Fossil Fuel & Biofuel Combustion by C Kohn, Waterford WI

Partner Names: Hour Date

Date Due: Why late? Score: + ✓ -

**Directions**: With your partner at your desk, complete the activities and questions below. NOTE: It is EXTREMELY important not to lose any pieces. Be sure that all parts of your molecules are returned to their plastic bags at the end of this lab. Please be courteous and completely disassemble your molecules before ending the lab.

**Objective**: Determine which fuels are best for combustion based on their molecular rearrangement and bond energy changes and consideration for enthalpy and Hess’s law.

**Overview**: 85% of energy in the United States is provided by non-renewable fossil fuels. However, biofuels such as ethanol and biodiesel are increasingly meeting a greater percentage of our nation’s energy needs. Both kinds of fuel provide energy in the same way – the rearrangement of high-energy molecules (such as ethanol or octane) into low-energy molecules (such as carbon dioxide and water) causes energy to be “leftover” from each reaction. This “leftover energy” (or **enthalpy**) is released as excess energy in **exothermic reactions**. These “rearrangement reactions” that occur with oxygen and release energy are known as combustion.

**The more energy that is released during combustion, the higher the value of that fuel.** The change in energy to a system because of a reaction is known as **enthalpy**. Other considerations include –

* By-products of combustion, particularly pollutants and particulates that can affect human and environmental health. Particularly of concern are greenhouse gasses such as carbon dioxide, and incompletely broken down carbon molecules from combustion (soot, smog, and carbon monoxide)
* Stability of fuel, and rate of combustion (smaller molecules combust more easily than large molecules)
* Sustainability of the fuel – for how long can we use this fuel before it runs out or negatively impacts society?
* Availability of the fuel – how easily and widely available is the fuel for mass consumption? Is there enough of the fuel to meet society’s needs?

In this lab, we will be considering 3 different kinds of fuel – octane (the major component of gasoline), ethanol (made from fermented plant matter), and methane (a major component of both the fossil fuel-based natural gas and of renewable biogas from manure and decomposing organic matter).

1. To begin, make sure that you have enough materials to create your molecules. In your kits, you should have the following (or equivalent to the following; your instructor may provide you with different materials):
	1. 10 carbon atoms (black sphere with 4 pegs)
	2. 20 hydrogen atoms (white with one peg)
	3. 26 oxygen atoms (blue with two pegs)
	4. 45 covalent bonds (white tubes)
2. Next, hypothesize which of the three (octane, ethanol, or methane), would make the best fuel:

I think that will prove to be the best fuel. I think this because
3. Your first task will be to create the methane and ethanol (we’ll leave octane for later because it is a more complicated molecule).
	1. To create methane (CH4), you will need one carbon and four hydrogen atoms. Be sure to use the white tubes (or equivalent) to form the covalent bonds between carbon and hydrogen.
	2. To create ethanol (C2H6O), you will need to combine two carbons, six hydrogens, and one oxygen atom.
4. Raise your hand when finished so that your instructor can check your molecules. After they approve, draw both of your molecules in the space below. Instructor Initials:

Methane Ethanol
5. Now we will model combustion. In the space below, write down everything that you think is necessary for something to combust (or burn).
6. In this particular example, we need three things – 1) the combustible substance, either methane or ethanol, 2) oxygen, and 3) the initial energy to break the bonds of the carbon molecules (a “spark” if you will).

In this case, you already have your combustible molecules, methane and ethanol. We will now add oxygen. In the air, oxygen mostly exists as O2. Create an O2 molecule. Notice that O2 has to have a double bond (we never want open “pegs” or binding sites). Create two O2 molecules. Draw them below
7. We now have two of the three things we need for combustion – our combustible substances and our oxygen. The last thing we need is a “spark”, or the energy needed to break the bonds of our products so that they can be recombined into different molecules. The spark in this case will be … you! Take apart your methane molecule and your O2 molecules. Recombine them into CO­2 and H2O molecules. Complete the chemical reaction below:

 1 CH4 + 2 O2 🡺 \_\_\_\_ CO2 + \_\_\_\_H2O
8. Methane has 1,656.8 kilojoules per mole of energy stored in its chemical bonds (a kilojoule, or kJ, is just a way to measure energy, just like pounds are a way to measure weight). O­2 has 494.9 kJ/mol in its bonds (keep in mind you have two O2 molecules in your reaction, so you will have to multiply this amount by 2). After combustion occurred, you had one molecule of CO­2­ with 1606.6 kJ/mol and 2 molecules of water with 928.8 kJ/mol.
	1. Using the information above, complete the following table:

	Reactants Bond Energy Products Bond Energy

	Methane kJ/mol Carbon Dioxide kJ/mol

	O2 (x2) kJ/mol H2O (x2) kJ/mol

	TOTAL kJ/mol TOTAL kJ/mol
	2. Subtract your Product’s bond energy from your Reactant’s bond energy. ENERGY RELEASED:

- **Note** – energy is released when bonds are formed; your Products should have a higher energy total.  **- Enthalpy** = ∆H = energy released from a reaction.
 **-** This reaction involves **Hess’s Law** which states that the enthalpy from a reaction is equal to the sum of the individual components of that reaction.

- ∆**H** = energy of bonds broken minus the energy of bonds made. If you have energy released, you have negative change to the enthalpy (∆H) of a reaction.

1. Now perform the same exercise with your ethanol molecule. For this, create 5 O2 molecules for your 1 ethanol molecule. Model this reaction by breaking apart your ethanol and O2 and recombining it to make only CO2 and H2O.
 1 C2H6O + 3 O2 🡺 \_\_\_\_ CO2 + \_\_\_\_H2O
2. Ethanol has 3,240.4 kilojoules per mole of energy stored in its chemical bonds. O­2 has 494.9 kJ/mol in its bonds (keep in mind you have multiple O2 molecules in your reaction). CO­2­ has a bond energy of 1606.6 kJ/mol and water has a bond energy of 928.8 kJ/mol.
	1. Using the information above, complete the following table (note – if multiple molecules of a substances were needed or produced, you will have to multiply its bond energy):

	Reactants Bond Energy Products Bond Energy

	Ethanol kJ/mol CO2 x 2 kJ/mol

	O2 x3 kJ/mol H2O x3 kJ/mol

	TOTAL kJ/mol TOTAL kJ/mol
3. Subtract your Product’s bond energy from your Reactant’s bond energy. Then compare your result to Methane.

ENERGY RELEASED FROM COMBUSTION OF ETHANOL: CO2 molecules produced:

ENERGY RELEASED FROM COMBUSTION OF METHANE: CO2 molecules produced:
4. The last molecule you will be creating is Octane (C8H18). Create this structure from your disassembled molecules from earlier steps. When you feel you have created an accurate model of octane, raise your hand so that your instructor can inspect your work. Instructor Initials:
5. With a partner, predict how many O2 molecules will be needed to combust octane. Write your estimate below:

*We think that molecules of O2 will be needed to combust one molecule of octane. We think

this because…*
6. “Combust” your octane molecule by rearranging octane and O2 into CO2 and H2O until no more pieces remain.

Was your estimate for the amount of O2 needed correct? Explain:
7. Octane has 9,886.7 kilojoules per mole of energy stored in its chemical. O­2 has 494.9 kJ/mol in its bonds (keep in mind you have multiple O2 molecules in your reaction). CO­2­ has a bond energy of 1606.6 kJ/mol and water has a bond energy of 928.8 kJ/mol.
	1. Using the information above, complete the following table (note – if multiple molecules of a substances were needed or produced, you will have to multiply its bond energy):

	Reactants Bond Energy Products Bond Energy

	Octane kJ/mol CO2 x 8 kJ/mol

	O2 x 13 kJ/mol H2O x 9 kJ/mol

	TOTAL kJ/mol TOTAL kJ/mol
8. Subtract your Product’s bond energy from your Reactant’s bond energy. Compare all three fuels.

ENERGY RELEASED FROM COMBUSTION OF ETHANOL: CO2 molecules produced:

ENERGY RELEASED FROM COMBUSTION OF METHANE: CO2 molecules produced:

ENERGY RELEASED FROM COMBUSTION OF OCTANE: CO2 molecules produced:

1. List the 3 fuels from greatest to least in terms of energy released during combustion (include energy released):

*Greatest* *Least*
2. List the 3 fuels from greatest to least in terms of CO2 released during combustion (include molecules of CO2 released per molecule of fuel combusted)

*Greatest* *Least*
3. Which do you think is the best fuel? Explain:
4. Some aspects not covered by this activity include availability of supply, ease of use, pollutants, and CO2 removed or added during production (e.g. ethanol production removes CO2 through photosynthesis when its sources are growing but this is not reflected here). How might this affect your answer to the question above? Explain: