Experiment 30A

FV 6/23/2016

ENERGY CONTENT OF FUELS

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 fuel value for several fuels.
3. 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:

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

Propane:

$$C_3H_8\,(l) + 5\,O_2\,(g) \rightarrow 3\,CO_2\,(g) + 4\,H_2O\,(l) \qquad \Delta H_{comb} = -2200 \text{ kJ/mol } C_3H_8$$

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):

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

Glycerol tristearate (fat):

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

Comparing Fuels

Comparing fuels can be difficult. Enthalpies 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 measure the fuel value.

Enthalpy of Reaction (Δ_rH)

It is very useful to know and understand enthalpy changes for various chemical and physical processes. Calorimetry has allowed us to use Hess's Law (Indirect Method) to determine enthalpy values for several chemical reactions. Another method to calculate the enthalpy of reaction known as the indirect method uses standard molar enthalpies of formation ($\Delta_f H°$) to calculate enthalpies of reactions. When using standard values, the following equation is used for calculations:

$$\Delta_r H° = \Sigma n\, \Delta_f H^o\ (products) - \Sigma\, n\, \Delta_f H^o\ (reactants)$$

Calorimeter Efficiency

Since our calorimeter systems are not perfect, not all the energy produced by the fuel is absorbed by the water in the calorimeter. To compensate for this, the efficiency of the calorimeter can be determined prior to calculating fuel values. By burning **n** moles of a fuel with a known $\Delta H_{combustion}$, the $\mathbf{q_{combustion}}$ can be calculated.

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

The $\mathbf{q_{water}}$ is determined from the mass of the water, its specific heat, and the temperature increase.

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

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.

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

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

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

$$\text{Fuel value} = \frac{q_{combustion}}{\text{mass of fuel used}} \qquad (4)$$

PROCEDURE:

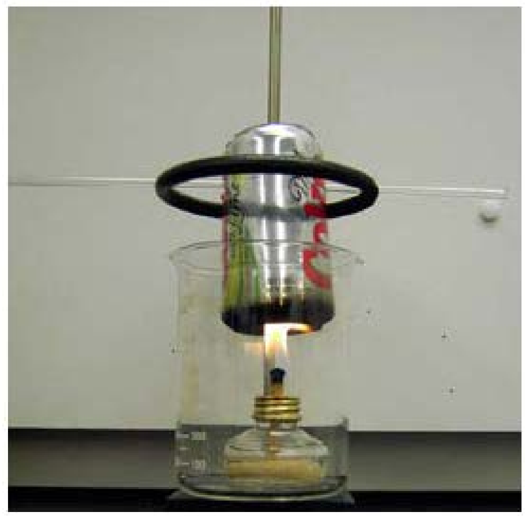
SAFETY: The fuels used in this experiment are very flammable and care must be taken to avoid spillage. Aluminum cans also may have sharp edges.

Determining the Efficiency of Energy Transfer using Ethanol

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. In the data table provided on page **E30A-4** perform a quick calculation of temperature change of the water (ΔT) divided by change in mass of fuel (Δm). If the first and second measurements are not within 10%, repeat the measurement a third time.

Determining the Energy Content of *n*-Octane and 2-Pentanol

1. Repeat steps 1-9 using the *n*-octane burner.

2. Repeat steps 1-9 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

Determining the Fuel Value of Ethanol, *n*-Octane and 2-Pentanol

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 (ºC)			
Initial temperature of water (ºC)			
ΔT (ºC)			
$\frac{\Delta T\ (^{o}C)}{\Delta m\ (g)}$			

***n*-octane Burner**	Trial 1	Trial 2	Trial 3
Final mass of burner and cap (g)			
Initial mass of burner and cap (g)			
Mass of n-octane combusted (g)			
Final temperature of water (ºC)			
Initial temperature of water (ºC)			
ΔT (ºC)			
$\frac{\Delta T\ (^{o}C)}{\Delta m\ (g)}$			

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 (ºC)			
Initial temperature of water (ºC)			
ΔT (ºC)			
$\frac{\Delta T\ (^{o}C)}{\Delta m\ (g)}$			

DATA TREATMENT
Experiment 30A

Determining the Efficiency of Energy Transfer and Fuel Value of ethanol

(1) Calculate the q_{water} for each ethanol trial using the equations from the discussion. Show the calculations for trial 1. Record all trials in the table below.

