CALORIMETRY

- When a substance undergoes a change in temperature, heat/**thermal energy** is being transferred (either lost or gained)

- There is a simple equation that can be used to calculate the amount of heat/thermal energy transferred when a substance undergoes a temperature change

$$Q = mc\Delta t$$

where Q is the thermal energy gained or lost by a substance (J)
m is the mass of the substance (g)
c is the specific heat capacity of the substance ($\frac{J}{g \cdot {}^{o}C}$)
$\Delta t = t_f - t_i$ → Δt is the change in temperature of the substance

- The *thermal energy* is also a measure of the *kinetic energy* of the molecules in the substance

* important to understand meaning

- The **specific heat capacity** (c) is the quantity of energy required to raise the temperature of one gram of substance by one degree Celsius. This explains the units of J/(g·°C).
    - A substance with a large specific heat capacity requires a large amount of energy to be transferred before it has a change in temperature
    - The specific heat capacity varies for each different substance. A list of specific heat capacities for certain substances is found on pg. 3 of data book.

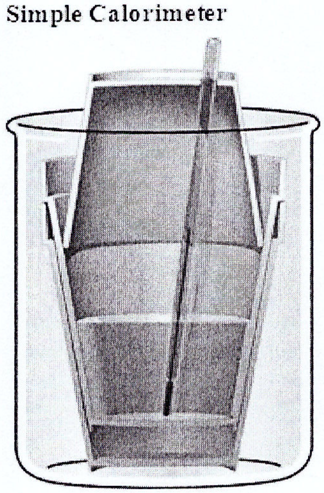
*****Now try Specific Heat Capacity Practice Problems*****

- **Calorimetry** is a technological process of measuring the energy changes of an isolated system. In other words, calorimetry is a process that is used to measure the heat released or absorbed by a reaction.
    - This process is carried out using an instrument called a **calorimeter**

How a Calorimeter Works

- The calorimeter is the isolated system (or as close to an isolated system as possible)
- The *chemical system* (reactants) being studied is placed inside the calorimeter
- The chemical system is surrounded by a known amount of liquid (usually water) which is called the *surroundings*
- As the reaction occurs, the potential energy of the chemical bonds is being converted into the kinetic energy/thermal energy of the surroundings (ie. indicated by a change in temperature)
- The assumption is that the calorimeter is an isolated system and that all the energy is transferred between the chemical system and the surrounding liquid.
    - Ultimately, calorimeters will lose some energy to the outside

Simple Calorimeter



A simple laboratory calorimeter consists of an insulated container made of three nested polystyrene cups, a measured quantity of water, and a thermometer. The chemical system is placed in or dissolved in the water of the calorimeter. Energy transfers between the chemical system and the surrounding water are monitored by measuring the changes in the water temperature.

Analyzing Energy Changes from a Calorimeter

- As the chemical reaction takes place, there are two options for energy transfer
    - The chemical system loses energy to the surroundings, causing the surroundings to gain energy and increase in temperature. exothermic
    - The chemical system absorbs energy from the surroundings, causes the surroundings to lose energy and decrease in temperature. endothermic

* • The chemical/potential energy (ΔH) lost by the chemical system is *equal* to the kinetic/ thermal energy (Q) gained by the surrounding liquid (and vice versa)

CHEMICAL SYSTEM		CALORIMETER/SURROUNDINGS
$\Delta E_{potential}$	=	$-\Delta E_{kinetic}$
ΔH	=	$-Q$
$n\Delta rH$	=	$-(mc\Delta t)$

$\Delta H = n\Delta_r H$

$Q = mc\Delta t$

where Q is the thermal energy of the surroundings (J)

m is the mass of the surroundings (g)

remember 1.0mL of water = 1.0g of water

$\Delta t = t_f - t_i$

Δt is the change in temperature of the surrounds (°C)

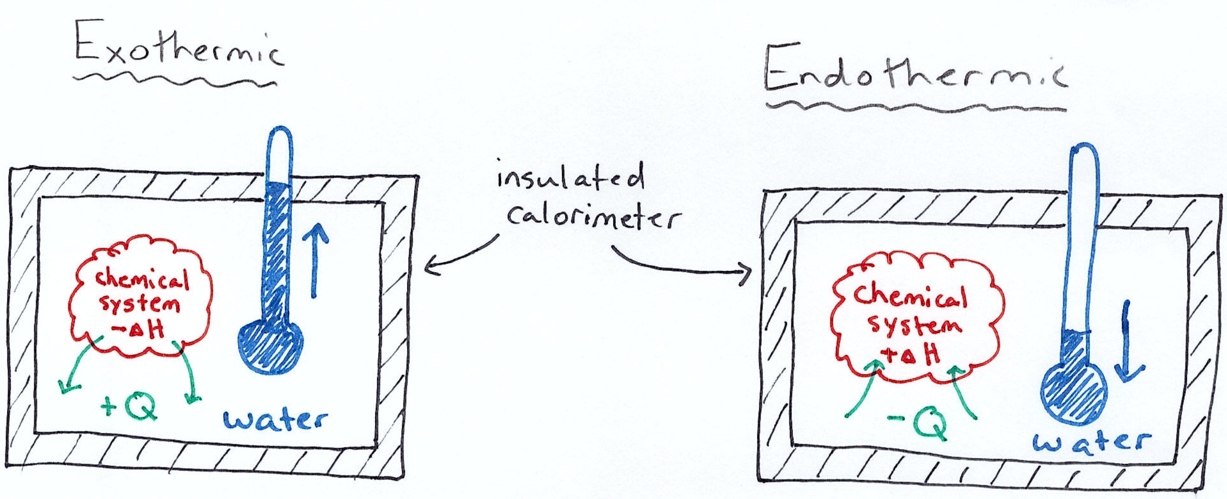
c is the specific heat capacity of the surroundings [J/(g·°C)]

