

# **Thermodynamics - Energy Relationships** **in Chemical Reactions:**

**energy - The capacity to do work.**

**Types of Energy:**  
**radiant-Energy from the sun.**

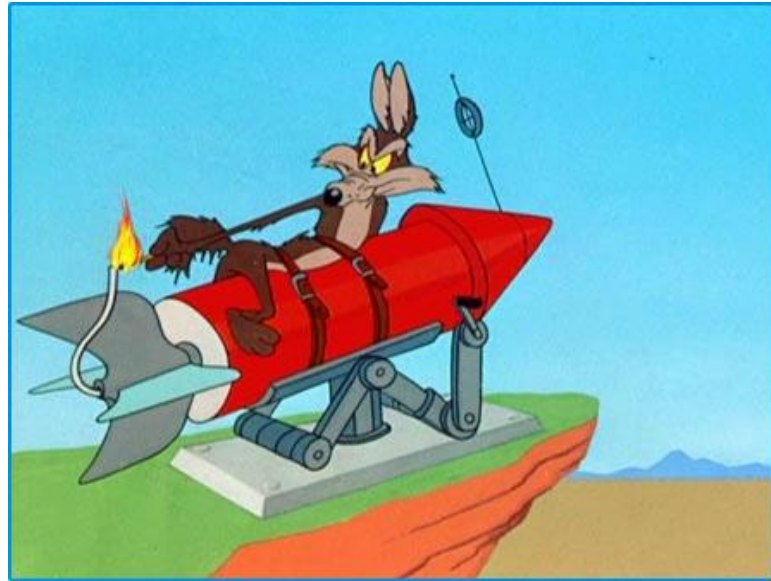


**potential-Energy due to an objects position.**



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**chemical-Energy stored in the structural bonds of substances.**



## **Types of Energy cont...**

**thermal-Energy associated from the random motion of atoms and molecules.**

**heat- The transfer of thermal energy between two bodies at different temperatures.**

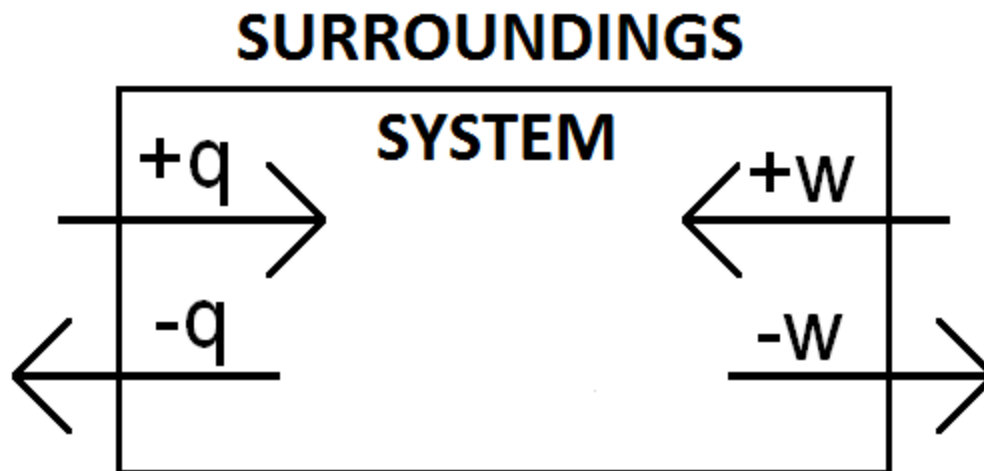
**Law of Conservation of Energy - Energy is  
neither created nor destroyed but simply  
converted from one form to another.**

**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).

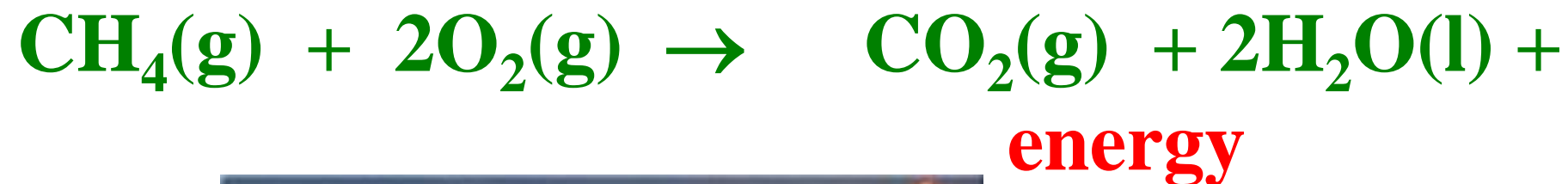
**closed system-** System that allows the exchange of energy(heat) but not mass.

