# UNIT 7

# The Energy in Chemical Reactions

## Thermochemistry and Reaction Energies

## Unit Overview

Unit 7 introduces students to thermochemistry, the study of energy in chemical reactions. After completing this unit, students should be able to understand the various types of energy along with basic thermodynamic terms: system, surroundings, heat, and work. The unit also explains the first law of thermodynamics and provides an introduction to enthalpy and calorimetry.

### Learning Objectives and Applicable Standards

Participants will be able to:

1. Describe the different kinds of energy: kinetic, potential, chemical, and thermal.
2. Describe how the internal energy of a system can increase or decrease through heat and work.
3. Explain the First Law of Thermodynamics (the conservation of energy) and give examples of one form of energy converting into another.
4. Define the terms endothermic and exothermic.
5. Describe Hess's Law and what it is used for.
6. Explain how bond enthalpies can be used to find the change in enthalpy ($\Delta H$) of a reaction.

### Key Concepts and People

1. **Kinetic and Potential Energy:** Objects can possess both kinetic and potential energy. When a substance goes through a phase change, its particles gain or lose kinetic energy.
2. **Energy Changes in Chemistry:** Chemical energy is potential energy that is stored in chemical bonds. This energy can be released by exchanging high energy bonds for low energy bonds. Thermal energy is the total energy of all of the molecules in a substance.
3. **Heat, Work, and Internal Energy:** Heat and work are not energy, but rather describe ways energy enters and leaves a system.

4. **Energy Conversions and Conservation:** While energy frequently changes from one form to another, the total amount of energy in an isolated system does not change. The First Law of Thermodynamics explains that energy cannot be created nor destroyed.

5. **Enthalpy:** The change in enthalpy ($\Delta H$) expresses the amount of heat produced or absorbed by a chemical reaction at constant pressure. Chemical reactions either release heat (exothermic) or absorb heat (endothermic).

6. **Calorimetry:** A calorimeter is a container that can measure the amount of heat released or absorbed by any chemical process.

7. **Hess's Law:** Hess's Law explains that the change in enthalpy of a chemical reaction depends only on the enthalpy of the reactants and the products, and not on anything that happens in between.

8. **Bond Enthalpies:** Breaking chemical bonds requires enthalpy (endothermic), and forming chemical bonds releases enthalpy (exothermic). Determining if a chemical reaction as a whole is endothermic or exothermic depends on whether or not breaking the old chemical bonds requires more enthalpy than the enthalpy released from forming the new chemical bonds. The amount of enthalpy associated with a bond is called "bond enthalpy."

9. **Standard Enthalpies of Formation:** The standard enthalpy of formation refers to the enthalpy change of the formation of one mole of a compound from its constituent elements in their most stable state at 1 atm of pressure and 25°C.

## Video

The phrase "chemical reaction" conjures up images of explosions, bubbling gases, flames, and smoke. So many chemical reactions have visible results because energy is being transferred from one form to another—the realm of thermodynamics. Thermodynamics provides rules for predicting the progress of a reaction and for harnessing the energy released. It is key to solving pressing engineering problems, such as making the next generation of cleaner, and more efficient, automobile engines.

### VIDEO CONTENT

#### Host Introduction
#### "Exothermic Reactions"

Dr. Nicole Labbe, a chemical engineer at the Argonne National Laboratory at the University of Chicago, introduces the first law of thermodynamics and enthalpy by describing what happens when a log burns.

### Laboratory Demonstration
### "Flame Tornado"
Daniel Rosenberg, Lecture Demonstrator at Harvard University, creates a flame tornado to show a visually stunning example of an exothermic reaction.

### Laboratory Demonstration
### "Heat Absorber"
To show an endothermic reaction, Daniel Rosenberg reacts barium hydroxide and ammonium chloride. To prove the reaction is endothermic, he places water between a block of wood and the beaker in which the reaction is taking place. Because the reaction is endothermic, it takes enthalpy from the surroundings, which means it takes some enthalpy from the water between the beaker and the block. This causes the water to freeze.

### Host Science Explanation
### "Measuring Heat"
Dr. Nicole Labbe demonstrates how a coffee cup calorimeter is used to measure the heat released when sodium hydroxide dissolves in water.

### Real World Application
### "Mississippi Biofuels"
By using bomb calorimetry, Dr. Gretchen Sassenrath is leading a study for the U.S. Department of Agriculture to determine which crops in Mississippi could best serve as a potential alternative fuel source.

