**Fossil Fuel and Biofuel Combustion Comparison**

*Adapted from Fossil Fuel & Biofuel Combustion activity by C. Kohn, Waterford Union High School, Waterford WI*

**Objective:**

The purpose of this activity is to compare fuels based on their energy production during combustion and the amounts of carbon dioxide produced.

**Introduction:**

A large percentage of energy in the United States is produced by fossil fuel sources (coal and oil). However, these resources are limited and biofuels (ethanol and biodiesel) are now being studied in hopes of finding a suitable replacement. Both fossil fuels and biofuels provide energy by breaking high energy bonds during combustion reactions. Combustion reactions are exothermic, which means the energy of the products is less than the energy of the reactants. What makes a fuel source valuable is that it releases more energy than was used to produce it. Some other considerations for alternative fuel sources include human health and environmental impacts and supply and availability for our needs.

* Do the by-products of the combustion reaction produce pollutants?
* How much carbon dioxide and carbon monoxide are contributing to greenhouse gas levels?
* How long can the United States produce this fuel and will it ever run out?
* How available is the starting material and can it meet our fuel needs in the future?

In this lab, we will be considering four types of fuel – Octane (the major component of gasoline), ethanol (made from fermented plant matter, usually corn in the US), methane (a major component of natural gas and the biofuel from decomposing landfill garbage), and the biodiesel prepared in lab. Biodiesel, like gasoline, is a mixture of compounds. Data from the National Biodiesel Board indicates the major component of biodiesel is linoleic acid methyl ester. This is the compound we will use to compare to the other fuels.

**Procedure**

1. Make a hypothesis before you begin. Which of the four fuels would be the best? Remember to include your reasoning behind your choice.
2. Using the materials at your desk, first build a methane molecule. The formula for methane is CH4. The carbon is the central atom and it is covalently bonded to four hydrogen molecules.
3. Next, build ethanol. The formula for ethanol is C2H6O. Ethanol is an alcohol, formed by a carbon backbone, five covalently bonded hydrogen atoms, and one hydroxyl (-OH) group. Begin by bonding the two carbon atoms together. Add the five hydrogen atoms to the carbon with single bonds. Finally, bond an oxygen atom to a carbon, and then bond the hydrogen to the oxygen.
4. Draw and label the molecules once you have them built. Raise your hand so the teacher can check your structures before you go any further.
5. When a fuel burns, this represents a combustion reaction. What is necessary for combustion? (Hint: think about what a combustion reaction looks like)
6. Now build two oxygen molecules. Remember oxygen is a diatomic molecule, so its formula is O2 and it contains a double covalent bond. Draw and label one of the oxygen molecules below.
7. Now that the fuel and the oxygen required for combustion are present, perform the combustion reaction for methane by rearranging your reactants into products. The products of a combustion reaction are carbon dioxide (CO2) and water (H2O). Some energy is required to break the C-H bonds in the methane and the double bonded oxygen as products are formed. Once you have made models of the carbon dioxide and water, balance the chemical reaction below:

\_\_\_\_\_CH4 + \_\_\_\_\_O2 🡪 \_\_\_\_\_CO2 + \_\_\_\_\_H2O

1. Methane has 1656 kilojoules per mole of energy stored in its chemical bonds. The double bond in an oxygen molecule has 498 kJ/mol. Remember that energy is required to break these bonds. After combustion, you formed carbon dioxide and water. The energy stored in these molecules is 1606.6 kJ/mol for CO2 and 928 kJ/mol for water. We need to calculate the net amount of energy released during this reaction. Complete the following table. Notice that the oxygen and water values are multiplied by 2 because of the coefficients you used when you balanced the equation in #7. To get the total reactant energy, add the energy values for the reactants. Do the same for the products.

|  |  |  |  |
| --- | --- | --- | --- |
| **Reactants** | **Bond Energy (kJ/mol)** | **Products** | **Bond Energy (kJ/mol)** |
| Methane |  | Carbon dioxide |  |
| O2 (x2) |  | H2O (x2) |  |
| TOTAL REACTANTS ENERGY |  | TOTAL PRODUCTS ENERGY |  |

NET ENERGY RELEASED = TOTAL PRODUCTS ENERGY – TOTAL REACTANTS ENERGY

NET ENERGY FOR METHANE \_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_

1. Now perform the same exercise with your ethanol molecule. For this reaction, create several oxygen molecules to react with your one ethanol. You may use as many oxygen molecules as necessary, but only react one ethanol molecule. Perform this combustion reaction by rearranging the atoms to form carbon dioxide and water again. Balance the chemical reaction once you have completed the models.

