

Experiment 30A

ENERGY CONTENT OF FUELS

FV 11/29/2010

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), 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 fuel values of various materials.

OBJECTIVES: By the end of this experiment, the students 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 fuel value for several fuels.
4. Compare the fuel value of an oxygenated and non-oxygenated fuel.

PRE-LAB: Read the entire lab guide and instructions and complete **page E30A-8**.

DISCUSSION:

Fuels

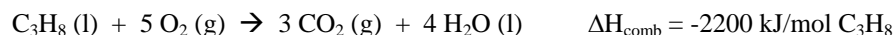
Combustion reactions are utilized in converting stored chemical energy into other forms of energy. Although mechanical and biological systems are quite different, they both utilize combustion.

Combustion of a hydrocarbon produces carbon dioxide and water. This reaction releases energy (exothermic reaction). Mechanical systems utilize this energy to do work. Two specific combustion reactions are shown below.

Natural gas, methane:

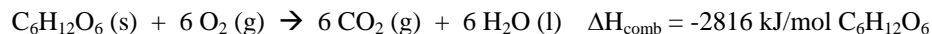


Propane:

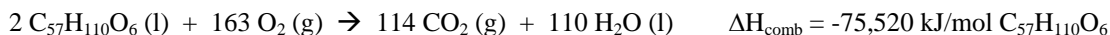


Carbohydrates and fats are examples of biological fuels (food). Although these are not hydrocarbons since they contain oxygen, they both undergo the same type of combustion reactions.

Glucose (simple carbohydrate):



Glycerol tristearate (fat):



Comparing Fuels

Comparing fuels can be difficult. Enthalpy of combustion (ΔH_{comb}) are given in the previous examples, but these depend on the moles of CO_2 and H_2O formed. Therefore, when comparing fuels it may be more useful to compare the *energy content* or *fuel value* of each fuel. **Fuel value (kJ/g)** is defined as the amount of energy released per gram of fuel. The fuel value for methane is 55.5 kJ/g while that of glucose is 15.6 kJ/g.

In this lab, the heat of combustion can be measured by a constant pressure calorimeter. The amount of fuel burned can be determined by difference in mass. Both of these measurements will be used to get the fuel value.

Calorimeter Efficiency

The efficiency of the calorimeter must be determined prior to calculating fuel values. By burning n moles of ethanol with a known fuel value, the $q_{\text{combustion}}$ can be calculated.

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

The q_{water} is determined from the mass of the water, it's specific heat, and the difference in temperature.

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

In this experiment it will become evident that all of the heat released by the burning of the fuels is not transferred to the water. The ratio of the heat absorbed by the water to the heat from the combustion of the fuel is the **efficiency** of the calorimeter.

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

Once the efficiency of the calorimeter has been determined, the q_{water} can be measured for other fuels, and then the $q_{\text{combustion}}$ of each fuel is calculated.

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

The ratio of the $q_{\text{combustion}}$ to the mass of fuel used is the **fuel value (kJ/g)**.

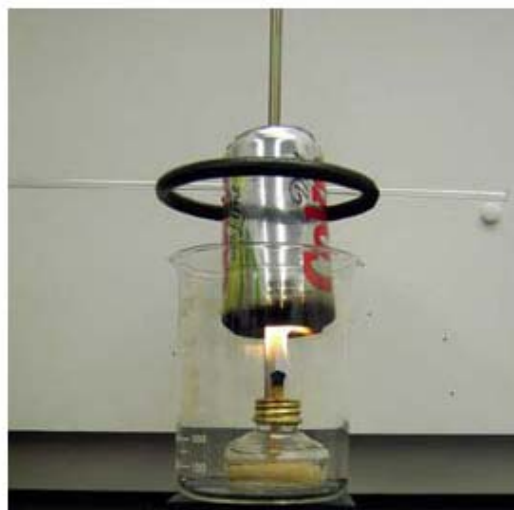
$$\text{Fuel value} = \frac{q_{\text{combustion}}}{\text{mass of fuel used}} \quad (5)$$

PROCEDURE:

SAFETY: The fuels used in this experiment are very flammable and care must be taken to avoid spillage. Aluminum cans also may have sharp edges.

Part A. Determining the Efficiency of the Calorimeter

1. Light the ethanol burner. If the flame is not an inch or less in height, extinguish the flame by replacing the burner cap. Adjust the height of the wick and recheck the flame. Once the wick is adjusted properly, extinguish the flame by recapping.
2. Weigh the ethanol burner with cap on a top-loading balance and record the mass. Place the capped burner in an 800 mL beaker.
3. Use steel wool to clean your aluminum can if it is sooty. Gently push a glass rod through the pre- drilled holes in the 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. **AFTER THIS ADJUSTMENT DO NOT CHANGE THE HEIGHT OF THE RING.**
4. Using a graduated cylinder, place 100 mL of water into the aluminum can. Record the temperature of the water.
5. Lift out the can, remove the burner cap, and then light the burner and replace the can as quickly as possible. Stir the water with the thermometer. (Don't just leave it sitting on the bottom!)
6. When the water temperature is about 40°C above its initial temperature, remove the can and quickly cap the burner using tongs. Re-suspend the can and keep stirring the water; record the highest temperature the water reaches.
7. Weigh and record the mass of the ethanol burner and cap.
8. Pour the water out of the can. If the can is sooty, clean it with steel wool. Readjust the wick if necessary.
9. 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. Determining the Fuel Value

1. Repeat Part A using the *n*-octane burner.
2. Repeat Part A using the 2-pentanol burner.

Clean-up:

1. Use the steel-wool to remove soot from the aluminum can. Do not discard the can!
2. Clean-up any spilled fuels and recap all burners.

