FLAMES, HEAT, AND CALORIES:
An Introduction to Thermodynamics

OBJECTIVES

The objectives of this experiment are to:

- ➤ Learn to operate a Bunsen burner at optimal conditions.
- ➤ Be able to select the optimum height above the burner for quick heating.
- ➤ Understand the concepts of calories and specific heat.
- ➤ Perform heat-gain and heat-loss calculations.
- ➤ Compare the energy changes involved in chemical and physical changes.
- ➤ Determine the amount of heat released by combustion of a peanut.

BACKGROUND

The Bunsen Burner

One of the more useful tools in a chemistry laboratory is the Bunsen burner. Developed by the German chemist Robert W. Bunsen in the mid 1800s, this burner has provided a source of hot, controlled flame for several generations of chemists and chemistry students. In this experiment we will review briefly the combustion process used to produce heat, and we will determine the optimum operating conditions for a Bunsen burner.

A Bunsen burner produces heat by combustion of natural gas. The reaction of the gas (methane) with oxygen to produce carbon dioxide and water vapor can be written in words:

Methane plus oxygen gives carbon dioxide plus water vapor plus heat,

or with chemical formulas:

$$CH_4 + 2\,O_2 \rightarrow CO_2 + 2\,H_2O + \text{heat}$$

Air is approximately 20% oxygen. The Bunsen burner mixes air with the fuel gas in appropriate proportions for the fuel to react completely. If the flame is fuel-rich, it is bright yellow and sooty. The bright yellow color is caused by hot, unburned carbon particles. An air or oxygen-rich flame is blue and unstable. It tends to strike back down inside the mixing tube and burn at the gas jet. A correctly adjusted burner has a violet flame that surrounds a small blue cone (see Figure 1). In part I of the experiment, you will explore the temperature gradients in the flame of a Bunsen burner.

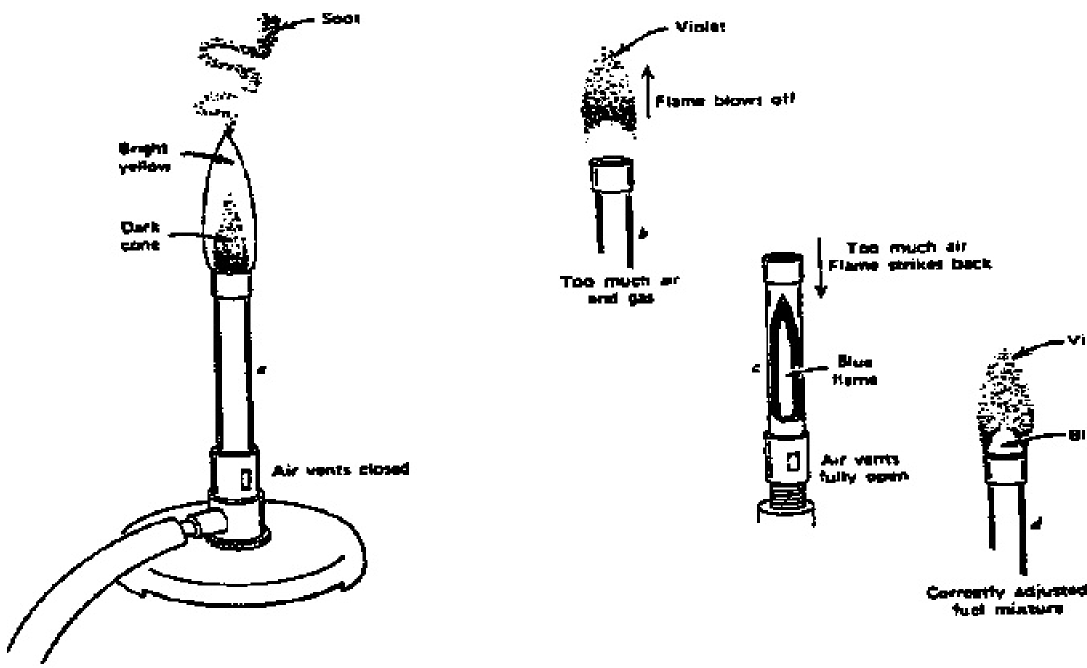

Figure 1. The Bunsen burner. Fuel and air mix in the vertical tube. A proper fuel-air mixture is important for maximum heat production by the burner.

