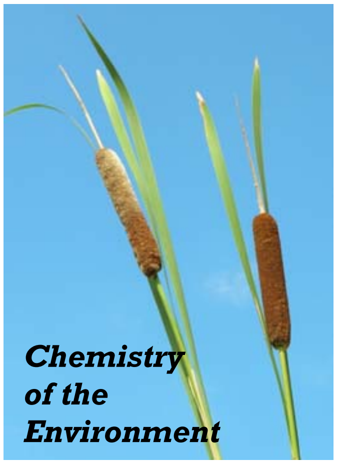

Comparison of the Energy Content of Fuels

Introduction

The ready availability of high-quality energy sources is one of the most important issues facing the global community. In this experiment, you will investigate the energy content of two hydrocarbon fuels by using them to heat water. The data collected will enable you to compare the energy provided by ethanol and a lamp fuel in order to see with one provides more energy for a given mass of fuel burned. You will be able to draw conclusions about the energy implications of oxygenated fuels such as ethanol.

Hydrocarbons are among the most energy-rich and cleanest-burning fuels. These compounds contain only hydrogen and carbon in their molecular structure. The simplest example of a hydrocarbon fuel is natural gas, or pure methane. When methane is burned completely, the products are carbon dioxide and water.

$$CH_4 + 2\ O_2 \rightarrow CO_2 + 2\ H_2O$$

Other hydrocarbons such as propane (C_3H_8) and butane (C_4H_{10}) will also burn to produce the same products, carbon dioxide and water. The lamp oil you will be using today is also a hydrocarbon and has the approximate composition of $C_{12}H_{26}$.

Ethanol, the other fuel that will be used today, contains oxygen in addition to carbon and hydrogen. In effect, this fuel is a hydrocarbon that has already partially reacted with oxygen:

$$2\ C_2H_6 + O_2 \rightarrow 2\ C_2H_5OH$$

In this experiment you will measure the amount of heat given off by known amounts of the two mentioned fuels, ethanol and lamp oil. The experimental procedure is to heat some water by burning a measured amount of the fuel samples. *It takes exactly 1 calorie (cal) of heat to raise the temperature of 1 gram of liquid water by 1° C.* Therefore if you know the mass of water and how many degrees the temperature goes up, then the total amount of heat absorbed by the water can be calculated as follows:

$$\text{Heat absorbed (cal)} = \text{mass of water} \times \text{temp. change} \times 1.00\ \text{cal/g}^{\circ}C$$

Theoretically, the amount of heat liberated by the burning fuel should equal the heat absorbed by the water, but in practice, some of the heat will be lost to the surroundings. **It is important to perform the experimental runs as consistently as possible to minimize differences in heat loss!**

Experimental Procedure

***Important note: Success of this experiment relies heavily on the accuracy of collecting mass measurements using laboratory balances.

1. Obtain a dry soda can apparatus. The can will have extended metal paper clips through four small holes in the can that allow it to be suspended on a ring stand.
2. Obtain a fuel burner with either lamp oil or ethanol in it. Label the data section in correspondence to the fuel you are using. Place the burner under the can and adjust the height of the ring stand so that the can is suspended about 2 centimeters above the top of the wick.
3. Take the soda can with the extended metal paper clips to the balance. Check to be sure the empty balance reads 0.000 gram (If it doesn't, press the button marked T, which will tare the instrument to make it read 0.000; then weigh the can and clips and **record the mass to the nearest 0.01 gram.**
4. Using a graduated cylinder, add approximately 100 mL of water to the can. The graduated cylinder is read below the meniscus, which is the curved top of a column of liquid in a small tube. **Reweigh the can + water to the nearest 0.01 gram.** Using subtraction, calculate the mass of the water in the can.
5. Suspend a thermometer from the ring stand so that the thermometer is in the water but is not touching the bottom of the can. The thermometer is marked at 1 degree intervals increasing from the red glass bulb at the bottom of the thermometer. The 0.1 degree should be estimated. **Record the initial temperature of the water** to the nearest 0.1 degree Celsius.
6. Take your particular fuel burner to the balance. Remove the burner cap. Check to be sure the empty balance reads 0.00 gram. Weigh the burner and **record the fuel**

type and its mass to the nearest 0.01 gram. **Be sure you record the type of fuel you are using.**

7. Place the fuel burner under the soda can and ***carefully*** light the burner. Observe the flame. If necessary, adjust the height of the can so that the top of the flame is just below the bottom of the can.
8. Heat the water until the temperature increases about 20 degrees Celsius, then extinguish the flame by blowing it out. Stir gently with the thermometer until the temperature stops rising; record the highest temperature, again estimating in to the nearest 0.1 degree. Calculate temperature change by subtracting the initial temperature from the final temperature; **record the temperature change.**
9. Take your **cooled burner** (be careful!) and data sheet to the balance. **Record the new mass of the burner**. Calculate the mass of the fuel burned by subtracting the final weight of the burner from the original weight of the burner.
10. You can now repeat the procedure with the other fuel and 100 mL more of water. Reweigh 100 mL of water in the can. The mass of the empty can should not change. ***Make sure you label the data so you know to what fuel it belongs.***

Calculations

To calculate the amount of heat liberated by 1 gram of burning fuel, you need the following information from both trials and of which should be on your data sheet:

- the mass of the water that was heated
- the change in temperature of the water
- The mass of fuel that was burned.

It takes 1.00 calorie of energy to raise the temperature of 1 gram of water 1 degree Celsius, the total heat absorbed by the water is found using the calculation:

$$\text{Heat absorbed (cal)} = \text{mass of water} \times \text{temp. change} \times 1.00\ \text{cal/g}^{\circ}C$$
$$= m \times \Delta T \times 1.00\ \text{cal/ g}^{\circ}C$$

1. **For both trials**, calculate the total calories of heat absorbed by the water. This will be assumed to be the same as the amount of heat liberated by the burning fuel.

$$\text{Heat absorbed (cal)} = \text{mass of water} \times \text{temp. change} \times 1.00\ \text{cal/g}^{\circ}C$$

2. **For both trials,** calculate the calories of heat per 1 gram of fuel burned.

   Calories per gram of fuel = Heat absorbed (cal) / grams of fuel burned

3. Average the results of Calculation 2 to find the average number of calories of heat produced per gram of fuel burned

Sample Calculation

102.54 g water $\Delta T = 24.5° C$ Fuel Burned = 1.47 g fuel

1. Heat absorbed (cal) = mass of water × temp. change × 1.00 cal/g°*C*

   Heat absorbed (cal) = (102.54 *g water*)(24.5° *C*)(1.00 *cal/ g°C*)

   Heat absorbed (cal) = 2512.2 calories

2. Calories per gram of fuel = Heat absorbed (cal) / grams of fuel burned

   Calories per gram of fuel = 2512.2 calories / 1.47 g fuel

   Calories per gram of fuel = 1709 cal/g fuel

Results

1. Average calories per gram of fuel produced
2. Identify which fuel was used