Thermochemical Systems and Calorimetry

When we study energy changes, the item in question is called a **system.** Since energy always flows from high to low, every system has its s**urroundings.**

System + Surroundings = Universe

Energy exchange between a system and its surroundings can be represented as **heat lost or gained.**

Since the system and the surrounding are in thermal contact, they undergo a process represented by the following equation:

$q_{system} = - q_{surroundings}$

This is explained using the **First Law of Thermodynamics- any change in the energy of a system is accompanied by an equal and opposite change in the energy of the surroundings.**

To calculate heat energy gained or lost we have two equations we can use (so far!)

The formulae for q are:

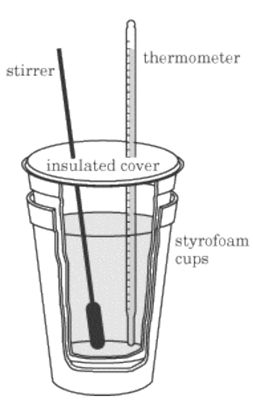
$q = mc\Delta T$ Use when a mass is referred to or a specific heat capacity is given

$q = C\Delta T$ Use when heat capacity is given or asked for

Sometimes we cannot directly measure the change in energy of a system. When this is the case we have to use the First Law of Thermodynamics and look at the surroundings. This is a branch of thermochemistry called **CALORIMETRY.**

Calorimetry is the study of the change in kinetic energy between a system and its surroundings. A calorimeter is the device used to measure the change. There are two types of calorimeters:

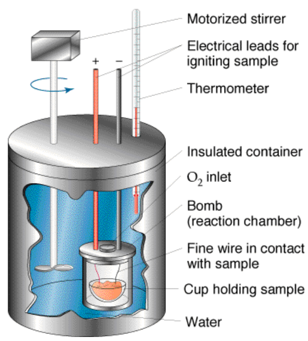
1. **Coffee cup calorimeter (aka constant pressure calorimeter)**
   It is used to measure heat changes associated with heating, cooling, phase changes, solution formation, and chemical reactions that occur in aqueous solution, like acid-base neutralization.



Important assumptions about coffee cup calorimeters

- system is isolated, no heat gets in or out
- the calorimeter itself doesn't absorb or release any heat
- the reaction or process is complete
- constant properties of the water (density, specific heat capacity)

2. **Bomb Calorimeter (aka constant volume calorimeter)**
   It is used for combustion reactions. The heat measurements are more precise and the total heat capacity of a bomb calorimeter includes the heat capacities of all the separate parts.



Important assumptions about bomb calorimeters

- No energy is lost outside
- The calorimeter absorbs all the energy of combustion

Examples:

A simple calorimeter contains 150.0 g of water. A 5.20 g piece of aluminum alloy at 525°C is dropped into the calorimeter causing the temperature of the calorimeter water to increase from 19.30°C to 22.68°C. Calculate the specific heat capacity of the alloy. (Note that the alloy sample will lose heat to the calorimeter water until a thermal equilibrium occurs. Therefore the final temperature of the alloy sample is equal to the final temperature of the calorimeter water)

Write out your "givens":

m_{Al} = 5.20g

m_{H2O} = 150.0 g

c_{H2O} = 4.184 J/g °C

T_i Al = 525 °C

T_f Al = 22.68 °C

T_i H_2O = 19.30 °C

T_f H_2O = 22.68 °C

C_{Al} = ?

$$\mathbf{q_{sys} = -q_{cal}}$$

$$(mc\Delta T)_{Al} = -(mc\Delta T)_{H2O}$$

$$c_{Al} = \frac{-(mc\Delta T)_{H2O}}{(m\Delta T)_{Al}}$$

$$= \frac{-(150.0\text{g} \times 4.184\ \text{J/g}^{o}\text{C} \times [22.68-19.30]\ ^{o}\text{C}}{5.20\ \text{g} \times [22.68-525]\ ^{o}\text{C}}$$

$$= 0.812\ \text{J/g}\ ^{o}\text{C}$$

The temperature in a simple calorimeter with a heat capacity of 1.05 kJ/°C changes from 25.0°C to 23.94°C when a very cold 12.8 g piece of copper was added to it. Calculate the initial temperature of the piece of copper. The specific heat capacity of copper is 0.385 J/g°C

Given:

C_{cal} = 1.05 kJ/°C*

T_i cal = 25.0 °C

T_f cal = 23.94°C

T_f Cu = 23.94 °C

m_{Cu} = 12.8 g

c_{Cu} = 0.385 J/g°C*

NOTE: C is in kJ/°C, but

C is in J/g°C, so you convert

$C_{cal} = 1.05 \times 10^3 J/^oC$

$\mathbf{q_{sys} = -q_{cal}}$

$(mc\Delta T)_{Cu} = -(C\Delta T)_{cal}$

$$\Delta T = \frac{-(C\Delta T)_{cal}}{(mc)_{Cu}}$$

$$= \frac{-(1.05 \times 10^3 J/^oC \times [23.94-25.0]}{(12.8\ g \times 0.385\ J/g^oC)}$$

$= 225.9\ ^oC$

If $\Delta T = T_f - T_i$

then $T_i = T_f - \Delta T$

$T_i = 23.94^oC - 225.9\ ^oC$

$= -202.0\ ^oC$

Now you try these:

1. A very cold piece of silver with a mass of 78.41 g is added to a simple calorimeter that contains 150.0 g of water. The temperature of the calorimeter water changes from 19.73°C to 16.11°C. (The specific heat capacity of silver is 0.24 J/g°C.) How cold was the piece of silver?

2. A new ceramic material underwent for use as an insulator. Part of the analysis involved determining its specific heat capacity. A 20.00 g sample was heated to 200.00°C and added to a simple calorimeter with a heat capacity of 1.46 kJ/°C. The temperature in the calorimeter changed from 24.87°C to 27.15°C. Calculate the specific heat of the ceramic material.