Thermochemistry: Heat and Chemical Changes

Strand Phases of Matter and Kinetic Molecular Theory

Topic Investigating properties of matter

Primary SOL CH.5 The student will investigate and understand that the phases of matter are explained by kinetic theory and forces of attraction between particles. Key concepts include
- a) pressure, temperature, and volume;
- b) partial pressure and gas laws;
- c) vapor pressure;
- d) phase changes;
- e) molar heats of fusion and vaporization;
- f) specific heat capacity;
- g) colligative properties.

Related SOL CH.1 The student will investigate and understand that experiments in which variables are measured, analyzed, and evaluated produce observations and verifiable data. Key concepts include
- a) designated laboratory techniques;
- b) safe use of chemicals and equipment;
- c) proper response to emergency situations;
- e) accurate recording, organization, and analysis of data through repeated trials;
- g) mathematical manipulations including SI units, scientific notation, linear equations, graphing, ratio and proportion, significant digits, and dimensional analysis;
- h) use of appropriate technology including computers, graphing calculators, and probeware for gathering data, communicating results, and using simulations to model concepts.

Background Information

Colligative properties are properties of solutions that depend on the amount of solute particles in the solution (concentration) and are independent of the nature of the solute. Freezing point depression, boiling point elevation, vapor pressure lowering, and osmotic pressure are all colligative properties. The word colligative is derived from the Latin *colligatus* meaning *bound together*, since these properties are bound together by the fact that they all depend on the number of solute particles and not on the type of chemical species present.

Materials
- Large rubber bands
- 13 × 18 cm piece of sheet metal
- 13 × 18 cm piece of Styrofoam
- Approximately 40 × 20 cm piece of wood
- Thermometers

- Ring stands and clamps
- Metal rods
- Bunsen burners
- Heat-resistant gloves
- Erlenmeyer flasks
- Water
- Food coloring
- One-hole rubber stopper
- Glass tubing
- Ice
- Two-hole rubber stopper
- Hot plate
- Large beaker or bucket
- Safety goggles
- Tongs
- Snack-food packages
- Soda crackers and box
- Empty soda can
- Dissecting needles
- Pictures of frogs, fish, or reindeer
- Large picture of a car
- Flat-bottomed Florence flasks
- Bag of raisins
- Bag of dried cranberries
- Pint-size, unopened glass bottle of soda water
- Rock salt
- Styrofoam cup

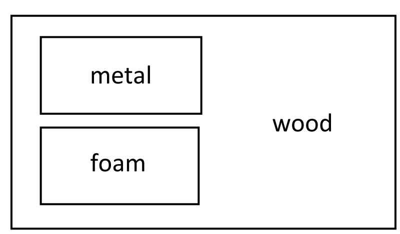
Vocabulary

absorb, calorie, Calorie (kcal), Celsius, conduction, convection, endothermic, exothermic, heat, Kelvin, momentum, solute, solvent, thermochemistry

Student/Teacher Actions (what students and teachers should be doing to facilitate learning)

In this lesson, students focus on describing some physical properties of solutions. They learn why some solutions conduct electricity while others do not. They also learn about three colligative properties of solutions.

Introduction



1. Before the introductory activity, make a "temperature-comparison board" by gluing a 13 × 18 cm piece of sheet metal and a 13 × 18 cm piece of Styrofoam to a piece of wood that is about 40 × 20 cm, as shown at right.
2. After cautioning students about misuse, give each student a clean, medium-sized rubber band.
3. Have students do this procedure, following your verbal instructions:

   a. Hook your index fingers through the ends of the rubber band. Without stretching the rubber band, place it against your upper lip or forehead, and note its temperature.
   b. Move the rubber band away from your skin. Quickly stretch and hold it, and again place it against your skin. Note any temperature change.
   c. Fully stretch the rubber band, and then allow it to return to its original state. Once more, place it against your skin, and note any temperature change.
   d. Repeat steps 3.b and 3.c until you are certain of the temperature change in each step.
4. Ask students the following questions:
   - Did the rubber band feel cool or warm after it was stretched in step 3.b?
   - Did the rubber band feel cool or warm after it returned to its original shape in step 3.c?
   - What is heat?
   - In what direction does heat flow?
5. Students will discover that the rubber band feels warmer after it is stretched and cooler after it is allowed to relax. Students may infer that heat is related to the observed temperature changes. From their experience, students may realize that heat flows from a warm object to a cool object. The stretching of rubber is exothermic; the reverse process—relaxing the stretched rubber—is endothermic. Define exothermic reactions and endothermic reactions. Stress that in an exothermic reaction, heat is released, and the energy of the products is less than that of the reactants. Draw a graph on the board to depict this. Emphasize that in an endothermic reaction, the energy of the products is greater than that of the reactants. Graph this on the board.
6. Review thermochemistry as the study of heat changes occurring during chemical reactions.
7. Pass around the piece of wood with the sheet metal and Styrofoam attached, and ask students to put their hands on each of the three surfaces and describe the temperature of each. Usually, students respond that the metal feels the coolest and the Styrofoam feels the same temperature as their hands. Explain that all three surfaces are actually the same temperature—the temperature of the room, which is usually considerably cooler than their hands—and that when they touch those surfaces, heat is being transferred from their hands to them. Metals conduct heat away from the hand more rapidly than Styrofoam does, so the metal "feels" cooler than the Styrofoam.

Procedure

Heat versus Temperature; Conducting Heat Energy through a Solid

1. Begin with a review of the arrangement of atoms in a crystal, emphasizing that the atoms are held together in an orderly arrangement that gives the crystal a definite shape.
2. Ask students to consider what would happen to the atoms if they began to absorb energy from, for example, an open flame. Point out that the atoms would move around more vigorously, slamming into one another with greater momentum. This could change the arrangement of atoms in the crystal and alter its shape. The crystal could melt and eventually vaporize. Explain that this change would be the result of an increase in the momentum of the atoms in the system. The transfer of energy from atom to atom in a solid is called *conduction.*
3. Ask students to consider what would happen if a thermometer were to be placed against the crystal while it is being warmed. Have student analyze the system at the atomic level

and higher and write a brief statement that explains why the liquid inside the thermometer reads higher and higher as the crystal is warmed. Guide students to the conclusion that the thermometer does not measure heat directly; instead, it reflects the average kinetic energy of the atoms in the system. *Heat* is a measure of the total energy of a system. The heat energy released during a chemical change in a substance can be measured using a *calorimeter.* The unit of heat energy is the *calorie*: one calorie is the amount of energy needed to raise the temperature of 1 gram of pure water 1 degree Celsius.

4. Distribute thermometers that read both degrees Fahrenheit and degrees Celsius. Remind students that each thermometer uses several scales, Fahrenheit and Celsius, to read the same amount of average kinetic energy; therefore, for example, 32°F = 0°C = 273K. Show students how to use the temperature conversions to change temperatures from Fahrenheit to Celsius to Kelvin.
5. <u>Conducting Heat Energy through a Solid</u>: In the lab, have students track the change of temperature of a metal being heated, as follows: Have them clamp a thermometer to a ring stand and also clamp a metal rod to the same ring stand. Direct them to use heat-resistant gloves while heating the metal rod with a Bunsen burner, reading and recording the temperature every 15 seconds for 2 minutes. Then, ask them to use this data to make a line graph of the change of temperature over time.

Transferring Heat though a Liquid or Gas

1. Pour a small amount of water and several drops of food coloring into a flask, and stopper the flask with a one-hole rubber stopper. Place a piece of glass tubing through the stopper and into the colored water. Rub your hands together, place them over the bottom of the flask, and have students watch closely and record their observations. Next, rub ice over the bottom of the flask in the same place you positioned your hands, and have students watch and record their observations.
2. Fill a small Erlenmeyer flask with water and several drops of food coloring, and stopper the flask with a two-hole rubber stopper. Warm the flask on a hot plate on a low setting for two minutes. While the flask is warming, fill a large beaker or bucket with cold water. When the water in the flask is hot, put on goggles and use tongs or heat-resistant gloves to transfer the flask to the bottom of the large beaker or bucket. Have students record their observations. *CAUTION! Use great care with laboratory glassware; although it is tempered to withstand drastic temperature changes, it is possible that it will shatter. Follow the MSDS General Safety Precautions.*
3. Discuss students' recorded observations from both labs. In the first, they should have observed the colored water rising in the tube as the air in the flask expands due to heat convection and then the colored water sinking in the tube as the air contracted. In the second lab, they should have observed the colored, hot water escaping from the flask when it was placed in the large beaker or bucket of cold water.
4. Review the results of the previous activity, repeating the distinction between heat energy and temperature. Explain that the metal bar used in the previous activity did not hold heat very well but transferred the heat very quickly; in other words, the metal was an excellent conductor of heat but did not have the capacity to hold or store heat. Scientists can

measure the capacity of a substance to hold or store heat. The capacity of a substance to store chemical energy is called *specific heat.* Water has a specific heat equal to 1 because it takes one calorie of energy to raise the temperature of 1 gram of water 1°C. The specific heat of iron, on the other hand, is only 0.11; that is, it takes 0.11 calories to raise the temperature of 1 gram of iron 1°C—only about one-tenth the amount of energy needed to raise the temperature of an equal amount of water.

