How Can Energy Change?

In our everyday experience, energy seems to disappear all the time. Batteries go dead, a wind-up toy stops moving, fires run out of fuel. But the reality is that energy is never destroyed, it only changes from one form to another. In this section we present a number of examples of how energy transfers account for where energy goes in chemical and physical reactions. We also present a career profile that profiles a chemist who works for an energy company.

- Lab Investigation—The Energy of Evaporation
- Lab Investigation—The Energy Efficiency of Heating Water
- Lab Investigation—Exothermic, Endothermic, & Chemical Change
- Video—Exothermic & Endothermic Reactions
- Video—Meet a BP chemist

The Energy of Evaporation | A Lab Investigation

Summary

In this investigation, students test evaporation rates for different liquids. Next, students use a thermometer to measure the temperature change during evaporation.

Objective

Students will explore the energy change associated with evaporation and the differences in evaporation rates of different liquids.

Safety

- Be sure you and the students wear properly fitting goggles.
- Isopropyl alcohol and acetone are flammable and should be handled with care. Avoid flames or sparks, and work in a well-ventilated area. Avoid body tissue contact.
- Do not substitute plastic cups for glass beakers in this experiment, as acetone will dissolve some plastic cups.

Materials for Each Group

- 4 student thermometers
- 5 paper towels
- 4 small rubber bands
- Tape or sticky labels
- 3 droppers
- Water
- 99% isopropyl alcohol
- Acetone
- 3 small glass beakers (50–150 mL)
- 3 graduated cylinders (10–100 mL)

Time Required

One class period, approximately 45–50 minutes.

Lab Tips

Because water can take a long time to evaporate, you may wish to have students move on to the second part of the investigation after recording evaporation times for acetone and isopropyl alcohol. Note that evaporation times may vary according to humidity and air currents.

Integrating into the Curriculum

This investigation could be incorporated into a unit on phase changes, chemical and physical changes, and energy.

PREPARING TO INVESTIGATE

In this activity, you will explore the energy change that accompanies the process of evaporation. Evaporation, like melting or freezing, is an example of a *phase change*—a change from one physical form of a substance to another.

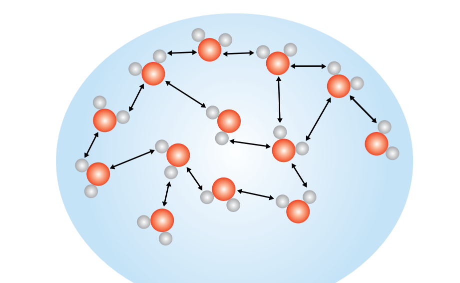
During evaporation, energetic molecules leave the liquid phase, which lowers the average energy of the remaining liquid molecules. The remaining liquid molecules can then absorb energy from their surroundings. This process can take place at any temperature because some of the molecules in a liquid will *always* have enough energy to enter the gas phase.

Phase changes release or use energy because they bring particles closer together or cause them to move farther apart. To understand why these processes release or use energy, recall that all atoms or ions have at least some attraction for one another. Overcoming these attractions, as particles move farther apart, requires energy. When particles come back together, energy is released.

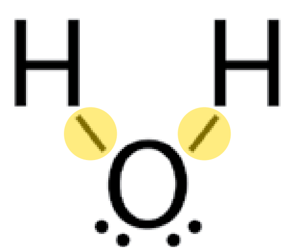
But are all attractions the same?

Intermolecular vs. Intramolecular Forces

During a phase change, the attractive forces between *whole molecules* are disrupted or restored. These are called *intermolecular* forces. Conversely, during a chemical change, the bonds between *atoms within* a molecule or ion are disrupted or restored. These are called *intramolecular* forces.



Intermolecular forces
Hydrogen bonding between water molecules



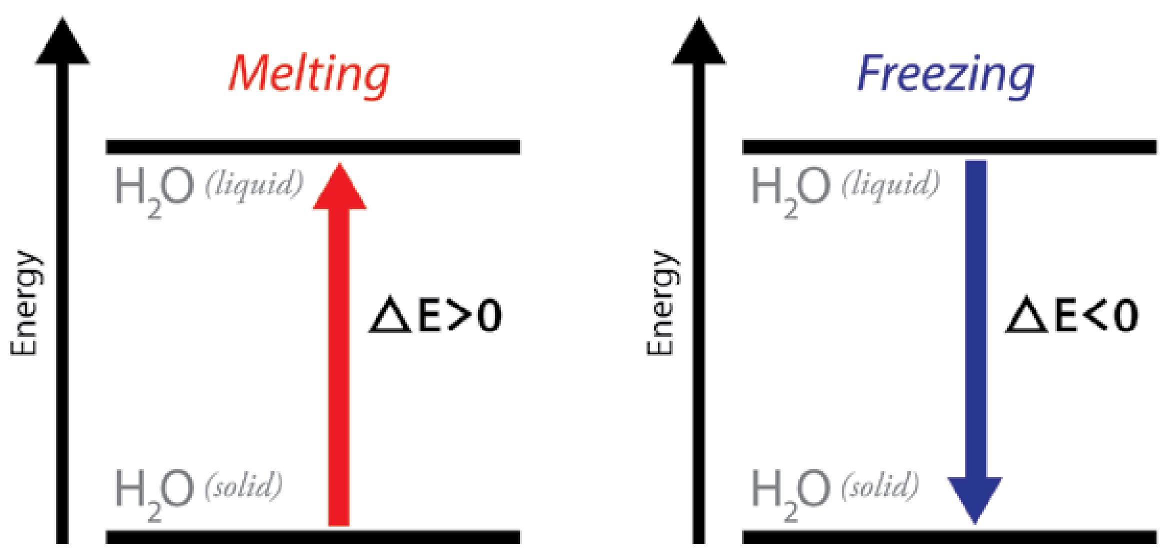
Intramolecular forces
Covalent bonds between hydrogen and oxygen

The relative strength of the intermolecular forces of a substance determines how much energy is required for two molecules of that substance to move further apart. Substances with very strong intermolecular forces (like water) require a comparatively greater amount of energy to separate.

For example, a great deal of energy is needed to convert liquid water to water vapor because water molecules have a particularly strong form of intermolecular attraction called hydrogen bonding. The energy added to the liquid water has to be enough to overcome the attraction that individual water molecules have for one another. The molecules of other liquids, however, may not be as strongly attracted to one another as water molecules are, and therefore require less energy to vaporize.

Energy Diagrams

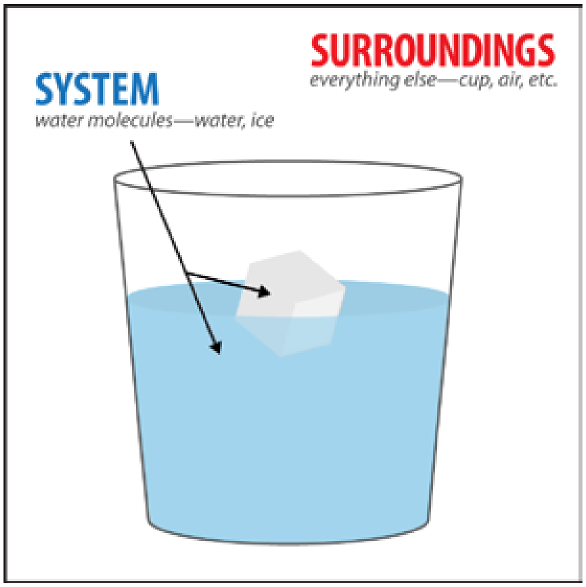
You can visualize the energy implications of phase changes using energy diagrams like the ones shown below. The horizontal lines represent the energies of substances in particular states. The higher horizontal line represents a substance in a physical state at a higher energy level, and the lower horizontal line represents a substance in a physical state at a lower energy level. Energy diagrams can help us keep track of energy *differences* even when we don't have numeric values. The diagram on the left shows energy being *absorbed by* water molecules as it would be when ice melts to form liquid water. The energy diagram on the right shows energy being *released from* water molecules as it would be when liquid water freezes to form ice.



System and Surroundings

But where does the energy needed to melt ice come from? Where does the energy released as ice is formed go? Because we can't create or destroy energy, we know that the *surroundings* play an important role in energy changes.

Energy may be absorbed from the surroundings to provide the energy needed to melt ice, or the surroundings may receive the energy released when water freezes. In the examples above, the surroundings can be defined as everything other than the water molecules involved in these phase changes. The water molecules are the *system*—the molecules we are studying. If these phase changes were taking place in a cup, even the cup would be considered the surroundings. So too would the air, the table the cup is sitting on, and even the rest of universe. Because we're studying water molecules in this phase change, only the water molecules are part of the *system*.



Considering the surroundings also helps us to make sense of the sign convention for energy changes. When energy is *absorbed by the system*, ΔE is positive (+). When energy is *released from the system*, ΔE is negative (–).

In this investigation you will explore the energy change associated with evaporation and the differences in evaporation rates of different liquids.

GATHERING EVIDENCE

Evaporation Test

1. Fold a paper towel into thirds. When unfolded, the fold lines divide the towel into equal sections.
2. Label the 3 sections of the paper towel "water," "isopropyl alcohol," and "acetone." The sections should be at least 3–4 inches apart.
3. Add 3 drops each of water, isopropyl alcohol, and acetone to the corresponding section of the paper towel.
4. Record the evaporation time in minutes for each liquid in Table 1 below.

Energy Change During Evaporation

1. Add 20-mL samples of water, isopropyl alcohol, and acetone into three small glass beakers labeled "water," "isopropyl alcohol," and "acetone."
2. Fold a paper towel along the long edge several times until it is 1 inch wide.
3. Wrap the folded paper towel around the base of a thermometer, ensuring that it covers the bulb at the bottom, but doesn't obscure numbers higher than 0 °C. Fasten the folded paper towel around the thermometer with a small rubber band.

4. Prepare a total of four thermometers according to the procedure detailed in step 3. Label the 4 thermometers "water," "isopropyl alcohol," "acetone," and "control" using tape or sticky labels.
5. Record the initial temperature (T_i) for each thermometer in Table 2 below.
6. With the help of a lab partner, dip the 3 thermometers labeled "water," "isopropyl alcohol," and "acetone" into the beakers containing the corresponding liquids so that the paper towels are moist but not dripping wet. Do not dip the control apparatus into any liquid.
7. With a thermometer in each hand, gently swing the thermometer back and forth using a gentle pendulum motion for 2 minutes. Coordinate with your lab partner to use the same motion and intensity.
8. After 2 minutes, record the final temperature (T_f) in Table 2.

Table 1. Evaporation test results

Liquid	Water	Isopropyl alcohol	Acetone
Time (minutes)			

Table 2. Energy change during evaporation results

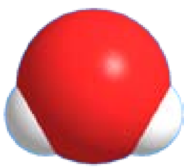
Liquid	T_i	T_f	ΔT	ΔE (+/−)
1. Water				
2. Isopropyl alcohol				
3. Acetone				
4. Control				

ANALYZING EVIDENCE

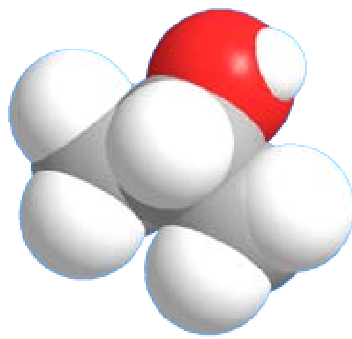
Complete Table 2 using your observations from the experiment. Enter either a plus (+) or minus (–) sign in the last column of the table for the change in energy of the liquid-dampened paper towel in each case.

