes they should be swirled to mix.
5. Dispose of the solutions as directed by the teacher.

Appendix 4.1: Polar and Non-polar Substances (continued)

Observations

Substance	Copper(II) Sulphate	Water	Vinegar	Solid Iodine	Vegetable Oil	Kerosene
Copper(II) Sulphate	X	X	X	X	X	X
Water		X	X	X	X	X
Vinegar			X	X	X	X
Solid Iodine				X	X	X
Vegetable Oil					X	X
Kerosene						X

Discussion Questions

1. What did you observe regarding the mixing of polar and non-polar substances?
2. Based on the general rule about solutions created from polar and non-polar substances, determine whether the following substances are polar or non-polar.
 - a) cobalt(II) chloride crystals, which are soluble in water
 - b) ammonia gas, which is soluble in water
 - c) trichlorotrifluoroethane (TTE) liquid, which is not miscible in water
3. Write a general rule that can be applied to polar and non-polar substances when they are mixed.

Appendix 4.1: Polar and Non-polar Substances (Teacher Notes)

Purpose

The purpose of this lab is to have students to discover the relationship between polar and non-polar substances when mixed (i.e., “like dissolves like”).

Procedure and Precautions

Review the current Material Safety Data Sheet (MSDS) for each of the substances with students and have the data sheets available during the lab in case of an accident. First aid treatment instructions are provided on each MSDS.

The approximate amount of solid iodine required for the lab should be placed in a beaker that is, in turn, placed on a clearly labelled paper towel in a fume hood. Any solid iodine left over after the experiment should be disposed of using Workplace Hazardous Materials Information System (WHMIS) procedures.

By placing a few crystals of each solid on a watch glass, students will understand how much to use for each solution. Have students transfer the crystals directly into their test tubes at the dispensing site.

Samples of the other substance should also be placed on clearly labelled paper towels. If an automatic dispensing pipette is available, the liquids should be dispensed in this manner with the set screw adjusted to 10 mL. Otherwise, a clean 10 mL graduated cylinder should be available for each liquid to be measured by students.

Note: Placing food dye into the liquid mixtures will identify the polar solvents.

Answer Key

1. Student observations:
 - Typical observations for the immiscible liquid combinations will be:
 - They do not mix.
 - They separate into layers.
 - They mix when shaken but gradually separate out into distinct layers.
 - Iodine should not dissolve and go into solution with water, as iodine and water are “unlike” components.
 - The same is true for kerosene and copper sulphate.
2. a) Cobalt(II) chloride must be polar, as it dissolves in polar water.
b) Ammonia gas must be polar, as it dissolves in water.
c) TTE must be non-polar, as it does not dissolve in polar water.

Appendix 4.1: Polar and Non-polar Substances (Teacher Notes) *(continued)*

3. Possible statements:

- If *both* substances are polar or non-polar, then they will dissolve or be miscible.
- If a polar substance is mixed with a non-polar substance, they will not dissolve. They will be immiscible.
- Like substance will dissolve like substance.

Appendix 4.2: Why Don't Water and Oil Mix? (Demonstrations—Teacher Notes)

Use the following demonstrations to help students develop an understanding of the molecular interactions between polar and non-polar molecules.

