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ENERGY FOUNDATIONS
for High School Chemistry

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American Chemical Society

Energy Foundations for High School Chemistry

The activities described in this book are intended for students under the direct supervision of teachers. The American Chemical Society cannot be held responsible for any accidents or injuries that might result from conducting the activities without proper supervision, from not specifically following directions, or from ignoring the cautions contained in the text.

American Chemical Society Staff

Authors
Michael J. Tinnesand
Erica Jacobsen
Melanie Rock
Adam M. Boyd

Education Division Director, Mary Kirchhoff
Assistant Director of K–12 Education, Terri Taylor
Content Reviewer, Dr. Jerry Bell
Copy Editor, Kelley Carpenter

Safety Review, David Katz

SAFETY

Important Notice

Energy Foundations for High School Chemistry is intended for use by high school students in the classroom laboratory under the direct supervision of a qualified chemistry teacher. The experiments described in this book involve substances that may be harmful if they are misused or if the procedures described are not followed. Read cautions carefully and follow all directions. Do not use or combine any substances or materials not specifically called for in carrying out investigations. Other substances are mentioned for educational purposes only and should not be used by students unless the instructions specifically so indicate.

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Safety and Laboratory Activity

In *Energy Foundations for High School Chemistry*, you will frequently complete the laboratory investigations. While no human activity is completely risk-free, if you use common sense, as well as chemical sense, and follow the rules of laboratory safety, you should encounter no problems. Chemical sense is just an extension of common sense. Sensible laboratory conduct won't happen by memorizing a list of rules, any more than a perfect score on a written driver's test ensures an excellent driving record. The true "driver's test" of chemical sense is your actual conduct in the laboratory.

Students will see the goggle icon (at right) when they should be wearing them. Students should continue to wear goggles as they work through an investigation and until they are completely finished in the laboratory. The term "goggles" used in the safety section of each lesson refers to chemical splash goggles as shown in the icon accompanying every activity in the lessons.

Students will see the red caution icon (at right) when there are substances or procedures requiring special care. It is especially important for the instructor to supervise the usage of materials or execution of procedures noted with this icon.

ABOUT THIS TEXT

The Energy Foundations Project

Energy Foundations for High School Chemistry is a collection of resources that you can use to teach the big ideas about energy in your high school chemistry classroom. Featuring lab investigations, demos, student readings, and multimedia, it provides educational pieces that will support your existing curriculum.

Although the resources are organized into four sections (listed on page 6), this is only one suggested way of stepping through the content. Please be encouraged to skip around and choose the pieces from each section that best meet the educational needs of your specific classroom.

Likewise, the organization of resources within a section is meant only as a guide. Taken together, every piece of content is intended to be used *à la carte*, with sufficient flexibility to accommodate varying approaches and needs.

The Authors

Energy Foundations was developed and written by Michael J. Tinnesand, Erica Jacobsen, Melanie Rock, and Education Division staff at the American Chemical Society. Michael, Erica, and Melanie have decades of experience in chemistry teaching and science education consulting. You can find more resources from the ACS Education Division at *www.acs.org/education*.

The Sponsor

All of the resources from *Energy Foundations* are made possible by generous funding from BP.

WELCOME FROM THE AUTHORS

Dear Fellow Educators,

Welcome to the new High School Energy Module from the American Chemical Society! In this booklet and its companion website—highschoolenergy.acs.org—you will find content for high school chemistry teachers, including animations, videos, and articles that teach concepts of energy in the context of laboratory investigations.

The idea for this project started with a grant from BP and included teacher surveys and interviews. Teachers were asked about which topics and content types would be most beneficial.

Using this input, our curriculum developers came up with activities divided into four areas:

1. What is Energy? (Types of energy and how we experience them)
2. How Do We Use Energy? (The energy in chemical processes and everyday life)
3. How Can Energy Change? (Conservation of energy and energy transfer)
4. What Theories Explain Energy? (Thermodynamics—the theory of energy)

For each of these areas, we developed a set of lab investigations, videos, demonstrations, and background reading materials.

In order to make these materials as widely useful as possible, we sought to align them with the Next Generation Science Standards, which were based on the National Research Council's Framework for K–12 Science Education. Although we've presented our ideas in a slightly different order and description, there is a close match between the ACS High School Energy Module and the NGSS.


Michael J. Tinnesand, lead curriciulum developer

Each of the elements of the module can be used alone or as a complete set of activities, depending on your curriculum needs. In addition, many of the activities are presented in a format that can be downloaded and modified to meet your specific requirements.

We are confident you will find these materials useful, and we hope you enjoy using them. If you have any comments about these materials, please feel free to contact us at *education@acs.org*.

Cheers!
Michael

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What is Energy?

What is energy? Most of us have a feeling that we understand energy and recognize it when we see it, but coming up with a formal definition might be harder for us to do. In this section we introduce the concept of energy by having students investigate a variety of toys. They are given a list of the various categories and types of energy and are asked to form a working definition of what energy means. A series of videos and reading selections help them solidify their definition into an authoritative version.

- Lab Investigation—Exploring Energy
- Video—What is Energy?
- Video—The Energy of Toys
- *ChemMatters* Reading—The Amazing Drinking Bird
- *ChemMatters* Reading—Thermometers

Exploring Energy | A Lab Investigation

Summary

Energy as a concept is both familiar and vague. Most students have a good operational definition of energy, but may not be able to verbalize a good formal definition. The purpose of this lab is to solidify student ideas of energy by providing several examples of objects whose use or operation involve energy and provide clear definitions that describe what energy is and is not.

Objective

Students will explore the definition of energy by making careful observations about simple toys that illustrate basic principles of energy.

Safety

- Goggles or safety glasses should be worn when working with any materials that can spill or splash or those that can pop or jump with any force.
- Do not cut or tear open light sticks, chemical hand warmers, or cold packs. If any materials leak from these items, clean up spills with water. Wash any chemicals from your skin with water.
- Jumping discs and poppers can react with significant force to cause injury to the face and eyes. Keep your face away from these objects.

Materials for Each Group

Here is a sample list. Collect your own devices based on what you have available and your budget. Consider asking students to bring in a household device that involves energy.

- Battery-operated flashlight
- Clock with "glow in the dark" face
- Small mechanical toy (like a wind-up car)
- Drinking bird toy
- Chemiluminescent glow or light stick
- Chemical hand-warmer
- Emergency cold packs
- Bi-metallic jumping discs
- Polymer or "poppers" toy
- A hand-wound or spring-operated music box

Time Required

One class period, approximately 45–50 minutes.

Lab Tips

A number of household objects and toys have been suggested for use in this investigation, but the final selection can be modified to whatever you have on hand. Try to choose objects that display a variety of forms of energy. Since this is a formative type of investigation, it is not critical that students are perfect in their assessments of the type of energy present or the transitions that take place. Rather, use their observations and class discussion to clarify their understanding of the concept.

Pre-Lab Discussion

This lab is basically an informal formative assessment of what students know about energy. It is a good chance to clarify student ideas about energy by identifying what is and what is not considered to be energy.

Integrating into the Curriculum

This investigation could fit into a unit on thermochemistry.

PREPARING TO INVESTIGATE

What is energy? Most of us have a feeling that we understand energy and recognize it when we see it, but coming up with a formal definition might be harder for us to do. Here are some of the basic concepts associated with a definition of energy:

- Energy is required to make things change.
- Energy is the ability to do work. The kind of work we are talking about is not like going to a job or doing chores. We are talking about work as defined in the physical sciences.
- Work is the application of a force to move an object in the direction of the force, such as when you pedal a bicycle or an electric motor lifts an elevator.
- A force is an influence that can cause an object to move or stop an object that is already moving. Think about how brakes must be applied to slow an automobile, or how you throw a ball hard to get it to go as far as possible.
- Motion is a change in distance over time.

One of the easiest ways to recognize energy is to know the various forms it can take.
All energy falls into two categories, kinetic and potential energy. Kinetic energy is the energy of motion. Think about the energy in a brick. Which has more energy, a brick lying on your foot, or a brick dropped on your foot? The dropped brick has more kinetic energy.

Energy is not an object in itself, rather, energy refers to a condition or state of an object.

The main types of kinetic energy are:

- **Electrical Energy**—The energy associated with the movement of electrons. When electrons flow through wires, we call it electricity.
- **Thermal Energy**—The energy that results from the movement of atoms and molecules and is related to their temperature. The faster the particles move, the greater the amount of energy and the higher the temperature.
- **Movement Energy**—When objects or materials flow or move from one place to another they produce energy, such as water flowing through a dam.
- **Sound Energy**—Produced by the periodic movement of matter in a medium. Sound can travel through air, solids and liquids, but not through a vacuum, because there is no matter there.
- **Radiant Energy**—A type of kinetic energy that includes light rays, X-rays, radio waves, microwaves, and any other part of the electromagnetic spectrum. Electromagnetic waves are a result of the vibration of charged particles such as electrons. Microwave ovens use radiant energy to heat food by causing the water molecules in the food to vibrate.

Potential Energy is energy that is stored. In the example above, before it fell on your foot, the raised brick had potential energy because of its position. When it fell, this energy was converted to kinetic energy of motion.

Here are the main types of potential energy:

- **Chemical Energy**—Chemical bonds hold atoms together. It takes energy to break these bonds, and energy is released when new bonds form.
- **Mechanical Energy**—Examples include a clock that is powered by a wound-up spring, or an arrow shot from a bow. With this type, energy is stored in the mechanical device by the application of a force, such as when we wind the clock or pull the bow back.
- **Nuclear Energy**—When the nucleus of an atom splits or is fused to another nucleus, energy can be released. It is the type of energy that powers our sun and is found in nuclear power plants.
- **Gravitational Energy**—This is the energy associated with an object's position in a gravitational field. A ball resting at the top of a ramp has higher potential energy than when it has rolled to the bottom. Water behind a dam has higher potential energy than when it has flowed to the river below. In each case the potential energy is due to relative position in the gravitational field.

GATHERING EVIDENCE

You will be provided with a number of toys and other common household objects. The objects can include a flashlight, a clock with a "glow in the dark" face, a small mechanical toy, a drinking bird, a light stick, a hand-warmer or emergency cold packs, poppers, or a music box, among others. Your teacher may ask you to bring in a household object that involves energy.

For each of four objects, operate the device and write a complete description about what you observe. Be sure to include details such as sounds, movements, timing, and any changes you notice. If you are unclear about how to operate any of the objects, ask your teacher to show you how. Your teacher may choose to demonstrate some of the examples for the entire class.

Name of object	Description of operation
1.	
2.	
3.	
4.	

ANALYZING EVIDENCE

To analyze what you saw for each of the objects, you need to consider what kinds of energy were involved. Using the types of energy listed in the Preparing to Investigate section on page 10, list all the kinds of energy you believe are involved in the operation of each of the objects.

Your teacher may ask you to work in small, collaborative groups to complete this section.

Name of object	Kind(s) of energy involved
1.	
2.	
3.	
4.	

INTERPRETING EVIDENCE

Although some of the toys or objects may have appeared to “run out” of energy, one of the basic rules of science is that energy cannot be created or destroyed. It only changes from one form to another. Think back to your observations and analysis of the household items and speculate about how energy changed in each of the objects. In some cases there may have been multiple transitions.

Name of object	Describe the kinds of energy transitions you observe
1.	
2.	
3.	
4.	

REFLECTING ON THE INVESTIGATION

1. How did you detect the types of energy you observed in this investigation? Which types did you feel? Which types did you hear? Which types did you see? List any other ways you detected energy.
2. One definition of energy given in this lab was "energy is the ability to do work." Considering all the objects you observed in this lab, record any examples of work that was done.
3. Given the definition of force (an influence that can cause an object to move or stop moving) identify some examples of where force was applied in this investigation.
4. Based on your experience, write a definition of energy in your own words. Also include some examples of things that are NOT energy.
5. Starting with the sun and ending with you making a piece of toast at home, write down as many types of energy and energy transitions as you can that go into your ability to have a piece of bread and toast it.
6. Often when the word "energy" is used, it refers to commercial production to help power our homes. List the major types of energy used in commercial power plants.

TEACHER'S KEY

Gathering Evidence

Only one sample is given to demonstrate typical answers. The results will vary, based on the collection of objects you provide to the class.

Name of object	Description of operation
Bi-metallic jumping disc	*The jumping disc is a shallow dome made from a special bi-metallic piece of metal, like the mechanism of many thermostats. When it is warmed to about body temperature, you can "click it" into its loaded position (dome inverted) and then carefully place it on a hard surface. When it cools down to room temperature, it suddenly snaps back into the old position, and simultaneously jumps high up in the air.*

Analyzing Evidence

Name of object	What kind(s) of energy do you observe?
Bi-metallic jumping disc	*Thermal energy is transferred to the cool disc as it is warmed by my thumb and fingers until it is warmed to about body temperature. I observed mechanical energy as I "snap" the disc into its loaded position and then carefully place it on a hard surface.* *When it cools down to room temperature, it suddenly releases its stored potential energy and snaps back into the old position, and simultaneously jumps high up in the air. It jumps several hundred times its own height, and I observed kinetic or motion energy.*

Interpreting Evidence

Name of object	Describe the kinds of energy transitions you observe
Bi-metallic jumping disc	*I observed chemical energy from my metabolism creating thermal energy, which was transferred to the disc. I used mechanical energy to click the disc into its loaded position, creating potential energy. As the disc "popped" it converted potential energy into kinetic energy as it jumped off the desk. When the disc fell back to the desk it transferred its kinetic energy into thermal energy by slightly warming up the desktop (although this is difficult to observe, I assumed it occurred).*

Reflecting on the Investigation

1. How did you detect the types of energy you observed in this investigation? Which types did you feel? Which types did you hear? Which types did you see? List any other ways you detected energy.

   Students will likely report examples of all the types of energy listed above, including nuclear energy (some glowing clock faces use tritium or promethium to create the glow).

2. One definition of energy given in this lab was "energy is the ability to do work." Considering all the objects you observed in this lab, record any examples of work that was done.

   Work is any force applied over a distance. In chemical systems it involves electrical work (moving electrons in a wire as by a battery) or work of expansion (volume expanding as a result of a chemical reaction, as when bubbles are formed in a reaction of baking soda and acid).

3. Given the definition of force (an influence that can cause an object to move or stop moving) identify some examples of where force was applied in this investigation.

   Many examples are presented, such as the spring driving the mechanical toy, using your fingers to snap the popper into its loaded position, and so on.

4. Based on your experience, write a definition of energy in your own words.

   Student answers will likely echo the definitions given in the first part of this lab. Encourage students to put the ideas in their own words.

5. Starting with the sun and ending with you making a piece of toast at home, write down as many types of energy and energy transitions as you can, that go into your ability to make toast.

   The sun provides radiant energy for wheat plants to grow, and the wheat plant stores the solar *energy in seeds via photosynthesis to create chemical potential energy. Radiant solar energy powers the water cycle, giving rivers that can turn the kinetic energy of moving water into electrical energy. Household electrical energy moves through high-resistance heating elements in the toaster and is converted into thermal energy.*

6. Often when the word "energy" is used, it refers to commercial production to help power our homes. List the major types of energy used in commercial power plants.

   The main sources of our electricity are:

   - *Coal and natural gas*
   - *Hydroelectric dams*
   - *Nuclear*
   - *Wind*
   - *Solar*

Post-Lab Discussion

Use the post-lab discussion to listen to student observations and use a group discussion to clarify any misconceptions about energy.

Additional Resources

- Thermodynamics, Real-Life Applications
  http://bit.ly/highschoolnrg1

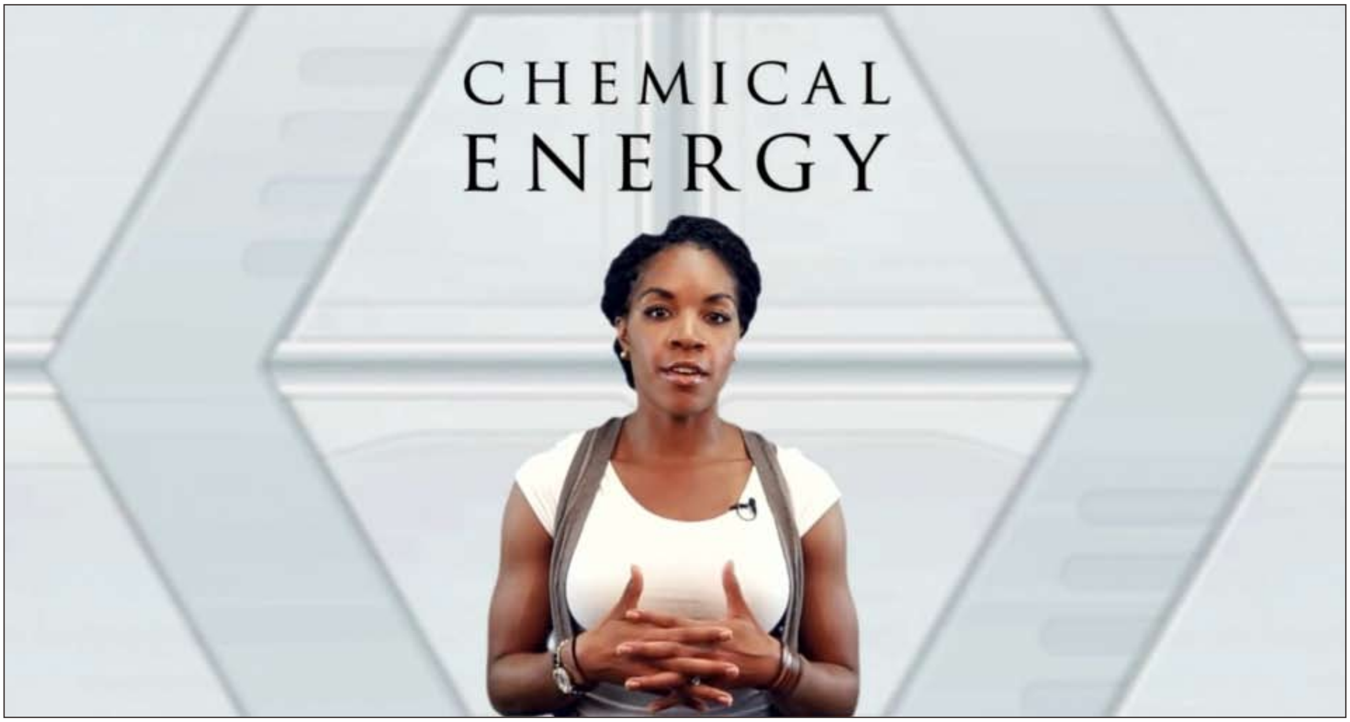
- U.S. Department of Energy, Secondary Energy Infobook, Activities and Background for High School Students
  http://bit.ly/highschoolnrg2

Note: Even in some of these official websites, there is a tendency to misrepresent heat as something that is contained by an object and flows between one object and another.

Objects/systems do not contain "heat." The motion of molecules is not heat. The term "heat content" is historic and archaic (even though it unfortunately continues in use) and comes from a time when heat was visualized as a fluid called "caloric."

As a corollary, heat does not "flow." Heat is a measure of the energy that is transferred from one object to another due to a difference in their temperatures. The energy transferred is often called "thermal energy," since it is dependent on there being a temperature difference. The form of the transfer is conduction and radiation. It's okay to call the amount of energy transferred "heat" as long as we are clear that "heat" is a number (with units) that is not found within the objects themselves, but is a measure of a process. It's best to call attention to the process by using "heat" as a verb instead of a noun, as in "When I heated the soup, it got hotter; its temperature went up."

What is Energy? | A Video



This video defines energy and explores some of its most common forms. It can be used as a supplement to an investigation, or may stand on its own to introduce a lesson or extend student learning.

highschoolenergy.acs.org/what-is-energy/what-is-energy.html

Video Transcript

Brittny
What is energy? Most of probably understand the concept of energy, or at least we can recognize it when we see it.

Chris
Here are some basic concepts that will help us define energy. First, energy is required to make things change. For example, to change this No. 2 pencil from one useful writing utensil into two less useful pieces, we need to use energy.

Brittny
Energy is the ability to do work. But we're not talking about going to work or doing your chores, we're talking about work as defined in the physical sciences. Work is the application of a force to move an object in the direction of the force. Like, when you pedal a bicycle, or when an electric

motor lifts an elevator. Energy is conserved. So, what does that mean? It means that one form of energy may be transformed into another form, but the total amount of energy remains the same. In other words, energy cannot be created, nor can it be destroyed.

Chris
One of the easiest ways to recognize energy is to know the different forms it can take. All energy falls into two categories: potential and kinetic energy. Potential energy depends on the position of an object or the arrangemet of its constituent parts. Kinetic energy is the energy of motion. Think about the energy of a brick resting on your foot compared to a brick being dropped on your foot. The moving brick has more kinetic energy, which you become painfully aware of when it transfers some of that energy onto your foot. When both bricks are resting on your foot, they have the same potential energy.

Brittny
There are four main types of potential energy. First, we have chemical energy. Chemical bonds holds atoms together. It takes energy to break these bonds and to move atoms further apart. Energy is released when new bonds form and atoms come closer together. Chemical energy is where the energy in our food comes from. Our body digests food, like this sandwich, and uses the energy to do things.

Another type of potential energy is mechanical energy. Mechanical energy is stored in a device by the application of a force, like the stored energy in a pulled bowstring. Once it is released, stored energy is converted into kinetic energy.

Two other types of potential energy include nuclear energy and gravitational energy. Gravitational energy is related to an object's position in a gravitational field. This tennis ball has a higher potential energy up here than when it falls to earth.

Chris
OK, enough about potential energy. Let's talk about kinetic energy. First of all, there's hydro or wind energy. All objects in motion have kinetic energy that can be transferred to other objects by collisions. For example, the movement of air can turn windmills to pump water or produce electricity.

Speaking of electricity, there's also electrical energy. Electrical energy is the energy associated with the movement of ions and electrons. When electrons flow through wires, we call it electricity. In addition, there's radiant energy, which comes from light waves, x-rays, and microwaves, and it's all around us.

And then, there's thermal energy. Thermal energy is the energy that results from the movement of atoms and molecules and is related to their temperature. The faster the particles move, the greater the amount of energy, and the higher the temperature.

Sound energy is produced by the periodic movement of matter in a medium. Sound can travel through gases, solids, and liquids. Pretty much any matter, but not through the vacuum of space, because there's no matter there. So you know that old sci-fi movie line about how "in space, no one can hear you scream"? Totally true. And a little terrifying.

So there you have it, the main types of potential and kinetic energy. All this talk about energy makes me want to absorb some.

Brittny
Is that my sandwich?!

Chris
Uhh ... potentially ...

The Energy of Toys | A Video

This video explores the chemistry behind simple toys that showcase energy changes. It can supplement an investigation, or may stand on its own to introduce a lesson or extend student learning.

highschoolenergy.acs.org/what-is-energy/what-is-energy.html

Video Transcript

What, exactly, is energy? Because we can't see it, coming up with a formal definition for energy can be tough. However, we can see how energy is transferred from one form to another. For example, put a battery in a flashlight and soon we get light out the other end.

Energy is the ability to do work and make things change.

Best of all, we can understand energy by playing with some classic toys.

Meet my friend Trevor, the amazing drinking bird. The amazing drinking bird is a toy that has been around for over 50 years. It looks like a bird, and will bob up and down and appears to drink out of a cup of water placed in front of it.

But how does it work?

Well, the amazing drinking bird is made out of glass and contains a red liquid. The head and tail are bulbs connected by a narrow glass tube. The only way for liquid to move between the head and the tail is through this narrow tube. A felt-like fuzzy material covers the head and beak.

The red liquid inside the bird is actually methylene chloride, also known as CH2Cl2. Methylene chloride is a very volatile liquid, which means it evaporates easily—easier than water. Most of the air has been removed from the inside of the bird before the liquid is added, making it even easier for the liquid to vaporize.

Wetting the head of the bird starts the action. After wetting the head, the red liquid inside the bird slowly begins to rise from the bottom bulb up to the head. This changes the center of gravity and the bird begins to tip down into the cup of water. As the bird tips over, the liquid then flows back into the bottom bulb, lowering the center of gravity and the bird stands upright. Then the process repeats.

So, where's the chemistry here?

Well, when the bird's beak gets wet, it causes the inside of the head bulb to cool. As the inside of the head cools, the methylene chloride in the head changes from a gas to a liquid, lowering the vapor pressure inside the head. Because the vapor pressure in the tail bulb is now greater than the vapor pressure in the head bulb, the liquid is pushed up into the head bulb.

Another neat toy that can teach us about energy is called a jumping disc.

The disc is about an inch in diameter and made of thin layers of two different types of metals. It is slightly concave. To make it jump, the disc is taken in the hand and rubbed between the thumb and fingers. This warms the disc and makes it click into a convex shape.

When the disc is then placed on the table, it is still for a moment, then it jumps into the air a foot or more.

So where does the energy come from to make the disc jump?

These discs are made from two different types of metal, sandwiched together into a very thin layer. One of the layer tends to expand when its temperature increases while the other layer does not. As the disc is warmed, the temperature sensitive strip of metal grows slightly larger which causes the disc to snap from a concave shape to a convex shape.

When the disc is placed on the slightly cooler countertop, this same layer begins to contract, and the disc snaps back into its original shape. The force of this transformation launches the disc into the air.

So how is energy involved in this toy?

