

## Experiment 30B

MEB 11/5/2022

# 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 stir rod, ring stand, cork, 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:

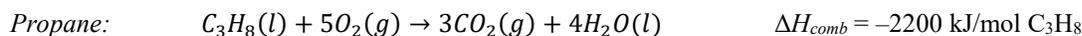
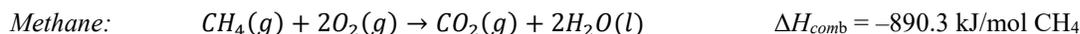
1. Construct and use an aluminum can calorimeter.
2. Calculate the fuel value for several fuels.
3. Compare the fuel value of an oxygenated and non-oxygenated fuel.

**PRE-LAB:** Read the discussion & procedure below and complete the attached pre-lab assignment.

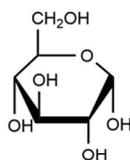
### DISCUSSION:

#### Fuels

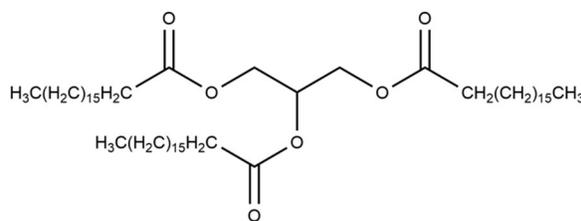
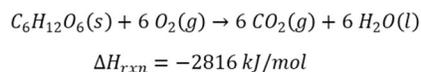
Combustion reactions involve a fuel, a substance which is oxidized<sup>1</sup>, reacting with molecular oxygen to form one or more oxygen-containing compounds. Combustion reactions are utilized to convert stored chemical energy into other forms of energy. Complete combustion of a hydrocarbon produces carbon dioxide and water as products and releases energy, mainly as heat (it is an *exothermic* reaction). Combustion reactions of common fuels methane (CH<sub>4</sub>) and propane (C<sub>3</sub>H<sub>8</sub>) are shown below:



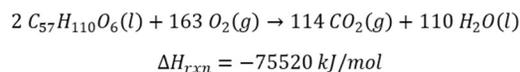
where  $\Delta H_{\text{comb}}$  is the enthalpy (heat) of combustion. Carbohydrates and fats are examples of biological fuels (food). Though they contain some oxygen, they both undergo the same type of reactions as hydrocarbons:



Glucose, simple carbohydrate



Glycerol Tristearate, a fat



**Figure 1.** Biological Fuels.

The two reactions shown above lead us to the topic of metabolism, the process by which foods undergo reaction to provide energy (and to enable a host of other functions within our bodies). Although the reactions of glucose and glycerol tristearate appear to be one-step processes, in our bodies these molecules are actually oxidized in a series of steps, or “metabolic pathway,” to generate the enthalpy of reaction. Because enthalpy is a state function, the overall enthalpy change is the same whether the reaction involves a direct combustion or a more complicated metabolic

<sup>1</sup> Oxidation processes involve the transfer of electron(s). In the above reactions, electrons are transferred from the fuels to the oxygen. This topic will be discussed in SC112.

pathway.<sup>2</sup> The energy-generating reactions in our bodies are not called combustion reactions; instead, they are simply called oxidation reactions (that proceed via exceptionally complicated metabolic pathways).

### Comparing Fuels: Fuel Value

Because the purpose of a fuel is to deliver energy, comparisons between fuels focus on the relative amount of energy released, which is normally expressed as a *positive* quantity (note that this is opposite to the usual thermochemical convention that heat released by an exothermic process is given a negative sign!). Also, comparing fuels can be difficult because molecules can contain vastly different numbers of carbon and hydrogen atoms, and thus form different amounts of CO<sub>2</sub> and H<sub>2</sub>O upon combustion (compare reactions above). When comparing different fuels, it is therefore most useful to compare the *energy content* or **fuel value (kJ/g)**, which corresponds to the *amount of energy released per gram of fuel*. For example, 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 (kJ) for several different fuels will be measured using a simple constant pressure calorimeter. The mass of fuel burned (g) will be determined by weighing the burner before and after reaction. Both of these measurements will be used to calculate the fuel value.