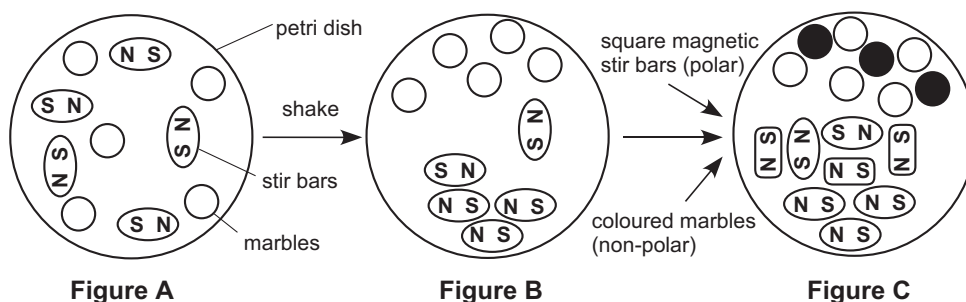
Demonstration A

- Cut ovals (2 cm x 3 cm) and rectangles (2 cm x 3 cm) from an acetate sheet.
 - Using black markers, label one end of the oval positive (+), and the other end negative (-). The ovals represent polar water molecules.
 - Using a different colour of marker, label the rectangles as neutral. The rectangles represent oil molecules.
- Place the ovals and rectangles on the overhead projector and ask students to predict what will happen. Students will likely predict that the negative end of one oval will attract the positive end of another oval. Matching the ovals in this way results in the expulsion of the rectangles, producing two layers – one layer of polar molecules and a second layer of non-polar molecules.
- Relate this demonstration to the attraction of water, a polar substance, to an electrostatically charged comb.
- Discuss hydrogen bonding and dipole-dipole interactions.

Demonstration B

- Place the magnetic stir bars and clear marbles into a petri dish (see Figure A).
 - The stir bars represent polar molecules.
 - The marbles represent non-polar molecules.

Ask students to predict what will happen when the dish is agitated. Shake the petri dish (see Figure B). This motion results in a separation of the two phases (i.e., the magnets attract each other and exclude the marbles).
- Add different shaped stir bars (another polar substance) and coloured marbles (another non-polar substance) to the petri dish. Ask students to predict what will happen when the dish is agitated. Shake the petri dish (see Figure C). Again, the polar molecules will be attracted to each other and exclude the non-polar molecules, thus creating two distinct regions.
- This demonstration illustrates the concept of “like dissolves like.”



Appendix 4.3: Constructing a Solubility Curve

Pre-Lab Discussion

The amount of solute required to achieve a saturated solution in water depends upon the temperature. Most solid solutes increase in solubility as temperature increases in a liquid solution. A solubility curve can be created by varying the amount of solute, keeping the amount of solvent constant, and determining at what temperature all the solute dissolves (or conversely, we can see when the solute particles start to reappear as we slowly cool the solution).

In this experiment, student groups will be assigned a specific amount of solute to dissolve. There are two procedures that will introduce a significant error into the results: the measurement of both the solid and the temperature when precipitation occurs. The temperature can be checked, whereas the solid measurement can be done only ONCE. After the graphs have been drawn from the class data, it will be very obvious whose data are incorrect.

Problem

At what temperature does the given amount of solute precipitate?

Materials

- ammonium chloride, NH_4Cl
- distilled water
- graduated cylinder
- balance
- filter paper
- scoopula
- Bunsen burner
- matches
- 2 thermometers
- thermometer clamps
- 30 cm length of 18-gauge copper wire
(Bend the wire into a large loop to go around the thermometer to stir the solution. The end of the wire should extend well above the top of the test tube.)
- test tubes (20 x 200 mL)
- beaker (400 mL)
- tap water

Procedure

1. Mass out exactly _____ g of NH_4Cl .
2. Measure out exactly 20 mL of distilled water in a graduated cylinder.

Appendix 4.3: Constructing a Solubility Curve (*continued*)

- Place both the water and the solute into a large test tube and stir with the copper wire instrument you created. Place one thermometer into the test tube and a second thermometer into a water bath. Wherever possible, use a specialized thermometer clamp to hold the thermometer off the bottom of the test tube.
- Use a hot water bath to heat the solution about 10°C higher than the point at which the solute looks entirely dissolved.
- Allow the solution to cool, with constant stirring, and record the temperature when the solute begins to precipitate out. The precipitate may be very obvious, or the solution may just begin to look cloudy.
- Repeat steps 4 and 5 two more times. Remember, you should heat solutions to about 10°C higher than the temperature you will actually record each time. Take the temperature after a cooling period has elapsed and a precipitate just becomes visible.
- Record the data in the table provided.
- Examine the temperature data. Discard unreasonable data, and perform steps 4 and 5 until three sets of comparable data are achieved. Average the three best temperatures used.

