Thermodynamics - Energy Relationships in Chemical Reactions:

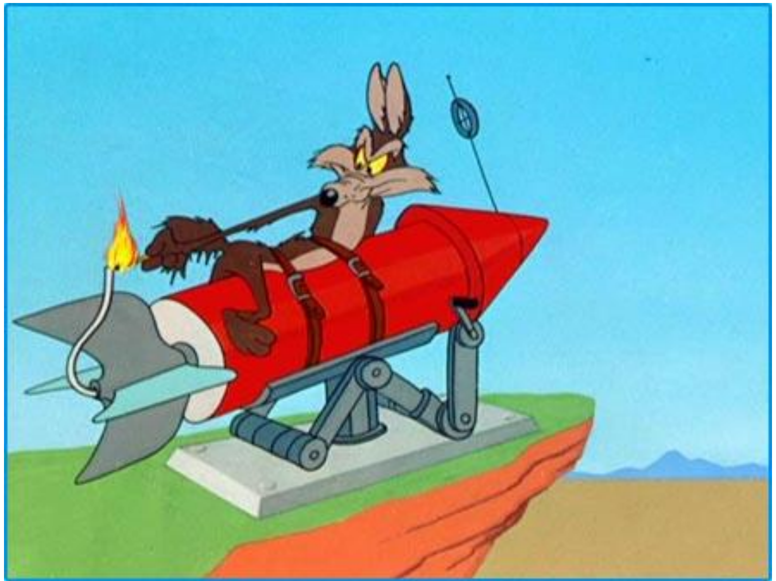
energy - The capacity to do work.

Types of Energy:

radiant-Energy from the sun.

<u>potential</u>-Energy due to an objects position.

<u>chemical</u>-Energy stored in the structural bonds of substances.



Types of Energy cont…

<u>thermal</u>-Energy associated from the random motion of atoms and molecules.

<u>heat</u>- The transfer of thermal energy between two bodies at different temperatures.

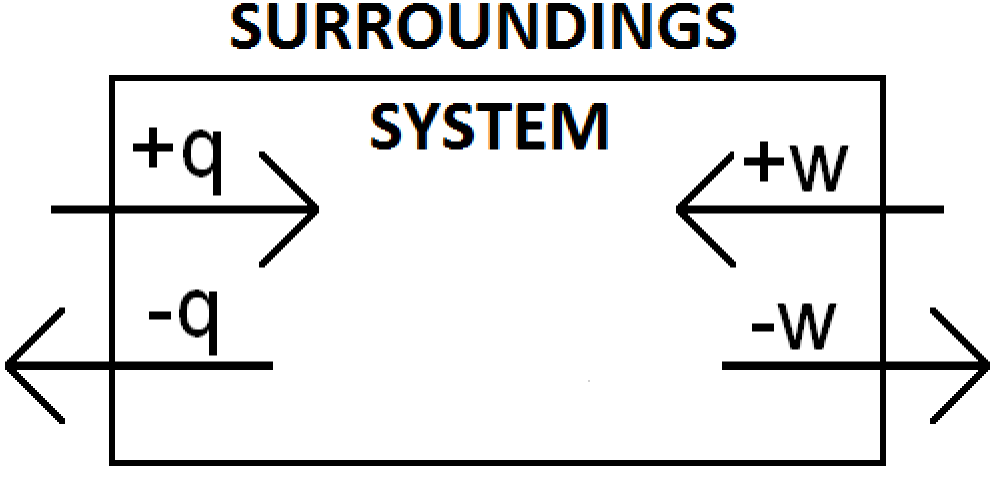
Law of Conservation of Energy - Energy is niether created nor destroyed but simply converted from one form to another.

Most energy changes involve the exchange of heat(q) or work(w).

System:

system- Part of universe of interest or under study.

surroundings- The universe outside of the system.



Types of systems:

open system- System that exchanges mass and energy(usually heat).

closed system- System that allows the exchange of energy(heat) but not mass.

isolated system- System that does not allow the transfer of energy or mass.

Exothermic/Endothermic Processes:

exothermic- Any process that gives off heat to the surroundings.

Ex: Combustion of natural gas.

$$CH_4(g) + 2O_2(g) \rightarrow CO_2(g) + 2H_2O(l) + \text{energy}$$

endothermic- Any process in which heat must be supplied.