Thermal energy is transferred from your warm hand to the cooler disc. This warms the metal strips in the disc itself. As the temperature increases, the atoms in the metal move faster and spread further apart. The inert layer resists change, due to the way its atoms are arranged.

So, when the warm disc changes its shape, it has a higher potential energy. When it snaps back to its original shape, this potential energy is converted into kinetic energy and the disc jumps.

Although we're more likely to focus on the batteries in our smart phones, these simple toys are a fun reminder that energy is all around us.

Isn't that right, Trevor?

The Amazing Drinking Bird!

By Brian Rohrig

The drinking bird is a mesmerizing little science toy that has fascinated young and old for more than 50 years. It is also known as the happy bird, the dippy bird, the happy dippy bird, and other variations on this same theme. It looks like a bird and will bob up and down as it appears to drink out of a glass of water. As long as it can reach the water, it will bob up and down indefinitely. The drinking bird consists of two elongated glass bulbs that are connected by a straight glass tube extending well into the interior of the bottom bulb. The only way any substance can pass between the two bulbs is through this narrow glass tube. The top bulb is covered with a porous feltlike material that also makes up the beak. On top of the head is a plastic top hat, which is only for decoration. Taped to the bottom chamber are tail feathers, which help it to maintain balance. The whole thing is suspended from plastic legs, with a horizontal piece of metal that acts as a pivot, allowing it to bob up and down.

Inside the drinking bird is a highly volatile liquid known as methylene chloride (CH_2Cl_2). Since methylene chloride is colorless, coloring must be added to enhance the visual effect. This liquid is also highly volatile, meaning it evaporates rapidly due to weak intermolecular bonds in the liquid state. Its boiling point is 39.7 °C (103.5 °F), and its vapor pressure at room temperature is 46 kilopascals (compared to only 3 kPa for water). Methylene chloride is somewhat toxic, so if a drinking bird breaks, care must be exercised in cleaning it up. Methylene chloride is commonly used as an industrial cleaner, degreaser, and paint remover.

After the methylene chloride is added by the manufacturer, most of the remaining air is then vacuumed out. Because a near vacuum now exists within the bird, the highly volatile liquid readily evaporates until the space above the liquid is saturated with vapor. At this point, a dynamic equilibrium is established within the bird between the liquid and the vapor above it. Once equilibrium is established, anytime a molecule evaporates, another molecule will condense, resulting in an overall constant amount of vapor within the bird as long as the temperature stays constant.

To activate the drinking bird, his head is dipped into a glass of water, and he is then set upright in such a position that when he tips his beak, he will be able to reach into the glass of

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water. Once the head is wet, a strange thing immediately begins to happen. Like magic, the fluid begins to rise upward into the head, until his head fills with liquid. The head then becomes top heavy as the center of gravity of the bird is raised. The bird then topples over, takes another drink. As the bird tips over, the liquid flows back to the bottom bulb, restoring the low center of gravity. The bird resumes its upright position, beginning the whole process all over again.

To understand what makes the fluid rise within the bird, think about what happens whenever your own head gets wet. As long as the relative humidity is not 100%, the water will immediately begin to evaporate. And evaporation always causes cooling, because it is an endothermic process. That's why you sweat when you get hot; it's not the sweating itself that cools you, but rather the evaporation of the sweat from your body. Any phase change that requires bond breaking will be endothermic, because energy is required to break bonds. This energy is drawn from the surroundings, thereby causing the temperature of the surroundings to decrease.

Because water evaporates from the head of the bird, the head immediately begins to cool. This is the most crucial point in understanding how the drinking bird works. If you could cool the head another way, the drinking bird would work just the same. When the head begins to cool, some of the vapor within the head will condense into tiny droplets of liquid. A similar process occurs at night when water vapor condenses out of the air as it cools, forming dew on the ground. Because some of the vapor condenses within the top chamber of the bird, there is now less vapor pressure in the top bulb. Less vapor means less pressure. But the vapor pressure within the bottom bulb has not changed. Because the vapor pressure in the bottom bulb is now greater than the pressure in the top bulb, the liquid is forced upward into the top chamber. (Don't say the liquid is sucked up into the top chamber—science never sucks!) Once the bird tips over, vapor from the bottom travels to the top until the pressure in both spheres equalizes and the bird begins the process all over again.

To understand how this pressure differential can cause the fluid within the bird to rise, consider what happens when you use an ordinary drinking straw. When you suck fluid up into the straw, you are creating a region of reduced pressure within the straw. Because outside air pressure is greater, it pushes downward on the surface of the fluid, forcing it up the straw.

Not only is the drinking bird educational, but it can also provide hours of entertainment. Many science museums feature displays of drinking birds. No science classroom would be complete without one. The amazing drinking bird has even appeared in a 1995 episode of *The Simpsons*, where Homer positions a drinking bird in front of his keyboard to help monitor the controls at the Springfield nuclear power plant. The artist Daniel Reynolds spent a small fortune and several years developing an art exhibit comprising a whole flock of giant 6 1/2 feet tall drinking birds, each weighing 3,000 times more than an original drinking bird. They had to be made with a special vacuum attachment in order to work properly.

There are many variations on the drinking bird. They come in a variety of styles and sizes. There is even a drinking giraffe! The very popular, but falsely named "hand-boiler" is nothing more than a drinking bird that is stripped down to the bare essentials. It works either by cooling the top or warming the bottom.

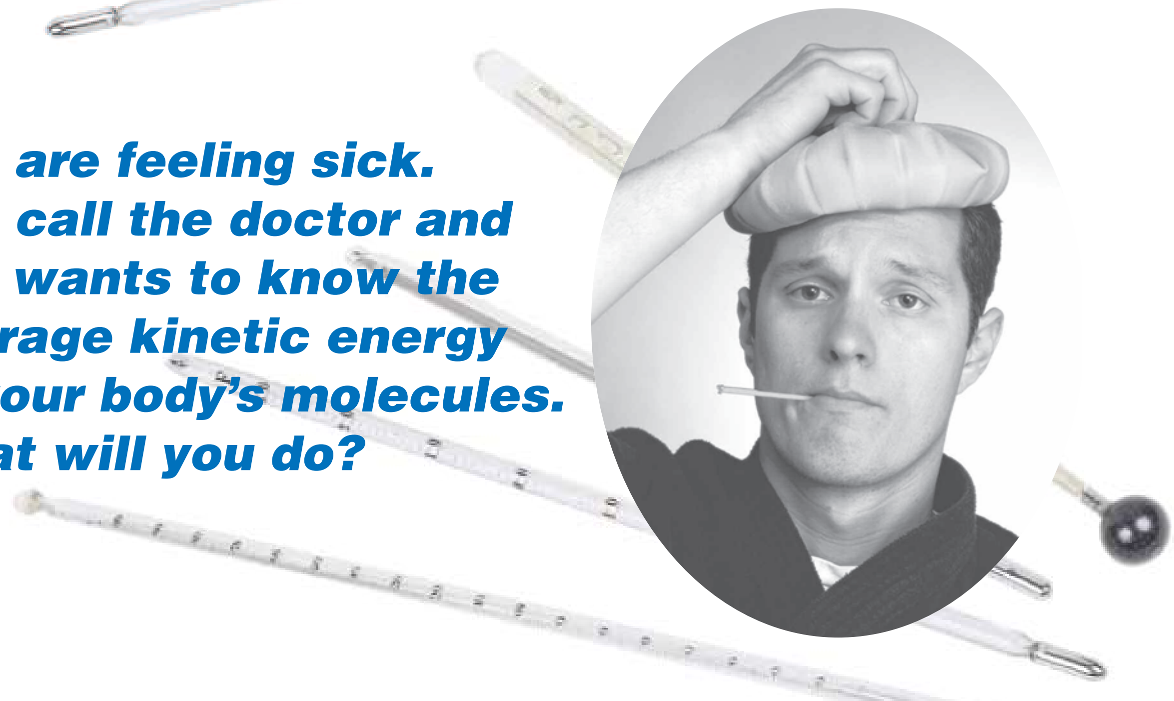
The next time you need a unique gift for the person who has everything, consider a drinking bird. A quick search on the Internet will reveal a plethora of sources. They especially make great gifts for science teachers! ▲

Additional experiments you can perform with the drinking bird:

1. Place a gallon-size freezer bag over an operational drinking bird. He will almost immediately stop drinking. Can you explain why?
2. Have a drinking bird "drink" hot water and then ice-cold water. Which will make the bird drink faster?
3. Have the bird "drink" a liquid with a higher evaporation rate than water, such as rubbing alcohol. Does the drinking bird drink faster?
4. Instead of cooling the head, heat the body, either with your hand or a heat lamp. Does the drinking bird work?
5. If you have a spare drinking bird that you don't mind disfiguring, paint the top bulb silver and the bottom bulb black. Place in a sunny windowsill and watch it bob up and down! No water required.

Brian Rohrig is a chemistry teacher at Jonathan Alder High School in Plain City, OH. His article "There's Chemistry in Golf Balls!" also appears in this issue.

THERMOMETERS



JUPITERIMAGES

You are feeling sick. You call the doctor and she wants to know the average kinetic energy of your body's molecules. What will you do?

By Brian Rohrig

We have all used a thermometer—to check for a fever, record data during a chemistry lab, or to help us decide how to dress before leaving for school in the morning. But have you ever thought about how a thermometer works? And when you measure temperature, just what exactly are you measuring?

The prefix thermo- refers to heat. Thermodynamics is the study of heat. A thermos either keeps heat in or out. You wear thermal underwear to prevent body heat from escaping. Despite its name, however, a thermometer does not actually record heat, but rather temperature. Temperature and heat are two radically different concepts.

Temperature is a measure of the average kinetic energy of the molecules within a substance. When you record the temperature of something, you are making a statement about how fast the molecules are moving. When you are waiting for a bus in the morning in the middle of January, instead of saying, “Boy, its cold out here this morning,” it would be more accurate to say, “Boy, the molecules in the air are moving quite slow this morning!”

Heat vs. temperature

Heat is a little trickier to define. Heat refers to the movement of energy from a substance of high temperature to one of low temperature. Heat always refers to energy in transit. A substance can have a high temperature, but little heat available to transfer. A drop of boiling water contains less actual heat than a bathtub full of water at a lower temperature. Temperature is a measure of only the average kinetic energy of molecules, but because heat depends on the total energy, there is not a simple, universal relation between the two.

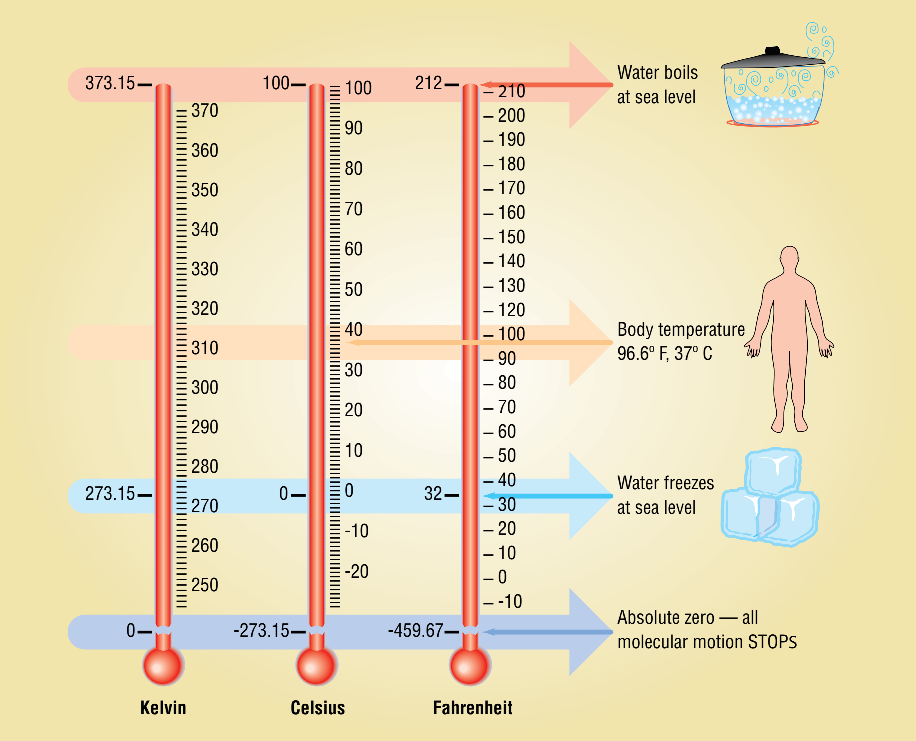
Here’s an everyday example that helps to illustrate the difference between heat and temperature. Consider ice: when you cool a drink using ice, a lot of heat flows from the drink into the ice (so the drink’s temperature falls). But the temperature of the ice does not rise, it stays at 0 °C—the heat goes into breaking the interactions between water molecules to melt the ice (at 0°) to form water (still at 0°). Ice and water at 0° have the same temperature but very different amounts of heat.

Temperature scales

In the United States, most thermometers for everyday use are calibrated in degrees Fahrenheit. Most of the rest of the world measures temperature in degrees Celsius. At one point during the 18th century, there were nearly 35 different temperature scales in use! Many scientists felt the need to devise a uniform temperature scale that would meet widespread acceptance.

One temperature scale that met with some success was the Romer scale, which was first used in 1701. This temperature scale was invented by Ole Christensen Romer, a Danish astronomer whose biggest claim to fame was measuring the speed of light in 1676. His temperature scale set the boiling point of water at 60° and the freezing point at 7.5°. The lowest temperature you could achieve with a mixture of salt and ice was 0°. Because most people from that time period were not too concerned about the temperature of ice and salt, this scale was destined for the dustbin of history.

Daniel Gabriel Fahrenheit, a German physicist, published an alternate scale in 1724. Borrowing from the work of Romer, he set 0 °F as the lowest temperature that could be achieved with a mixture of salt, ice, and ammonium chloride. (It is unclear whether Romer also used ammonium chloride in his experiments, as many of his records were destroyed in a fire.) Fahrenheit set the freezing point of water at 32° and the body temperature of a person at 96°, which he determined by measuring the temperature under his wife's armpit. Each degree of his scale corresponded to one ten-thousandth the initial volume of mercury used in his thermometer. To this day, there is considerable controversy as to how Fahrenheit actually arrived at his temperature scale. He never did reveal exactly how he arrived at the reference points for his thermometer, as he did not want others to construct and sell the thermometers he had spent much of his life perfecting.



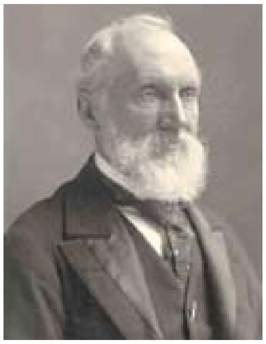
CESAR CAMINERO

A comparison of three temperature scales.

His scale met widespread acceptance because everyone could relate to it, since 0 °F and 100 °F were the lowest and highest temperatures typically experienced on any type of regular basis in Western Europe. If the temperature rose above 100°, you knew it was really hot. If the temperature dipped below 0°, you knew it was quite cold. Whether these points were intentionally chosen to represent these extremes or just happened to work out this way is still being debated today. The biggest problem with this scale was the freezing and boiling points of water were set at 32° and 212°, not exactly round numbers. This was an issue not so much with the general public, but rather with scientists, who tend to obsess over such things. However, others have postulated that placing 180 degrees between the freezing and boiling points of water was not arbitrary but quite rational, as this number represents the number of degrees in half a circle.

To counter this problem, Swedish astronomer Anders Celsius came up with another scale in 1742, setting the freezing and boiling points of water at 0° and 100°, with 100 divisions in between. Hence, it was termed the Centigrade scale, since the prefix centi- represents one-hundredth. Celsius had initially set the freezing point of water at 100° and the boiling point at 0°. This was later reversed after his death. Most countries that have adopted the metric system of measurement use this temperature scale, as it is conveniently broken down into units of 10. In 1948, the Centigrade scale was officially designated the Celsius scale, although some people still use the outdated term.

Anders Celsius



WWW.NPS.GOV

Lord Kelvin

The most scientific scale in use today is the Kelvin, or absolute, temperature scale. It was devised by British scientist William Thomson (Lord Kelvin), in 1848. Because temperature is a measure of molecular motion, it only makes sense that the zero point of your scale should be the point where molecular motion stops. That is exactly what the Kelvin scale accomplishes. 0 Kelvin (K) is the point at which all molecules stop moving. 0 K is known as absolute zero, which has never actually been reached. In 2003 at MIT, scientists came very close to reaching absolute zero, obtaining a frosty temperature of 4.5×10^{-9} K.

The Kelvin scale is primarily used in science, and temperature must be expressed in Kelvin when solving many equations involving temperature, such as the gas laws. But it tends to be too cumbersome for everyday use, since the freezing point of water is 273 K and the boiling point is 373 K.

Types of thermometers

Early thermometers

The first thermometer in modern times was a crude water thermometer believed to have been invented by Galileo Galilei in 1593. In 1611, Sanctorius Sanctorius, a colleague of Galileo's, numerically calibrated the thermometer. Many of these first thermometers used wine, as its alcohol content prevented it from freezing and its red color made it easy to read. However, these first thermometers were very sensitive to air pressure, and functioned as much as a barometer as they did as a thermometer. So eventually, all thermometers were constructed of a sealed glass tube that had all the air removed. Because these vacuum tubes were shut

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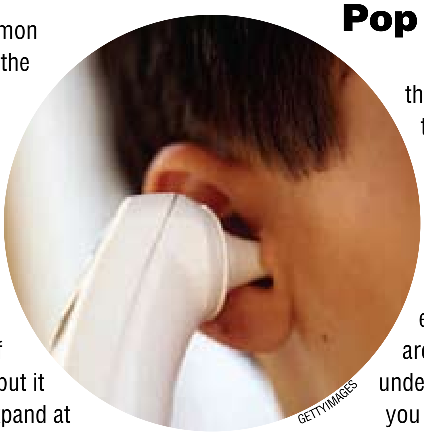
off from the outside atmosphere, changes in air pressure would not affect the temperature reading. In 1709, Fahrenheit invented the alcohol thermometer, and in 1714, he invented the first mercury thermometer. All thermometers work according to the same basic principle: objects expand when heated and contract when cooled.

Bulb thermometers

The most common thermometer is the bulb thermometer, which comprises a large bulb filled with a liquid and a narrow glass tube through which the liquid rises. All liquids expand when heated and contract when cooled (with the exception of H_2O near its freezing point; ice-cold H_2O at 0 °C contracts until 4 °C where it expands like other materials), which explains why the liquid within a thermometer rises as the temperature increases and falls when it decreases. Mercury was the liquid of choice for many years, because it expands and contracts at a very constant rate, making mercury thermometers very accurate. However, because of concerns about mercury toxicity, mercury has often been replaced with alcohol that is colored red. Mercury has a silver color. It freezes at –39 °C, so it cannot be used if temperatures get colder than this.

Bimetallic strip thermometers

Another very common type of thermometer is the bimetallic strip thermometer. This thermometer comprises two different metals, such as copper and iron, which are welded together. Each of the metals used has a different coefficient of linear expansion, or to put it simply, these metals expand at different rates. Connected to this bimetallic strip is a pointer, which points to the correct temperature on the face of the thermometer. Because these metals expand at different rates, when heated, the welded strip of metal will bend. When cooled, it will bend in the opposite direction. A variation of the bimetallic strip thermometer is the thermostat used in homes and automobile engines. These thermostats are made of a thin bimetallic strip, which is fashioned into a coil, making it more sensitive to minor temperature fluctuations.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Infrared thermometer

Infrared thermometers

A fascinating thermometer is the infrared thermometer. This handheld device is used by simply pushing a button as you point it toward an object. A digital readout tells you the temperature. All objects above absolute zero are emitting infrared radiation (IR)—an invisible (to human eyes) form of electromagnetic energy. The infrared radiation we emit is commonly known as body heat. The infrared thermometer has a lens that focuses the infrared energy into a detector, which measures the IR intensity and converts that reading to temperature. Infrared thermometers have a wide variety of applications. They are used by firefighters to detect hot spots in buildings and in restaurants to ensure that served food is still warm. Infrared thermometers are also used for determining the temperature of a human body, automobile engines, swimming pools, hot tubs, or whenever a quick surface temperature is needed.


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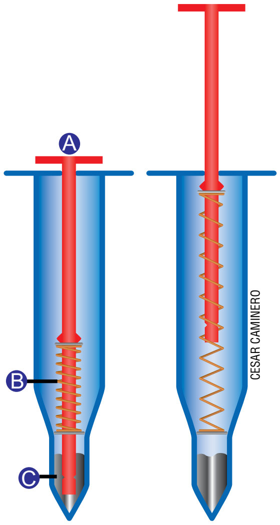
Pop ups

You are cooking that Thanksgiving turkey, and you want to make sure that the inside of the turkey is completely done. To ensure that you are not feasting on undercooked bird, you can use an ingenious device known as the pop-up turkey timer. This instrument is simply stuck into the turkey, and when the turkey is done, a red indicator pops up (A). The little red indicator is spring loaded (B) and is held in place by a blob of solid metal (C). When this metal reaches a temperature of 85 °C, which is the temperature of a fully cooked turkey, it melts, causing the red indicator to pop up.



CESAR CAMINERO

Popup turkey timer

This technology is similar to that used in sprinklers found on the ceilings of many buildings, which actually served as the inspiration for the pop-up turkey timers. When a certain temperature is reached, a metal component within these sprinklers melts, activating the sprinkler. By mixing together different metals, a particular alloy can be created with a desirable melting point. Pop-up timers can be purchased for a wide variety of different types of meat, from ham to hens. You can even buy a pop-up timer for steak, which pops up in increments indicating rare to well done.

And now for something completely different…

Perhaps the most unusual thermometer ever invented is the Galileo thermometer, based on a similar device invented by Galileo. This instrument does not look like a thermometer at all, as it is composed of several glass spheres containing different colored liquids that are suspended in a cylindrical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Galileo thermometer

column of a clear liquid. Attached to each of the colored spheres is a little dangling metal tag with an engraved temperature. The temperature is determined by reading the tag on the lowest floating sphere. As the temperature rises, the spheres will begin to fall one by one. When the temperature falls, the spheres will then rise one by one.

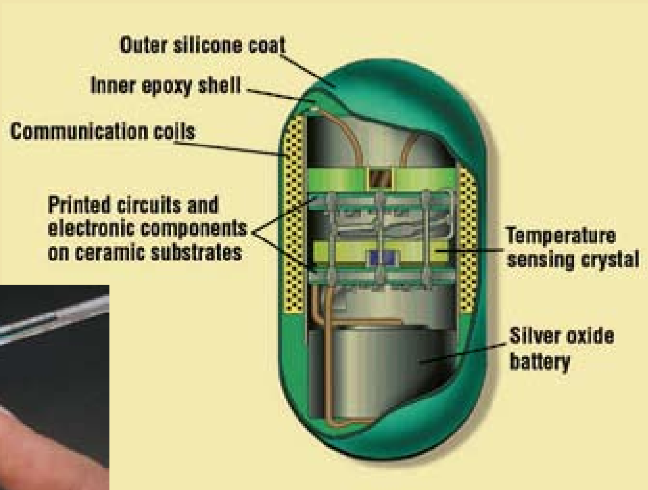
The liquid within each glass sphere is composed of either colored water or alcohol. Each of the spheres is of a slightly different mass, and thus a slightly different density, since the volume of each sphere is the same. Each sphere differs in mass by about 0.006 grams. This difference is accomplished by making each tag a slightly different mass. The clear liquid surrounding the spheres is an inert hydrocarbon-based oil, similar to mineral oil. When this liquid is heated, it expands, becoming less dense. Less dense liquids exert a lesser buoyant force, so the most dense sphere will then sink. If the temperature continues to rise, the molecules of the surrounding liquid will continue to spread apart from one another, causing more spheres to fall. As the liquid cools, its molecules come closer together, exerting a greater buoyant force, causing the spheres to rise. The spheres themselves do not expand or contract nearly as much as the surrounding liquid when heated or cooled, since they are composed of glass, which hardly expands at all when heated.

Even though it looks nothing like a conventional thermometer, the Galileo thermometer still functions according to the same basic principle as most other thermometers: substances expand when heated and contract when cooled.

What's the future for thermometers?

Technology has come a long way since Galileo's day, but his thermometer to this day has a futuristic look to it. Another futuristic thermometer that is available today is the CorTemp thermometer. Developed by Dr. Leonard Keilson of the Applied Physics Laboratory of the Johns Hopkins University in conjunction with NASA, the CorTemp thermometer is swallowed, allowing accurate temperature readings while it travels through, or is stationed at some particular spot in the body. The probe is enclosed in a small pill that is taken internally, while the temperature readings are recorded on a device that is monitored externally.

No matter what device you use to take your temperature when you have a fever, none will make you feel better. But in this technologically advanced world today, your choice of thermometer might bring you a bit of welcomed distraction while measuring the average kinetic energy of your body's molecules.



COURTESY OF CORTEMP

The CorTemp system can measure and record the body temperature and/or heart rate of many athletes on the field during practices or competition. Once the probe is inside the gastrointestinal tract, a crystal sensor vibrates at a frequency relative to the temperature of the body tissues surrounding it. These data are then transmitted harmlessly through the body to the monitor.

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Brian Rohrig teaches at Jonathan Alder High School in Plain City, OH. His most recent *ChemMatters* article, "Glass: More Than Meets the Eye", appeared in the October 2006 issue.

How Do We Use Energy?

