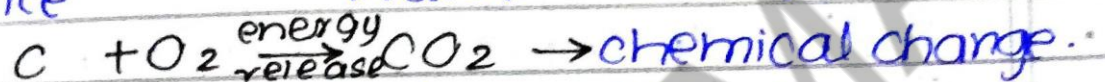
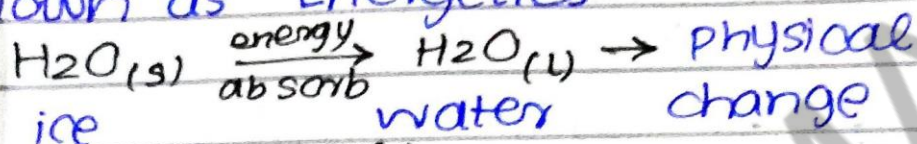


Unit #6 : Energetics

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Energetics: The study of energy changes associated with all types of physical and chemical changes is known as Energetics.



Energy in chemical reactions:

The reaction in which old bonds are broken and new bonds are formed is termed as chemical reactions. Following are bond terms.

- (i) **Bond breaking** = Absorbs Energy / or Bond consumes energy. It is always be an endothermic reaction.

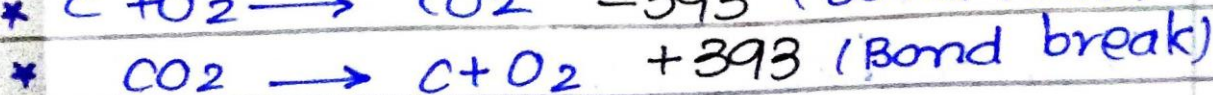
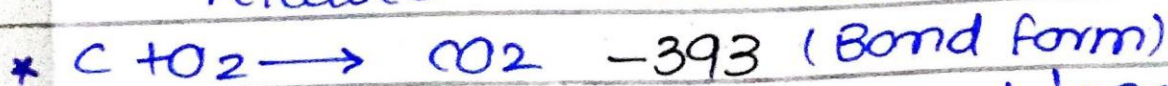
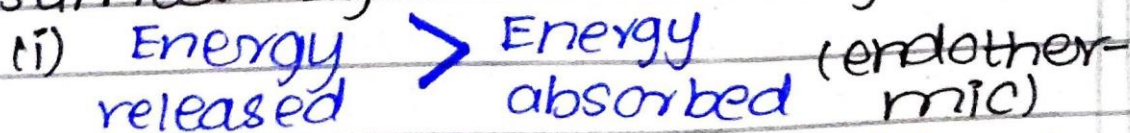
$$\Delta H = +ve$$

- (ii) **Bond formation** = Releases or emit the energy. It is always exothermic reaction, then: $\Delta H = -ve$

* **System**: The part of universe which is under observation / study.

* **Surrounding**: All the universe except the system or observation.

"When the energy released by bond formation is greater than energy consumed by bond breaking."



The energy of the bond breaking is released from system to surroundings.

(ii) Energy absorbed $>$ Energy released

The energy is transferred from surrounding to system.

Exothermic reaction

Units of thermal energy:

The unit of heat in SI system is the joule.

Thermal energy is energy of system due to movement of molecules (kinetic energy).

Non-SI unit: $1 \text{ calorie} = 4.18 \text{ J}$

The heat required to raise the temperature of 1 g of water by 1°C is called Calorie.

(1°C calorie) $14.5^\circ\text{C} \rightarrow 15.5^\circ\text{C}$ ($\Delta H = 1^\circ\text{C}$)

Thermochemical reaction:

The balanced chemical reaction which shows heat change is called thermochemical reaction.

$\Delta H = +H_1 - H_2 \rightarrow$ Final heat
heat enthalpy $\rightarrow$ initial heat

$\Delta H =$ Heat change of reaction

$\Delta H = +ve$ (Endothermic reaction)

$\Delta H = -ve$ (exothermic reaction).

The branch of chemistry that deals with heat / thermal energy changes is called the Thermochemistry.

ΔH° = standard enthalpy.

$P = 1 \text{ atm}$, $T = 25^\circ\text{C} / 298 \text{ K}$.

The reactants are products are in standard physical states.

$\text{H}_2\text{O} (\text{l})$, $\text{SO}_2 (\text{g})$.

Example: $\text{C} (\text{s}) + \text{O}_2 (\text{g}) \rightarrow \text{CO}_2 (\text{g})$ ($\Delta H^\circ = -393.5 \text{ kJ}$)
(Exothermic reaction) $\uparrow$

* $\text{CO}_2 (\text{g}) \rightarrow \text{C} (\text{s}) + \text{O}_2 (\text{g})$ ($\Delta H^\circ = +393.5 \text{ kJ}$)
(Endothermic reaction) $\uparrow$

Exothermic and Endothermic Rxn:

There are two types of reactions:

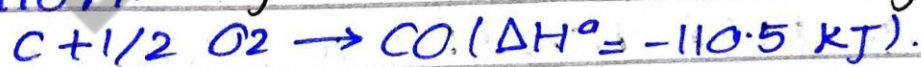
Types of thermochemical:

Following are its types:

(i) **Exothermic reaction:** A reaction chemical rxn that proceeds with evolution/release of heat.

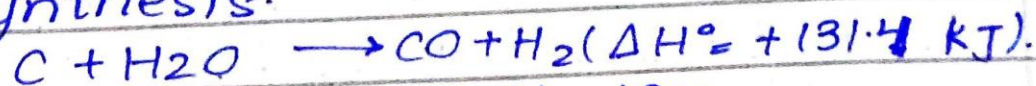
Release energy, leading to an increase in temperature of surroundings.

Such as combustion of fuels is a common example of an exothermic reaction. System $\rightarrow$ surroundings.



(ii) **Endothermic reaction:** A chemical reaction that proceeds with the absorption of heat. surroundings $\rightarrow$ system

Absorbs energy from surroundings, as a result temperature of surroundings is decreased. such as the process of photosynthesis.