An Introduction to Thermodynamics

A Model for Energy

It may be difficult to visualize a reasonable model for heat or energy. Perhaps it can best be visualized in terms of what it does. For example, when water is heated what happens to the water molecules? Scientists believe that the warmer molecules move at a greater velocity and collide with their neighbor molecules more frequently than cooler molecules do. The rate of evaporation of the molecules from a warm water sample is accordingly greater, and the vapor pressure of the sample (determined by the number of water molecules in gaseous form above the liquid) is greater. Therefore, it is possible to think of a reasonable model at the molecular level to help us visualize and develop an intuitive appreciation of the concept of heat.

Solid and Liquid forms of Water

As is readily apparent from the differences in properties of ice and liquid water, there must be a difference in the molecular behavior of these two forms of H_2O. An acceptable model should account for the changes observed when ice is heated to form liquid water and when liquid water is cooled to form ice. A chemist visualizes water molecules as extremely tiny balls of matter of approximately spherical shape that are in close proximity to each other. These molecules, however, are not static but are moving extremely rapidly from side to side and up and down, and are colliding frequently with neighboring molecules. As a liquid changes to a solid or freezes, much of the energy of the molecules is lost. The movement of molecules and frequency of collisions are diminished. The intermolecular forces of attraction in ice (the attractive forces between neighboring molecules) are strong enough to hold the water molecules relatively immobile.

Heat of Fusion

When ice is melted by the addition of heat, part of the heat goes toward breaking the forces that hold the water molecules together as a solid. The amount of energy required to melt a gram of solid (called **heat of fusion**) has a definite characteristic value for each compound. In Part IV of this experiment you will determine the amount of heat required to melt a gram of ice.

Energy Changes in Chemical Reactions

What happens when energy is absorbed or released in a chemical reaction? In any chemical reaction bonds are broken and new ones are formed. When you break bonds it always costs energy; a bond is a stable system in comparison to the separated atoms, so you have to expend energy to separate the atoms. Conversely, when you form a bond you release energy. So whether a reaction gives off or absorbs energy depends on the relative energies of the bonds that are broken and formed in the reaction. In part III of this experiment you will burn a candle, and measure the amount of heat the reaction transfers to a measured amount of water.

Specific Heat

In all of the experiments which follow, you will be measuring temperature changes, and using these changes to calculate changes in the heat content of various components of the experiment. These calculations will relate the terms *heat*, *energy*, and *temperature*. For any given substance (water, for instance), a specific and measurable amount of energy is required to change the temperature by a certain amount. **Specific heat** is the energy in calories required to raise the temperature of one gram of a compound or element by one degree centigrade. For water this number is especially important, and we will use it a lot in these calculations. The amount of heat required to raise the temperature of 1.000 gram of water by 1.000 degree centigrade is 1.000 calorie. The units of measurement for specific heat are cal/g-^{o}C. Changes in water temperature provide a convenient way to measure heat transfer and will be used in parts II-IV of this experiment. Remember that energy must always be conserved, so heat lost in one part of a system must be gained in another part.

SAFETY PRECAUTIONS

It is sometimes difficult to see at a glance whether a Bunsen burner is operating. Always exercise caution when working around a Bunsen burner by keeping your clothes, hair, and papers away from the flame area. Also, keep the burner away from suspended computer cables and electrical wiring. When you do part IV of the experiment, there should be no flames in the lab (operating Bunsen burners, burning candles, matches, or peanuts). As always, wear eye protection in lab when experimental procedures are in progress.

MATERIALS

Equipment:

- Bunsen burner, metal cup, thermometer, thermocouple, 2 thermistors (check out from stockroom)
- Ring stand, clay triangle, wire screen (from your locker)
- Styrofoam coffee cups, ring stand, candle (provided in lab)

Supplies:

➤ Index cards, peanuts, ice, ethyl alcohol, pentane

EXPERIMENTAL PROCEDURES **Part I. The Bunsen Burner Flame**

1. Connect your Bunsen burner to the gas valve, turn on the fuel supply, and light the burner by holding a match near its top. It is easier to light the burner if the air supply is restricted, causing a fuel-rich mixture. After the burner is lighted, adjust the air and gas controls until a violet flame with a blue interior cone results (see Figure 1).

