CHEMICAL REACTIONS AND HEAT

1. Endothermic and Exothermic Reactions

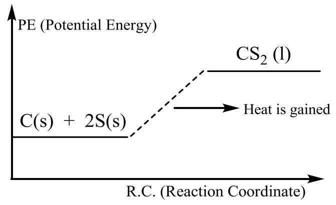
- Thermo chemistry is the study of heat changes that accompany chemical reactions and phase changes.
- In chemical reactions energy is either absorbed or released. According to this there are two types of reactions; endothermic and exothermic.

a. ***Endothermic Reactions***

- Energy is absorbed by reactants and total potential energy of reactants is smaller than that of products.

$$87.9\ kJ + C(s) + 2S(s) \longrightarrow CS_2(l)$$

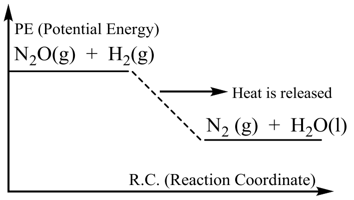
The reaction above is an example for endothermic reactions.



b. ***Exothermic Reactions***

- Energy is released by reactants and total potential energy of reactants is greater than that of products.

$$N_2O(g) + H_2(g) \longrightarrow N_2(g) + H_2O(l) + 368\ kJ$$



2. Enthalpy Change of Reactions (ΔH)

- Enthalpy (H) is the heat content of a substance at constant pressure.
- The change in enthalpy for a reaction is called the enthalpy of reaction (ΔH).

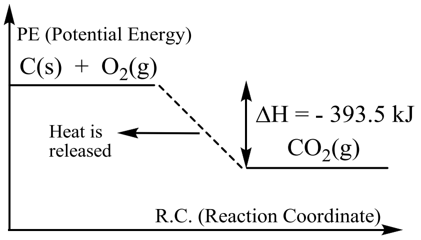
$\Delta H = \Sigma H_{products} - \Sigma H_{reactants}$

- If $\Sigma H_{products} > \Sigma H_{reactants}$, then $\Delta H > 0$ so the reaction is endothermic. Similarly,
- If $\Sigma H_{products} < \Sigma H_{reactants}$, then $\Delta H < 0$ so the reaction is exothermic.

Example 1

$$C(s) + O_2(g) \longrightarrow CO_2(g) + 393.5\ kJ$$

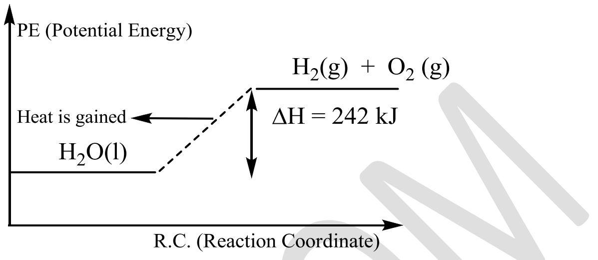
$\Delta H = -393.5\ kJ < 0$ then the reaction is exothermic.



Example 2

$$H_2O(l) + 242\ kJ \longrightarrow H_2(g) + O_2(g)$$

$\Delta H = 242\ kJ > 0$ then the reaction is endothermic.



Standard Heat of Formation (ΔH^o_f)

- The heat change when 1 mole compound is produced

from its elements in their most stable states (under 1 atm pressure and at 25°C is called as standard heat of formation, and shown by $\boldsymbol{\Delta H^o_f}$.

- $\boldsymbol{\Delta H^o_f}$ of the free atoms (K, Fe, Na, S, P, Cu…etc) and free simple molecules (O_2, N_2, Cl_2, P_4, …etc) are accepted as zero.

Heat of a reaction, ΔH^o can be calculated by using ΔH^o_f values.

$$\Delta H^o = \Sigma H^o_{f(products)} - \Sigma H^o_{f(reactants)}$$

Example 3

Find the heat of the reaction (ΔH^o)

$SO_2(g) + ½\, O_2(g) \rightarrow SO_3(g)$

by using ΔH^o_f of the compounds given

$\Delta H^o_{f(SO2)} = -297$ kJ/mol, $\Delta H^o_{f(SO3)} = -396$ kJ/mol

$\Delta H^o_{f(O2)} = 0$ kJ/mol

Solution

$\Delta H^o = \Sigma H^o_{f(products)} - \Sigma H^o_{f(reactants)}$

$\Delta H^o = \Delta H^o_{f(SO3)} - [\Delta H^o_{f(SO2)} + ½\ \Delta H^o_{f(O2)}]$

$\Delta H^o = (-396) - [(-297) + ½ \times (0)]$

$\Delta H^o = -99$ kJ

Example 4

When 2.4 g of graphite burnt with O_2 completely, 78.70 kJ heat is released. What is the molar enthalpy of the formation of CO_2 ?

