*How Hot is Too Hot? : Using Thermodynamics to Study Fuels*

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South Mecklenburg High School

This curriculum unit is recommended for:

Chemistry (any level)
Physical Science

**Keywords:** Thermodynamics, Energy, Enthalpy

**Teaching Standards:** See Appendix 1 for teaching standards addressed in this unit.

**Synopsis:** This curriculum unit is designed to help high school students relate more closely with the topic of thermodynamics by working with fuels and doing an inquiry lab with fabrics. The two big connections are quantifying the energy that is stored in certain fuels in relation to things that students can grasp. An activity demonstrating such a connection is an inquiry lab about the safety of fabrics.

*I plan to teach this unit during the coming semester to* approximately 80 students in standard chemistry.

Introduction

Thermochemistry is an area of chemistry that gets a lot of lecture time, but not much lab or in-depth study due to the difficult nature of the topic. In chemistry, the topic of thermo normally rests after midterm exams and about halfway through the second quarter; so it is natural for the students to become easily frustrated with the topics, as well as the teachers, since there is so much to cover. Normally this unit is done as an actual unit but in reviewing the essential standards for chemistry, practically all of objective two and three are devoted to energy changes (i.e. thermochemistry). This means that this 'unit' of thermochemistry actually needs to be integrated throughout the entire chemistry course.

My goal with this curriculum unit is to integrate thermodynamics, energy, and chemistry into everyday items we use and with the varied reactions that can be seen outside of the chemistry classroom. We will be taking a focus on globalization and the economy so we will also be looking at the safety of some fabrics that are supposed to be very fire safe. This will be done in a qualitative, inquiry based way using some materials that Fire/EMS use as well as baby clothes and blankets found in stores. Some demonstrations and/or labs that will be performed will be finding the internal temperature of different classes of fires by looking at the starting fuel source (qualitative only). Teaching on a high school level allows the teacher to show demonstrations like the thermite reaction in order to get the students thinking about energy, the high temperatures that can be produced, and how fast a reaction can take place. This will be a good introduction into how extremely exothermic (large heat release) a reaction can actually behave. These reactions will also allow students to see what can happen when different fuel sources combust and how the heat and elevated temperatures can directly impact how fast it spreads, as well as its potential to hurt us. This curriculum unit will work with many types of sciences including, but not limited to: chemistry, medical science, fire science, biology and apparel/design.

I teach chemistry at South Mecklenburg High School in Charlotte, North Carolina, which has a student population of 44% White; 28% African-American; 22% Hispanic; 5% Asian and 1% other. Forty-three percent of our student-body receives free and reduced lunch and 7% of our student body is classified as limited English proficient.

My target audience will be students in grades tenth and eleventh enrolled in either honors or standard chemistry; however this curriculum unit can be modified up or down grade levels depending on the students' understanding of chemistry. This unit could be chunked into labs or demonstrations that could be useful in teaching other concepts within the broad topic of thermodynamics. Students in the standard class would most likely need more time and help learning and integrating prior concepts in this unit, but it is certainly not out of the question; it would just mean this curriculum unit would be longer for standard chemistry, which makes sense because this thermodynamic unit would most likely be spread out and integrated throughout a three week period of time anyway.  If this curriculum unit is taught to an honors or pre-AP (Advanced Placement) chemistry class, the integration and assumption of prior knowledge will be strong. Students will have to integrate thermodynamics and how electrons move seamlessly with knowledge acquired in the pre-AP chemistry class.  The calculations will include enthalpy, Gibbs free energy and some entropy.  If this curriculum unit is taught to a standard chemistry class, the amount of actual calculations would probably only be limited to enthalpy of reactants, products, and the overall reaction.  From this we should be able to make a connection to how the reaction will or will not have the desire result based on if the result of the calculation is either positive or negative $\Delta H$ (enthalpy or heat of reaction).  The discussions of Gibbs free energy and entropy would certainly happen, but extensive calculations of these quantities would not.  These calculations could be used for our standard 'high flyer' students.  Additional calculations could be included as a way to differentiate instruction from our 'super high flyers' and our 'high flyers.'

In conclusion, thermodynamics is used throughout the chemistry curriculum and students need to determine how and why it is important to understand the energy behind reactions.  This curriculum unit can be linked to many other areas of science and even into some electives.  The study of electrons when they bond and rearrange are why chemical reactions take place, so some energy has to move when this happens.  This, in part, is what thermodynamics is all about; the study of the electrons moving around during reactions and how energy is used, released or transferred in those apparent reactions.