INTERPRETING EVIDENCE

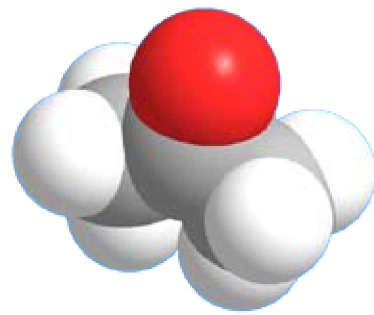
1. Using the space-filling models below as a guide, write the chemical formulas for (1) water, (2) isopropyl alcohol, and (3) acetone.



Water



Isopropyl alcohol



Acetone

2. Using the language of intermolecular forces, explain the order of the evaporation rates you observed in the first part of your experiment.
3. What differences in intermolecular forces might explain the differences in the time it takes water, isopropyl alcohol, and acetone to evaporate?
4. Using the language of intermolecular forces and energy, explain why you observed the temperature changes that you did in the second part of your experiment. Refer to the energy diagrams in Preparing to Investigate to review sign conventions for phase changes.
5. Draw and label energy diagrams for the evaporation of equal amounts (same number of moles) of water, acetone, and isopropyl alcohol. How are the energy diagrams different from one another?

REFLECTING ON THE INVESTIGATION

1. Is evaporation a cooling process or a heating process? Explain your answer in terms of energy and intermolecular forces.
2. Given what you've learned in this investigation, why do you think people sweat?
3. In this investigation, you have seen that evaporation is a cooling process. Condensation is the opposite of evaporation—water in the vapor phase condenses to form liquid water. Using the language of energy and intermolecular attraction, explain why condensation is considered a warming process. Draw an energy diagram for the process of condensation.

TEACHER'S KEY

Analyzing Evidence

1. Complete Table 2 using your observations from the experiment. Enter either a plus (+) or minus (–) sign in the last column of the table for the change in energy of the liquid dampened paper towel in each case.

Table 1. Evaporation test results

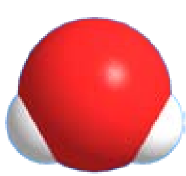
Liquid	Water	Isopropyl alcohol	Acetone
Time (minutes)	10:00	3:00	0:52

Table 2. Energy change during evaporation results

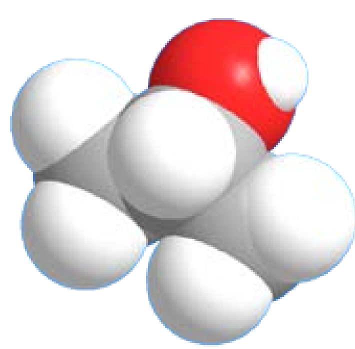
Liquid	T_i	T_f	ΔT	ΔE (+/–)
1. Water	24 °C	18 °C	–6 °C	–
2. Isopropyl alcohol	24 °C	15 °C	–11 °C	–
3. Acetone	24 °C	6 °C	–18 °C	–
4. Control	24 °C	24 °C	0 °C	N/A

Interpreting Evidence

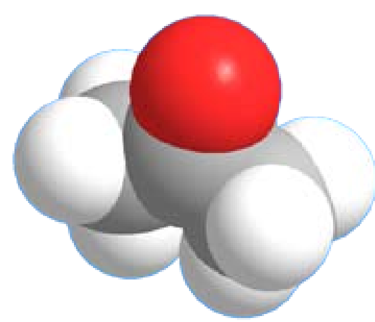
1. Using the space-filling models below as a guide, write the chemical formulas for (1) water, (2) isopropyl alcohol, and (3) acetone.



Water—H_2O
HOH



Isopropyl alcohol—C_3H_8O
$(CH_3)_2CHOH$



Acetone—C_3H_6O
$(CH_3)_2CO$

2. Using the language of intermolecular forces, explain the order of the evaporation rates you observed in the first part of your experiment.

   Acetone has the weakest intermolecular forces, so it evaporated most quickly. Water had the strongest intermolecular forces and evaporated most slowly. The strength of the intermolecular forces in isopropyl alcohol are in between water and acetone, but probably closer to acetone because the water took much longer to evaporate.

3. What differences in intermolecular forces might explain the differences in the time it takes water, isopropyl alcohol, and acetone to evaporate?

   Water evaporates most slowly because its molecules are attracted to one another by hydrogen bonding. Acetone does not participate in hydrogen bonding, so its intermolecular forces are comparatively weaker, and it evaporates most quickly. Isopropyl alcohol can also participate in hydrogen bonding, but not as successfully as water because it has a non-polar region, so it evaporates at an intermediate rate.

4. Using the language of intermolecular forces and energy, explain why you observed the temperature changes that you did in the second part of your experiment.

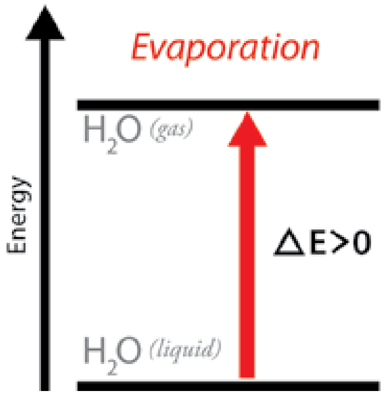
   The temperature decreased for each of the liquid-soaked paper towels in the experiment because evaporation is a phase change that absorbs energy.

   During evaporation, some fast-moving, highly energetic molecules have enough energy to overcome the attractions that individual molecules have for one another and enter the gas phase. As these high-energy molecules leave the liquid phase, the average energy of the remaining liquid molecules is lowered and the temperature goes down. Because the liquid is now at a lower temperature than its surroundings, it continues to absorb energy from its surroundings to keep evaporation going.

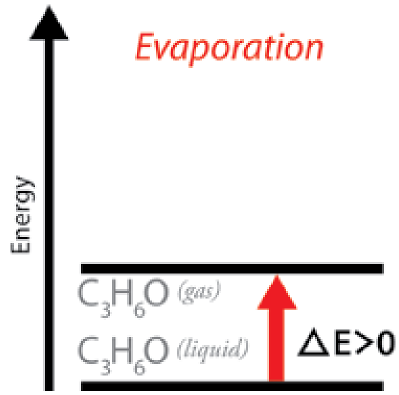
   We observe energy being absorbed by a liquid as it evaporates as a decrease in temperature of the liquid and in its surroundings as they transfer energy to the liquid.

5. Draw and label energy diagrams for the evaporation of equal amounts (same number of moles) of water, acetone, and isopropyl alcohol. How are the energy diagrams different from one another?

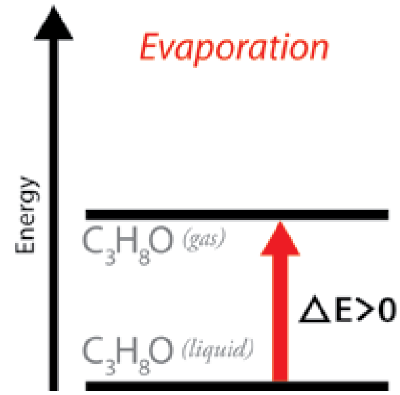
   These diagrams show the relative energy levels for the same number of moles of each liquid and its vapor at the same temperature. All of the diagrams show that the process of evaporation absorbs energy. Water absorbs the most energy, acetone absorbs the least energy, and isopropyl alcohol is somewhere in between.



Water (H_2O)
HOH



Isopropyl alcohol
(C_3H_8O)



Acetone
(C_3H_6O)

Reflecting on the Investigation

1. Is evaporation a cooling process or a heating process? Explain your answer in terms of energy and intermolecular forces.

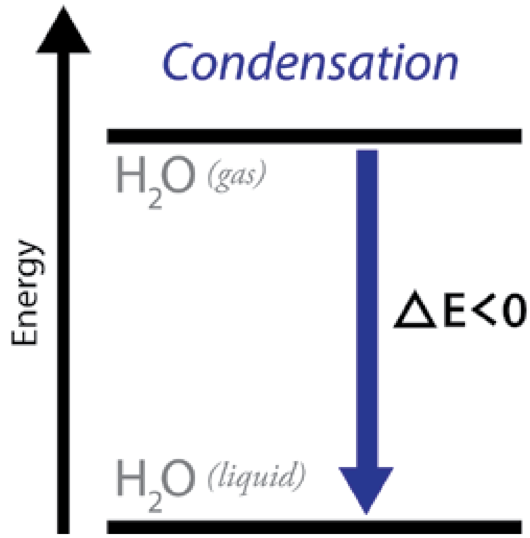
   Evaporation is a cooling process. In the liquid state, molecules are closer together than they are in the gas state. Energy is required to separate these molecules of a liquid as they become farther apart in the gas phase. The amount of energy absorbed by the liquid must be enough to overcome the attractions that the molecules of the liquid have for each other.

2. Given what you've learned in this investigation, why do you think people sweat?

   People sweat as a means of temperature regulation. Sweating keeps us cool. As our sweat evaporates, it absorbs energy from our skin, which cools us down. Sweating is a form of "evaporative cooling."

3. In this investigation, you have seen that evaporation is a cooling process. Condensation is the opposite of evaporation—water in the vapor phase condenses to form liquid water. Using the language of energy and intermolecular attraction, explain why condensation is considered a warming process. Draw an energy diagram for the process of condensation.

   Condensation is a warming process because energy is released as molecules in the gas phase come closer together in the liquid phase. Because molecules are attracted to one another, they are at lower energy when they are closer together (as they are in the liquid phase) than when they are further apart (in the gas phase). This transition from high energy to low energy releases heat to the surroundings.



Post-Lab Discussion

This investigation can also be helpful to illustrate the relationship between temperature and energy. Consider showing the video "Temperature and Energy" from the *Energy Foundations* collection.

Extensions

To more fully explore phase changes in water, you may consider pairing this investigation with constructing a heating curve for water. Several different procedures are available online:

- *http://bit.ly/highschoolnrg11*
- *http://bit.ly/highschoolnrg12*

Additional Resources

"Water: A Natural Wonder," *Chemistry: A Project of the American Chemical Society*. New York: W. H. Freeman and Company, 2005, pp 1–71.

The Energy Efficiency of Heating Water | A Lab Investigation

Summary

Students use a Bunsen burner, microwave oven, and hot plate to determine which instrument heats water most efficiently. Students perform detailed calculations to support their conclusions.

Objective

Students will use observations to calculate the energy efficiency of various heating methods.

Safety

- Be sure you and the students wear properly fitting goggles.
- Review safe use of Bunsen burners or hot plates. Caution students about handling hot water to avoid scalds or burns.
- Methane (natural gas) is flammable. It can be explosive if mixed with air in certain proportions. Avoid any sparks or flames when collecting the gas. Methane is toxic by inhalation. Work in a well-ventilated area.