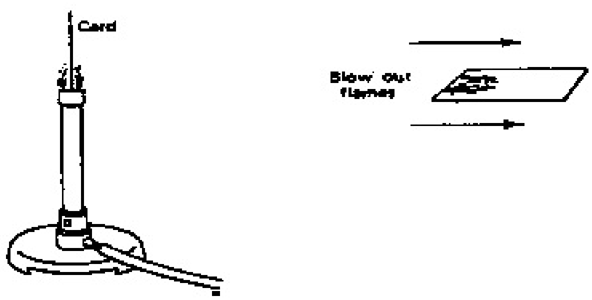

Figure 2

2. A scorch card provides an easy way to visualize the temperature distribution in a flame. A scorch card is simply a 3 x 5 index card that is held vertically in the flame with the bottom resting on the top of the burner, until it starts to catch fire. At this point, quickly remove the card, turn it sideways, and blow out the flame on both sides of the card at the same time (Figure 2). The scorch marks clearly show the temperature gradients within the flame. Attach this scorch card to your lab report.

3. Try another scorch card test, this time with the air supply to the burner restricted so that a yellow flame results. What conclusions can you draw concerning both the distribution and the magnitude of the flame temperature when operating under these conditions?

A more direct approach to measuring the temperature distribution in a Bunsen burner flame is to use a temperature probe called a thermocouple. The thermocouple you will use in lab is less sensitive than an integrated circuit (IC) temperature sensor, but it can be used to measure a wide range of temperatures. It can also be placed directly in the flame of a Bunsen burner, which would destroy an IC temperature sensor. **Be sure that you are using the right probe for this experiment**.

Use the blue voltage cable with alligator clips to connect the thermocouple to your MicroLab. The black alligator clip lead connects to the red thermocouple wire; the red alligator clip lead connects to the yellow thermocouple wire. Plug your voltage probe cable into one of the MicroLab CAT-5 inputs.

4 Write a MicroLab data acquisition program that will measure voltage and time, and set the program to make measurements every half second. Display your data as a graph with time on the horizontal axis, as digital numbers, and in the spreadsheet.

Because the thermocouple temperature conversion table reads in millivolts and the MicroLab voltage input displays its value in volts, let's write a formula to convert volts to millivolts. Click on "Formula", and using the mouse operated calculator, write : Millivolts = Volts *1000.

Drag the formula to the digital display. You can now read millivolts directly. Since the table below converts millivolts to degrees Celsius, you may wish to use this variable on your graph axis as well.

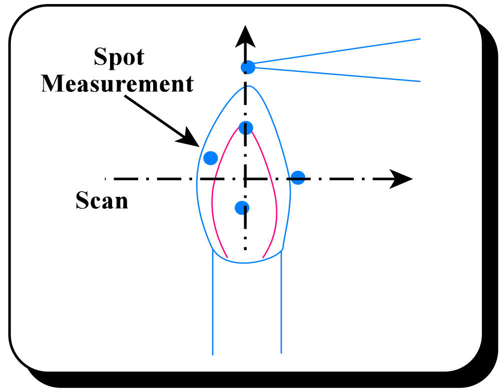

Figure 3. Use the thermocouple to determine the temperature in and around your flame.

5. Start the program and position the thermocouple tip at several points around and within the flame as illustrated in the figure. At each point tested, have your lab partner watch the computer display and record a representative thermocouple reading for that part of the flame. When you have tested the points shown, scan the thermocouple slowly through the flame horizontally and vertically. Use Table 1 to convert the thermocouple readings to approximate temperatures in centigrade, to the nearest 10 degrees.

Your lab report should include a drawing of the flame showing the approximate temperature at each measurement point and your scan graphs. How do these results compare to what you saw on the scorch card?

Thermocouple Table

Type "K" Thermocouple, 0^o Reference - Voltage in millivolts.

