**Bomb Calorimetry**

**Objectives**

In this experiment, bomb calorimetry will be used to determine the heat of combustion and standard enthalpy of formation of a pure substance (sucrose) as well as the calorie content of a food product.

**Introduction/Theory**

A bomb calorimeter is a sealed container capable of holding several atmospheres of gas pressure. A weighed sample of substance is placed in contact with an ignition wire inside the bomb. The bomb is filled to about 20 atm of pressure with O$_2$, sealed, and placed in a known amount of water. An electric current is passed through a wire to ignite the mixture. As the combustion takes place, the heat evolved raises the temperature of the calorimeter and its surrounding water, as measured by a thermometer. In order to prevent heat loss from the calorimeter system, some calorimeters are surrounded by a second water bath, whose temperature is continuously adjusted (by the experimenter) to match that of the calorimeter. Thus, the heat transfer between the system (calorimeter and contents) and the surroundings (the water jacket, primarily) is zero, making the process adiabatic: $q = 0$. Our calorimeters use good thermal insulation instead of a second water bath to prevent loss of heat to the surroundings.

Since the combustion takes place in a sealed container with constant volume, the work done on or by the system is also zero, so that $\Delta U = q + w = 0$ for the overall process shown below.

![Figure 1: Since internal energy is a state function, we can break the overall process into two steps.](image)

The initial temperature of the system is $T_1$ and the final temperature is $T_2$ after combustion of reactants to products and heating of the calorimeter. Since $U$ is a state function, the pathway from initial to final state does not affect the value of $\Delta U$. This allows us to separate the two processes (combustion and heating), although, in reality, they occur simultaneously.

With this in mind, it is true that

$$\Delta U_{overall} = \Delta U_a + \Delta U_b = 0 \quad [1]$$

or,

$$-\Delta U_a = \Delta U_b \quad [2]$$
But, since Step b simply involves heating up the calorimeter and contents from $T_1$ to $T_2$,

$$\Delta U_b = C_v \Delta T = C_v (T_2 - T_1) \tag{3}$$

Note that $C_v$ is the heat capacity of the bomb calorimeter and its contents. From equations [2] and [3], the change in internal energy of combustion can be readily found from the change in internal energy of heating.

The molar enthalpy of combustion $\Delta H_m$ is given by the equation:

$$\Delta H_{\text{comb, m}} = \Delta U_{\text{comb, m}} + RT_{\text{comb}} \Delta n_{\text{gas}} \tag{4}$$

where $T_{\text{comb}}$ is the temperature at which combustion occurs, $\Delta n_{\text{gas}}$ is the change in the number of moles of gas (in the balanced reaction representing the combustion of one mole of compound). To ensure that units match, $\Delta n_{\text{gas}}$ should be treated as unitless.

Determining the heat evolved during the combustion from Equation 3 requires knowledge of the heat capacity of the calorimeter often called the calorimeter constant. By combusting a sample with a known molar heat of combustion, we can use the measured temperature change and Equation 3 to calculate the calorimeter constant.

**Procedure**

1. Watch the video posted on the web site.
2. Turn on the calorimeter.
3. Use a top-loading balance to weigh out approximately 0.9 g of special grade benzoic acid not exceeding 1 gram. Use this to make a pellet in the pellet pressure. Make three of these pellets, and then two sucrose pellets.
4. Tare an analytical balance with the metal combustion cup, and then place a pellet into the cup and obtain its mass to ± 0.01 mg.
5. Attach one strand of cotton string to the wire connecting the electrodes. Make sure that the string touches the pellet, and that the ignition wire is attached firmly to the electrodes. Gently wiggle the electrodes; they should not be loose.
6. When closing bomb, tighten down the cap by HAND only. Fill the bomb with oxygen, allowing the pressure to build up slowly to 25 atm. Purge the bomb with pure oxygen twice. Fill a third time, and then connect the ignition head to the top of the bomb.
7. Put the bomb in the calorimeter. Make sure that it is centered on the three pins at the bottom of the water bath.
8. Fill the reserve tank to the line with tap water.
9. Close the lid, and hit the button under “prepare an experiment”.

10. When the calorimeter has finished filling with water, take temperature readings every 30 seconds for two minutes before igniting the bomb.

11. Hit the button under “ignite”.

12. After a delay of 10 or 15 seconds, the temperature should begin to rise. Read the temperature of the calorimeter every 15 seconds for 3 minutes after igniting the bomb.

13. The temperature will begin to level off. Take temperature readings every 30 seconds for two more minutes.

14. Open the lid, wait for the water to drain, and then remove the bomb. Release the pressure in the bomb. Open the bomb, and check to make sure that no sample is left. Wash and dry the bomb vessel and head.

15. Combust at least three samples of the benzoic acid standard, at least two sucrose samples, and two food product samples.

**Calculations**

For each combustion, determine $\Delta T$.

During ignition, the cotton string combusts, thereby adding heat to the system. The extra heat must therefore be subtracted from the total. The amount of heat liberated (50 J) is stated on the bag of string.

$$\Delta U = -C_v \Delta T + q_{\text{string}}$$  \[5\]

**Heat Capacity of the Calorimeter**

The balanced equation for the combustion of one mole of benzoic acid [122.1 g mol$^{-1}$] is:

$$C_6H_5CO_2H(s) + \frac{15}{2}O_2(g) \rightarrow 7CO_2(g) + 3H_2O(l)$$  \[6\]

The standard molar enthalpy of combustion for benzoic acid is $\Delta H_{\text{comb, m}} = -3227 \text{ kJ mol}^{-1}$. The standard molar energy of combustion for benzoic acid can thus be calculated.

$$\Delta U_{\text{comb, m}}^o = \Delta H_{\text{comb, m}}^o - RT\Delta n_{\text{gas}}$$

$$= -3227 \text{ kJ mol}^{-1} - \big(8.314 \text{ E-3} \text{ kJ mol}^{-1}\text{K}^{-1}\big)\big(298 \text{ K}\big)\big(7 - 15/2\big)$$  \[7\]

$$= -3227 \text{ kJ mol}^{-1} + 1.24 \text{ kJ mol}^{-1}$$

$$= -3226 \text{ kJ mol}^{-1}$$
Using $\Delta U_{\text{comb,m}}$, the heat capacity for the calorimeter, $C_{\text{cal}}$ can be calculated using equation 5.

**Molar Enthalpy of Combustion for Sucrose**

Write the chemical reaction for the complete combustion of one mole of sucrose. Calculate $\Delta U^\circ$ for the combustion of your sample from

$$\Delta U_{\text{comb}} = -C_{\text{cal}}(T_2 - T_1) + q_{\text{string}}$$

[8]

Determine the *molar* change in internal energy in kJ mol$^{-1}$. Then calculate the molar enthalpy of combustion from equation [4] using $T=298K \sim T_1$.

Compare your value with that listed in the literature.

Now calculate the standard enthalpy of formation of sucrose from your enthalpy of combustion and the enthalpies of formation of CO$_2$ and H$_2$O.

**Molar Enthalpy of Combustion for Food Product**

Since these substances are complex mixtures, one can only calculate the heat of combustion per gram, assuming that $\Delta U \sim \Delta H$. Find $\Delta H$ (in kJ/gram) for each substance tested. Compare to the stated "calorie content" on the product label. If using food without a clearly labeled caloric content such as a nut, you may look up the calorie content on the USDA nutrient database.