

Experiment 30A

FV 11/13/06

HOW MUCH ENERGY DO YOU HAVE?

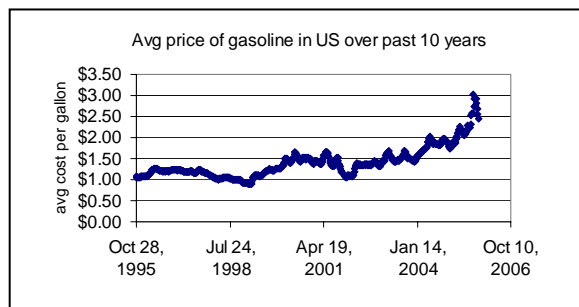
MATERIALS: 12-oz. aluminum beverage can with top cut out and holes on side, thermometer, 100 mL graduated cylinder, 800 mL beaker, long-stem lighter, three fuel burners (filled with ethanol, n-octane, or 2-pentanol), cashew nuts, steel wool, glass rod, ring stand, rubber cork, paper clip, room-temperature water.

PURPOSE: The purpose of this experiment is to determine and compare the energy content values of various materials.

LEARNING OBJECTIVES: By the end of this experiment, the student should be able to demonstrate the following proficiencies:

1. Construct and use an aluminum can calorimeter.
2. Calculate the efficiency of heat transfer to the calorimeter.
3. Calculate the energy content per gram for two fuels and a food item.
4. Compare the energy content of an oxygenated and non-oxygenated fuel.
5. Determine the Calorie content of a food item.

PRE-LAB: The pre-lab for this experiment is on page 7.
Complete the pre-lab **prior** to coming to lab.



DISCUSSION:

Fuels

Combustion reactions are utilized in converting chemical energy to more useful forms of energy. This chemical energy is used both by mechanical systems, such as cars, ships and planes, and by biological systems. The fuels used by mechanical and biological systems are, however, quite different.

Mechanical systems can use a variety of fuels. Common examples are coal, petroleum and natural gas. These three are all known as fossil fuels and contain primarily hydrocarbons (compounds containing carbon and hydrogen). They were formed over millions of years from the decomposition of plants and animals. Coal is the most abundant of the fossil fuels. It is a complex mixture of many carbon-based compounds and is used in this crude form. Petroleum is also a complex mixture of many organic compounds. Unlike coal, petroleum is refined prior to use. In the refining process, the various components are separated into a variety of commercially useful products. Gasoline, which Americans consume at a rate of nearly 130 billion gallons/year and whose cost has been steadily rising in recent months (see graph above), is the fraction of petroleum which boils between 30°C and 200°C. Kerosene contains compounds with boiling points between 175°C and 300°C. These broad fractions can be further refined for specific uses. Diesel fuel and home heating oil are both prepared by further distillation of kerosene.

Jet fuels are quite similar to kerosene in chemical composition. JP-4 (jet propellant-4), mainly used by the Air Force, is similar to Jet B, a commercial propellant of kerosene and gasoline, except it contains special corrosion inhibitors and anti-icing agents. JP-5, used by Navy aircraft on carriers, is similar to JP-4 but has a higher flash point so that it is safer to use aboard ship. The Air Force has recently replaced JP-4 with JP-8 which is safer (higher flash point and less carcinogenic), and burns more efficiently.

The combustion of the hydrocarbon fuels results in the formation of carbon dioxide and water with a release of energy (exothermic reaction). Two specific examples of such reactions are shown below.

Natural gas, methane:



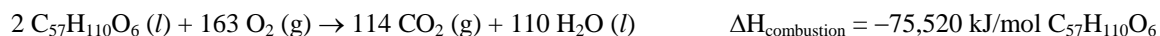
Propane:



For biological systems, fuel is more commonly known as food. The energy needed to sustain life generally comes from two classes of foods, carbohydrates and fats. One of the simplest carbohydrates is glucose, $C_6H_{12}O_6$. Most complex carbohydrates are broken down into glucose during digestion. Glucose is converted to carbon dioxide and water in a number of biochemical steps, with the overall process being the combustion of glucose.