Strategies:

Some possible directions for this curriculum unit can include the introduction to thermodynamics by doing a fun lab of burning magnesium ribbon.  After the introduction and some other useful thermodynamic labs, a basic calorimetry lab could be done in which they can make a simple calorimeter and determine the energy of fuels.  This would give students a great introduction into figuring out which fuel has the potential to produce the most energy.  Students could then determine how to make this lab more effective by completing simple calculations to compare their experimental data with their calculated data.  This exercise would give students the basis for writing a lab report with multiple trials going on.  North Carolina and Charlotte Mecklenburg schools most likely will allow all of the reactions in this curriculum unit to be done inside a chemistry/science classroom except the thermite reaction.  The easy thermite demonstration uses basic laboratory equipment and chemicals that are allowed on a high school level.  All of the materials for these labs and demos should be able to be found in a high school science lab or bought from the store.  For the globalization and 21st century skills link to this unit, students could find different types of materials and test them by making small fires and determining the heat/energy released from the fire (this might have to be more on a qualitative level instead of quantative). Students could also test different cloths as a function of the heat generated by different chemical fuels. Some that could possibly be used would be a fireman's jacket and the ear flaps (off the back of the firefighter's helmet) just to name a few. This would also be a good way to link into different careers and how the chemistry behind something common is used in the real world. These experiments/reactions would need to be investigated by the instructor first, before the students go at it on their own.

## Content Objectives:

Within this curriculum unit, objective two and three are discussed heavily. This is due to the fact that most of the thermodynamic study is listed in these two sections of the North Carolina Essential standards. A full list of the objectives from this curriculum unit can be found in Appendix 1.

## Unit Content:

This unit would most likely fit into most units of your semester/year-long plan. However I will be emphasizing the thermodynamic study of chemistry, but as I stated earlier, everything we do has to do with energy so the topic is always relevant. These would be great activities to get students excited about thermochemistry and then bring in the math and other concepts. These can be used as "touch stones"[4] or pieces of chemistry that you continually go back to and reference so the students have a broad understanding that everything is connected.

## Background Information/Teaching Strategies

Thermochemistry is the branch of chemistry which deals with only the macroscopic quantities, which means it is not a molecular branch of chemistry. What scientists look for in this area of chemistry is the exchange of energy between the system and the surroundings during a physical or chemical process. The 'system' is the part of the universe which is under investigation in an experiment. In contrast, the 'surroundings' are all the other objects that may act on the system. In working with this curriculum unit, we will try to find a way to link the content of thermochemistry to what students have already done; which will provide a solid framework for them to build upon what is happening in the reactions, and in turn, the energy of those reactions when they take place.  "Thermodynamics calls attention to the two macroscopic properties which are most fundamentally responsible for the behavior of matter".[1]  These properties are energy and entropy, which lead us to our two basic laws.  Law 1 states 'the energy of the universe is conserved' and law 2 states 'the entropy of the universe increases'.  In thermochemistry, the students need to understand where that energy goes and how it gets there from the use of enthalpy, entropy, Gibbs free energy and other calculations.  The Law of Conservation of Energy is one that students are generally familiar with and the two thermodynamic laws help to define and quantify the energy.

If we analyze the Law of Conservation of Energy further, we realize that energy can change form.  Every time we breathe, eat, walk and just plain live, reactions are going on and energy is changing form.  To most of us, we do not realize it or think on it during the day; if air is going in and out and blood goes round and round in our bodies, we are good!  If we allow the students to understand why this happens and how much our life relies on the numerous thermochemical reactions we take for granted, their excitement level will increase and they will be able to build upon that framework that I pointed out earlier in the introduction.  The study of chemistry is always building on itself, and we are gaining a greater understanding about what is happening on a molecular level; to be honest, it boils down to the motion and activity of those tiny electrons we cannot even see.  This is how reactions take place, by transferring, stealing and/or sharing these electrons during chemical reactions.  The by-product of such action at the molecular level dictates the produced or released energy multiplied by astronomical numbers (on the order of $10^{23}$ values), allowing us to see what is happening at the macroscopic and thermodynamic scale.