* always in joules b/c of

n is the moles of chemical undergoing the reaction (mol)

$\Delta_r H$ is the molar enthalpy change for a specific chemical undergoing the reaction (kJ/mol or J/mol)

watch conflict of units!



EXAMPLES: calorimetry

1. A 0.160 g sample of ethanol (C_2H_5OH) is burned and all the heat released is absorbed by 75.42 g of water. The temperature of the water rises from 26.9°C to 41.2°C. Use this information to calculate the molar heat of combustion of ethanol ($\Delta_c H$).

Chemical	Surroundings
$m = 0.160\ g$	$m = 75.42\ g$
$\Delta_c H = ?$	$t_i = 26.9°C$
	$t_f = 41.2°C$
	$c = 4.19\ J/(g \cdot °C)$

$$\Delta H = -Q$$

$$n\Delta_c H = -(mc\Delta t)$$

aside

$$n = \frac{m}{M} = \frac{0.160 g}{46.08 g/mol}$$

$$n = 0.00347\bar{2}\ mol$$

$$\therefore \Delta_c H = \frac{-(mc\Delta t)}{n}$$

$$\Delta_c H = \frac{-(75.42 g)(4.19)(41.2°C - 26.9°C)}{0.00347\bar{2}\ mol}$$

$$\Delta_c H = -1.3014\ldots \times 10^6\ J/mol$$

$$\boxed{\Delta_c H = -1.30 \times 10^6\ J/mol \quad \text{or} \quad -1.30 \times 10^3\ kJ/mol}$$

2. A sample of NH_4NO_3 having a mass of 30.5 g is dissolved in water in an insulated cup to make 500.0 mL of solution. The heat of solution ($\Delta_{sol}H$) is +25.7kJ/mol.

   a) Will the solution increase or decrease in temperature during the dissolving process?

   decrease in temp. b/c rxn is endothermic

   b) If the initial temperature of the water is 21.5°C, determine the final temperature after the dissolving process is complete. Assume no loss of heat to the surroundings, the density of the solution is equal to 1.00 g/mL and the specific heat capacity of the solution is 4.19 J/g°C.

   3 major assumptions we can usually make!

   Calorimetry question!

Chemical	Surroundings
$m = 30.5g$	$m = 500.0\,mL \times \left(\frac{1g}{1mL}\right) = 500.0g$
$\Delta_{sol}H = 25.7\,kJ/mol$	
$\Delta_{sol}H = 25700\,J/mol$	$t_i = 21.5°C$
	$t_f = ?$
	$c = 4.19\,J/(g \cdot °C)$

watch units!

$$\Delta H = -Q$$

$$n\Delta_{sol}H = -(mc\Delta t)$$

$$\frac{n\Delta_{sol}H}{-mc} = \Delta t$$

aside

$$n = \frac{m}{M} = \frac{30.5g}{80.06g/mol}$$

$$n = 0.38096...\,mol$$

$$\frac{(0.38096...\,mol)(25700\,J/mol)}{-(500g)(4.19)} = \Delta t$$

$$-4.6734...\,°C = \Delta t$$

but $\Delta t = t_f - t_i \Rightarrow t_f = \Delta t + t_i$

$$t_f = (-4.634...\,°C) + 21.5°C$$

$$\boxed{t_f = 16.8°C}$$

3. Aqueous copper(II) sulphate, $CuSO_{4(aq)}$, reacts with sodium hydroxide, $NaOH_{(aq)}$, in a double replacement reaction. A precipitate of copper(II) hydroxide, $Cu(OH)_{2(s)}$, and aqueous sodium sulphate, $Na_2SO_{4(aq)}$, is produced:

$$CuSO_{4(aq)} + 2NaOH_{(aq)} \longrightarrow Cu(OH)_{2(s)} + 2Na_2SO_{4(aq)}$$

limiting ∴ controls length of rxn

50.00mL of 0.300mol/L $CuSO_{4(aq)}$ is mixed with 50.00mL containing an excess of $NaOH_{(aq)}$. The initial temperature of both solutions is 21.40°C. After mixing the solutions in a simple calorimeter, the highest temperature reached is 24.60°C. Determine the enthalpy change for the reaction.