5. Have students use reference materials to create their own Specific Heat of Substances Chart. Assign some substances that they should know about, and have them choose some that they want to learn about. Many periodic tables give the specific heat of each element. You may wish to have them graph each substance listed on the chart.
6. Ask students also to determine the factors that would make a substance a good insulator. Ask, "Would such a substance have a low or high specific heat?" *(Insulators have a high specific heat. They store heat energy and prevent its transfer.)* Have students use the reference materials to make a list of good insulators.
7. Have students present their Specific Heat of Substances Charts and insulator lists.

Measuring Calories

1. Before this lab, have students bring in examples of the outside packaging from their favorite snack foods.
2. Remind students that *conduction* is the transfer of energy through a solid. Now, point out that energy can also be transferred from one particle to another in a liquid or a gas, and define *convection* as the transfer of energy through a fluid (i.e., a liquid or gas).
3. Review the definition of a *calorie*—the amount of energy needed to raise the temperature of 1 gram of pure water 1 degree Celsius. Therefore, raising the temperature of 100 grams (100 mL) of water 1 degree Celsius would require 100 calories.
4. Explain the difference between a food Calorie and an energy calorie: 1 food Calorie is equal to 1,000 energy calories, as measured using a thermometer or calorimeter. Have students read the ingredients label on the packaging of their favorite snacks, paying particular attention to the calories-per-serving information. Then, have them read the ingredients label on a soda cracker box. Example: "Calories per serving = 140; serving size 10 crackers." Ask, "How many food Calories are there in one soda cracker?" *(14)* "How many energy calories are in one cracker?" *(14,000)* Burning one cracker and using *all* of the released energy to heat 100 mL of water would heat the water to about 140°C—quite a lot of heat! Explain that in the lab that students are about to perform, they will see that most of the heat from the burning cracker will be lost.
5. Have students conduct the following lab:
6. Pour 100 mL of water from a beaker into an empty soda can.
7. Secure the soda can with two ring clamps, and lower a thermometer into the can just below the surface of the water.
8. Skewer a soda cracker onto the end of a dissecting needle, and clamp it to the ring stand below the soda can. Record the temperature of the water. Hold a lighted match to the soda cracker until the cracker burns on its own. Record the temperature of the water when

the cracker has completely burned. Have students determine the difference between the theoretical rise in temperature (about 140°C) and what they actually observed. Have them explain why this happened.

9. At the board, perform some calorimetry calculations. Have students practice more calorimetry calculations for homework.

The Colligative Properties of Solutions

1. Before students enter the room, write on the board or make a banner that reads:

Colligative property depends on the *concentration* of particles, not their identity!

2. Also, place a large picture of a car at the front of the room.
3. As students arrive, give each of them a picture of a frog, fish, or reindeer, and point out the large picture of a car at the front of the room. Have students try to figure out the following riddle: What do these three animals have in common with a car? When they have guessed long enough, give them a clue: antifreeze. If no one "gets it," tell them these animals, like a car, can survive being frozen. Scientist believe that a substance in the cell of these animals act as a natural antifreeze, which prevents their cells from freezing. Although fluids surrounding their cells may freeze, the cells themselves do not.
4. Ask students, "What happens to the molecules in water when water freezes?" *(The molecules form a crystalline lattice.)* "How does a solute change the freezing point of a solvent?" *(It might slow down the formation of the crystal lattice.)* "Why do we throw salt down on ice in the winter?" *(It lowers the freezing point, causing the ice to melt.)*
5. Point out that there are special properties of solutions. The physical properties of solutions differ from those of the pure solvent used to make the solution. These properties depend only on the number of particles dissolved in a given mass of solvent. The properties, called *colligative properties,* include vapor pressure lowering, boiling-point elevation, and freezing-point depression. Remind students that *vapor pressure* is the pressure exerted by vapor that is in dynamic equilibrium with its liquid in a closed system. Point out that a solution with particles that are not easily vaporized always has a lower vapor pressure than the pure solvent.
6. Place three beakers of water in front of students, and inform them that at this point, all three beakers have the same vapor pressure. The vapor pressure of a nonvolatile solution—a solution filled with particles that are not easily vaporized—is less than the vapor pressure of the pure solvent. Make sure students are writing the key points down in their lab books. Equilibrium is established between the liquid and the vapor in the pure solvent.
7. Add three raisins to the first beaker to represent the particles of the solute glucose. Now, point out that solvent particles form shells around the solute particles, thus reducing the number of free solvent particles able to escape the liquid. Equilibrium is eventually re-established at a lower vapor pressure. Use more raisins and some dried cranberries to show additional examples: in the second beaker, place three raisins and three cranberries

to represent the particles in sodium chloride; and in the third beaker, place three raisins and six cranberries to represent the particles in calcium chloride. As a review, have students determine the number of particles that belong in each solvent, based on the molecular formulas of sodium chloride and calcium chloride.

8. Ask students, "How many particles in solution are produced by each formula unit of aluminum bromide?" *(4)* "How many moles of particles would 3 mol Na_3PO_4 give in a solution?" *(12 mol of particles)* Stress that adding a nonvolatile solute to a solvent deceases the vapor pressure.
9. Define *boiling point* of a substance as the temperature at which the vapor pressure of the liquid phase equals the atmospheric pressure. Ask students, "If the vapor pressure is now lower because of the addition of the solute, then what happens to the boiling point of the solution?" *(It rises.)* Students will probably understand more readily if this is explained in terms of particles. Attractive forces exist between the solvent and solute particles. It takes additional kinetic energy for the solvent particles to overcome the attractive forces that keep them in the liquid. Thus, the presence of a solute elevates the boiling point of the solution. The magnitude of the boiling point elevation is proportional to the number of solute particles dissolved in the solvent. For example, the boiling of water increases by 0.51°C for every mole of particles that the solute forms when dissolved in 1,000g of water. This is why boiling point is a colligative property.

Freezing Point

1. When water freezes, the particles of the solid take an orderly pattern. The presence of a solute in the water disrupts the formation of this pattern. Point out that the solution will still freeze, but at a lower freezing point.
2. Have students observe the freezing point depression of water by the addition of a substance such as rock salt (NaCl) to ice-water mixture. You might have each group of students use a different salt. Provide each group with a thermometer, Styrofoam cup, and a mixture of ice and water, and ask them to measure the initial temperature of the ice-water mixture. Then have them measure the lowest temperature reached after the addition of the salt. Also, have them investigate the rate at which the frozen salt water melts to discover whether the salt water melts faster or slower than plain water.

Observations and Conclusions

1. Ask students, "Will 1 mole of sugar have the same effect as 1 mol of table salt, NaCl, in lowering the freezing point of water? Explain your answer." *(One mol of table salt will have approximately twice the effect in lowering the freezing point of water as 1 mol of sugar, because 1 mol of NaCl produces 2 moles of solute particles in solution, while 1 mole of sugar produces only 1 mole of solute particles.)*
2. Have students
   - define *thermochemistry*
   - define specific heat capacity
   - identify colligative properties of solutions
   - calculate energy changes, using specific heat capacity

- perform calorimetry calculations.

Assessment

- **Questions**
  - How did the temperature of the ice bath compare to the normal freezing point of water? *(The temperature was lower than 0°C because of freezing point depression caused by the salt.)*
  - Why did the soda water remain a liquid while in the ice bath? *(The CO_2 gas dissolved in the water lowered its freezing point.)*
  - What happened to the CO_2 when the bottle was opened, and why? *(The CO_2 came out of solution because the pressure was lowered when the bottle was opened.)*
  - Why did the water freeze when the bottle was opened? *(When the CO_2 came out of solution, the freezing point of the water was raised to 0°C, but the temperature of the water was already below 0°C.)*
- **Other**
  - Perform the following lab and complete the lab write-up in one class period. This lab is a good example to link the ideas of freezing point depression and gas solubility.
    - Each student will need a pint-size, unopened glass bottle of soda water, ice, a large beaker, and rock salt. Have students remove the label on their bottles, pack the bottle in a large beaker of ice, and sprinkle rock salt on top of the ice. Approximately 10 minutes later, have students remove their bottles from the ice. (The soda water should still be liquid.) Next, have students open their bottles. As the gas effervesces from the water, the water in the bottle will instantly freeze. The students should select the appropriate tools they need to plan and conduct this investigation.

Extensions and Connections (for all students)

- Have students research how solubility and colligative properties play a role in human physiology, environmental science, plant growth, and industry.

Strategies for Differentiation

- Have students use graphic organizer software to create a concept map to tie together the vocabulary terms and show their interrelationships.
- Have students use vocabulary word cards to manipulate and form graphic organizer webs showing the interconnectedness of the terms listed in vocabulary.
- Have students visit a state or county highway department and find out what type of ice melt is used and why. Ask the maintenance supervisor for the schools to explain what ice melt the school uses and why.
- Have students calculate the energy calories in their lunch, using the food labels.
- Have students create a personal vocabulary list of the target words as well as any other unknowns.
- Have students create prompt cards for the equations to convert between Fahrenheit and Celsius and between Celsius and Kelvin temperatures.