(2) Using the ΔH value calculated in pre-lab question 3 and the mass of ethanol combusted, calculate $q_{combustion}$ for each trial. Show the calculation for $q_{combustion}$ for Trial 1. Record all trials in the table below.

(3) Calculate the efficiency for each trial. Show the calculations for trial 1. Determine the average efficiency of the trials. Record all values in the table below.

(4) Use the $q_{combustion}$ and mass of ethanol combusted to calculate the fuel value for each ethanol trial and summarize the results.

ethanol	Trial 1	Trial 2	Trial 3
q_{water} (kJ)			
$q_{combustion}$ (kJ)			
Efficiency			
Average Efficiency			
Fuel value (kJ/g)			
Average fuel value (kJ/g)			

Determining the Fuel Value of *n*-octane

(1) Calculate the q_{water} for each n-octane trial. Record in the table below.

(2) Use the average efficiency calculated for the burner and q_{water} for each n-octane trial to calculate the $q_{combustion}$ for each n-octane trial. Record in the table below. Show one example calculation here.

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

n-octane	Trial 1	Trial 2	Trial 3
q_{water} (kJ)			
$q_{combustion}$ (kJ)			
Fuel value (kJ/g)			
Average fuel value (kJ/g)			

Determining the Fuel Value of pentanol

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

(2) Use the average efficiency calculated for the burner and q_{water} for each 2-pentanol trial to calculate the $q_{combustion}$ for each 2-pentanol trial. Record in the table.

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

2-pentanol	Trial 1	Trial 2	Trial 3
q_{water} (kJ)			
$q_{combustion}$ (kJ)			
Fuel value (kJ/g)			
Average fuel value (kJ/g)			

Mass Percent of Oxygen in each Fuel

(4) Calculate the mass percent of oxygen in ethanol (C_2H_5OH), *n*-octane (C_8H_{18}), and 2-pentanol ($C_5H_{11}OH$).

QUESTIONS
Experiment 30A

1. Oxygenated fuels (compounds containing C, H, and O) are used as motor vehicle fuels or fuel additives because they burn cleaner, thereby reducing air pollution. They also affect the miles per gallon. Compare the fuel value and mass percent oxygen for each fuel, and decide if oxygenation results in an increase or decrease in miles per gallon. Explain why.

2. Use the fuel values to calculate the maximum kJ/gal you could get from ethanol (0.789 g/mL) and n-octane (0.703 g/mL).

3. Developing alternative liquid fuels from renewable resources is an area of high interest to the Navy, which is one of the largest consumers of fossil fuels in the world. One possible approach is to develop algae that can be used to produce 2-butanol, which could replace gasoline. An approach to use algae to produce bio-diesel is to harvest fats and convert them to "fatty acids." If a typical chemical formula of a fatty acid is $C_{18}H_{38}O_2$, which fuel would have a higher fuel value?

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., $q_{combustion} + q_{water} \neq 0$. Because of this, the efficiency of the energy transfer must be used to determine the amount of energy generated by the fuel from the amount of energy absorbed by the water:

a. In this experiment, what will be the sign of q_{water}?

b. What is the sign of $q_{combustion}$ for the fuel in the burner?

c. Some of the energy of the fuel in the burner DOES NOT heat the water. Explain why NOT.

2. Write a balanced chemical equation for the complete combustion of one mole of liquid ethanol, C_2H_5OH. (Assume the water produced is in the gaseous state.)

3. Using the following enthalpies of formation, calculate ΔH for this reaction.

$\Delta_f H°$	(kj/mol)
C_2H_5OH (*l*)	-276.98
CO_2 (g)	-393.50
H_2O (g)	-241.80

4. The fuel value of a substance is kilojoules of energy released per gram of fuel burned. Calculate the fuel value for ethanol.