Materials for Each Group

- 1 Bunsen burner
- 1 hot plate
- 1 one-meter hose sections of tubing
- 1 tub or bucket for holding water
- 1 stopwatch or clock with a second hand
- 1 2-L soda bottle (clear with label removed)
- 1 400-mL beaker
- 1 alcohol thermometer
- 1 ring stand support with ring
- 1 pair of beaker tongs

Materials for the Whole Class

- 1 microwave oven (shared by all groups)
- A variety of volumetric containers, such as large graduated cylinders for each group

Time Required

Two to three class periods, approximately 45–50 minutes each.

Lab Tips

It may be difficult for some students to come up with a satisfactory scheme for collecting and measuring the amount of methane collected from the gas outlet. Lend appropriate hints and tips as required. If your lab doesn't have a typical standard gas system piped in permanently, consider using portable gas burners, alcohol lamps, or whatever other alternative system you generally use. Also, consider using portable electric immersion heaters in the place of hot plates. They are much more efficient and at $10 per heater, much less expensive.

If presented as a "lab challenge" this basic investigation could be adapted as a lab practical exam to test how students are able to apply what they have learned. You could vary the complexity of the task by adjusting the amount of information you provide (such as the enthalpy of combustion for methane). This investigation also provides an opportunity for considering the total environmental cost involved in a simple lab procedure.

Pre-Lab Discussion

Make sure students are familiar with the proper operation of Bunsen burners, hot plates, and microwaves. You may direct students to read the procedure and make up appropriate data tables before they go to the lab. Their data tables can be their "passport" to begin the investigation.

Incorporating into the Curriculum

This investigation could be incorporated into a unit on stoichometry, chemical changes, or thermochemistry.

PREPARING TO INVESTIGATE

Taking into account the energy requirements for a process and finding ways to minimize the energy required are important ways to make the process as environmentally friendly or "green" as possible.

One way to do this is to make the process as efficient as possible. But what do we mean when we say efficient? Usually it means the fraction or percent of all the energy that goes into a process that actually is used for the desired effect; the rest is "lost." Although we say "lost," we know that energy is never lost, it just moves out of the system we are interested in or gets converted into a less useful form.

Consider a simple example that does involve energy. Suppose our process is to catch the water leaking from a pipe in our house. If 1000 mL of water leaks and our container holds only 650 mL, we would say our catchment process is 65% efficient. That is, because

$$\left(\frac{650\text{ mL}}{1000\text{ mL}}\right)100\%$$

was caught in the container. The other 35% was "lost." Of course, this 35% wasn't really lost either, it is on the floor under the leak, but it was "lost" to our system (the container).

In this lab you will consider the energy efficiency of heating water in a typical high school laboratory setting. You will consider three ways this is typically done and calculate the efficiency of each method.

GATHERING EVIDENCE

Part A—Heating with a Bunsen Burner

One common way to heat water in the laboratory is to use a Bunsen burner. The typical setup is to use natural gas that is piped into the classroom from a commercial supplier and mixed with air in the burner to produce a flame. Mixing with air ensures a more complete combustion. Natural gas is composed primarily of methane (CH_4). The complete combustion of methane is represented in the following equation:

$$CH_4\,(g) + 2\,O_2\,(g) \longrightarrow 2\,H_2O\,(g) + CO_2\,(g)$$

This reaction is very exothermic (gives off a lot of thermal energy).

Of course, not all of the energy released by this reaction is actually absorbed by the material being heated. Some of the energy goes into warming the container or the immediate surroundings. The efficiency of the process can be calculated as follows:

How efficient is it to heat a sample of water using a Bunsen burner?

1. Using the balanced equation above, calculate the amount of energy (ΔH) for the combustion of methane. (Assume we are at STP.)

$$\% \, Efficiency = \left(\frac{\text{energy gained by water}}{\text{energy produced by combustion}} \right) 100\%$$

Table 1. Standard enthalpies of formation, ΔH°f, at 25 °C (kJ/mol)

Methane (CH_4)	−74.85	Water (H_2O)	−241.8
Oxygen (O_2)	0	Carbon dioxide (CO_2)	−393.5

The general equation for finding the standard enthalpy change of a chemical reaction is:

$$\Delta H^\circ rxn = \Sigma \, n\Delta H^\circ_f \,(\text{products}) - \Sigma \, n\Delta H^\circ_f \,(\text{reactants})$$

2. Devise a method to determine the rate (L/s) natural gas is delivered from a fully opened gas valve connection on your desktop, through the Bunsen burner. You will have access to the following equipment: 1-meter hose, a tub or bucket for holding water, a stopwatch or clock with a second hand, a 2-L soda bottle, and a variety of volumetric containers, such as large graduated cylinders.
3. Write a plan, paying particular attention to safety, including how you will safely dispose of the natural gas you are measuring. When you have documented the plan, show it to your teacher and get approval before proceeding. Be sure to record your results during the procedure for use in completing this activity.
4. Fill a beaker with a carefully measured (+ or – 1 mL) amount of tap water (somewhere between 175 and 225 mL). Set the beaker on a ring stand or support suitable for heating it with a Bunsen burner. Measure the initial temperature of the water with a thermometer to the nearest 0.1 °C, and begin to heat it with the Bunsen burner. Be sure the stopcock of the gas fitting is wide open, as before. Start tracking the time.
5. Heat the water until the temperature rises by 30–50 °C. Measure the final temperature to the nearest 0.1 °C. Note the time elapsed during the heating.

Part B—Heating with an Electric Hot Plate

Another typical means of heating in the laboratory is to use an electric hot plate. In this part you will do an experiment similar to the one described in Part A, but this time you will determine the efficiency of using an electric hot plate to heat water.

1. In this case, the energy released by the hot plate depends on its energy rating. Look at the bottom or sides of the hot plate for its power rating, measured in watts.
2. Since a watt = 1 J/s , we can calculate the total amount of energy released from the hot plate using the following equation:

$$\textit{Energy released by the hot plate (J)} = \textit{total hot plate wattage (watts)} \times \textit{time (s)} \times \left(\frac{1\ J/s}{watt}\right)$$

3. Using a procedure similar to the activity above, heat a sample of water similar in size to the amount used in part A, and use the results to calculate the efficiency of the electric hot plate for heating water. Be sure to use the hot plate on the highest setting to be sure it is operating at its highest power output.

Part C—Heating with a Microwave Oven

Microwave ovens have long been used in homes, but now they are being used more widely in scientific laboratories. Use the experience and information in the previous two investigations to determine the efficiency of heating a sample of water with a microwave oven.

1. Look at the bottom, sides, or elsewhere on the microwave oven for its power rating, measured in watts. This information is usually on a small sticker or plate.
2. Since a watt = 1 J/s , we can calculate the total amount of energy released from the microwave using the following equation:

$$\textit{Energy released by the microwave (J)} = \textit{total microwave wattage (watts)} \times \textit{time (s)} \times \left(\frac{1\ J/s}{watt}\right)$$

3. Using a procedure similar to the activity above, heat a sample of water and use the results to calculate the efficiency of the microwave oven for heating water. Be sure to use the microwave oven on the highest setting to be sure it is operating at its highest power output. It is best to not leave the thermometer in the microwave while it is operating. Stop the microwave and take the temperature at intervals until the desired temperature change is achieved.

ANALYZING EVIDENCE

Analyzing Part A

Calculate the amount of heat absorbed by the water, the amount of heat released by the burning natural gas, and the percent efficiency of the heating process. Use the following equations in your calculations (assume STP).

The equation for determining the amount of thermal energy absorbed by a substance (where a change in state is not involved) is given by the equation:

$$\textit{Energy absorbed} = \left(\begin{matrix}\textit{mass of the}\\\textit{substance}\end{matrix}\right) \times \left(\begin{matrix}\textit{the change in}\\\textit{temperature of the}\\\textit{substance}\end{matrix}\right) \times \left(\begin{matrix}\textit{the specific heat}\\\textit{capacity of the}\\\textit{substance}\end{matrix}\right)$$

In our investigation, the energy absorbed by the water is given by the equation:

$$\textit{Energy absorbed by water (J)} = \textit{mass of the water (g)} \times \Delta T\ (^\circ C) \times \left(\frac{4.18\ J}{g \cdot {}^\circ C}\right)$$

$$\begin{matrix}\textit{Energy released}\\\textit{by burning}\\\textit{natural gas (J)}\end{matrix} = \textit{time (s)} \times \left(\frac{\textit{vol gas (L)}}{\textit{time (s)}}\right) \times \left(\frac{\textit{1 mol gas}}{\textit{22.4 L}}\right) \times \left(\frac{\Delta H^\circ_{rxn}\ (J)}{\textit{1 mol gas}}\right)$$

$$\%\ \textit{Efficiency} = \left(\frac{\textit{energy absorbed by water (J)}}{\textit{energy released by burning natural gas (J)}}\right)100\%$$

Analyzing Part B

$$\textit{Energy absorbed by water (J)} = \textit{mass of the water (g)} \times \Delta T\ (^\circ C) \times \left(\frac{4.18\ J}{g \cdot {}^\circ C}\right)$$

$$\textit{Energy released by the hot plate (J)} = \textit{total hot plate wattage (watts)} \times \textit{time (s)} \times \left(\frac{1\ J/s}{\textit{watt}}\right)$$

$$\%\ \textit{Efficiency} = \left(\frac{\textit{energy absorbed by water (J)}}{\textit{energy released by hot plate (J)}}\right)100\%$$

Analyzing Part C

$$\text{Energy absorbed by water (J)} = \text{mass of the water (g)} \times \Delta T\ (^\circ C) \times \left(\frac{4.18\ J}{g \cdot ^\circ C}\right)$$

$$\text{Energy released by the microwave (J)} = \text{total microwave wattage (watts)} \times \text{time (s)} \times \left(\frac{1\ J/s}{watt}\right)$$

$$\%\ \text{Efficiency} = \left(\frac{\text{energy absorbed by water (J)}}{\text{energy released by microwave (J)}}\right) 100\%$$

INTERPRETING EVIDENCE

1. When you have calculated the efficiency for heating by each method, record your results, along with the rest of the class, on the board in the front of the classroom, or as your teacher directs. Calculate the average efficiency for each of the methods and discuss the precision in the range of results of the data from the various lab groups.

   Discuss any result that differs significantly from the class average, and if time allows, repeat the experiment to improve the precision.

Cross Link
Students can get more practice with calculations like these in *Preparation & Combustion of Biodiesel* on page 43.

2. Compare the results of the three methods of heating. Which was the most efficient, and which was the least efficient?

REFLECTING ON THE INVESTIGATION

The data you gathered helps determine the efficiency of the various types of heating, but does not reveal the total cost of heating. Depending on how it is generated, the power plants that produce electricity may contribute significantly more (or less) pollution in the process than does the production of natural gas.

The total cost of energy should reflect not only the cost of production, but also the indirect costs associated with damage to the environment, regulations, and other factors.

1. Use your household electricity and natural gas bills to calculate the cost in dollars for each part of this investigation. Calculate the cost for heating a 200-g sample of water by 10 °C

for each type of heating. If natural gas is not available in your community, use the national average price or substitute the cost of propane gas (an alternative fuel to natural gas).

This cost in dollars is rarely a reflection of the total cost to the environment, since it only reflects the net cost (after government subsidies) to consumers for the energy. Which energy source was the most expensive? Which was the least expensive?

2. There are many ways to generate electricity and produce natural gas. Some require less energy to produce, and some give off less pollution. Using the Internet and other resources, investigate which power source tends to require less energy to produce and contributes less total pollution to the environment.

3. Using the information you collected regarding efficiency, dollar costs, and environmental costs, make a recommendation for how best to minimize the energy used to heat substances in your school laboratory. Explain the reasoning behind your recommendation.

TEACHER'S KEY

Analyzing Evidence

Sample data:

200 g water, $\Delta T = 50.5\,°C$ (16.0 °C to 66.5 °C)

Combustion of 4.86 L of methane, $\Delta H = -802.3$ kJ/mol CH_4

698-watt electric hot plate

1000-watt microwave

Electricity: 10.56¢ per kilowatt-hour

Natural gas: $11.80 per thousand cubic feet

Analyzing Part A

From sample data

$$\text{Energy absorbed by water (J)} = 200\text{ g} \times 50.5\,°\text{C}) \times \left(\frac{4.18\text{ J}}{\text{g}\cdot°\text{C}}\right) = 42.2\times10^3\text{ J}$$

$$\text{Energy released by burning natural gas (J)} = 245\text{ s} \times 0.0198\text{ L/s} \times \left(\frac{1\text{ mol gas}}{22.4\text{ L}}\right) \times \left(\frac{802\times10^3}{1\text{ mol gas}}\right) = 174\times10^3\text{ J}$$

$$\%\text{ Efficiency} = \left(\frac{42.2\times10^3\text{ J}}{174\times10^3\text{ J}}\right)100\% = 24\%$$

Analyzing Part B

$$\text{Energy absorbed by water (J)} = 200\text{ g} \times 50.5\,°\text{C}) \times \left(\frac{4.18\text{ J}}{\text{g}\cdot°\text{C}}\right) = 42.2\times10^3\text{ J}$$

$$\text{Energy released by hot plate (J)} = 698\text{ watt} \times 378\text{s} \times \left(\frac{1\text{ J/s}}{\text{watt}}\right) = 264\times10^3\text{ J}$$

$$\%\text{ Efficiency} = \left(\frac{42.2\times10^3\text{ J}}{264\times10^3\text{ J}}\right)100\% = 16\%$$

Analyzing Part C

$$\textit{Energy absorbed by water (J)} = 200\,g \times 50.5\,°C) \times \left(\frac{4.18\,J}{g \cdot °C}\right) = 42.2 \times 10^3\,J$$

$$\textit{Energy released by microwave (J)} = 698\,watt \times 378s \times \left(\frac{1\,J/s}{watt}\right) = 264 \times 10^3\,J$$

$$\%\ \textit{Efficiency} = \left(\frac{42.2 \times 10^3\,J}{264 \times 10^3\,J}\right) = 16\%$$

Sample Data

Heat source	Energy released (kJ)	Energy absorbed (kJ)	Efficiency	Time (sec)	Cost (¢)	Savings*
Bunsen burner	174	42.2	24%	245	0.203	0.571¢
Hot plate	264	42.2	16%	378	0.774	-
Microwave	62	42.2	68%	62	0.181	0.593¢

**as compared to the most expensive method tested.*

Interpreting Evidence

1. When you have calculated the efficiency for heating by each method, record your results along with the rest of the class on the board in the front of the classroom, or as your teacher directs. Calculate the average efficiency for each of the methods and discuss the precision in the range of results of the data from the various lab groups.

   Discuss any result that differs significantly from the class average, and if time allows, repeat the experiment to improve the precision.

   This investigation typically yields consistent results. The discussion of results that differ from the average can lead to an appreciation of care and accuracy in doing lab investigations. Consider having students redo their investigation after a discussion of ways to control variability.

2. Compare the results of the three methods of heating. Which was the most efficient, and which was the least efficient? Did the results follow what you expected?

   The microwave method is likely to be the most efficient method of heating, followed by the Bunsen burner and then the hot plate. This is reasonable because the microwaves heat by directly increasing the motion of the water molecules, while the other methods have a more indirect transfer, which can be subject to inefficient heating.

Reflecting on the Investigation

1. Use your household electricity and natural gas bills to calculate the cost in dollars for each part of this investigation. Calculate the cost for heating a sample of water by 10 °C for each type of heating. If natural gas is not available in your community, use the national average price or substitute the cost of propane gas (an alternative fuel to natural gas).

   Although the cost in dollars is not a perfect reflection of the total cost to the environment., it does reflect how difficult it is to bring the energy source to market. Which energy source was the most expensive? Which was the least expensive?

   Using the data from the experiment above, we can divide the cost to raise the temperature of 200 g of water by 50.5 °C to get the cost per degree. Multiplying this result by 10 °C gives the desired cost for the 10 °C rise. It is about equal in expense to heat with cheaper natural gas in an inefficient Bunsen burner, as it is to use more expensive electricity in a relatively efficient microwave oven. The inefficient hot plate was much more expensive.

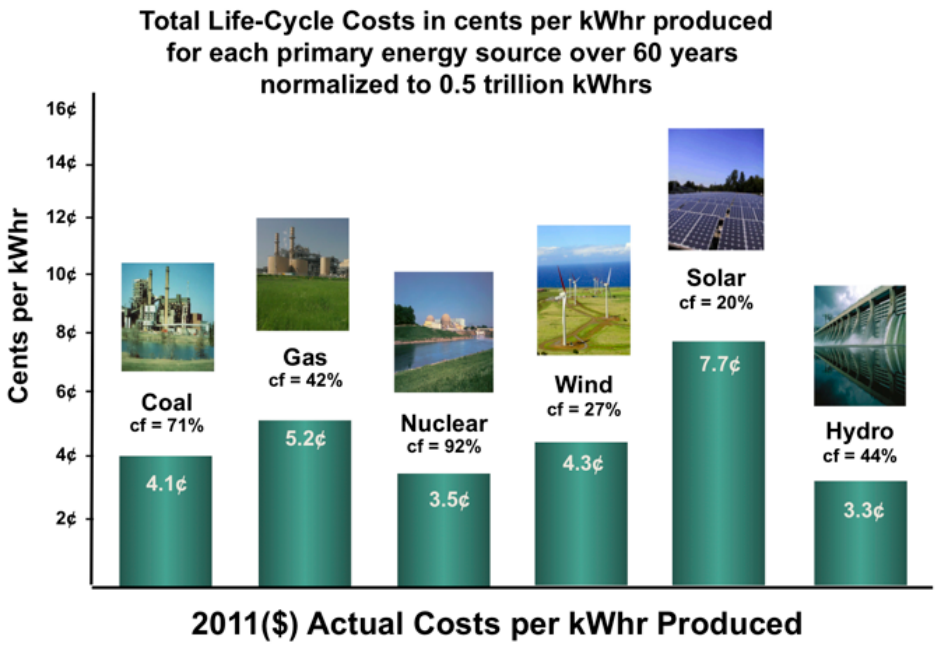
Heat source	Exp ΔT	Cost (¢)	Cost/10 °C
Bunsen burner	50.5 °C	0.203¢	0.040¢
Hot plate	50.5 °C	0.77¢	0.153¢
Microwave	50.5 °C	0.181¢	0.035¢

2. There are many ways to generate electricity and produce natural gas. Some require less energy to produce, and some give off less pollution. Using the Internet and other resources, investigate which power source tends to require less energy to produce and contributes less total pollution to the environment.

   This is obviously a huge question. There is an interesting discussion of this issue at the Canadian website www.iclei.org.

   "All energy use has some negative impact on the environment. Burning fossil fuels such as coal and oil produces emissions of greenhouse and acid gases, which result in global warming and acid rain, respectively. Fossil fuels are also responsible for urban air pollution and its associated health hazards. Nuclear power plants expose the environment to low levels of radiation during many stages of the nuclear fuel cycle, and also pose risks of nuclear accidents. Storage of spent nuclear fuel continues to be a problem. Even renewable technologies using energy from the sun have some negative impacts on the environment. Hydroelectric dams, for example, can flood vast areas, and damage aquatic ecosystems."

Below is an example of one analysis of total cost for energy production. Note: in this analysis "costs do not include electrical grid upgrade, transportation issues, connectivity of renewables, and buffering of their intermittency by rapid cycling of fossil fuel plants as presently practiced in this country, and externalities such as any carbon-tax, pollution, and health care costs associated with energy production and use. Also, these costs are not levelized but are actual direct costs."



Source: http://forbes.com

3. Using the information you collected regarding efficiency, dollar costs, and environmental costs, make a recommendation for how best to minimize the energy used to heat substances in your school laboratory.

   Students can make suggestions based on your lab results. Students will likely suggest the most efficient heating method that was the lowest cost. A more advanced response will include the "total cost" of energy sources.

Post-Lab Discussion

- Based on the above data, if you lowered the temperature of your 50-gallon hot water tank from 99 °F to 90 °F you would save approximately 20% on the cost of heating your water.
- Heating of larger volumes of water (1 L) with a microwave has been found to be more efficient, about 80% compared to the 68% found for 200 mL of water.
- Heating 1-L volumes with a source that submerses the heating coils in the water, like an electric kettle, has about 90% efficiency.
- The extremely low efficiency of the hot plate is most likely due to the difference in area of the hot plate and the beaker of water. A large percentage of the heat provided by the hot plate is released into the air rather than absorbed by the water.

Extensions

The most complex level of this lab involves the Life Cycle Assessment (LCA) of all the energy costs associated with energy use and production. An LCA considers all the material and energy inputs in a product or process. According to the EPA, LCA is a technique to assess environmental impacts associated with all the stages of a product's life from cradle to grave (i.e., from raw material extraction through materials processing, manufacture, distribution, use, repair and maintenance, and disposal or recycling). LCAs can help prevent a narrow outlook on environmental concerns by:

- compiling an inventory of energy and material inputs and environmental releases;
- evaluating the potential impacts associated with identified inputs and releases; and
- interpreting the results to help make a more informed decision.

Students could extend the activity by researching the LCA of fossil fuel sources of energy, means of generating electricity, and other energy sources in the realm of transportation, communication, and leisure.

Additional Resources

- *Introduction to Green Chemistry*. Ryan, M. A.; Tinnesand, M., Eds. American Chemical Society: Washington, DC, 2002.
- Jansen, M. P. "How Efficient is a Laboratory Burner in Heating Water?" *J. Chem. Educ.* 1997, 74, 213–215.