Observations

Individual Data	
Trial #	Temperature (°C)
#1	
#2	
#3	
#4	
#5	
Average	

Group Data						
Solubility (g/20 g of Water)						
Solubility (g/100 g of Water)						
Temperature (°C)						

Appendix 4.3: Constructing a Solubility Curve (*continued*)

Questions

1. Why is it important to measure out exactly the given amount of solute and solvent?
2. Use the class data to construct a graph with temperature on the horizontal axis and solubility (g of solute/100 g of water) on the vertical axis.
3. Does the class data produce a smooth-flowing curve? Are there any points that are not on the curve? Explain these anomalies.
4. What is the relationship between temperature and solubility on this graph?
5. What does a point on the line represent?
6. Determine the solubility of ammonium chloride at 10°C and at 90°C. Explain how these values were determined.

Appendix 4.3: Constructing a Solubility Curve (Teacher Notes)

Having students collect data for all the points on the solubility curve is unnecessary. For this experiment it is suggested that student groups be assigned a specific amount of solute to dissolve. Students can carefully collect data for one of the points on the curve and still develop the necessary lab skills. Another option would be to have student groups collect data for more than one point on the curve.

Teachers have the option of providing students with the following information to encourage them to produce their best work and to strive for accuracy and precision. There are, of course, two procedures that will introduce a significant error into the results: the measurement of the solid and the temperature when precipitation occurs. The temperature can be checked and averaged, whereas the solid measurement can be done only ONCE. After the graphs have been drawn from the class data, it will be very obvious whose data are incorrect.

Below are the recommended amounts of solute to assign to lab groups to produce the best solubility curve. If there are more than six groups, then multiples of one mass of solute will provide a greater degree of precision. Note that three significant figures are suggested.

Recommended Amounts of Solute to Dissolve		
Lab Group	Grams of NH_4Cl / 20 g of Water	Temperature (°C)
#1	8.00 g	~28°C
#2	9.00 g	~39°C
#3	10.00 g	~50°C
#4	11.00 g	~60°C
#5	12.00 g	~70°C
#6	13.00 g	~78°C

The solubility curve for ammonium chloride, NH_4Cl , can be found in most chemistry texts.

Extensions

1. Have students discuss the instances of experimental errors.
2. Students could use Vernier's *Graphical Analysis*[®] software (provided under a provincial site licensing agreement) to graph the curve and determine a mathematical relationship for solubility and temperature.
3. Have one group conduct several trials when stirring and when not stirring the solution.

Appendix 4.3: Constructing a Solubility Curve (Teacher Notes) *(continued)*

Answers to Questions and Discussion

1. If an incorrect amount of solute is measured out, the temperature results will not produce a point that will fit a smooth-flowing curve.
2. Students draw the graph either manually or by using plotting software.
3. Answers will vary. Usually, several points are not close to the best-fit curve due to carelessness. Students will usually say that either the solid was not massed accurately or the temperature was incorrect. If students do not stir continuously, super cooling will occur and the solid will not precipitate until the solution reaches a lower than expected temperature. (See Extensions.)
4. The solubility of ammonium chloride is directly proportional to the temperature of the solvent.
5. The solubility of ammonium chloride at that temperature.
6. The solubility of ammonium chloride at 10°C is approximately 33 g/100g H₂O, and at 90°C it is 73 g/100g H₂O. Both values were determined by extrapolation from the data.

Appendix 4.4: Unsaturated, Saturated, and Supersaturated Solutions

Problem

To distinguish between unsaturated, saturated, and supersaturated solutions.

Materials

- sodium thiosulphate pentahydrate, $\text{Na}_2\text{S}_2\text{O}_3 \cdot 5\text{H}_2\text{O}$, crystals
- beaker (400 mL)
- 3 test tubes (18 x 150 mm)
- hot plate
- stirring rod
- water

Procedure

1. Make a water bath by pouring about 250 mL of water into a 400 mL beaker. Place the beaker on a hot plate, and heat the water to just boiling.
2. Place about 2.5 to 3 g (1.5 to 2 cm) sodium thiosulphate crystals into an 18 x 150 mm test tube. Add 1 mL of water. Heat in the boiling water bath until all crystals have dissolved.
3. Remove the test tube from the water bath and label it as test tube A.
4. Allow test tube A and its contents to cool to room temperature (about half an hour or overnight).
5. Place about 2.5 to 3 g (1.5 to 2 cm) sodium thiosulphate crystals into an 18 x 150 mm test tube. Add 1 mL of water. Shake and stir to dissolve as much of the solid as possible.
6. Label this test tube as B. Allow test tube B and its contents to settle and come to room temperature.
7. Place enough sodium thiosulphate crystals into an 18 x 150 mm test tube to cover just the bottom of the test tube (approximately 0.2 to 0.3 g). Add 3 mL of water. Shake and stir to dissolve as much of the solid as possible.
8. Label this test tube as C. Allow test tube C and its contents to settle and come to room temperature.
9. When test tube A and its contents are cool, carefully add one small crystal of sodium thiosulphate to each of test tubes A, B, and C. Describe your observations.
10. Hold the bottom of each test tube in your hand. Describe your observations.