**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.



**endothermic- Any process in which heat must be supplied.**

**Ex: Melting of Ice.**





# **Enthalpy:**

$$\mathbf{H = E \quad + \quad PV}$$

**H: heat content**

**E: Internal energy**

**Can not measure H. Can measure the change in enthalpy( $\Delta H$ )**

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

**$\Delta H$  is a measure of the heat absorbed or given off.**

$$\mathbf{\Delta H = +\#}$$

**endothermic**

$$\mathbf{\Delta H = -\#}$$

**exothermic**

**Ex:**

**When 1 mol H<sub>2</sub>(g) reacts with 1/2 mol O<sub>2</sub>(g) to produce 1 mol H<sub>2</sub>O(l) and 286 kJ of heat is evolved.**

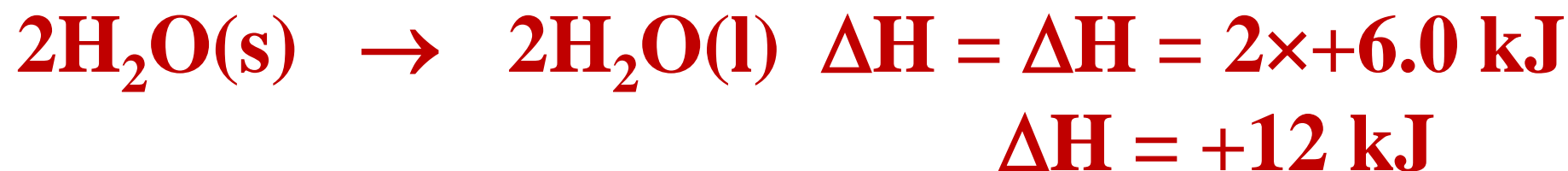


**When the equation is reversed, the sign of  $\Delta H$  is changed.**

**Ex:**



**If the coefficients in the equation are multiplied by a factor, the  $\Delta H$  value must be also multiplied by the factor.**



**Ex:**

**If 36.0 g of HI(g) is reacted to produce H<sub>2</sub>(g) and I<sub>2</sub>(s), how much heat is liberated?**



# Hess's Law:

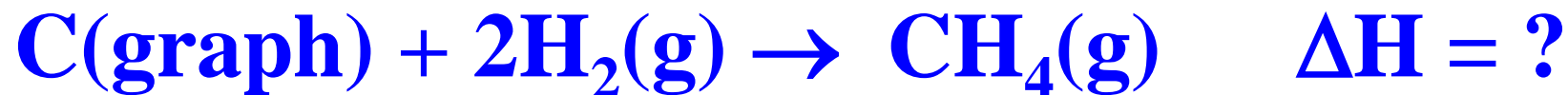
Hess's law states that the change in enthalpy for any chemical reaction is the same whether the reaction occurs in one step or in several steps.



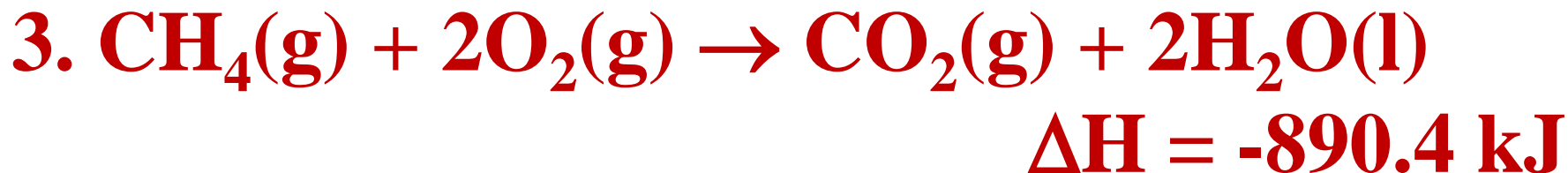
Two steps

**Can be used to find  $\Delta H$  for reactions that can not be measured directly.**

**Ex:**



**Given data:**



# **Calorimetry:**

**Method of measuring heat changes( $q$ ).**

**$q$ : heat given off or absorbed by a system.**

**$C$ : Heat Capacity**

**$\Delta t$ : final temperature - initial temperature**

**Heat capacity- The amount of heat required to raise the temperature of a given mass of a substance by  $1^{\circ}\text{C}$ .**

