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Science 30

Society:

http://swileman.wix.com/alternativeenergies

Technology:

https://www.lucidpress.com/documents/edit/08c6400a-02db-4bd9-9fba-0316d70ba903#

Science:

Determining the Experimental Value of Heat of Combustion of Two Different Fuels (Ethanol (C_2H_6O) and Butane (C_4H_{10})) by Measuring the Temperature Change of 20 mL of Water (H_2O) Using a Hand Built Calorimeter Until Each Fuel to Loses 5g of Mass to Figure Out Which Fuel Produces More Energy

Abstract:
The purpose of this lab was to compare the heat combustions (enthalpy) of a two different fuel sources to determine which is the most effective at producing energy. The efficiency of the fuels should have an effect on the enthalpy that the system produces in the combustion reaction. The calorimeter was set up with the ethanol burner and butane lighter to heat 20 mL and loss 5g of mass. These are the manipulated and controlled variables respectively and the responding variable is the temperature to determine the enthalpy/energy produced. Experiment 1 with the ethanol burner and ethanol inside had an experimental enthalpy of 1.31985kJ, and the theoretical value of the enthalpy was -1234.8 kJ. Experiment 2 with the butane had an experimental enthalpy of 0.6704kJ, and a theoretical enthalpy of -2657.3 kJ. The errors in this experiment included loss of heat in the open system calorimeter, assuming the exact loss of weigh of fuels is accurate to the substance inside and the inaccuracy of measuring distance between the can and the flame produced by the fuels.

Background Information:

The first thing we should know about for this lab would be enthalpy which is a thermodynamic quantity equivalent to the total heat content of a system (Bissell, 2003). In simpler terms it's the overall change in energy in a system. So for example in an endothermic (reaction that absorb heat from surroundings) the enthalpy would be positive because heat is being added to the reaction. Whereas exothermic reactions release heat to it's surroundings making it have a negative enthalpy. Enthalpy is the same thing as heat of combustion except heat of combustion shows the value of the energy released during a combustion reaction only. To theoretically calculate the enthalpy of a specific reaction you need to use the formula ΔH_{rxn} =

$\sum(nHm)_{products}$ - $\sum$ $(nHm)_{reactants}$. This does not include any measurements that can vary the overall enthalpy.

Calorimetry however requires careful measurements of a variety of things. Such as the exact volume of water, exact mass of everything, change in temperature, and heat capacity. Heat of combustion only measures exothermic change/ loss of potential energy from reactants to products. A way that we can measure the energy lost in an exothermic or combustion reaction is by using a calorimeter that bases its measurements on heat transfer. In a calorimetry experiment, set up the experiment so all of the energy from the reaction (the system) is exchanged with the surroundings. A calorimetry works when the substance within the calorimeter undergoes combustion, scientists can then measure the energy given off by combustion. Calorimetry is the idea that enthalpy of a reaction will change the temperature, which can be calculated as a change in kinetic energy (ΔQ). This uses the equation $\Delta Q = -mc\Delta T$. (These values will be shown in calculations.)

The reason we are using ethanol and butane specially for the combustion reactions to is determine which fuel is more effective because they are both the main indigents in bio-ethanol and gasoline. Fuels like butane (C_4H_{10}) theoretically produces quite a bit more energy than fuels like ethanol ($C_2H_6O_6$) because it is a bigger and has more carbons. Butane has a theoretical molar enthalpy of 2874 kJ/mol^{-1}, and ethanol has a theoretical molar enthalpy of 1368 kJ/mol^{-1} when it combusts (*AUS-e-TUTE, 2015*) Butane has a larger molar enthalpy than ethanol. The main reason for this is the combustion reaction and moles needed. When using the formula for reaction enthalpy ($H_{rxn}=\sum(H_f$ products)-$\sum(H_f$ reactants), the sum of the formation enthalpy of the products will be much higher, and the reaction will release much more energy (standard enthalpy of formation).

