

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 the late twentieth century. Chapter 4 in *Chemistry in Context* examines some of the societal choices we face and the advantages/disadvantages of various fossil fuels. Hydrocarbons (and mixtures of hydrocarbons) are among the most energy-rich and cleanest-burning fuels. Many of them are liquids, which make them attractive for transportation fuels.

In this experiment, you will investigate the energy content of several hydrocarbon fuels by using them to heat water. The data that you and your classmates obtain will enable you to compare fuels to see which ones provide more energy for a given mass of fuel burned. You will also be able to draw conclusions about the energy implications of using oxygenated fuels such as ethanol.

Background Information

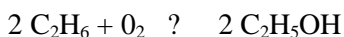
Fuels are substances that burn to give off relatively large amounts of heat. In an overall sense, such burning is simply a combustion reaction between the fuel and oxygen. The amount of heat generated depends on what kind of fuel is used and how much of it is burned.

The simplest example of a common fuel is natural gas, which is almost pure methane (CH₄). When methane is burned completely, the only products are carbon dioxide and water.



Methane is an example of a hydrocarbon fuel. Hydrocarbons are compounds that contain only hydrogen and carbon. Other hydrocarbons, such as propane (C₃H₈) and butane (C₄H₁₀), will burn to produce the same products, carbon dioxide and water.

Some fuels such as ethanol (ethyl alcohol, C₂H₅OH) contain oxygen in addition to carbon and hydrogen. In effect, they are hydrocarbons that have already partially reacted with oxygen:



In this experiment, you will measure the amount of heat given off by known amounts of several fuels. Your instructor will tell you exactly which fuels will be studied and how the class assignments will be made. Some of the fuels will be alcohols such as methanol (CH₃OH), ethanol (C₂H₅OH), isopropanol (C₃H₇OH), or butanol (C₄H₉OH). Some will be hydrocarbons, such as lamp oil or candle wax. Each of the latter is actually a mixture of hydrocarbons, but we will approximate their compositions as C₁₂H₂₆ (lamp oil) and C₄₀H₈₂ (candle wax).

The experimental procedure is to heat some water by burning a measured amount of a fuel sample. It takes exactly 1 calorie (cal) of heat to raise the temperature of 1 gram of liquid water by 1° C. (In Section 5.9 of the text, this is called the "specific heat.") Therefore, if you know the mass of water or how many degrees the temperature goes up, then the total amount of heat absorbed by the water can be calculated as follows:

$$\begin{aligned} \text{heat absorbed (calories)} &= \text{mass of water (grams)} \times \text{temp. change (C)} \times 1.00 \text{ cal/g}^\circ\text{C} \\ &= m \times \Delta T \times 1.00 \text{ cal/g}^\circ\text{C} \end{aligned}$$

ΔT is a shorthand way of saying "change of temperature." In this equation, the last term (1.00 cal/g° C) is needed to make the units come out right. (You should convince yourself that combining and canceling the units on the right

side will leave only calories.) 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.

Overview of the Experiment

1. Assemble the apparatus and obtain a burner containing a known fuel.
2. Add a measured volume of water to the can and then determine the mass of the water.
3. Weigh the burner.
4. Record the initial temperature of the water.
5. Light the burner and heat the water until the temperature increases about 20°C.
6. Record the highest temperature of the water.
7. Weigh the burner again, in order to find the mass of fuel used.
8. Repeat with two or more additional trials.
9. For each trial, calculate the amount of heat released per gram of fuel burned.

Important Note About Use of Laboratory Balances:

The success of this experiment depends heavily on the accuracy of many mass measurements. All weighing will be done on a laboratory balance, which you will be sharing with other students. Your instructor will explain the use of the particular balances in your lab. For each weighing, it is important to know that the balance reads *exactly zero* with nothing on the balance pan. For some kinds of laboratory balances, the "zero" can be changed easily, and perhaps unintentionally, so your measurements might be in error. With a balance of that type, it is important to check the "zero" *each* time you make a weighing.