### Host Science Explanation
### "Bond Enthalpy"
Dr. Nicole Labbe explains how to calculate values of enthalpies of a given reaction by looking up the bond enthalpies of the products and reactants of that reaction. Chemists have been doing experiments for hundreds of years to determine bond enthalpies for various chemicals, which are now compiled in tables. Bond enthalpies can help predict the mechanism of a given reaction.

### History of Chemistry
In thermodynamics, work is any transfer of energy in a system that is not a heat transfer. Work can create heat, and heat can create work. During the days of steam power, scientists wondered if an engine could perform a heat–to-work conversion with 100% efficiency. Using a theoretical piston engine, Nicolas Leonard Sadi Carnot (1798-1832), a French military engineer and physicist, was the first to realize that it is impossible. Carnot's discovery led to the second law of thermodynamics.

### Current Chemistry Research
### "Fighting the 2$^{nd}$ Law"
Kevin Cedrone, a graduate student at MIT's Sloan Automotive Lab, studies how to get as much work out of an engine as possible. He strives to get engines to run as cleanly and efficiently as possible.

# Unit Text

### Content Overview

Energy is the ability to do work, and it comes in many forms. The unit begins with a discussion of kinetic energy, potential energy, thermal energy, latent heats of fusion/vaporization, and chemical energy. These last three forms of energy make up the internal energy of a system. A system can exchange energy with the surroundings in the forms of heat and/or work.

Next, the unit discusses the work of Thomson and Joule, which demonstrated that energy cannot be created nor destroyed (the First Law of Thermodynamics), but it can be converted from one form to another.

The First Law of Thermodynamics is followed by a discussion of enthalpy. The enthalpy change ($\Delta H$) of a system is the change in internal energy plus the work done by or to the system. $\Delta H$ can be positive (endothermic) or negative (exothermic). Calorimeters are used to measure the enthalpy change of a reaction. The final sections of the unit cover bond enthalpy and the energy values of food.

### Sidebar Content

1. **State Functions:** This sidebar explains the difference between state and path-dependent functions.
2. **Food Energy:** Gram for gram, fat stores more energy than protein and carbohydrates.
3. **Nitrogen's Triple Bond:** The Haber-Bosch process is an industrial process that converts nitrogen gas ($N_2$), into fertilizer, which contains nitrogen that plants can use to grow. The Haber-Bosch process has also made the production of explosives more economical.

# Interactives

### Historical Timeline of Chemistry

This interactive illustrates how different discoveries build upon, disprove, or reinforce previous theories. This not only reinforces basic chemistry concepts, but also emphasizes the nature of science. Scientists mentioned in this unit are listed on the timeline.

### Control a Haber-Bosch Ammonia Plant Interactive

This interactive allows students to control an ammonia plant to maximize profit. Students will be able to observe the effect of changing the temperature on the production of ammonia. The background information also discusses the role of bond enthalpy in the productions of artificial fertilizer. Please note that a lesson plan and student worksheet are available online with this interactive.

### Chemistry of Running Interactive

Students will try to maximize the performance of their avatar in a marathon using information from this unit. The body can extract the most energy from food using oxygen in aerobic respiration (glycolysis). When oxygen delivery is maxed out, the body also performs anaerobic respiration, which causes the buildup of lactic acid.

## During the Session

### Before Facilitating this Unit

There are many concepts in this unit that are traditionally difficult for students to understand. For instance, there are a lot of common misconceptions that make learning this material challenging. Pay close attention to the tips and suggestions section and use the activities and demonstrations to help reinforce these concepts. Be sure to leave plenty of time for class discussion. The video highlights the main ideas in this unit and can provide students with real world applications to these challenging concepts.

### Tips and Suggestions

1. **Students often think heat and temperature are the same.** Be sure to clearly and explicitly explain the difference between heat (thermal energy measured in Joules) and temperature. This guide provides a demonstration that helps show the difference. (See In-Class Demonstrations.)

2. **A common misconception is that breaking bonds releases energy, and making bonds takes energy.** In fact, breaking bonds takes energy and making bonds releases energy. To drive this point home, use the analogy of breaking a stick; breaking the stick requires effort (an input of energy).

3. **Some students find it hard to remember that exothermic reactions warm up and endothermic reactions cool down.** Because we describe exothermic reactions as releasing energy, students conclude that if something loses energy, it should cool down. Emphasize that the chemical bonds lose energy in an exothermic reaction, and the surrounding solution then heats up.

4. **Some students are confused by the definition of "moles of reaction."** The units of enthalpy ($\Delta H$) are kilojoules per mole, and students sometimes ask, "per mole of what?" The moles in this case are moles of the reaction as written. In other words, the $\Delta H$ for the reaction 2Ag2B will be double the $\Delta H$ for the reaction $A \rightarrow B$.