Heat of Reaction / Enthalpy:

The standard conditions are:

- (i) $T = 25^{\circ}\text{C} / 298\text{ K}$
- (ii) $P = 1\text{ atm}$
- (iii) Reactant and products are in standard physical state.

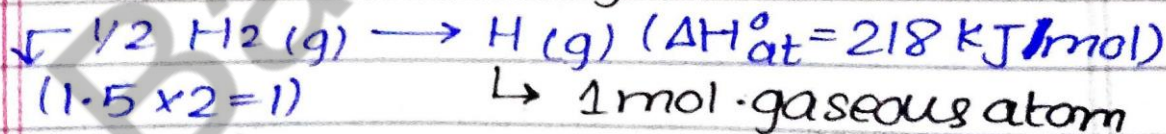
→ Heat of reaction is defined as:
 "Amount of heat released or absorbed when 1 moles of reactant (standard state) indicated by balanced chemical equation reacts with each other to form 1 mole of product."

(i) **Standard enthalpy of atomization:**

"Amount of energy absorbed when 1 mole of gaseous atom is formed from it's element under standard conditions."

$$\Delta H_{\text{at}}^{\circ} = +ve \text{ (endothermic)}$$

Product = 1 mole gaseous atom.

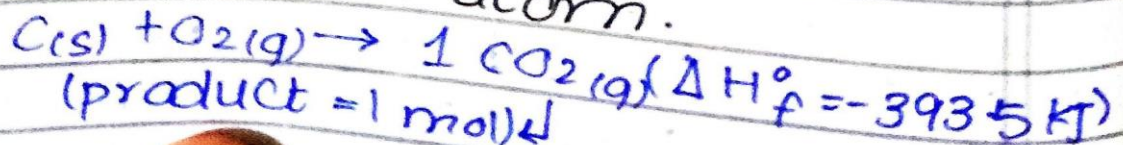


(ii) **Standard enthalpy of Formation:**

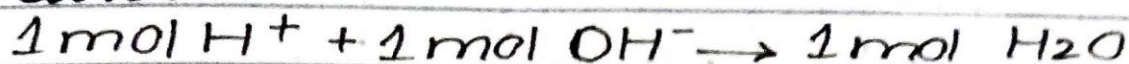
"Amount of heat released or absorbed when no. of moles of reactants (standard state) react with each other and form 1 mole of product".

$$\Delta H_f^{\circ} = +ve \text{ or } -ve \text{ reactions both}$$

Product = 1 mol atom.



(iii) Standard Enthalpy of Neutralization:

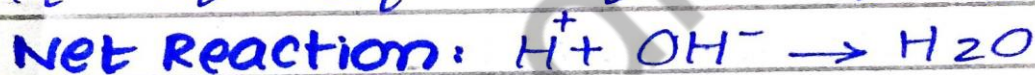
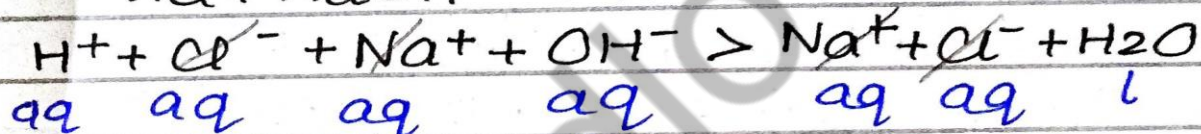
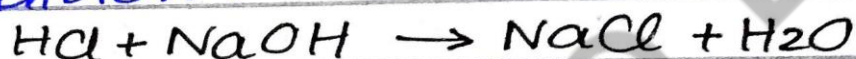


$\Delta H_n^\circ = -ve$ (Exothermic reaction).

"Amount of energy released when 1 mole of acid is and base is then neutralized and 1 mol water is formed as product".

OR

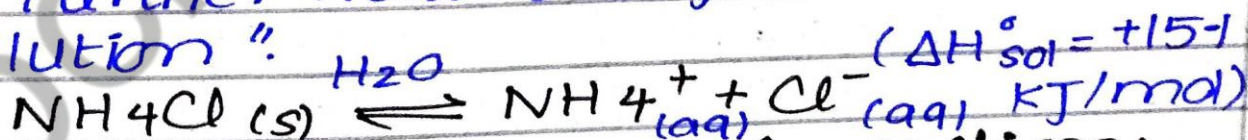
"Amount of heat released when 1 mole H^+ ion from an acid reacts with 1 mol OH^- ion from base to form 1 mol of H_2O under standard condition".



(ionic equation) ($\Delta H_n^\circ = -57.4 \text{ kJ/mol}$).

(iv) Standard enthalpy of solution:

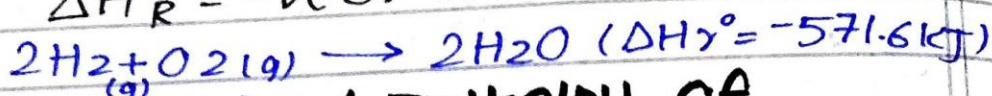
"Amount of heat absorbed/ released when 1 mol of compound is dissolved in so much solvent that there is no further heat change on further dilution".



(v) Standard enthalpy of reaction:

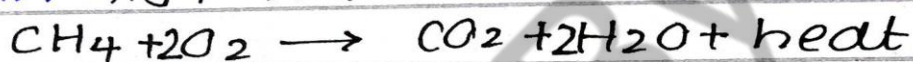
"Amount of heat change that occurs when a chemical reaction happens with all reactants and products in their standard states (298 K, 1 atm and 1 M). It can be either

$\Delta H_R^\circ = -ve$ or $+ve$ both reactions



(vi) Standard Enthalpy of combustion:

"Amount of heat change or energy released when one mole of a substance is completely burnt in oxygen under standard conditions (298 K, 1 atm). It's measured in kJ/mol. Such combustion CH_4

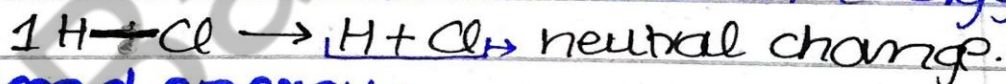


Heat is released as a result.

$$\Delta H_c^\circ = -ve \text{ (Exothermic)}$$

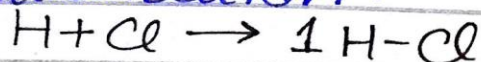
Bond dissociation energy:

"Amount of energy required to break one mole of particular bond to form neutral atoms is called bond dissociation energy".