Degrees C	0	10	20	30	40	50	60	70	80	90
0	0	0.397	0.798	1.203	1.611	2.022	2.436	2.850	3.266	3.681
100	4.095	4.508	4.919	5.327	5.733	6.137	6.539	6.939	7.338	7.737
200	8.137	8.537	8.938	9.341	9.745	10.151	10.560	10.969	11.381	11.793
300	12.207	12.623	13.039	13.456	13.874	14.292	14.712	15.132	15.552	15.974
400	16.395	16.818	17.241	17.664	18.088	18.513	18.938	19.363	19.788	18.214
500	20.640	21.066	21.493	21.919	22.346	22.772	23.198	23.0624	24.050	24.476
600	24.902	25.327	25.751	26.176	26.599	27.022	27.445	27.867	28.288	28.709
700	29.128	29.547	29.965	30.383	30.799	31.214	31.629	32.042	32.455	32.866
800	33.277	33.686	34.095	34.502	34.909	35.314	35.718	36.121	36.524	36.925
900	37.325	37.724	38.122	38.519	38.915	39.310	39.703	40.096	40.488	40.879
1000	41.269	41.657	42.045	42.432	42.817	43.202	43.585	43.968	44.349	44.729
1100	45.108	45.486	45.863	46.283	46.612	46.985	47.356	47.726	48.095	48.462
1200	48.828	49.192	49.555	49.916	50.276	50.633	50.990	51.344	51.697	52.049

Figure 1. Output voltages from a "type K" thermocouple at 10 degree increments. To use the table, find the value closest to your thermocouple reading and add the column heading to the row heading to obtain the temperature to the nearest 10 degrees Celsius.

Part II. Heat Loss and Heat Gain

In this part of the experiment and in the parts that follow, you will use the lab interface to collect several sets of temperature data. Be sure that each time you save data to a disk file, you give the data file a new name that is different from the file names you have already used.

Our thermochemical experiments will be carried out in "thermocups," commonly known as styrofoam coffee cups. These cups are excellent insulators, and have such low mass that they absorb practically no heat themselves. *Please do not discard these cups. Rinse and save them for the next lab section.*

We will begin with a cup of a cold water and a cup of hot water, with a thermistor in each cup to monitor its temperature. Then we will pour the water from one cup into the other, transferring the thermistor as well, and use both thermistors to monitor the temperature of the combined systems.

1. Connect two temperature sensors to your MicroLab. Calibrate both sensors using ice water and hot tap water.

2. Write a MicroLab program that will measure time, Temp A, and Temp B. You will use this program in this and other parts of the experiment. Because the temperature changes rather slowly in some of the procedures, insert a 2 second delay in your program, so that you will collect a manageable amount of data.

3. Weigh a dry, empty styrofoam cup. Put about 30 mL of cold tap water in the thermocup. Add a little ice to the water so that when it is melted the temperature will be somewhat below room temperature, perhaps 10 or 15°C. Re-weigh the cup to determine the total mass of the cold water and the cup.

4. Put about 30ml of hot tap water in another thermocup. Put one of the temperature sensors in each cup (A in cold, B in hot) and **start your program**. Watch for a few seconds to see that the temperature readings are fairly stable. (The ice in the cold cup should be melted.) Now pour the hot water into the cold water, and put both sensors in the mixed water sample. Stir the water with the sensors until the temperature looks fairly stable, then stop the program. Save your data.

5. Remove the temperature sensors from the cup, touching the sensor tips to the inside of the cup to drain the water off them, and weigh the cup of mixed water samples to find the mass of the hot water that was added to the cold water. Why pour the hot water into the cold water, rather than the reverse? Would it have been more or less accurate to have weighed the hot water as well as the cold water before mixing the two samples?

6. To get the temperatures you will need for the heat loss and heat gain calculations, go to the **spreadsheet**, load your data file, and label the columns. Graph the data with time on the X-axis, Temp A on Y-axis 1, and Temp B on Y-axis 2. This graph will be easier to read if you **Scale** the Y1 and Y2 axes the same; they will not come out with the same scales unless you make them so. Do the first graph without doing anything about scaling, look at it, and then decide where to set the scaling. For the scale minimum pick a temperature that is a few degrees lower than the coldest temperature in your data, and for the maximum use a temperature a few degrees higher than the highest data point. Set the minimum and maximum for both Y1 and Y2 at the chosen values and add grid lines to your graph. Print the re-scaled graph, and use the graph and the spreadsheet data to find the water temperatures before and after mixing.

