Activity 3  Cooking Fuels

GOALS
In this activity you will:
• Make quantitative observations about different fuels.
• Understand where the energy comes from when a fuel is burned.
• Understand the relationship between heat and temperature change.
• Determine the amount of heat released from the combustion of various fuels.

What Do You Think?
Cooks use different fuels for different reasons.
• Which fuel do you think will cook foods the fastest?
• Where does the energy in the fire come from?
Record your ideas about these questions in your Active Chemistry log. Be prepared to discuss your responses with your group and the class.

Investigate
In this investigation, you will compare the energy content of various fuels. Fuels that you may be testing may include methanol, ethanol, kerosene, lamp oil, butanol, paraffin (candle wax), and jellied petroleum.

Part A: Designing Your Own Investigation
1. A common characteristic of most fuels is that they are compounds made of carbon and hydrogen (hydrocarbons) or carbon, hydrogen, and oxygen (alcohols). The amount of energy that a fuel releases has to do with how much and how completely the fuel is burned. But where do you think this energy comes from? You discussed this with your class in the What Do You Think? section.
2. To determine the amount of energy released from the combustion of a fuel, you will heat some water. By measuring the amount of fuel used and the change in temperature of the water, you will be able to make conclusions about the energy of each fuel.

Consider the fuels that your teacher will be asking you to test. Predict which fuel will have the best heating ability.

a) Record your prediction in your Active Chemistry log.

3. Using this strategy, design an experiment that can be conducted to determine the amount of energy released from the combustion of a fuel.

Include in your design:
- What apparatus and supplies you will need.
- What measurements you will make.
- What data you will record.
- How you will analyze the data.
- How you will draw a conclusion based on the data.

4. If your teacher has the materials in your design and approves your design, then you may proceed to carry out your investigation. Your teacher may decide to give you credit for your design and ask you to use the steps outlined in Part B.

Part B: Energy Content of Various Fuels

If your teacher does not have the supplies you need to do the investigation you designed, follow this procedure. In this procedure, you will use a soda can as a container for the water. While the soda can will not provide completely accurate results (some heat energy will be lost to the surroundings), it will give you some idea of the energy content of fuels. A cook will want to choose the right fuel for the job.

1. Your teacher will provide you with a soda can that has two holes at the top. Set up a ringstand and ring. Place a glass stirring rod through the holes in the can. Set the can and stirring rod on the ring, as shown in the diagram.

2. Carefully, measure a specific amount of water (somewhere between 100 mL and 150 mL) in a graduated cylinder. Pour it into the can, being careful not to spill any.

   a) Record the amount of water you use in your log.

   b) Record the starting temperature of the water.
3. Obtain an alcohol burner. Measure the mass of the burner and fuel.
   a) Record the type of fuel in the burner.
   b) Record the total mass of the burner and fuel as “mass before.”

4. Set the burner under the can of water. Adjust the height of the ring so the flame will hit the bottom of the can. This should be about 2–3 cm from the wick.

5. Light the burner and readjust the height of the can, if necessary. Let the fuel burn for four minutes. Blow out the flame. Measure the temperature of the water using a thermometer or temperature probe.
   a) Record the color of the flame.
   b) Record the highest temperature of the water.

6. Let the burner cool for a while and then measure its mass again.
   a) Record this as “mass after.”

7. Now that you have learned the techniques, you can complete the activity. Design an experiment, using the apparatus available, to compare three fuels and their heating ability.

   Each group will now work with one fuel and then the class will share data.

   You may want to prepare a table similar to the one below in your Active Chemistry log. Record your measurements for each fuel in the table.

   Note: If you use a candle, record the mass of the candle and foil-covered square base. For the jellied petroleum, record the mass of the open can.

Part C: Analyzing Your Data

1. The thermal energy required to change the temperature of a given amount of water by 1°C is the specific heat capacity. The specific heat capacity of water is one calorie per one gram per one degree Celsius. It is written as:

   \[
   \frac{1 \text{ cal}}{\text{g} \cdot \degree \text{C}} \quad \text{or} \quad 1 \text{ cal g}^{-1} \degree \text{C}^{-1}
   \]

   In other words, it takes one calorie of heat energy to raise the temperature of one gram of water by one degree Celsius.

   If you know the amount of water and the change in temperature through which it goes, the specific heat capacity allows you to determine how much heat energy the water absorbs.

   The equation to determine heat energy is:

   \[
   \text{Heat} = (\text{mass of water}) \times (\text{specific heat capacity of water}) \times (\text{temperature change of water})
   \]

   Using mathematical symbols, this equation can be written as:

   \[
   Q = mc\Delta T
   \]

   Where
   - \( Q \) = the change in heat
   - \( m \) = the mass of water (in grams)
   - \( c \) = the specific heat capacity of water
   - \( \Delta T \) = the change in temperature \( (T_{\text{final}} - T_{\text{initial}}) \)
2. For each fuel you tested, you will need to complete the calculations shown below. Prepare a table, similar to the one shown above, to display your results in your Active Chemistry log.

<table>
<thead>
<tr>
<th>Fuel</th>
<th>Mass of water (g)</th>
<th>Temp. change of water (°C)</th>
<th>Heat absorbed by water (cal)</th>
<th>Mass of fuel burned (g)</th>
<th>Heat per gram of fuel burned (cal/g)</th>
</tr>
</thead>
</table>

a) Mass of water: Since the density of water is 1 g/mL, you simply change the volume of water to mass by changing the units of milliliters to grams. Mass of 150 mL of water = 1 g/mL × 150 mL = 150 g

b) Temperature change of water: Subtract to find the temperature change.

c) Heat absorbed by water (lost by the fuel): Use the equation \( Q = mc\Delta T \) to find heat.

d) Mass of fuel burned: Subtract to find how much of the fuel you burned.

e) Heat per gram of fuel burned: Divide the heat absorbed by the mass of fuel burned.

3. What conclusions can you draw from your investigation? Record the following in your Active Chemistry log:

a) Compare your results to those of other groups. Which category is best for comparison purposes? Why?

b) Which fuel did you find produced the most heat per gram? The least?

c) Rank the fuels and compare the ranking to your prediction.

d) Look at the flame colors and the amount of heat produced per gram of fuel. Is there any correlation between them? If so, what is it?

**ChemTalk**

**THERMOCHEMISTRY**

**Specific Heat Capacity**

When a fuel is burned, the energy released can be put to good use. In this activity, you used the energy to heat water and raise its temperature. As you learned previously, temperature and heat are related, but they do not measure the same thing. Temperature measures the kinetic energy or motion of molecules in a system. When you added heat energy to the water in the soda can, the water molecules began moving faster. As a result, the temperature went up. Heat is a measure of the total energy in a system. It is measured with a unit called a calorie or a joule.
A calorie is the amount of heat needed to raise the temperature of one gram of water by one degree Celsius. You are probably familiar with the term Calorie as it relates to foods. A food Calorie (note the capital C) is equal to 1000 cal (calories) or 1 kcal (kilocalorie). It is also a measure of the energy available from the food. The joule is also used as the unit of energy.

In chemistry, as in all science, the international community has agreed to use joules as a unit of energy. However, in this activity, you used calories in your calculations. This is because in the USA, the energy content of foods is still given in kilocalories (kcal or C). A food that is 100 C (Calories) has an energy content of 100,000 cal (calories). Since 1 cal is equal to 4.184 J (joules), 100 C food has an energy content of 418,400 J.

To calculate the heat per gram of fuel consumed you needed to use the energy required to raise the temperature of 1 g of water by 1ºC. This is the specific heat capacity of water. The specific heat capacity of water is one calorie per one gram per one degree Celsius. It is written as:

$$\frac{1 \text{ cal}}{\text{g} \cdot \text{°C}} \text{ or } 1 \text{ cal g}^{-1} \text{ °C}^{-1}$$

In other words, it takes one calorie of heat to raise the temperature of one gram of water by one degree Celsius.

In the activity, you measured the amount of water and the change in temperature of the water. Then, using specific heat capacity, you determined how much heat energy the water absorbed using the following equation:

Heat = (mass of water) \times (specific heat capacity of water) \times (temperature change of water)

Using mathematical symbols, this equation can be written as:

$$Q = mc\Delta T$$

Where Q = the change in heat  
    m = the mass of water (in grams)  
    c = the specific heat capacity of water  
    \Delta T = the change in temperature  
    (T_{\text{final}} - T_{\text{initial}})

**Terms Used in Thermochemistry**

**Thermochemistry** is the study of heat effects that accompany chemical reactions. When you measure the quantities of heat gained or lost in chemical reactions, you are investigating the thermochemistry of
those reactions. In studying thermochemistry, you use the term **system** to describe the reactants, solvent, and products of a reaction. You use the word **surroundings** to indicate everything outside of the chemical reaction: the can, the room, building, and so on, out into the universe.

When energy is released from the system to the surroundings, it is called an **exothermic reaction**. An **endothermic reaction** is one where energy is absorbed by the system from the surroundings. In the activity, energy was released as the fuel burned in an exothermic reaction.

When energy is released in an exothermic reaction, where does the energy come from? To understand exothermic reactions, you have to look at the same events from the viewpoint of the molecules and atoms. In *Active Chemistry*, we refer to this as the nanoscopic scale. That is because one nanometer is the size of some molecules. The energy in exothermic and endothermic reactions relates to the chemical bonds of the reactants. It takes energy to break a chemical bond. Consequently, when a chemical bond is formed, energy is released to the surroundings. Energy **in** to break a bond. Energy **out** to create a bond.

When your fuel was burned in the presence of oxygen, it underwent a combustion reaction. Combustion of any hydrocarbon is an exothermic process, which produces carbon dioxide and water. The combustion of candle wax you observed earlier was the same exothermic reaction.

Methane is the chief component of natural gas used in kitchens for cooking. Methane undergoes combustion according to the following balanced equation:

\[
\text{CH}_4(\text{g}) + 2\text{O}_2(\text{g}) \rightarrow \text{CO}_2(\text{g}) + 2\text{H}_2\text{O}(\text{g}) + \text{energy}
\]

You see that + energy has been added to this reaction. It indicates that this is an exothermic process. The energy that is released from this reaction comes from the stored energy in the molecular bonds. The energy to break the bonds of the reactants is less than the energy released to create the bonds in the products. The extra energy is released to the surroundings in this exothermic process.

All molecules have potential energy resulting from the bonds that hold the atoms together. In this reaction, if 1 mol of CH₄ were used, the potential energy of the reactants is equal to 3450 kJ. The total potential energy of the products is equal to 2642 kJ. The heat energy that is released from this reaction is the difference between the potential energies of the reactants and products:

\[
2642 \text{ kJ} - 3450 \text{ kJ} = -808 \text{ kJ}
\]

These 808 kJ of heat energy are released in this exothermic reaction.
One way to graphically represent the energy changes of a reaction is with an energy diagram. Typical energy diagrams for both an exothermic reaction and an endothermic reaction are shown below.

In the exothermic reaction, the potential energy (P.E.) of the products is lower than that for the reactants. The energy difference (in the form of heat energy, or Q) is negative because it leaves the system and is given off to the surroundings.

In the endothermic reaction, the P.E. of the products is higher than that of the reactants. Energy must be supplied to the reaction to make it occur. The energy difference is positive because it is provided to the system from the surroundings.

When weaker bonds (less stable, higher energy) of the reactants are broken and stronger bonds (more stable, lower in energy) in the products are formed, energy is released to the surroundings in an exothermic reaction. This is because the potential energy of the system (the substances involved in the reaction) has been lowered.

A chemical reaction where stronger bonds of the reactants are broken and weaker bonds are formed in the products is an endothermic reaction that requires energy in order to take place.

In both cases, there is an energy that must be supplied in order for the reactants to have the proper collision energy for the reaction to occur. This is called the activation energy ($E_a$). When you light the propane burner of a camp stove, a spark is needed to start the reaction, supplying the activation energy. Even though the combustion of propane gas is an exothermic reaction, the spark must be supplied before the reaction will proceed. An exothermic reaction can be sustained by the heat liberated by the reaction.
Alcohol Fuels

Alcohol fuels contain the hydroxyl group (–OH). If you replace one of the hydrogens of methane with a hydroxyl group, you will now have a compound that is called an alcohol. In this case, the alcohol is called methanol or methyl alcohol. Methane is CH₄ and methanol is CH₃OH.

The combustion of methanol is shown as:

\[ 2\text{CH}_3\text{OH}(l) + 3\text{O}_2(g) \rightarrow 2\text{CO}_2(g) + 4\text{H}_2\text{O}(g) \]

As in the combustion of methane, energy is released as the strong carbon-oxygen bonds are formed in CO₂. As a general rule, the fewer carbon-oxygen bonds in the reactants, the more energy released. Larger hydrocarbon compounds with no oxygen present should release more heat on burning than smaller, partially oxygenated hydrocarbons.

Here are the formulas of some possible fuels:

<table>
<thead>
<tr>
<th>Name</th>
<th>Molecular formula</th>
<th>Condensed structural formula</th>
</tr>
</thead>
<tbody>
<tr>
<td>methanol</td>
<td>CH₄O</td>
<td>CH₃OH</td>
</tr>
<tr>
<td>ethanol</td>
<td>C₂H₆O</td>
<td>CH₃CH₂OH</td>
</tr>
<tr>
<td>1-propanol</td>
<td>C₃H₈O</td>
<td>CH₃CH₂CH₂OH</td>
</tr>
<tr>
<td>2-propanol (isopropyl alcohol)</td>
<td>C₃H₈O</td>
<td>CH₃CH(OH)CH₃</td>
</tr>
<tr>
<td>n-butanol</td>
<td>C₄H₁₀O</td>
<td>CH₃CH₂CH₂CH₂OH</td>
</tr>
<tr>
<td>kerosene</td>
<td>A mixture averaging C₁₀H₂₂</td>
<td></td>
</tr>
<tr>
<td>lamp oil</td>
<td>A mixture of paraffin and kerosene</td>
<td></td>
</tr>
<tr>
<td>candle wax (paraffin)</td>
<td>A mixture of C₂₀H₄₂ and larger</td>
<td></td>
</tr>
</tbody>
</table>

Balanced equations show relationships in molar quantities. A **mole** is a specific number of particles (6.022 × 10²³ particles). The combustion of methanol shown in equation above can also be described in words as:

Two moles of methanol react with three moles of oxygen to produce two moles of carbon dioxide and four moles of water, with a certain amount of energy released.

In the investigation, you determined the amount of heat per gram of fuel. In the equation just described, the heat released would be per mole(s) of fuel burned.
**What Do You Think Now?**

At the beginning of the activity you were asked:

- Which fuel do you think will cook foods the fastest?
- Where does the energy in the fire come from?

Now that you have completed this activity, where does the energy released in the exothermic process of burning fuels come from?

### Chem Essential Questions

#### What does it mean?

Chemistry explains a macroscopic phenomenon (what you observe) with a description of what happens at the nanoscopic level (atoms and molecules) using symbolic structures as a way to communicate. Complete the chart below in your Active Chemistry log.

<table>
<thead>
<tr>
<th>MACRO</th>
<th>NANO</th>
<th>SYMBOLIC</th>
</tr>
</thead>
<tbody>
<tr>
<td>What did you observe in this activity that led you to believe that the combustion of a fuel is an exothermic reaction?</td>
<td>When a fuel burns, compare the bonds of the molecules on the product and reactant sides.</td>
<td>Use an energy diagram as a symbolic structure to describe what happens during an exothermic reaction.</td>
</tr>
</tbody>
</table>

#### How do you know?

Take a look at the class data for the energy content of the different fuels used. In the *Chem Talk* section, you read the hypothesis that partially oxygenated fuels (those with some carbon-oxygen bonds present) will release less heat than similar hydrocarbons without oxygen. Does the class data support this hypothesis?

#### Why do you believe?

Your everyday experiences make it easy to accept that when a fuel is burned, it is an exothermic process. What examples of fires would you use to explain to someone that fires require fuel?

#### Why should you care?

Consider some of the fuels that are sources of heat for cooking. You cannot always use fuels with the highest heat output per gram. Consider why not in terms of cost, convenience, and safety.

### Reflecting on the Activity and the Challenge

Think about the advantages of gas and electric stoves in terms of the control of the amount of heat, the response to the change in heat, and safety. You may want to describe these as well as the combustion reaction of the gas stove in your show. Making the discussion entertaining is one of the tougher parts of your cooking show because it requires creativity.
Activity 3 Cooking Fuels

Inquiring Further
Cost of fuels and energy content
Investigate whether the cost of fuels is related to the energy content of the fuels.

1. How would your data be different if you used only 50 mL of water instead of 150 mL in this experiment?

2. Write balanced chemical equations for the fuels that you tested.

3. a) Explain the difference between an endothermic and an exothermic reaction.
   b) List some technologies that use endothermic and some that use exothermic reactions.

4. Convert the following heat quantities recalling 1 cal = 4.184 J:
   a) 350 cal to joules  
   b) 515 J to calories  
   c) 1.6 kcal to joules

5. Suppose you have two containers of water. One contains 150 mL at 80ºC and the other has 75 mL at 60ºC. Is the heat content of both containers equivalent? If not, which has the lesser heat content?

6. How much heat is required to change the temperature of 150 g of water by 20ºC?

7. How much heat will be given off when 1500 g of water cools down by 20ºC?

8. Which statement describes the characteristics of an endothermic reaction?
   a) The sign of $Q$ is negative, and the products have less potential energy than the reactants.
   b) The sign of $Q$ is positive, and the products have less potential energy than the reactants.
   c) The sign of $Q$ is negative, and the products have more potential energy than the reactants.
   d) The sign of $Q$ is positive, and the products have more potential energy than the reactants.

9. What is the total number of joules of heat energy absorbed by 15 g of water when it is heated from 30ºC to 40ºC?
   a) 10  
   b) 63  
   c) 150  
   d) 630

10. Whenever bonds between atoms are broken and rearranged:
    a) Energy is involved.  
    c) Melting has occurred.
    b) An exothermic reaction has taken place.  
    d) Energy is supplied to the system.

11. **Preparing for the Chapter Challenge**
    In preparing for your Chapter Challenge, you may want to explain what source of heat is used for cooking and why it is used. Make a list of possible sources in your Active Chemistry log.

Inquiring Further
Cost of fuels and energy content
Investigate whether the cost of fuels is related to the energy content of the fuels.