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

$$C = \text{mass} \times \text{sp\_heat}$$

$$q = C\Delta t \quad \text{or} \quad q = (\text{mass} \times \text{sp\_heat})\Delta t$$

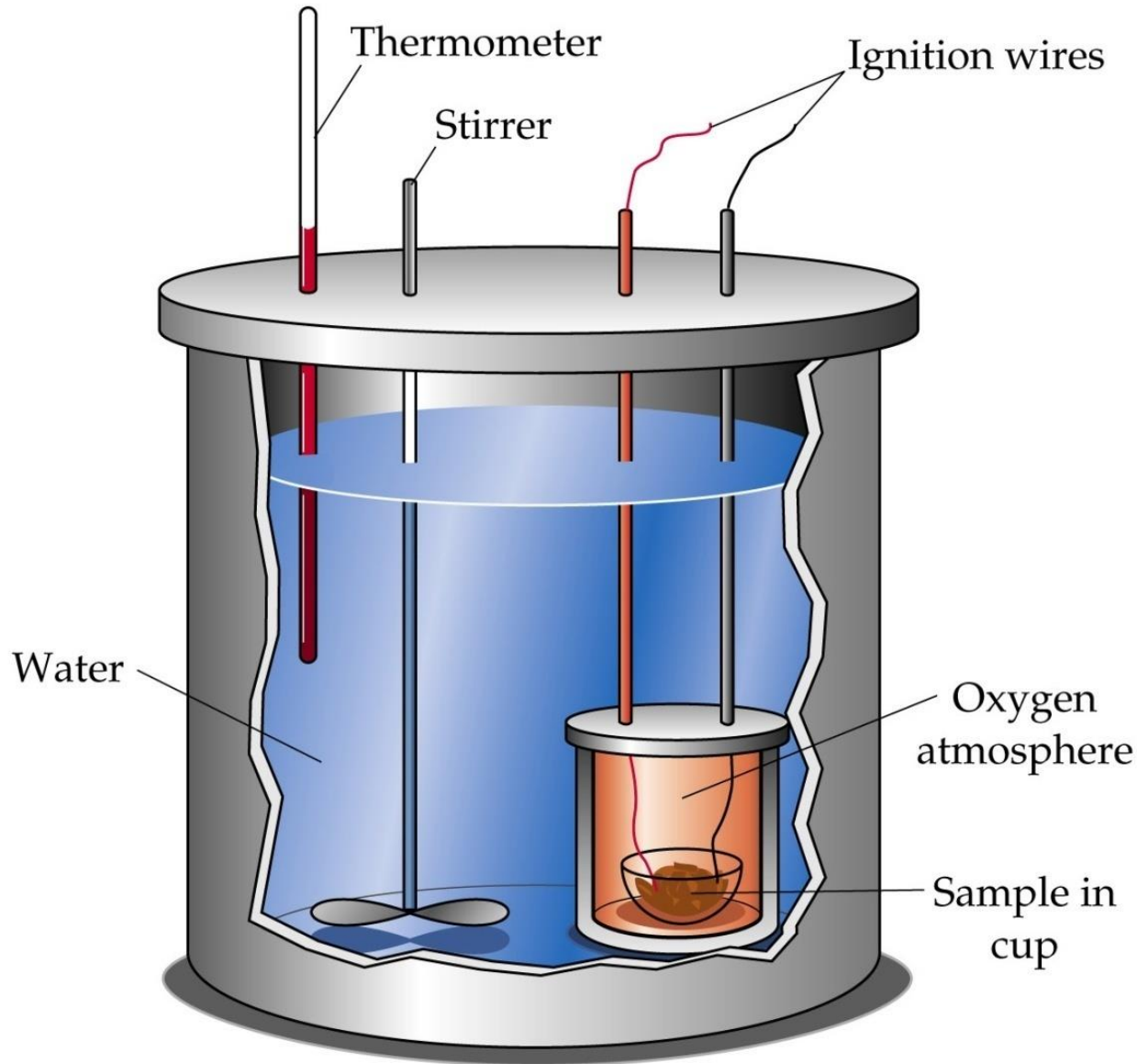
**For water: sp\_heat = 4.184 J/g°C**

**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.

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

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

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

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

**Ex: Combustion of glucose.**

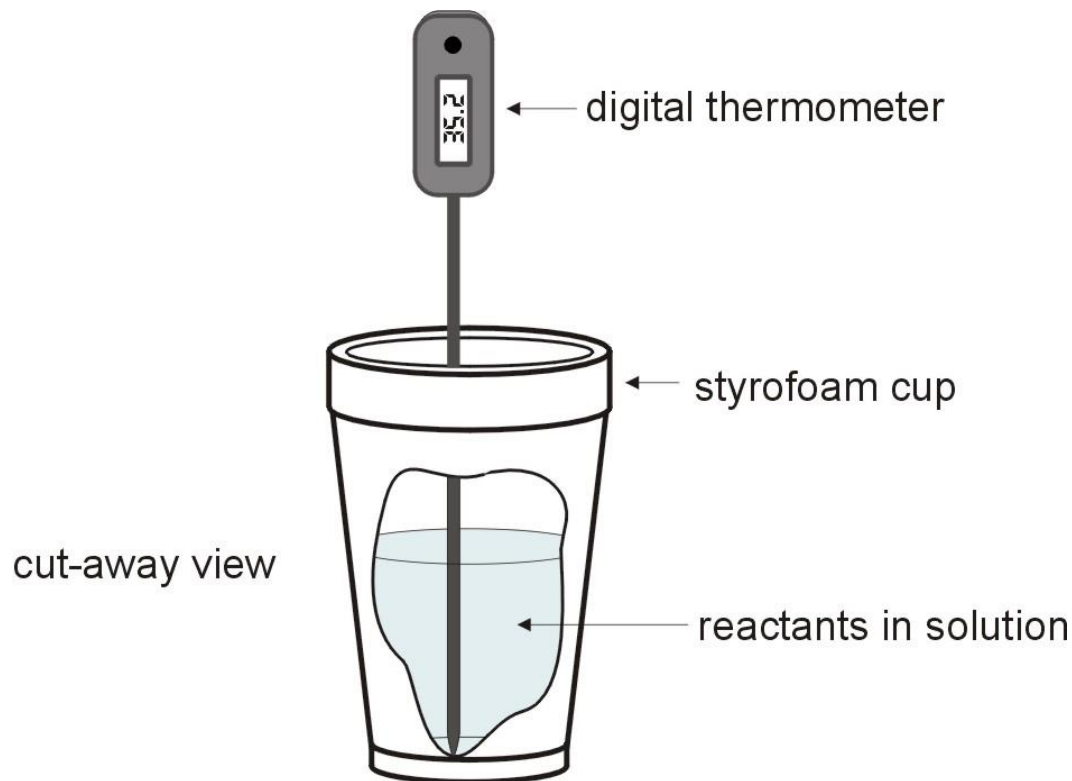


**A 3.00 g sample of glucose,  $\text{C}_6\text{H}_{12}\text{O}_6$ , is placed in the bomb calorimeter. The combustion causes the temperature to increase by  $6.50\text{ }^\circ\text{C}$ . The heat capacity of the calorimeter is  $7.23\text{ kJ}/^\circ\text{C}$ . Calculate the molar heat of combustion of glucose in units of  $\text{kJ}/\text{mole}$ .**

**Ex2: A 1.567 g sample of  $\text{C}_{10}\text{H}_8(\text{s})$  is burned in a calorimeter and a temperature increase of  $8.37^\circ\text{C}$  is noted. The heat of combustion of  $\text{C}_{10}\text{H}_8$  is 5153.9 kJ/mole. Determine the heat capacity of the calorimeter.**

**When 1.227 g of  $\text{C}_{10}\text{H}_{14}\text{O}(\text{s})$  is burned in the same calorimeter, a temperature increase of  $6.12^\circ\text{C}$  is observed. What is the heat of combustion in kJ/mole of  $\text{C}_{10}\text{H}_{14}\text{O}$ ?**

# Coffee Cup Calorimetry(constant pressure):



**Standard Enthalpy of Formation( $\Delta H_f^\circ$ ):**  
**'''°''' Indicates standard conditions (1 atm and reference temperature, usually 25°C)**

**$\Delta H_f^\circ$ : The value of  $\Delta H$  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,  
 $\Delta H_f^\circ = 0$**

**If the enthalpy of formation( $\Delta H_f^\circ$ ) of ethane is 52.30 kJ/mole, write the reaction equation this enthalpy change would correspond.**

**Can use  $\Delta H_f^\circ$  values to determine  $\Delta H^\circ$  of any reaction.**

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

**Ex:**



**Given:**

<u>Compound</u>	$\Delta H_f^\circ(\text{kJ/mole})$
HCl(g)	-92.3
N <sub>2</sub> (g)	0
NH <sub>3</sub> (g)	-46.19
Cl <sub>2</sub> (g)	0