Data Analysis

We know from experience that if we mix cold water with hot water we will get water at an intermediate uniform temperature. We also know that the final temperature will be determined by the relative amounts of cold water and hot water we mix. How do we quantify these observations in a way that is consistent with a molecular model for heat transfer? Our molecular model says that the molecules in the hot water are moving with greater velocity and more collisions than the molecules of cold water. When the two samples are mixed, energy is transferred from the hot water molecules to the cold water molecules. We can calculate the amount of energy transferred in the form of heat in units of calories.

How much heat did the cold water gain? You had a certain number of grams of cold water at a measured temperature. After mixing, that water ended up at a higher temperature, which you also measured. From the specific heat of water, you know that one calorie of heat will raise the temperature of one gram of water by one degree centigrade. Therefore you can compute the number of calories of heat it took to raise the temperature of your cold water sample to its mixed temperature, using this equation:

Eq. 1 (Heat gain) = (specific heat of water) x (mass of cold water) x (temperature change)

> *Example:*
>
> If you started with 27.5 grams of cold water at a temperature of 15 °C, and the temperature after mixing with hot water was 26 °C, the heat gain would be:
>
> Heat gain = (1.000 cal/g-°C) x (27.5 g) x (11 °C) = 302.5 cal

What about the hot water? Its temperature decreased. You can also compute the number of calories that the hot water lost to decrease its temperature from its original to its final temperature.

(Heat loss) = (specific heat of water) x (mass of hot water) x (temperature change)

To perform these calculations use the data summary sheet for "Heat Loss and Heat Gain" at the end of this experiment.

Compare the number of calories of heat gained by the cold water to the number of calories of heat lost by the hot water. Compute the ratio of calories gained to calories lost. What conclusion can you draw?

Part III. Heat Associated with Chemical Change

<u>A. Combustion of Candle Wax</u>

In this experiment, you will measure the heat released during the combustion of candle wax, a chemical change. You will use water in a small metal cup to absorb the heat from the burning candle, and use the change in temperature of the water to calculate the amount of heat transferred. You will weigh the candle before and after heating the water to determine how much candle wax has reacted with oxygen (burned). Since some of the candle wax will melt, use a small piece of paper towel to catch the melted wax.

1. Set up a ring stand with a clay triangle on an iron ring to hold a metal cup above a candle. The cup will serve as a water container. Using a clay triangle to hold the metal cup will allow more efficient heat transfer from the flame to the cup than if you used a wire screen.

2. Clean as much carbon off the metal cup as you can, then weigh it. Put about 100 mL of cold water in the cup, then re-weigh the cup to find the mass of the water in the cup.

3. Weigh the candle together with a small piece of paper towel. Place the candle on the piece of paper towel under the metal cup. Adjust the height of the cup to about 1 inch above the candle wick.

4. Place your Temp A sensor in the water in the cup. Modify your program so only Temp A data is taken. Start the program and watch for a few seconds to see if the temperature reading is stable, and then light your candle. Stir the water gently with the temperature sensor as it warms up. When the temperature has risen about 15^{o}C (but not higher than the high calibration temperature in part II), blow the candle out, but leave the program running until the water temperature has reached a maximum.

5. Allow the molten wax to solidify so that drops are not lost in transit. Weigh the candle and towel to determine the grams of candle burned.

Data Analysis

Use Equation 1 from part II to calculate the amount of heat gained by the water in the metal cup. Your data and calculations should be recorded on the data summary sheet for "Heat Associated with Chemical Change."

The amount of heat gained by the water is approximately the amount of heat given off by the combustion of the candle, if we assume that all heat given off was absorbed by the cup of water. Actually, there would be some heat loss to the atmosphere, metal cup, clamp, and clay triangle but those losses are relatively small and will be disregarded in this experiment.

The amount of heat liberated per gram of candle burned can be estimated by taking the amount of heat gained by the water in the metal cup and dividing by the mass of the candle wax that was consumed in combustion. This value, in calories per gram, represents the heating ability of the candle and is also a measure of the energy change involved in a chemical reaction (the reaction of candle wax with oxygen from the air).

B. Combustion of a Peanut

Food provides two things for us -- molecules and atoms which act as building blocks for new and replacement cells, and energy to keep our metabolism going. The amount of energy that can be produced by metabolism of a relatively small amount of food is sometimes quite surprising. The problem involved in this part of this experiment is to determine the amount of heat that is produced as a peanut is burned. While chemical reactions in our bodies consume the peanut in a little different manner, the overall effect is the same: relatively complex molecules are broken into simpler molecules with a concurrent liberation of heat.