Websites

- National Average Rate of Cost for Electricity
  http://bit.ly/highschoolnrg13

- City of Toronto Energy Efficiency
  http://bit.ly/highschoolnrg14

- U.S. Department of Energy, Efficiency Page
  http://bit.ly/highschoolnrg15

- Forbes Total Cost of Electricity Article
  http://bit.ly/highschoolnrg16

Exothermic, Endothermic, & Chemical Change
A Lab Investigation

Summary

In this investigation, students classify chemical reactions as exothermic or endothermic. Next, students explore the relationship between an observed change in temperature and the classification of a change as chemical or physical.

Objective

Students will explore energy changes during chemical reactions, heat of reaction (ΔH), and the connection between energy changes and chemical changes.

Safety

- Be sure you and the students wear properly fitting goggles.
- Acetic acid (vinegar) vapors can be irritating. Work in a well-ventilated area. In the event of eye contact, flush with water. The concentration of acetic acid in this experiment does not present any significant hazards.
- Calcium chloride can be an irritant to body tissues. In the event of contact, wash affected areas with water. Dispose of calcium chloride solutions according to local regulations.

Materials for Each Group

- Vinegar
- Baking soda
- Calcium chloride
- Water
- Thermometer
- 4 small clear plastic cups
- 1 cup measuring cup
- Measuring spoons (1 tablespoon, ½ teaspoon)

Time Required

One class period, approximately 45–50 minutes.

Lab Tips

After students explore one example of an endothermic change and one example of an exothermic change, they are then asked to explore the connection between energy changes and chemical reactions. To do this, students may need some guidance to arrive at the idea that temperature changes may also accompany dissolving.

Students will have an easier time devising a fair test if they are well versed in the definitions of physical changes and chemical changes. Students should propose an experiment to you before they test their hypothesis. To observe a temperature change during a physical change, students should devise a procedure such as:

Add 10 mL of water to a small plastic cup and place a thermometer in the water. Record the initial temperature (T_i).

Add ½ teaspoon of calcium chloride to the water and swirl the cup. After it has stopped changing, record the final temperature (T_f).

Pre-Lab Discussion

This investigation introduces the concepts of enthalpy (heat) of ΔH in the context of exothermic and endothermic reactions. To give students a deeper grounding in the basics and reinforce basic concepts covered previously, you may wish to review the mechanics of chemical changes, how to write balanced chemical equations, and the law of conservation of energy.

Incorporating into the Curriculum

This investigation could be incorporated into a unit on chemical changes or thermochemistry.

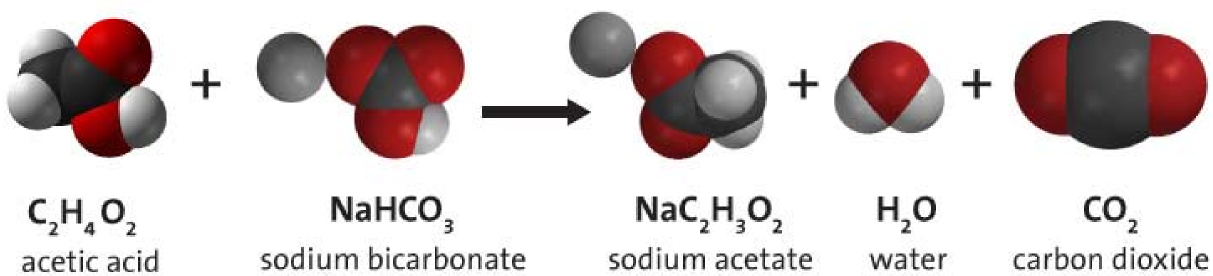
PREPARING TO INVESTIGATE

Energy Changes in Chemical Reactions

In this activity, you will explore the energy changes that accompany chemical reactions. To understand the energy implications of chemical reactions, it's important to keep in mind two key ideas:

1. It takes energy to break bonds.
2. Energy is released when bonds are formed.

To understand this, consider the chemical reaction between vinegar (also known as acetic acid to chemists) and baking soda (known as sodium bicarbonate). Before the atoms of acetic acid and sodium bicarbonate can be rearranged to form the products, the bonds between the atoms in those molecules must be broken, and because the atoms are attracted to one another, it takes energy to pull them apart.



Then, when the products are formed (sodium acetate, water, and carbon dioxide) energy is released because atoms that have an attraction for one another are brought back together. Not every bond between atoms in the reactants is necessarily broken during a chemical reaction, but some bonds are.

By comparing the energy used when bonds in the reactants are broken with the energy released when bonds in the products are formed, you can determine whether a chemical reaction releases energy or absorbs energy overall.

Chemical reactions that release energy are called exothermic. In exothermic reactions, more energy is released when the bonds are formed in the products than is used to break the bonds in the reactants. Chemical reactions that absorb (or use) energy are called endothermic. In endothermic reactions, more energy is absorbed when the bonds in the reactants are broken than is released when new bonds are formed in the products. If a chemical reaction absorbs as much energy as it releases, it is called isothermic—there is no net energy change.

But because we can't observe bonds breaking or being formed, how can we distinguish between exothermic and endothermic chemical reactions?

Identifying Exothermic & Endothermic Reactions

There are two methods for distinguishing between exothermic and endothermic reactions.

1. **Monitor temperature change**
   When energy is released in an exothermic reaction, the temperature of the reaction mixture increases. When energy is absorbed in an endothermic reaction, the temperature decreases. You can monitor changes in temperature by placing a thermometer in the reaction mixture.

2. **Calculate the enthalpy of reaction (ΔH)**
   To classify the net energy output or input of chemical reactions, you can calculate something called the enthalpy change (ΔH) or heat of reaction, which compares the energy of the reactants with the energy of the products.

   Enthalpy is a measure of internal energy. So, when you calculate the difference between the enthalpy of the products and the enthalpy of the reactants, you find the enthalpy change (ΔH), which can be represented mathematically as:

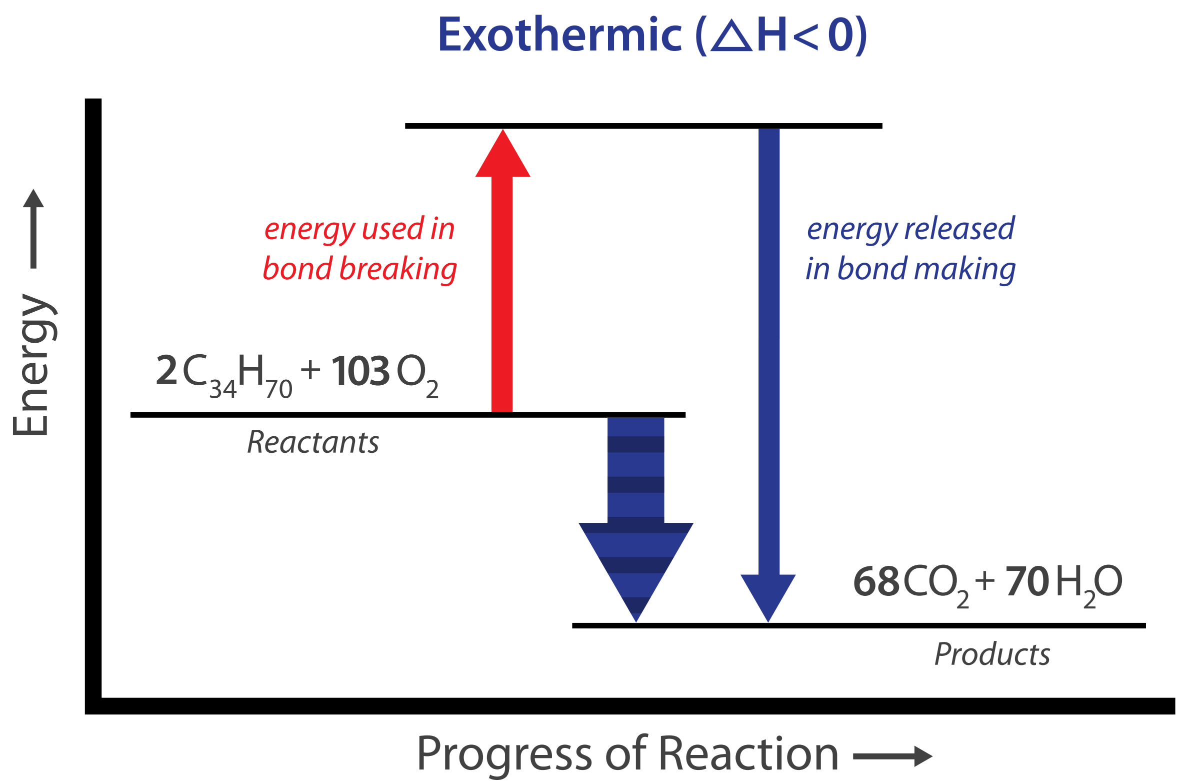
   ΔH = energy used in reactant bond breaking + energy released in product bond making

   Wait, how can you find a difference by adding? The enthalpy values are added in the equation above because, by definition, energy used in reactant bond breaking is always positive (+) and energy released in product bond making is always negative (–).

   If ΔH is negative (–) then the chemical reaction is exothermic, because more energy is released when the products are formed than energy is used to break up the reactants. If ΔH is positive (+) then the chemical reaction is endothermic, because less energy is released when the products are formed than the energy is used to break up the reactants.

   You can also use energy level diagrams to visualize the energy change during a chemical reaction as a result of the energies used and released according to the above equation for ΔH. To understand these diagrams, compare the energy level of the reactants on the left-hand side with that of the products on the right-hand side.

   The graph below charts the energy change when a candle burns. The wax ($C_{34}H_{70}$) combusts in the presence of oxygen (O_2) to yield carbon dioxide (CO_2) and water (H_2O). Because more energy is released when the products are formed than is used to break up the reactants, this reaction is exothermic, and ΔH for the reaction is negative.



In this investigation, you will observe whether energy is absorbed or released in two different chemical reactions and categorize them as exothermic and endothermic. You will also explore the relationship between energy changes and chemical reactions.

GATHERING EVIDENCE

Baking Soda and Vinegar

1. Pour about 10 mL of vinegar into a small plastic cup. Then, place a thermometer into the vinegar. Record the initial temperature (T_i) in the table below.
2. While the thermometer is in the cup, add about ½ teaspoon of baking soda to the cup.
3. Watch the thermometer for any change in temperature. After it has stopped changing, record the final temperature (T_f) and any other observations you made in the table below.

Baking Soda and Calcium Chloride

1. Make a baking soda solution by dissolving about 2 tablespoons of baking soda in 1 cup of water. Stir until no more baking soda will dissolve.
2. Place about 10 mL of baking soda solution in a small plastic cup. Then, place a thermometer into the baking soda solution. Record the initial temperature (T_i) in the table below.
3. While the thermometer is in the cup, add ½ teaspoon of calcium chloride to the cup.
4. Watch the thermometer for any change in temperature. After it has stopped changing, record the final temperature (T_f) and any other observations you made in the table below.

Process	T_i	T_f	ΔT	Exothermic or endothermic?	Other observations?	ΔH (+/−)
Baking soda + vinegar						
Baking soda solution + calcium chloride						

ANALYZING EVIDENCE

1. Calculate the temperature change for both chemical reactions. To do this, subtract the initial temperature (T_i) from the final temperature (T_f), and record the difference in the column labeled ΔT. You may see this calculation expressed elsewhere as $\Delta T = T_f - T_i$.
2. Based on your observations of the baking soda and vinegar reaction, is the reaction exothermic or endothermic? Apply your knowledge of energy changes in chemical reactions to complete the table above.
3. Based on your observations of the baking soda solution and calcium chloride reaction, is this chemical reaction exothermic or endothermic? Apply your knowledge of energy changes in chemical reactions to complete the table above.

