Experiment 30

FV 9/18/03

ENERGY CONTENT OF FUELS

MATERIALS: 12-oz. aluminum beverage can with cut out top, thermometer, 100 mL graduated cylinder, 400 mL beaker, 800 mL beaker, ice, matches, wood splints, plastic pipets, three fuel burners - filled with ethanol, kerosene and 1-pentanol, respectively, steel wool, three-finger clamp, ring stand.

PURPOSE: The purpose of this experiment is to determine the energy content of some oxygenated and non-oxygenated fuels.

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.
4. Compare the energy content of an oxygenated and non-oxygenated fuel.

DISCUSSION:

Fuels

Many reactions used to convert chemical energy to more useful forms of energy are combustion reactions. 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 system 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.[1]

Gasoline 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.

The combustion of the hydrocarbon fuels results in the formation of carbon dioxide and water. Two specific examples of such reactions are shown below.

Natural gas:

$$CH_4\,(g) + 2\,O_2\,(g) \rightarrow CO_2\,(g) + 2\,H_2O\,(l) \qquad \Delta H_{combustion} = -890.3\ \text{kJ/mol}\ CH_4$$

Octane (a component of gasoline):

$$2\,C_8H_{18}\,(l) + 25\,O_2\,(g) \rightarrow 16\,CO_2\,(g) + 18\,H_2O\,(l) \qquad \Delta H_{combustion} = -5470\ \text{kJ/mol}\ C_8H_{18}$$

[1]For more information about refining and the various fuels prepared from petroleum, see "Chemistry", Brown, LeMay & Bursten, pp. 172-173 and pp. 970-971. For Naval applications of fuels, see http://www.chemistry.usna.edu/navapps/Fuels/default.htm

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.

$$C_6H_{12}O_6\ (s) + 6\ O_2\ (g) \rightarrow 6\ CO_2\ (g) + 6\ H_2O\ (l) \qquad \Delta H_{combustion} = -2816\ kJ/mol\ C_6H_{12}O_6$$

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:

$$2\ C_{57}H_{110}O_6\ (l) + 163\ O_2\ (g) \rightarrow 114\ CO_2\ (g) + 110\ H_2O\ (l) \qquad \Delta H_{combustion} = -75{,}520\ kJ/mol$$

Comparing Fuels

Comparing fuels can be difficult. While the enthalpy of combustion values, $\Delta H_{combustion}$, 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 $\Delta H_{combustion}$ of –2816 kJ/mol really a better fuel than methane, with a $\Delta H_{combustion}$ of –890.3 kJ/mol?

A more practical way to compare fuels is to use the *energy content* of the fuel. Energy content is defined as the amount of energy released per gram of fuel. For the examples above, the energy content of methane is –55.5 kJ/g while that of glucose is –15.6 kJ/g. For pure substances, these values can be calculated by dividing the $\Delta H_{combustion}$ by the molar mass of the substance. The energy content values of commercial fuels are used as a comparison value because the fuels are mixtures and therefore have an unknown or indefinite molar mass.

The energy content of a fuel can be determined using calorimetry. Determination of exact 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 cold 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.

Calibration of the Calorimeter

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

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

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

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

where m is the mass of the water in the calorimeter, s is the specific heat of water, 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.

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

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:

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

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 two fuels, kerosene and 1-pentanol. Burners containing these two 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 $q_{combustion}$, using equation (4). The energy content is then calculated by dividing $q_{combustion}$ by the mass of the fuel consumed.

$$\text{Energy content} = \frac{q_{combustion}}{\text{mass of fuel}} \qquad (5)$$

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.

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 less 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. Once the wick is properly adjusted, extinguish the flame until you reach Step 5.

✎ Answer question #1 on page E30-9.

2. Weigh the ethanol burner and cap on a top-loading balance and record the mass.

3. Place the ethanol burner in a 800 mL beaker.

4. Using steel wool, clean your aluminum can if it is sooty. Mount the can vertically with a three-finger clamp on the ring stand so that it is about 2 inches above the tip of the flame when the burner is lit.

5. Prepare water that is about 10°C below room temperature by melting some ice in about 250 mL of water in a beaker. With a graduated cylinder, measure 100 mL of this water and pour it into the aluminum can. Be sure that there is no ice in the water. Record the temperature of the water.

✎ Answer question #2 on page E30-9.

6. Light the ethanol burner (that is in the 800 mL beaker) and quickly slide the beaker under the can of water, centering the flame under the can. Because the burner is in a beaker, you might have to use a wood splint to light the burner. The goal is to have a narrow, consistent flame which continuously heats the water in the can. (The burner minimize air currents which may lead to irregular or wandering flames.)

7. Stir the water with the thermometer, and when the temperature has risen to about 20°C above room temperature, remove the ethanol burner and cap it to extinguish the flame. Keep stirring the water and record the highest temperature the water reaches.

8. Record the mass of the ethanol burner and cap.

9. 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 steel wool. Readjust the wick if necessary.

10. 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.

Part B. Determination of the Energy Content of Kerosene

1. Adjust the wick of the kerosene burner using the same procedure as in Step 1 of Part A. Adjust the height of the can so that it is about 2 inches above the tip of the flame when the burner is lit.

2. Weigh and record the mass of the kerosene burner and cap.

3. Using the same procedure as in Step 4 of Part A, measure 100 mL of cold water and pour it into the aluminum can. Carry out the temperature measurements, heating and weighing as before. Record the initial and maximum temperatures of the water and the final mass of the kerosene burner and cap.

