Experiment 30B

10/2/2018

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, 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 the pre-lab.

DISCUSSION:

Fuels

Combustion reactions are between a substance, a fuel, which is oxidized[1] and molecular oxygen to form one or more oxygen-containing compounds and are utilized in converting stored chemical energy into other forms of energy. Complete 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:

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

Propane:

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

where $\Delta_{comb}H$ is the enthalpy of combustion. 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 reactions:

Glucose (simple carbohydrate):

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

Glycerol tristearate (fat):

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

The above two reactions 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 oxidizedin a series of steps to generate the enthalpy of reaction. Because enthalpy is a state function, the enthalpy change is the same whether the reaction involves a direct combustion or a more complicated metabolic pathway.[2] The energy-generating reactions in our bodies are not called combustion reactions; instead, they are simply called oxidation reactions (that proceed via complicated metabolic pathways).

Comparing Fuels

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[1] 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.

[2] 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.

The purpose of a fuel is to deliver energy, so comparisons made between fuels measure their energy <u>released</u>. Therefore, an exothermic reaction delivers a <u>positive</u> amount of energy. Also, comparing fuels can be difficult because molecules can package wildly different number of carbons and hydrogens to form CO_2 and H_2O upon combustion (compare reactions above). When comparing fuels it may be more useful to compare the *energy content* or ***Fuel value (kJ/g)*** *as defined as the amount of energy released per gram of fuel. is* 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) can be measured by a constant pressure calorimeter. The mass of fuel burned (g) can be determined by difference. Both of these measurements will be used to calculate the fuel value.

Enthalpy of Reaction ($\Delta_r H$)

It is very useful to know and understand enthalpy changes for various chemical and physical processes. A convenient way is to use tabulated standard molar enthalpies of formation ($\Delta_f H°$). When using these standard values, the following equation is used for the calculations:

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

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

Calorimeter Efficiency

We can experimentally determine ΔH for a chemical reaction by calorimetry, a method for measuring heat changes in chemical and physical processes. In our calorimeter system, water in an aluminum can is heated by the combustion of a fuel. Not all the heat goes into the water (you can hold your hand over the set-up and feel some heat given off). We will make the assumption that the same fraction of heat goes into the water—so to speak, the efficiency of the calorimeter. To determine the efficiency of the calorimeter, we will burn fuel with a known $\Delta_{comb}H$ and measure the temperature rise. Knowing the number of moles of the fuel, **n,** the heat of combustion, $\mathbf{q_{combustion}}$, can be calculated using the equation

$$q_{combustion} = n \cdot \Delta_{comb}H \tag{1}$$

With the temperature rise (ΔT), the mass of the water (m), its specific heat (C_{sp}), the heat absorbed by the water, $\mathbf{q_{water}}$, is determined.:

$$q_{water} = m \cdot C_{sp} \cdot \Delta T \tag{2}$$

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

$$\text{efficiency} = -\frac{q_{water}}{q_{combustion}} \tag{3}$$

By rearranging Equation 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.

$$q_{combustion} = -\frac{q_{water}}{\text{efficiency}} \tag{4}$$

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

$$\text{fuel value} = \frac{q_{combustion}}{\text{mass of fuel combusted}} \tag{5}$$

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

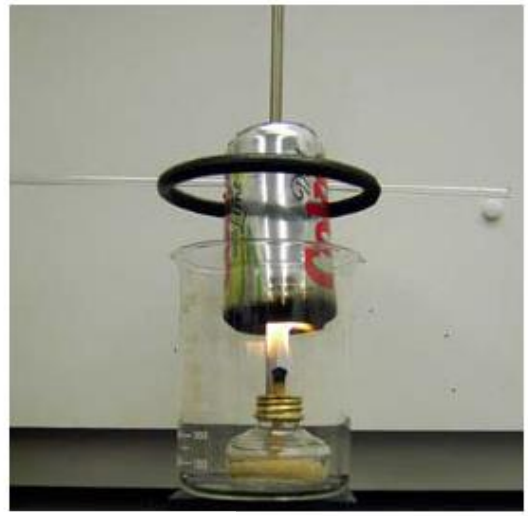
1. Light the ethanol burner. If the flame is not an inch or less in height, extinguish the flame by replacing the burner cap. Adjust the height of the wick and recheck the flame. Once the wick is adjusted properly, extinguish the flame by recapping.