So in essence, everything we do has energy.  "It is the study of thermodynamics that allows us to determine if that reaction will be exothermic (release heat/energy) or endothermic (absorb heat/energy)."[2]  In this curriculum unit, we will be analyzing different fuel sources for their energy content and relating this to the theme of "how hot is too hot".  We will use cross curricular ideas to bring in different types of fabrics to see what will burn at what approximate temperatures (this will be mainly qualitative since the equipment needed to determine the exact temperature is not available).

The beginning experiment that students will do is to complete a magnesium ribbon burn.  This experiment is utilized so that students can see a large release of heat by way of light.  If students look at the elements magnesium and oxygen, they do not immediately react with each other, but we can discuss that thermodynamically, the reaction really wants to go; which means it has a negative free energy.  Once the magnesium molecules are heated enough by the oxygen (by way of the burner), the electrons are so excited that they start to bond with the oxygen.  This allows the new compound, magnesium oxide, to be produced.

Students will also witness the "whoosh bottle" in which we will move ethanol (ethyl alcohol) into the gaseous state by swishing/moving it around in the bottle.  After approximately five minutes or so, we will dump all the alcohol out and light the top gas on fire. This will cause a somewhat expected reaction, a fire ball, which is an extremely exothermic reaction. You must be very careful with this experiment, because you do not want to place the bottle on top of a table, because you might burn the ceiling. Also make sure it is on either a low chair or the floor. Students will be able to feel the bottle and notice that it is warm to the touch, giving them qualitative information that the reaction was indeed exothermic and allowing us to talk about the combustion reaction that takes place. Students should recognize this type of reaction because of the water that is produced.[2] This demonstration could further be done by potentially putting additional fuels (safe ones…methanol, butanol) into the whoosh bottle and seeing if the visual energy output would be the same. This is a wonderful 21st century skill where students can analyze and compare the different fuels to understand what is happening. This demonstration will also give students an idea on how hot the combustion of certain fuels can be. A great addition to this demonstration is to add certain kinds of salts to make the flame a different color. Borax is wonderful for turning the flame a brilliant green color.[2] You will have to dissolve the borax into water and let it dry in order for it to stick to the sides of the bottle.

The next experiment that would be conducted will be to do a simple synthesis reaction of an element and study the enthalpy of this reaction as we begin to delve into the quantitative nature of thermodynamics. Students would again light a piece of magnesium ribbon and collect the product, magnesium oxide. They would start to investigate and discuss if they thought this was exothermic, give reasons why they might think this, and also start to introduce the math behind this thermodynamic study. The math behind this quick experiment is to understand that all elements have an enthalpy of zero because they come from the Earth. Producing magnesium oxide is an incredibly exothermic reaction, but the real question is; does this reaction want to go? This is where we can introduce positive and negative change in enthalpy ($\Delta H_{rxn}$) reaction values and start to really see the power of thermodynamics in predicting everyday reactions and life processes. Students will learn that the quantity of energy given off, the $\Delta H$, corresponds to the energy change when going from the higher energy pure magnesium metal and oxygen gas to the more stable and lower energy form we now call magnesium oxide. This lab is a long one and will need to be done during two lab periods most likely. This should include discussions and data analysis; which also makes this a great lab to teach the students the correct method of writing a full blown lab report. I have done this during multiple enrichment periods and students get a more in depth understanding of why they did the lab.

The next lab that students would do in their study of thermodynamics is a calorimetric one. Students will make their own calorimeter out of a soda can. This is not the best method, but it is the easiest and the students can manipulate the mechanics of their home-made systems, and it is a fairly cheap laboratory setup. Students would be using the difference in water temperature which is inside the can calorimeter while taking a difference in the mass of the fuel needed to produce the temperature change in order to determine the total change in enthalpy. The energy number that students get will be in Joules and eventually converted into kilojoules. The analysis depends on using the specific heat of water to find the kJ of energy produced by our fuels. For a quick example, if we compare ethanol and gasoline, 1 gram of ethanol produces about 11.76 kJ of energy and 1 gram of gasoline gives about 19.39 kJ of energy. The values needed to complete the calculation can be found in a thermodynamic table.[3] This will probably not mean a lot to our students, but if we tell them that based on their results, we can now figure the if we used ethanol in our cars we can get 5 mi/gal, but we can get approximately 10 mi/gal with the gasoline by looking at the ratio of their energy outputs per gram, all of a sudden this lab becomes important to them, because they want to save money and drive. During the lab, you might also point out that while burning the ethanol and gasoline, students will see that ethanol will burn cleaner (less soot) then gasoline, but still, the gasoline stores more energy per gram. This all boils down to the fact that the "larger the molecule the larger the energy output is."[4], but having oxygen in the molecule helps it burn cleaner. Pointing out that the octane rating at the gas pump is usually directly related to the amount of oxygen containing molecules in the fuel (the more oxygen in the molecules the higher the octane rating) again links a common experience to the chemistry.

<u>Classroom Activities:</u>

*All of these activities must be done with goggles on in order to stay safe!

**Activity 1 - Magnesium ribbon burn**

<u>Materials:</u>

Evaporating dish or watch glass (something to catch the magnesium oxide, MgO, in)

Magnesium Ribbon, cut into about 3-4 in segments

Fire source (Bunsen burner is best, but a BIC lighter can work as well)

Water (a few drops to add to the watch glass or evaporating dish to test for acid/base)

Universal Indicator paper

Steel wool to clean the magnesium ribbon

Cardboard to clean on (otherwise you will scratch your tables)

Tongs/Forceps to hold the end of the magnesium ribbon with

Procedure:

1. Have all students place on safety equipment
2. Gather non-chemical materials at tables
3. Have students get a piece of magnesium ribbon and clean each side with steel wool on a piece of cardboard.
4. Have them light the fire source
5. Holding one end of the magnesium ribbon (with forceps or tongs) place the end of the magnesium ribbon in the fire source until it fully lights (you will see a brilliant white light)
6. Move the burning magnesium over the evaporating dish/watch glass
7. After completion of burn, observe the difference between the magnesium before hand and what has now been formed.
8. Add a few droplets of water (8-10) to the evaporating dish/watch glass and test the new compound (magnesium oxide) with the pH paper
9. Disposal:
    a. All unused magnesium ribbon needs to be collected so it can be reused.
    b. All spent magnesium oxide and water needs to go into a waste container; evaporated off and then thrown in the trash.

**Demo - Whoosh bottle with Ethyl Alcohol and (possibly other chemical fuels)**

Materials:

Water jug (from an office watering system like Crystal Farms, etc.)

Ethyl Alcohol (Ethanol) ~about 30-35 mL

Fire source (long handled BIC lighters work really well)

Procedure:

1. Place the ethyl alcohol into the jug discretely (so students don't know it is not water)
2. Using a spinning motion (see picture: appendix ___), talk to the students about what they smell coming out of the end of the bottle.
    a. Questions to use could be:

i. What does regular water smell like?

ii. Does this have an odor?

iii. Can you place the odor to something we have either done in class or have you smelled this someplace before?

3. Ask students why you are spinning the bottle?

   a. Answer: In order to move the ethyl alcohol particles out of the aqueous state and into the gaseous state. They are heavier then the water molecules, so they push them out leaving just the ethyl alcohol gaseous molecules.

4. You will want to continue to do this for about 3-5 minutes (absolutely no more than about 7 minutes).

5. Dump the excess aqueous ethyl alcohol out and place the bottle on a chair or floor wherever you feel comfortable.

6. Counting down (for effect and safety), bring the long BIC lighter to the top of the container. It should immediately shoot a fire ball out of the top of the bottle.

7. The "liquid" left inside is water from the combustion reaction

8. Disposal: All ethyl alcohol can go down the sink.

9. Additions to be made to this demo (**Note: You will have to use different bottles. Since the reaction is exothermic, it cannot be used again until it has cooled)

   a. Place borax (bought at the grocery store) into the bottle with water. Allow the water to evaporate totally before doing this. It creates the emerald green color from The Wizard of Oz.

   b. We might be able to try this with other fuels. I have not attempted this yet, so the results are unknown, but we need the fuels to be somewhat heavier than the air so their vapors stay in the bottle. It is unknown what this reaction would do with different fuels, since I have not attempted it. I would suggest doing anything other than ethyl alcohol in a well-ventilated area (ex: outside) and have the students stand back just in case.

**Activity 2 - Determination of $\Delta H$ for reactions using $\Delta H_f$ (enthalpy of formation) tables.**

In this calculation activity, we will demonstrate the predictive power of thermodynamics and calculate the energy produced in the previous Activity and the demo.

