**Chemistry Lab: *Specific Heat of a Metal***

Laura E. Bishop

**Introduction - (also includes Pre-Lab information):**

One of the vast many characteristic properties of water (H20) is its unusual ability to absorb large quantities of heat without significant change in temperature. This characteristic is called the specific heat capacity (Cp). When water absorbs exactly 4.184 joules of heat, the temperature of one gram of H20 will increase by one degree Celsius, basically stating that this is an enormous amount of heat energy. An example can be seen when looking at the State in which we live! Florida, a coastal state surrounded by water, maintains relatively stable climates with moderate temperatures. This is the case because water is able to absorb and/or release a great deal of heat without experiencing enormous fluctuations in temperature.

18th century Scottish scientist [Joseph Black](http://www.britannica.com/EBchecked/topic/67460/Joseph-Black" \o "Joseph Black), noticed that equal masses of different substances required different amounts of heat to raise them through the same temperature interval, thus founding the concept of specific heat. The specific heat of a substance is the heat required to raise the temperature of one gram of substance one degree Celsius. The molar heat capacity is the amount of energy required to raise one mole of the substance by one degree. Quite often, it is applied to the metallic elements, where it can be used as a basis for comparing energy absorption and transfer. In this lab, the specific heat of a metal (Cu) is calculated by using the equation:

**(q) = (mass)(Cp)(° ∆T)**

**q= quantity of heat (in Joules)**

**m=mass of object (in grams)**

**Cp=specific heat capacity (in Joules/gram °C)**

**∆T=change in temperature of the object ( Tf-Ti in °C)**

As seen in our lab with Cu and H20, when any two substances (which have different temperatures) come into contact, energy in the form of heat is exchanged until they both reach a “common temperature”. The joules of energy gained by the water will be equal to the joules of heat lost by the copper. Thus, the heat transferred (q), is found using the equation:

**heat energy lost by Copper (q) = heat energy gained by the H20 + heat gained by beaker**

Calculating the heat gained by the water allows us to configure how many joules were lost by the copper. For this, we use the equation:

**heat gained by water (J) = (mass of H20 in grams)(∆T)(Cp of the H20= 4.184 J/g °C)**

With this experiment, we also learned about a calorimeter. This is a device used to measure the amount of energy evolved or absorbed in a chemical or physical process. Simply put, a calorimeter is a container with insulating walls, so that basically no heat is exchanged between the contents of the calorimeter and the surroundings.Inside the calorimeter, chemical reactions can occur or heat may pass from one part of the content to another. However, that being said, no heat flows in or out of the calorimeter from or to the surroundings. In this experiment, aweighed amount of copper is heated to a certain temperature and then quickly placed into a calorimeter which contains a measured amount of H20 at a known temperature. Heat flows from the Copper to the H20 and the two “equilibrate” at a temperature between the initial temperatures of the copper and the H20. The amount of heat that flows from the metal as it cools is equal to the amount of heat absorbed by the water and the calorimeter.

In terms of non-SI heat unit of calorie, the specific heat of water is 1 calorie/(g °C), which allows the specific heats of other substances to be easily compared to it. I noticed that the specific heat of H20 is quite large when compared to many other substances used in chemistry.

**Materials Used:**

\*2 x 250mL beakers

\*About 250mL distilled water. (Note: De-ionized water can have many dissolved substances that do not result in ions, yet are present nonetheless, potentially altering any future calculations. However, distilled water is very pure and contains only water molecules, the very reason it was used in this experiment.)

\*Sample of a metal – Copper pellet

\*Hot plate

\*Tongs

\*Ring stand

\*Iron ring

\*Calorimeter: the exchange between the copper sample and the water that occurred inside the calorimeter, minimizing the heat loss to the environment.

\*Thermometer: inserted into the Calorimeter

\*Boiling water bath - the copper sample was originally submerged here and an initial temp of @ 100ºC was achieved.

\*Bunsen burner: heats the water

\* “flicker” to light the Bunsen burner with

\*Balance/scale: measures to 1/1000th of a gram in order to measure the mass of the Copper sample and the mass of the water.