Fats are also converted in many steps to carbon dioxide and water in the body. The overall reaction for a major component of body fat, glycerol tristearate, is shown in the following reaction:



Comparing Fuels

Comparing fuels can be difficult. While the enthalpy of combustion values, $\Delta H_{\text{combustion}}$, are given for each of the examples above, these values are difficult to compare since they are strongly dependent on the number of moles of carbon dioxide and water formed from the various fuels. For example, is glucose with a $\Delta H_{\text{combustion}}$ of -2816 kJ/mol really a better fuel in terms of energy output than methane, with a $\Delta H_{\text{combustion}}$ of -890.3 kJ/mol ?

A more practical way to compare fuels is to use the *energy content* or *fuel content* of the fuel. Energy content is defined as the amount of energy released per gram of fuel. For simplicity in comparing values of released energy in this experiment, all energy content or fuel content values will be positive values. For the examples above, the energy content of methane is 55.5 kJ/g while that of glucose is 15.6 kJ/g of released energy.

The energy content of a fuel can be determined using calorimetry. Determination of accurate values would require use of a constant volume calorimeter known as a bomb calorimeter. To obtain approximate values, this experiment will utilize a constant pressure calorimeter composed of an aluminum beverage can filled with 100 mL of water. The combustion reactions cannot be done directly in the calorimeter. Instead, a burner containing the fuel will be placed underneath the calorimeter. The heat produced by lighting the burner will partially be transferred to the water in the calorimeter. You conducted a calorimeter experiment earlier in the semester (Expt. 12F – Combustion of Magnesium) which relied on similar principles.

Calibration of the Calorimeter

The calorimeter must first be calibrated. This will be accomplished by burning ethanol, an oxygenated fuel with a known $\Delta H_{\text{combustion}}$. The amount of energy released in the form of heat is related to $\Delta H_{\text{combustion}}$, where n is the number of moles of ethanol consumed.

$$q_{\text{combustion}} = n \cdot \Delta H_{\text{combustion}} \quad (1)$$

The energy transferred to the water in the calorimeter can be determined using the following relationship:

$$q_{\text{water}} = m \cdot s \cdot \Delta T \quad (2)$$

where m is the mass of the water in the calorimeter, s is the specific heat of water ($4.184 \text{ J/}^\circ\text{C}\cdot\text{g}$), and ΔT is the change in temperature of the water. The heat capacity of the can is assumed to be negligible.

In this experiment, not all of the energy is transferred to the water, so the efficiency of the energy transfer must be determined. The efficiency factor can be calculated using the following definition.

$$\text{Efficiency} = \frac{-q_{\text{water}}}{q_{\text{combustion}}} \quad (3)$$

Assuming the experiment is set up and performed in the same way, it is possible to determine q for other reactions using this efficiency factor and the equation:

$$q_{\text{combustion}} = \frac{-q_{\text{water}}}{\text{Efficiency}} \quad (4)$$

Energy Content of Fuels

Once the efficiency of energy transfer is known for a particular calorimeter, the calorimeter can be used to determine the energy content of other fuels, such as n-octane and 2-pentanol in this experiment. Burners containing these fuels will be used to heat the water in the calorimeter using the same procedure followed with the ethanol. The energy content of the fuel is calculated in two steps. The first step is determining the $q_{\text{combustion}}$, using equation (4). The energy content is then calculated by dividing $q_{\text{combustion}}$ by the mass of the fuel consumed. Note that we will be defining energy contents as positive values.

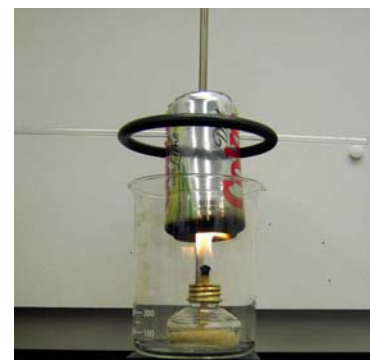
$$\text{Energy content} = \left| \frac{q_{\text{combustion}}}{\text{mass of fuel}} \right| \quad (5)$$

PROCEDURE: (work in pairs)

Safety: The fuels used in this experiment are very flammable and care must be taken to avoid spilling them. The aluminum cans may also have sharp edges. Use the same aluminum can for the entire experiment.