One of the best ways to establish a concept is to provide multiple examples of the concept in action. In this section we present a number of activities where students investigate energy through chemical and physical reactions. The activities present chemical processes that will be familiar to students through their everyday lives, but focus on how energy is involved in the changes they observe. These ideas are reinforced with a chemical demonstration, a video, and an article on the chemistry of explosives.

- Lab Investigation—Combustion & Burning
- Demo—Soda Can Steam Engine
- Lab Investigation—Preparation & Combustion of Biodiesel
- Demo—Electrolysis of Water
- Video—Temperature & Energy
- *ChemMatters* Reading—The Explosive History of Nitrogen

Combustion & Burning | A Lab Investigation

Summary

In this investigation, students are challenged to make careful observations about a burning candle to discover the chemical and physical changes that make it work.

Objective

Students will desribe their observations about a burning candle to discover the chemical and physical changes that make it work.

Safety

- Be sure to wear safety goggles while making close observations, tie your hair back if you have long hair, and always be aware of where the flame is with respect to your clothing —do not lean over the flame.
- Burning candles are open flames that can cause burns. Liquid wax is hot and can cause burns to the skin.

Materials for Each Group

- Tea light candle
- Matches
- Small beaker or evaporating dish (such as 50 mL or 100 mL)
- Balance
- Ruler or meter stick
- String (for measuring circumference)
- Stopwatch

Time Required

One class period, approximately 45–50 minutes.

Lab Tips

Tea candles can be purchased in large quantities, as can tapered candles from hardware stores. Candles of different colors, sizes, and shapes can enrich the observations made by the class.

Integrating into the Curriculum

This investigation could fit into a unit on chemical reactions.

PREPARING TO INVESTIGATE

We are all familiar with burning candles, but how much have we really seen while looking at one? In this exercise you will be challenged to observe a burning candle through the lens of chemistry. You may be surprised by how much is really happening in what seems like a familiar process!

First we will try to observe details beyond our usual way of looking at a candle, and then we will look for evidence of the physical or chemical nature of any changes that are taking place. Take care in recording your observations—try to imagine that you have never seen a burning candle before!

GATHERING EVIDENCE

Obtain a tea light candle, matches, and any measuring equipment available to you. Before lighting the candle, make three qualitative observations and three quantitative observations of the candle.

Qualitative observations describe qualities or characteristics, quantitative observations involve quantities, or measurements. You may choose among many properties to observe: color, texture, mass, physical dimensions, density, aroma… or any others you can think of. Use as many of your physical senses as possible, and be sure to include units for any of your quantitative observations.

Record your observations below:

Qualitative observations	Quantitative observations

Use a match to carefully light the candle.

Repeat the process of making qualitative and quantitative observations of the burning candle, this time making six of each kind of observation. (Some of your observations may relate to properties you have already recorded, that may be changing now that the candle is burning.) Once again use all your senses to make qualitative observations. Ask yourself what you can measure for the quantitative observations.

For your observations and measurements, consider not only the physical candle, but also consider its influence on the air in the space surrounding it. Record your observations below:

Qualitative observations	Quantitative observations

Now sketch a detailed drawing of the burning candle, an enlargement of the portion about a centimeter from the top of the candle to the topmost tip of the flame:

Label the following regions of the candle in your drawing: solid wax, liquid wax, region of wick without flame, region of wick with flame, and region of the wick that glows.

Now distinguish as many different aspects of the flame itself and label them in the drawing. When you have completed recording your observations and drawings, prepare to blow out the candle, and be ready to make further observations as you do this!

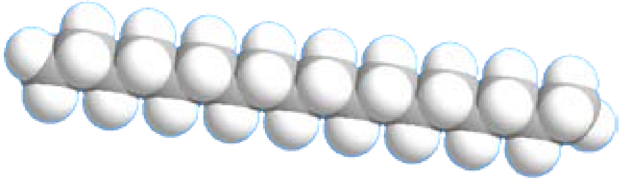
Blow out the candle and record two qualitative and two quantitative observations of the candle immediately after it is extinguished:

Qualitative observations	Quantitative observations

INTERPRETING EVIDENCE

Now that we have paid close attention to the properties of a burning candle, we can begin to interpret what we've seen in terms of phase changes and chemical reactions.

In a phase change (from solid to liquid or from liquid to gas) the particles of a substance (in this case the molecules of wax) will change the arrangement they have with each other. As a solid they will be closely packed together and unable to move freely. In the liquid they will be able to flow around each other but will still be close together, and as a gas they will fly freely apart but they will still be wax molecules and will have the same structure they had as molecules in the solid and the liquid phases:



Waxes from plants and animals have different structures. Here is a molecule similar to wax, $C_{18}H_{38}$*— a material that is normally a major component of the candle wax.*

In the boxes below, use small, single shapes to represent entire wax molecules, and sketch their arrangement in the three physical phases:

Solid	Liquid	Gas

Based on your observations of the burning candle, did you find evidence for the change from the solid wax phase to the liquid wax phase?

$$C_{18}H_{38}\ (s) \longrightarrow C_{18}H_{38}\ (l)$$

Describe this evidence as specifically as possible:

Did you find evidence for the change from the liquid wax phase to the gaseous wax phase?

$$C_{18}H_{38}\ (l) \longrightarrow C_{18}H_{38}\ (g)$$

Describe this evidence as specifically as possible:

A physical candle is composed of wax surrounding a wick. Look up the definition of the verb “to wick,” and write it below:

Describe the wicking process in the burning candle, and be sure your description includes the physical state of the wax and the wick:

As a candle burns, the wax gradually disappears from sight. Is there a physical change in the wax that would account for this disappearance?

If the wax was simply undergoing physical change, would you expect the candle to produce heat and light as it does when a candle burns?

The process of burning (as opposed to evaporating) is a chemical reaction, a chemical change. The wax molecules are undergoing a chemical change; they are changing into different molecules by reacting with a substance in the air. What is this substance in the air that a candle needs to burn?

Interestingly, a candle releases energy through the same kind of reaction that your body uses to obtain energy. Hold your breath for a moment. What is your body craving? This is the same substance the candle needs to burn in a chemical change that we call combustion.

We can represent the combustion of a candle like this:

$$2\,C_{18}H_{38}\ (s) + 55\,O_2\ (g) \longrightarrow 36\ CO_2 + 38\ H_2O$$

Complete the following sentence that restates this equation:

Two molecules of wax react with ______________ molecules of oxygen to produce
____________ molecules of ___________________________________ and _____________
molecules of ______________________.

We are not aware of oxygen in the air because it is an odorless and colorless gas, and at the temperature of a burning candle, both the CO_2 and the H_2O are also gases that are invisible to us. We can see evidence of chemical change, however, if we restrict the amount of oxygen that is allowed to react with the wax.

If less oxygen is available, one possible outcome in the combustion of wax is

$$2\,C_{18}H_{38}\ (s) + 37\,O_2\ (g) \longrightarrow 18\,C\ (s) + 18\,CO_2\ (g) + 38\,H_2O\ (g)$$

Relight the candle and lower the bottom of a small beaker or a clear watch glass into the flame so that the flame touches the surface of the glass. Hold it there for a moment (where it restricts some of the air flow to the flame) and then remove it.

What do you observe on the bottom of the beaker? What could this substance be and where did it come from?

How does this provide evidence that a chemical reaction is taking place?

Wipe the bottom of the beaker clean and place an ice cube inside it. Hold the beaker several centimeters above (not in) the flame and allow the candle to burn. Can you see a substance forming on the bottom of the beaker? What could this substance be and where did it come from? Could it have come from the melting ice inside the beaker?

Consider the flame itself. Is the flame a substance produced by the candle? What observations lead you to think that the flame is matter? In what ways does the flame seem like energy?

Is the following statement true or false: "As a candle burns, wax turns into fire."

If this isn't an accurate statement, can you restate it correctly below?

Finally, let us consider the physical phase of the wax as it reacts with the oxygen to produce carbon dioxide and water. Prepare to observe carefully! Blow out the candle and very quickly bring a match close to the wick. Try to notice exactly where the lit wooden splint is with respect to the wick when the candle reignites.

Do this several times to see if you observe something interesting. Note any evidence you have observed that helps to determine the physical phase of the wax when it reacts with oxygen.

REFLECTING ON THE INVESTIGATION

1. What phase is the wax in when it mixes with oxygen in the air and undergoes a chemical change?
2. Based on your observations, in which phase (solid, liquid, or gas) is the wax most likely to undergo a chemical change?
3. Explain why this is the case based on your sketches of wax in solid, liquid, and gas phases.
4. How might this (your answer to question 3 above) explain a sign that says NO SMOKING near the pumps at a gas station?
5. Look back at your diagrams of the candle flame. Can you identify the region where the wick itself is undergoing combustion? Where is this region?
6. What is fire? Is it matter or is it energy? Explain your answer.
7. In a short paragraph, summarize the changes that a candle undergoes when it burns. Incorporate the following words into the paragraph in a meaningful way: wax, wick, phase change, chemical reaction, vaporize, melt, transforms, solid, liquid, gas, energy, light, and heat.

Extension

Further analysis problem using stoichiometry:

Given the reaction that we are using for the combustion of wax:

$$2\,C_{18}H_{38}\ (s) + 55\,O_2\ (g) \longrightarrow 36\,CO_2 + 38\,H_2O$$

Use stoichiometry relationships to calculate:

1. The volume of oxygen gas needed at STP to burn 37.5 g of wax.
2. The mass of carbon dioxide released when 37.5 g of wax burns.

If less oxygen is available, one possible outcome in the combustion of wax is

$$2\,C_{18}H_{38}\ (s) + 37\,O_2\ (g) \longrightarrow 18\,C\ (s) + 18\,CO_2\ (g) + 38\,H_2O\ (g)$$

If this were to happen, what mass of soot (carbon) is produced for every 1.0 g of wax that burns?

Can you think of another gas that could be produced when there is insufficient oxygen to produce CO_2?

The following links provide further inquiry in different aspects of burning candles:

- This site from NASA helps students gain insight into the characteristic shape of a candle flame by observing how a candle burns in zero gravity.
  - *http://bit.ly/highschoolnrg3*
  - *http://bit.ly/highschoolnrg4*

- Diagrams and explanations of candle combustion from the National Candle Association:
  - *http://bit.ly/highschoolnrg5*

- A *Journal of Chemical Education* article on observing candles that includes a demonstration of the scorch marks left on an index card held in a flame:
  - *http://bit.ly/highschoolnrg6*

TEACHER'S KEY

Reflecting on the Investigation

1. What phase (solid, liquid, or gas) is the wax in when it mixes with oxygen in the air and undergoes a chemical change?

   Gas

2. Based on your observations, in which phase (solid, liquid, or gas) is the wax most likely to undergo a chemical change?

   In the gas phase surrounding the wick.

3. Explain why this is the case based on your sketches of wax in solid, liquid, and gas phases.

   In the gaseous phase the wax molecules have more freedom to mix and collide with oxygen molecules in the air, and react with them to produce water and carbon dioxide.

4. How might this (your answer to question 3 above) explain a sign that says NO SMOKING near the pumps at a gas station?

   Gasoline that has vaporized near the pumps may react with oxygen if exposed to an open flame or a lit cigarette.

5. Look back at your diagrams of the candle flame. Can you identify the region where the wick itself is undergoing combustion? Where is this region?

   At the very tip of the wick, where it is glowing orange, the wick itself is burning, or reacting with oxygen in the air. Melted wax (in the process of changing into gaseous wax) coats the rest of the wick.

6. What is fire? Is it matter or is it energy? Explain your answer.

   Both. What we call fire is what we see and feel when the gaseous reactants and products of combustion release energy in the form of light and thermal energy that heats the surroundings.

Extension

Further analysis problem using stoichiometry:

Given the reaction that we are using for the combustion of wax:

$$2\ C_{18}H_{38}\ (s) + 55\ O_2\ (g) \longrightarrow 36\ CO_2 + 38\ H_2O$$

Use stoichiometry relationships to calculate:

1. The volume of oxygen gas needed at STP to burn 37.5 g of wax.

   90.8 L O_2

2. The mass of carbon dioxide released when 37.5 g of wax burns.

   117 g CO_2

If less oxygen is available, one possible outcome in the combustion of wax is

$$2\ C_{18}H_{38}\ (s) + 37\ O_2\ (g) \longrightarrow 18\ C\ (s) + 18\ CO_2\ (g) + 38\ H_2O\ (g)$$

If this were to happen, what mass of soot (carbon) is produced for every 1.0 g of wax that burns?

0.42 g C

Can you think of another gas that could be produced when there is insufficient oxygen to produce CO_2?

CO

Soda Can Steam Engine | A Demonstration

Summary

In this demonstration, students consider the concepts of energy transformations and conservation as they observe a simple steam engine made from an aluminum soda can.

Objective

Students learn about energy transformations and the concept of conservation of energy.

Safety

- Be sure you and the students wear properly fitting goggles.
- Handle the hot soda can with tongs. Steam from the can could scald your hand and skin.

Materials

- Unopened 12-oz aluminum soda can (using diet soda will help avoid potential stickiness)
- Thumbtack
- Access to a sink
- Wash bottle
- Water
- 10-mL graduated cylinder
- String
- Ring stand
- Ring
- Bunsen burner or hot plate
- Tongs

Time Required

Part of one class period, approximately 10–15 minutes.

Integrating into the Curriculum

This demonstration could fit into a unit on phase changes, thermochemistry, or technology.

PREPARATION

1. Working over a sink, create a hole in the middle of the side of an unopened soda can using a thumbtack. Continually shake the can, using the pressure generated by carbonation to force the soda out through the hole.
2. When the can is empty, create a second hole on the side opposite the first.
3. Rinse the can with a wash bottle, squirting water into one of the two holes.
4. Once you have emptied and rinsed the can, use the wash bottle to add about 10 mL water (or enough so you hear it sloshing around inside the can).
5. Use the thumbtack to slant each of the two holes in the same direction, tangential to the can's surface. These holes act as "jets" in order to propel the can.

DEMONSTRATION

1. Shake the can so that students can hear the water sloshing around inside the can. Allow them to observe that the pull tab is unopened and that there are two holes in opposite sides of the can.
2. Tie a string to the can's unopened pull tab and secure the can to a ring attached to a ring stand at a height that allows a Bunsen burner or hot plate to warm it.
3. While waiting for the can and water to heat, ask students to predict what will happen.
4. Gently warm the can and water. Water vapor will eventually exit both holes and should result in a net force that sets the can spinning.

DISCUSSION WITH STUDENTS

1. Ask students to describe any energy transformations present in the demonstration.
2. Challenge students to explain why the can behaves as it does. As long as the water is heated gently, no condensed water vapor is visible as steam exits the can.

EXPLANATION

Energy may go through several conversions before we actually use it to do work. Energy is not "used up" in any conversion; it is simply changed from one form to another. Like matter, the total energy is conserved.

Several energy conversions are involved in the demonstration. Chemical energy from the fuel used in the Bunsen burner or the source that produced the electricity (some possibilities are hydroelectric, solar, wind, geothermal, nuclear, and petroleum energy sources) powering the hot plate is used to warm the water and can. Some of the energy goes to overcoming the attractions among the liquid water molecules so they vaporize to produce warm water vapor. The molecules of the warm water vapor have high kinetic energy, which increases the pressure inside the can and thus forces some of the gas out through the holes in the sides of the can. As these jets of gas leave the can, they push on the air outside the can and create an opposite push on the can, which then is partially converted to kinetic energy as the escaping water vapor causes the can to rotate. Some of the energy from the escaping water vapor is also partially converted to potential energy as the string attached to the spinning can twists.

EXTENSIONS

Instructors may also wish to discuss the idea that no energy conversion is 100% efficient in transforming one form of energy to another desirable form. Thus, some useful energy is always "lost" whenever energy is converted from one form to another. The energy itself is not destroyed, but becomes unavailable to do useful work. For example, some of the electrical energy used in the operation of a hair dryer is converted to sound energy, which does not help to perform the work of drying your hair.

Students can investigate energy conversions in everyday items. Some suggestions are light bulbs, batteries, toys, car engines, etc.

Students could research the early steam engine invented by Hero of Alexandria in the first century A.D. and compare it to the soda can steam engine.

ADDITIONAL RESOURCE

"Demonstration Idea," *Chemistry in the Community*, 6th ed., Teacher's ed., New York: W. H. Freeman and Company/BFW, 2012, p 361.

Preparation & Combustion of Biodiesel | A Lab Investigation

Summary

In this investigation, students use canola oil, typically used in the kitchen, to produce biodiesel by reacting it with methanol using a potassium hydroxide catalyst. They then burn a sample of the biodiesel produced. The sample is placed underneath a soda can that contains a measured amount of cold water. The initial and final temperatures of the water are recorded. The collected data is used to calculate the heat of combustion of the biodiesel, which is compared to the theoretical heat of combustion of diesel fuel from crude oil. The concepts of renewable and non-renewable fuels are also discussed.

Objective

Students learn about a reaction used to produce biodiesel, heat of combustion, and the concepts of renewable and non-renewable fuels.

Materials for Each Group

- 25-mL graduated cylinder
- 10-mL graduated cylinder
- Canola oil (25 mL)
- Methanol (4 mL)
- Plastic container with a screw-top lid (needs a capacity of at least 50 mL)
- Dropper
- 9 M potassium hydroxide (KOH) (12.5 g KOH in 25-mL solution)
- Sodium chloride (0.5 g)
- Large test tube (needs a capacity of at least 30–35 mL)
- Cork or stopper to seal large test tube
- Test tube rack
- 12-oz empty, clean aluminum soft drink can with pull tab
- Ring stand
- Ring
- Thermometer
- Stirring rod
- Matches
- Balance
- Tea light candle with metal cup and wick
- Watch glass

Safety

- Be sure you and the students wear properly fitting goggles.

- Burners must not be used, since many reagents are flammable.

- Methanol particularly poses a serious fire hazard, and its flame is almost invisible. Avoid flames or sparks. Methanol is also toxic by ingestion. Skin contact causes dermatitis. Work in a well-ventilated area.

- When working with alcohols such as methanol, the National Science Teachers Association recommends that "the primary reagent alcohol container be kept in the chemical storeroom. The minimum quantity of alcohol needed for the experiment should be available to students":

  http://bit.ly/highschoolnrg7

- *Extreme caution* should be used with 9 M KOH. It is caustic and corrosive. The instructor must dispense drops of the concentrated base directly into each group's container. 9 M potassium hydroxide is caustic and corrosive. Avoid skin contact. In the event of skin or eye contact, rinse well with water. Have contact area evaluated by qualified medical personnel.

- Instructors should follow their own state's rules for proper disposal of waste. The waste layer in the investigation contains mostly glycerol, excess methanol, and potentially unreacted potassium hydroxide. A useful document is "Biodiesel Safety and Best Management Practices for Small-Scale Noncommercial Use and Production" from Penn State's College of Agricultural Sciences:

  http://bit.ly/highschoolnrg8

Time Required

Two class periods, approximately 45–50 minutes each. The biodiesel reaction mixture must sit for at least 30 minutes or overnight.

Pre-Lab Discussion

This investigation introduces the concept of heat of combustion of a fuel. It also highlights the difference between renewable and non-renewable fuel sources, in particular, biodiesel and diesel obtained from crude oil. A discussion of the basic organic structures presented in the investigation would be helpful.

Lab Tips

Instead of shaking the bottle with the reaction mixture for 10 minutes, a magnetic stir bar apparatus could be used.

One option for the plastic containers with screw-top lids are "baby soda bottles," also known as soda bottle preforms. These are sold by science supply companies and look like large plastic test tubes with 2-liter soda bottle screw-top lids.

The reaction mixture is transferred to a test tube after mixing so it is easier to see the two layers separated and to decant the top biodiesel layer. Separatory funnels could be used instead if available. The time needed for the investigation can be shortened if a centrifuge is available; portions of the reaction mixture can be centrifuged to obtain the two separated layers rather than leaving them to separate overnight.

Chilled water should be used in the combustion portion of the investigation. Ice can be added to cool the water if needed. However, unmelted ice should be removed from the water before using it in the soft drink can.

Instructors may wish to use only a few 25-mL graduated cylinders that would be reserved for this activity, as they will become coated with oil and may be difficult to clean thoroughly. Student groups can share the cylinders.

Integrating into the Curriculum

This investigation would fit into units on chemical reactions, thermodynamics, combustion, and green chemistry.

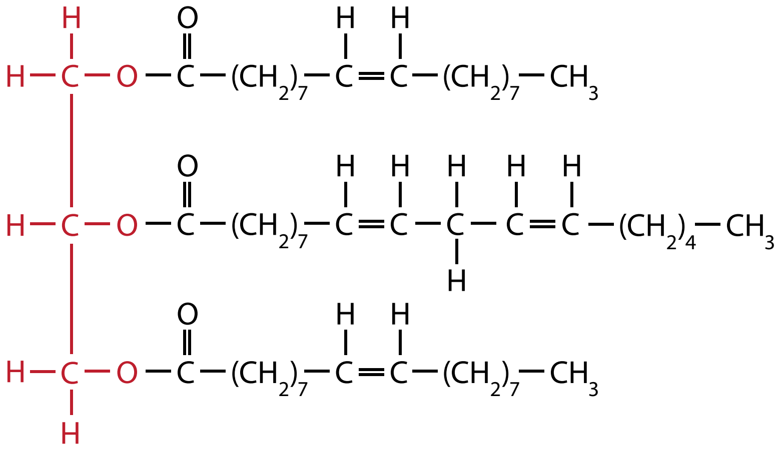
PREPARING TO INVESTIGATE

How did you get to school today? Walk? Ride a bike? Catch a bus? Drive a car? What fuel powered your method of transportation? If you walked or biked, the fuel might have been a bowl of breakfast cereal. Many school buses use diesel fuel, which is made from crude oil that was formed over an incredibly long time by the decomposition of plants and animals. What if instead you could fill your vehicle's tank and run it with recycled cooking oil? Sewage sludge? Pond scum? It's not as crazy as it may at first sound.

These sources of fat or plant oil are potential starting materials for biodiesel, a type of renewable fuel that can be used in diesel engines. Renewable energy sources are those that can be replenished in a short period of time. The materials listed above can be obtained on a much smaller time scale than crude oil; for example, this can be as short as a single growing season for plants. Crude oil resources, on the other hand, generate over such a long time that they are considered non-renewable.

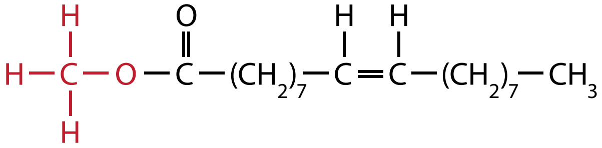
There are various issues to consider with the different possible materials. For example, certain products that would normally be considered waste, such as used cooking oil headed for the landfill, could be turned into fuel instead. However, even if all the used cooking oil was turned into biodiesel, it would be only a tiny portion of the diesel fuel used in a year. The reaction to form biodiesel from a starting product might create its own difficulties. For example, the reaction that you will use in this investigation produces both biodiesel and glycerol. But, what to do with all the leftover glycerol? It can be used in an application such as soapmaking, but with the large amounts being generated, research is being done for other ways to use it.

Canola oil, which comes from the crushed seeds of the canola plant, can be used as a starting material for biodiesel. The structure below shows a typical fat found in canola oil:



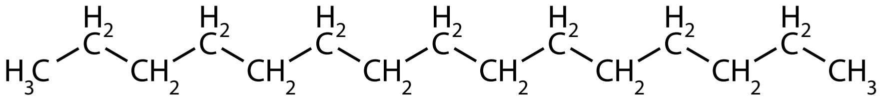
The red portion of the structure represents the glycerol backbone of the fat, while the long portions are the fatty acids that serve as a source of carboxylic acids in the preparation of biodiesel. Biodiesel synthesis breaks the bonds between the glycerol and fatty acids and adds a methyl group (one carbon atom bonded to three hydrogen atoms, —CH_3) to the end of the fatty acid.

The other reaction product is glycerol, which can be used in soapmaking. One of the three biodiesel molecules formed from the canola oil molecule above would be:



$$H{-}\underset{\underset{H}{|}}{\overset{\overset{H}{|}}{C}}{-}O{-}\overset{\overset{O}{\|}}{C}{-}(CH_2)_7{-}\overset{\overset{H}{|}}{C}{=}\overset{\overset{H}{|}}{C}{-}(CH_2)_7{-}CH_3$$

Engines designed to run on diesel fuel made from crude oil can also run on biodiesel fuel with little to no modification, although a blend of diesel and biodiesel is typically used. Diesel fuel is a mixture of hydrocarbons, molecules composed of hydrogen and carbon. Diesel contains molecules with carbon chains of 10–24 carbons. For example, one such molecule might be:



In this investigation, you will use canola oil to prepare biodiesel fuel by combining it with methanol and a potassium hydroxide catalyst. Then, you will combust a sample of the fuel.

GATHERING EVIDENCE

Biodiesel Preparation

1. Use a graduated cylinder to measure 25 mL of canola oil. Pour it into a plastic container with a screw-top lid. Record your observations of the oil's appearance (color, viscosity, odor).
2. Add 4 mL of methanol to the oil in the container. (Caution: Methanol is flammable and toxic. No sources of open flame should be nearby.)
3. Ask your instructor to slowly add 5 to 6 drops of 9 M potassium hydroxide to the liquid in the container. (Caution: Potassium hydroxide is corrosive.)
4. Tightly cap the plastic container. Shake vigorously for 10 minutes.
5. Add 0.5 g sodium chloride to the container. Re-cap and shake vigorously for several seconds.
6. Transfer the liquid to a large test tube (should be able to hold at least 30–35 mL). Stopper the top of the test tube with a cork or rubber stopper.
7. Allow the test tube to sit in a rack for 30 minutes or overnight. What happens to the contents of the test tube?