### Enthalpy of Reaction ( $\Delta H_{rxn}$ )

It is very useful to know and understand enthalpy changes for various chemical and physical processes. A convenient way to calculate enthalpy changes is to use tabulated standard molar enthalpies of formation ( $\Delta H_f^\circ$ ). The total enthalpy change for a reaction is simply the difference between the total enthalpies of formation of the products and that of the reactants:

$$\Delta H_{rxn}^\circ = \sum n \Delta H_f^\circ (\text{products}) - \sum n \Delta H_f^\circ (\text{reactants}) \quad (1)$$

where the  $n$  represents the stoichiometric coefficient of each reactant and product.

### Calorimeter Efficiency

We can experimentally determine  $\Delta H$  for a chemical reaction by calorimetry, a method for measuring heat changes in chemical and physical processes. In the simple calorimeter used in this lab, water in an aluminum can is heated by the combustion of a fuel. Note that not all the heat released by the burner is absorbed by the water (you can hold your hand over the set-up and feel some heat that escapes to the surroundings). In this lab, we will assume that the *same fraction* of the heat released by the burner is absorbed by the water in all trials — this will be referred to as the *efficiency*. To determine the efficiency of the calorimeter, we will burn a known amount of a fuel with a known  $\Delta H_{comb}$  and measure the temperature rise. Knowing the number of moles ( $n$ ) of the fuel that was burned, we can calculate the heat of combustion,  $q_{combustion}$ , using the equation:

$$q_{combustion} = n \Delta H_{comb} \quad (2)$$

With the temperature rise ( $\Delta T$ ), the mass of the water ( $m$ ), its specific heat ( $c_p$ ), the heat absorbed by the water,  $q_{water}$ , is determined.:

$$q_{water} = m c_p \Delta T \quad (3)$$

The ratio of the heat absorbed by the water in the calorimeter to the heat released by combustion of the fuel is the **efficiency** of the calorimeter:

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

By rearranging Equation (4) 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.

$$q_{combustion} = - \frac{q_{water}}{efficiency} \quad (5)$$

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

$$fuel\ value = \frac{q_{combustion}}{mass\ of\ fuel\ combusted} \quad (6)$$

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<sup>2</sup> Recall that Hess's law shows that if a process is the sum of two or more reactions, the enthalpy change for the overall process is the sum of the enthalpy changes for the constituent reactions.

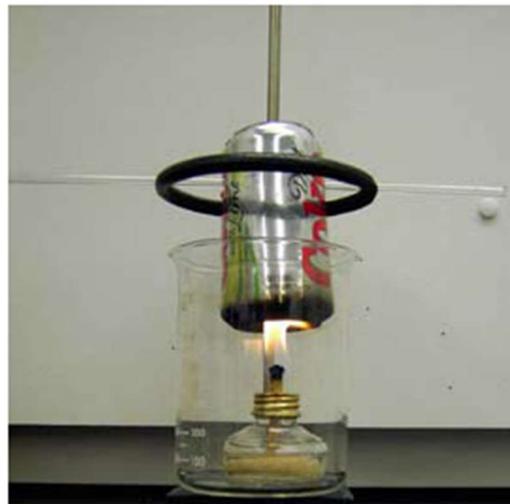
As stated above, the fuel value is reported as a positive quantity by definition. Fuel values for fuels with differing oxygen contents will be measured, and a correlation between oxygen content and fuel value will be deduced.

### PROCEDURE:

**SAFETY:** The point of fuels is that they are *flammable*. Care must be used in this experiment to avoid uncontrolled flames and 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 using the tongs. 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 given to the right. 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 **E30B-4** perform a quick calculation of temperature change of the water ( $\Delta T$ ) divided by change in mass of fuel ( $\Delta m$ ). If the first and second measurements are not within 10%, repeat the measurement a third time.