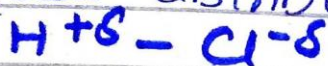


Bond energy:

"Amount of energy required to make 1 mole of bond from neutralization".



unequal distribution of electron pairs.



(Theoretical < Experimental)

As different electronegativity value. Bond strength of polar compound is greater than theoretical values of bond dissociation.

reactions
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Heat and temperature:

It can also be stated by the First law of thermodynamics.

"The law of conservation of energy states that energy cannot be produced in any physical or chemical process out can only be changed from one form to another state".

The impact of heat on body indicated by its temperature.

Heat capacity:

"Amount of heat required to raise the temperature of given amount of a substance by 1 kelvin is called heat capacity."

Unit: J/K or JK^{-1} .

$$\text{Heat capacity} = \frac{\text{heat}}{\Delta T} = \frac{q}{\Delta T}$$

Specific heat capacity:

"Amount of heat required to raise the temperature of 1g of substance by 1 kelvin is specific heat capacity".

unit: $\text{Jg}^{-1}\text{K}^{-1}$

$$\text{S.H.C} = q / \text{mass} \times \Delta T$$

Molar heat capacity:

"Amount of heat required to raise the temperature of 1 mole of substance by 1 kelvin".

unit: $\text{JK}^{-1}\text{mol}^{-1}$

$$\text{M.H.C} = q / \text{moles} \times \Delta T$$

As we know that:

→

$$q \propto \Delta T$$

$$q = \text{heat capacity} \times \Delta T$$

And now take:

→

$$\text{Heat capacity} = q / \Delta T$$

*

$$\text{specific heat } c = q / m \cdot \Delta T$$

$$q = mc\Delta T$$

so heat could be found by this equation.

Measurement of heat of reaction:

The device that measures the flow of heat is called calorimeter.

Types of calorimeter:

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There are two types:

(i) Coffee cup calorimeter:
constant pressure calorimeter.

(ii) Bomb calorimeter:

constant volume calorimeter.

Coffee cup calorimeter:

→ All enthalpies can be measured except enthalpy of combustion.

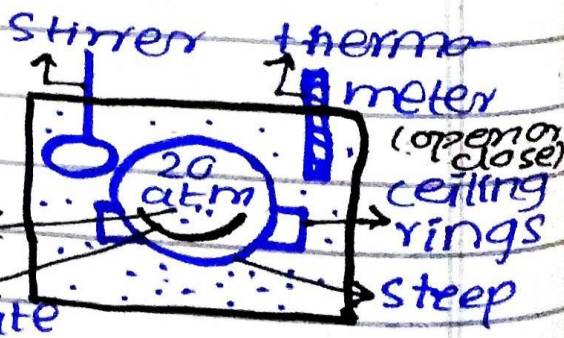


→ Most used to measure heat of neutralization.

It is called constant cup pressure calorimeter because external pressure is kept constant / same.

Bomb calorimeter:

It only measures the enthalp of combustion. food fuel organic compounds



CONCEPT ASSESSMENT # 6.3 :-



$$87.5 \text{ cm}^3 : 25 \text{ cm}^3$$

$$0.1 \text{ M} : 0.35 \text{ M}$$

$$\rho = 1 \text{ g/cm}^3$$

$$c = 4.2 \text{ J g}^{-1} \text{ K}^{-1}$$

$$\Delta T = 1.07^\circ \text{C}$$

$$T = 26.07^\circ \text{C} - 25^\circ \text{C}$$

$$\text{Volume}_{\text{net}} = V_1 + V_2 = 87.5 + 25$$

$$V = 112.5 \text{ cm}^3 / m = 112.5 \text{ g}$$

By using the formulae of heat change:

$$q = mc \Delta T$$

Put the values, now:

$$q = 112.5 \times 4.2 \times 1.07$$

$$q = 503.57 \text{ J}$$

$$[q = 0.503 \text{ kJ}]$$

Now convert cm^3 to dm^3 :

$$V_1 = 0.0875 \text{ dm}^3$$

$$V_2 = 0.025 \text{ dm}^3$$

For calculating moles of each:

$$\text{No. of moles} = \text{Molarity} \times \text{Volume}$$

$$\text{No. of moles of HCl} = M \times V$$

$$= 0.1 \times 0.0875$$

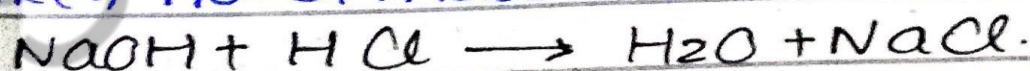
$$[\text{No. of moles} = 8.75 \times 10^{-3} \text{ mol HCl}]$$

$$\text{No. of moles of NaOH} = M \times V$$

$$= 0.35 \times 0.025$$

$$[\text{No. of moles} = 8.75 \times 10^{-3} \text{ mol NaOH}]$$

Since, no. of moles of HCl, NaOH are equal:



$$\text{So: } 8.75 \times 10^{-3} \text{ mol} = 0.503 \text{ kJ}$$

$$1 \text{ mol} \xrightarrow{\quad} x$$

As ΔH_n° always ex

$$8.75 \times 10^{-3} x = 0.503$$

$$[x = 57.48 \text{ kJ/mol}]$$

$$q_p = \Delta H_n^\circ = -57.78 \text{ kJ/mol}$$

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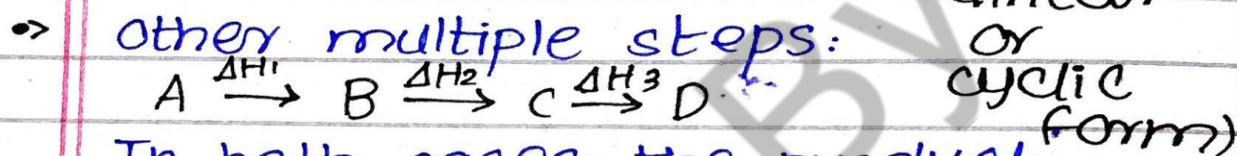
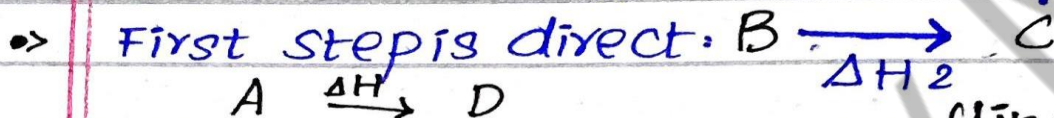
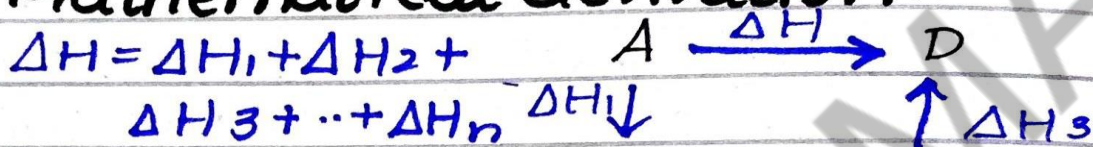
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"HESS LAW:-"

It states that:

"According to this law, enthalpy change during chemical reactions is same whether the reaction takes place in the single or more than one step".

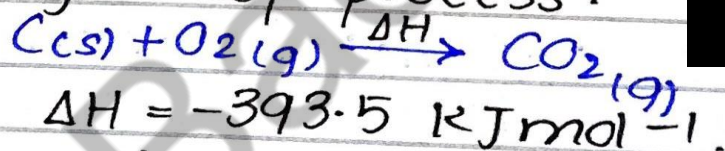
Mathematical derivation:



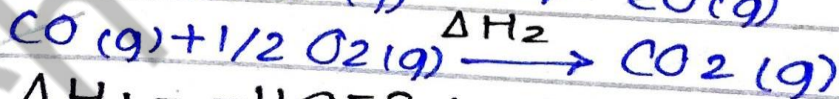
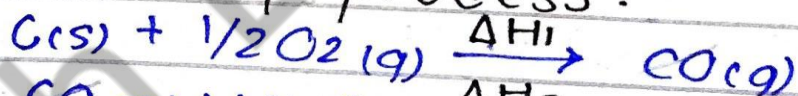
In both cases, the product and the enthalpy of the reaction remains the same.

EXPLANATION:

* Single step process:



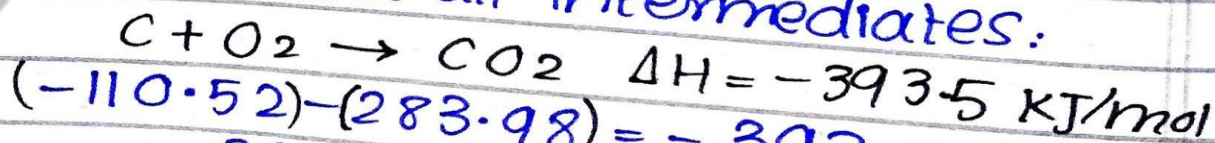
* Two step process:



$$\Delta H_1 = -110.52 \text{ kJ/mol}$$

$$\Delta H_2 = -283.98 \text{ kJ/mol}$$

Now cancel all intermediates:



$$(-110.52) - (283.98) = -393.5$$

$$-393.5 + 393.5 = 0 \quad \therefore \Delta H = \Delta H_1 + \Delta H_2$$

$$\boxed{0 = 0}$$

So, the cyclic process has $\Delta H = 0$

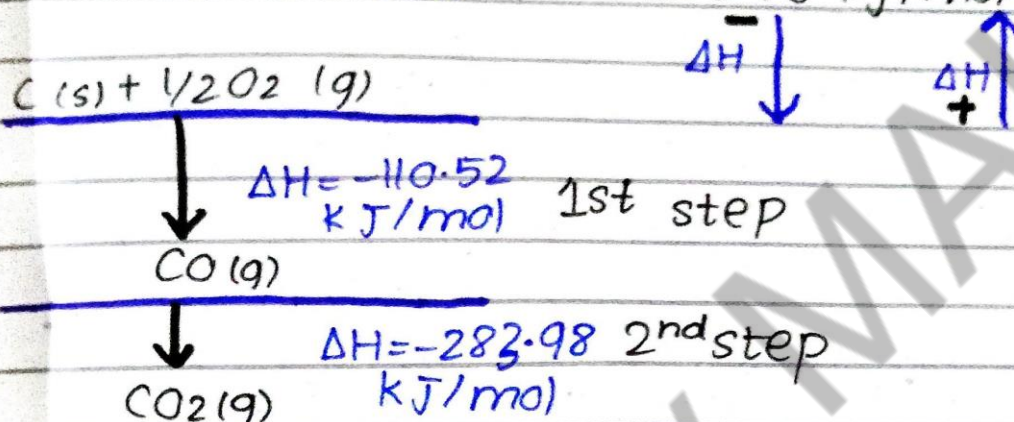
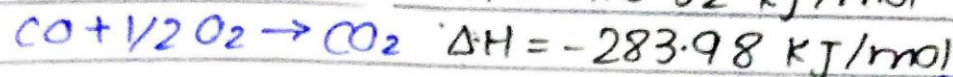
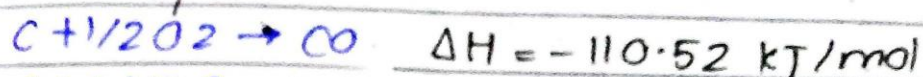
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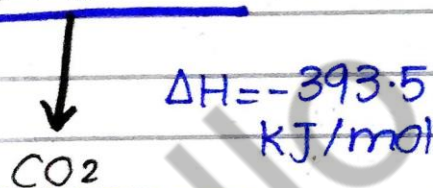
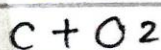
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ENERGY CYCLE OF FORMATION OF CO₂:

* 1st step:



* 2nd step:



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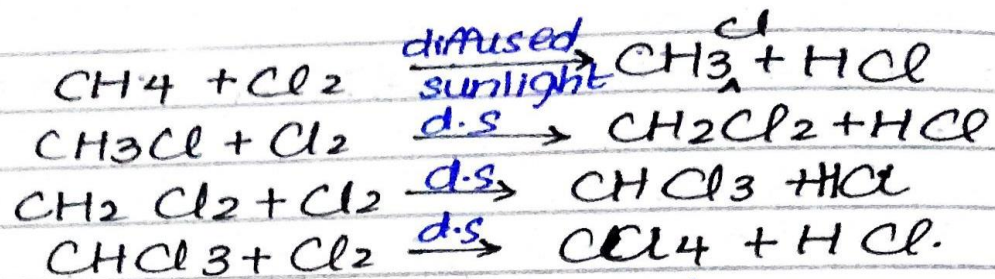
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* For endothermic process (+ve) the reaction is upwards and For exothermic process (-ve) it's downwards

Importance of Hess's law:

There are many reactions for which ΔH cannot be calculated by the calorimetric method for such reaction hess's law is used.

(i) Formation of CCl₄ cannot be determined by calorimetric method because when CH₄ reacts with Cl₂ reaction takes place in 4 step so calorimetric method with calculate ΔH for each step but cannot calculate for direct CCl₄ (carbon tetrachloride) product



Hess's law is used.

- (ii) Formation of B_2O_3 (Boron oxide) cannot be determined by calorimetric method because it contains boron in unreacted form and reaction does not complete.

So Hess's law is applied.

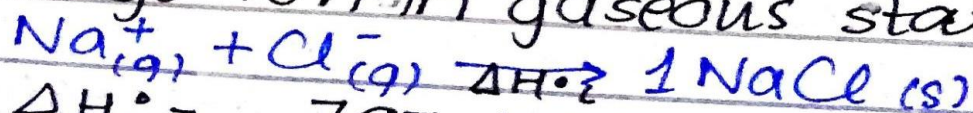
- (iii) Formation of Al_2O_3 cannot be determined by calorimetric method because it contains unreacted Al left behind.

[Al]: Al_2O_3 (tough coats)

"BORN HABER'S CYCLE:-"

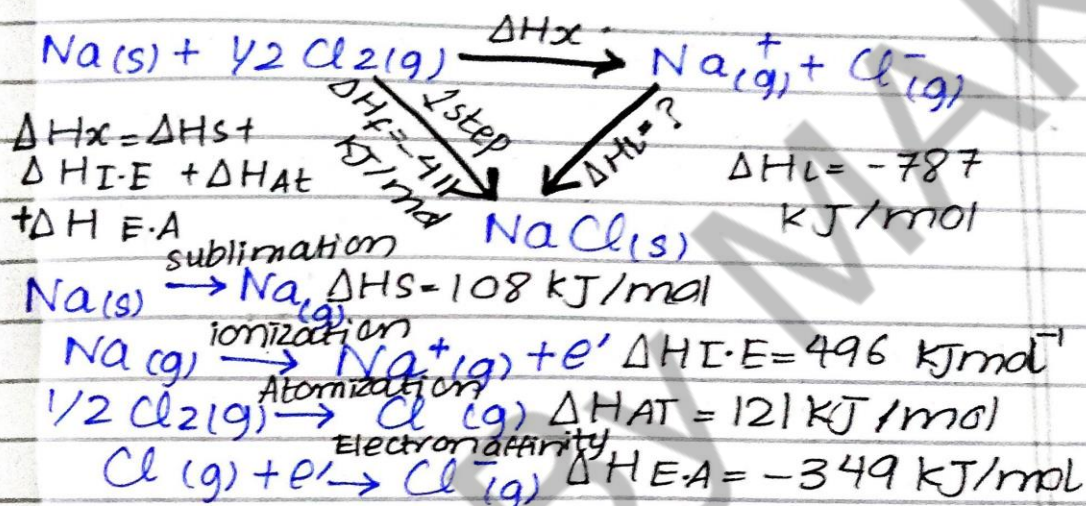
- 1- It is the application of Hess's law. ($\Delta H = \Delta H_1 + \Delta H_2 + \Delta H_n$)
- 2- Complete energy cycle.
- 3- Energy change in a cyclic process is always zero.
- 4- It is used to find lattice energy of Binary ionic compound.

Lattice energy: The statement is: "Amount of Energy released when 1 mole of ionic compound is formed from their opposite charge ion in gaseous state."



$$\Delta H_f^\circ = -787 \text{ kJ/mol.}$$

* The lattice energy is always the energy releasing like exothermic.
Born Haber's cycle for NaCl Form:



