Lab: How Much Heat Is Released Upon Fuel Combustion?

Purpose: In this laboratory, you will measure the amount of heat released during combustion of some common liquid fuels and then a food item. From these experimental measurements the heat of combustion (ΔH_{comb}) for each of these fuels will be calculated. ***During this lab activity you need to be aware of your ability to reproduce your experimental results.***

General Procedure

1. Assemble a calorimeter as demonstrated in the pre-lab introduction. Include a diagram of the apparatus in your lab notebook.

2. Working in pairs, you will use a small burner to combust the fuels; practice lighting and extinguishing the burner.

3. For each combustion experiment, three experimental parameters required: a) the mass of fuel combusted, b) the mass of water heated as a result of combustion, and c) the change in temperature of the water. All masses are to be measured and recorded to ± 0.01 g; initial and final temperatures are to be measured and recorded to ± 0.1°C. The following are some suggestions to obtain the experimental parameters.

Mass of fuel: It is suggested that the burner be massed (weighed) with the cap on at the beginning of an experiment and then again at the end of the experiment.

Mass of the water: The mass of the water used during the experiment is best determined by first measuring the volume in a graduated cylinder (~50 mL) and then determining the exact mass by taring the "soda can" on the balance and then adding the water.

Temperature change: We will use the Vernier stainless steel temperature probes to measure an initial temperature (T_i) and a final temperature (T_f), then do the math. We suggest holding the thermometer at all times for three reasons: a) if you sit the thermometer on the very bottom of the can, it may be much hotter than the water itself, b) by holding the thermometer, you can gently agitate the water to allow the heat to dissipate throughout the entire water sample, and c) you are much less likely to "tip over" the apparatus.

4. To initiate the experiment, obtain the mass of the burner with cap, and then do the following quickly: a) remove the burner cap, b) light the burner, c) place the water-filled soda can/temperature shield on top of the burner. Monitor the temperature while the fuel combusts, extinguishing the flame once the temperature has risen ~ 20 °C. Measure and record the final temperature of the water.

Experiment 1: Calibration Experiment; Burning Methanol

It is unreasonable to expect that all of the heat generated from a reaction will be captured by the water. For this reason, we conduct this "calibration" experiment. Following the above procedure, we can collect data to calculate a "heat loss factor" (*f*) by burning a fuel with a known heat of combustion (ΔH_{comb}), in this case methanol (CH_3OH). You will use the change in mass of the CH_3OH burner, the temperature change, and the amount of heat that should be generated (calculate) to calculate the *f*; see below. Collect data for three separate experiments, calculate *f* for each experiment and then average the *f* values. Record all data and calculations in your lab notebook. Please check your individual *f* results with your instructor before proceeding to the next experiment.

Calculating *the heat loss factor*: The *f* of the calorimeter accounts for the fact that not all heat generated by the combusted fuel goes into the water. The *f* is the ratio of the heat absorbed by the water (q_w) to the heat generated (q_f) by burning *n* moles of fuel.

$$f = \frac{\text{heat absorbed by the water } (q_w)}{\text{heat generated from fuel } (q_f)} \qquad \text{(eq 1)}$$

q_f is calculated by multiplying the moles of fuel burned (calculated from the grams of methanol burned) by ΔH_{comb} of methanol (-726 kJ/mol).

$$q_f = n\,(\Delta H_{comb}) \qquad \text{(eq 2)}$$

q_w is calculated using the mass of the water, the specific heat of water, and the change in temperature.

$$q_w = (\text{mass } H_2O) * (\text{specific heat } H_2O) * \Delta T \qquad \text{(eq 3)}$$

You have recorded the mass of water and the can determine the ΔT from your data. The specific heat of water is $4.184 \frac{J}{g \cdot {}^\circ C}$.

Ideally, the water would absorb all the heat generated in the combustion reaction and $f = 1$. In reality, there will be heat lost to the environment and the numerator of eq 1 will be smaller than the denominator. Therefore, the *f*-factor should be less than 1.

***Experiment 2:* Assigned Fuel Experiment:**
You and your partner will be assigned one other fuel to combust. Repeat the experiment as described above using your fuel (3 trials).

The heat of combustion (aka: molar enthalpy change for the combustion) of your assigned fuel can be calculated using the following equation:

$$\Delta H_{comb} = \frac{q_w}{f * n} \qquad \text{(eq 4)}$$

where, *f* is the heat loss factor (average value from three methanol experiments), q_w is calculated from eq 3 and *n* is the moles of fuel combusted (calculated from the mass).

Clearly show all calculations in your lab notebook.

***Experiment 3:* Food Fuel Experiment:**
Repeat experiment 2, but instead combust a food item provided (3 trials). *Note: the ΔT just needs to be greater than 5 °C*. Calculate the "heat of combustion" in units of J/gram and Calories/gram and then compare to the food packaging.

There is no reporting sheet for this lab. Turn in all carbon copies, which contain all data and calculations associated with the lab activity.

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