Repeat the experimental procedures and calculations that you just did in Part IIIa, but this time replace the candle with a peanut. Start with cold water again, and substitute a peanut on a wire screen for the candle on a piece of paper towel. You will need to adjust the height of the iron ring. Weigh the peanut and wire screen together before burning the peanut, and re-weigh the screen with the unburned peanut residue afterwards. Record your data and calculations in the peanut column on your data summary sheet. Which is the better fuel, gram-for-gram, candle wax or peanuts?

Part IV. Heat Associated with a Physical Change

A. Heat of Fusion

One can determine the heat of fusion of ice by placing a dry sample of known mass in a sample of water at known temperature, and watching the temperature change as the ice completely melts.

1. Weigh a styrofoam cup, then add about 100 mL of warm water from the tap and weigh again to determine the mass of the water exactly. Place your Temp A sensor in the warm water.

2. Obtain about one-third of a 50 mL beaker of ice. Start the temperature measurement program. Remove any water from the ice by spreading it on a paper towel and blotting with another towel. Now dump the ice into the styrofoam cup that contains the 100 mL of warm water.

Note that the ice must be dry when you add it to the water. Any ice that has melted will already have gained its heat of fusion.

3. Monitor the temperature of the water/ice sample, and stir it with the temperature sensor until the ice has melted and the temperature curve looks fairly flat. It may not ever get as flat as the one you had in the first experiment today. Then stop the program, take your cup to the balance room and weigh it to determine the weight of the ice you added.

4. Use the spreadsheet to determine the temperature of the water before the ice was added and after all the ice had melted.

Data Analysis

Perform the heat-loss and heat-gain calculations outlined below. It is a little more complicated this time, since the heat gain side consists of the heat required to melt the ice at 0 °C, and another amount of heat required to bring the melted ice to the final temperature.

1. *Heat loss calculation.* The heat loss calculation is fairly straightforward.

(Heat loss) = (specific heat of water) x (mass of warm water) x (temperature change)

2. *Heat gain calculation.* The amount of heat gained by the ice sample will equal the amount of heat lost by the original warm water sample. However, the heat gained by the ice has two components: a) the amount of heat required to change the ice to water at 0 ^{o}C, and b) the amount of heat required to warm the melted ice from 0 ^{o}C to the final temperature of the sample.

Calculate these two components separately. The amount of heat required to warm the melted ice from 0^{o}C to the final temperature is:

(Heat gained by melted ice) = (specific heat of water) x (mass of ice) x (temperature change).

The difference between the heat lost by the warm water sample and the heat gained by the melted ice is the amount of heat that went into melting the ice.

(Heat gained to melt ice) = (heat loss from warm water) - (heat gain by melted ice)

The heat required to melt one gram of ice can be calculated by dividing by the mass of ice that was melted. This value is known as the **heat of fusion.** Compute *your* value for the heat of fusion of water in the space provided on the data summary sheet for "Heat Associated with a Physical Change." Compare it with the accepted value of 79.7 cal/gm. With careful technique you should be able come fairly close to the accepted value. You may wish to do your calculations before you leave the lab. If your results don't agree with the accepted value, you will probably want to run this part of the experiment again.

Example:

An ice sample weighing 10.4 g was added to 100.2 g of water. The water temperature just before the ice is added was 25.4 ^{o}C. After all the ice had melted the temperature of the combined sample was 15.5 ^{o}C.

Heat loss calculation:

(Heat loss) = (1.000 cal/g-^{o}C) x (100.2 g) x (25.4 ^{o}C - 15.5 ^{o}C) = 992 cal

Heat gain calculations:

(Heat gained by melted ice) = (1.000 cal/g-^{o}C) x (10.4 g) x (15.5 ^{o}C - 0.0 ^{o}C) = 161 cal

(Heat gained to melt ice) = (992 cal) - (161 cal) = 831 cal

Heat of fusion calculation:

Heat of fusion = (831 cal) / (10.4 g) = 79.9 cal/g

Compare the amount of heat required to change the physical state of 1 gram of water from solid to liquid to that amount of heat liberated by the combustion of 1 gram of candle wax or peanut. Based on today's experiment, what generalization can you make about the relative amounts of energy involved in physical and chemical changes.