Biodiesel Combustion

1. Set up the apparatus pictured in the top photograph to measure the heat of combustion of the biodiesel you prepared.
    - Slightly bend up the pull tab of an empty, clean aluminum soft drink can.
    - Slide a glass stirring rod through the top hole of the pull tab.
    - Hold the glass stirring rod horizontally and set it on a ring attached to a ring stand so the soft drink can is suspended underneath it.
    - Raise or lower the ring stand so the bottom of the can is ~2 cm above the wick of the metal sample cup you will use to burn the biodiesel.

2. Prepare the sample of biodiesel for combustion as pictured in the bottom photograph.
    - Take a tea light candle in its metal cup. Remove the candle from the cup.
    - Remove the metal circle and its attached wick from the bottom of the candle.
    - Set the candle aside. Place the metal circle and its attached wick back in the metal cup, so the wick stands upright.
    - Decant ~5 mL of the biodiesel layer (top layer) from the test tube or remove ~5 mL of the top layer using a dropper. Place this sample in the metal cup.
    - Dispose of the remaining layer according to your instructor's directions.

3. Record your observations of the sample's appearance (color, viscosity, odor). Measure and record the initial weight of the biodiesel sample with the cup and wick to the nearest 0.1 g.
4. Weigh ~100 g of cold water, recording the weight to the nearest gram. Pour the water into the soft drink can. Measure and record the initial temperature of the water to the nearest degree Celsius.
5. Ignite the biodiesel sample using the wick. Once it is ignited, immediately move the metal cup underneath the soft drink can.
6. As the water in the can heats, stir it gently. Allow the biodiesel sample to burn for ~5 min.
7. Extinguish the flame by placing a watch glass over the metal cup.
8. Measure and record the highest temperature reached by the heated water to the nearest degree Celsius.

9. Allow the metal cup and sample to cool. Measure and record the final weight of the biodiesel sample with the cup and wick to the nearest 0.1 g.

ANALYZING EVIDENCE

1. Using the temperature and weight data from heating the water in the can, calculate how much thermal energy was used to heat the water. The specific heat capacity of water is 4.18 J/(g · °C), meaning it takes 4.18 J to raise the temperature of 1 g of water by 1 °C.
2. Calculate the heat of combustion in kJ/g for the sample of biodiesel you burned. The heat of combustion is the quantity of thermal energy given off when a certain amount of a substance burns. Assume that all of the energy released by the burning biodiesel is absorbed by the water.
3. Petroleum diesel (from crude oil) produces 43 kJ/g of thermal energy when burned. Compare this to the thermal energy your biodiesel sample produced when it was burned.
4. Compare your calculated heat of combustion with those calculated by the rest of the class. What is the class mean?

INTERPRETING EVIDENCE

1. What evidence was there that a reaction occurred during the biodiesel preparation?
2. Look again at the structures of two possible biodiesel and diesel molecules in the Preparing to Investigate section. What similarities do the molecules have that would help to explain how biodiesel can be used with little to no modification in diesel engines?
3. Draw the balanced equation for the reaction to form biodiesel from canola oil.

Cross Link
Students can get more practice with calculations like these in *The Energy Efficiency of Heating Water* on page 75.

REFLECTING ON THE INVESTIGATION

1. Is biodiesel a “better” fuel than diesel fuel from crude oil? Explain and support your answer.
2. The combustion of biodiesel produces carbon dioxide (CO_2), a greenhouse gas. However, biodiesel is labeled as “carbon-neutral,” meaning it does not increase the net amount of carbon (carbon dioxide) in the atmosphere. Why?
3. What is a potential concern with diverting food crops to the production of biodiesel?
4. Suppose that you wanted to make biodiesel for your own car using the method in this investigation. What challenges might you encounter in scaling up the process?

TEACHER'S KEY

Analyzing Evidence

1. Using the temperature and weight data from the heating of the water in the can, calculate how much thermal energy was used to heat the water. The specific heat capacity of water is 4.18 J/(g · °C), meaning it takes 4.18 J to raise the temperature of 1 g of water by 1 °C.

   Answers will vary. A sample calculation is:

   Data:

   Mass of biodiesel burned: 3.4 g (difference between biodiesel sample, metal cup, and wick before and after burning)

   Mass of water: 1.00 x 102 g water

   Initial water temperature: 5 °C

   Final water temperature: 67 °tC

   $E = mC\Delta T = (1.00 \times 102\ g)(4.18\ J/(g \cdot {}^\circ C))(67\ {}^\circ C - 5\ {}^\circ C) = 26{\times}10^3\ J$

2. Calculate the heat of combustion in kJ/g for the sample of biodiesel you burned. The heat of combustion is the quantity of thermal energy given off when a certain amount of a substance burns. Assume that all of the energy released by the burning biodiesel is absorbed by the water.

   Answers will vary. The method of gathering data for the heat of combustion is somewhat inefficient. A sample calculation, using the data from Analyzing Evidence question 1, $26{\times}10^3$ *J, or 26 kJ, is given off by the burning biodiesel. The heat of combustion = 26 kJ / 3.4 g = 7.6 kJ/g*

3. Petroleum diesel (from crude oil) produces 43 kJ/g of thermal energy when burned. Compare this to the thermal energy your biodiesel sample produced when it was burned.

   Gram for gram, biodiesel produces less energy.

4. Compare your calculated heat of combustion with those calculated by the rest of the class. What is the class mean?

   Answers will vary, depending on class data.

Interpreting Evidence

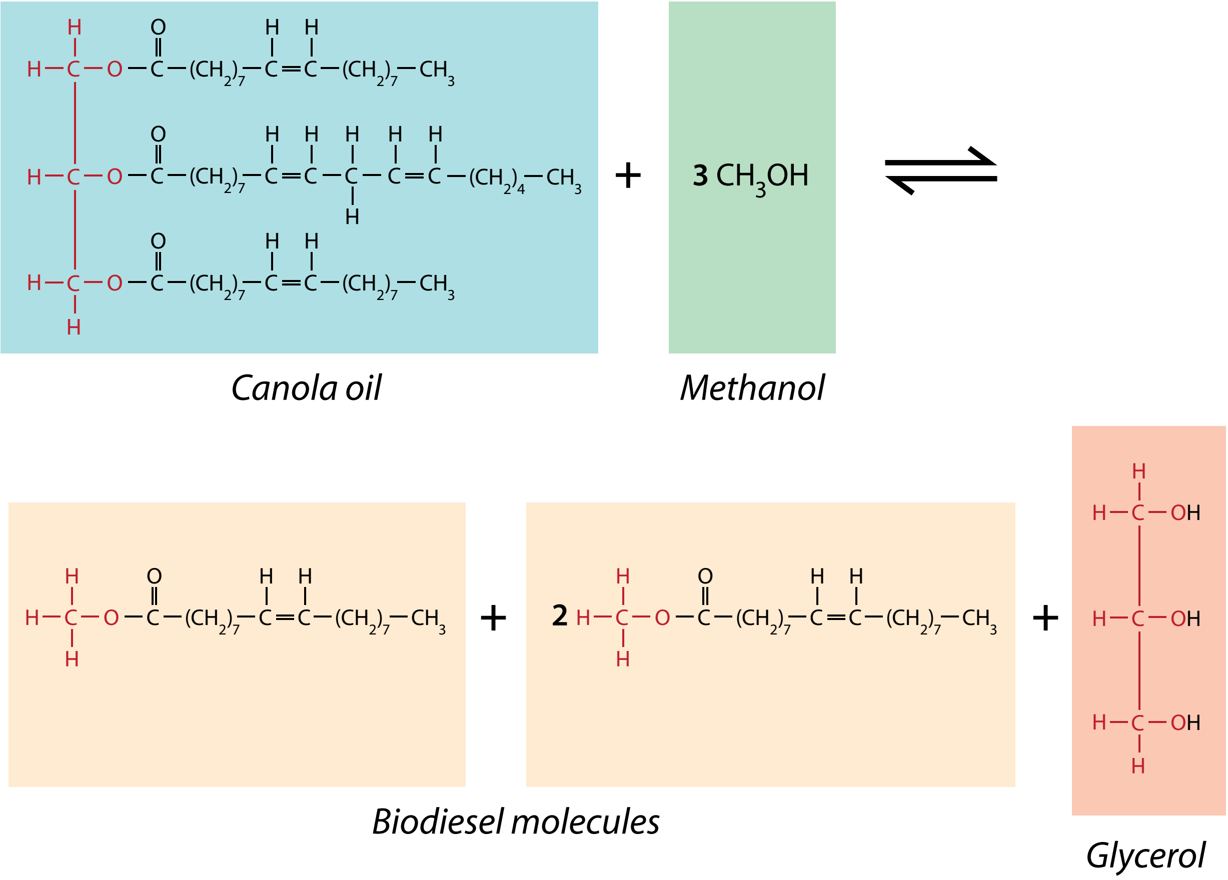
1. What evidence was there that a reaction occurred during the biodiesel preparation?

   Two layers formed in the reaction mixture.

2. Look again at the structures of two possible biodiesel and diesel molecules in the Preparing to Investigate section. What similarities do the molecules have that would help to explain how biodiesel can be used with little to no modification in diesel engines?

   They are both largely chains of carbon atoms of similar lengths. The biodiesel molecule has mostly hydrogen atoms bonded to the carbon atoms like the diesel molecule, but with some differences.

3. Draw the balanced equation for the reaction to form biodiesel from canola oil.



Reflecting on the Investigation

1. Is biodiesel a "better" fuel than diesel fuel from crude oil? Explain and support your answer.

   Answers will vary. A "yes" answer could be supported with the idea that it is a renewable source of fuel. A "no" answer could be supported with the idea that it has a smaller heat of combustion.

2. The combustion of biodiesel produces carbon dioxide (CO_2), a greenhouse gas. However, biodiesel is labeled as "carbon-neutral," meaning it does not increase the net amount of carbon (carbon dioxide) in the atmosphere. Why?

   The combustion does release carbon dioxide. However, the carbon from the biodiesel was previously absorbed from the atmosphere by the plants used to produce the biodiesel. The release is offset by the absorption.

3. What is a potential concern with diverting food crops to the production of biodiesel?

   Diverting food crops to the production of biodiesel can raise food prices for the consumer as well as possibly creating shortages of food.

4. Suppose that you wanted to make biodiesel for your own car using the method in this investigation. What challenges might you encounter in scaling up the process?

   Some challenges could be difficulties in shaking a large mixture; storage of the starting and finished materials and by-products, including a corrosive, highly basic solution; and obtaining necessary amounts of the starting materials.

Post-Lab Discussion

The viability and practicality of using biodiesel as a replacement fuel could be at the center of a rich discussion, particularly if students did further research on the topic. Additional biodiesel starting materials could be discussed, such as soybeans.

Students may be interested in learning more about how vegetable oils such as canola oil can be used directly in diesel engines. Some modifications are made to the engines. For example, the oil must be heated before it enters the engine, due to its higher viscosity than diesel fuel. This could be combined with a discussion of gelling that can occur in diesel and biodiesel fuels at low temperatures.

Students could explore the reaction stoichiometry semi-quantitatively, including an estimation of how many moles of oil and methanol are used, and which is in excess.

Extensions

1. The American Chemical Society (ACS) textbook *Chemistry in the Community* has an excellent series of questions where students decide whether it is viable for biodiesel made from soybeans to completely replace diesel made from crude oil. (*Chemistry in the Community*, 6th ed., New York: W. H. Freeman and Company/BFW, 2012, pp 374–375)
2. Both biodiesel and diesel fuels can "gel" at low temperatures, making modifications necessary for areas that experience cold weather. Students could determine the temperature at which their biodiesel sample gels and decide if it would be a reasonable fuel for their local area. Place a small sample (~1 mL) in a test tube. Leave the test tube in a freezer until the biodiesel gels, approximately 15 minutes. Remove the test tube. Gently stir the sample with a thermometer and observe the temperature at which the sample becomes liquid again.
3. Use a "fire syringe" to demonstrate the principle of the diesel engine, that thermal energy is produced when a gas is rapidly compressed and can ignite fuel. The product is available from science supply companies, such as *www.teachersource.com/product/fire-syringe-demo*. The link includes a video showing the fire syringe in action; see another example at *www.youtube.com/watch?v=MnpvQvCTj90*.
4. The ACS has a sustainable energy section online with links and podcasts about different biofuel sources, such as chicken feathers, and more at *http://bit.ly/highschoolnrg9*.
5. Students could read articles related to the topic from *ChemMatters*, an ACS magazine for high school chemistry students:
    - Hill, M. "From Fish Tank to Fuel Tank," *ChemMatters*, 2012, 30 (2), pp 12–14.
    - Nolte, B. "Tanking Up with Cooking Oil," *ChemMatters*, 2011, 29 (2), pp 5–7.
    - Kirchhoff, M. "Do You Want Biodiesel with That?" *ChemMatters*, 2005, 23 (2), pp 7–9.

Additional Resources

- "D.8 Synthesizing and Evaluating Biodiesel Fuel," *Chemistry in the Community*, 6th ed., New York: W. H. Freeman and Company/BFW, 2012, pp 371–373.
- "C.4 Combustion," *Chemistry in the Community*, 6th ed., New York: W. H. Freeman and Company/BFW, 2012, pp 337–341.
- "Biodiesel: Using Renewable Resources," *Introduction to Green Chemistry*, Washington, DC: American Chemical Society, 2002, pp 13–22.

Electrolysis of Water | A Demonstration

Summary

In this demonstration, students investigate the idea that energy from a battery can be used to drive a chemical reaction that does not happen spontaneously, such as the splitting of water molecules to produce hydrogen and oxygen gases.

Objective

Students learn about using an outside energy source to drive a chemical reaction.

Safety

Be sure you and the students wear properly fitting goggles.

Materials

- 9-volt battery
- Two metal thumbtacks
- Water
- Epsom salt ($MgSO_4 \cdot 7\ H_2O$)
- Clean, empty, clear and colorless plastic water bottle with cap with the label removed
- Scissors
- Black permanent marker
- Beaker or plastic cup
- Paper towels or modeling clay

Time Required

Part of one class period, approximately 10–15 minutes.

Demo Tips

Epsom salt can be purchased locally in drugstores, where it is sold as a laxative and as a material to add to a warm water bath for treating aches and pains. Plastic containers other than a water bottle can be used, as long as the container has a flat bottom for easy connection to the battery contacts and the plastic is thin enough to easily push the thumbtacks through. For example, small plastic condiment cups could be used.

Integrating into the Curriculum

This demonstration could fit into a unit on chemical reactions or thermochemistry.

PREPARATION

1. Remove the cap from a clean, empty, clear and colorless plastic water bottle. Turn the lid over so that the top of the lid touches the two contacts of a 9-volt battery. Center the lid over the two contacts. Using a black permanent marker, make two dots on the inside of the lid, one over the center of each contact.
2. Place the lid on a hard surface with the top of the lid facing up. Push a metal thumbtack into the top of the lid directly over one of the dots. Push a second thumbtack into the lid directly over the second dot. The two thumbtacks should not touch.
3. Using scissors, cut off the top half of the bottle, so that it looks like a funnel. Screw the lid back onto the bottle.

DEMONSTRATION

1. Show students the top half of the bottle, particularly the two metal thumbtacks pushed through the lid.
2. Fill the bottom half of the cut water bottle approximately half full of water. Add about a teaspoon of Epsom salt. Swirl to stir until most of the salt dissolves.
3. Pour the Epsom salt solution into the top half of the bottle (hold so the lid faces down).
4. Ask a volunteer student to observe the solution, particularly the two points of the metal thumbtacks. Is there any evidence of a reaction occurring?
5. Place the two metal thumbtacks so that each thumbtack touches one of the contacts on the 9-volt battery. Ask the volunteer to observe the solution again. Is there any evidence of a reaction occurring?
6. The battery can be placed into the bottom of a beaker or clear, colorless plastic cup, held upright with paper towels or modeling clay. The top of the bottle can then be rested on top of the battery with more stability if the instructor wishes to show it up close to the rest of the class.

DISCUSSION WITH STUDENTS

1. What evidence is there that a reaction occurs when the battery contacts touch the metal thumbtacks?

   Bubbles of one or more gases are produced and rise from each of the thumbtacks.

2. Is there any difference in the amount of gas produced at each battery contact?

   Yes, more bubbles are produced at one of the connections.

3. Could the difference described in question 2 be due to a difference in the tacks? What could one do with the apparatus to potentially provide support for the idea that the battery contact is responsible?

   One could rotate the battery contacts so that each contact is now touching the opposite thumbtack. The battery contact that produced more bubbles the first time should still produce more bubbles when touching either thumbtack.

4. What is the purpose of the battery?

   The battery drives a reaction that we did not observe happening spontaneously without this additional outside energy source.

5. Why is Epsom salt added to the water?

   Epsom salt is an electrolyte and serves to conduct electricity through the solution. Electricity from the battery is able to travel through the solution and between the two thumbtacks.

6. What is the balanced equation for the breaking apart of water molecules?

   $$2\,H_2O\,(l) \longrightarrow 2\,H_2\,(g) + O_2\,(g)$$

7. If the reaction we are observing is the breaking apart of water molecules, which battery contact might be producing which gas and why?

   The reaction for the breaking apart of water molecules shows two gases being produced, H_2 *and* O_2*. In the equation, twice as much hydrogen gas is produced than oxygen gas. The battery contact that produces more bubbles would be the one producing hydrogen gas.*

8. What could be done to gather further evidence that there are two gases produced and that they are H_2 and O_2?

   One could collect the gases produced and test them for known properties of H_2 *and* O_2 *gases.*

EXPLANATION

Separating the word “electrolysis” into its component parts summarizes its meaning—using electricity (electro-) to break apart (-lysis) something. In this demonstration, the electricity supplied by a 9-volt battery is used to break apart water molecules, overall producing hydrogen and oxygen gases.

The idea of including energy as a reactant or a product in a chemical equation can be used to illustrate the need for an additional outside energy force to drive a reaction that normally does not happen spontaneously and to connect the demonstration to the concepts of exothermic and endothermic reactions.

For example, *Chemistry in the Community* describes:

> …if a particular chemical reaction is exothermic (releasing thermal energy), then the reverse reaction is endothermic (converting thermal into potential energy). For example, burning hydrogen gas—involving the formation of water—is exothermic. The energy released by formation of H–O bonds in water molecules is greater than that required to break bonds in H_2 and O_2 molecules:
>
> $$2\ H_2 + O_2 \longrightarrow 2\ H_2O + \text{energy}$$
>
> Therefore, the separation of water into its elements—the reverse reaction—must be endothermic, the quantity of energy equal to the quantity released when water is formed from gaseous H_2 and O_2.
>
> $$\text{Energy} + 2\ H_2O \longrightarrow 2\ H_2 + O_2$$

EXTENSIONS

Challenge students to design a method for collecting the gases in the demonstration and to describe how they could be tested to determine if they are H_2 and O_2, and which is which. Glass tubes filled with the Epsom salt solution can be inverted over the two thumbtacks to collect the gases; it can take a substantial amount of time to collect an appreciable amount of gas.

Showing the reverse reaction, that of mixing hydrogen and oxygen gases and combusting them, could be used in connection with this electrolysis demonstration. This explosive, exothermic reaction should be used with appropriate safety precautions and equipment.

Another common setup used for showing water electrolysis uses two graphite pencils sharpened on both ends as electrodes. Directions written for students to carry out this investigation are available for free as part of the American Chemical Society curriculum "Middle School Chemistry" in Lesson 4.4—*www.middleschoolchemistry.com/lessonplans/chapter4/lesson4*.

ADDITIONAL RESOURCES

- *Chemistry in the Community*, 6th ed., New York: W. H. Freeman and Company/BFW, 2012, pp 335–336.
- ChemTeacher Demonstrations: Demonstrate Water Electrolysis in a Simple Way. *http://bit.ly/highschoolnrg10*

Temperature & Energy | A Video

This video explores evaporation and the definition of temperature. It can be used as a supplement to an investigation, or may stand on its own to introduce a lesson or extend student learning.

highschoolenergy.acs.org/how-do-we-use-energy/temperature-energy.html

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Video Transcript

Chris
Ah, the humble thermometer. Whether we want to know if our holiday turkey is cooked or if we need concrete proof that we're too sick to attend school, a thermometer is the exact scientific tool we need. But what, exactly, does a thermometer measure?

In other words, what is temperature, really?

Brittny
Whether it's solid, liquid, or gas, all matter is made up of atoms and molecules. These particles are constantly in motion. They collide with each other and with the walls of anything they are contained in. We quantify the motions of the particles by their kinetic energies.

In chemistry, we define the temperature of a substance as the average kinetic energy of all the atoms or molecules of that substance.

Not all of the particles of a substance have the same kinetic energy. At any given time, the kinetic energy of the particles can be represented by a distribution. Some atoms or molecules have a lot of kinetic energy and move very fast. Other atoms or molecules have a little kinetic energy and move very slowly. It is the *average* kinetic energy of the particles that thermometers measure and we record as the temperature.

One process that illustrates varying kinetic energies particularly well is evaporation. As you probably know, evaporation is a phase change where particles of a substance move from the liquid phase into the gas phase.

But have you ever wondered how a puddle of water can evaporate at room temperature? Keep that in mind—we'll come back to that question in a bit.

Chris

When we think about a phase change from liquid to gas, we often think about adding thermal energy to a liquid by heating it up. When we do, the molecules of the liquid move faster and spread slightly farther apart, until they acquire enough energy to overcome the attractions they have for other molecules of the liquid and enter the gas phase.

Think about this: If someone asked you to turn a beaker full of water into water vapor, what would you do?

Being the brilliant young scientist that you are, you might put it on a hot plate and crank up the heat. Now you can kick back and relax until the hot plate transfers enough energy to get all the water molecules to transition from the liquid phase to the gas phase. In less science-y terms, you would boil the water.

Boiling is a special example of a liquid-to-gas phase change that occurs at a specific temperature called the boiling point, where the vapor pressure of the substance is equal to one atmosphere pressure. Boiling is usually carried out using a continuous input of energy from an external source (like a hot plate) to keep the temperature constant.

The obvious conclusion is this: If you continuously add thermal energy from a hot plate you can cause a phase change from liquid to gas.

But how can a puddle of water evaporate at room temperature?

When water evaporates at room temperature, 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 their temperature decreases. This liquid is at a lower energy than its surroundings, so it absorbs energy from those surroundings. The cycle continues as the puddle slowly disappears.

Boiling is a faster process because the surroundings (the hot plate) heat the liquid to a higher temperature where more molecules have high energy, so vaporization is faster. The hot plate is hotter than the liquid, so thermal energy transfer is fast enough to keep the liquid temperature constant at the boiling temperature.

Evaporation can take place at any temperature because some of the molecules in a liquid—the ones at the higher end of the distribution—will always have enough energy to enter the gas phase.

Chris
To sum up, temperature represents the average kinetic energy of the particles of substance. But it's the spread of kinetic energies among the individual particles that explains why puddles dry up.

The Explosive History of Nitrogen

By Tim Graham

April 16, 1947. It started with a small fire on the S.S. *Grand Camp*, a French cargo ship anchored off Texas City, TX. The ship had recently taken on 2380 tons of ammonium nitrate (NH_4NO_3) fertilizer.

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Early efforts by the crew to extinguish the fire were unsuccessful. The fire soon raged out of control. At 9:12 a.m., the ammonium nitrate exploded, sending the 7200-ton ship 20 feet in the air—the first of a series of catastrophic events. Burning debris reached surrounding oil refineries and chemical plants. A 15-foot tidal wave caused two other ships anchored in the harbor to collide. Both were soon ablaze. One of impacted vessels also contained ammonium nitrate. By the time the last flame had been extinguished, 576 people were dead and Texas City was in ruins.

PHOTODISC

Before 1850. By the mid-1800s, much research on potential explosives had already been done, most of which focused on nitrating solid substances that contained carbon…wood, coal, etc.

1846. Nitroglycerin is developed by the Italian chemist Asconio Sobrero. Believing that organic liquid substances might show more promise than solids, he nitrated glycerol. The resulting oily yellow liquid, called nitroglycerin, was a powerful explosive that was quite sensitive to shock. But he failed to develop a controlled and effective way to ignite it.

1863. The blasting cap is invented by Alfred Nobel as a means for exploding nitroglycerin. A blasting cap is a smaller primary explosive that is used to initiate the larger explosion. Nobel placed the nitroglycerin in an insulated container, added a gunpowder primer with safety fuse, and sealed the entire container. Moments after the fuse was lit, the gunpowder cartridge would explode and this, in turn, created enough energy to detonate the nitroglycerin.

1863. Trinitrotoluene, or TNT, is discovered to be a powerful explosive by the German chemist J. Wilbrand. It proves to be too expensive to manufacture in large quantities.

April 19, 1995. It was one of the most devastating acts of domestic terrorism ever to hit our nation. A truck loaded with roughly two tons of a mixture of ammonium nitrate fertilizer and fuel oil was detonated with a blasting cap on a street just outside the Alfred P. Murrah Federal Building in Oklahoma City, OK. A total of 168 people, many of them children, lost their lives. In 1997, two U.S. citizens, Timothy McVeigh and Terry Nichols, were convicted. McVeigh was sentenced to death and executed on June 11, 2001. Nichols is serving a life sentence.