Part A. Calibration of the Calorimeter

1. Light the ethanol burner and check the characteristics of the flame. A proper flame should be narrow, an inch or so in height and as free of soot as possible. If the flame is not the proper height, extinguish the flame by capping the burner with its top. Adjust the height of the wick and retest the flame. Extinguish the flame.
2. Remove the cap on the ethanol burner and place the burner in an 800 mL beaker.
3. Gently push a glass rod through the pre-drilled holes in a clean, aluminum can. Set up an iron ring on a ring stand to suspend the can assembly as in the figure. Adjust the height such that the bottom of the can is approximately an inch above the burner.
4. Light the burner using the long-stem lighter, tipping the can slightly to the side. When lit, the flame should be in the center of the can with much of the flame touching the can. In theory, you want all of the energy from the flame (combustion reaction) to be transferred to can (which will heat the water inside later). In practice, the energy transfer is not 100% efficient with losses due to inconsistent flames, buildup of soot, air currents, etc. Adjust the position of the can as necessary. Remove the can by lifting it off the ring (don't change the height of the ring), then extinguish the flame by carefully replacing the cap.
5. Weigh the ethanol burner and cap on a top-loading balance and record the mass. Place it in the 800 mL beaker. Remove the cap.
6. Remove the rod from the can. Using a paper towel first (then steel wool if necessary), clean your aluminum can if it is sooty. Record the mass of the empty can on a top-loading balance.
7. With a graduated cylinder, add 100 mL of room-temperature water to the aluminum can. Record the mass of the can containing the water. Obtain the initial temperature of the water. Replace the rod on the can without spilling the water.
8. Light the ethanol burner and quickly place the can of water on the iron ring over the flame. Center the can over the flame. Try to minimize air currents which may lead to irregular or wandering flames. The goal is to have a narrow, consistent flame which continuously heats the water in the can.
9. Stir the water with a thermometer measuring the temperature in the middle of the water, not with the thermometer touching the bottom of the can. When the temperature has risen to about 40°C above the initial water temperature, remove the can and quickly cap the burner to extinguish the flame. Keep stirring the water and record the highest temperature the water reaches.



- Record the mass of the ethanol burner and cap to determine the amount of fuel burned.
- Pour the water out of the can and inspect the bottom of the can. If it is sooty, clean the bottom of the can with a paper towel or if necessary, use steel wool.
- Repeat the measurement. If the temperature change of the water divided by the mass change of the burner does not agree with the first measurement within 10%, repeat the measurement a third time. Note, consistent flame and can positions lead to more consistent results.
- Using the $\Delta H_{\text{combustion}}$ value for ethanol (-1366.83 kJ/mol), calculate the efficiency factor for your calorimeter. Report your average value to your instructor.

Part B. Determination of the Energy Content of an Assigned Fuel

- From your instructor, obtain your assigned fuel, n-octane or 2-pentanol. Following similar procedures as above, determine the average energy content of your assigned fuel (in kJ/g).
- Record your procedure and data clearly.
- Have your instructor check your results.
- From a classmate, obtain the energy content value for the other fuel you did not test.

Part C. Determination of the Energy Content of a Cashew

- From your instructor, obtain cashew samples (do not eat any food in lab). Following similar procedures as above, determine the average energy content of a cashew (in kJ/g). Using a rubber stopper and a paperclip, devise a way to hold the cashews as they are burned. Remember the cashew flame needs to be consistent and remain under the can. Burn 2 cashews for each trial, and have your instructor check your setup before proceeding. Be careful, the cashews remain hot even after they have been extinguished.
- Record your procedure and data clearly.