### Starting the Session: Checking Prior Thinking

You might assign students a short writing assignment based on the following questions, and then spend some time discussing prior thinking. This will help elicit prior thinking and misconceptions.

1. What is energy? What different kinds of energy can you name?
2. What happens to the energy of gasoline after it burns inside a car engine?
3. What chemical reactions can you think of that release energy? What chemical reactions can you think of that absorb energy (or require energy to happen)?
4. Why do you think some reactions release energy, and some absorb energy?
5. Why does fatty food contain a lot of calories? And why does your body store extra calories as fat?

## Before Watching the Video

Students should be given the following questions to consider while watching:

1. What is the First Law of Thermodynamics?
2. What is "enthalpy"?
3. What is an exothermic reaction? Is the change in enthalpy positive or negative?
4. What is an endothermic reaction? Is the change in enthalpy positive or negative?
5. What are calorimeters used for?
6. What is the difference between food calories and chemistry calories?
7. What is bond enthalpy, and how are the values used?
8. What is work?
9. What is the Second Law of Thermodynamics?
10. What is a PV diagram used for?
11. If an engine produces a lot of heat, what does that tell you about its mechanical efficiency?

## Watch the Video

## After Watching the Video

Use these additional questions as follow-up, either as a group discussion or as short writing assignments:

1. What practical uses can you think of for exothermic chemical reactions? For endothermic chemical reactions?
2. Why is it technically incorrect to say that a power plant or engine "produces" energy?

3. Electrical energy is generated in different ways, such as burning coal, hydroelectric dams, nuclear power plants, wind turbines, and solar panels. What forms of energy do these use as their sources?

4. What other uses could you think of for calorimetry?

5. Why might a chemist use bond enthalpies to calculate a $\Delta H$ value instead of calorimetry?

6. As you saw in the video segment about the car engine, some energy is always lost in the form of heat. How could a car engine or a power plant make use of this energy instead of letting it go to waste?

# Group Learning Activities

## Finding the Enthalpy of Reaction ($\Delta H$)

### Objective

In this activity students will determine the $\Delta H$ values for two chemical reactions.

### List of Materials

Each group requires:

- 1 Styrofoam coffee cup
- 50 mL of 1 M HCl
- 50 mL of 1 M NaOH
- 15 g sodium acetate
- Thermometer
- Scale
- Weigh boats

### Procedure (Part 1: The reaction of NaOH(aq) and HCl(aq))

1. Write the balanced equation for the following reaction:

   $NaOH(aq) + HCl(aq) \rightarrow$

2. Weigh an empty coffee cup. Record the mass.

3. With a graduated cylinder, obtain 50 mL of 1 M HCl and pour it into the coffee cup.

4. Rinse the graduated cylinder and obtain 50 mL of 1 M NaOH.

5. With a thermometer, measure the temperature of the HCl.

6. Pour the NaOH into the coffee cup and stir gently with the thermometer.

Record the highest temperature that the thermometer reaches.

7. Weigh the cup and the liquid.

## Discussion

The following data table can help students when recording information for this experiment.

| | |
|---|---|
| Mass of empty cup = ___________ | Starting temperature = ___________ |
| Mass of cup with liquid = ___________ | Ending temperature = ___________ |
| Mass of liquid = ___________ | Change in temperature = ___________ |

The following questions can help guide students thinking about this activity:

1. Was the reaction exothermic or endothermic? How do you know?
2. Describe the flow of energy in the system.
3. Calculate the amount of heat the reaction gave off using $q = mc\Delta T$ ($c = 4.18$)
4. Calculate the moles of HCl and NaOH that reacted.
5. Calculate the $\Delta H$ of this reaction using $q = n\Delta H$
6. If the solutions you used were more concentrated, would the value of q that you calculated in question 2 be higher or lower?
7. If the solutions you used were more concentrated, would the value of $\Delta H$ that you calculated in question 4 be higher or lower?

## Procedure (Part 2: The Dissolving of Sodium Acetate ($NaC_2H_3O_2$))

1. Write the balanced equation for this reaction.

   $NaC_2H_3O_2 + H_2O \rightarrow$
2. Obtain 15 g of $NaC_2H_3O_2$ (sodium acetate) on a weighing boat.
3. Break any large chunks of the $NaC_2H_3O_2$ into pea-size pieces. (It will not dissolve fast enough if the pieces are too large.)
4. With a graduated cylinder, obtain 100 mL of tap water. Pour it into the coffee cup and measure the temperature.
5. Place the $NaC_2H_3O_2$ in the water, stir gently with the thermometer, and record the lowest temperature reached.

6. Weigh the cup and the liquid.

### Discussion
The following questions can help guide students thinking about this activity:

1. Was the reaction exothermic or endothermic? How do you know?
2. Calculate the $\Delta H$ of the reaction of sodium acetate with water.
3. What practical use might this reaction have?

### Hazards
It is good lab practice to review a chemical's Material Safety Data Sheet (MSDS) before working with any chemical. Follow instructions on the MSDS and encourage students to review them. Wear proper protective gear at all times: chemical splash goggles, chemical-resistant apron, lab coat, and gloves. This activity should be performed in a fume hood. Hydrochloric acid: Toxic by ingestion or inhalation; severely corrosive to skin and eyes. Sodium hydroxide: Corrosive liquid; skin burns are possible; very dangerous to eyes; wear gloves. Sodium acetate: Skin, eye and respiratory irritant. LD50 3530 mg/kg.

### Disposal
Check local regulations for proper disposal of the chemicals.

# In-Class Chemical Demonstrations

## Difference Between Heat (Thermal Energy) and Temperature

### Objective
In this demonstration the same amount of thermal energy is added to different volumes of water. The temperature of each volume of water will be different even though the same amount of heat is added to each container of water. The smaller volume will have a higher temperature than the larger volume because you are adding more heat per molecule into the system.

### List of Materials
- 1 large heating surface (large enough to hold two large beakers)
- 2 large glass beakers of water
- 2 thermometers
- Heat resistant gloves to handle hot beakers

### Procedure
1. Fill beakers with water so that one has twice as much water as the other.
2. Place each beaker on the heating surface and turn it on.
3. Indicate that the same amount of thermal energy is being added to each beaker of water.

Ask students what they think will happen to the temperature of each beaker of water. Will they be the same temperature? Will one heat up faster than the other?

4. Before the water boils, take the temperature of both beakers of water. The smaller beaker should be at a lower temperature than the larger beaker.

### Discussion

The following questions can help students understand what they observe.

1. If the thermal energy being added to this system is the same, why does one have a higher temperature than the other?
2. Describe the flow of energy in this system.
3. What is the difference between heat and temperature?
4. How are heat and temperature related?
5. Can these two beakers of water ever reach the same temperature? If so, under what conditions?
6. What would happen to the temperature of each beaker of water if you added the same amount of ice to each beaker? Would the ice melt faster in one beaker than in the other? (You may wish to demonstrate this as well)
7. What would happen to the temperature of the water if you had the same volume of water, but one container was made out of copper and the other was made of glass? Describe the energy flow in this system. (You may wish to perform this demonstration as well.)

### Hazards

Be careful when using the burners. Wear thermal gloves to handle hot beakers.

### Disposal

There are no special disposal considerations.

## The Conversion of Mechanical Work to Thermal Energy

### Objective

In this demonstration running a blender causes the temperature of the water inside to increase, showing an example of mechanical work converting to thermal energy.

### List of Materials

- Blender
- Water
- Thermometer

### Procedure

1. Fill a blender half way with room temperature water.
2. Take the temperature of the water.
3. Run the blender on the highest speed for a few minutes.
4. Take the temperature of the water.

### Discussion

The following questions can help students understand what they observe.

1. In this example, what is the mechanical work?
2. Why does the temperature of the water increase when the blender is turned on?
3. Describe the flow of energy in the system.

### Hazards

Be careful when operating the blender. Make sure the lid is on before starting the blender. Do not operate the blender when the thermometer is in it.

### Disposal

There are no special disposal considerations.

## Endothermic Reaction

### Objective

This demonstration provides a safe example of an endothermic reaction of citric acid solution and sodium bicarbonate:

$$H_3C_6H_5O_7(aq) + 3\ NaHCO_3(s) \rightarrow 3\ CO_2(g) + 3\ H_2O(l) + Na_3C_6H_5O_7(aq)$$

### List of Materials

- 8 g sodium bicarbonate
- 25 mL of 1 M citric acid
- 400 mL beaker
- Thermometer

### Procedure

1. Add 8 g of sodium bicarbonate to 25 mL of 1 M citric acid in a 400-mL beaker.
2. Add the solid slowly to avoid bubbling over.
3. Monitor the temperature with a thermometer, and/or have students feel the bottom of the beaker to see that it is cold.

