Name: __*Answer Key*________ Date: ___/___/___

# Energy Content of Foods - Calorimetry

Energy content is an important property of food. The energy your body needs for running to exciting chemistry classes, talking about awesome chemistry principles, and thinking about how cool chemistry is, comes from the food you eat. Energy content for every food is quantifiable; defining the amount of heat produced by the burning of one gram of that food, and is measured in joules per gram (J/g), or even calories per gram (cal/ g).

In the laboratory, you can determine energy content by burning a portion of food and capturing the heat released to a known mass of water in a calorimeter. If you measure the initial and final temperatures, the energy released can be calculated using the equation:

$$\Delta H = mC\Delta T$$

where $\Delta H$ = heat energy absorbed (in J), $m$ = mass (in g), $C$ = specific heat capacity (4.18 J/g°C for water), and $\Delta T$ = change in temperature (in °C). Dividing then, the resulting energy value ($\Delta$H) by grams of food burned provides the energy content (in J/g).

There are some assumptions we make when performing this experiment. We assume that all of the heat released *(*$\Delta H$) as the consequence of setting the food on fire, is absorbed directly by the water in the calorimeter (or Erlenmeyer flash, for our purposes). As an example, if the food burned releases 15 joules of energy, we assume the water absorbs 15 joules of energy. In calorimetry, the only way we can quantify the heat released however *(*$\Delta H$), is by observing the temperature change *(*$\Delta T$) of a known mass ($m$) of water – which is why we use the heat equation.

**Objectives**
In this experiment, you will:

• Empirically determine energy content of two foods per gram.
• Calculate percent error comparing empirical value with accepted value for each food
• Compare the energy content of different foods.

## Methods

**Materials**

Erlenmeyer flask (125 mL capacity)
Cylindrical metal stand
Metal suspension plate
Cork-and-nail sample holding assembly
Rubber stopper
Matches or a lighter
Thermometer or temperature probe
Triple-beam balance (0.01 g resolution)
Two food samples (peanut, popcorn,
potato chip, marshmallow)

*Figure 1*. Components of the Eisco Calorimeter.

**Procedure**

1. Obtain and wear goggles (you are about to burn something!).

2. Get a sample of food and a food holder like the one shown in Figure 1 (cork-and-nail).

3. Record the *name* of the food in Table 1 (a).

4. Find and record the *initial* mass of the food sample and food holder, in Table 1 (b). CAUTION: Do not eat or drink in the laboratory.

5. Determine and record the mass of the *empty* Erlenmeyer flask in Table 1 (d).

6. Place about 125 mL of cold water into the flask.

7. Determine and record the mass of the *flask plus water* in Table 1 (e).

**Set up the apparatus shown in Figure 2.**

8. Insert the top of the flask through the metal plate and turn by 90 degrees so the metal clamp on the flask rests on the plate.

9. Insert a thermometer into the flask through the hole in the rubber stopper, making sure it is submerged in the water, but not touching the glass at the bottom of the flask.

10. Place the plate on top of the metal cylinder to suspend the flask.

11. Firmly affix the food sample to the cork-and-nail holder (make sure the food will not fall off during burning).

*Figure 2*. Calorimeter Assembly.

**Start the experiment.**

11. Record the *initial* temperature of the water in Table 1 (f).

12. Slide the food holder out from under the flask, and using a wooden splint, light the food, and quickly place the burning food sample directly under the center of the flask. Allow the water to be heated until the food sample stops burning. CAUTION: Keep hair and clothing away from an open flame.

13. Gently swirl the flask periodically while the food is burning, until the temperature stops rising. Record the highest, *final* temperature in Table 1 (g).

14. Determine the *final* mass of the food sample and food holder, by weighing them both, and record in Table 1 (c).

15. Repeat the procedure for a second food sample. Use a new 125 mL portion of cold water.

16. When you are finished, place burned food, used matches, and partly-burned wooden splints in the container supplied by the teacher.

## Data

Table 1

*Data Recorded from Calorimetry Protocol*

| | | Sample 1 | Sample 2 |
|---|---|---|---|
| a. | Food Used | crouton | peanut |
| b. | Mass of food and holder (initial) | 4.91 g | 4.59 g |
| c. | Mass of food and holder (final) | 4.28 g | 4.12 g |
| | Mass of the food that was burned (c – b) | 0.63 g | 0.47 g |
| d. | Mass of empty flask | 96.03 g | 96.03 g |
| e. | Mass of flask plus water | 205.59 g | 210.08 g |
| | Mass of water only (e – d) | 109.56 g | 114.05 g |
| f. | Initial water temperature | 19.9℃ | 20.1℃ |
| g. | Final water temperature | 25.2℃ | 41.03 ˚C |
| | Temperature change (g – f) | 5.3℃ | 21.2℃ |
| | Empirical* energy content per gram (J/g) | 3,856.37 J/g | 21,524.19 J/g |
| | Theoretical energy content per gram (J/g) | 20,920.00 J/g | 24,685.60 J/g |
| | $\% \text{ error} = \frac{\text{theoretical value} - \text{empirical value}}{\text{theoretical value}} \text{x } 100$ | 81.57 % | 12.81 % |

## Analysis

For the below analysis, *show your work*, where appropriate.

1. Calculate change in water temperature, $\Delta T$, for each sample, by subtracting the initial temperature from the final temperature ($\Delta T = t_{final} - t_{initial}$), or g – f, and record in Table 1.

2. Calculate the mass (in g) of the water heated for each sample. Subtract the mass of the empty flask from the mass of the flask plus water, (e – d), and record in Table 1.

3. Use the results of Steps 1 and 2 to determine the heat energy gained by the water (in J).

Use the equation $\Delta H = mC\Delta T$, where $\Delta H$ = heat absorbed (in J), $C$ = specific heat capacity (4.18 J/g°C for water), $m$ = mass of the water heated (in g), and $\Delta T$ = change in temperature (in °C).

$$\Delta H_{Crouton} = (109.56g)\left(\frac{4.184\,J}{g \cdot °C}\right)(5.3\ °C) = 2{,}429.52\ J$$

$$\Delta H_{Peanut} = (114.05g)\left(\frac{4.184\,J}{g \cdot °C}\right)(21.2\ °C) = 10{,}116.33\ J$$

4. Calculate the mass (in g) of each food sample burned, by subtracting the initial mass from the final mass (c – b), and record in Table 1.

5. Use the results of Steps 3 and 4 to calculate the empirical energy content per gram (in J/g) of each food sample, and record in Table 1.

$$\frac{J_{Crouton}}{g} = \frac{2{,}429.52\ J}{0.63\ g} = 3{,}856.38\frac{J}{g} \qquad \frac{J_{Peanut}}{g} = \frac{10{,}116.33\ J}{0.47\ g} = 21{,}524.11\ \frac{J}{g}$$

6. For each food, calculate the percent error by comparing the empirical value you obtained with its accepted value (% error = [theoretical value - empirical value]/ theoretical x 100).

$$\%\ error_{Crouton} = \frac{20{,}920.00\frac{J}{g} - 3{,}856.38\frac{J}{g}}{20{,}920.00\frac{J}{g}}\ x\ 100 = 81.57\%$$

$$\%\ error_{Peanut} = \frac{24{,}685.60\ \frac{J}{g} - 21{,}524.11\ \frac{J}{g}}{24{,}685.60\ \frac{J}{g}}\ x\ 100 = 12.81\%$$

7. Is your empirical value close to theoretical for energy content per food? Why or why not? What could be done to obtain a lower % error?

*The empirical value for the crouton was very far off from its theoretical value. Perhaps a lot of the mass of the crouton is burned up before it actually catches on fire? The empirical value for the peanut was very similar to the theoretical value. The peanut caught on fire pretty quickly.*

8. Which of the foods has the greatest energy content? Why do these foods have the greatest energy content?

*The peanut had much more energy content per gram than the crouton did. This is because the peanut has more fat content, and dietary fat has more than double the amount of energy per gram than either protein or carbohydrate.*

9. Determine the number of *Calories* per gram each food you analyzed has.

$$Crouton = \left(3{,}856.38\frac{J}{g}\right)\left(\frac{1\ calorie}{4.184\ J}\right)\left(\frac{1\ kcal}{1{,}000\ calories}\right) = 0.922\frac{kcal}{g}$$

$$Peanut = \left(21{,}524.11\ \frac{J}{g}\right)\left(\frac{1\ calorie}{4.184\ J}\right)\left(\frac{1\ kcal}{1{,}000\ calories}\right) = 5.144\ \frac{kcal}{g}$$