For the calculations:

Since these are all combustion reactions, students will need to write the balanced chemical reactions.  This is a great review of writing reactions and predicting products.

2 Mg (s) + $O_2$ (g) => 2 MgO (s)

$CH_3CH_2OH$ (ethanol, l) + 3 $O_2$ (g)  => 2 $CO_2$ (g) + 3 $H_2O$ (l)

Using the thermodynamic charts[3] (these can be found in the Physical Chemistry book or online to determine the $\Delta H_{form}$ for each compound.  Elements are zero (0) because they are found on Earth.

$\Delta H_{form}$ for the compounds in the reactions above are found to be:

Mg(s) and $O_2$(g) are in their elemental state so $\Delta H_{form}$ = 0 kJ/mole

MgO(s)  $\Delta H_{form}$ = -601.6 kJ/mole

$CH_3CH_2OH$ (l)  $\Delta H_{form}$ = -277.6 kJ/mole

$H_2O$ (l)  $\Delta H_{form}$ = -285.8 kJ/mole

$CO_2$ (g)  $\Delta H_{form}$ = -393.5 kJ/mole

After these values are found, place into the mathematical formula $\Delta H_{reaction}$=Products – Reactants and solve for a value, making sure to multiply the $\Delta H_{form}$ by the coefficient before you put in the equation.

For the MgO reaction:

$\Delta H_{reaction}$ = 2 * (-601.6 kJ/mole) – 2 (0 kJ/mole, Mg) – 1* (0 kJ/mol, $O_2$) = -1,203.2 kJ/mole   This calculation confirms that the reaction is very exothermic because the value is negative.

For the combustion of ethanol:

$\Delta H_{reaction}$= 2 * (-393.5 kJ/mole) + 3 * (-285.8 kJ/mol) -  1* (-277.6 kJ/mole) – 3* (0 kJ/mol) = -1366.8 kJ/mol  Again, a very exothermic reaction.

Students can then take any balanced reaction, and if they can find the $\Delta H_{form}$ values, they can predict the energy content and if the reaction would be exothermic or endothermic. For advanced students this would be a great challenge for them to look up their own data for any reaction they think is important to them.

**Activity 3 - Tin can calorimeter experiment investigating fuel sources and $\Delta H$ (all of these are combustion reactions)**

Materials:

Tin Can (Soda)

Scissors

Glass stir rod

Ring (that will fit can through and stir rod will sit on opposite edges of ring)

Ring stand

Graduated cylinder

Water

Alcohol burners with different fuels (*Make sure to designate specific alcohol burners for each fuel, otherwise it will be a mess to clean*)

Fuel (methanol, ethanol, butanol, gasoline, isopropyl alcohol)

Electronic balance

Fire source (BIC lighter, matches, etc)

Timer

Thermometer (electronic)

Funnel

Fume Hood (this will become very important if you choose to test gasoline or anything with large hydrocarbons, since they put off so much soot)

Procedure:

1. Using the scissors make 2 holes equally across the can to slide the glass stir rod through to stabilize the can. You want to make these two holes about one-third of the way down the can. Insert the glass stir rod. (This is your tin can calorimeter)
2. Setup the ring stand. Place the ring about 6-10 inches from the bottom (you will have to lower this later).
3. Place the tin can calorimeter on the ring.
4. Using the graduated cylinder, place approximately 25-30 mL of water into the soda can. (A funnel will help). The amount needs to be the same for each run of the experiment.

5. Place the thermometer (electronic) into the can making sure it hits the water. Let the temperature stabilize.
    a. If you do not have an electronic thermometer, you will have to find a way to suspend the thermometer in the water without it touching the bottom of the tin can. If the thermometer touches the bottom of the can, you will have skewed temperature readings.
6. Have students create a data table to record the start time, end time, start mass of the fuel, end mass of the fuel, start temperature of water, end temperature of water, and $\Delta T$ of the water. A sample might look like:

| Fuel Source | Mass:<br>Fuel Start | Mass:<br>Fuel End | Temp:<br>Water Start | Temp:<br>Water End | $\Delta T$<br>Water | Start Time | End Time |
|---|---|---|---|---|---|---|---|
| | | | | | | | |
| | | | | | | | |

