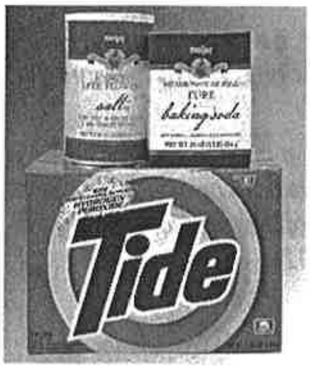
Dissolving Energy



Overview

In this activity, students investigate temperature changes that occur when various household substances dissolve in water.

Key Concepts

- anhydrous substances
- dissolving
- experimental design
- heat of solution
- solutions
- water of hydration

National Science Education Standards

Science as Inquiry

Abilities Necessary to Do Scientific Inquiry

- *Students use thermometers to determine the temperature changes that occur when common household substances are dissolved in water. (5–8, 9–12)*
- *Students use logic and evidence to formulate explanations about how dissolving various amounts of solid in the same amount of water affects the temperature of the solutions. (5–8, 9–12)*
- *Students gain experience in identifying and controlling variables as they plan and conduct experiments to investigate the heat of solution of various solids in water. (9–12)*
- *In Part A, students design and conduct an experiment to determine whether the starting temperature of the water affects the temperature change observed. (5–8, 9–12)*
- *In Part B, students design and conduct an experiment to determine whether or not the ratio of solid to water affects the temperature change. (9–12)*
- *In Part C, students use logic and evidence to formulate explanations about the differences in the temperature changes resulting from dissolving Lite Salt and table salt. (5–8, 9–12)*

Physical Science

Properties and Changes of Properties in Matter

- *In Part C, students observe that dissolving Lite Salt in water produces a different temperature change than dissolving table salt and learn that the different substances in these products account for that difference. (5–8)*

Structure and Properties of Matter

- *In Part C, by dissolving Lite Salt and table salt in water, students observe that these substances have different heats of solution. They learn that Lite Salt contains not only sodium chloride, but also potassium chloride, which has a heat of solution that is more endothermic than that of sodium chloride. (9–12)*

Transfer of Energy

- *Students discover that when a solute dissolves in water, a noticeable change in temperature can occur and they can feel this heat transfer as well as measure the temperature change with a thermometer. (5–8)*

Chemical Reactions

- *Students observe and conclude that some chemical reactions and other processes that involve the breaking of bonds and the formation of hydrated ions during the dissolving process can release or consume energy. (9–12)*

Part A: Student Exploration

Do temperature changes occur when common household substances dissolve in water?

Materials

- 100-mL graduated cylinder or measuring cup
- tablespoon measure
- 2–3 small plastic cups
- stirrer
- water
- baking soda (sodium bicarbonate, $NaHCO_3$)
- powdered laundry detergent (for example, Tide® or Cheer®)
- table salt (sodium chloride, NaCl)
- washing soda (sodium carbonate, Na_2CO_3)

⚠ *Washing soda is a severe eye irritant and minor skin irritant. Goggles must be worn at all times, and gloves are recommended. Washing soda may be harmful if swallowed or inhaled.*

- thermometer with a scale that includes room temperature and allows measuring temperature increments of 2°C
- other materials as needed for the student-designed experiment in step 3

Procedure

⚠ *Because some household solids are especially hazardous, confine your exploration to those listed in the Materials section.*

1. Pour 60 mL (¼ cup) room-temperature water into a cup and record the temperature. Add 60 mL (4 level tablespoons) baking soda to the water all at once. Stir and record the temperature. (Do not stir with the thermometer.) Continue stirring and recording the temperature until a maximum or minimum temperature is observed. When no further change in temperature is observed, discontinue stirring and recording after 2–3 minutes. *What evidence of change do you observe in the dissolving process?*

2. Repeat step 1 using powdered laundry detergent in place of baking soda. *How does dissolving the detergent differ from dissolving baking soda?* Repeat with table salt and then with washing soda. *Formulate a claim about one or more of the systems you have studied and substantiate your claim with evidence you have collected.*

3. Design an experiment to determine whether the starting temperature of the water affects the temperature change observed.

Part B: Student Exploration

Use Epsom salt and anhydrous magnesium sulfate to illustrate the heating and cooling effects that can result when a salt is dissolved in water.

Materials

- 60 mL (¼ cup) Epsom salt (magnesium sulfate heptahydrate, $MgSO_4{\cdot}7H_2O$)
- 60 mL (¼ cup) anhydrous magnesium sulfate ($MgSO_4$)
- 2 small zipper-type plastic bags
- water
- 100-mL graduated cylinder
- permanent marker
- other materials as needed for the student-designed experiment in step 3

Procedure

1. Label one plastic bag Epsom salt ($MgSO_4{\cdot}7H_2O$) and the other anhydrous magnesium sulfate ($MgSO_4$). Place 60 mL (¼ cup) of the appropriate solid into each labeled bag.
2. Measure out 60 mL (¼ cup) room-temperature water. Record its temperature. Pour the water into one of the solid-containing bags. Seal the bag and shake to mix. Record the highest or lowest temperature of the solution after mixing. Repeat with the other solid and compare results. *Make a claim about differences in your observations of the dissolution of the anhydrous and hydrated forms of magnesium sulfate and substantiate the claim with the data you have collected. What chemical species do you think are present in solution when anhydrous magnesium sulfate dissolves? What chemical species do you think are present in solution when Epsom salt dissolves?*
3. Design an experiment to determine whether or not the ratio of solid to water affects the change in temperature.

Part C: Student Exploration

Explore the differences between two flavor-enhancing products, table salt and Lite Salt™, by dissolving each in water.

Materials

- tablespoon, measuring cup, or graduated cylinder
- 30 mL (⅛ cup or 2 level tablespoons) table salt
- 30 mL (⅛ cup or 2 level tablespoons) Lite Salt
- water
- 2 small cups
- stirrer
- thermometer

Procedure

1. Place 60 mL (¼ cup) room-temperature water into a cup and record the temperature. Add 30 mL (⅛ cup) table salt to the water all at once and stir. (Do not stir with the thermometer.) *Record the highest or lowest temperature reached.*

2. Repeat step 1 using the Lite Salt in place of the table salt. *Read the label for the Lite Salt and record its composition. Use this information to interpret your observations. Why does the Lite Salt behave differently than table salt?*

Instructor Notes

Tips and Instructional Strategies

- This activity can be used when studying solutions and endothermic and exothermic processes. You may want to do Parts B and C with older students only.
- To prepare dehydrated Epsom salt for Part B, heat 120 mL (½ cup) Epsom salt (magnesium sulfate heptahydrate, $MgSO_4 \cdot 7H_2O$) in an aluminum pie pan for 30 minutes at 450°F (230°C). Remove from the oven and allow to cool for 10 minutes. Cover the dehydrated Epsom salt with plastic wrap and use a hard flat object to crush the salt into small pieces. Store the salt in a sealed container to minimize rehydration. This dehydration step can be done several days in advance of the activity or you may want to have older students do the dehydration as part of the activity. By doing this step themselves, students could determine the percentage of water in Epsom salt experimentally and compare their results with the actual percentage.
- Part C could also be done using plastic bags, as in Part B. The bag in which the Lite Salt dissolves will feel noticeably cooler than the one in which the table salt dissolves.
- As an extension to Part C, step 2, you may want to challenge students to predict what would happen if pure KCl were dissolved instead of the Lite Salt.
- You may want to challenge the students to calculate the amount of heat energy that is released or absorbed during solution formation. Have students record the volume of water used to the nearest milliliter. Knowing that the specific heat capacity of water is 1.00 cal/g·°C and using the water volume and the temperature change, students can calculate the amount of heat released or absorbed.
- For older students, you may want to demonstrate the dissolution of lithium chloride (LiCl) in water. Dissolving 42 g LiCl in 50 mL distilled water raises the temperature of the resulting solution to above 65°C (149°F). Follow appropriate safety precautions, as the hot solution can cause severe burns. You could also challenge students to find a periodic pattern for Li, Na, and K and ask them to predict what would happen with cesium chloride (CsCl).

Explanation

Many common household materials produce noticeable energy changes when dissolved in water. In this activity, students observe the energy released or absorbed by a solid dissolving in liquid water to form a solution.

In general, a solution consists of one substance, the solute, dissolved in another, the solvent. The solute is generally the component present in the smallest amount. In this activity, a variety of household solids (the solutes) were dissolved in water (the solvent). The temperature changes observed result from a transfer of energy that occurs during the dissolving process, either into (endothermic) or out from (exothermic) the resulting solution.

When baking soda ($NaHCO_3$) dissolves in water, the temperature noticeably decreases. For $NaHCO_3$ to dissolve in water, energy (in the form of heat) is transferred from the surroundings, such as your fingers, the thermometer, the container, or the air, into the $NaHCO_3(aq)$ solution. Since heat is leaving the fingers, a cooling sensation is perceived. In this endothermic process, the resulting solution has gained energy, and the energy change is represented by a positive number. When washing soda (Na_2CO_3) dissolves in water, heat is transferred from the solution to the surroundings. This is an exothermic process and the energy change is represented by a negative number.

Part B calls for the use of Epsom salt ($MgSO_4{\cdot}7H_2O$) and anhydrous $MgSO_4$. In Epsom salt, the $MgSO_4$ has water molecules (called water of hydration) in a definite ratio of water to salt—every magnesium sulfate formula unit has seven water molecules arranged about it, yielding crystals that have a characteristic shape different from that of anhydrous $MgSO_4$. The anhydrous $MgSO_4$ is prepared by heating the $MgSO_4{\cdot}7H_2O$, which drives off the water of hydration as water vapor. At 150°C, six of the water molecules are released. At 200°C, the seventh water molecule is released, leaving the anhydrous $MgSO_4$. In this activity, the Epsom salt is heated at about 230°C for 30 minutes, which removes all of the water of hydration.

The activity compares the heat changes that occur when $MgSO_4{\cdot}7H_2O$ and anhydrous $MgSO_4$ dissolve in water. While the same aqueous ions exist in the resulting solutions, there is a significant difference in the temperatures of the resulting solutions. These two situations can be described in the form of equations that include heat as a product in the first case and heat as a reactant in the second case.

Dissolving anhydrous $MgSO_4$ in water:

$$MgSO_4(s) \xrightarrow{H_2O} Mg^{2+}(aq) + SO_4^{2-}(aq) + heat$$

Dissolving Epsom salt in water:

$$MgSO_4{\cdot}7H_2O(s) + heat \xrightarrow{H_2O} Mg^{2+}(aq) + SO_4^{2-}(aq) + 7H_2O(l)$$

The following discussion uses the term hydration in two contexts. First, when a solid ionic compound dissolves in water, the ions are separated from one another. This is because the polar water molecules shield the positive charge on the cation from the negative charge on the anion, allowing the ions to dissolve. These ions are described as hydrated, meaning water molecules surround each oppositely charged ion, which keeps the ions from re-crystallizing as the solid salt. The symbols for hydrated ions are often followed by the (aq) designation.

In Epsom salt crystals, "water of hydration" exists within the solid crystal structure; these water molecules are a definite part of the crystal, seven per each magnesium and sulfate ion. To understand the cause of the temperature differences observed when these two salts are dissolved, we have to be careful not to mix up "hydrated ions in solution" with the "water of hydration" in the crystal.

When anhydrous $MgSO_4$ dissolves in water, the resulting ions, $Mg^{2+}(aq)$ and $SO_4^{2-}(aq)$, are surrounded by water molecules, which allow these ions to stay separated in solution; this dissolving process releases energy in the form of heat. (The heat of solution, ΔH, is –84.9 kJ/mol.) When $MgSO_4{\cdot}7H_2O$ dissolves, once again the $Mg^{2+}(aq)$ and $SO_4^{2-}(aq)$ are surrounded by water molecules and energy is released. But another factor is involved—energy is required in order to overcome the attractive forces between the water of hydration and the Mg^{2+} and SO_4^{2-} ions in the Epsom salt crystal. The evidence for the overall endothermic nature of the dissolution of Epsom salt (ΔH = +16.2kJ/mol) is the cooling of the resulting solution. This is because more energy must be taken in from the surroundings to accomplish the loss of the water of hydration from the crystalline structure than the energy that is released by the formation of hydrated ions in solution.

In Part C, table salt is mostly sodium chloride (NaCl). The container label of the Lite Salt lists NaCl as the first ingredient and potassium chloride (KCl) as the second. The heat of solution of KCl is more endothermic (+17.24 kJ/mole) than the heat of solution of NaCl (+3.9kJ/mole). Thus, a slightly larger cooling effect is expected for the Lite Salt.

If you demonstrated the dissolution of lithium chloride (LiCl), you noted that the heat of solution is highly exothermic (ΔH = –37.1kJ/mole). The difference between the heat of solution of NaCl, KCl, and LiCl is due to the smaller size of Li^+ compared to Na^+ or K^+.

Answers to Student Questions

Part A

Step 1

Evidence of change is the temperature decreases (energy is absorbed) and the amount of solid baking soda decreases.

Step 2

a. *At least some of each solid dissolves in both cases; with the detergent, a temperature increase occurs (energy is released).*
b. *As different solids dissolve in water, different amounts of heat may be absorbed or released. Whether heat is released or absorbed depends on the identity of the solid. For example, dissolving baking soda absorbs more heat than dissolving table salt. Dissolving washing soda releases heat.*

Part B

Step 2

a. *A temperature increase occurs when the anhydrous $MgSO_4$ dissolves, and a temperature decrease occurs when Epsom salt dissolves. Due to the presence of water of hydration in the Epsom salt, it has a different heat of solution than anhydrous $MgSO_4$.*

b. *When anhydrous $MgSO_4$ dissolves, $Mg^{2+}(aq)$ and $SO_4^{2-}(aq)$ are present. (Note that these are hydrated ions.)*

c. *The same species are present as when anhydrous $MgSO_4$ dissolves.*

Part C

Step 1

For table salt, a decrease of about 2°C is typically observed.

Step 2

For Lite Salt, a decrease of about 8°C is typically observed. The Lite Salt behaves differently from the table salt, which is sodium chloride (NaCl), because in addition to NaCl, the Lite Salt also contains potassium chloride (KCl). In addition, students can infer that the K^+ is responsible for the greater degree of cooling.

References

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