**Ice Creamer**

(Taken in part from [www.carolina.com](http://www.carolina.com))

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**Introduction**

Demonstrate the colligative property freezing-point depression and make edible ice cream at the same time. This safe, inexpensive, and engaging experiment works well for a science night, open house, kids' group or anywhere fun and science meet. Using table salt, reduce the temperature of a mix of ice and liquid water to below the freezing point of water. After 5 minutes, open the cup for a tasty treat.

**National Science Education Standards**

* Grades 9-12: Structure and Properties of Matter

**Teacher Background:**

The physical properties of solutions differ from those of pure solvents. Pure water freezes at 0°C, but aqueous solutions freeze at lower temperatures than the pure solvent. Like vapor pressure and boiling point, freezing point is a colligative property. Colligative properties are determined by the number of particles involved rather than the type of particle. The particles may be molecular or ionic as long as they are dissolved and nonvolatile.

Antifreeze (ethylene glycol) is often added to the cooling system of cars to lower the freezing point and raise the boiling point of the aqueous solution within. Ethylene glycol, the solute alcohol, extends the liquid range of the water in the cooling system to well below 0°C and well above 100°C. The extreme heat produced by the engine is absorbed and released by the aqueous solution circulating through the radiator, and the solution is much less vulnerable to freezing in extremely cold temperatures and from boiling at very high temperatures, thereby protecting the automobile from damage.

Colligative properties must be considered in the context of how dissolved solute particles interact with solvent particles. For instance, for a solution to freeze, molecules of the solvent must cluster to form a solid. The temperature at which this solidification occurs is the freezing point of the solution. When a solute is added to a solvent, the solute particles prevent the solvent molecules from clustering as readily. The solvent molecules bounce off the solute particles rather than sticking to each other. One way to make the solution freeze is to lower its temperature further, taking kinetic energy away from the solvent molecules. As the solvent molecules slow down, they become more likely to form clusters. When the temperature is lowered sufficiently and the clusters form, the new freezing point has been reached. Thus, the depression of the freezing point of the solution is a result of the number of solute particles in the solution.

To boil, a liquid’s vapor pressure must equal the atmospheric pressure above it. In a solution, the solute particles block some of the solvent molecules from escaping at the surface. To energize the solvent molecules enough to create vapor pressure to equal atmospheric pressure, the temperature of the solution must be increased. When the temperature is raised sufficiently to create vapor pressure to equal atmospheric pressure, the new boiling point has been reached. Thus, the boiling point of the liquid has been elevated by the solute particles in it.

The equations used to find the freezing-point depression and boiling-point elevation are as follows:

freezing-point depression: ΔTf = *i*Kf*m*

boiling-point elevation: ΔTb = *i*Kb*m*

ΔT = the change in temperature

*i* = the van’t Hoff factor (the number of particles from dissociation of ionic compounds)

For molecular compounds, because there is no dissociation, *i* = 1.

Kf = the molal freezing-point constant (for water, Kf = 1.86°C/m)

Kb = the molal boiling-point constant (for water, Kb = 0.52°C/m)

*m* = molal concentration (moles of solute/kg of solvent)

The molal freezing-point constant, Kf, is the number of degrees Celsius that a liquid compound is depressed below its freezing point for every mole of particles dissolved per kilogram of solvent. For example, water has a molal freezing-point constant of 1.86°C/*m*, which means that water can be depressed 1.86°C below 0°C for every mole of solute particles dissolved per kilogram of water. Ionic compounds depress the freezing point and elevate the boiling point of a liquid much more than molecular compounds. This is based on the van’t Hoff factor (*i*), the number of particles from dissociation of the solute. The molal boiling-point constant, Kb, is the number of degrees Celsius that a liquid compound is elevated above its boiling point for every mole of particles dissolved per kilogram of solvent. Here is the freezing-point depression calculation for a 1 *m* aqueous solution of sodium chloride. Sodium chloride dissociates into a mole of Na+ ions and a mole of Cl– ions, so the van’t Hoff factor is 2.

ΔTf = (2)(1.86°C/*m*)(1 *m*) = 3.72°C

Therefore, a 1 *m* aqueous solution of sodium chloride should freeze at –3.72°C.