Questions

Answer the following questions, based on your observations.

1. Identify which solutions are unsaturated, saturated, and supersaturated.
2. Describe the differences between unsaturated, saturated, and supersaturated solutions.

Appendix 4.4: Unsaturated, Saturated, and Supersaturated Solutions (Teacher Notes)

The experiment is intended to show the difference between the following solutions:

- an *unsaturated solution* (test tube C)—an added crystal dissolves in this solution
- a *saturated solution* (test tube B)—a solid is present together with the solution, and adding another crystal has no effect
- a *supersaturated solution* (test tube A)—an added crystal causes the “extra” solute to crystallize out around the seed

Students should also observe the following:

- The test tube and contents become warmer as the solid crystallizes.
- When making the saturated solution, dissolving requires energy, as evidenced by test tube B and its contents becoming colder as solute dissolves.

Variations

1. Prepare the three solutions (saturated, unsaturated, and supersaturated) in advance. Give corked and labelled test tubes to students and ask them to observe the effect of adding one crystal of sodium thiosulphate pentahydrate to each test tube. These solutions can be reused several times.
2. For a microscale version of the experiment, prepare the supersaturated solution by placing 3 and 4 cm (about a third to half full) of sodium thiosulphate in a 10 x 70 mm test tube and adding 5 drops of water. Use a 100 mL beaker and 50 mL of water for the water bath. Heat the solution in the boiling water bath to dissolve the solid, and allow the solution to cool to room temperature before adding a single crystal of sodium thiosulphate.

Note: Sodium thiosulphate is also known as “hypo,” and should be readily available from camera stores.

Solubility of Sodium Thiosulphate Pentahydrate, $\text{Na}_2\text{S}_2\text{O}_3 \cdot 5\text{H}_2\text{O}$, in Water		
Temperature (°C)	Grams of $\text{Na}_2\text{S}_2\text{O}_3 \cdot 5\text{H}_2\text{O}$ /100 g Saturated Solution	Grams of $\text{Na}_2\text{S}_2\text{O}_3 \cdot 5\text{H}_2\text{O}$ /100 g Water
0	33.40	50.15
10	37.37	59.66
20	41.20	70.07
25	43.15	75.90
30	47.17	91.24
45	55.33	123.87

Reference: Solubility data from: Seidel, A., ed. *Solubilities of Inorganic and Organic Compounds*. Vol. 1. New York, NY: D. van Nostrand Co., Inc., New York, 1911.

Appendix 4.5: Crystals and Crystal Growing

A. Rock Candy Sugar Crystals

This is a popular activity because students can grow large edible sugar crystals. This activity can also be done at home.

Safety Precautions

If crystals are to be eaten, use only clean household equipment. If crystals are to be observed only, laboratory glassware could be used.

Materials

- ½ cup (125 mL) water
- small saucepan
- short length of white string
- pencil
- 1 cup (250 mL) granulated sugar
- small nut or paper clip (to act as a weight)
- wooden spoon
- measuring cup
- tall glass
- hot plate

Procedure

1. Clean all equipment thoroughly before use.
2. Tie a small nut or paper clip to one end of a clean piece of white string and tie the other end of the string to a pencil. The string should be just long enough to touch the bottom of the glass when the pencil is resting across the top of the glass.
3. Pour the sugar and water into the saucepan and stir to dissolve as much of the sugar as you can at room temperature. Heat the mixture on the hot plate, stirring constantly. Continue heating until all the sugar is dissolved.
4. Pour the hot syrupy solution into the glass. Hang the string, weighted down with a nut or paper clip, into the solution.
5. Set the glass and contents aside, covered loosely with a piece of paper, in a place where it won't be disturbed. Check the solution daily. Break up any crust that forms on the surface so that water can continue to evaporate as crystals grow along the string.