INTERPRETING EVIDENCE

1. In the chemical reaction between baking soda and vinegar, what did you observe other than a temperature change? What might this tell you about one of the products of this chemical change?
2. In the chemical reaction between baking soda solution and calcium chloride, what did you observe other than a temperature change? What might this tell you about one of the products of this chemical change?
3. Use your answers from questions 1 and 2 to help you write the chemical equation for:
   - the chemical reaction between baking soda and vinegar
   - the chemical reaction between baking soda and calcium chloride
4. Using the language of breaking and making bonds, explain the net energy change for the chemical reaction between baking soda and calcium chloride.
5. Draw energy profiles for both chemical reactions. Refer to the exothermic energy profile shown previously as an example. Are they the same or different?
6. What is the sign of the heat of reaction (ΔH) for an exothermic reaction? Why?

REFLECTING ON THE INVESTIGATION

1. Based on your investigation so far, do you think that energy changes only accompany chemical reactions? Using only the materials from the first two reactions, design an experiment that would test this idea. Propose a procedure and have it approved by your teacher before you continue experimenting.
2. Is dissolving calcium chloride in water a chemical change? Explain your reasoning.
3. Using the language of breaking and making bonds, how can you describe the temperature change you observed when you dissolved calcium chloride in water?
4. How might you use exothermic or endothermic processes to solve a real-world problem? Are there any instances when it would be useful to quickly make something hot or cold? Explain how it is useful to know which processes absorb or release energy.

TEACHER'S KEY

Analyzing Evidence

Please note that initial and final temperature readings may vary slightly from student to student. The baking soda and vinegar reaction should produce a termpature decrease of approximately 7 °C. The baking soda and calcium chloride reaction should produce a temperature increase of approximately 15–20 °C.

Process	T_i	T_f	ΔT	Exothermic or endothermic?	ΔH (+/−)
Baking soda + vinegar	~22 °C	15 °C	−7 °C	Endothermic	+
Baking soda solution + calcium chloride	~22 °C	42 °C	+20 °C	Exothermic	−

1. Based on your observations of the baking soda and vinegar reaction, is the reaction exothermic or endothermic? Apply your knowledge of energy changes in chemical reactions to complete the table above.

   Endothermic, see table above.

2. Based on your observations of the baking soda solution and calcium chloride reaction, is this chemical reaction exothermic or endothermic? Apply your knowledge of energy changes in chemical reactions to complete the table above.

   Exothermic, see table above.

Interpreting Evidence

1. In the chemical reaction between baking soda and vinegar, what did you observe other than a temperature change? What might this tell you about one of the products of this chemical change?

   The reaction mixture bubbled, so one of the products must have been a gas. By considering the reactants, students may be able to infer that carbon dioxide was produced.

2. In the chemical reaction between baking soda solution and calcium chloride, what did you observe other than a temperature change? What might this tell you about one of the products of this chemical change?

   The reaction mixture bubbled, so one of the products must have been a gas. By considering the reactants, students may be able to infer that carbon dioxide was produced.

3. Write the chemical equation for:
   - the chemical reaction between baking soda and vinegar

     $C_2H_4O_2 + NaHCO_3 \longrightarrow CO_2 + H_2O + NaC_2H_3O_2$

   - the chemical reaction between baking soda and calcium chloride

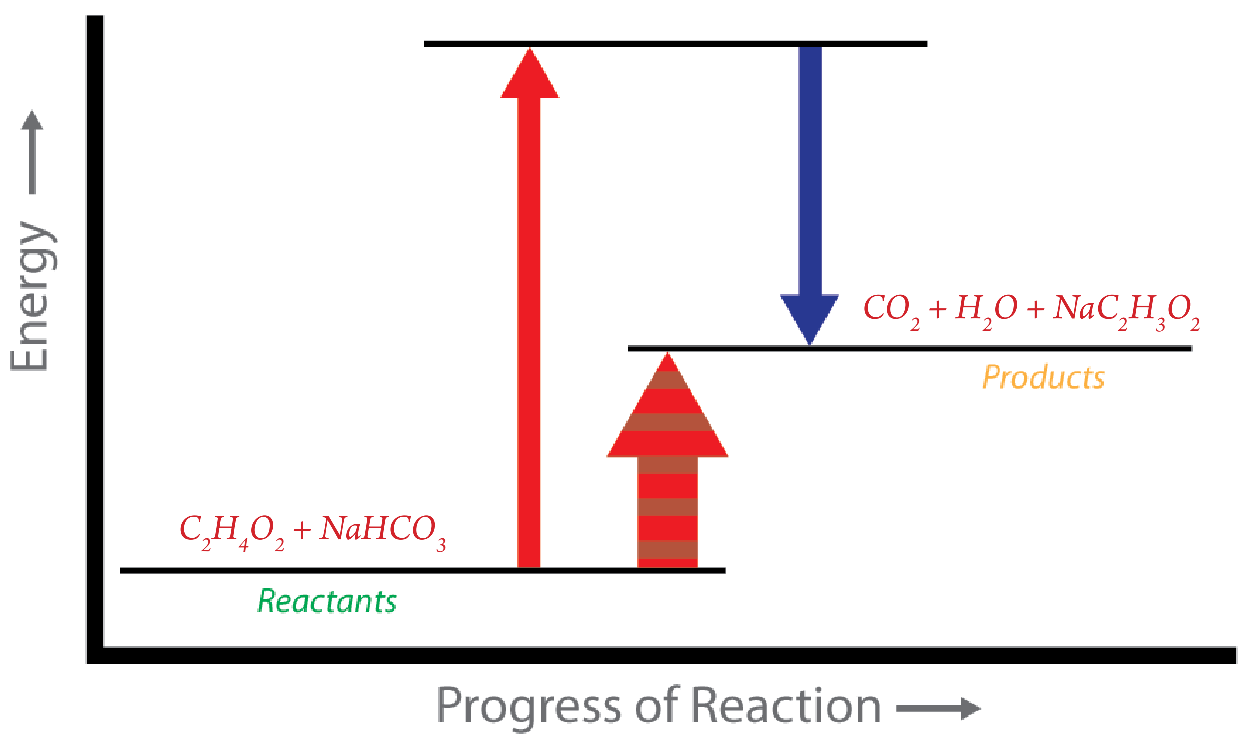
     $2NaHCO_3 + CaCl_2 \longrightarrow 2NaCl + CaCO_3 + H_2O + CO_2$

4. Using the language of breaking and making bonds, explain the net energy change for the chemical reaction between baking soda and calcium chloride.

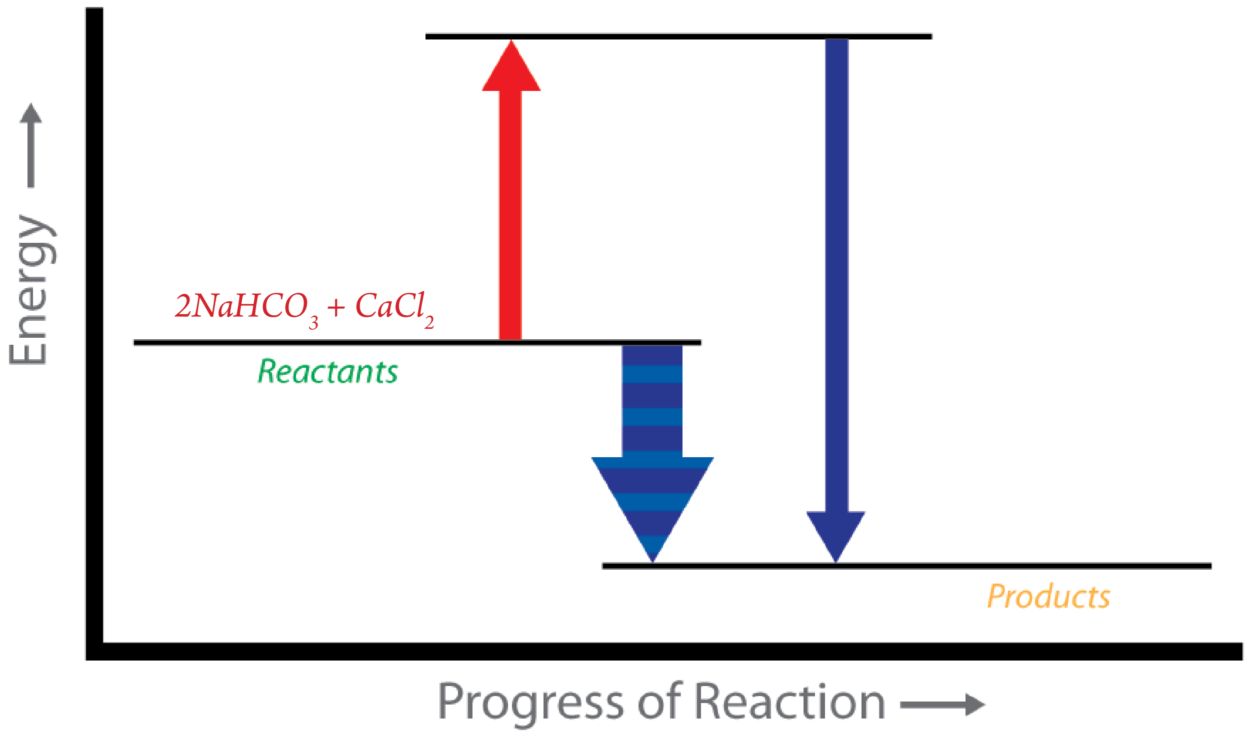
   In the chemical reaction between baking soda and calcium chloride, more energy is released when the products ($2NaCl + CaCO_3 + H_2O + CO_2$) are formed than energy was used to break bonds in the reactants ($2NaHCO_3 + CaCl_2$). There is therefore a net release of energy to the surroundings and the reaction is exothermic.

5. Draw energy profiles for both chemical reactions. Are they the same or different?

Baking soda + vinegar
Endothermic
$\Delta H>0$



Baking soda + calcium chloride
Exothermic
$\Delta H<0$



6. What is the sign of the enthalpy of reaction (ΔH) for an exothermic reaction? Why?

   The enthalpy (heat) for an exothermic reaction must be negative because more energy is released when the products are formed than is used when the reactants are broken up.

REFLECTING ON THE INVESTIGATION

Students should propose an experiment to you before they test their hypothesis. To observe a temperature change during a physical change, students should devise a procedure such as:

1. Based on your investigation so far, do you think that energy changes only accompany chemical reactions? Using only the materials from the first two reactions, design an experiment that would test this idea. Propose a procedure and have it approved by your teacher before you continue experimenting.

   Students should suggest a procedure like the following:

   Exothermic

   - Add 10 mL of water to a small plastic cup and place a thermometer in the water. Record the initial temperature (T_i).
   - Add ½ teaspoon of calcium chloride to the water and swirl the cup. After it has stopped changing, record the final temperature (T_f).