Ex: Melting of Ice.

$$H_2O(s) \quad + \quad \text{energy} \quad \rightarrow \quad H_2O(l)$$

Enthalpy:

$$H = E + PV$$

H: heat content E: Internal energy

Can not measure H. Can measure the change in enthalpy(ΔH)

$$\Delta H = H(\text{products}) - H(\text{reactants})$$

ΔH is a measure of the heat absorbed or given off.

$\Delta H = +\#$ endothermic

$\Delta H = -\#$ exothermic

Ex:

When 1 mol $H_2(g)$ reacts with 1/2 mol $O_2(g)$ to produce 1 mol $H_2O(l)$ and 286 kJ of heat is evolved.

$$H_2(g) + 1/2O_2(g) \rightarrow H_2O(l) \qquad \Delta H = -286\ kJ$$

When the equation is reversed, the sign of ΔH is changed.

Ex:

$H_2O(s) \rightarrow H_2O(l)$ $\Delta H = +6.0$ kJ

$H_2O(l) \rightarrow H_2O(s)$ $\Delta H = -6.0$ kJ

If the coefficients in the quation are multiplied by a factor, the ΔH value must be also multiplied by the factor.

$$H_2O(s) \rightarrow H_2O(l) \quad \Delta H = +6.0\ kJ$$

$$2H_2O(s) \rightarrow 2H_2O(l) \quad \Delta H = \Delta H = 2\times{+6.0}\ kJ$$

$$\Delta H = +12\ kJ$$

Ex:

If 36.0 g of HI(g) is reacted to produce $H_2(g)$ and $I_2(s)$, how much heat is liberated?

$$2HI(g) \rightarrow H_2(g) + I_2(s) \qquad \Delta H = -51.8\ kJ$$

Hess's Law:

Hess's law states that the change in enthalpy for any chemical reaction is the same whether the reaction occurs in one step or in several steps.

$$A \rightarrow D \quad \Delta H = 3\ kJ$$

One step

$$A \rightarrow C \quad \Delta H = 1\ kJ$$

$$\underline{C \rightarrow D \quad \Delta H = 2\ kJ}$$

$$A \rightarrow D \quad \Delta H = 3\ kJ$$

Two steps

Can be used to find ΔH for reactions that can not be measured directly.

Ex:

$C(graph) + 2H_2(g) \rightarrow CH_4(g) \quad \Delta H = ?$

Given data:

1. $C(graph) + O_2(g) \rightarrow CO_2(g) \quad \Delta H = -393.5\ kJ$
2. $H_2(g) + 1/2O_2(g) \rightarrow H_2O(l) \quad \Delta H = -285.9\ kJ$
3. $CH_4(g) + 2O_2(g) \rightarrow CO_2(g) + 2H_2O(l)$

   $\Delta H = -890.4\ kJ$

Calorimetry:

Method of measuring heat changes(q).

q: heat given off or absorbed by a system.
C: Heat Capacity
Δt: final temperature - initial temperature

Heat capacity- The amount of heat required to raise the temperature of a given mass of a substance by 1°C.

Specific heat- The amount of heat required to raise the temperature 1°C for 1 g of substance.