Appendix 4.5: Crystals and Crystal Growing (continued)

B. Crystals on Glass

Rather than growing large crystals in this activity, students produce small crystals rapidly, for observations on shape and form.

Materials

- petri dish, beaker, or drinking glass
- magnesium sulphate, MgSO_4
- sodium chloride, NaCl
- sodium carbonate, Na_2CO_3
- copper sulphate, CuSO_4
- hot plate
- stirring rod

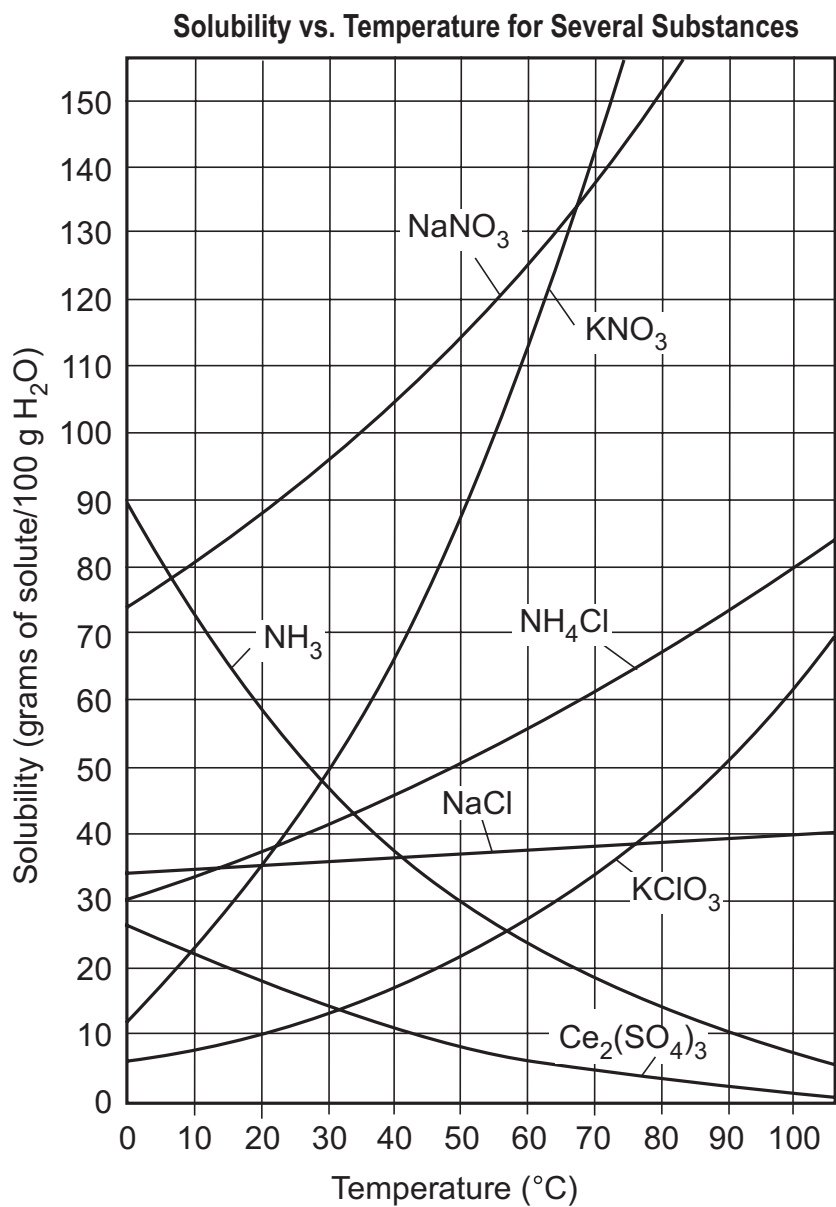
Procedure

1. Mix the appropriate amounts of solid compound and boiling water (see table below). Stir to dissolve the solid completely. (If necessary, add more boiling water. However, if the solution is too diluted it may take much longer for crystals to appear.)
2. When all the solute has dissolved, pour just enough solution to completely cover the surface of a flat-bottomed glass container (e.g., petri dish, beaker, or drinking glass). Set the container aside. Do not disturb while the water evaporates.
3. Check the solution occasionally. Describe the shape of the crystals as they start to form on the glass. Do they all have the same shape?