Solution

$C_{(graphite)} + O_2(g) \rightarrow CO_2(g)$

2.4 g C releases 78.70 kJ

$\underline{\text{12 g (1mol) C releases}\ \ x}$

x = 393.5 kJ

$\Delta H^o = \Sigma H^o_{f(products)} - \Sigma H^o_{f(reactants)}$

$\Delta H^o = \Delta H^o_{f(CO2)} - [\Delta H^o_{f(C)} + \Delta H^o_{f(O2)}]$

$-393.5 = \Delta H^o_{f(CO2)} - [(0)+(0)]$

$\Delta H^o_{f(CO2)} = -393.5$ kJ

Example 5

The combustion reaction of propane is

$C_3H_8(g) + 5O_2(g) \rightarrow 3CO_2(g) + 4H_2O(l) \quad \Delta H^o = -2220.3$ kJ

If $\Delta H^o_{f(CO2)}$ and $\Delta H^o_{f(H2O)}$ values are -393.5 kJ/mol and -286 kJ/mol respectively find $\Delta H^o_{f(C3H8)}$?

Example 6

$Fe_2O_3(s) + 3CO(g) \rightarrow 2Fe(s) + 3CO_2(g)$

If $\Delta H^o_{f(Fe2O3)}$, $\Delta H^o_{f(CO)}$ and $\Delta H^o_{f(CO2)}$ values are -826 kJ/mol, -110.5 kJ/mol and -393.5 kJ/mol respectively find ΔH^o for the reaction.

According to types of reactions, there are different types of enthalpies of reactions. For example heat of combustion, heat of neutralization and heat of dissolution.

Example 7

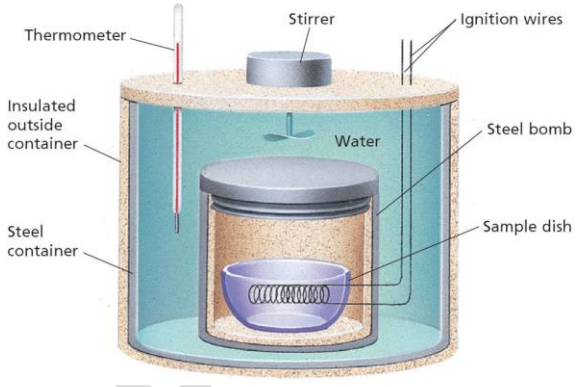
What is the molar heat of combustion of propene (C_3H_6), if 343
kJ heat is liberated when 7 g of C_3H_6 is burnt?

Example 8

When some amount of NaOH is dissolved in 200 mL of water, 8.6 kJ heat is produced. What is the concentration of the solution in M, if $\Delta H^o_{dissolution} = -43$ kJ/mol?

3. Calorimeters

Calorimeters are devices to measure heats of reactions (enthalpy change)



In a bomb calorimeter heat of the reaction is calculated by the following formula.

Heat of reaction = Heat absorbed by bomb + Heat absorbed by water

$Q_{reaction} = -[Q_{bomb} + Q_{water}]$ exothermic process

$\Delta Q_{bomb} = C_{bomb} . \Delta T$

$\Delta Q_{water} = C_{water} . \Delta T$

$\Delta Q_{reaction} = -(C_{bomb}. \Delta T + C_{water} . \Delta T)$

$= -(C_{bomb} + C_{water}) . \Delta T$

$Q_{reaction} = -C_{calorimeter} . \Delta T$

Example 9

When a 0.5 g of benzene (C_6H_6) is burnt in a bomb calorimeter, the temperature of the calorimeter rises from 25°C to 33.52°C. If the molar heat of benzene is 3189.9 kJ , find the heat capacity of the calorimeter?

Solution

$C_6H_6 + 15/2\ O_2 \rightarrow 6CO_2 + 3H_2O$

$\Delta H^o_{combustion} = -3189.9$ kJ/mol

78 g (1mol) C_6H_6 releases 3189.9 kJ

0.5 g releases x kJ

x = 20.45 kJ

$Q_{reaction} = -C_{calorimeter} . \Delta T$

- $20.45 = -C_{calorimeter} . (33.52 - 25)$
- $C_{calorimeter} = 2.4$ kj/°C

Example 10

When a 10 g of glucose ($C_6H_{12}O_6$) is burnt in a bomb calorimeter, what will be the change in the temperature of the calorimeter? (*$\Delta H^o_{combustion}$ of glucose = -2808 kJ/mol, and $C_{calorimeter}$ = 6.1 kJ/°C)*

4. Bond Energies

- The energy that is used to break up 1 mol of covalent
bond of a molecule in gaseous state is called Bond Energy.