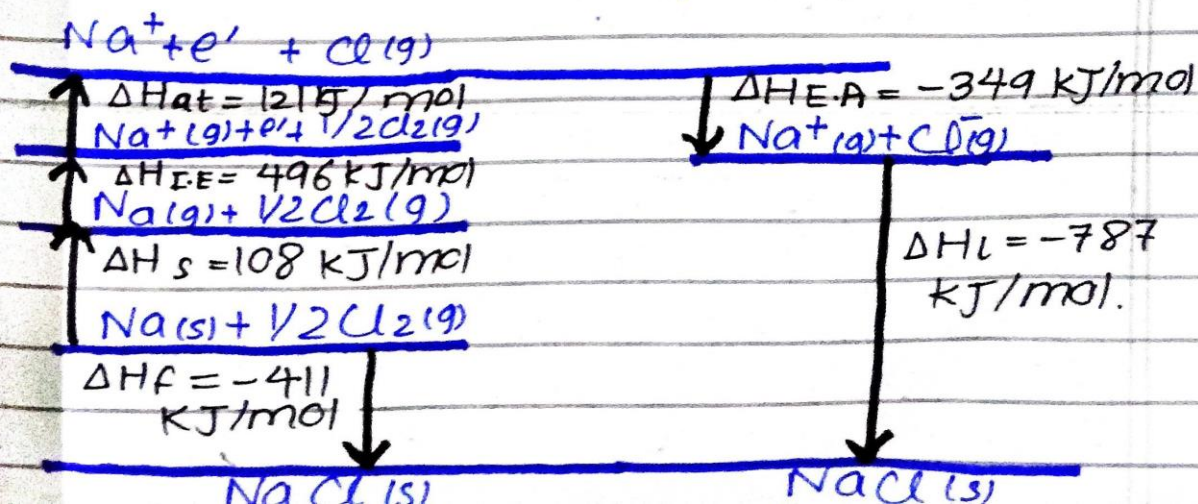
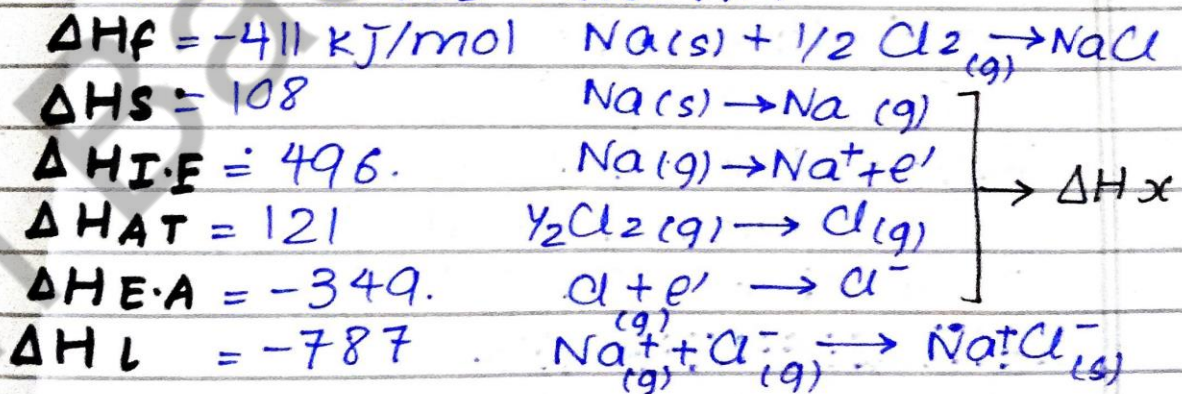
According to Hess's law:

$$\Delta H_f = \Delta H_x + \Delta H_L$$

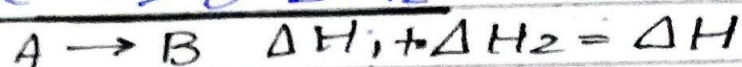
$$\begin{aligned}
 \Delta H_f &= \Delta H_s + \Delta H_{\text{I.E.}} + \Delta H_{\text{AT}} + \Delta H_{\text{E.A.}} + \Delta H_L \\
 -411 &= 108 + 496 + 121 - 349 + \Delta H_L
 \end{aligned}$$

$$\Delta H_L = -787 \text{ kJ/mol}$$

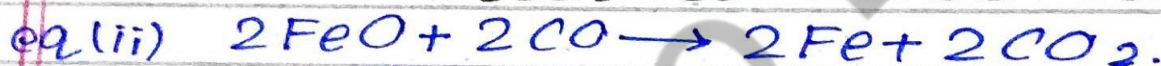
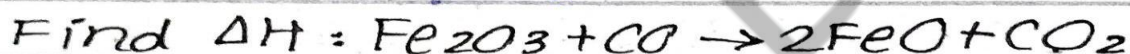
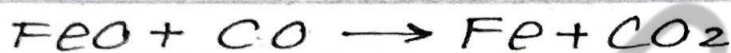
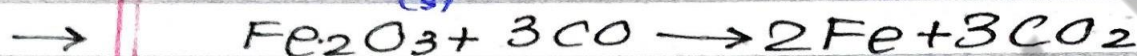
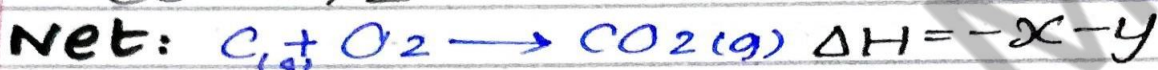
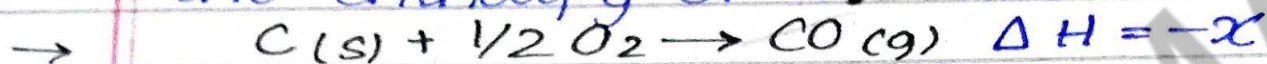
ENERGY CYCLE DIAGRAM:



"HESS'S LAW:-"



* The no. of steps do not influence the enthalpy of a reaction.

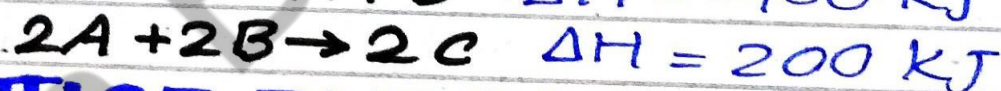
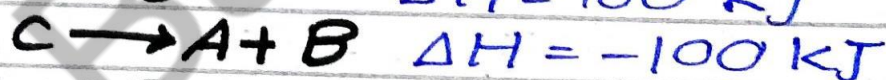
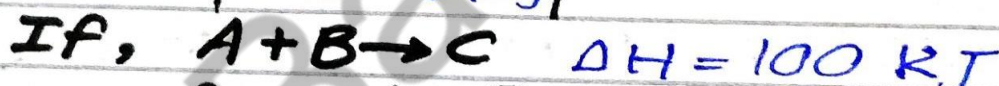


Now subtract eq (i) - eq 2 (ii).