\*Pen and Paper to record data

**Procedure:**

In summary, during this lab I used the principles of thermodynamics to determine the specific heat of a metal sample, Copper. Heat is never lost – but it can be transferred. In this lab, we will be transferring heat from a hot copper sample to a cool water sample. When a piece of Copper (Cu) is submerged in a sample of water, any heat exchanged from the metal to the water will result in a temperature change for the water until both water and the Copper sample reach equilibrium. In order to determine the specific heat of the copper sample, an observable temperature change was needed. Therefore, the copper was started in a boiling water bath and a thermometer was used to measure the temperature of the bath to a tenth of a degree Celsius.

**Step 1**: Obtain a sample of Copper, recording its color, mass, and kind of metal.

**Step 2**: Room temperature distilled water was added to the calorimeter and the mass of the water was found by taking the difference between the mass of the calorimeter with the water and the mass of the empty calorimeter.

**Step 3**: Having filled the 250ml beaker until it was half full with distilled water (125ml), now place the beaker of water on a hot plate or a ring stand with wire gauze and use the Bunsen burner to bring to a boil.

**Step 4**: As the temperature of the calorimeter and water eventually come to equilibrium, read and record this temperature; this was used as the initial temperature of the water*.* Place the Copper pellet into the beaker of boiling water and allow it to sit for at least 5 minutes, ensuring that it reaches 100 °C.

(\*\*\*IT IS VERY IMPORTANT TO MAKE SURE THE COPPER SAMPLE AND WATER EQUILIBRATE TO 100 °C\*\*\*)

**Step 5**: The temperature of the boiling water bath was recorded before removing the copper; this was used as the initial temperature of the copper.

**Step 6**: The hot metal copper was placed quickly into the calorimeter by carefully using tongs and the temperature of the water was monitored with the thermometer, while the contents inside were gently swirled every 20-30 seconds to avoid areas of differently heated water.

**Step 7**: Record the highest temperature the water reached, as this temperature was used as the finaltemperature of the water*.*

**Step 8**: The copper sample was removed from the calorimeter, dried, and its mass measured. Using the equations listed above, we were able to determine the specific heat capacity (Cp) of Copper.

**Step 9**: Using the percent error equation and the theoretical values of specific heat, determine the percent error for the copper sample tested and see how accurate we were in our measurements during the lab.

**Results:**

I determined the final temperature of the water by using the formula:

**∆T = Change in Temp of the Water =**

**( Final Temp of the Water – Initial Temp of the Water ) =**

**25.2ºC – 23.2ºC = 2.0 °C**

Since the copper was submerged in the water as they reached that final temperature, I assumed they were at the same temperature and thus used the final temperature of the water as the final temperature of the copper*.*

To determine the heat water gained, I used the formula**:**

**Heat water gained = (Mass of water) (Specific heat of water) (Change in Temperature of water)**

**Heat Water Gained** **= ( 149.6g ) ( 1 cal/g x Cº ) ( 2.0Cº )**

**Heat Water Gained =** 299.2cal

As stated earlier, the heat the water gained will be the same as the heat the metal lost*.*

**Heat lost by the copper sample (H) = (mass of copper)(specific heat of copper)(∆T)**

**H = Heat lost by the copper = 299.2cal**

**m = Mass of copper = 42.57grams**

**Specific heat of copper = ? cal/g x °C**

**∆T = Change in temperature of the copper sample =**

**(final temperature of copper sample – initial temperature of copper sample ) =**

**100.2ºC – 25.2ºC =**

**75.0 ºC**

**I calculated the heat copper lost by using the formula:**

**Heat copper sample lost = ( mass of copper ) (specific heat of copper ) (change in temperature of copper)**

**299.2cal = ( 42.57g ) ( Sp.Ht. ) ( 75.0Cº )**

**Specific heat of Copper = 0.094 cal/g x Cº**

We can now compare our measured specific heat of copper to the known specific heat of copper (which 0.39 J/g °C) to see how accurate we are in our measurements.

% Error = |(Lab Value) – (Known Value)| x 100

Known Value

( 0.094 cal/g x Cº ) – ( 0.095 cal/g x Cº ) x100 = 0.1 cal/g x Cº = 1.05% Error

A percent error of 1.05% is acceptable given the equipment used. Some possible sources of error could have been in the thermometers not being accurately read, the heat lost to the air was not factored in, also possibly altering the results.

**Conclusion:**

Since the Copper sample was fully submerged when it lost its heat, I can say that the heat flowed from the cylinder to only the water. The heat lost by the metal (copper) is equal to the heat gained by the water. The value I calculated for copper in this lab was very close to the value of copper in the table of specific heat. Overall, I feel this lab exercise should be considered successful.