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Ever since Alfred Nobel, the founder of the Nobel Peace Prize, developed a process to make dynamite in 1867, explosives have played a key role in both peace and war.

Today, when we think of explosives, substances like TNT (trinitrotoluene) and nitroglycerin come to mind. But ammonium nitrate? That's just a simple fertilizer! What is it about this simple inorganic compound that can cause it to react so violently? As you probably guessed, the answer is in the chemistry.

All explosions share some features. They all involve the rapid and violent release of large amounts of energy from a confined region of space. Particularly true for chemical explosions, they often involve the rapid expansion of gases generated during the explosion itself. Chemical explosions like those in Texas City and Oklahoma City are accompanied by a loud sharp report, flying debris, heat, light, and fire.

An explosive is a chemical compound or mixture that does the job. The explosive decomposition of nitroglycerin illustrates several features common to explosions:

$$4\,C_3H_5N_3O_9\,(l) \rightarrow 6\,N_2\,(g) + 12\,CO_2\,(g) + 10\,H_2O\,(g) + O_2\,(g) + \text{energy}$$

nitroglycerin

First, the reaction is exothermic, meaning that it releases energy. Second, it produces several gaseous products, all of which expand as the released energy raises the temperature. Third, even though the equation doesn't show it, the reaction is very rapid—once underway, all the energy is released in a very short time. Finally, the reactants include the element nitrogen.

Why do so many explosives contain the element nitrogen? Look at the products of explosive reactions and you'll find the same gas showing up over and over—ordinary nitrogen gas, N_2. The irony is that nitrogen gas is a very stable compound at a very low energy state. But when it is formed from reactants that start out in a very *high* energy state, a very large amount of energy is released in the process. Kaboom!

Why do explosive compounds react so rapidly? One way to speed up a reaction is to thoroughly mix the reactants. Mixing allows for immediate contact to occur. You may have read about explosions in flour mills and grain elevators. Even otherwise harmless substances like flour can explode violently if thoroughly mixed with air and ignited by a spark.

1867. Dynamite


PHOTODISC

is invented. Nitroglycerin is a highly unstable liquid likely to explode with the slightest shock. To reduce its obvious hazards, Nobel uses a finely powdered silicon-based absorbent called Kieselguhr to soak up the liquid nitroglycerin—thus, stabilizing the explosive without sacrificing its strength. Later, he replaces Kieselguhr with sawdust and sodium nitrate. He substitutes ammonium nitrate for some of the nitroglycerin to make a new, low-cost explosive, dynamite.

1900. TNT production costs drop.

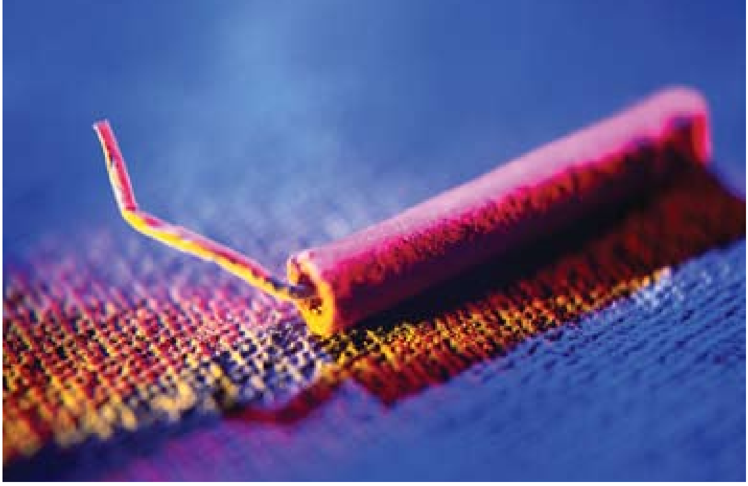
TNT is appreciated as a very stable solid that can be poured and even melted with relative safety.

1914. TNT is used as a weapon in World War I.

TNT's big advantage over dynamite is its capacity for producing shock waves that can rupture the steel on armor-plated vehicles.

1940. World War II

weaponry introduces two new explosives, **RDX** (hexahydro-1,3,5-trinitro-1,3,5-triazine) and **PETN** (pentaerythritoltetranitrate). With additions of wax, motor oil, and other stabilizing fillers, RDX is renamed *Composition Four*, or C-4 explosive. Stable within a large temperature range (–70 ° to 170°F), safe to handle, and easy to mold due to it's plastic-like properties, C-4 is attached to bridge supports, armored vehicles, or the hulls of ships. It is detonated with blasting caps.



PHOTODISC

Ammonium nitrate replaced unstable nitroglycerin to make the low-cost explosive—dynamite.

Molecules of explosive compounds like nitroglycerin or trinitrotoluene take the mixing step one step further. For these compounds all of the reactants are on board the same molecule. Immediate contact is assured.

Let's go back to the Texas City tragedy. What caused the ammonium nitrate in the holds of the ship to explode without the use of some other explosive? Chemists found that the answer was in the bag. The ammonium nitrate fertilizer was packaged in plain paper. The cellulose used to make paper contains a large amount of the element carbon. It was the carbon and ammonium nitrate mixture that reacted to unleash the tragic explosion.

COURTESY OF THE MOORE MEMORIAL PUBLIC LIBRARY, TEXAS CITY, TEXAS

By analyzing the circumstances surrounding the Texas tragedy, chemists began to appreciate the power and potential of ammonium nitrate-based explosives. An effective, relatively safe, and inexpensive explosive called ANFO (ammonium nitrate fuel oil) was developed. And there were no risky transport problems to be solved. To make ANFO, ammonium nitrate and fuel oil were mixed at the blast site.

MIKE CIESIELSKI

Ammonium nitrate, the key reactant in several explosives, is a common ingredient in many plant fertilizers.

But ANFO was limited as a commercial explosive. Ammonium nitrate is water-soluble. As it gains water, the energy necessary to initiate its reaction with fuel oil increases to levels making it useless as an explosive. Dupont chemists went to work to produce a form of ammonium nitrate that would detonate even in a wet environment. By adding sensitizers, they were able to detonate the mixture with less-energetic shock waves. Then by adding thickening agents, they produced a syruplike mixture called TOVEX—easy to pour into drill holes at the blasting site.

It's interesting that dynamite's inventor Alfred Nobel became a pacifist later in life. The man whose name will forever be known as the father of modern explosives dedicated much of his influence and fortune to opposing their use as weapons of destruction.

Tim Graham teaches chemistry at Roosevelt High School in Wyandotte, MI. His most recent article "Luminol—Casting a Revealing Light on Crime" appeared in the December 2001 issue of *ChemMatters.*

REFERENCES

Conkling, John A. Pyrotechnics. *Scientific American*, July 1990, pp 96–102.

Fisher, D. E. and Fisher, M. J. The Nitrogen Bomb. *Discover*, April 2001, pp 52–57.

1945. Ammonium nitrate is manufactured, stockpiled, and shipped to war-torn Europe as an inexpensive fertilizer for enriching depleted farm soil. It's potential explosive power is known, but its stability relative to other explosives seems adequate for production and storage.

1947. S.S. Grand Camp explodes in Texas City Harbor. Chemists reconsider the stability of ammonium nitrate.

1957. ANFO (Ammonium Nitrate Fuel Oil) explosive is developed by taking ammonium nitrate "prills" and mixing them with liquid fuel oil to make a "slush". Mixed at the blast site, ANFO is relatively safe to handle.

1988. Pan AM Flight 103 wreckage is found to contain RDX residue. Presumably, as a result of terrorism, the plane crashed in Lockerbie, Scotland, killing all 270 people on board.

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.

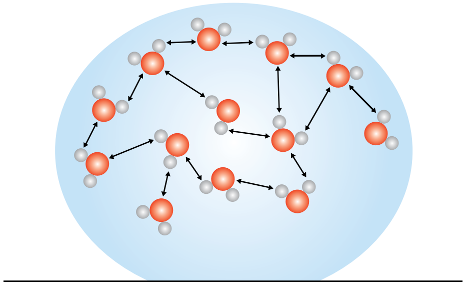
During evaporation, energetic molecules leave the liquid phase, which lowers the average energy of the remaining liquid molecules. The remaining liquid molecules can then absorb energy from their surroundings. This process can take place at any temperature because some of the molecules in a liquid will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

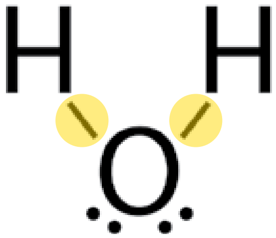
But are all attractions the same?

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



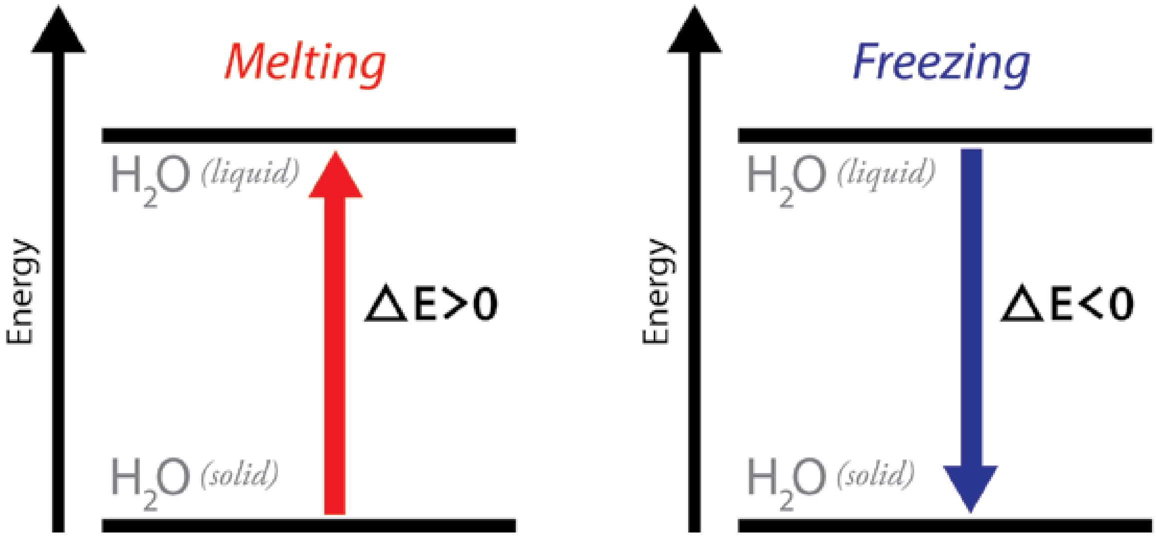
Intramolecular forces
Covalent bonds between hydrogen and oxygen

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

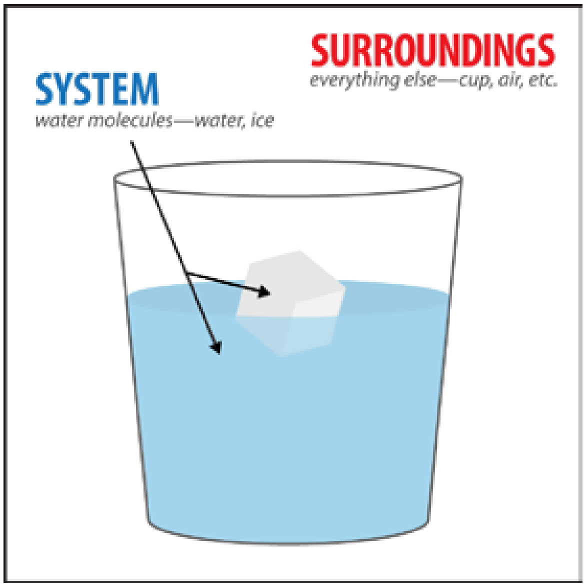
You can visualize the energy implications of phase changes using energy diagrams like the ones shown below. The horizontal lines represent the energies of substances in particular states. The higher horizontal line represents a substance in a physical state at a higher energy level, and the lower horizontal 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

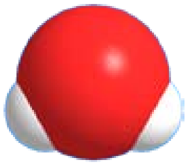
Liquid	**T_i**	**T_f**	**ΔT**	**ΔE (+/–)**
1. Water				
2. Isopropyl alcohol				
3. Acetone				
4. Control				

ANALYZING EVIDENCE

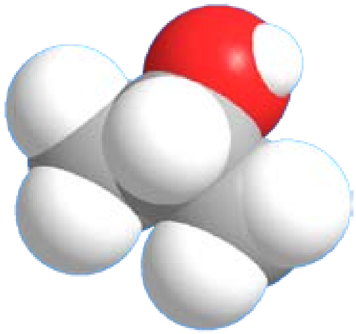
Complete Table 2 using your observations from the experiment. Enter either a plus (+) or minus (–) sign in the last column of the table for the change in energy of the liquid-dampened paper towel in each case.

INTERPRETING EVIDENCE

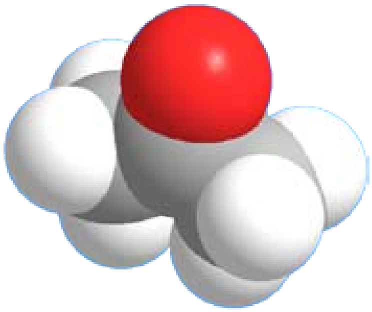
1. Using the space-filling models below as a guide, write the chemical formulas for (1) water, (2) isopropyl alcohol, and (3) acetone.



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

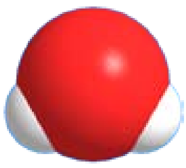
Liquid	Water	Isopropyl alcohol	Acetone
Time (minutes)	10:00	3:00	0:52

Table 2. Energy change during evaporation results

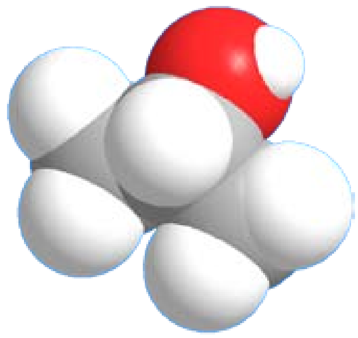
Liquid	T_i	T_f	ΔT	ΔE (+/−)
1. Water	24 °C	18 °C	−6 °C	−
2. Isopropyl alcohol	24 °C	15 °C	−11 °C	−
3. Acetone	24 °C	6 °C	−18 °C	−
4. Control	24 °C	24 °C	0 °C	N/A

Interpreting Evidence

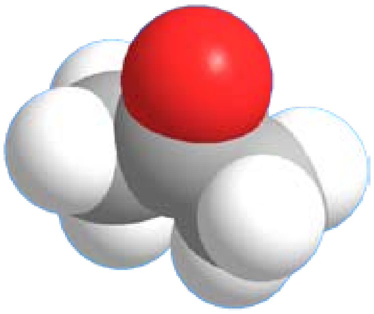
1. Using the space-filling models below as a guide, write the chemical formulas for (1) water, (2) isopropyl alcohol, and (3) acetone.



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.

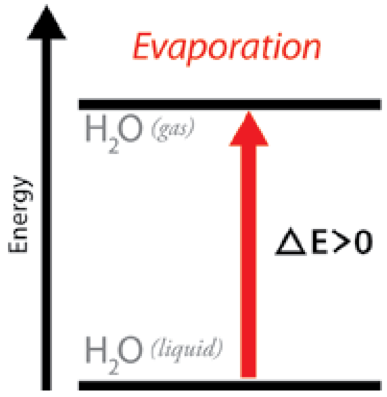
   The temperature decreased for each of the liquid-soaked paper towels in the experiment because evaporation is a phase change that absorbs energy.

   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.

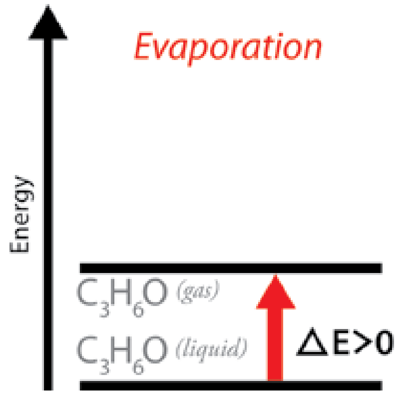
   We observe energy being absorbed by a liquid as it evaporates as a decrease in temperature of the liquid and in its surroundings as they transfer energy to the liquid.

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?

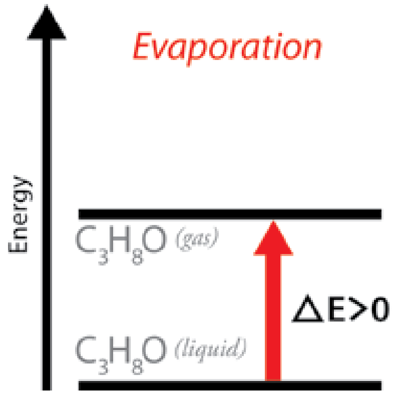
   These diagrams show the relative energy levels for the same number of moles of each liquid and its vapor at the same temperature. All of the diagrams show that the process of evaporation absorbs energy. Water absorbs the most energy, acetone absorbs the least energy, and isopropyl alcohol is somewhere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

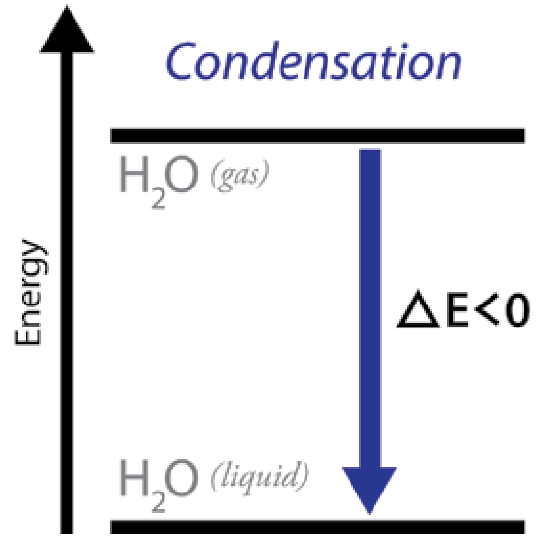
   Evaporation is a cooling process. In the liquid state, molecules are closer together than they are in the gas state. Energy is required to separate these molecules of a liquid as they become farther apart in the gas phase. The amount of energy absorbed by the liquid must be enough to overcome the attractions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

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

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:

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

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

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

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

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

Analyzing Part B

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

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

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

Analyzing Part C

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

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

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

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

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

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

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

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

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

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

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

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.

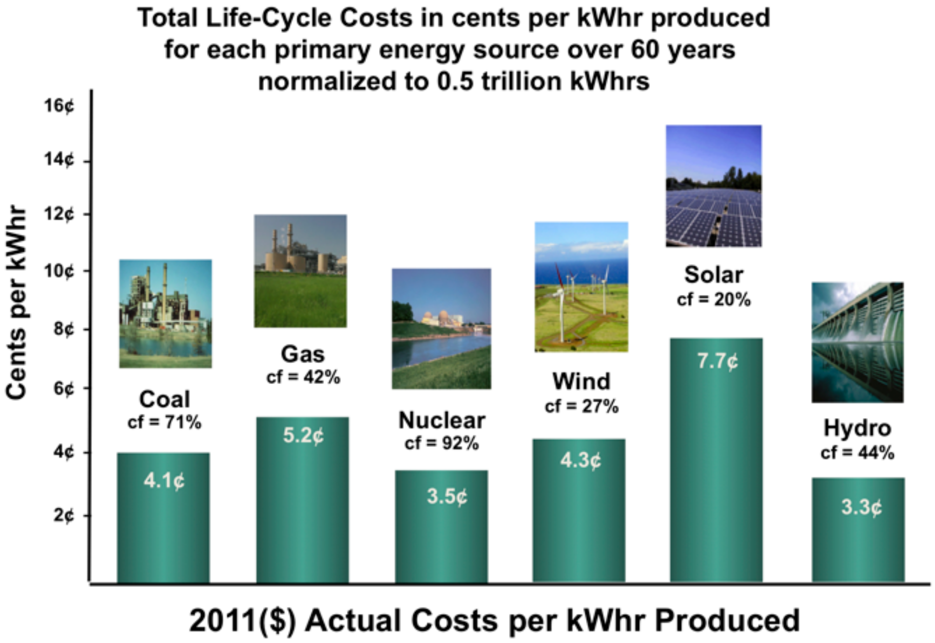
Heat source	Exp ΔT	Cost (¢)	Cost/10 °C
Bunsen burner	50.5 °C	0.203¢	0.040¢
Hot plate	50.5 °C	0.77¢	0.153¢
Microwave	50.5 °C	0.181¢	0.035¢

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.

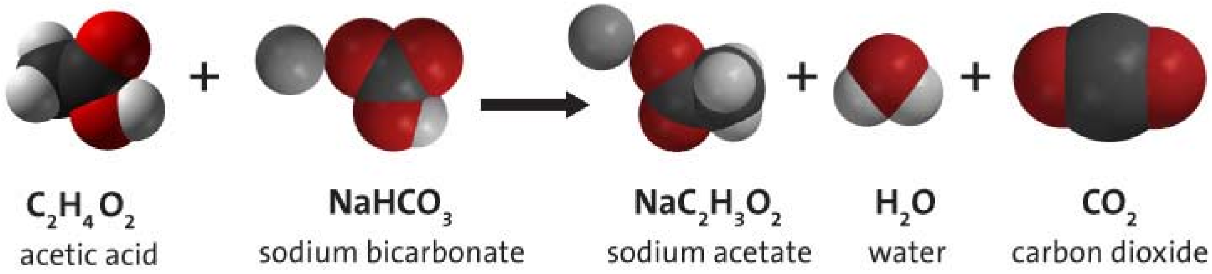
PREPARING TO INVESTIGATE

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:

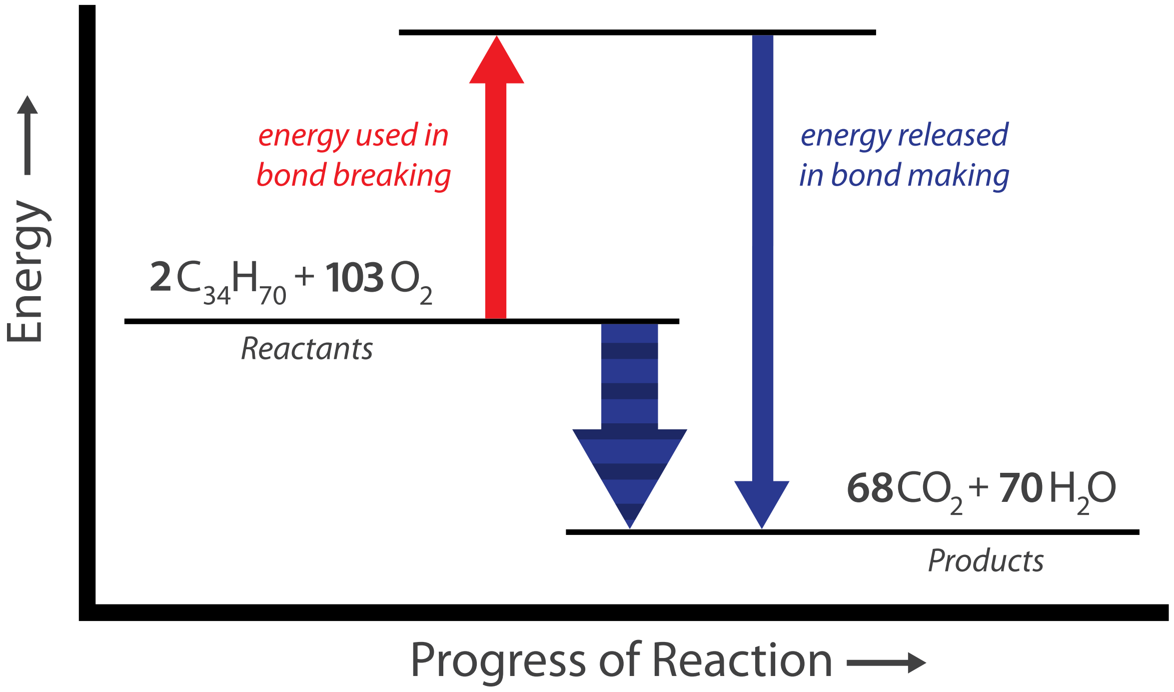
   ΔH = energy used in reactant bond breaking + energy released in product bond making

   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.