Name _____ Section _____

Name _____ Date _____

DATA SECTION
Experiment 30A

Part A. Determining the Efficiency of the Calorimeter using Ethanol

Ethanol Burner	Trial 1	Trial 2	Trial 3
Final mass of burner and cap, g			
Initial mass of burner and cap, g			
Mass of ethanol combusted, g			
Final temperature of water, °C			
Initial temperature of water, °C			
ΔT , °C			

Part B. Determining the Fuel Value of *n*-octane and 2-pentanol

<i>n</i>-octane Burner	Trial 1	Trial 2	Trial 3
Final mass of burner and cap, g			
Initial mass of burner and cap, g			
Mass of <i>n</i> -octane combusted, g			
Final temperature of water, °C			
Initial temperature of water, °C			
ΔT , °C			

2-pentanol Burner	Trial 1	Trial 2	Trial 3
Final mass of burner and cap, g			
Initial mass of burner and cap, g			
Mass of 2-pentanol combusted, g			
Final temperature of water, °C			
Initial temperature of water, °C			
ΔT , °C			

DATA TREATMENT
Experiment 30A

Part A. Determining the Efficiency of the Calorimeter, using Ethanol

(A.1) Using the relationship $q_{\text{water}} = \text{specific heat} \times \text{mass} \times \Delta T$, show the calculation for q_{water} for Trial 1.

(A.2) Using the fuel value for ethanol determined in Pre-Laboratory Assignment 2c and the mass of ethanol combusted, show the calculation for $q_{\text{combustion}}$ for Trial 1.

(A.3) Show the calculation for the efficiency of energy transferred for Trial 1.

(A.4) Calculate q_{water} , $q_{\text{combustion}}$, and efficiency for your other trial(s) and summarize the results.

Ethanol	Trial 1	Trial 2	Trial 3
q_{water} , kJ			
$q_{\text{combustion}}$, kJ			
efficiency			
Average efficiency			

NOTE: This average efficiency will be used to calculate the fuel value of other fuels.

Part B. Determining the Fuel Value of *n*-octane and 2-pentanol

Determining the Fuel Value of *n*-octane

(B.1) Calculate the q_{water} for each *n*-octane trial. Record in the table on next page.

(B.2) Use the average efficiency for the ethanol burned and q_{water} for each *n*-octane trial to calculate the $q_{\text{combustion}}$ for each *n*-octane trial. Record in the table on next page.

(B.3) Use the $q_{\text{combustion}}$ and mass of *n*-octane combusted to calculate the fuel value for each *n*-octane trial and summarize the results.

<i>n</i> -octane	Trial 1	Trial 2	Trial 3
q_{water} , kJ			
$q_{\text{combustion}}$, kJ			
Fuel value, kJ/g			
Average fuel value			

Determining the Fuel Value of pentanol

(B.4) Calculate the q_{water} for each 2-pentanol trial and record in the table below.

(B.5) Use the average efficiency for the ethanol burned and q_{water} for each 2-pentanol trial to calculate the $q_{\text{combustion}}$ for each 2-pentanol trial. Record in the table.

(B.6) Use the $q_{\text{combustion}}$ and mass of 2-pentanol combusted to calculate the fuel value for each 2-pentanol trial. Record in the table.

pentanol	Trial 1	Trial 2	Trial 3
q_{water} , kJ			
$q_{\text{combustion}}$, kJ			
Fuel value, kJ/g			
Average fuel value			

Mass Percent of Oxygen in each Fuel

(B.7) Calculate the mass percent of oxygen in ethanol ($\text{C}_2\text{H}_5\text{OH}$), *n*-octane (C_8H_{18}), and 2-pentanol ($\text{C}_5\text{H}_{11}\text{OH}$).

Name _____

Section _____

Date _____

PRE-LAB EXERCISES
Experiment 30A

1. In this experiment a liquid fuel burner is used to heat water in an aluminum beverage can. Not all of the energy from the burner is transferred to the water, i.e., $q_{\text{combustion}} + q_{\text{water}} \neq 0$. Because of this, the efficiency of the energy transfer must be determined using the relationship:

$$\text{efficiency} = \frac{-q_{\text{water}}}{q_{\text{combustion}}}$$

- In this experiment what will be the sign of q_{water} ?
 - What is the sign of $q_{\text{combustion}}$?
 - Why is the negative sign included with the q_{water} ?
2. Write a balanced chemical equation for the complete combustion of one mole of liquid ethanol, $\text{C}_2\text{H}_5\text{OH}$. (Assume the water produced is in the liquid state.)
3. Using standard enthalpy of formation values found in Appendix 3 of your textbook, calculate ΔH for this reaction.
4. The fuel value of a substance is kilojoules of energy released per gram of fuel burned. Calculate the fuel value for ethanol.