Bio-ethanol is a type of bio-fuel, which is also a type of biomass. The process of producing bio-fuels and biomass is very similar to other types of sources of energy in today's world: place the organic matter into a boiler where it burns and produces steam to turn a turbine, connected to a generator. Once the plant is burned then a new one is planted, creating a continuous cycle of sustainable and environment friendly energy. Bio-ethanol specially needs sugar syrups from plants like sugar cane or crops like grain using yeast for their carbohydrates so that with O_2 they can be part of fermentation then distillation/rectification and dehydration (*Vogelbusch, 2015*). Once this process is complete the bio-ethanol is added to specific amounts of gasoline to produce an alternative energy source in a fuel burning machine with usually very little modification needed.

Problem:

The purpose of this lab is to compare the heat combustions (enthalpy) of a two different fuel sources to determine which is the most effective at producing energy.

Hypothesis:

If the butane lighter heats used to heat the test tube with 20 mL of water in the calorimeter setup creates a greater temperature change with a 5g mass loss then the fuel ethanol used in a different trial it is because of the theoretical higher molar enthalpy of butane.

Theoretical Enthalpy of Butane

$C_4H_{10(l)} + 7O_{2(g)} \rightarrow 4CO_{2(g)} + 5H_2O_{(g)}$

-125.7 kJ/mol 0 -393.5 kJ/mol -241.8kJ/mol

$H_{rxn}=\sum(H_f \text{ products})-\sum(H_f \text{ reactants})$
H_{rxn}=(-393.5 kJ/mol x 4 mol + (-)241.8 kJ/mol x 5 mol) – (-125.7 kJ/mol x 1 mol)
H_{rxn}= (-1574(+) -1209 kJ) - (-125.7 kJ)
H_{rxn}= -2657.3 kJ per one mol

Theoretical Enthalpy of Ethanol

$C_2H_6O_{(l)} + 3O_{2(g)} \rightarrow 2CO_{2(g)} + 3H_2O_{(g)}$

-277.6 kJ/mol 0 -393.5kJ/mol -241.8kJ/mol

$H_{rxn}=\sum(H_f \text{ products})-\sum(H_f \text{ reactants})$
H_{rxn}=(-393.5 kJ/mol x 2 mol + (-)241.8 kJ/mol x 3 mol) – (-277.6 kJ/mol x 1 mol)
H_{rxn}= (-787 (+) -725.4 kJ) – (-277.6 kJ)
H_{rxn}= -1234.8 kJ per one mol

Variables:
Controlled:
1. Materials
2. SATP
3. Amount of Substances Being Burned (5g Being Lost)
4. Amount of Water (20mL)
5. Calorimeter Method (Setup of Calorimeter)

Manipulated:
1. Fuel Source Being Used to Heat the Water

Responding:
1. Temperature of Water Being Heated (Celsius)

Diagram:

Figure 1.1- Calorimeter Used to Burn Ethanol (C_2H_6O) Inside a Ethanol Burner and a Butane (C_4H_{10}) Lighter to Determine the Energy of the Reaction Based on Heat Transfer to the Water (20mL) in Test Tubes

Materials:

- Graduated Cylinder
- Electronic Balance
- Test Tube Clamp
- Matches
- Test Tube Rack
- Thermometer
- Ethanol Burner
- 2, 25 mm by 200 mm Test Tubes
- 40 mL of Distilled Water at Room at Temperature
- Thermometer Clamp
- Laboratory Stand
- Butane Lighter
- Tape (1m)

Procedure:

1. Use a graduated cylinder to fill each of the three test tubes with 20.0 mL of distilled water. Once you have filled the test tubes, place them in a test tube rack for later.

2. Use the electronic balance to measure the initial masses of each of the following: ethanol burner with ethanol inside and butane lighter. Record these values in the data table.