Experimental Procedure:

The following general procedure is to be used for each measurement. Your instructor will tell you which fuel or fuels you are to investigate. You may be asked to do one trial with each of three different fuels. Alternatively, you may be asked to do several trials with a single fuel to improve the level of your confidence about an average value for that particular fuel, in which case the results for the whole class will be assembled for comparison.

Safety Notes;

- ? You need to be aware constantly that you (and other students) are working with flammable solvents and open flames. Handle the fuel burners very carefully. Before starting the experiment, be sure you know where a fire extinguisher is located in your laboratory.
 - ? As with any experiment involving open flames, long hair should be tied back and extremely loose sleeves should be avoided.
1. Obtain a dry soda can with the top removed and two holes punched on opposite sides near the top. Slide a glass rod through the holes in the can so that it can be suspended from the ring attached to a ring stand as shown in the diagram on the next page.
 2. Obtain a fuel burner, place it under the can, and adjust the height of the ring so that the bottom of the can is about 2 centimeters above the top of the wick. (*Figure 11.1.*)
 3. Take the empty soda can plus your data sheet to a balance. Check to be sure that the empty balance reads 0.0 gram; then weigh the can and record the mass to the nearest 0.1 gram.
 4. Using a graduated cylinder, add approximately 100 mL of water to the can. Reweigh the can plus water to the nearest 0.1 gram. By subtraction, calculate the mass of water in the can.
 5. Put a thermometer in the can, stir the water for a few moments, and then measure the temperature of the water, trying to estimate to the nearest 0.1 degree Celsius. (If you are using a glass thermometer, it is probably marked

at 1 degree intervals, so it is not possible to measure the temperature accurately to a tenth of a degree. Nevertheless, it is useful to make the best estimate that you can.)

6. Take the fuel burner plus your data sheet to a balance. If necessary, check to be sure that the empty balance reads zero. Weigh the burner and record its mass on the data sheet, reporting it to the nearest 0.01 gram (or the nearest 0.001 gram if the balance permits this).

Note that the can and water need only be weighed to the nearest 0.1 gram, while the burner must be weighed to the nearest 0.01 or 0.001 gram.

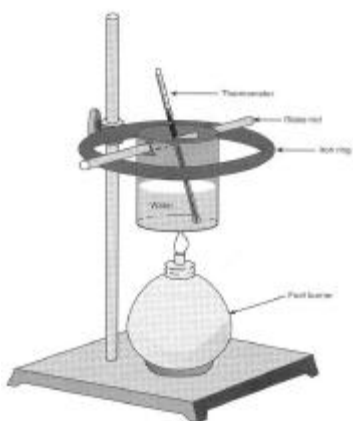


Figure 11.1 Diagram of can and burner setup

7. Place the fuel burner under the soda can and light the burner. Observe the flame. If necessary, cautiously adjust the height of the can so that the top of the flame is just below the bottom of the can.
8. Stir the water occasionally and continue heating the water until the temperature has increased about 20 degrees Celsius; then extinguish the flame. Do the next two steps quickly.
9. Continue stirring the water gently until the temperature stops rising; then record the highest temperature, again estimating it to the nearest 0.1 degree. Calculate the temperature change by subtracting the initial temperature from the final (highest) temperature.
10. Take the burner and your data sheet to a balance. If necessary, check that the balance reads zero and then weigh the burner and record the mass to the nearest 0.01 gram (or 0.001 gram). Calculate the mass of fuel burned by subtracting the final weight of the burner from the original weight of the burner.
11. Before doing another measurement, take a few moments to discuss the procedure you're your partner. Did you encounter any difficulties? Can you think of any desirable improvements in the procedure? "Also do the calculations, as described in the next section. This will give you a better idea of what you are looking for in subsequent trials.
12. Now repeat the procedure, either making additional measurements with the same fuel or switching to different fuels, as directed. If the water in the can has cooled to nearly room temperature, you can use the same water sample. If not, it is a good idea to empty the can and add a fresh 100-mL portion of water. Note that you recorded the mass of the dry empty can before you started. This will not change, except for possible buildup of soot on the bottom; therefore, you need to measure only the mass of the can with water in it.