Exothermic ($\Delta H < 0$)

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

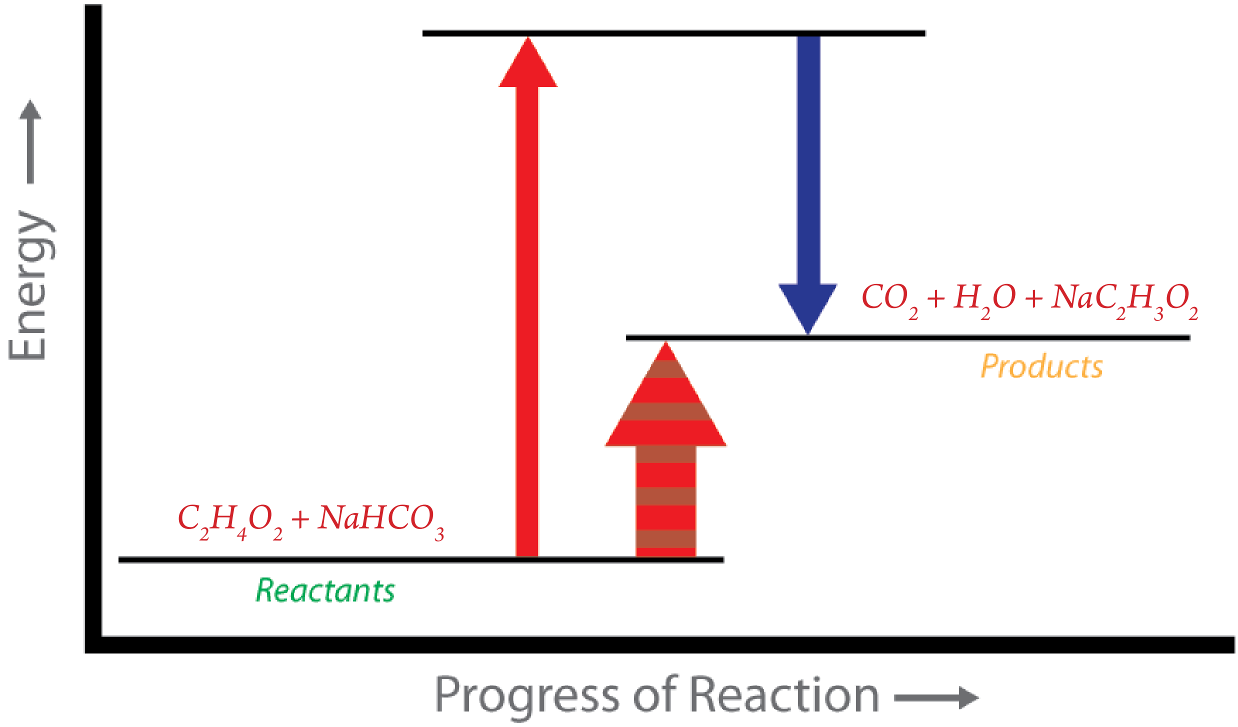
      $2NaHCO_3 + CaCl_2 \longrightarrow 2NaCl + CaCO_3 + H_2O + CO_2$

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

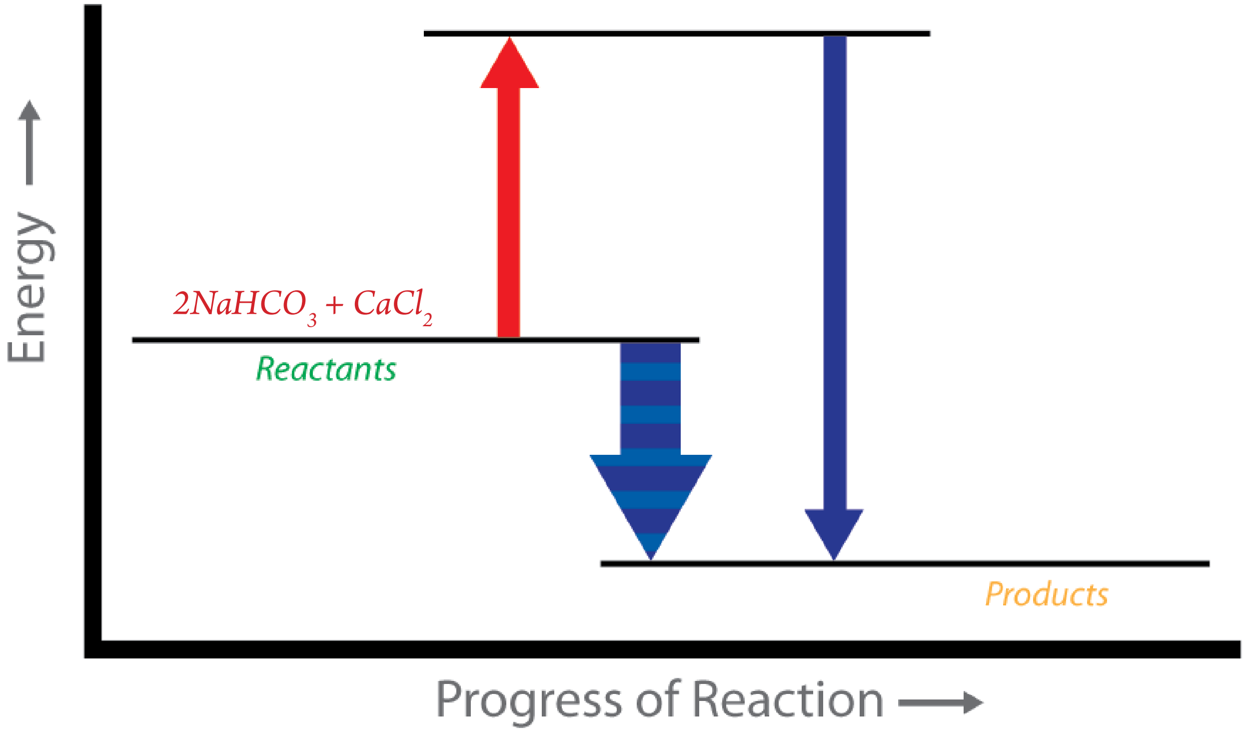
   In the chemical reaction between baking soda and calcium chloride, more energy is released when the products ($2NaCl + CaCO_3 + H_2O + CO_2$) are formed than energy was used to break bonds in the reactants ($2NaHCO_3 + CaCl_2$). There is therefore a net release of energy to the surroundings and the reaction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

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

Baking soda + vinegar
Endothermic
ΔH>0



Baking soda + calcium chloride
Exothermic
ΔH<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

$$\mathrm{Li\,(s)} + \tfrac{1}{2}\,\mathrm{F_2\,(g)} \longrightarrow \mathrm{LiF(s)}$$

$$\mathrm{Li^+\,(g)} + \mathrm{F^-\,(g)} \longrightarrow \mathrm{LiF(s)}$$

$$\tfrac{1}{2}\,\mathrm{F_2\,(g)} \longrightarrow \mathrm{F^-\,(g)}$$

$$\mathrm{Li\,(g)} \longrightarrow \mathrm{Li^+\,(g)} + \mathrm{e^-}$$

$$\mathrm{Li\,(s)} \longrightarrow \mathrm{Li(g)}$$

$$\mathrm{F\,(g)} + \mathrm{e^-} \longrightarrow \mathrm{F^-\,(g)}$$

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

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

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

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

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

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

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

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

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

1st $IE_{Na} = 496$ kJ

½ $BE_{Cl_2} = 75.3$ kJ

$EA_{Cl} = -349$ 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

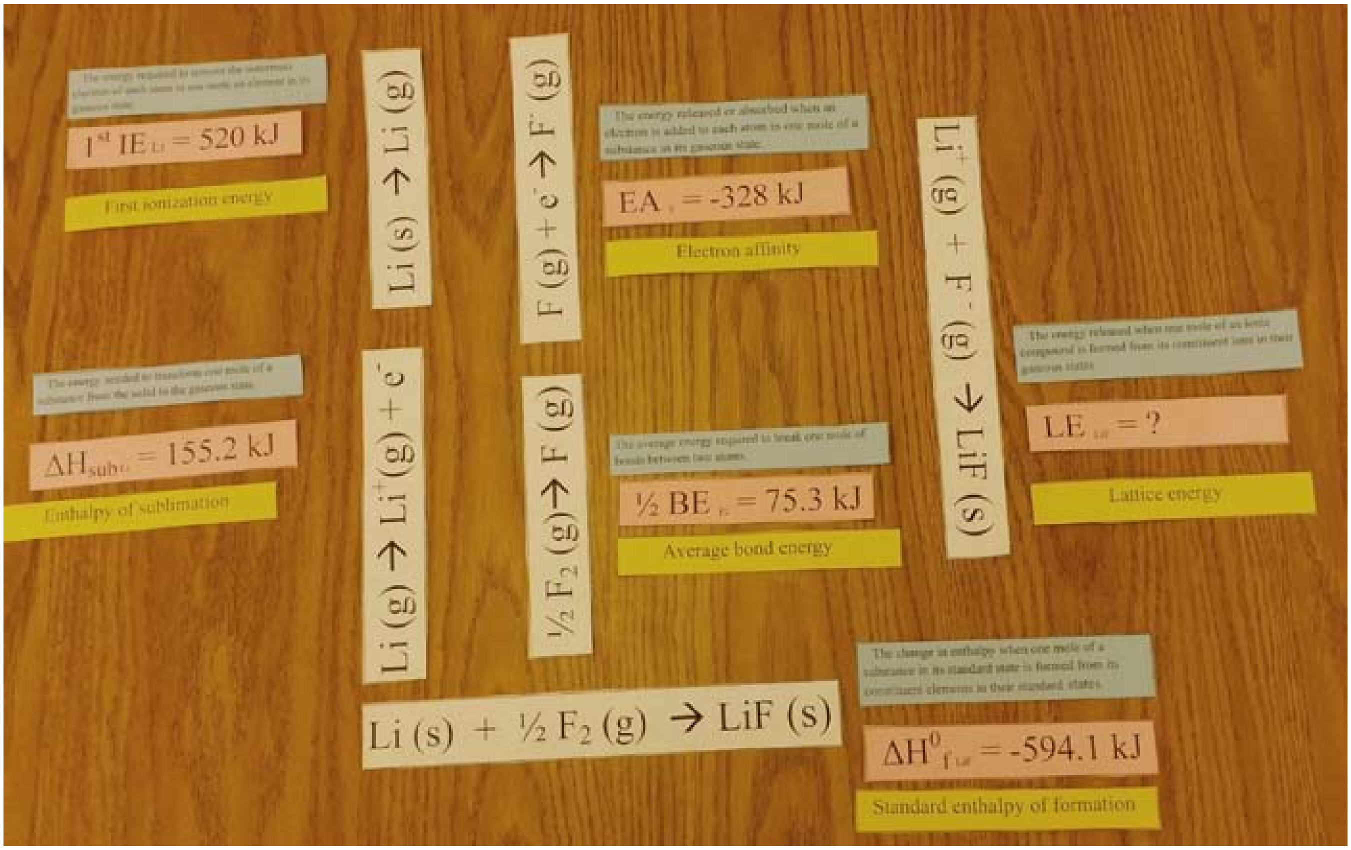
$EA_{Cl} = -349$ kJ

$LE_{CaCl_2} = ?$

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) → Li (g)	ΔH_{sub} = 155.2 kJ
Li (g) → Li^+(g) + e^-	1st Ionization Energy of Li = 520 kJ
½ F_2 (g) → F (g)	½ Bond Energy of F_2 = 75.3 kJ
F (g) + e^- → F^- (g)	Electron Affinity for fluorine = −328 kJ
Li^+ (g) + F^- (g) → LiF (s)	Lattice Energy of LiF = ??
Li (s) + ½ F_2 (g) → 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) → Na (g)	ΔH_{sub} = 108 kJ
Na (g) → Na^+(g) + e^-	1st Ionization Energy of Na = 496 kJ
½ Cl_2 (g) → Cl (g)	½ Bond Energy of Cl_2 = 122 kJ
Cl (g) + e^- → Cl^- (g)	Electron Affinity for chlorine = −349 kJ
Na^+ (g) + Cl^- (g) → NaCl (s)	Lattice Energy of NaCl = ??
Na (s) + ½ Cl_2 (g) → 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) → Ca (g)	ΔH_{sub} = 178.2 kJ
Ca (g) → Ca^+(g) + e^-	1st Ionization Energy of Ca = 590 kJ
Ca^+ (g) → Ca^{2+}(g) + e^-	2nd Ionization Energy of Ca = 1145 kJ
Cl_2 (g) → 2Cl (g)	Bond Energy of Cl_2 = 244 kJ
2Cl (g) + $2e^-$ → $2Cl^-$	Electron Affinity for chlorine = 2(−349 kJ)
Ca^{2+} (g) + $2Cl^-$ (g) → $CaCl_2$ (s)	Lattice Energy of $CaCl_2$ = ??
Ca (s) + Cl_2 (g) → $CaCl_2$ (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

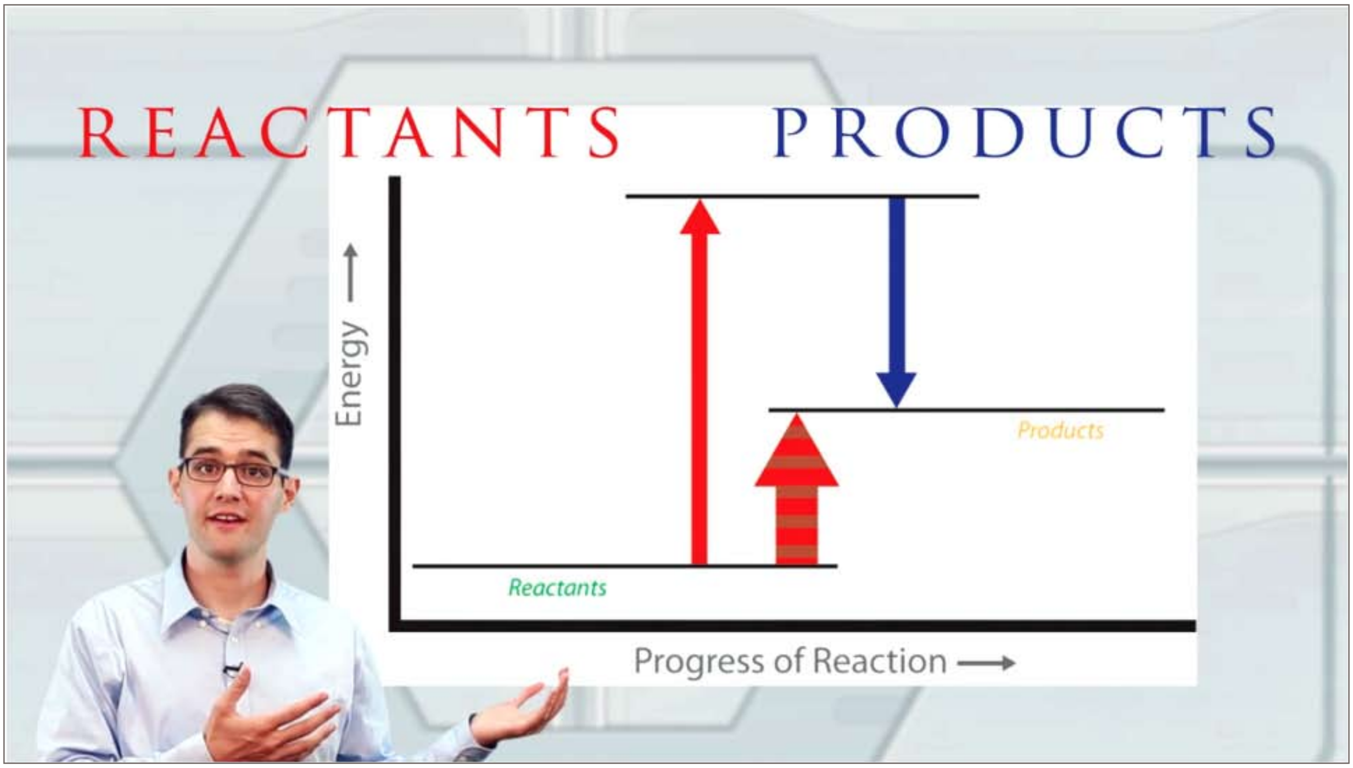
Further Analysis

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.

What Theories Explain Energy?

Thermodynamics is the study of energy change accompanying physical and chemical changes. In chemical systems, these changes are described through the ideas of heat and work, enthalpy, entropy, and Gibbs free energy. These explanations are best understood at the microscopic scale, where the energy can be characterized through the motions of particles and the relative position of particles. While these concepts are abstract, the investigations, videos, and readings in this section give students practical experience with the concepts and rules that govern how matter and energy interact.

- Lab Investigation—Entropy & Enthalpy Changes
- Lab Investigation—Energy & Entropy of a Stretched Rubber Band
- *ChemMatters* Reading—Why Cold Doesn't Exist

Entropy & Enthalpy Changes | A Lab Investigation

Summary

In this investigation, students will explore basic thermodynamic concepts, including spontaneity, entropy, and enthalpy through a series of guided questions and procedures.

Objective

Given prior knowledge of the thermodynamic terms entropy, enthalpy, and spontaneous processes, students will gain a deeper understanding of how $\Delta G = \Delta H_{sys} - T\,\Delta S_{sys}$ expresses the second law of thermodynamics by exploring energy transfer between system and surroundings as salts dissolve.

Safety

- Be sure you and the students wear properly fitting goggles.
- Ammonium chloride can be an irritant to body tissues. In the event of contact, wash affected areas with water.
- Ammonium nitrate is a strong oxidizer. May emit toxic vapors of NO_x and NH_3 when heated to decomposition. Can be an irritant to body tissues. In the event of contact, wash affected areas with water.
- Calcium chloride can be an irritant to body tissues. In the event of contact, wash affected areas with water.
- Acetone is flammable. Avoid flames or sparks. Irritating to body tissues. Avoid body tissue contact. Slightly toxic by ingestion. Skin contact causes dermatitis. Vapor may cause weakness, fatigue, nausea, and headache. Work in a well-ventilated area.
- Dispose of solutions according to local regulations.

Materials for Each Group

- 10 g ammonium chloride or ammonium nitrate
- 5 g calcium chloride
- Thermometer
- 100-mL graduated cylinder
- 3 150-mL beakers, stirring rod

Optional Materials for the Post-Lab Demo

- Digital thermometer
- 10-mL graduated cylinder
- 10 mL acetone

Time Required

One class period, approximately 45–50 minutes.

Lab Tips

This lab is designed for students to work together, discussing and answering the questions posed while proceeding through the step-by-step treatment of the second law.

Pre-Lab Discussion

What does spontaneous mean? What kinds of processes in your experience happen spontaneously? Are there any differences between them?

Incorporating into the Curriculum

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

PREPARING TO INVESTIGATE

Thermodynamics is a way of describing energy transformations when a system changes from one state to another. The entire architecture of thermodynamics is built on carefully defined terms, many of which have an everyday meaning that is not exactly what chemists mean when they use the term. For example, one way that chemists state the second law of thermodynamics is that in any *spontaneous* change, the *entropy* of the *universe* increases. The underlined words have a very particular meaning that we need to know before we can understand the second law.

The second law of thermodynamics may be expressed in many ways, and it has been used by chemists to understand everything from the work of a steam engine to the direction of time. It grew in the nineteenth century out of observations made about big things like steam engines, and today it is often used to illuminate the conceptual, chemical world of tiny things like atoms, ions, and molecules. In this activity we will use careful observations of the process of dissolving salts in water to more deeply understand the second law.

What is a Spontaneous Change?

A spontaneous change is any change that happens freely in time. For example, you can drop a ball from above your head and it falls to the floor (spontaneous) but you need to provide energy to the ball to place it over your head again. Being able to predict what processes will be spontaneous is how we apply the second law.

Which of the following processes are spontaneous?

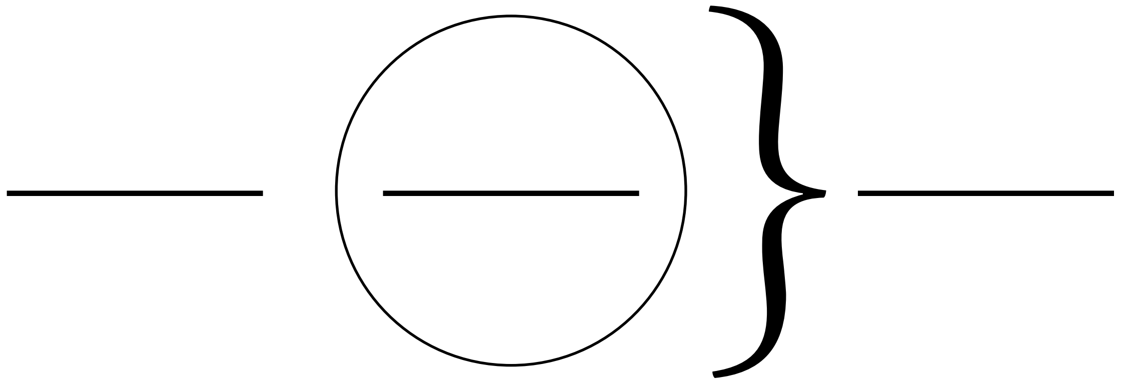
- Ice melts when dropped in a cup of warm water.
- Water evaporates when it is spilled on a hot surface.
- Water in a glass on your desk decomposes to hydrogen and oxygen.
- Iron rusts in air.
- The smell of perfume spray spreads throughout a room.
- Equal volumes of olive oil and vinegar dissolve together to make a salad dressing.

Good to know: spontaneous ≠ instantaneous! Even if a change is spontaneous, this doesn't necessarily mean that it happens quickly. The second law tells us that all diamonds are spontaneously turning into coal, but this process is so slow we will never observe it taking place.

System, Surroundings, Universe

The system is the specific part of the universe we are considering, where a change is taking place. It can be any size—a test tube, a beaker, a human body, or an ocean. The surroundings are everything outside the system. The universe consists of the system and the surroundings together.

Fill in the blanks below. Label the regions with the terms *system*, *surroundings,* and *universe*.



GATHERING EVIDENCE

1. Place 5.0 g of ammonium nitrate in a 150-mL beaker, then put the beaker on the portion of the diagram above that you have labeled as the system. Write the formula for ammonium nitrate and identify its solid type.

   Formula:
   Identify by underlining: Ionic Solid Covalent Solid Metallic Solid

2. Record your observations of the macroscopic properties of the salt:

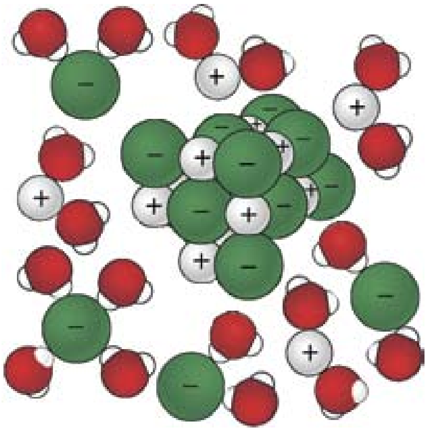
3. Complete the sentence: Ammonium nitrate is an ______________________ solid, and the individual particles in its lattice structure are _________________.

Defining Entropy and Looking at Entropy Changes in a System

Entropy is a mathematically defined property in thermodynamics. It can often help to understand it as a measure of the possible arrangements of the atoms, ions, or molecules in a substance.

The symbol for entropy is S, and a change in entropy is shown as “delta” S or ΔS. If the entropy of a system increases, ΔS is positive. If the entropy of a system decreases, ΔS is negative.

1. Pour 100 mL of water over the salt in the beaker and stir. Can you still see the ammonium nitrate?
2. Make and record your observations:



3. Using the terms *cation, anion, solute, solvent,* and *solution,* label the diagram on the right.

4. Given the physical state of ammonium nitrate before it dissolves, how do the possible arrangements of the ions in the salt compare to their possible arrangements when free to move within the solution?

5. Does your answer to the preceding item suggest that the entropy of the ammonium nitrate increased or decreased upon dissolving?

6. Would ΔS for this change be positive or negative?

But more is going on than just ions leaving the solid and moving about more freely. Note in the figure above that the polar water molecules are attracted to and oriented around the dissolved ions. The ions are solvated. This orientation of a lot of the water molecules reduces their freedom to move about in the liquid, so the number of possible arrangements of the water molecules is reduced when the ions are present.

7. Would ΔS for this change in the arrangement of water molecules be positive or negative?

The change in entropy for the reaction system ΔS_{sys} has to include both the positive change for the ions and the negative change for the water molecules. Which one predominates? For most salts with single charges on their cations and anions, like NaCl, KBr, or $LiNO_3$, the positive change in entropy for the ionic solid separating into its ions in solution will predominate.

8. Did dissolution of ammonium nitrate happen spontaneously?
9. If the overall ΔS_{sys} for the dissolution of ammonium nitrate is positive, which ΔS in the system predominates: ΔS for the ions, or ΔS for the water molecules?

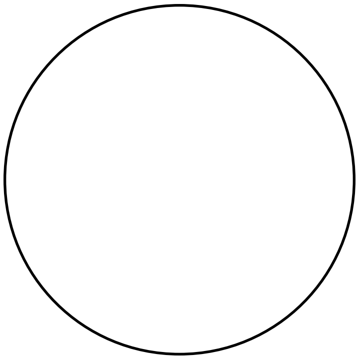
Looking at Entropy Changes in the Surroundings by Defining Enthalpy

We will now consider entropy changes in the surroundings by looking at another thermodynamic term, *enthalpy*. The enthalpy of a system has a definition in thermodynamics that relates to its internal energy, the pressure on the system, and the volume of the system. It is useful in understanding the second law, however, because at constant pressure and volume, a change in enthalpy is the same as the thermal energy transferred from the system to the surroundings, or from the surroundings to the system.

The symbol for the enthalpy of a system is H, and a change in enthalpy is shown as "delta" H or ΔH. If thermal energy transfers from the system to the surroundings during a physical or chemical change, the ΔH is negative and the change is *exothermic*. If thermal energy transfers from the surroundings to the system during a change, the ΔH is positive and the change is *endothermic*.