3. Assemble the apparatus by attaching a test tube filled with 20 mL of water to the laboratory stand 10 cm off the ground with tape and using the clamps of laboratory stand attach the thermometer. The bulb of the thermometer should not be touching the bottom or walls of the test tube. Measure the initial temperature of the water in the first test tube, and record it in the table.

4. Place the electronic balance under the test tube and place the ethanol burner with ethanol on top. Determine the desired final mass by adding 5 gram to the initial mass, and record it in the data table.

5. Ignite the ethanol burner, and quickly place it under the test tube so that the upper tip of the flame just touches the bottom of the test tube. Carefully monitor the mass of the fuel. Once it reaches the final desired mass quickly remove the ethanol burner and extinguish the flame.

6. Then read final temperature of thermometer and record in table.

7. Replace the test tube with another test tube from the test tube rack, and repeat steps 4 and 5 for butane lighter.

8. Use electronic balance to measure the final masses of each of the following: the alcohol burner containing ethanol, lighter containing butane, and the candle. Record these values in the data table.

Experimental Results:

Table 1.1- Quantitative Results of an Calorimeter Used to Burn Ethanol (C_2H_6O) Inside a Ethanol Burner and a Butane (C_4H_{10}) Lighter to Determine the Energy of the Reaction Based on Heat Transfer to the Water (20mL) in Test Tubes

	Experiment 1: Ethanol in Ethanol Burner (C_2H_6O)	**Experiment 2: Butane in Butane Lighter (C_4H_{10})**
Initial Mass (g)	110.6	30.5
Final mass (g)	105.5	25.5

Change in mass (g)	5	5
Initial Temperature of Water (Celsius)	22	22
Final Temperature of Water (Celsius)	85	54
Change in Temperature of Water (Celsius)	63	32
Volume of Water in Calorimeter (mL)	20	20

Calculations:

- Combustion of Ethanol: $C_2H_6O_{(l)} + 3O_{2(g)} \rightarrow 2CO_{2(g)} + 3H_2O_{(g)}$

  $\Delta Q = -mc\Delta T$

  $= (5g)(4.19J(gC)(-63C)$

  $= -1319.85J$

  $= 1.31985kJ$ per one mole

- Combustion of Butane: $C_4H_{10(l)} + 7O_{2(g)} \rightarrow 4CO_{2(g)} + 5H_2O_{(g)}$

  $\Delta Q = -mc\Delta T$

  $= (5g)(4.19J(gC)(-32C)$

  $= -670.4J$

  $= 0.6704kJ$ per one mole

Analysis of Results:
For experiment one, both the theoretical enthalpy was much higher than the experimental enthalpy, about 1200 times bigger than the other. For experiment two the theoretical was quite a bit higher than the experimental, approximately 2100 times bigger. This is because of the second law of thermodynamics, which states that no reaction is 100% efficient at converting energy, because energy will always be wasted as work. The second law also states that heat will flow from one system (the classroom) to another system (the calorimeter) until the two are in equilibrium (about 22 C). The types of chemicals that reacted in the calorimeter will effect what type of reaction happens in the calorimeter because of the formation enthalpies of

those chemicals. The efficiency of the calorimeter will affect how much heat remains within the system of the calorimeter, and how much is transferred to or from the environment outside of the calorimeter.

Conclusion:
The purpose of this lab was to compare the heat combustions (enthalpy) of a two different fuel sources to determine which is the most effective at producing energy. The efficiency of the fuels should have an effect on the enthalpy that the system produces in the combustion reaction. The calorimeter was set up with the ethanol burner and butane lighter to heat 20 mL and loss 5g of mass. These are the manipulated and controlled variables respectively and the responding variable is the temperature to determine the enthalpy/energy produced. Experiment 1with the ethanol burner and ethanol inside had an experimental enthalpy of 1.31985kJ, and the theoretical value of the enthalpy was -1234.8 kJ. Experiment 2 with the butane had an experimental enthalpy of 0.6704kJ, and a theoretical enthalpy of -2657.3 kJ. My hypothesis was incorrect but the space for error pretty much discounts the entire lab because butane would have a higher enthalpy and that explains why it's the main ingredient in gasoline whereas ethanol is the main ingredient in bio-ethanol which is an alternative source for many types of machines. If my experiment would prove that ethanol is a very useful source of energy and should be used everywhere. The errors in this experiment included loss of heat in the open system calorimeter, assuming the exact loss of weigh of fuels is accurate to the substance inside and the inaccuracy of measuring distance between the can and the flame produced by the fuels.