Calculations

To calculate the amount of heat liberated by 1 gram of burning fuel, you need the following items of information from each trial, all of which should be on the data sheet: (a) the mass of water that was heated, (b) the change in temperature of the water, and (c) the mass of fuel that was burned.

Since 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 equal to the temperature change x the grams of water x 1.00 cal/gram degree.

Heat absorbed = temperature change x grams of water x 1.00 cal/g°C

1. For each trial, 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.

2. Then calculate the calories of heat per 1 gram of fuel burned.

You should do these calculations for each trial as soon as it is finished to see how the results are coming out.

Interpretation of Results

Because it was only possible for you to do a few trials, it is desirable to assemble a much larger body of data from your whole class or lab section so that more reliable comparisons of fuels can be made. Your instructor will tell you how to post or report your results for the rest of the class to see. You may be asked to calculate class averages for each fuel that was used. The following questions should help to focus your interpretation of the results.

Optional Extensions

1. Investigate the effects of changes in procedure. These might include (a) adding some nonflammable insulation around the can, (b) adding a cylindrical shield of aluminum foil around the burner, (c) using 200 mL of water and only a 10°C temperature rise (or 50 mL of water and a 10°C temperature change).

2. Energy content of wood. See Performance-Based Assessment Activity #3.

3. Caloric content of nuts. See Performance-Based Assessment Activity #4.

Questions To Be Answered After Completing This Experiment

Write out answers to the following questions on a separate sheet of paper and hand it in along with the data sheet.

1. Rank the fuels in order from least to most “heat per gram of fuel.” Now identify each fuel in your ranking as either a hydrocarbon or an alcohol. What general rule can you propose about the heat content of hydrocarbons versus that of alcohols?

2. Which class of fuels, hydrocarbons or alcohols, contains oxygen? Based upon your answer to Question 1, how does the presence of oxygen in the structure of a fuel affect its heat content? Give a reasonable explanation for the difference.

3. Gasohol is a mixture of gasoline (hydrocarbons) and ethanol (an alcohol).

a. How would the energy content per gram of gasohol compare to that of plain gasoline?

b. How do you think this would affect a car's fuel efficiency (miles per gallon)?

c. Considering that gasoline comes from petroleum, while ethanol can be made from corn, explain why environmentalists might promote the use of gasohol. Now explain why farmers might promote the use of gasohol. Finally, explain how you think someone concerned with feeding the poor might feel about increasing use of gasohol.

4. Carbohydrates have the general formula CH_2O , while fats typically have a general formula of about $\text{C}_{10}\text{H}_{19}\text{O}$.

Which of these classes of foods has the greater oxygen content as a percentage of the molecule's mass? Since foods are simply fuels for the body, which of these will release more energy when metabolized in the body? Explain your reasoning. How does this conclusion compare to the relative calorie content of fats and carbohydrates? (If you are not sure how many dietary Calories are in a gram of fat or carbohydrate, see Section 11.8 in your textbook.)

5. Suppose you put 50 mL of water in the can instead of 100 mL and heated it 40 degrees instead of 20 degrees. In what ways, if any, would this affect the results of the experiment? What would happen if you used 200 mL of water and heated it only 10 degrees? Can you think of any advantages or disadvantages in using either 50 mL or 200 mL of water in this experiment?
6. There are many possible sources of error in this experiment. List three that you can think of. Would each error have a large effect, a medium effect, or a small effect on the calculated heat content of a fuel? Also indicate whether the calculated result would be too high, too low, or could go either way.

This exercise was adapted from the following reference:

Stratton, W., Steehler, G., Pienta, N., & Middlecamp, C. (2003). Laboratory Manual for Chemistry in Context, 4th Edition. McGraw Hill Higher Education, Boston, MA. pp. 67-73.