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!)



Experimental setup

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 <u>highest</u> 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 (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 ______________________

Partner ___________________________________ Date ________________________

DATA SECTION
Experiment 30B

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, **Δm (g)**			
Final temperature of water (°C)			
Initial temperature of water (°C)			
ΔT (°C)			
ΔT / Δm (°C/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, **Δm (g)**			
Final temperature of water (°C)			
Initial temperature of water (°C)			
ΔT (°C)			
ΔT / Δm (°C/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, **Δm (g)**			
Final temperature of water (°C)			
Initial temperature of water (°C)			
ΔT (°C)			
ΔT / Δm (°C/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 2, 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 ΔH value calculated in pre-lab question 3 and the mass of ethanol combusted, calculate $q_{combustion}$ for each trial (Eqn 1, p. E30B-2). Show the calculation for trial 1. Record the answer for this trial and the other trials in the table below.

   ΔH for combustion of ethanol from pre-lab question 3: _______________ kJ/mol

3. Calculate the efficiency for each trial (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. 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 4, p. E30B-2). Show the calculations for trial number 1, and show all answers in the table below.

3. Enter in the table below the mass of *n*-octane combusted for each trial. Then calculate the fuel value for each *n*-octane trial and the average fuel value. Enter the answers in the table below.

n-Octane	Trial 1	Trial 2	Trial 3
q_{water} (kJ)			
$q_{combustion}$ (kJ)			
Mass of *n*-octane combusted, Δm (g)			
Fuel value (kJ/g)			
Average fuel value (kJ/g)			

Determining the Fuel Value of 2-Pentanol

1. Calculate the q_{water} for each 2-pentanol 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 2-pentanol trial to calculate the $q_{combustion}$ for each 2-pentanol trial. Show the calculations for trial number 1, and show all answers in the table below.

3. Enter in the table below the mass of 2-pentanol combusted for each trial. Then calculate the fuel value for each 2-pentanol trial and the average fuel value. Enter the answers in the table below.

2-Pentanol	Trial 1	Trial 2	Trial 3
q_{water} (kJ)			
$q_{combustion}$ (kJ)			
Mass of 2-pentanol combusted, Δm (g)			
Fuel value (kJ/g)			
Average fuel value (kJ/g)			

Mass Percent of Oxygen in each Fuel

1. Determine the mass percent of oxygen in ethanol (C_2H_5OH), 2-pentanol ($C_5H_{11}OH$), and *n*-octane (C_8H_{18}). *Remember to show all work.*

   - ethanol

   - *n*-octane

   - 2-pentanol

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

Substance	**Ethanol (C_2H_5OH)**	**n-Octane (C_8H_{18})**	**2-Pentanol ($C_5H_{11}OH$)**
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

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, 2-butanol or the typical fatty acid? *Note*: Many factors, such as saturated vs unsaturated, affect the fuel value, but base your answer only on oxygen content.

4. Biofuels and renewable energy are of great interest to the U.S. Navy and continued global operations with minimal dependence on petroleum based fuels. Read the discussion at the provided link and give a brief description of the Great Green Fleet initiative and the purpose. Great Green Fleet

Name ________________________________ Section ____________________

Date ______________________

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., $q_{combustion} + q_{water} \neq 0$ or more completely:
2.

$$\text{efficiency} \cdot q_{combustion} + q_{water} = 0 \qquad (6)$$

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 (i.e. efficiency ≠ 1). Explain why NOT.

3. 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.)

4. Using the following enthalpies of formation, calculate ΔH for the combustion reaction in question 2. Also, record your result on p. E30B-5, question 2.

$\Delta_fH°$	(kJ/mol)
C_2H_5OH (l)	–276.98
CO_2 (g)	–393.50
H_2O (g)	–241.80

5. 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 $\Delta_{comb}H$ calculated in question 3. Also, record your result in the summary table at the bottom of p. E30B-6.