### Discussion

1. Is this an example of an endothermic or exothermic reaction?
2. Describe the flow of energy in this system.
3. What happens in an endothermic reaction?
4. What happens in an exothermic reaction?

### Hazards

It is good lab practice to review a chemical's Material Safety Data Sheet (MSDS) before working with any chemical. Follow instructions on the MSDS and encourage students to review them. Wear proper protective gear at all times: chemical splash goggles, chemical-resistant apron, lab coat and gloves. Citric acid: Severe eye irritant. Sodium bicarbonate: Slightly toxic by ingestion. LD50 4220 mg/kg.

### Disposal

Check local regulations for proper disposal of the chemicals.

## Going Deeper (In-Class Discussion or Reflection)

Instructors should allow up to 30 minutes for discussion at the end of the session, or students can use the time to reflect on one or more of these questions in journals.

1. How many different forms of energy can you find in the classroom right now?
2. The lights in the classroom release energy as light. Can you trace all of the energy conversions from the local power plant to the light? Can you trace all of the energy conversions from the power plant back to the sun?
3. If you charge an electric car using electricity from a coal-powered plant, is that better for the environment than a regular gas-powered car? What information would you need to answer this question?
4. How many real-life examples of exothermic and endothermic reactions can you think of?
5. If a calorimeter is not insulated well, what effect would that have on your calculated value of $\Delta H$? (Assume the water in the calorimeter starts at room temperature.)
6. Do you think that drinking ice water would be an effective way to lose weight because your body would expend energy to warm it up inside your body? Would it be possible to counteract the calories in a scoop of ice cream if the ice cream were made cold enough? What information would you need to answer these questions?
7. Suppose you put two pots of water on the stove. One is a large pot for pasta; the other

is a small pot for a cup of coffee or tea. Which one boils faster (assume that the burners are on the same setting)? What is the temperature of the boiling pasta water? What is the temperature of the boiling coffee/tea water?

8. Write the chemical equation for the combustion of octane (the principal hydrocarbon in gasoline). What will happen to the pressure if this reaction is performed in a closed container? What if the container is closed, but can change its volume?

9. Investigate how instant hot- and cold-packs (such as those available in first-aid kits) work.

## Before the Next Unit

Students should read the Unit 7 text if they haven't already done so. They may be assigned one or more reading assignments from the list below, or if you choose to have them use the course materials outside of class, they can watch the Unit 8 video and/or read the Unit 8 text as an assignment before the next session.

## References and Additional Resources

Lightman, Alan P. "The Conservation of Energy." *Great ideas in physics: the conservation of energy, the second law of thermodynamics, the theory of relativity, and quantum mechanics*. 3rd ed. New York: McGraw-Hill, 2000.

## For Professional Development

In addition to watching the videos, reading the text, and going through the activities listed in the course guide, participants taking this course for professional development should read the following papers and answer the corresponding reflection questions. Participants should then complete the accompanying professional development assignments.

### Further Reading and Reflection Questions

Barker, Vanessa. "Students' Ideas about Thermodynamics." *Beyond Appearances: Students' Misconceptions about Basic Chemical Ideas*. A report prepared for the Royal Society of Chemistry: 58-63. Accessed July 8, 2013. http://modeling.asu.edu/modeling/KindVanessaBarkerchem.pdf.

1. Were you surprised by these common misconceptions about thermodynamics? Have you previously held any of these views yourself or encountered students that held these view? How have you addressed them in the past?

2. Do you agree with the author's suggestions for progress? Do you foresee any difficulties in implementing any of her suggestions? Can you think of any other ways to improve the teaching of thermodynamic principles?

Le Merechal, Jean-Francois and Rania El Bilani. "Teaching and Learning Chemical Thermodynamics in School." *International Journal of Thermodynamics* 11:(2) 2008. 91-99. Accessed July 31, 2013. http://ijoticat.com/article/view/1034000216

1. Have you encountered students who had good problem solving skills but little conceptual understanding? What are some ways you could assess students to determine if they have both good problem solving skills and conceptual understanding?

2. How could you incorporate or adapt the "energy chain" model in your teaching in a way that you think could increase conceptual understanding of energy and thermodynamics?

3. Has this study influenced how you think about teaching energy and thermodynamics? Are there any other ways you could increase conceptual understanding of these topics? Explain.

### Professional Development Assignments

1. After reading the papers above and reflecting on the questions presented develop a lesson plan designed to teach material presented in this unit.

2. Using a group activity or classroom demonstration presented in this course guide, show how you would implement it into your classroom. Where would it fit into your curriculum or standards? Would you change the demonstration or activity in any way? How would you assess student learning?