Compounds that yield good crystals include the following:

Compounds That Yield Good Crystals		
Substance	Amount of Solute (mL)	Amount of Boiling Water (mL)
Magnesium sulphate, MgSO_4	10	15
Sodium chloride, NaCl	15	30
Sodium carbonate, Na_2CO_3	15	30
Copper sulphate, CuSO_4	15	30

Appendix 4.6: Solubility Curve



Appendix 4.7: The Effects of Salt and Antifreeze on the Melting Point of Ice

In this investigation you will study the effects of salt and antifreeze on the melting point of ice.

Materials

- beaker (250 mL)
- thermometer
- coarse table salt
- 10 mL measure
- finely crushed ice (as small as coarse sand)
- antifreeze (ethylene glycol)
- graduated cylinder (25 mL)
- stirring rod

Procedure A: Effect of Salt on the Melting Point of Ice

1. Pour crushed ice into the beaker until it is half full.
2. Take the temperature of the ice while you are stirring it. Do not stir with the thermometer. Use the stirring rod.
3. When the thermometer reads 0°C , add 1 level measure of salt to the ice. Continue stirring. Record the temperature after it stops changing.
4. Continue adding 1 level measure of salt at a time, with continuous stirring and temperature readings, until you have added 7 or 8 measures.
5. Record your results in a table.
6. Empty the beaker and wash it.

Procedure B: Effect of Antifreeze on the Melting Point of Ice

The antifreeze used in this investigation is ethylene glycol. It is the permanent type that comes in new cars. Repeat Procedure A, but use 15 mL of antifreeze each time instead of 1 level measure of salt.

Discussion

1. What evidence do you have that the ice is at its melting point, even after salt and antifreeze are added?
2. What effect does salt have on the melting point of ice?
3. What effect does antifreeze have on the melting point of ice?
4. Explain why salt is added to ice-covered roads.
5. Explain why antifreeze is added to the water in the radiators of cars in cold areas of the country.
6. Why is salt not used in the radiators of cars?
7. Where did the heat go in this experiment? In other words, how can something get colder when you add a warmer substance to it in a warm room?

The Effects of Salt and Antifreeze on the Melting Point of Ice: Adapted from William A. Andrews, T.J. Elgin Wolfe, and John F. Eix, *Physical Science: An Introductory Study* (Scarborough, ON: Prentice-Hall of Canada, Ltd., 1978) 79–80. Adapted by permission of Pearson Education Canada.

Appendix 4.8: The Effect of Antifreeze on the Boiling Point of Water

Most vehicle manufacturers recommend leaving the antifreeze in the radiator during the summer. Why would one do this?

Materials

- round-bottomed flask (250 mL)
- thermometer (-10°C to 110°C)
- distilled water
- boiling chips
- Bunsen burner
- ring stand
- iron ring
- wire gauze
- 2 adjustable clamps
- 50 mL antifreeze (ethylene glycol, undilute)

Safety Precaution

Wear safety goggles during this investigation.

Procedure

1. Set up the apparatus as demonstrated by your teacher. Begin with 100 mL of distilled water in the flask.
2. Heat the water until it is boiling. Note the temperature.
3. Add 10 mL of antifreeze to the water. Continue heating the mixture. Note the new boiling point.
4. Repeat step 3 until all the antifreeze has been used.

Discussion

1. Describe the effect of antifreeze on the boiling point of water.
2. Why do car makers recommend leaving the antifreeze in a car radiator during the summer?
3. Compare the effects of antifreeze on the boiling point and freezing point of water.
4. Does antifreeze help or hinder boiling? Does antifreeze help or hinder melting?

The Effect of Antifreeze on the Boiling Point of Water: Adapted from William A. Andrews, T.J. Elgin Wolfe, and John F. Eix, *Physical Science: An Introductory Study* (Scarborough, ON: Prentice-Hall of Canada, Ltd., 1978) 85–86.

Appendix 4.9: Heat Transfer: I Scream, You Scream, We All Scream for Ice Cream

Purpose

This lab activity is a practical, real-life application of the principles of calorimetry, heat transfer, heat exchange measurements, heat of fusion, specific heat, and freezing point depression.

Materials

- 2 sandwich-size and 2 large (4 L) resealable plastic bags to make into double-thickness bags (to prevent leaks)
- large bowl
- measuring spoons
- measuring cup
- spoon
- egg beater
- balance
- Celsius thermometer
- 3 cups crushed ice
- 1/2 cup rock salt

Recipe

- 1/2 cup cream (or half-and-half)
- 1/4 cup milk
- 2 tablespoons sugar
- 1/8 teaspoon vanilla
- 1 beaten egg

Procedure

1. Measure and record the mass of a sandwich-size plastic, zippered bag.
2. In the large bowl, mix together the cream, milk, sugar, vanilla, and beaten egg. Pour this ice cream mixture into a sandwich-size resealable bag.
3. Measure and record the mass of the ice cream mixture.
4. Measure and record the initial temperature of the ice cream mixture.
5. Place the ice cream mixture into the second sandwich-size bag, making sure the bags are sealed tightly.
6. Place the small bag of ice cream mixture inside the large bag of ice and salt, making sure everything is sealed tightly.
7. Gently rock the bag for about five minutes. (Don't rock too hard, or you'll end up with salty ice cream!)
8. When the ice cream looks frozen enough, remove the small bag from the ice-salt mixture and rinse the bag in cold water to remove any salt on the outside.

**Appendix 4.9: Heat Transfer: I Scream, You Scream, We All Scream
for Ice Cream** *(continued)*

9. Measure and record the final temperatures of the ice-salt mixture and the ice cream mixture.
10. Eat the ice cream!

Data Table	
• Mass of small plastic, zippered bag:	_____
• Mass of ice cream mixture + bag:	_____
• Mass of ice cream mixture:	_____
• Initial temperature of ice cream mixture:	_____
• Final temperature of ice cream mixture:	_____
• Final temperature of ice cream mixture:	_____

Appendix 4.10: The Effect of Salt on the Melting Point of Ice

Question

What is the effect of salt on the melting point of ice?

Prediction

Predict how the addition of salt will affect the melting temperature of ice.

Materials

- computer system and temperature sensor
- beaker (400 mL)
- table salt, NaCl (50 g—divided into 10 portions of 5 g each)
- finely crushed ice (200 mL)
- stirring rod
- support stand
- utility clamp

Safety Precautions

- This experiment must be done carefully to avoid burns and broken glass.
- Wear a lab apron and safety goggles.

Procedure

1. Set up the computer system with the temperature sensor.
2. Display the temperature sensor with a graph so that temperature is along the vertical axis and time is along the horizontal axis.
3. Add the crushed ice to the beaker and set this on the base of the support stand.
4. Clamp the thermometer into the utility clamp and attach the clamp to the support stand.
5. Lower the thermometer tip into the middle of the crushed ice.
6. Using the stirring rod, stir the ice continuously.
7. Start the data collection.
8. When the thermometer stabilizes, add 5 g of salt to the ice. Continue stirring.
9. Repeat step 8 until you have added 50 g of the salt to the ice.
10. Stop the data recording.
11. Clean up the materials as directed by the teacher.

Appendix 4.10: The Effect of Salt on the Melting Point of Ice *(continued)*

Questions**Analysis**

1. What evidence do you have that the ice is at its melting point even after the salt has been added?
2. What appears on the outside of the beaker?
3. What minimum temperature was attained?

Conclusion

4. What effect does salt have on the melting point of ice?

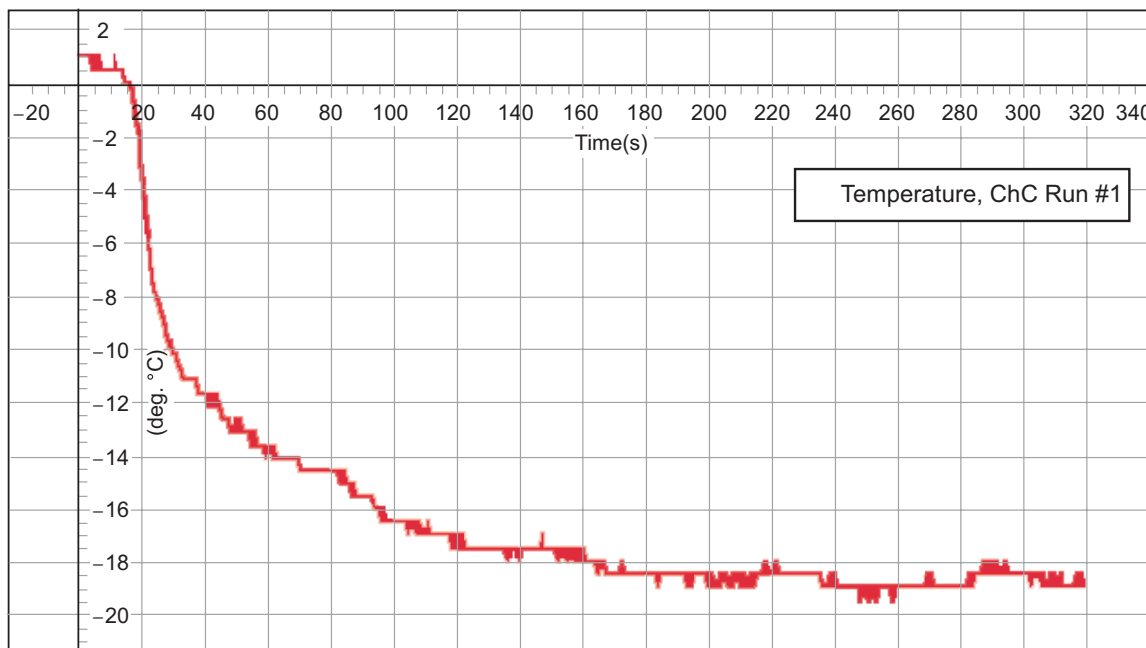
Applications

5. Explain why salt is added to ice-covered roads.
6. This experiment can be repeated using ethylene glycol (antifreeze) instead of salt.

Appendix 4.10: The Effect of Salt on the Melting Point of Ice (Teacher Notes)

Sample Data: Answer Key

Observations



Temperature vs. Time

Analysis

1. Both the solid and the liquid states are present.
2. Frost appears on the outside of the beaker.
3. The minimum temperature attained was -19°C .

Conclusion

4. Salt lowers the melting point of ice.

Applications

5. Salt dissolves in the fine layer of water found on the surface of ice. The melting point of the ice is then lowered. When the melting point of the ice is below the surrounding temperature, the ice melts.

Appendix 4.11: A WebQuest for Solubility Units

Work with a partner in the computer lab (or at home) to find answers to the following questions.

- The strength (or concentration) of a solution can be represented in several forms, dependent upon the intended usage or the nature of the components making up the solution. Remember, a solution has two components: a *solute* and a *solvent*. Therefore, these components are often included in the units to express the strength of a solution. Complete the following chart.

Unit to Express the Strength of a Solution	What the Unit Means in Long Form	Example of Where You Might Find This Unit Being Used
<i>Example:</i> g/L	Grams of solute in 1 litre of solution	Mixing powdered chemicals into a swimming pool
% w/v		
% v/v		
ppm		
ppb		
mol/L		

- Describe the process of treating a water supply.
- Identify the allowable concentrations of metallic and organic species in water suitable for human consumption. (Consider tap water, well water and/or bottled water.)

Appendix 4.12: Solutions: A Scavenger Hunt

You and your partner will search through your chemistry textbook to find answers to the following questions. If you cannot find all the answers there, use another source that your teacher has supplied for you or recommended to you. You will not find all the answers in the same section of the textbook, so it is important that you use the index in the text correctly.

Write a thorough, but precise, answer to each of the following questions. Don't write "stuff" you don't understand or have simply borrowed for the day. Make every effort to write in your own words and in a familiar style. This is an important skill that will help you in any advanced studies.

1. Describe and give examples of various types of solutions. Include all nine possible types.

Type of Solution	Description	Example