Extension:
A real life application of the reaction used in the experiment regarding ethanol is related to the two main bio-fuels, which are bio-ethanol and bio-diesel. Ethanol is currently produced from sugar crops such as sugar beet and sugarcane, or starch crops such as corn and wheat. Bio-diesel is made from plant oils such as soybean oil, palm oil, rapeseed oil as well as animal fast. Bio-diesel is an alternative fuel to petroleum based diesel fuel which according to the U.S. Energy Information Administration is used in trucks, trains, boats and barges to transport good as well as in most farm and construction equipment. Bio-ethanol could be used for the exact same kinds of of transport as long as it was blended with regular gasoline. In fact the U.S. sold 36,343,072,000 gallons of diesel in 2012. So it's clear to see in a very small number of years based on the rate of use the U.S. and world will run out of fossil fuels (non-renewable). Bio-fuels are gaining attention in the U.S. and is already quite successful in Europe. Not only are they renewable but also non-toxic and biodegradable. So very much like bio-diesel, bio-ethanol could be used as an alternative fuel source in daily life and help reduce all kinds of problems we have in today's day and age. This lab is very applicable to real life because we can view how different butane based gasoline and bio-ethanol are and what the advantages and disadvantages of each include. A new lab investigation could be done based on this information about how much mileage you could get from a tank of bio-ethanol vs

gasoline and comparing the fuel prices of each. However the point of bio-fuels in based more on the eco-friendly and responsible way of transport and is not so short sighted.

Error Analysis:

1. A possible error was when heat being released in the reaction does not go directly to the water but around and away. This is because this experiment was done in an open system, meaning the heat exchange of the reaction also occurred with the air and not the water, this leads to a loss in energy and decreases the accuracy of experimental values. This influenced the results by changing the temperature change measured in the lab because the energy produced was not effective in its purpose. A scientific way to fix this error is by performing the lab in a closed system like that of a bomb calorimeter.
2. Another possible error is assuming the exact loss of weigh of fuels is accurate to the substance because of the electrical balance scale. When combustion is occurring the fuel being burned could also be losing weight to burning of the burner or lighter they are in. This would affect the results by making the experimental enthalpy lower because the heat was again not going directly to heating the water also the values would be incorrect because the "container" the fuels were in each have a different heat capacity. A way to fix this error would have been to obtain the substances separately and place them in the exact same space and the temperature change of just the fuel can be measured.
3. A final error that was the inaccuracy of measuring distance between the can and the flame produced by the fuels. There was no exact measurement required or recorded in the experiment therefore the flame could have been very close to or far away from the test tube. The distances being so varied could effect the results by changing the temperature of the water because of the heat lost to the surroundings. Specific changes could be made to the lab steps or multiple trials could be concluded with the flame being closer or farther from the test tube and finding the middle ground for accurate results.

$$\% \text{ error} = \frac{\text{theoretical value} - \text{experimental value}}{\text{theoretical value}} \times 100\%$$

$$\% \text{ error for Butane} = \frac{-2657.3 \text{ kJ/mol} - -0.6704\text{kJ/mol}}{-2657.3 \text{ kJ/mol}} \times 100\%$$

$$= 99.97\%$$

$$\% \text{ error for Ethanol} = \frac{-1234.8 \text{ kJ/mol} - -1.31985\text{kJ/mol}}{} \times 100\%$$

-1234.8 kJ/mol
= 99.89%

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