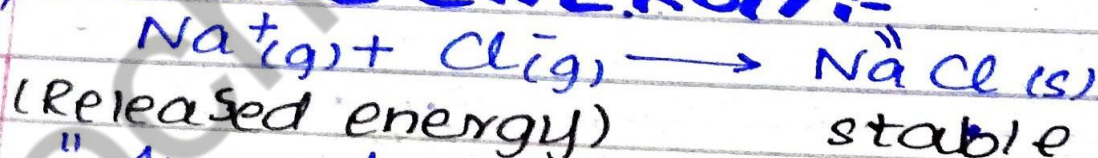


$$\Delta H = H_1 - 2\Delta H_2 = -26 - 2(-16)$$

$$[\Delta H = 6 \text{ kJ}]$$



"LATTICE ENERGY:-"



"Amount of energy released when 1 mole of a substance is formed from an anion and cationic compound". (gaseous)

It's called lattice energy. There's no direct method to

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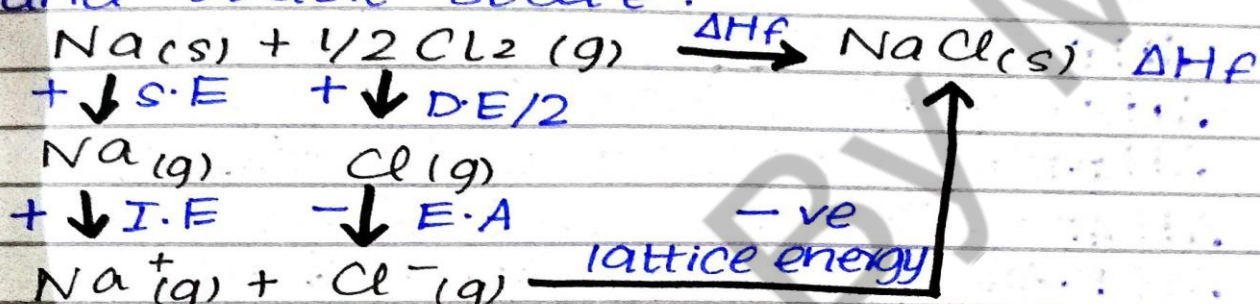
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determine the lattice energy...
This indirect method is called Born
Haber's cycle.

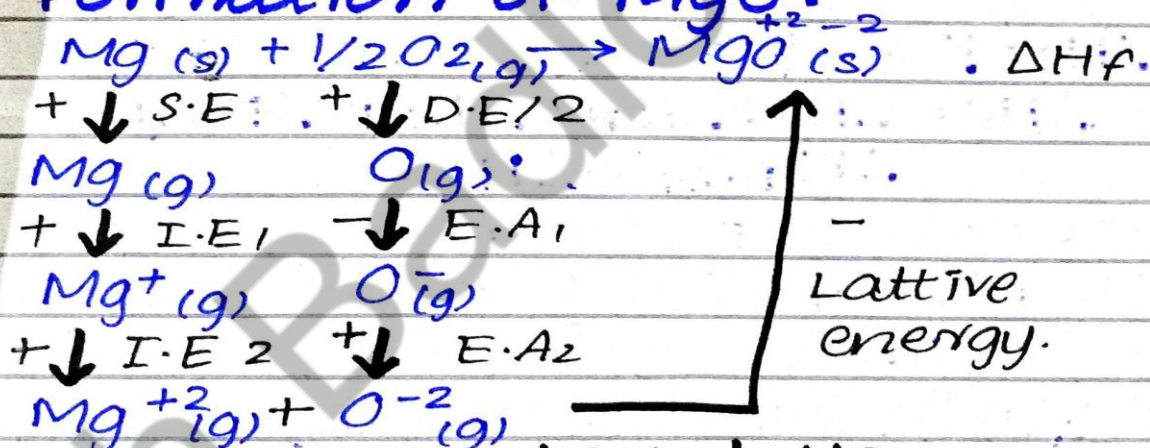
Heat of formation ΔH_f :

"The Amount of heat released when
1 mole of a compound is formed
when it's elements is in pure
and stable state".



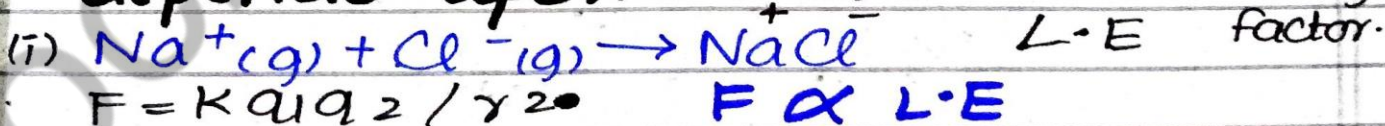
$$\Delta H_f = S.E + I.E + D.E/2 - E.A - L.E$$

Formation of MgO :