$C = mass \times sp\_heat$

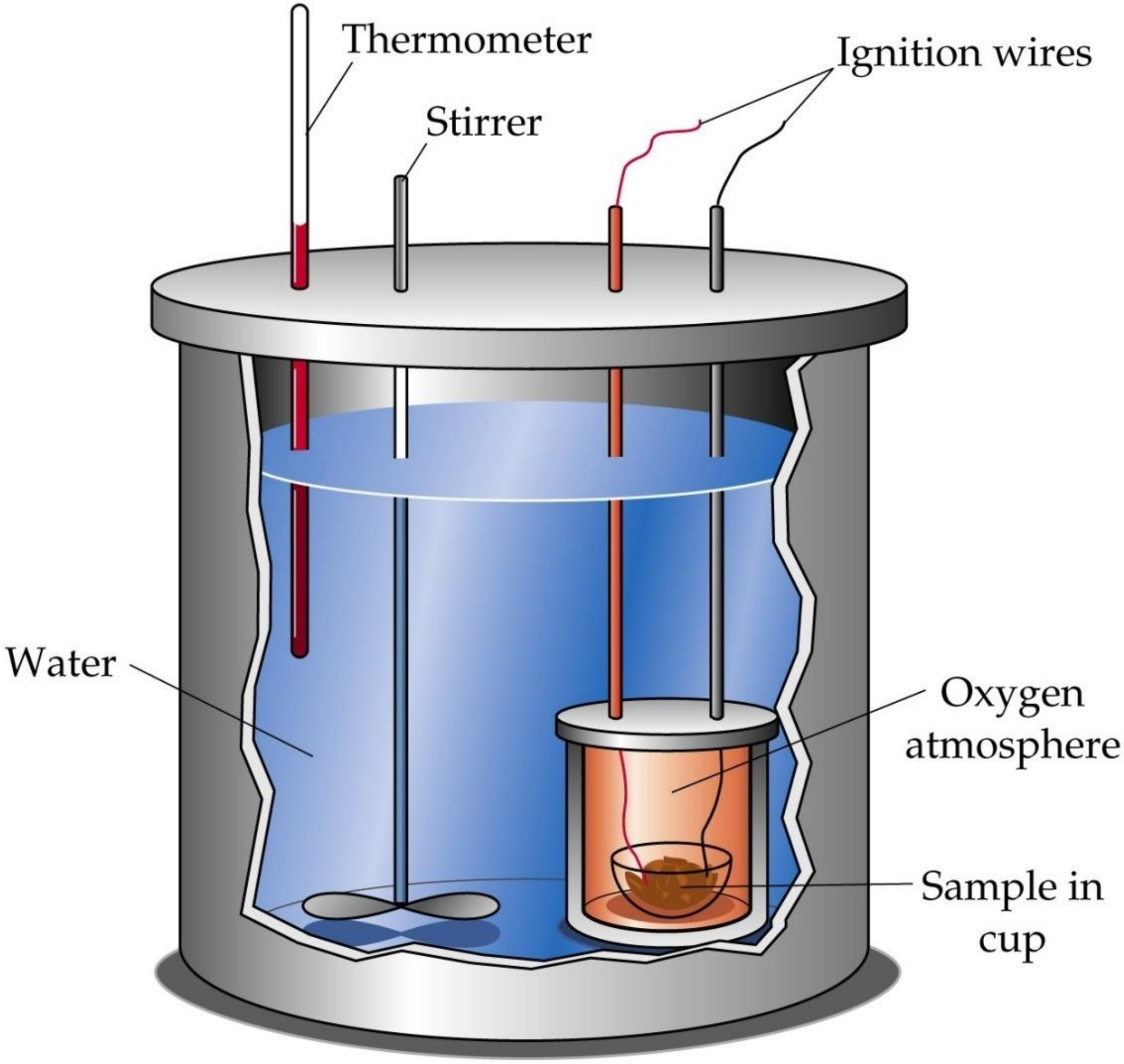
$q = C\Delta t$ or $q = (mass \times sp\_heat)\Delta t$

For water: $sp\_heat = 4.184\ J/g°C$

Ex:

How much heat is absorbed by a 125 g sample of water when it is heated from 20.00°C to 25.00°C.

Bomb Calorimetry(constant volume):



Constant Volume Calorimetry

Used to measure the heat evolved by combustion.

$$q_{sys} = q_{cal} + q_{rxn} = 0$$

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

Since $q = C\Delta t$, **thus to find** q_{rxn}

$-q_{rxn} = q_{cal}$ **where** $q_{cal} = C_{cal}\Delta t$

Ex: Combustion of glucose.