7. Using an electronic balance, get the mass of one of your fuels. This includes everything in the assembly that will be burning. (Some caps come off, some stay attached. It depends on your kind of burner). This is your fuel start.
8. By now, the water temperature should have leveled out, so take the water start temperature (in °C).
9. Place the burner under the tin can calorimeter and lower the calorimeter assembly until it is only about 1-2 inches above the flame. Again, try to be as consistent between runs as you can.
10. Light the burner, and record your time. You need to have a temperature degree jump of about 30 degrees (keep it approximately the same throughout all the trials) of the water inside of the calorimeter. Watch carefully…
11. When the temperature gets to about a 30 degree difference, take the burner out and extinguish the flame. Continue to take temperature readings and record the highest temperature the water finally achieves as the end temperature in the chart. Record your end time (it may be a good idea to record the time difference in the margins.)
12. Using the same electronic balance, mass the alcohol burner and record the ending mass of fuel.
13. Repeat steps 7-12 replacing the water each time so it starts at about the same temperature for all fuels.

14. Remind students that “good scientists reproduce their results twice.” Or you can just tell them to do two trials for each fuel.
15. Disposal:
    a. The water out of the tin can may go down the drain.
    b. All other glassware and alcohol burners need to be put away following the directions of the teacher.
16. For the calculations:
    a. Before (Theoretical) – following the template of Activity 2
        i. Since these are all combustion reactions, students will need to write the balanced chemical reactions. This is a great review of writing reactions and predicting products.
        ii. Using the thermodynamic charts[3] (these can be found online to determine the $\Delta H_{form}$ for each compound. Elements are zero (0) because they are found on Earth.
        iii. After these values are found, place into the mathematical formula $\Delta H_{reaction}$=Products – Reactants and solve for a value, making sure to multiply the $\Delta H_{form}$ by the coefficient before you put in the equation.
    b. Experimental (All data comes out of the chart created. See Activity 3 Procedure #6)
        i. Students need to find the difference in time for all of their trials.
        ii. Find the $\Delta H$ per mole of fuel for each reaction

$\Delta H_{reaction}$= (Volume of water) * (density of water, 1 g/ml)*(specific heat of water, 4.184 J/g-°C)*(temperature change, $\Delta T$ in table)*(1kJ/1000 J)*(1/(mass of fuel used in grams))*(Molecular weight of the fuel in g/mole)

        iii. Compare to the theoretical values. Because of the flaws and heat losses in the experiment, the calculated values will always be smaller than the theoretical values.
        iv. Discuss the findings, the experimental flaws, fuel comparisons, …
        v. Have students write a report on their findings. A formal report could be assigned giving them a template from a real scientific journal article.

**Demo – Thermite**

**(This was attempted with 15 g Aluminum powder and 50 g Iron (III) Oxide)**

<u>Materials:</u>

1 6” clay flower pot

2 4.24” clay flower pots

Large ring that will hold the 4.24” clay pot

Ring stand

Sand (can be any kind; best is the play sand though)

Long handled BIC lighter

Magnesium Ribbon ~8-10 cm

30 grams of Aluminum powder or as finely ground as possible (the more fine the better the reaction)

100 grams of Iron (III) Oxide ($Fe_2O_3$)

Kimwipe or small piece of tissue in bottom of inside 4.25” pot so no chemicals spill out

2 spatulas

4-5 beakers

Electronic balances

<u>Procedure:</u>

There are several youtube videos that one can find that have similar set-ups to this one for a reference.

1. Measure 100 grams of iron (III) oxide onto a weighing boat or a tarred beaker.
2. Measure 30 grams of aluminum powder or finely ground aluminum onto a weighing boat or tarred beaker.
3. Gently mix the iron (III) Oxide and aluminum into a beaker together and set aside.
4. Setup your ring stand. Place a ring (capable of holding 2 4.24” clay pots) onto the ring stand.
5. Place the 2 4.24” clay pots inside of each other. Place a small piece of tissue or Kimwipe over the innermost hole so the $Al/Fe_2O_3$ mixture does not fall out.
6. Place sand (can be play sand) into the larger clay pot directly under the smaller pots on the bottom of the ring stand.

7. At this point take the ring setup outside and reset it up.
8. Place the $Al/Fe_2O_3$ mixture into the smaller clay pot at the top of the ring stand setup.
9. Insert into the mixture a piece of magnesium ribbon. You will need to bend it some so that it is sticking straight out of the mixture.
10. Light the magnesium ribbon and make sure that it is fully lit (make sure you see the bright light).
11. As soon as that happens, stand back.

