Calorimetry

1. What is calorimetry?

   An experimental method that allows one to determine the amount of heat energy a system gives/takes during the course of a reaction.

2. What type of calorimeter maintains constant pressure?

   Coffee Cup Calorimeter

   a. $\Delta H = nCp\Delta T = q_p$ or $\Delta H = mCp\Delta T = q_p$

3. What type of calorimeter maintains constant volume?

   Bomb Calorimeter

   a. $\Delta E = nCv\Delta T = q_V$ or $\Delta E = mCv\Delta T = q_V$

4. A biological experiment requires water to be 37.0°C. The temperature of cold water is 22.0°C and the temperature of hot water is 55.0°C. If a student starts with 90.0g of cold water, what mass of hot water must be added to reach 37.0°C?

   This is a very common style of calorimetry problem. In it, we are putting a hot and cold object/solution together and allowing the temperatures to equalize. Whenever we are dealing with calorimetry our main focus is that on q – the heat transferred. What we must understand in this type of problem is that, assuming no heat loss to the surroundings, all of the heat energy lost by the hot system is transferred completely over to the cold system. . This leads us to the following, very important relationship:

$$q_C = -q_H$$

We put a negative sign in front of the q related to the hot system because it loses the same heat that the cold system gains. The value is of the same magnitude, but opposite sign.

Based on the fact that we are interested in q we will use the following equations to solve this particular problem.

$$q_C = m_C C \Delta T_C$$

$$q_H = m_H C \Delta T_H$$

We are interested in the equation hat relates q to mass because we are given and asked about the mass of substances present in solution.

Based on the original relationship established between the two q values we get the equation:

$$m_C C \Delta T_C = - m_H C \Delta T_H$$

We will now evaluate the information given and plug it into the equation above.

	COLD	HOT
initial	$T_{Cold} = 22.0^{o}C$	$T_{Hot} = 55.0^{o}C$
final	$T_{Cold} = 37.0^{o}C$	$T_{Hot} = 37.0^{o}C$
mass	$m_C = 90.0g$	$m_H = ?$

Remember that both systems will ultimately have the same final temperature.

Now we can plug out data into the equation.

$$(90.0\text{ g})(37.0^{\circ}\text{C} - 22.0^{\circ}\text{C}) = -(m_H)(37.0^{\circ}\text{C} - 55.0^{\circ}\text{C})$$

$$(90.0\text{ g})(15^{\circ}\text{C}) = (m_H)(18^{\circ}\text{C})$$

$$m_H = \boxed{75.0\text{g}}$$

5. A 0.1964g sample of $C_6H_4O_2$ is burned in a bomb calorimeter that has a heat capacity of 1.56kJ/°C. The temperature of the calorimeter increases by 3.2°C. Calculate the energy of combustion of $C_6H_4O_2$ per gram.

A bomb calorimeter is one in which no work can be done (the volume is constant). That means that all of the energy transferred is in the form of heat. So we can determine the heat of combustion by analyzing the amount of absorbed by the calorimeter.

An important point to not, as well, is that the equation for $q_{calorimeter}$ is a bit different - There is no mole or mass amount included.

$$q_{cal} = C_{cal}\Delta T$$

With the information provided we can plug into the equation and solve.

$$C_{cal} = 1.56\,\frac{\text{kJ}}{^{\circ}\text{C}} \qquad \Delta T_{cal} = +3.2\,^{\circ}\text{C}$$

$$q_{cal} = C_{cal}\Delta T = \left(1.56\,\frac{\text{kJ}}{^{\circ}\text{C}}\right)(3.2^{\circ}\text{C}) = 4.99\text{ kJ}$$

The same relationship that was established in the previous problem applies here. The heat absorbed by the calorimeter is equal to the heat given off by the reaction. Leading to :

$$q_{cal} = -q_{rxn}$$

Lastly we can calculate for the energy of combustion per gram of $C_6H_4O_2$

$$\text{Energy of combustion per gram} = \frac{q_{\text{combustion}}}{\text{g of } C_6H_4O_2} = \frac{-4.99\ \text{kJ}}{0.1964\ \text{g}} = -25.4\ \frac{\text{kJ}}{\text{g}}$$