Clean-up:

- Using steel wool, remove any soot from the aluminum can. It can be reused by another student.
- Recap all burners and return them to their proper locations.
- Make sure the cashews are completely extinguished (douse with water) before disposing in the trash.
- Clean up any spilled material, especially fuels, from the lab benches.

Data Analysis:

- For each of the fuels, ethanol ($\text{CH}_3\text{CH}_2\text{OH}$), n-octane (C_8H_{18}), and 2-pentanol ($\text{CH}_3\text{CHOHCH}_2\text{CH}_2\text{CH}_3$), calculate the mass % of oxygen in the fuel.
- Construct a table that compares the % O and energy content for each of the 3 fuels. What trend do you observe? Do fuels that are highly oxygenated produce more or less energy when burned? Explain your answer.
- Convert your average energy content (in kJ/g) for cashew nuts to units of Calories/g. Note that this is the dietary Calorie (capital "C") where 1 Calorie = 1 kcal. Based on the nutritional information for a cashew, determine the reported Cal/g value and compare this to your experimental result (calculate % difference).
- A person of average weight uses about 100 Cal/mi when jogging. Based on the nutritional information, how many servings of cashews would provide the energy content requirements for jogging 2 miles? How many cashews does this correspond to (use your average mass of 2 cashews in this calculation)?

Nutritional info for cashews.

Nutrition Facts	
Serving Size 1 oz (28g)	
Amount Per Serving	
Calories 156	Calories from Fat 104
% Daily Value*	
Total Fat 12g	19%
Saturated Fat 2g	11%
Cholesterol 0mg	0%
Sodium 3mg	0%
Total Carbohydrate 9g	3%
Dietary Fiber 1g	4%
Sugars 2g	
Protein 5g	
Vitamin A 0%	Vitamin C 0%
Calcium 1%	Iron 10%
*Percent Daily Values are based on a 2,000 calorie diet. Your daily values may be higher or lower depending on your calorie needs.	
NutritionData.com	

4. Look up some information about fuel cells. What are the pros and cons of using fuel cell technology to power cars? Currently what is the biggest obstacle in utilizing fuel cells in cars? Include citations for your references.

5. A student followed the procedure of this experiment to determine the Calorie content of a peanut. Before burning, the peanut had a mass of 0.705 g, and after combustion, it had a mass of 0.035 g. The energy released during combustion caused a 15.5°C increase in the temperature of 250.0 mL of water in the calorimeter. Determine the mass of the peanut burned and its energy content.

Expt. 30A PRE-LAB

1. What do coal, petroleum, and natural gas all have in common? Circle **all** that apply.

- a) They all consist of primarily hydrocarbons.
- b) They all can participate in combustion reactions.
- c) They are all fossil fuels.
- d) They will be tested in this experiment.
- e) They all have the same boiling point.
- f) They all release the same amount of energy per mole when burned in the presence of oxygen.

2. How is jet fuel on aircraft carriers different from aviation fuels used in other applications?

- a) It has a higher octane rating.
- b) It has a lower flash point.
- c) It has a higher flash point.
- d) It has a higher rate of combustion.

3. Three liquid fuels and one food item will be burned in this lab. Circle at least three of the correct materials.

- | | | | | |
|-------------|----------|---------------|----------|-----------|
| cashew | ethane | ethanol | gasoline | isooctane |
| marshmallow | methane | n-octane | peanut | pentane |
| 2-pentanol | soda pop | vegetable oil | | |

4. What type of experimental method will be applied to determine the energy content of the fuels in this experiment?

- a) calorimetry
- b) elemental analysis
- c) spectrophotometry
- d) titration
- e) viscosity measurements

5. To compare the different fuels, we will determine the “energy content” (also called “fuel content”) of the fuels we study. Define this quantity.

6a. The $\Delta H_{\text{combustion}}$ for ethanol is -1366.83 kJ/mol. If 1.00 g of ethanol ($\text{C}_2\text{H}_5\text{OH}$) is burned, how much energy (in kJ) is released?

b. If **all** of this energy in part a was transferred to 100.0 g of water, what is the expected change in temperature, ΔT , of the water? The specific heat of water is 4.184 J/g°C.