   Endothermic

   - Add 10 mL of water to a small plastic cup and place a thermometer in the water. Record the initial temperature (T_i).
   - Add ½ teaspoon of sodium bicarbonate to the water and swirl the cup. After it has stopped changing, record the final temperature (T_f).

2. Based on the experiment you conducted in the question above, do you think dissolving is a chemical change? Explain your reasoning.

   Student answers will vary. Some students may argue that nothing chemically distinct is made, so dissolving is a physical change. Other students may see the disruption of intermolecular forces as "bond breaking" (and solvating ions as "bond making") and therefore regard dissolving as a chemical change.

3. Using the language of breaking and making bonds, how can you describe the temperature change you observed when you dissolved calcium chloride in water?

   More energy was released when the calcium and chloride ions were combined with water than was required to break them apart.

4. How might you use exothermic or endothermic processes to solve a real-world problem? Are there any instances when it would be useful to quickly make something hot or cold? Explain how it is useful to know which processes absorb or release energy.

   Exothermic reactions could be harnessed to power machines or heat homes, while endothermic reactions could be used for treating injuries or cooling. By classifying reactions as exothermic or endothermic, we understand which reactions are best suited to meet specific challenges.

Calculating Lattice Energies Using the Born-Haber Cycle
An Extension Activity for AP Chemistry Students

A particular set of equations known as a Born-Haber cycle demonstrates how chemists are able to use the first law of thermodynamics—that the energy of the universe is conserved in any chemical or physical change—to find an unknown energy value that is difficult or impossible to measure experimentally. Some energy quantities, such as the lattice energy of a mineral or the electron affinity of an atom, can be difficult to measure in the lab. Examining a Born-Haber cycle we see that there is more than one path to the formation of a substance in a particular state, and that if we use consistent definitions, an energy value that we seek can be calculated from energy values that we already know.

The following exercise will help us see the way these energy values relate to one another, give us practice with their definitions and symbols, and deepen our insight to their meaning when we see them in other types of problems.

Each physical or chemical change represented has:

- an equation that represents a clearly defined physical or chemical change;
- a definition of the particular type of energy change;
- a symbol or abbreviation for the energy change equal to a value for the change in enthalpy (ΔH), the energy that is released or absorbed during the change, expressed in kJ/mol; and
- a name by which that change in enthalpy is commonly known.

To set up the equations in a Born-Haber cycle, cut out the cards for names, equations, definitions, and symbols with energy values. Arrange them to show the two alternate pathways to forming the ionic solid, linking the sequence of changes using the equation cards, and placing the definitions, names, and values near each equation. Pay close attention to the physical phases noted in the definitions.

When the cards have been arranged, examine the way the equations fit together. Can you clearly trace two paths to a final product? If so, according to the first law of thermodynamics, the energy changes along one path will be equal to the energy changes along the other path. By setting the sum of energy changes from one path equal to the energy changes of the other path, find the unknown value for the lattice energy of the solid.

CYCLE 1

$$Li(s) + \frac{1}{2} F_2(g) \longrightarrow LiF(s)$$

$$Li^+(g) + F^-(g) \longrightarrow LiF(s)$$

$$\frac{1}{2} F_2(g) \longrightarrow F^-(g)$$

$$Li(g) \longrightarrow Li^+(g) + e^-$$

$$Li(s) \longrightarrow Li(g)$$

$$F(g) + e^- \longrightarrow F^-(g)$$

$\Delta H^{o}_{f_{LiF}} = -594.1$ kJ

$\Delta H_{sub_{Li}} = 155.2$ kJ

1st IE = 520 kJ

½ $BE_{F_2} = 75.3$ kJ

$EA_{F} = -328$ kJ

$LE_{LiF} = ?$

The change in enthalpy when one mole of a substance in its standard state is formed from its constituent elements in their standard states.

The energy required to remove the outermost electron of each atom in one mole of an element in its gaseous state.

The energy released or absorbed when an electron is added to each atom in one mole of a substance in its gaseous state.

The energy needed to transform one mole of a substance from the solid to the gaseous state.

The energy released when one mole of an ionic compound is formed from its constituent ions in their gaseous states.

The energy required to break one mole of bonds between two atoms.

Standard enthalpy of formation

First ionization energy

Lattice energy

Bond energy

Enthalpy of sublimation

Electron affinity

CYCLE 2

$$Na(s) + \frac{1}{2} Cl_2(g) \longrightarrow NaCl(s)$$

$$Na^+(g) + Cl^-(g) \longrightarrow NaCl(s)$$

$$\frac{1}{2} Cl_2(g) \longrightarrow Cl(g)$$

$$Na(g) \longrightarrow Na^+(g) + e^-$$

$$Na(s) \longrightarrow Na(g)$$

$$Cl(g) + e^- \longrightarrow Cl^-(g)$$

$\Delta H^{o}_{f_{NaCl}} = -411 \text{ kJ}$

$\Delta H_{sub_{Na}} = 108 \text{ kJ}$

$1\text{st } IE_{Na} = 496 \text{ kJ}$

$\frac{1}{2}\, BE_{Cl_2} = 75.3 \text{ kJ}$

$EA_{Cl} = -349 \text{ kJ}$

$LE_{NaCl} = ?$

CYCLE 3

For cycle 3, set up the Born-Haber cycle to find lattice energy using known energy values as before. This time you will also need to find the second ionization energy of calcium where:

Second ionization energy
The energy required to remove the *second* outermost electron of each atom in one mole of an element in its gaseous state.
$Ca^{+}(g) \longrightarrow Ca^{2+}(g) + e^{-}$
$2nd\ IE_{Ca} = 1145\ kJ$

$$Ca\,(s) + Cl_2\,(g) \longrightarrow CaCl_2(s)$$

$$Ca^{2+}\,(g) + 2Cl^-\,(g) \longrightarrow CaCl_2(s)$$

$$Cl_2\,(g) \longrightarrow 2Cl\,(g)$$

$$Ca\,(g) \longrightarrow Ca^+\,(g) + e^-$$

$$Ca\,(s) \longrightarrow Ca(g)$$

$$2Cl\,(g) + 2e^- \longrightarrow 2Cl^-\,(g)$$

$\Delta H^{o}_{f_{CaCl_2}} = -795.8\ kJ$

$\Delta H_{sub_{Ca}} = 178.2\ kJ$

1st $IE_{Ca} = 590\ kJ$

½ $BE_{Cl_2} = 244\ kJ$

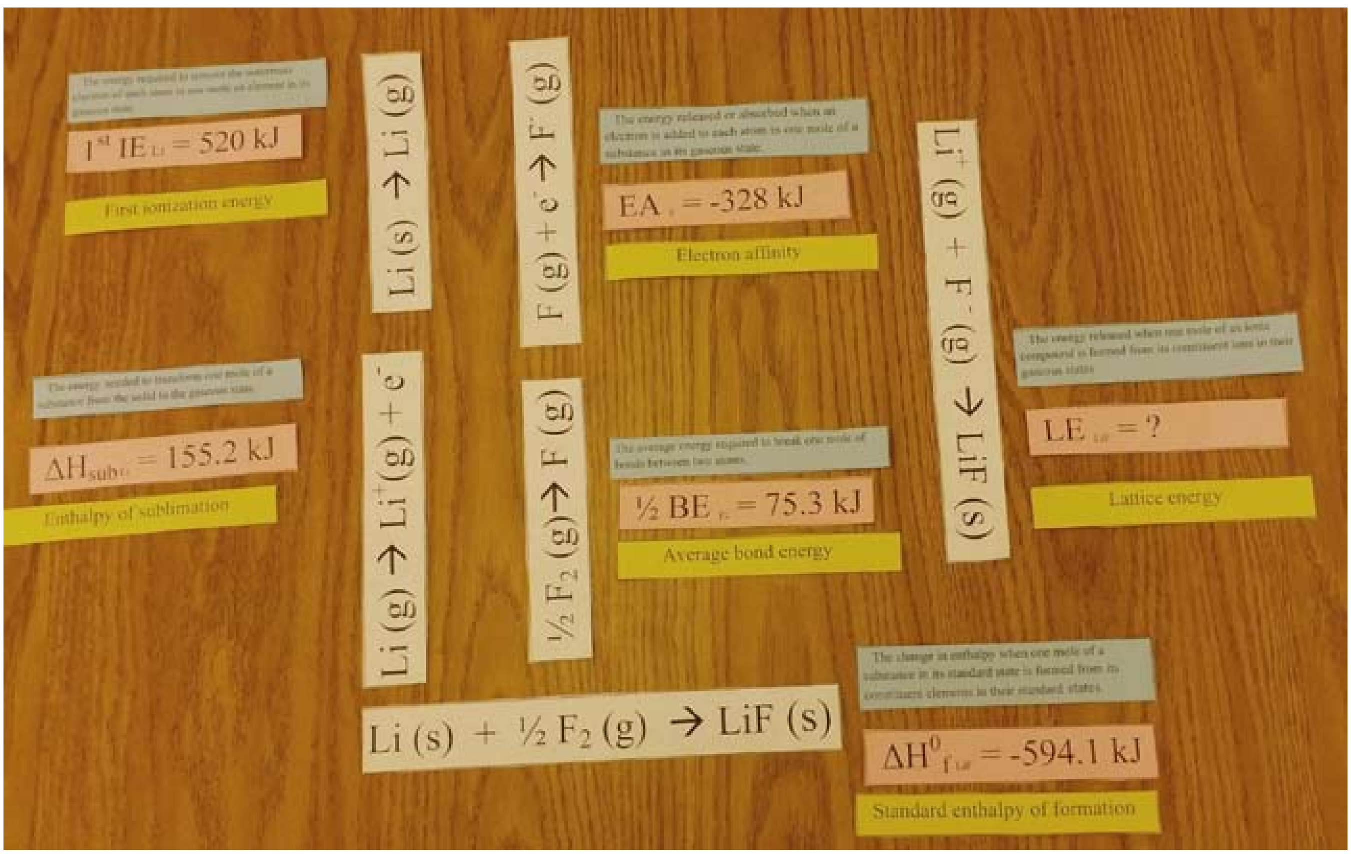
$EA_{Cl} = -349\ kJ$

$LE_{CaCl_2} = ?$

TEACHER'S KEY

- In this exercise, students will identify ionization energy, electron affinity, standard enthalpy of formation, energy of sublimation, bond energy, and lattice energy as ΔH values for particular reactions.
- Constructing a Born-Haber cycle will demonstrate how Hess's law (and the first law of thermodynamics) can be used to find one of the energy values in the cycle, if the others are known.
- Creating a visual arrangement with cards gives students a visual context to help them see how the specific energy quantities can be used to determine the unknown value.
- As shown below, definitions, symbols, equations, and energy values could each be copied in a different color, to help with recognizing and organizing the cards.
- Students should work with one complete set of equations at a time.

Sample of student arrangement:



Cycle 1

$Li (s) \rightarrow Li (g)$	ΔH_{sub} = 155.2 kJ
$Li (g) \rightarrow Li^{+}(g) + e^{-}$	1st Ionization Energy of Li = 520 kJ
$\frac{1}{2} F_2 (g) \rightarrow F (g)$	½ Bond Energy of F_2 = 75.3 kJ
$F (g) + e^{-} \rightarrow F^{-} (g)$	Electron Affinity for fluorine = −328 kJ
$Li^{+} (g) + F^{-} (g) \rightarrow LiF (s)$	Lattice Energy of LiF = ??
$Li (s) + \frac{1}{2} F_2 (g) \rightarrow LiF (s)$	Standard Enthalpy of Formation of LiF (s) = −594.1 kJ

−594 kJ = LE + −328 kJ + 75.3 kJ + 520 kJ + 155.2 kJ

Lattice Energy for LiF = −1016 kJ

Cycle 2

$Na (s) \rightarrow Na (g)$	ΔH_{sub} = 108 kJ
$Na (g) \rightarrow Na^{+}(g) + e^{-}$	1st Ionization Energy of Na = 496 kJ
$\frac{1}{2} Cl_2 (g) \rightarrow Cl (g)$	½ Bond Energy of Cl_2 = 122 kJ
$Cl (g) + e^{-} \rightarrow Cl^{-} (g)$	Electron Affinity for chlorine = −349 kJ
$Na^{+} (g) + Cl^{-} (g) \rightarrow NaCl (s)$	Lattice Energy of NaCl = ??
$Na (s) + \frac{1}{2} Cl_2 (g) \rightarrow NaCl (s)$	Standard Enthalpy of Formation of NaCl (s) = −411 kJ

−411 kJ = LE + −349 kJ + 122 kJ + 496 kJ + 108 kJ

Lattice Energy for NaCl = −788 kJ

Cycle 3

$Ca (s) \rightarrow Ca (g)$	ΔH_{sub} = 178.2 kJ
$Ca (g) \rightarrow Ca^{+}(g) + e^{-}$	1st Ionization Energy of Ca = 590 kJ
$Ca^{+} (g) \rightarrow Ca^{2+}(g) + e^{-}$	2nd Ionization Energy of Ca = 1145 kJ
$Cl_2 (g) \rightarrow 2Cl (g)$	Bond Energy of Cl_2 = 244 kJ
$2Cl (g) + 2e^{-} \rightarrow 2Cl^{-}$	Electron Affinity for chlorine = 2(−349 kJ)
$Ca^{2+} (g) + 2Cl^{-} (g) \rightarrow CaCl_2 (s)$	Lattice Energy of $CaCl_2$ = ??
$Ca (s) + Cl_2 (g) \rightarrow CaCl_2 (s)$	Standard Enthalpy of Formation of $CaCl_2$ (s) = −795.8 kJ

−795.8 kJ = LE + 2(−349 kJ) + 244 kJ + 1145 kJ + 590 kJ + 178.2 kJ

Lattice Energy for $CaCl_2$ (s) = −2255 kJ

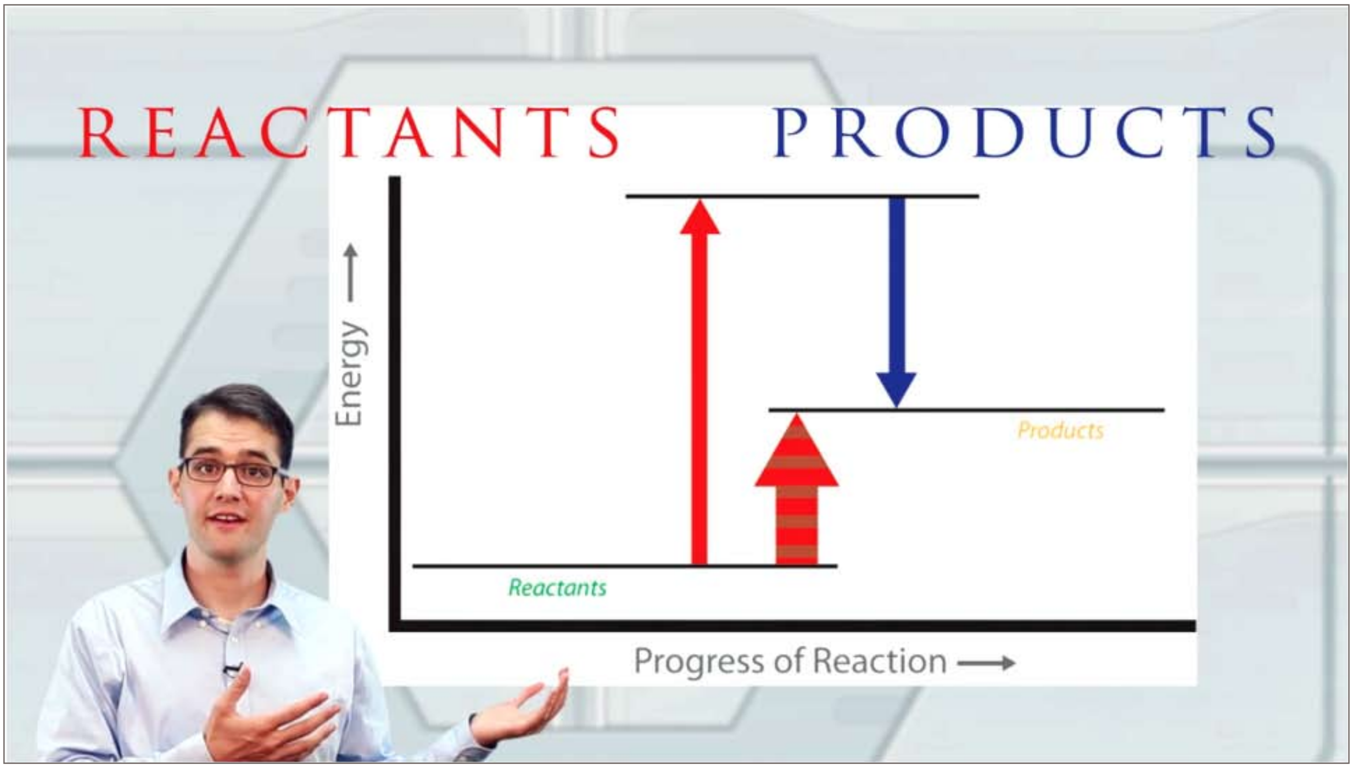
Further Analysis

Lattice energy is often used to estimate the strength of an ionic bond. Comparing the lattice energies of the three salts, students could now be asked to look at the relative strengths of the bonds in the compounds to relative sizes of ions and relative charge on ions.

According to Coulomb's law, the strength of the bond should increase with increasing charge on the ion, and with decreasing size of the ion.

Do students see any evidence for this in their results? Can they explain their reasoning?

Exothermic & Endothermic Reactions | A Video



This video explores the energy implications of chemical change. It can be used as a supplement to the investigation on "Exothermic, Endothermic, & Chemical Change" or may stand on its own to introduce a lesson or extend student learning.

highschoolenergy.acs.org/how-can-energy-change/exothermic-endothermic.html

Video Transcript

Chris
Chemical reactions: They're fundamental to chemistry; they make new things by rearranging other things. They can blow stuff up ... or freeze things quickly. In short, they are awesome.

Brittny
But why do some chemical reactions release massive amounts of energy, while others absorb energy? In a chemical reaction, the main change that occurs relates to the way atoms are connected (or bonded) to each other. In order to change those connections, bonds must be broken and new bonds must be formed. Let's break down how energy is transferred in these reactions.

Chris
To understand the energy implications of chemical reactions, it's important to keep in mind two key ideas:

1. It takes energy to break bonds.
2. Energy is released when bonds are formed.

To understand this, consider the chemical reaction between vinegar and baking soda. That's right—the classic baking soda volcano experiment. The chemical reaction behind this science fair favorite involves baking soda—also known as sodium bicarbonate to chemists—and vinegar, otherwise known as acetic acid.

These compounds react to form the molecules sodium acetate, water, and carbon dioxide. The baking soda and vinegar are called the reactants. The sodium acetate, water, and carbon dioxide that are formed are called the products.

Before the atoms in acetic acid and sodium bicarbonate can be rearranged to form the products, some of the bonds between the atoms in those molecules must be broken, and because the atoms are attracted to one another, it takes energy to pull them apart.

Then, when the products are formed (sodium acetate, water, and carbon dioxide) energy is released because atoms that have an attraction for one another are brought back together.

By comparing the energy absorbed when bonds in the reactants are broken with the energy released when bonds in the products are formed, you can determine whether a chemical reaction releases energy or absorbs energy overall.

Brittny
Chemical reactions that release energy are called exothermic. In exothermic reactions, more energy is released when the bonds are formed in the products than is used to break the bonds in the reactants. Exothermic reactions are accompanied by an increase in temperature of the reaction mixture.

Chemical reactions that absorb (or use) energy overall are called endothermic. In endothermic reactions, more energy is absorbed when the bonds in the reactants are broken than is released when new bonds are formed in the products. Endothermic reactions are accompanied by a decrease in temperature of the reaction mixture.

Chris

You can use energy level diagrams to visualize the energy change during a chemical reaction. To understand these diagrams, compare the energy level of the reactants on one side with that of the products on the other side.

Consider, for example, a diagram that charts the energy change when a candle burns. Wax ($C_{34}H_{70}$) combusts in the presence of oxygen (O_2) to yield carbon dioxide (CO_2) and water (H_2O). Because more energy is released when the products are formed than is used to break up the reactants, this reaction is exothermic.

Brittny

All of this stuff relates to thermodynamics—the study of heat and its relationship to energy and work. Using thermodynamics, you'll learn how to calculate the precise amount of energy used or released by chemical reactions.

Classifying a chemical reaction as exothermic or endothermic is simple. It comes down to weighing the energy needed to break bonds in the reactants with the energy released when the products are formed.

It's a simple idea, but one with a lot of power.

Meet a BP Chemist | A Video

This video profiles the work of Sarah, a chemist at BP. It can be used as a supplement to give students an idea of what a chemist who works on energy issues does on a day-to-day basis.

highschoolenergy.acs.org/how-can-energy-change/career-profile.html

---

Video Transcript

As a scientist, sometimes it's hard to tell people what you're actually working on, because you say, "I'm a scientist," and they sort of tune out.

But the research we're doing now is really relevant to everybody, because everybody understands gasoline—they put it in their car (almost) every single day. We're trying to take feedstocks, which come from land that wouldn't be producing food and convert them into (fuel) that you can put in your tank—to drive your car, your truck, and maybe even airplanes at some point.

So, the work that I'm doing to make biofuels will directly impact everybody, and when I tell people I work at BP Biofuels and alternative energy, they get really excited about it, and they want to talk to me to find out what's happening.

Being a part of that technology makes you feel like your invention, or your discoveries, or your technology will actually be used someday, and that's one of the benefits that not all scientists get to realize.

Somebody who wants to join BP Biofuels is somebody who wants to make a difference. We need to be invested in renewable energy, and so it's one of the very few places where you can work and your technology will actually be transplanted into a large-scale technology that's going to touch everyone's lives.