Chemical	Surroundings
$V = 50.00mL$	$m = (50.0mL + 50.0mL) \times \left(\frac{1g}{1mL}\right)$
$C = 0.300mol/L$	$m = 100.0g$
$\Delta H = ?$	$t_i = 21.40°C$
	$t_f = 24.60°C$
	$c = 4.19\ J/(g \cdot °C)$

$\Delta H = -Q$

$\Delta H = -(mc\Delta t)$

$\Delta H = -(100g)(4.19)(24.60°C - 21.40°C)$

$\Delta H = -1340.8\ J$

$\boxed{\Delta H = -1.34 \times 10^3\ J \quad or \quad -1.34\ kJ}$

*****Now try pg. 363 #13, Practice Problems #1-5 & pg. 355 #7, 9 (#7 typo: *molar* enthalpy change, in kJ/mol)*****

- For simple calorimeters made of Styrofoam cups, it is *not* necessary to account for the heat transferred to the Styrofoam calorimeter because Styrofoam is an insulator and does not transfer heat
- For metal made calorimeters, *it is* necessary to account for the heat transferred to the metal calorimeter because metal will conduct heat
    - The enthalpy change for the reaction is equal to the *total* thermal energy of both the water and the metal (ie. $\Delta H = Q_{water} + Q_{metal}$)

EXAMPLE:

When 0.75g of octane, C_8H_{18}, is burnt inside a calorimeter, the temperature of 175mL of water in a 12.0g aluminium calorimeter increases from 10°C to 45°C. Assuming the water and the calorimeter both start and end with the same temperatures, what is the molar enthalpy change for the combustion of octane?

Chemical	Surroundings	
	water	aluminium
$m = 0.75g$	$m = 175g$	$m = 12.0g$
$n = 0.00656... mol$	$t_i = 10°C$	$t_i = 10°C$
$\Delta_c H = ?$	$t_f = 45°C$	$t_f = 45°C$
	$c = 4.19\ J/(g \cdot °C)$	$c = 0.897\ J/(g \cdot °C)$

aside

$n = \frac{m}{M}$

$n = \frac{0.75g}{114.26\ g/mol}$

$n = 0.00656...\ mol$

$$\Delta H = -Q_{total} = -(Q_{water} + Q_{metal})$$

$$n\Delta_c H = -[(mc\Delta t)_{water} + (mc\Delta t)_{metal}]$$

$$\Delta_c H = \frac{-[(mc\Delta t)_{water} + (mc\Delta t)_{metal}]}{n}$$

$$\Delta_c H = \frac{-[(175g)(4.19)(45°C - 10°C) + (12.0g)(0.897)(45°C - 10°C)]}{0.00656...\ mol}$$

$$\Delta_c H = -3967181.847\ J/mol$$

$$\boxed{\Delta_c H = -4.0 \times 10^6\ J/mol \text{ or } -4.0 \times 10^3\ kJ/mol}$$

*****Now try pg.363 #14, 17 & Practice Problem #6*****

Practice Problems

1. The molar enthalpy of combustion of ethanol is -1234.8 kJ/mol. How much water can be heated from 13.5°C to 24.7°C by the combustion of 22.0g of ethanol? **[12.6kg]**

2. Butane, $C_4H_{10(g)}$, is commonly used in lighters. The molar enthalpy of combustion of butane is -2657.3 kJ/mol. If you were lost in the woods in the spring with only a butane lighter, what mass of butane would you need to heat a cup (250 mL) of water from a 10°C to 80°C? **[1.6g]**

3. A 1.75g sample of acetic acid, CH_3COOH, was burned in a calorimeter that contained 925g of water. If the temperature of the water increased from 22.2°C to 26.5°C, what is the molar heat of combustion of acetic acid? **[-572 kJ/mol]**

4. A 15.7g sample of sucrose, $C_{12}H_{22}O_{11}$, was burnt in a calorimeter that contained 850g of water initially at 21.5°C. If sucrose has a molar heat of combustion of -5639.7 kJ/mol, what will be the final temperature of the water in the calorimeter? **[94.1°C]**

5. In planning a camping trip, you decide to conduct a series of experiments to determine which food would provide the most energy for the mass that you would have to carry. List three sources of experimental error that can occur in such a calorimetry experiment. **[Sources of experimental error: heat loss to the surroundings resulting in too small of a temperature change in the water and calorimeter; failure to stir the water results in inaccurate temperature readings; evaporation of water results in incorrect mass of water used in the calculation.]**

6. A 1.000 g sample of the rocket fuel hydrazine, N_2H_4, is burned in an 850g aluminum calorimeter which contains 1200 mL of water. The temperature of both the water and the calorimeter rise from 24.62°C to 28.16°C. Calculate the molar enthalpy of combustion for hydrazine. **[- 657kJ/mol]**