B. Heat of Vaporization

Even more energy is required to change a substance from liquid to gaseous state than to change it from solid to liquid. The energy involved in this change of state is called **heat of vaporization**. This part of the experiment will consist of a simple demonstration of the energy changes involved in vaporization.

1. Make sure there are no flames in the lab before starting this part of the experiment. Change the time delay in your program to ½ second. Start data collection, and after the graph has recorded about ten seconds of data at room temperature, put a drop of ethyl alcohol on the tip of the temperature sensor.

2. Wave the temperature sensor back and forth in the air, and watch the temperature as the drop of liquid evaporates from the thermistor tip. Continue to wave the temperature sensor until the temperature returns to room temperature. Why did the temperature of the thermistor change?

3. Save the data to a disk file.

Try this experiment twice again, using water and pentane. Can you draw any conclusions about the relative heat of vaporization of these three materials? Use the spreadsheet to graph the three sets of vaporization data and print the graphs for your report.

PRE-LAB QUESTIONS

Make sure you can answer these before you enter the lab!

1. Describe how the flame of a properly adjusted Bunsen burner should appear.

2. Distinguish between **heat** and **temperature**.

3. In the melting of ice portion of this experiment, suppose you had 101.2 grams of water which cooled down by 10.2 ^{o}C upon the addition of 9.7 grams of ice. The final temperature of the ice/water mixture was 20.0 ^{o}C. Use this data to calculate the heat of fusion for water. (Remember, consider the temperature of ice to be 0.0 ^{o}C, and remember that heat lost equals heat gained!)

4. In part III of this experiment, you will measure the heat associated with a burning candle. Does this represent a physical change or a chemical change? Describe the change occurring.

Name ____________________

Date _____________________

Lab Section _______________

DATA SUMMARY

Part I. The Bunsen Burner Flame

Attach your scorch card and your sketch of a Bunsen burner flame with thermocouple temperature measurements.

Part II. Heat Loss and Heat Gain

Mass of cup and cold water	_____________ g
Mass of cup	_____________ g
Mass of cold water	_____________ g

Mass of cup and mixed hot and cold water	_____________ g
Mass of cup and cold water	_____________ g
Mass of hot water	_____________ g

	Cold water sample	Hot water sample
Temperature before mixing	_____________ ºC	_____________ ºC
Temperature after mixing	_____________ ºC	
Change in temperature	_____________ ºC	_____________ ºC

Heat gain calculation:

Heat loss calculation:

Ratio of (Heat gain) / (Heat loss):

Part III. Heat Associated with Chemical Change

	Candle	Peanut
Mass of cup and water	_____________ g	_____________ g
Mass of empty cup	_____________ g	_____________ g
Mass of water	_____________ g	_____________ g
Water temperature before heating	_____________ ºC	_____________ ºC
Water temperature after heating	_____________ ºC	_____________ ºC
Temperature change	_____________ ºC	_____________ ºC
Mass of candle and paper towel or peanut and screen before burning	_____________ g	_____________ g
Mass of candle and paper towel or peanut residue and screen after burning	_____________ g	_____________ g
Mass of candle or peanut burned	_____________ g	_____________ g

Calculate calories of heat from the burning candle absorbed by the water.

Calculate the calories of heat produced per gram of candle wax burned.

Calculate calories of heat from the burning peanut absorbed by the water.

Calculate the calories of heat produced per gram of peanut burned.

Part IV. Heat Associated with a Physical Change

A. Heat of Fusion

Mass of cup and water	_____________	g
Mass of cup	_____________	g
Mass of water	_____________	g
Mass of cup and water and ice	_____________	g
Mass of cup and water	_____________	g
Mass of ice	_____________	g
Temperature of water before adding ice	_____________	ºC
Temperature after ice melted	_____________	ºC
Temperature change of water sample	_____________	ºC
Temperature after ice melted	_____________	ºC
Initial temperature of ice	0.0	ºC
Temperature change of melted ice sample	_____________	ºC

Heat loss calculation:

Heat gain and heat of fusion calculations:

From your data, the heat required to melt one gram of ice is _____________ calories.

B. Heat of Vaporization

Attach temperature graphs for vaporization of ethyl alcohol, water, and pentane.