**Ice Cream Lab:**

**Materials**

* Individual Packet (sealed cup) of Liquid Non-Dairy Creamer
* 16-oz Deli Cup and Lid
* Ice
* Water (tap, deionized, or distilled)
* [Sodium Chloride](http://www.carolina.com/catalog/detail.jsp?prodId=888880), about 50 g
* Taster Spoon

**Safety**

No special safety precautions are required in using these materials. It may be helpful to perform this activity in a non-laboratory setting so as not to confuse students about the basic rule of not eating or drinking anything in the lab.

**Preparation and procedure**

1. Pour the 50 g of sodium chloride (salt) into the deli cup. If you do not have a scale, simply pour a layer of salt from .5 to 1 cm deep.
2. Place the unopened creamer cup in the deli cup.
3. Fill the deli cup with ice.
4. Fill the deli cup with water, leaving just enough room to attach the lid.
5. Place the lid on the cup.
6. Swirl the cup on a flat surface for 5 min to mix.
7. Remove the creamer and, without opening it, check for firmness either by squeezing the sides or shaking the cup near your ear. If you feel or hear liquid sloshing, place the creamer back in the deli cup and swirl for 5 more minutes.
8. Open and enjoy!

**Conclusion**

Freezing-point depression is an example of a colligative property. A colligative property of a solution is one that depends on the number of solute particles but not on the solute's identity. For example, the freezing point of an aqueous 1 molal sodium chloride solution is the same as the freezing point of a 1 molal potassium chloride solution because sodium chloride (NaCl) and potassium chloride (KCl) dissociate into the same number of solute particles, thus lowering the freezing point of the solvent (water) the same amount. Calcium chloride (CaCl2) lowers the freezing point to a greater extent because it dissociates into 3 ions (1 calcium and 2 chloride) rather than 2. Other colligative properties include boiling-point elevation and osmosis and diffusion.

In this experiment, sodium chloride is added to ice and lowers the freezing point of water from 0 to –10°C. With its surroundings reduced to this temperature, the non-dairy creamer freezes. Packing the creamer cup in ice, water, and salt and then swirling to maintain temperature evenly through the cup results in ice creamer. Enjoy this tasty application of a colligative property.

**Extension**

The creamer itself consists of water as a solvent and contains particles of various types. While its own freezing point is not depressed to the extent of the ice-water-salt mix, it is depressed. Your students may enjoy developing experiments to find the freezing point of the creamer or to compare the freezing point among creamers with different ingredients.

**Pre-Lab Questions:**

Define COLLIGATIVE PROPERTIES:

Give four examples of COLLIGATIVE PROPERTIES

1. 2. 3. 4.

Increasing the concentration of a solution will \_\_\_\_\_\_\_\_\_\_ the boiling point of the solution

Increasing the concentration of a solution will \_\_\_\_\_\_\_\_\_\_ the freezing point of the solution

**Analysis:**

1. Calculate Δ Tf (change in the freezing point): \_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_
2. Using the equation ΔTf = kf m, calculate m (molality) \_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_

Remember kf = 1.86

**Post lab Questions:**

1. Why was salt added to the ice? (Why did your milk/cream/sugar mixture freeze?)
2. Where did the heat the flow from (What happened to the temperature of the ice after salt was added?)

1. What is the minimum temperature that pure water can exist as a liquid at standard pressure?
2. What do you think would happen to the temperature of the ice if you added 6 tablespoons of salt instead of 2 tablespoons?

5. If there are charged particles that can move around in a substance, that substance is able to conduct electricity. So, do ionic compounds or covalent compounds conduct electricity when dissolved in water? Explain. Use the idea of what happens when these two different types of compounds dissolve in water in your answer.

6. Which compound types (ionic or covalent) produce more particles when dissolved in water and why?

7. Answer the following questions:

How many particles are produced when a mole of sugar dissolves? \_\_\_\_\_\_\_

How many particles are produced when a mole of CO2 dissolves? \_\_\_\_\_\_\_

How many particles are produced when a mole of sodium chloride (NaCl) dissolves? \_\_\_\_\_

How many particles are produced when a mole of copper (II) chloride (CuCl2 ) dissolves? \_\_\_\_\_

How many particles are produced when a mole of iron (III) sulfate [Fe2(SO4)3] dissolves? \_\_\_\_\_

8. Which type of compound (ionic or covalent) will have a greater affect on the colligative properties of a solution? Explain. Use the answers to the previous questions in your answer.

 9. Why was salt used in the outside bag in this lab? Be as specific as possible. Use the idea of colligative properties and the idea of covalent and ionic compounds in your answer.

10. Explain in detail, using the idea of colligative properties, why salt is used to make icy roads safe for driving on in the winter.