4. Repeat the measurement. Remember to keep the wick trimmed and to scour the bottom of the can to remove any soot after each 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.

Part C. Determination of the Energy Content of 1-Pentanol

1. Adjust the wick of the pentanol burner using the same procedure as in Step1 of Part A. Adjust the height of the can so that it is about 2 inches above the tip of the flame when the burner is lit.

2. Weigh and record the mass of the pentanol burner and cap.

3. Using the same procedure as in Step 4 of Part A, measure 100 mL of cold water and pour it into the aluminum can. Carry out the temperature measurements, heating and weighing as before. Record the initial and maximum temperatures of the water and the final mass of the vegetable oil burner and cap.

4. Repeat the measurement. Remember to keep the wick trimmed and to scour the bottom of the can to remove the soot after each 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.

Clean-up:

1. Using steel wool, remove any soot from the aluminum can.

2. Recap all burners and return them to their proper locations.

3. Return all other items to their proper locations.

Name _____ Section _____

Partner _____ Date _____

DATA SECTION
Experiment 30

Part A. Calibration of the Calorimeter

	Trial 1	Trial 2	Trial 3
Initial mass of burner and cap (g)			
Final mass of burner and cap (g)			
Volume of water (mL)			
Initial temperature of water (°C)			
Maximum temperature of water (°C)			

Part B. Determination of the Energy Content of Kerosene

	Trial 1	Trial 2	Trial 3
Initial mass of burner and cap (g)			
Final mass of burner and cap (g)			
Volume of water (mL)			
Initial temperature of water (°C)			
Maximum temperature of water (°C)			

Part C. Determination of the Energy Content of 1-Pentanol

	Trial 1	Trial 2	Trial 3
Initial mass of burner and cap (g)			
Final mass of burner and cap (g)			
Volume of water (mL)			
Initial temperature of water (°C)			
Maximum temperature of water (°C)			

DATA TREATMENT
Experiment 30

Part A. Calibration of the Calorimeter

	Trial 1	Trial 2	Trial 3
Mass of ethanol burned (g)			
Change in temperature of the water, ΔT (°C)			
Heat absorbed by the water, q_{water} (kJ)			
Heat generated by combustion of ethanol, $q_{combustion}$ (kJ)			
Efficiency of heat transfer			

(A.1) For Trial 1, show your calculation of the heat absorbed by the water, q_{water} . Remember that $q_{water} = m \cdot s \cdot \Delta T$. Assume that the density of the water is 1.00 g/mL. The specific heat for water is 4.18 J/g·°C.

(A.2) For Trial 1, show your calculation of the heat generated by the combustion of the ethanol, $q_{combustion}$. For ethanol, $\Delta H_{combustion}$ is −1366.83 kJ/mol.

(A.3) For Trial 1, show your calculation of the efficiency of heat transfer from the burner to the water in the can.

(A.4) Determine the average efficiency of heat transfer.

Average efficiency =

Part B. Determination of the Energy Content of Kerosene

	Trial 1	Trial 2	Trial 3
Mass of kerosene burned (g)			
Change in temperature of the water, ΔT (°C)			
Heat absorbed by the water, q_{water} (kJ)			
Heat generated by combustion of kerosene, $q_{combustion}$ (kJ)			
Energy content of kerosene (kJ/g)			

(B.1) For each trial, use the average efficiency calculated in (A.4) and the mass of kerosene burned to calculate the energy content for kerosene, in kJ/g. Show your calculation for Trial 1.

(B.2) Calculate the average energy content of kerosene.

Part C. Determination of the Energy Content of 1-Pentanol

	Trial 1	Trial 2	Trial 3
Mass of 1-pentanol burned (g)			
Change in temperature of the water, ΔT (°C)			
Heat absorbed by the water, q_{water} (kJ)			
Heat generated by combustion of 1-pentanol, $q_{combustion}$ (kJ)			
Energy content of 1-pentanol (kJ/g)			

(C.1) For each trial, use the average efficiency calculated in (A.4) and the mass of pentanol burned to calculate the energy content for pentanol, in kJ/g. Show your calculation for Trial 1.

(C.2) Calculate the average energy content of 1-pentanol.

QUESTIONS
Experiment 30

1. Ethanol (C_2H_5OH), 1-pentanol ($C_5H_{11}OH$), and vegetable oil are oxygenated fuels while kerosene is a non-oxygenated fuel. Vegetable oils are mixtures of various components with a chief constituent being glycerol trioleate, $C_{57}H_{104}O_6$. Calculate the percent by mass of oxygen in glycerol trioleate, 1-pentanol, and ethanol.

Glycerol trioleate _____ Pentanol _____ Ethanol _____

2. In the table below, enter the appropriate values as determined in question #1 and in the Data Analysis section.

Fuel	Percent Oxygen	Energy Content (kJ/g)
kerosene		
glycerol trioleate		– 45.3
1-pentanol		
ethanol		– 29.67

3. For an organic compound, combustion is the process of combining with oxygen to form water and carbon dioxide. Compare the energy contents of the fuels studied in this experiment. Comment on the correlation between the percent of oxygen in the fuel and the amount of energy available per gram of fuel.

Name _____ Section _____

Partner _____ Date _____

IN-LAB QUESTIONS
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Answer the following questions using complete sentences.

1. Give two reasons why it is important to minimize the amount of soot produced by any of the flames.

2. Why is it important that the water contain no ice?