\_\_\_\_\_C2H6O + \_\_\_\_\_O2 🡪 \_\_\_\_\_CO2 + \_\_\_\_\_H2O

1. Ethanol has 3241 kilojoules per mole of energy stored in its chemical bonds. The oxygen molecule has 498 kJ/mol (remember you have multiple oxygen molecules in your reaction). CO2 has 1606.6 kJ/mol and water has a bond energy of 928 kJ/mol. Complete the following table using this information and the coefficients from the balanced equation. ***Remember, as with the methane, if multiple molecules of a substance were needed or produced, you will have to multiply its bond energy!***

|  |  |  |  |
| --- | --- | --- | --- |
| **Reactants** | **Bond Energy (kJ/mol)** | **Products** | **Bond Energy (kJ/mol)** |
| Ethanol |  | Carbon dioxide |  |
| O2 |  | H2O |  |
| TOTAL REACTANTS ENERGY |  | TOTAL PRODUCTS ENERGY |  |

NET ENERGY RELEASED = TOTAL PRODUCTS ENERGY – TOTAL REACTANTS ENERGY

NET ENERGY FOR ETHANOL \_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_

1. The next molecule you will be creating is octane (C8H18). Create this molecule with the materials at your table. Octane has a long chain carbon backbone with single bonded hydrogen atoms to the carbon atoms. Once you have this molecule built, let the teacher check your structure before you go any further.

1. When octane combusts in our car engines, it undergoes a similar reaction to the ones you have modeled for methane and ethanol. The octane reacts with oxygen to form carbon dioxide and water. Write the balanced chemical equation for the combustion of octane.
2. Octane has 9881 kilojoules per mole of energy stored in its chemical bonds. The oxygen molecule has 498 kJ/mol (remember you have multiple oxygen molecules in your reaction). CO2 has 1606.6 kJ/mol and water has a bond energy of 928 kJ/mol. Complete the following table using this information and the coefficients from the balanced equation. ***Remember, as with the methane, if multiple molecules of a substance were needed or produced, you will have to multiply its bond energy!***

|  |  |  |  |
| --- | --- | --- | --- |
| **Reactants** | **Bond Energy (kJ/mol)** | **Products** | **Bond Energy (kJ/mol)** |
| Octane |  | Carbon dioxide |  |
| O2 |  | H2O |  |
| TOTAL REACTANTS ENERGY |  | TOTAL PRODUCTS ENERGY |  |

NET ENERGY RELEASED =TOTAL PRODUCTS ENERGY – TOTAL REACTANTS ENERGY

NET ENERGY FOR OCTANE \_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_

1. The final molecule we will consider is the biodiesel prepared in lab. The balanced chemical equation is below. We will assume the biodiesel is composed of linoleic acid methyl ester (C19H34O2). The heat content of biodiesel is 21979.94 kJ/mol. The oxygen molecule has 498 kJ/mol (remember you have multiple oxygen molecules in your reaction). CO2 has 1606.6 kJ/mol and water has a bond energy of 928 kJ/mol. Complete the following table using this information and the coefficients from the balanced equation. Remember, as with the methane, if multiple molecules of a substance were needed or produced, you will have to multiply its bond energy!

2C19H34O2 + 53O2 🡪 38CO2 + 34H2O

|  |  |  |  |
| --- | --- | --- | --- |
| **Reactants** | **Bond Energy (kJ/mol)** | **Products** | **Bond Energy (kJ/mol)** |
| Biodiesel |  | Carbon dioxide |  |
| O2 |  | H2O |  |
| TOTAL REACTANTS ENERGY |  | TOTAL PRODUCTS ENERGY |  |

NET ENERGY RELEASED = TOTAL PRODUCTS ENERGY – TOTAL REACTANTS ENERGY

NET ENERGY FOR BIODIESEL \_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_

1. Is the combustion of a fuel endothermic or exothermic? Explain how you know.
2. Using your calculations, list the fuels in decreasing order (greatest to least) in terms of net energy released during combustion. Indicate how many moles of each fuel were used in the calculation by looking at the coefficients from the balanced equations.
3. Using the balanced equations, list the fuels in decreasing order in terms of carbon dioxide released during combustion. Indicate how many moles of each fuel were used in the calculation by looking at the coefficients from the balanced equations.
4. Which do you think is the best fuel? Evaluate in terms of energy released as well as CO2 production.
5. The US Energy Information Administration estimates that U.S. gasoline and diesel fuel consumption for transportation in 2012 resulted in the emission of about 1,089 and 422 million metric tons of CO2, respectively. Calculate the pounds of CO2 produced from each fuel. (1 metric ton = 1000 kg = 2204.6 pounds)
6. About 22.38 pounds of CO2 are produced by burning a gallon of diesel fuel. Burning a gallon of “B10” (diesel fuel containing 10% biodiesel by volume) results in emission of about 20 pounds of CO2. If the US uses 4 million barrels of diesel per day, how many pounds of CO2 could be eliminated *each year* if we converted to B10? (1 barrel = 42 gallons) From #19, what percentage of total diesel CO2 production would this represent?
7. Evaluate biodiesel as a green fuel source. Provide at least one argument for its use and one argument against its use. You are not limited to lab data to answer this question.
8. What would your recommendation be to President Obama as the best fuel source for the United States to move towards mass production of?