**Figure 2.** Experimental Setup

#### 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 \_\_\_\_\_

Partner \_\_\_\_\_

Date \_\_\_\_\_

**DATA SECTION**  
**Experiment 30B**

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

<b>Ethanol Burner</b>	Trial 1	Trial 2	Trial 3
Final mass of burner and cap (g)			
Initial mass of burner and cap (g)			
Mass of ethanol combusted, $\Delta m$ (g)			
Final temperature of water ( $^{\circ}\text{C}$ )			
Initial temperature of water ( $^{\circ}\text{C}$ )			
$\Delta T$ ( $^{\circ}\text{C}$ )			
$\Delta T / \Delta m$ ( $^{\circ}\text{C}/\text{g}$ )			

<b><i>n</i>-Octane Burner</b>	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, $\Delta m$ (g)			
Final temperature of water ( $^{\circ}\text{C}$ )			
Initial temperature of water ( $^{\circ}\text{C}$ )			
$\Delta T$ ( $^{\circ}\text{C}$ )			
$\Delta T / \Delta m$ ( $^{\circ}\text{C}/\text{g}$ )			

<b>2-Pentanol Burner</b>	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, $\Delta m$ (g)			
Final temperature of water ( $^{\circ}\text{C}$ )			
Initial temperature of water ( $^{\circ}\text{C}$ )			
$\Delta T$ ( $^{\circ}\text{C}$ )			
$\Delta T / \Delta m$ ( $^{\circ}\text{C}/\text{g}$ )			

**DATA TREATMENT**  
**Experiment 30B**

**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 (Eqn. 3, p. E30B-2). Show the calculations for trial number 1. Record the answer for this trial and the other trials in the table below.
2. Using the  $\Delta H$  value calculated in pre-lab question 3 and the mass of ethanol combusted, calculate  $q_{combustion}$  for each trial (Eqn. 2, p. E30B-2). Show the calculation for trial 1. Record the answer for this trial and the other trials in the table below.

$\Delta H$  for combustion of ethanol from pre-lab question 3: \_\_\_\_\_ kJ/mol

3. Calculate the efficiency for each trial (Eqn 4, p. E30B-3). Show the calculations for trial number 1. Record the answer for this trial and the other trials in the table below. Then, calculate the average efficiency.

Ethanol	Trial 1	Trial 2	Trial 3
$q_{water}$ (kJ)			
$q_{combustion}$ (kJ)			
Efficiency			
Average efficiency			

**Determining the Fuel Value of *n*-Octane**

1. Calculate the  $q_{water}$  for each *n*-octane trial. Show the calculations for trial number 1, and record the answer for this trial and the other trials in the table below.
2. Use the average efficiency calculated for the calorimeter and  $q_{water}$  for each *n*-octane trial to calculate the  $q_{combustion}$  for each *n*-octane trial (Eqn. 5, p. E30B-2). Show the calculations for trial number 1, and show all answers in the table below.



2. Summarize your calculated data on these fuels in the table below.

Substance	Ethanol (C <sub>2</sub> H <sub>5</sub> OH)	n-Octane (C <sub>8</sub> H <sub>18</sub> )	2-Pentanol (C <sub>5</sub> H <sub>11</sub> OH)
Average fuel value (kJ/g)			
% Oxygen by mass			

**QUESTIONS**  
**Experiment 30B**

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). 1 gallon = 4.00 quarts. 1 L = 1.056710 quarts

- ethanol

- *n*-octane



**PRE-LAB EXERCISES**  
**Experiment 30B**

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.,  $q_{\text{combustion}} + q_{\text{water}} \neq 0$  or more completely:

$$\text{efficiency} \cdot q_{\text{combustion}} + q_{\text{water}} = 0 \quad (7)$$

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  $q_{\text{water}}$ ? \_\_\_\_\_
  - b. What is the sign of  $q_{\text{combustion}}$  for the fuel in the burner? \_\_\_\_\_
  - c. Some of the energy of the fuel in the burner DOES NOT heat the water (i.e. efficiency  $\neq 1$ ). Explain why NOT.
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 gaseous state.)

3. Using the following enthalpies of formation, calculate  $\Delta H$  for the combustion reaction in question 2. Also, record your result on p. E30B-5, question 2.

$\Delta H_f^\circ$	(kJ/mol)
$\text{C}_2\text{H}_5\text{OH} (\text{l})$	-276.98
$\text{CO}_2 (\text{g})$	-393.50
$\text{H}_2\text{O} (\text{g})$	-241.80

4. The fuel value of a substance is kilojoules of energy released per gram (kJ/g) of fuel burned. Calculate the fuel value for ethanol using the  $\Delta H_{\text{comb}}$  calculated in question 3. Also, record your result in the summary table at the top of p. E30B-7.