**Activity 4 - How hot is too hot test on fabrics with different chemical fuels**

Materials:

Different fabrics cut into 2 inch X 2 inch squares (could be larger or smaller depending on your watch glass/glass petri dish

Fuel sources (methanol, ethanol, butanol, gasoline, etc)

Watch glass/glass petri dish

Timer

Electronic Balances

Procedure: (This would be done as inquiry at the end of my unit, but this is one procedure example.)

1. Have students make a hypothesis and a resulting data table determining what they want to look for in the burning of their fabrics. (The best one would be how long does it take to burn all the way through, how much alcohol does it take and time)
2. Cut out a 2 inch X 2 inch square (depending on watch glass size) of at least 3 different fabrics.
3. Weigh each fabric
4. Place each one on a watch glass (weigh watch glasses first)
5. Decide on a set amount of alcohol to start with using the calculations made prior in the unit.
6. Add alcohol to the watch glass and light on fire.
7. Time how long it takes for the fire to burn out and put into a data table

8. Weigh each fabric at the end as well (watch glass plus fabric then minus the watch glass)

Conclusion of Curriculum Unit

This curriculum unit makes students dive deeper into the understanding of how thermodynamics and energy relate to their everyday life. By using fabrics they can test and labs with fuels they are familiar with will give them a greater understanding of how the macromolecule study of thermodynamics fits into our lives.

End Notes:

[1] Mahan, Bruce H. Elementary Chemical Thermodynamics. W.A.Benjamin Inc: New York. 65.

[2] Arim, Mike. Interview by author. Matthews, North Carolina.

[3] Atkins, Peter 6$^{th}$ (1998) W.H. Freeman and Company. New York: 921-929

[4] Durwin, Striplin. Interview by author. Davidson, North Carolina.

Bibliography:

Mahan, Bruce H. *Elementary Chemical Thermodynamics*. New York: W.A. Benjamin, Inc., 1964

> This resource is a wonderful book in which it gives the ideas of thermodynamics in a neat, concise way. Someone who has been out of the thermodynamic arena for a while will appreciate how this book conveys what thermodynamics is and then works through example problems in detailed discussion with enthalpy, entropy and Gibbs free energy at the forefront.

Durwin, Striplin. Interview by author. Davidson, North Carolina, October 3, 2013.

> The discussions that happened with this individual were both enlightening and enriching to my mind. He worked through many thermodynamic calculations with me as well as teaching ways in order to bridge a very large topic; like thermodynamics into touch stones and ways to link content into the students' current lives.

Arim, James M. Interview by author. Matthews, North Carolina, November 16, 2012.

> The interview with this individual allowed me to understand the real reason that thermodynamics and kinetics work together. Working under this individual, I

was able to see how he linked content together to make multiple difficult thermodynamic concepts fit seamlessly together.

North Carolina Department of Public Education. “Chemistry Essential Standards” ncpublicschools.org.

These standards give a general and a very specific (unpacked) outline of what each student coming out of chemistry should have an understanding of. It allows teachers across the state of North Carolina to prepare each student in the best way possible and keep the accountability at a decent level with a specific state of standards.

Atkins, Peter Physical Chemistry. New York: W.H. Freeman and Company, 1998

This textbook has a wealth of information in it, included but not limited to the thermodynamic charts. Students will work to look up enthalpy values from this chart and integrate with large charts they are not used to looking at.

## Appendix 1: Implementing Common Core Standards

Chm.2.1.4 Infer simple calorimetric calculations based on the concepts of heat lost equals heat gained and specific heat.

Once students understand that heat must be transferred and it cannot disappear, they can then understand the concept of calorimetry.

Chm.2.2.1 Explain the energy content of a chemical reaction.

Energy is the needed in order to make sure that everything is annotated correctly and that energy is neither created nor destroyed.

Chm.2.2.3 Analyze the law of conservation of matter and how it applies to various types of chemical equations (synthesis, decomposition, single replacement, double replacement, and combustion).

Chm.2.2.4 Analyze the stoichiometric relationships inherent in a chemical reaction.

These standards are demonstrated in the writing and balancing of reactions, and calculations used in the activities.

Chm.3.2.1 Classify substances using the hydronium and hydroxide concentrations.

This is touched briefly when looking at the base properties of the MgO made in activity 1.