Case one for enthalpy

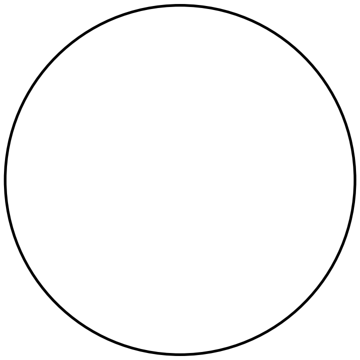
1. Measure out 100 mL of water in a clean 150-mL beaker. Once again, label the "system" and the "surroundings" in the diagram below.
2. Place the beaker on your paper in the region labeled "system" below, then measure the temperature of its contents with a thermometer and record.
3. Add 5.0 g of ammonium chloride or ammonium nitrate.
4. While holding the beaker at its base, stir, and make and record your observations, including the final temperature of the mixture.
5. Your hand at the base of the beaker can be considered part of the surroundings. Was thermal energy transferred to your hand from the beaker, or away from your hand to the beaker?
6. Label the diagram below with an arrow, showing the direction of the transfer of thermal energy. Is thermal energy transferred from the surroundings to the system, or from the system to the surroundings?



7. Is the change endothermic or exothermic?
8. Is ΔH of the system positive or negative?

Case two for enthalpy

1. Measure 100 mL of water in a clean 150-mL beaker.
2. Place the beaker in the “system” below, and measure and record the temperature of its contents. Add 5.0 g of calcium chloride.
3. While holding the beaker at its base, stir, and make and record your observations, including the final temperature of the mixture.
4. Was thermal energy transferred to your hand from the beaker, or away from your hand to the beaker?
5. Label the diagram below with an arrow, showing the direction of the transfer of thermal energy. Is thermal energy transferred from the surroundings to the system, or from the system to the surroundings?



6. Is the change endothermic or exothermic?
7. Is ΔH of the system positive or negative?

Relating the change in enthalpy (ΔH of the system) to a change in entropy of the surroundings (ΔS of the surroundings)

We are going to look at how entropy changes in the surroundings, depending on the sign of ΔH of the system. Before we do, however, let’s review.

1. When thermal energy is transferred from region A to region B, the molecules in region B, on average, ________________________ (speed up or slow down?).
2. When the average speed of the molecules in a region is high, there are ________________ (more or fewer?) high-speed molecules. When the average speed of the molecules in a region is low, there are ____________________ (more or fewer?) high-speed molecules.

3. When the average speed of the molecules in a region is high, there are ________________ (more or fewer?) possible arrangements for the molecules among the molecular speeds. When the average speed of the molecules in a region is low, there are _________________ (more or fewer?) possible arrangements for molecules among the molecular speeds.

Now consider what happens in the *surroundings* during the process of thermal energy transfer.

4. When thermal energy is transferred in an endothermic change, will the molecules of the surroundings speed up or slow down?
5. How will this affect the entropy of the surroundings, will it increase or decrease?
6. In this case, is ΔS of the surroundings positive or negative?
7. When thermal energy is transferred in an exothermic change, will the molecules of the surroundings speed up or slow down?
8. How will this affect the entropy of the surroundings, will it increase or decrease?
9. In this case, is ΔS of the surroundings positive or negative?

Looking at entropy changes in the universe and the second law

In any spontaneous change, the entropy of the universe increases, or ΔS universe > 0.

We have seen that we can consider the entropy of the system and the entropy of the surroundings separately. Since the universe is composed of the system and the surroundings, then we will consider both when determining the change in entropy of the universe:

$$\Delta S \text{ universe} = \Delta S \text{ surroundings} + \Delta S \text{ system}$$

- When both the change in entropy of the system and the change in entropy of the surroundings are positive, the change in entropy of the universe must be positive, so the reaction will always be spontaneous, under the given conditions.
- If both terms are negative, the reaction is never spontaneous under the given conditions.
- If one term is positive and the other is negative, then whichever has the largest absolute value determines whether ΔS universe is positive (spontaneous reaction) or negative (reaction is not spontaneous) under the given conditions.

Complete the following:

1. In an exothermic reaction, ΔS surroundings is ________________ (positive or negative?).
2. In an endothermic reaction, ΔS surroundings is _______________ (positive or negative?).
3. In a spontaneous, endothermic reaction, thermal energy is spontaneously ______________________(absorbed or released?) by the system from the surroundings.
4. In a spontaneous, exothermic reaction, thermal energy is spontaneously ______________________ (absorbed or released?) to the surroundings from the system.

ANALYZING EVIDENCE

Application of Free Energy

Thermodynamics also defines a term known as free energy, G; the change, ΔG, when a system undergoes a change, is often considered the energy of the system available to do work. Free energy also relates to the second law, since *in any spontaneous process the change in free energy (ΔG) for the system is negative*. The definition for the change in free energy summarizes our discussion of entropy changes in the system and in the surroundings:

$$\Delta G = \Delta H_{sys} - T\,\Delta S_{sys}$$

We have seen that whenever ΔH_{sys} is negative (exothermic) the ΔS_{sur} increases, or is positive.

1. So if ΔH_{sys} is negative and ΔS_{sys} is positive, what will the sign of ΔS universe always be?
2. Under these conditions, is the reaction spontaneous?
3. And under these conditions, what will the sign of ΔG always be?

This famous relationship of free energy change to changes in enthalpy and entropy shows us the balance between entropy changes in the system and the surroundings, and how that balance depends on temperature. Note that T stands for the absolute temperature in Kelvin, so its value is always positive.

Use the relationship $\Delta G = \Delta H_{sys} - T\,\Delta S_{sys}$ to complete the chart below:

If ΔH_{system} is	Then $\Delta S_{surroundings}$ will be (+ or –)	And if ΔS_{system} is	And $\Delta S_{universe}$ will be (+ or –)	The ΔG will be (+ or –)	And the reaction is (spontaneous or not spontaneous)
negative or exothermic		positive (any temperature conditions)			
positive, or endothermic		negative (any temperature conditions)			
negative or exothermic		negative (high temperature conditions)			

negative or exothermic		negative (low temperature conditions)			
positive, or endothermic		positive (high temperature conditions)			
positive, or endothermic		positive (low temperature conditions)			

INTERPRETING EVIDENCE

Decomposition of Baking Soda

The mysteries of bread making began to be simplified in the 1800s with the use of baking soda. Although it was also combined with sour milk to lighten the texture of heavy bread, it could produce a lightening effect in bread dough on its own when it decomposed:

$$2\ NaHCO_3\ (s) \longrightarrow Na_2CO_3\ (s) + CO_2\ (g) + H_2O\ (g)$$

1. Gases are produced in the decomposition, therefore the sign of ΔS_{sys} for the process is likely to be ______________.
2. Heat is absorbed in the process so the sign of ΔH_{sys} is ______________.
3. Apply this information to $\Delta G = \Delta H_{sys} - T\ \Delta S_{sys}$. When is it likely that ΔG will be negative, in other words, when will the process be spontaneous, at high T or at low T?

   ____

4. If the values of ΔH_{sys} and ΔS_{sys} are found, we can predict the temperature needed for the baking soda to decompose, and adjust our baking temperature accordingly!

Melting Point of Water

Consider the process:

$$H_2O\ (s) \longrightarrow H_2O\ (l)$$

1. Determine the sign for ΔH_{sys}. Explain your choice.
2. Determine the sign for ΔS_{sys}. Explain your choice.
3. Will ΔG be negative at high or at low temperatures?
4. We can use $\Delta G = \Delta H_{sys} - T\ \Delta S_{sys}$ to predict the temperature at which a substance will melt, if we have values for ΔH_{sys} and ΔS_{sys}. At some temperature (T), ΔG will change from positive to negative. What will be the value of ΔG at the temperature of this change?

5. Substitute this value for ΔG and rearrange $\Delta G = \Delta H_{sys} - T\,\Delta S_{sys}$ to solve for T. This value of T will be what?

Production of Ozone

Ozone (O_3) is an unstable form of oxygen that is formed in the stratosphere. The ozone layer in the upper atmosphere protects life on the earth's surface from high-energy ultraviolet light from the sun. Ozone is produced from oxygen gas:

$$3\,O_2\,(g) \longrightarrow 2\,O_3\,(g)$$

1. In this process, __________________ moles of oxygen gas produce _________________ moles of gaseous ozone. Are there more possible arrangements (higher entropy) for the reactants or the products? __________________Therefore, is ΔS_{sys} positive or negative?
2. The process is endothermic, so ΔH_{sys} is _________________________.
3. Use $\Delta G = \Delta H_{sys} - T\,\Delta S_{sys}$ to consider: At what T will the reaction be spontaneous?
4. If a system will never spontaneously absorb thermal energy from the surroundings to undergo a change, does that mean the change is impossible? Could energy somehow be forced into the system, for example, by doing "work" on the system?
5. In the case of the production of ozone in stratosphere, which goes on every day, what kind of energy could be "working" on the oxygen gas?

Decomposition of Hydrogen Peroxide

If you have ever used hydrogen peroxide to disinfect an open cut, you may have seen bubbles form when the hydrogen peroxide decomposes:

$$2\,H_2O_2\,(l) \longrightarrow 2\,H_2O\,(l) + O_2\,(g)$$

1. What must be forming in the bubbles?
2. When a gas is formed from a liquid, is the sign of ΔS_{sys} positive or negative?
3. This decomposition is exothermic, so what is the sign of ΔH?
4. So the decomposition of hydrogen peroxide is spontaneous at ___________temperatures.
5. Consider that the bottle you bought from the store has been sitting on the shelf for some time, and still contains H_2O_2. Does the conclusion that a reaction is spontaneous at all temperatures mean that the reaction happens quickly? What could have caused the reaction to speed up when the hydrogen peroxide was applied to your open cut?

REFLECTING ON THE INVESTIGATION

Demo

1. Place the measuring end of a digital thermometer into a 10-mL graduated cylinder full of acetone.
2. Have students note the temperature, then remove the thermometer, waving it a bit in the air to rapidly evaporate the acetone. The temperature quickly falls.
3. Have students analyze the process from the point of view of thermodynamic properties (the endothermic ΔH_{vap}, the positive ΔS_{system}, and the spontaneity of the process).

Cross Link
Students can explore the basic mechanics of evaporative cooling in the investigation *The Energy of Evaporation* on page 65.

Practice Problem

Now that we've observed the qualitative relationships between ΔH, ΔS, and ΔG, let's calculate the quantitative value of $\Delta G°_{rxn}$ for the dissolution of solid ammonium nitrate in water at 25 °C:

$$NH_4NO_3\ (s) \longrightarrow NH_4^+\ (aq) + NO_3^-\ (aq)$$

	ΔH (kJ/mol)	So (J/mol K)
NH_4NO_3 (s)	−365.6	151
NH_4^+ (aq)	−132.80	112.8
NO_3^- (aq)	−206.57	146.4

Remember that $\Delta H°_{rxn} = \Sigma\, n\, \Delta H°f$ (products) $- \Sigma\, n\, \Delta H°f$ (reactants) and $\Delta S°_{sys} = \Sigma\, n$ So (products) $- \Sigma\, n$ So (reactants). Then use $\Delta G = \Delta H_{rxn} - T\,\Delta S_{sys}$ to find the change in free energy. Pay attention to units as $\Delta H°_{rxn}$ will be calculated in kJ and $\Delta S°_{sys}$ will be calculated in J. Be sure to reconcile units before finding ΔG.

- $\Delta H°_{rxn}$ =
- $\Delta S°_{sys}$ =
- $\Delta G°_{rxn}$ =

Do your lab observations support your results? Explain.

TEACHER'S KEY

Which of the following processes are spontaneous?

- Ice melts when dropped in a cup of warm water.

  Spontaneous

- Water evaporates when it is spilled on a hot surface.

  Spontaneous

- Water in a glass on your desk decomposes to hydrogen and oxygen.

  Not spontaneous

- Iron rusts in air.

  Spontaneous

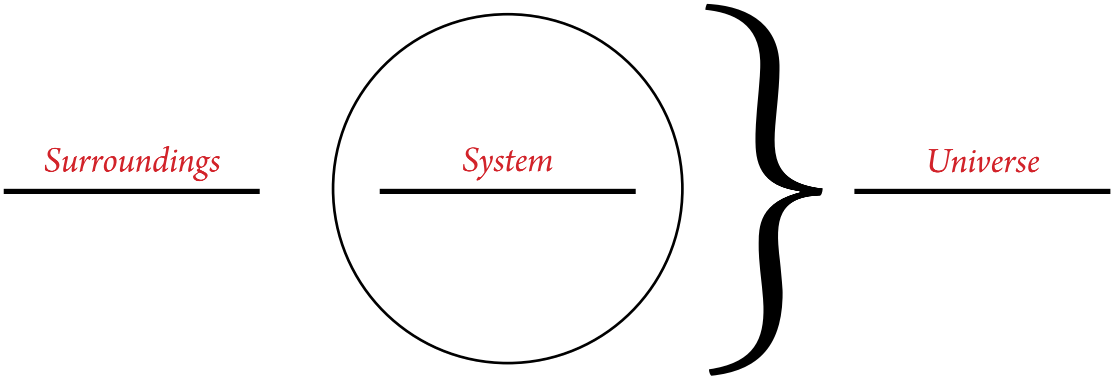
- The smell of perfume spray spreads throughout a room.

  Spontaneous

- Equal volumes of olive oil and vinegar dissolve together to make a salad dressing.

  Not spontaneous

Fill in the blanks below. Label the regions with the terms *system*, *surroundings,* and *universe*.



Gathering Evidence

1. Place 5.0 g of ammonium nitrate in a 150-mL beaker, then put the beaker on your lab sheet where you have designated the system. Write the formula for ammonium nitrate and identify what type of solid it is.

   Formula: NH_4NO_3
   Identify by underlining: *Ionic Solid* *Covalent Solid* *Metallic Solid*

2. Observe the macroscopic properties of the salt:

   White crystalline solid, small pellets 2–3 mm in diameter

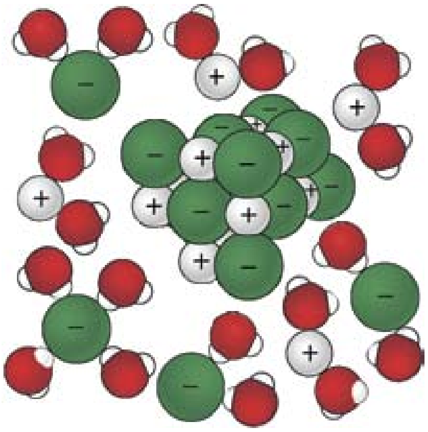
3. Complete the sentence: Ammonium nitrate is an *ionic* solid, and the individual particles in its lattice structure are *ions*.

Defining Entropy and Looking at Entropy Changes in a System

1. Pour 100 mL of water over the salt in the beaker and stir. Can you still see the ammonium nitrate?

   After stirring, the ammonium nitrate disappears; the solution appears transparent.

2. Make and record your observations.
3. Using the terms *cation, anion, solute, solvent,* and *solution,* label the diagram at right:



4. Given the physical state of ammonium nitrate before it dissolves, how do the possible arrangements of the ions in the salt compare to their possible arrangements when free to move within the solution?

   There are many more possible arrangements or possible positions for the ions after they have dissolved in water.

5. Does your answer to the preceding item suggest that the entropy of the ammonium nitrate increased or decreased upon dissolving?

   Entropy for the ions in ammonium nitrate has increased.

6. Would ΔS for this change be positive or negative?

   Positive

7. Would ΔS for this change in the arrangement of water molecules be positive or negative?

   There are fewer arrangements possible for the water molecules so the ΔS for the water molecules is negative.

8. Did dissolution of ammonium nitrate happen spontaneously?

   Yes

9. If the overall ΔS_{sys} for the dissolution of ammonium nitrate is positive, which ΔS in the system predominates: ΔS for the ions, or ΔS for the water molecules?

   Since ΔS for the ions is positive and ΔS for the water molecules is negative, the ΔS for the ions must predominate if the ΔS_{sys} is positive.

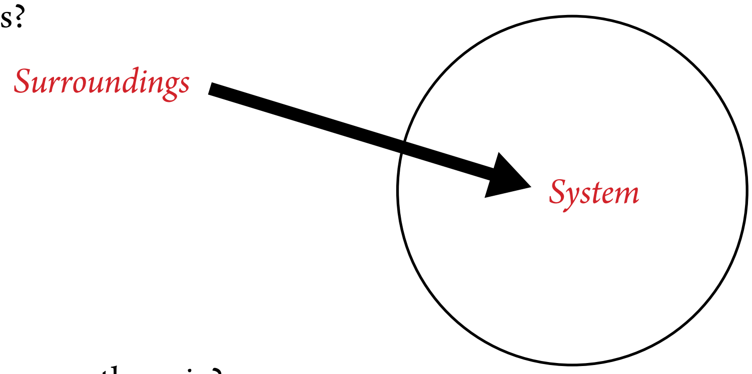
Defining Enthalpy and Looking at Entropy Changes in the Surroundings

Case one for enthalpy

1. Measure out 100 mL of water in a clean 150-mL beaker. Once again, label the "system" and the "surroundings" in the diagram below.
2. Place the beaker on your paper in the region labeled "system" below, then measure the temperature of its contents with a thermometer and record.
3. Add 5.0 g of ammonium chloride or ammonium nitrate.
4. While holding the beaker at its base, stir, and make and record your observations, including the final temperature of the mixture.
5. Your hand at the base of the beaker can be considered part of the surroundings. Was thermal energy transferred to your hand from the beaker, or away from your hand to the beaker?

   Since my hand feels cooler, thermal energy is being transferred from my hand to the beaker.

6. Label the diagram below with an arrow, showing the direction of the transfer of thermal energy. Is thermal energy transferred from the surroundings to the system or from the system to the surroundings?



7. Is the change endothermic or exothermic?

   Endothermic

8. Is ΔH of the system positive or negative?

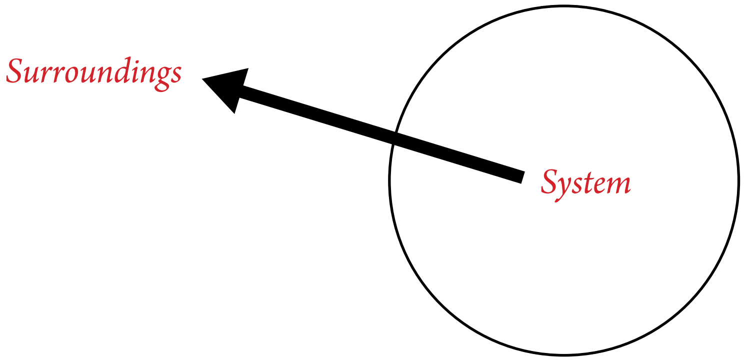
   Positive

Case two for enthalpy

1. Measure 100 mL of water in a clean 150-mL beaker.
2. Place the beaker in the "system" below, and measure and record the temperature of its contents. Add 5.0 g of calcium chloride.
3. While holding the beaker at its base, stir, and make and record your observations, including the final temperature of the mixture.
4. Was thermal energy transferred to your hand from the beaker, or away from your hand to the beaker?

   Since my hand feels warmer, thermal energy is being transferred from the beaker to my hand.

5. Label the diagram below with an arrow, showing the direction of the transfer of thermal energy. Is thermal energy transferred from the surroundings to the system or from the system to the surroundings?



6. Is the change endothermic or exothermic?

   Exothermic

7. Is ΔH of the system positive or negative?

   Negative

Relating the change in enthalpy (ΔH of the system) to a change in entropy of the surroundings (ΔS of the surroundings)

We are going to look at how entropy changes in the surroundings, depending on the sign of ΔH of the system. Before we do, however, let's review.

1. When thermal energy is transferred from region A to region B, the molecules in region B, on average, *speed up*.
2. When the average speed of the molecules in a region is high, there are *more* high-speed molecules. When the average speed of the molecules in a region is low, there are *fewer* high-speed molecules.
3. When the average speed of the molecules in a region is high, there are *more* possible arrangements for the molecules among the molecular speeds. When the average speed of the molecules in a region is low, there are *fewer* possible arrangements for molecules among the molecular speeds.

Now consider what happens in the *surroundings* during the process of thermal energy transfer.

4. When thermal energy is transferred in an endothermic change, will the molecules of the surroundings speed up or slow down?

   Slow down

5. How will this affect the entropy of the surroundings, will it increase or decrease?

   Decrease

6. In this case is ΔS of the surroundings positive or negative?

   Negative

7. When thermal energy is transferred in an exothermic change, will the molecules of the surroundings speed up or slow down?

   Speed up

8. How will this affect the entropy of the surroundings, will it increase or decrease?

   Increase

9. In this case, is ΔS of the surroundings positive or negative?

   Positive

Looking at entropy changes in the universe and the second law

Complete the following:

1. In an exothermic reaction, $\Delta S_{surroundings}$ is *positive*.
2. In an endothermic reaction, $\Delta S_{surroundings}$ is *negative*.
3. In a spontaneous, endothermic reaction, thermal energy is spontaneously *absorbed* by the system from the surroundings.
4. In a spontaneous, exothermic reaction, thermal energy is spontaneously *released* to the surroundings from the system.

Analyzing Evidence

1. So if ΔH_{sys} is negative and ΔS_{sys} is positive, what will the sign of ΔS universe always be?

   Positive

2. Under these conditions, is the reaction spontaneous?

   Yes

3. And under these conditions, what will the sign of ΔG always be?

   Negative

If ΔH_{system} is	Then $\Delta S_{surroundings}$ will be (+ or –)	And if ΔS_{system} is	And $\Delta S_{universe}$ will be (+ or –)	The ΔG will be (+ or –)	And the reaction is (spontaneous or not spontaneous)
negative or exothermic	+	positive (any temperature conditions)	+	–	spontaneous
positive, or endothermic	–	negative (any temperature conditions)	–	+	not spontaneous
negative or exothermic	+	negative (high temperature conditions)	–	+	not spontaneous

negative or exothermic	+	negative (low temperature conditions)	+	−	spontaneous
positive, or endothermic	−	positive (high temperature conditions)	+	−	spontaneous
positive, or endothermic	−	positive (low temperature conditions)	−	+	not spontaneous

Interpreting Evidence

Decomposition of baking soda

1. Gases are produced in the decomposition, therefore the sign of ΔS_{sys} for the process is likely to be *positive*.
2. Heat is absorbed in the process so the sign of ΔH_{sys} is *negative*.
3. Apply this information to $\Delta G = \Delta H_{sys} - T\,\Delta S_{sys}$. When is it likely that ΔG will be negative, in other words, when will the process be spontaneous, at high T or at low T?

   High T

Melting point of water

1. Determine the sign for ΔH_{sys}. Explain your choice.

   Positive. Energy must be provided to a system in order to melt a substance.

2. Determine the sign for ΔS_{sys}. Explain your choice.

   Positive. The liquid state has higher entropy than the solid state.

3. Will ΔG be negative at high or at low temperatures?

   Negative

4. We can use $\Delta G = \Delta H_{sys} - T\,\Delta S_{sys}$ to predict the temperature at which a substance will melt, if we have values for ΔH_{sys} and ΔS_{sys}. At some temperature (T), ΔG will change from positive to negative. What will be the value of ΔG at the temperature of this change?

   Zero

5. Substitute this value for ΔG and rearrange $\Delta G = \Delta H_{sys} - T\,\Delta S_{sys}$ to solve for T. This value of T will be what?

   $T = \Delta H_{sys}/\Delta S_{sys}$. This value of T will be the melting point.

Production of ozone

1. In this process, *three* moles of oxygen gas produce *two* moles of gaseous ozone. Are there more possible arrangements (higher entropy) for the reactants or the products? Therefore, is ΔS_{sys} positive or negative?

   Reactants, negative.

2. The process is endothermic, so ΔH_{sys} is *positive*.
3. Use $\Delta G = \Delta H_{sys} - T\,\Delta S_{sys}$ to consider: At what T will the reaction be spontaneous?

   At no T.

4. If a system will never spontaneously absorb thermal energy from the surroundings to undergo a change, does that mean the change is impossible? Could energy somehow be forced into the system, for example, by doing "work" on the system?

   No, change is not impossible. Energy is absorbed from another source.

5. In the case of the production of ozone in the stratosphere, which goes on every day, what kind of energy could be "working" on the oxygen gas?

   High-energy radiant energy from the sun.

Decomposition of hydrogen peroxide

1. What must be forming in the bubbles?

   Oxygen

2. When a gas is formed from a liquid, is the sign of ΔS_{sys} positive or negative?

   Positive

3. This decomposition is exothermic, so what is the sign of ΔH?

   Negative

4. So the decomposition of hydrogen peroxide is spontaneous at *all* temperatures.
5. Consider that the bottle you bought from the store has been sitting on the shelf for some time, and still contains H_2O_2. Does the conclusion that a reaction is spontaneous at all temperatures mean that the reaction happens quickly? What could have caused the reaction to speed up when the hydrogen peroxide was applied to your open cut?

   Something in the blood or on your skin—a catalyst.

Practice Problem

- ΔH°_{rxn} = *26.2 kJ*
  ΔH°_{rxn} = *−132.80 kJ + (−206.57 kJ) − (−365.6 kJ) = 26.2 kJ*

- ΔS°_{sys} = *108 J/mol K*
  ΔS°_{sys} = *112.8 J/mol K + 146.4 J/mol K – 151 J/mol/K = 108 J/mol K*

- ΔG°_{rxn} = *−5.98 kJ*
  ΔG°_{rxn} = *26.2 kJ – (298 K)(108 x 10-3 kJ/K) = −5.98 kJ*

Do your lab observations support your results? Explain.

> *The change in enthalpy has a positive value, and our observation was that the reaction was endothermic. The entropy of the system increased, which is consistent with a salt dissolving when the positive entropy change for the ions predominates over the negative entropy change for the water molecules. And the reaction happened spontaneously, which is consistent with a negative delta G.*

Energy & Entropy of a Stretched Rubber Band
A Lab Investigation

Summary

In this investigation, students work with a real-world item, a rubber band, to explore the concepts of Gibbs free energy, enthalpy, and entropy and their relation to the spontaneity of a physical process. They make observations of any thermal energy absorbed or given off by the rubber band as it is stretched and as it is allowed to contract by touching the band to their forehead. Then, they heat a stretched rubber band and observe whether its length stretches further, contracts, or stays the same.

Objective

Students learn about Gibbs free energy, enthalpy, and entropy, and the idea of assigning positive or negative values to each. They are then related to the spontaneity of stretching or contracting a rubber band.

Safety

- It is recommended that students wear goggles for this investigation, as rubber bands can break and could do so in a student's face.
- When heating the rubber bands, students should be careful not to hold the hair dryer or heat gun on one part of the rubber band for an extended period of time, as the rubber could melt and/or burn.
- If you elect to use the optional high-intensity light bulbs, be aware they get very hot. Also, without some sort of shade, they can be bright for the eyes.

Materials

- Wide rubber band for each student
- Several hair dryers or heat guns for class to share
- Materials for students to design set-up in step 6 (see Lab Tips below)

Time Required

One class period, approximately 45–50 minutes.

Lab Tips

Various lab equipment items could be placed in the laboratory area to inspire students as they design their method in step 6 to test the stretch/contraction of the rubber band as it is heated. These could include: ring stands, clamps, digital balances, weights, rulers, paper clips, safety pins, binder clips, and scissors. Different lab set-ups to measure a change in the rubber band could be:

1. Attach one end of a cut rubber band to a clamp that is attached to a ring stand. Attach a weight to the other end of the band. Adjust so the weight is partially resting on a digital balance, so only part of its weight registers. Any change in the length of the band will result in an increase or decrease in the balance reading.
2. Attach one end of a cut rubber band to a clamp that is attached to a ring stand. Attach a weight to the other end of the band. Use a ruler or other measurement scale next to the weight to observe any change in length. A wooden splint could be attached to the weight, with the end sticking out from the side of the weight, to more easily view any changes.
3. An uncut band could be looped around a doorknob or drawer pull, with a weight or heavy object looped through the bottom of the band. Use a ruler or other measurement scale next to the weight to observe any change in length. A wooden splint could be attached to the weight or object, with the end sticking out from the side of the weight, to more easily view any changes.

The best results are obtained if the band is stretched close to its maximum length. Instead of a hair dryer or heat gun, high-intensity light bulbs could be used.

Pre-Lab Discussion

This investigation introduces the concepts of Gibbs free energy (G), enthalpy (H), entropy (S), and spontaneity. These terms, their symbols, and their connection to each other in the equation $\Delta G = \Delta H - T\Delta S$ could be discussed before the lab; however, the investigation can be done without this discussion.

It can be pointed out that whether a process is spontaneous or not is not related to how quickly the process occurs. Other examples of spontaneous physical changes could be discussed. One common example is the melting of ice, which at ordinary pressures is spontaneous at temperatures above 0 °C.

Integrating into the Curriculum

This investigation would fit into units on thermodynamics and polymers.

PREPARING TO INVESTIGATE

In this investigation, you will make observations of an everyday item, a rubber band. Based on what you already know about rubber bands and how they behave, consider the following:

1. Picture holding a rubber band by looping your thumbs through the band so the band sits on your thumbs, but is unstretched. If you wait, will the band ever spontaneously stretch and fall off your thumbs?
2. Picture holding a rubber band between your thumbs as before, but this time stretched to nearly its limit. If you quickly pull a thumb out of the band, will the band spontaneously remain stretched out?

The two situations above may sound ridiculous. They describe the opposite of your past experiences with rubber bands. Anyone who has ever let fly a rubber band from her fingers or had a rubber band suddenly break knows that she can reasonably expect a stretched rubber band to contract when released. This change is spontaneous, and at everyday temperatures and pressures happens quite quickly. The reverse change, for a rubber band to go from a contracted position to a stretched position without an outside force acting on it, such as the pull of your thumbs, is not spontaneous.

Is there a way to determine if a change will be spontaneous or not? There is a quantity called Gibbs free energy (G) that can help. For a change in a system, if the amount of free energy decreases, that is, if $\Delta G < 0$, the change is spontaneous.

Different variables contribute to ΔG. Expressed as an equation, $\Delta G = \Delta H - T\Delta S$, where ΔH is change in enthalpy, T is temperature in Kelvin, and ΔS is change in entropy. All quantities are from the viewpoint of the system. Enthalpy is a measure of the internal energy of the system. We cannot measure the enthalpy of a system directly, but can measure changes in enthalpy as the thermal energy released or absorbed when changes in the system occur at constant pressure. From the viewpoint of the system, if thermal energy is released by the system, it has been transferred out of the system, and ΔH would have a negative sign. Entropy is a way of expressing how many different ways the particles and energy of a system can be arranged. For changes, nature favors increases in entropy ($\Delta S > 0$) that will maximize the number of different possible arrangements of particles and energy in the system.

In this investigation, you will observe whether thermal energy is absorbed or released from a rubber band as it is stretched and allowed to contract. Then, you will design a way to measure any stretching or contraction of a rubber band when it is heated. Using your observations, you will consider the changes in enthalpy and entropy that occur in the system of the rubber band, and how they contribute to the system's free energy changes.

GATHERING EVIDENCE

1. Obtain a wide rubber band. Hook both your thumbs into the rubber band, so you can use your thumbs to stretch out the band.

2. Hold the rubber band hooked on your thumbs so that the band is just snug enough to remain on your thumbs, but is not stretched out. Hold the wide, flat part of the band to your forehead. Does it feel warmer, cooler, or the same as your forehead? Repeat this step a few times to be sure of your initial observation.
3. Hold the rubber band hooked on your thumbs. Hold it near your forehead but not touching it. Quickly stretch the band out to the sides and immediately touch the wide, flat part of the band to your forehead. In the “Observations” column of the data table below, record your observations of whether it feels warmer, cooler, or the same as your forehead. Repeat this step a few times to be sure of your initial observation. In between the different tries, allow the band to contract to its original size and hold it unstretched for about 20 seconds to be sure it is at its original temperature.
4. Hold the rubber band hooked on your thumbs. Stretch the band out to the sides. Keep the band stretched out for about 20 seconds to be sure it is at its original temperature. While still looping your thumbs into the band, quickly allow the band to contract to its original position and immediately touch the wide, flat part of the band to your forehead. In the “Observations” column of the data table below, record your observations of whether it feels warmer, cooler, or the same as your forehead. Repeat this step a few times to be sure of your initial observation.
5. Based on your observations in steps 3 and 4, predict what will happen if you heat a stretched rubber band.
6. Design and set up a method, using materials your instructor has made available, to qualitatively measure any change, stretching, contracting, or staying the same, in the length in a stretched rubber band as it is heated with a hair dryer or heat gun. You may cut the rubber band into a single long strip if you wish. Keep the hair dryer or heat gun moving slowly to avoid overheating and melting the rubber band.
7. Use your method to observe any change in the length of the stretched rubber band as it is heated with a hair dryer or heat gun. In the “Observations” column of the data table below, record your observations.

Process	Observations	System absorbs or gives off thermal energy?	ΔH +/–	Process spontaneous? yes/no	ΔG +/–
stretching band (step 3)					
contracting band (step 4)					
heating band (step 7)					

ANALYZING EVIDENCE

1. Based on the temperature changes you observed when you held the stretched rubber band and the contracted rubber band to your forehead in steps 3 and 4, determine whether the system (the rubber band) absorbs or gives off thermal energy in these cases. Use this information to fill in the column "System absorbs or gives off thermal energy?" for "stretching band (step 3)" and "contracting band (step 4)."
2. Did the system (the rubber band) absorb or give off thermal energy when you used the hair dryer/heat gun in step 7? Use this information to fill in the column "System absorbs or gives off thermal energy?" for "heating band (step 7)."
3. Is the change in enthalpy (ΔH) for a system positive or negative if the system absorbs thermal energy? If the system gives off thermal energy? Remember, the sign is from the perspective of the system. Has the thermal energy in the system increased or decreased? Use this information to fill in the column "ΔH +/–" with the appropriate signs.
4. Fill in the column "Process spontaneous? yes/no," depending on whether the change that occurred was spontaneous.
5. Is the change in free energy (ΔG) for a system positive or negative when the system undergoes a spontaneous reaction? Use this information to fill in the column "ΔG +/–" with the appropriate signs.

INTERPRETING EVIDENCE

1. If the process of a stretched rubber band returning to its original size by contracting was written as the equation: stretched band $\longrightarrow$ contracted band, would "thermal energy" be shown in the equation as a reactant or product? Explain.
2. Is allowing a stretched band to contract an exothermic or endothermic reaction? Describe the direction of the flow of thermal energy between the system of the rubber band and its surroundings, which include your forehead.
3. As discussed in Preparing to Investigate, a rubber band contracting is a spontaneous process, meaning ΔG for that process is negative. Using the equation below and based on your observations during the investigation, label the variables ΔG, ΔH, and T with the appropriate sign, + or –, for the contraction process.

$$\Delta G = \Delta H - T\Delta S$$

4. What should the sign for ΔS be in the equation above for the contraction process to be spontaneous?
5. How did your observations in steps 3 and 4 when you touched the rubber band to your forehead relate to the result you saw in step 7 when you heated the rubber band?

REFLECTING ON THE INVESTIGATION

1. Based on the sign for ΔS, positive or negative, that you answered for question 4 in Interpreting Evidence, explain how the arrangement of the cross-linked polymer chains in the rubber band can result in an increase or decrease in entropy when it is contracted compared to when it is stretched. Explain or draw what happens on the molecular level.
2. How does the action of the heated rubber band in step 7 compare with what happens when other everyday materials, such as metal, are heated?
3. Picture using the method and set-up you designed in step 6 of Gathering Evidence, but instead of heating the stretched band, you cool it. Predict what would happen. Explain.

TEACHER'S KEY

In this investigation, students work with a real-world item, a rubber band, to explore the concepts of Gibbs free energy, enthalpy, and entropy, and their relation to the spontaneity of a physical process.

Analyzing Evidence

Process	**Observations**	**System absorbs or gives off thermal energy?**	**ΔH +/–**	**Process spontaneous? yes/no**	**ΔG +/–**
stretching band (step 3)	*feels warm*	*gives off thermal energy*	–	*no*	+
contracting band (step 4)	*feels cool*	*absorbs thermal energy*	+	*yes*	–
heating band (step 7)	*band contracts*	*absorbs thermal energy*	+	*yes*	–

Interpreting Evidence

1. If the process of a stretched rubber band returning to its original size by contracting was written as the equation: stretched band → contracted band, would "thermal energy" be shown in the equation as a reactant or product? Explain.

   Thermal energy would be shown as a reactant. Thermal energy is added to the system, the rubber band, as it contracts. This is why the band feels cool on the forehead, since energy flows from the forehead to the band.

2. Is allowing a stretched band to contract an exothermic or endothermic reaction? Describe the direction of the flow of thermal energy between the system of the rubber band and its surroundings, which include your forehead.

   It is an endothermic reaction. Thermal energy flows from the surroundings (forehead) to the system (rubber band). This makes the band feel cool on the forehead.

3. As discussed in Preparing to Investigate, a rubber band contracting is a spontaneous process, meaning ΔG for that process is negative. Using the equation below and based on your observations during the investigation, label the variables ΔG, ΔH, and T with the appropriate sign, + or –, for the contraction process.

$$\Delta G = \Delta H - T\Delta S$$

$$(-) = (+) - (+)(\Delta S)$$

4. What should the sign for ΔS be in the equation above for the contraction process to be spontaneous?

   The sign for ΔS must be positive in order for the right-hand side of the equation to be negative to agree with the sign of the left-hand side.

5. How did your observations in steps 3 and 4 when you touched the rubber band to your forehead relate to the result you saw in step 7 when you heated the rubber band?

   When the band stretched, it gave off thermal energy, which made the forehead feel warm. When the band contracted, it absorbed thermal energy, which made the forehead feel cool. When the band was heated with a hair dryer, the system was absorbing thermal energy. The system absorbing thermal energy is connected with the process of contracting.

Reflecting on the Investigation

1. Based on the sign for ΔS, positive or negative, that you answered for question 4 in Interpreting Evidence, explain how the arrangement of the cross-linked polymer chains in the rubber band can result in an increase or decrease in entropy when it is contracted compared to when it is stretched. Explain or draw what happens on the molecular level.

   When the band is stretched, the cross-linked polymer chains straighten out and are more aligned with each other, like a bundle of uncooked spaghetti noodles. There are fewer potential arrangements of particles in this situation. When the band is allowed to contract, the chains are not as straight and become more tangled, like a mass of cooked spaghetti noodles. There are more potential arrangements of particles in this situation.

2. How does the action of the heated rubber band in step 7 compare with what happens when other everyday materials, such as metal, are heated?

   Typically, materials such as metal expand when they are heated, rather than contracting as the rubber band did.

3. Picture using the method and set-up you designed in step 6 of Gathering Evidence, but instead of heating the stretched band, you cool it. Predict what would happen. Explain.

   Since thermal energy is being given off by the system, or removed from it, the rubber band should stretch.

Post-Lab Discussion

A review of the equation $\Delta G = \Delta H - T\Delta S$, its variables, and what a positive or negative sign for each variable means could be useful. Depending on their answers to Reflecting on the Investigation question 1, students might need a further discussion of what happens on the molecular level as a rubber band stretches/contracts.

Extensions

Students could also investigate the effect of heating the rubber band to different temperatures, such as using low/high settings on a hair dryer or various settings on a heat gun. Does the temperature change the amount of contraction?

An additional polymer to explore is a toy called Shrinky Dinks. Pieces of this plastic can be decorated with colored pencils or markers, then baked. When heated, the plastic contracts dramatically, resulting in a piece approximately one-third its original length and width, but much thicker. The change of this material is not easily reversible and could be contrasted with that of rubber bands, which can be stretched and contracted over and over. Instead of buying commercially available kits, transparent deli containers with the recycle code #6 can also be used.

As an advanced, challenging extension, students could construct a rubber band heat engine using directions available online. The spokes on a bicycle wheel rim are replaced with rubber bands. Each spoke is cut, and the ends remaining on the wheel are bent into hooks so the rubber band can be looped around them. The wheel is then mounted so it can rotate freely. A heat lamp is placed near one area of the rubber bands. As the lamp heats the rubber bands, they contract, which moves the wheel. The heated rubber bands rotate away and cool down, with different rubber bands coming close to the lamp to heat up and continue the process. Directions are available online:

http://bit.ly/highschoolnrg17

Videos of constructed rubber band heat engines can also be found online. The engine was described by physicist Richard Feynman (Feynman, R. P.; Leighton, R. B.; Sands, M. *The Feynman Lectures on Physics*, Vol. I: The New Millennium Edition: Mainly Mechanics, Radiation, and Heat. New York: Basic Books, 2010, p 44-2):

http://bit.ly/highschoolnrg18

Additional Resources

- Summerlin, L. R.; Ealy, J. L., Jr. *Chemical Demonstrations: A Sourcebook for Teachers*, Volume 1, second ed. Washington, DC: American Chemical Society, 1988, p 73.
- Hirsch, W. "Rubber Bands, Free Energy, and Le Châtelier's Principle." *J. Chem. Educ.*, 2002, 79 (2), pp 200A–B.

Why COLD Doesn't Exist

By Michael Tinnesand

Pop Quiz!

Q: What happens when an ice cube is added to warm soda?

- **a. Heat is transferred from the soda to the ice.**
- **b. The ice cube's "cold content" is transferred to the soda.**
- **c. Both heat and cold are transferred between the ice and soda.**

In our daily lives, we see many examples of warm objects that become colder, but it may not be obvious how it happens. Let's see how things become cold (and get the answer to our pop quiz).

Collisions in your soft drink

In the case of the ice cubes in the soft drink, the soft drink becomes colder and the ice cubes warmer because energy is transferred from the soft drink to the ice, not the other way around, so the best answer is "a." The reason is that energy is always transferred from a body at higher temperature to a body at lower temperature and, as a result, there is no such thing as "cold" flowing from the ice to the beverage.

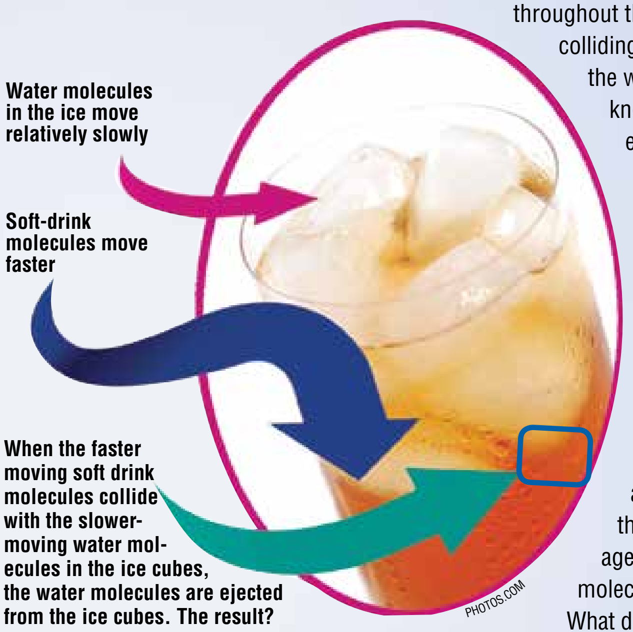
To understand how this works, let's consider the definitions of "temperature" and "kinetic energy." In any sample of matter, particles—atoms, ions, or molecules—can move in three different ways: They can vibrate (wiggle about a fixed position), translate (move from one location to another), or rotate (spin around).

At any time, these particles are constantly moving—slowly in solids, and more noticeably in liquids and gases, and as they do, they collide with one another. Each time these collisions occur, the speeds and directions of the colliding particles change. So the particles have a wide range of speeds and, similarly, a wide range of kinetic energies (because kinetic energy is proportional to the square of the speed). The values of the kinetic energies of all the particles follow a distribution that looks like the one shown in Fig. 1.

A common way to discuss the kinetic energy of all the particles in a sample is to use their average kinetic energy, which lies around the peak of the distribution. This average kinetic energy varies with temperature (Fig. 1). The more the particles vibrate, translate, or rotate, the greater the temperature of the sample.

Let's return to the example of ice in a warm soft drink. The particles in the soft drink move throughout the space of the container, colliding with each other and with the walls of the container. This is known as translational kinetic energy, and it is the main form of kinetic energy for gases and liquids. The water molecules in the ice vibrate about a fixed position and, as a result, their kinetic energy is mainly vibrational. The molecules in the soft drink move faster than the water molecules in the ice cubes because they are at a higher temperature and, therefore, have a higher average kinetic energy than the water molecules.

What does the energy transfer look like? When the faster-moving soft-drink molecules collide with the slower-moving water molecules, part of the kinetic energy of the soft-drink molecules is used to move the water molecules out of their crystal lattice. The net effect is that the kinetic energy of the soft-drink molecules is reduced and the water molecules have the same kinetic energy that they initially had in the ice crystal, except that this kinetic energy is translational rather than vibrational. The total amount of energy transferred from the soft-drink molecules to the water molecules in the ice cubes is called heat.



When the faster moving soft drink molecules collide with the slower-moving water molecules in the ice cubes, the water molecules are ejected from the ice cubes. The result? The ice melts and the soft drink cools.

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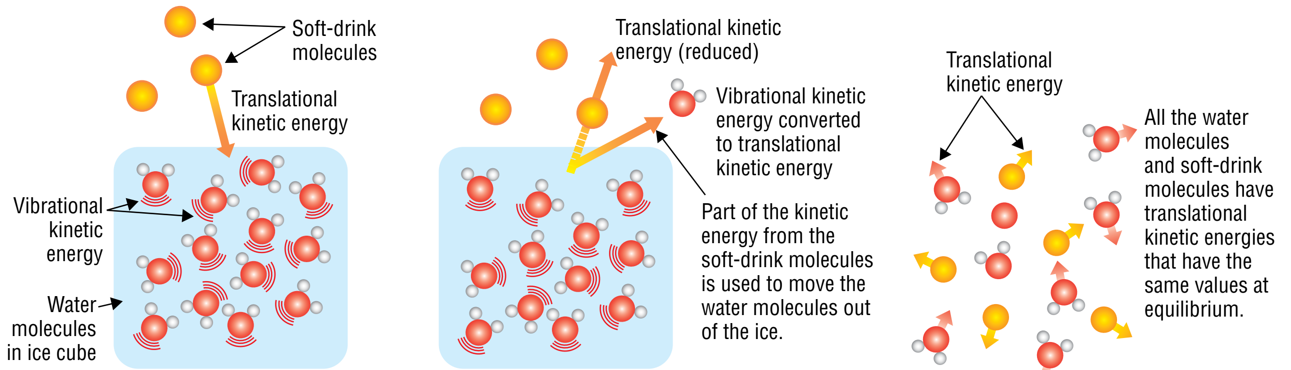
How Ice Cubes Melt and Cool Your Soft Drink

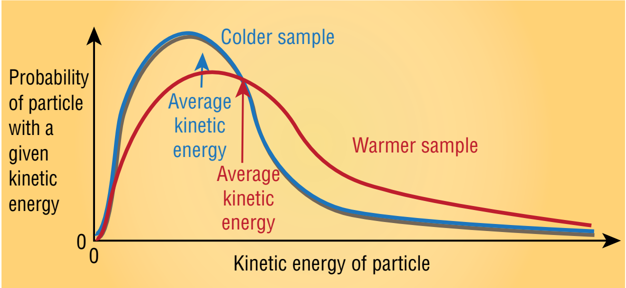


A The molecules of water in the ice cubes vibrate (they have vibrational kinetic energies), and the molecules in the soft drink move along straight lines in between collisions with other water molecules (they have translational kinetic energies).

B When a molecule in the soft drink collides with a water molecule, part of the translational kinetic energy is used to move the water molecule out of the ice, and the water molecule's vibrational kinetic energy is converted to translational kinetic energy.

C As more collisions between soft-drink molecules and water molecules occur, the soft-drink molecules keep transferring energy to the water molecules, until the temperatures of the soft-drink and water molecules are identical. This is known as equilibrium, in which case most of the ice cubes have melted, and both the soft-drink and water molecules are in liquid form.

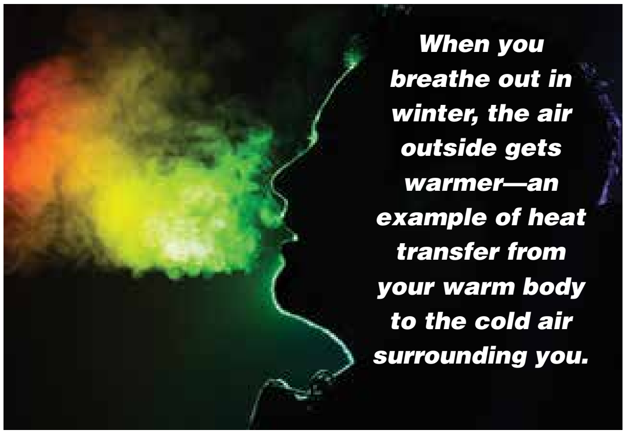
ANTHONY FERNANDEZ



RS GRAPHX

Figure 1. Distribution of the kinetic energies of particles in a sample at two different temperatures. The fraction of particles in the sample that have a given kinetic energy is shown by the height of the curve for that value of kinetic energy.

The collisions will continue to transfer energy until the temperatures of the ice molecules and soft-drink molecules are identical. This is known as thermal equilibrium. When thermal equilibrium has been reached, the average kinetic energy of both types of particles is equal, and the final temperature—assuming there is still ice left—will be about 0 °C.



When you breathe out in winter, the air outside gets warmer—an example of heat transfer from your warm body to the cold air surrounding you.

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Cold finger

Thinking about heat and cold in this way can be useful in explaining how heat transfer works in everyday life. It can help us explain common occurrences such as evaporation.

If you put your finger in water at the same temperature as your finger and then wave it in the air, it feels cooler. Why? Some of the water molecules with higher energy have enough energy to escape as water vapor, leaving behind water molecules with a slightly lower average energy, or a little lower temperature. Molecules in your finger collide with the molecules of water and transfer some of their kinetic energy to the water, thus slightly lowering the finger temperature. This is heat transfer due to the evaporation of water molecules on the surface of your skin.

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This happens over and over again, as higher energy water molecules vaporize, leaving water molecules with lower average kinetic energy, so heat transfer continues from your finger to the cooler water molecules. This net transfer of energy goes from your finger to the water, and your finger feels cooler.

So, does cold exist? Can we say that it is really cold outside? Yes, we can. Cold is a perfectly fine adjective to describe when something is not hot, or when its temperature is low. We sometimes use words to describe conditions that reflect the absence of something rather than the presence of something else. For example, there is no such thing as darkness. There is only light or lack of light. But that's a story for another time! CM

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Michael Tinnesand is a science writer and education consultant who lives in Portland, Ore. His latest *ChemMatters* article, "Drivers, Start Your (Electric) Engines!" appeared in the February 2013 issue.