Comparison of Energy Content of Fuels[1]

Objective: Knowing that different fuels have greater heat values you will investigate which are best to use and determine the chemical properties of an energy rich fuel.

Background: All fuels, when combusted, are reacted with oxygen from the air. You may have heard of switching from a heavy polluting fuel like gasoline or diesel to methane or hydrogen. Being in the Midwest you've certainly heard the benefits of using ethanol as an oxygenate. Butane is commonly used in lighters while propane is used in grilling. Calorimetry is a technique by which you can determine the amount of heat in a fuel per gram of fuel. You will use calorimetry to study ethanol (C_2H_5OH), butanol (C_4H_9OH), propanol (C_3H_7OH), kerosene (a mixture of hydrocarbons but approximated to be $C_{10}H_{22}$), and candle wax (again approximated to be a saturated hydrocarbon with 20 carbons).

It takes 1 calorie (cal) to raise the temperature of 1 gram of liquid water 1°C. Therefore, the total amount of heat absorbed by the water equals:

$$\text{Heat absorbed} = \text{g}\,(H_2O) \times \Delta T \times \frac{1\ \text{cal}}{\text{g}\ ^{\circ}\text{C}}$$

This is also the amount of heat liberated by the fuel and furnished by it to heat the water.

Procedure:

1. Obtain a dry, cut-off soda can with two holes punched in the side and weigh the empty can.
2. Add approximately 50 mL of water to the can and reweigh it to determine the mass of the water added.
3. Thread a glass rod through the holes in the can and suspend it from the ring attached to a ring stand as shown in the diagram.

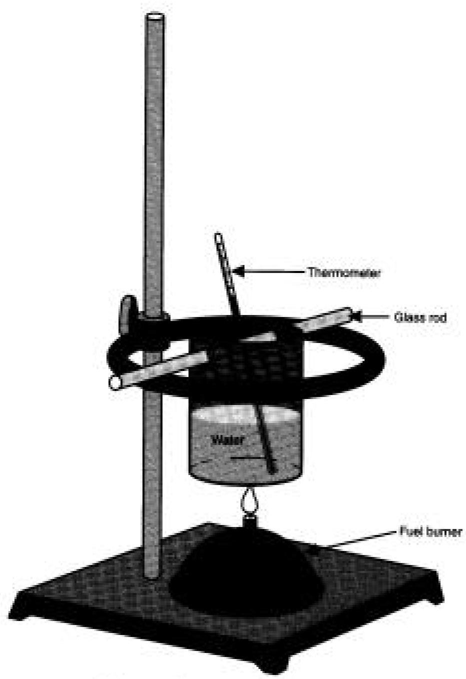

Diagram of can and burner setup

[1] Adapted from *Chemistry in Context Laboratory Manual*, ACS, Wm. C. Brown, Dubuque, IA, 1994.

4. Put a thermometer in the can and measure the temperature of the water to the nearest 0.1 °C.
5. Place a fuel burner under the can and adjust the ring so that the bottom of the can is about 2 cm from the top of the wick of the burner. **Never carry a lit burner. Imagine what would happen should it drop and break!**
6. Weigh the fuel burner and determine its mass.
7. Place the fuel burner under the soda can and light the burner.
8. Heat the water until the temperature has increased about 20 °C then extinguish the flame.
9. Stir the water gently until the temperature stops rising and record the temperature to the nearest 0.1 °C
10. Determine the mass of the fuel used by reweighing the fuel burner and subtracting that weight from the original weight of the burner.
11. Repeat the experiment with two additional fuels. You should not do all three alcohols.

Write Up

Complete the write up sheet posted at the web page.

Questions:

1. Calculate the heat liberated from burning each fuel.
2. Calculate the heat liberated from burning a gram of each fuel.
3. Average the class data.
4. Calculate the heat liberated from burning a mole of each fuel.
5. What is the relationship between the heat liberated from each fuel and the percentage of oxygen in the fuel?
6. Using the ethanol and butanol data, could you have predicted the propanol data with any accuracy? Can you predict the approximate amount of heat liberated by burning a gram of butane?
7. Grain alcohol is 95% C_2H_5OH and 5% H_2O. How does this affect the amount of heat per gram of alcohol?
8. Compare your results of the hydrocarbons to the standard accepted value of 11.5 kcal/g. There are many sources of error in this experiment. List five of them and suggest their order of magnitude. Will the error cause the value to be too large or too small?