$\Delta H = \Sigma H_{bond\ breakage} + \Sigma H_{bond\ formation}$

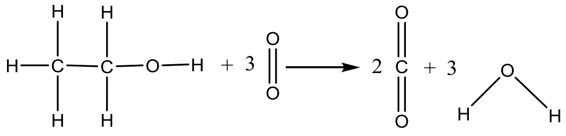
- Bond breakage is endothermic and bond formation is
exothermic processes.

Example 10

Find the molar heat of combustion of CH_3CH_2OH by using bond energies?

Solution

$CH_3CH_2OH + 3O_2 \rightarrow 2CO_2 + 3H_2O \qquad \Delta H^o_{combustion} = ?$



Solution

bond breakage	
C-C ⟶	347
C-H ⟶	5 x 414
C-O ⟶	335
O-H ⟶	463
O=O ⟶	3 x 551
	4868 kJ

bond formation	
C=O ⟶	2 x 2 x 715
H-O ⟶	2 x 3 x 463
	5638 kj

$\Delta H = \Sigma H_{\text{bond breakage}} + \Sigma H_{\text{bond formation}}$

$= 4868 - 5638$

$= -770$ kj/mol

5. Hess's Law

Hess's law states that the overall enthalpy change in a reaction is equal to the sum of the enthalpy changes for the individual steps in the process.

$P_4(s) + 10Cl_2(g) \rightarrow 4PCl_5(g) \quad \Delta H = -1596$ kJ

The reaction may occur in two steps as follows.

Step 1: $P_4(s) + 6Cl_2(g) \rightarrow 4PCl_3(g) \quad \Delta H = -1224$ kJ

Step 2: $PCl_3(g) + Cl_2(g) \rightarrow PCl_5(g) \quad \Delta H = -93$ kJ

- When equations are added or subtracted, enthalpy changes must be added or subtracted.
- When equations are multiplied by a constant, the enthalpy changes must also be multiplied by that constant.

Example 11

Calculate the enthalpy of the formation of pentane, C_5H_{12},

$5C(s) + 6H_2(g) \rightarrow C_5H_{12}(g) \quad \Delta H = ?$

using the given information below.

(1) $C(s) + O_2(g) \rightarrow CO_2(g) \quad \Delta H = -393.5$ kJ/mol

(2) $H_2(g) + \frac{1}{2}O_2(g) \rightarrow H_2O(l) \quad \Delta H = -285.8$ kJ/mol

(3) $C_5H_{12}(g) + 8O_2(g) \rightarrow 5CO_2(g) + 6H_2O(l) \quad \Delta H = -3535.6$ kJ/mol

Solution

(1) $C(s) + O_2(g) \rightarrow CO_2(g) \quad \Delta H = -393.5$ kJ/mol

(2) $H_2(g) + O_2(g) \rightarrow H_2O(l) \quad \Delta H = -285.8$ kJ/mol

(3) $5CO_2(g) + 6H_2O(l) \rightarrow C_5H_{12}(g) + 8O_2(g) \quad \Delta H = 3535.6$ kJ/mol

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$5C(s) + 6H_2(g) \rightarrow C_5H_{12}(g)$

$\Delta H = 5x-393.5 + 6x-285.8 + 3535.6$

$= -145.7$ *kJ*

Example 12

Given the reactions;

(1) $C_2H_2(g) + 5/2O_2(g) \rightarrow 2CO_2(g) + H_2O(l) \quad \Delta H = -1300$ kj/mol

(2) $C(s) + O_2(g) \rightarrow CO_2(g) \quad \Delta H = -394$ kJ/mol

(3) $H_2(g) + \frac{1}{2}O_2(g) \rightarrow H_2O(l) \quad \Delta H = -286$ kJ/mol

Calculate the enthalpy of the formation of acetylene, C_2H_2,

$2C(s) + H_2(g) \rightarrow C_2H_2(g) \quad \Delta H = ?$

Solution

(1) $2CO_2(g) + H_2O(l) \rightarrow C_2H_2(g) + 5/2O_2(g) \quad \Delta H = 1300$ kj/mol

(2) $C(s) + O_2(g) \rightarrow CO_2(g) \quad \Delta H = -394$ kJ/mol

(3) $H_2(g) + \frac{1}{2}O_2(g) \rightarrow H_2O(l) \quad \Delta H = -286$ kJ/mol

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$2C(s) + H_2(g) \rightarrow C_2H_2(g)$

$\Delta H = +1300 + 2x(-394) + (-286)$

$= -226$ kJ/mol