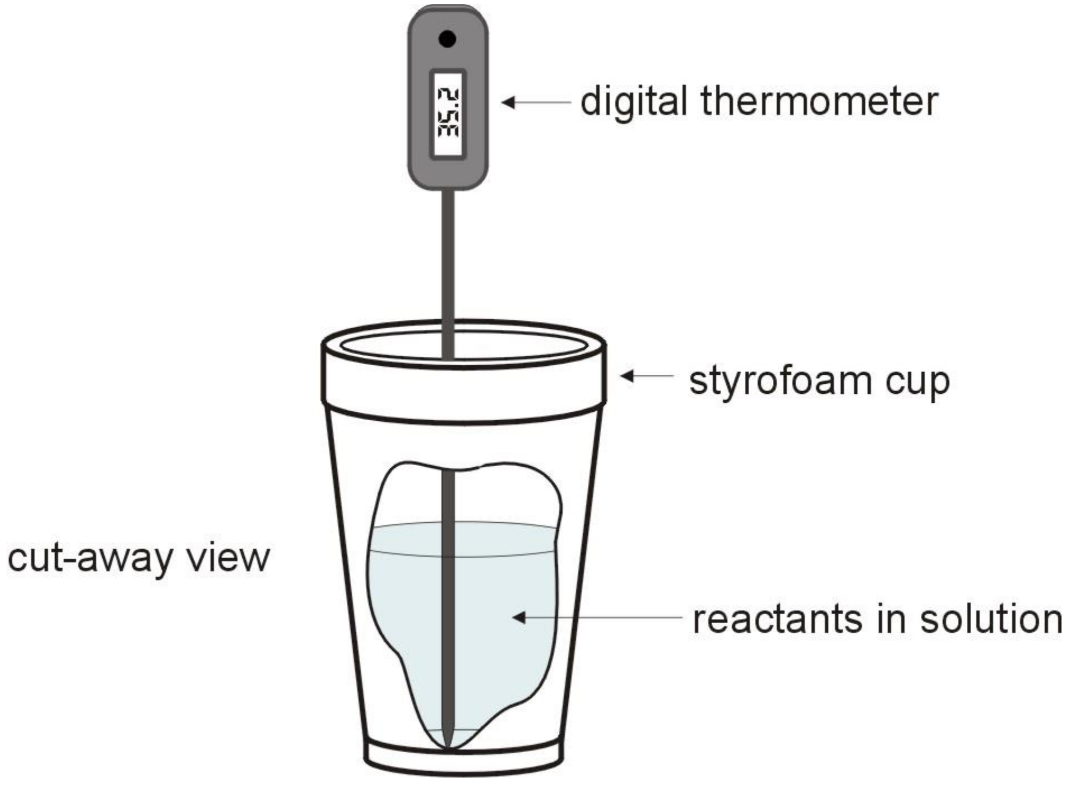
$$C_6H_{12}O_6(s) + 6O_2(g) \rightarrow 6CO_2(g) + 6H_2O(l)$$

A 3.00 g sample of glucose, $C_6H_{12}O_6$, is placed in the bomb calorimeter. The combustion causes the temperature to increase by 6.50 °C. The heat capacity of the calorimeter is 7.23 kJ/°C. Calculate the molar heat of combustion of glucose in units of kJ/mole.

Ex2: A 1.567 g sample of $C_{10}H_8$(s) is burned in a calorimeter and a temperature increase of 8.37°C is noted. The heat of combustion of $C_{10}H_8$ is 5153.9 kJ/mole. Determine the heat capacity of the calorimeter.

When 1.227 g of $C_{10}H_{14}O$(s) is burned in the same calorimeter, a temperature increase of 6.12°C is observed. What is the heat of combustion in kJ/mole of $C_{10}H_{14}O$?

Coffee Cup Calorimetry(constant pressure):



Standard Enthalpy of Formation($\Delta H_f°$):

"°" Indicates standard conditions (1 atm and reference temperature, usually 25°C)

$\Delta H_f°$: The value of ΔH that corresponds to a reaction in which 1 mole of a substance is formed from its elements in their most stable form.

For all elements in their most stable form, $\Delta H_f° = 0$

If the enthalpy of formation($\Delta H_f°$) of ethene(C_2H_4) is 52.30 kJ/mole, write the reaction equation this enthalpy change would correspond.

Can use $\Delta H_f°$ values to determine $\Delta H°$ of any reaction.

$\Delta H° = \Sigma\Delta H_f°(\text{products}) - \Sigma\Delta H_f°(\text{reactants})$

Ex:

$2NH_3(g) + 3Cl_2(g) \rightarrow N_2(g) + 6HCl(g)$ $\Delta H° = ?$

Given:

Compound	$\Delta H_f°$(kJ/mole)
$HCl(g)$	-92.3
$N_2(g)$	0
$NH_3(g)$	-46.19
$Cl_2(g)$	0