Factors on which lattice energy

depends upon: charge is dominating



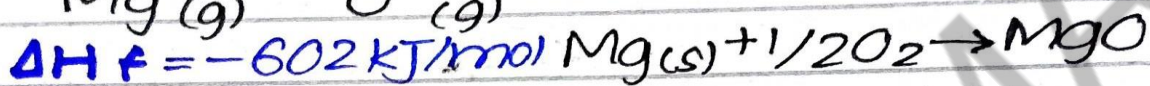
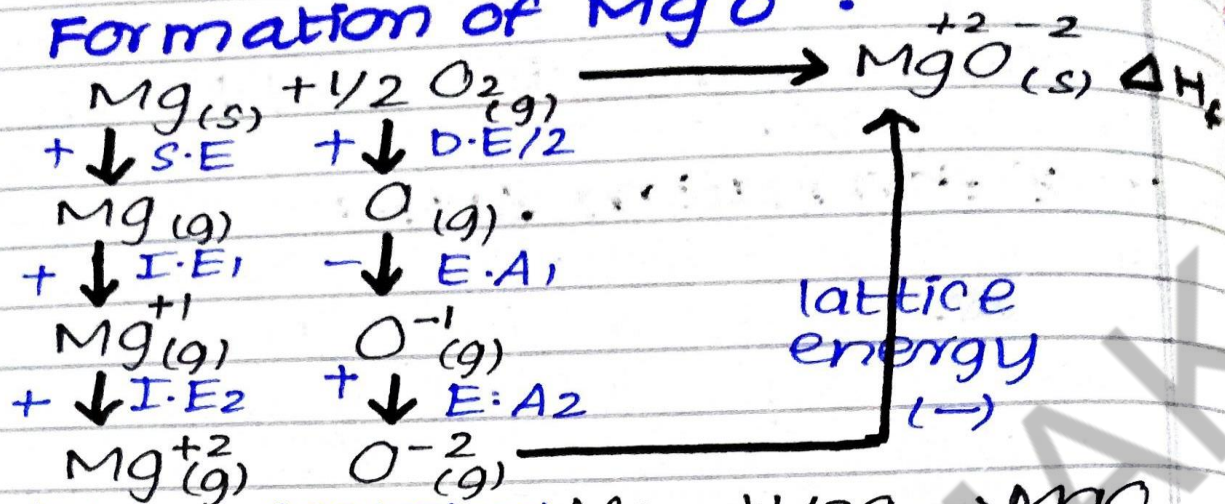
If Force of attraction increases
between the compound, lattice
energy also increases. $L.E \propto \text{charges}$

$$L.E \propto 1/r^2 \rightarrow \text{NaF} < \text{MgCl}_2$$

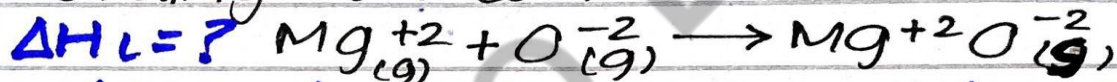
$$L.E \propto \text{charge: } \text{NaF} < \text{MgF}_2 < \text{AlF}_3$$

$$L.E \propto 1/\text{size: } \text{NaCl} > \text{KCl}, \text{NaF} > \text{NaCl} > \text{MgCl}_2$$

Formation of $Mg^{2+}O^{2-}$:



According to the Hess's law:



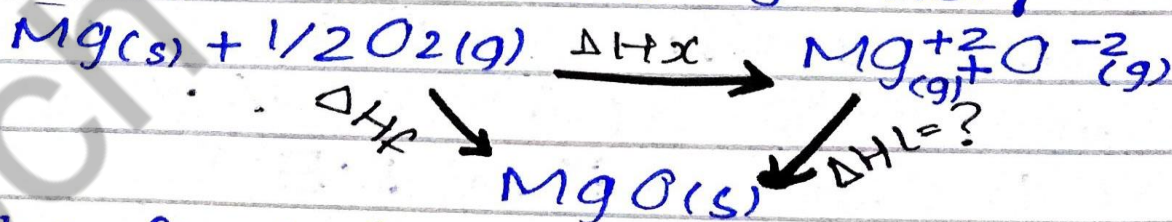
$$\Delta H_f = \Delta H_s + \Delta H_{AT} + \Delta H_{I.E} + \Delta H_{E.A_1} + \Delta H_{E.A_2} + \Delta H_L$$

Put values in it now:

$$\Delta H_f = 150 + 247 + 2180 - 141 + 878 + \Delta H_L$$

$$-602 = 3314 + \Delta H_L$$

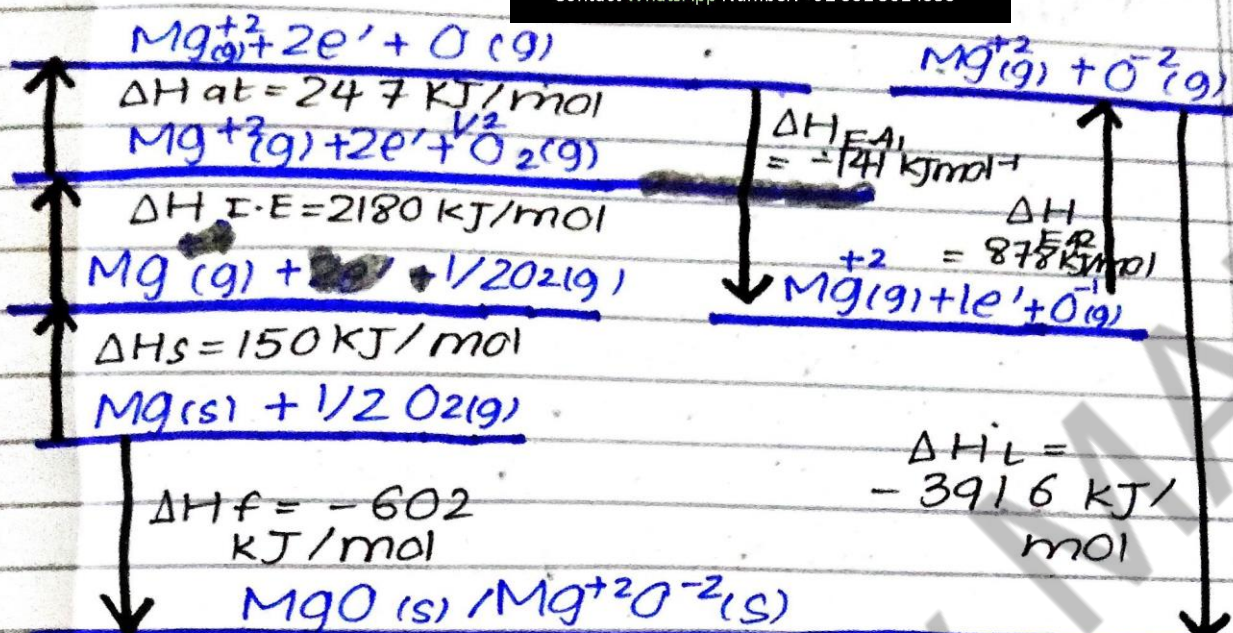
$$\boxed{\Delta H_L = -3916 \text{ kJ/mol}}$$



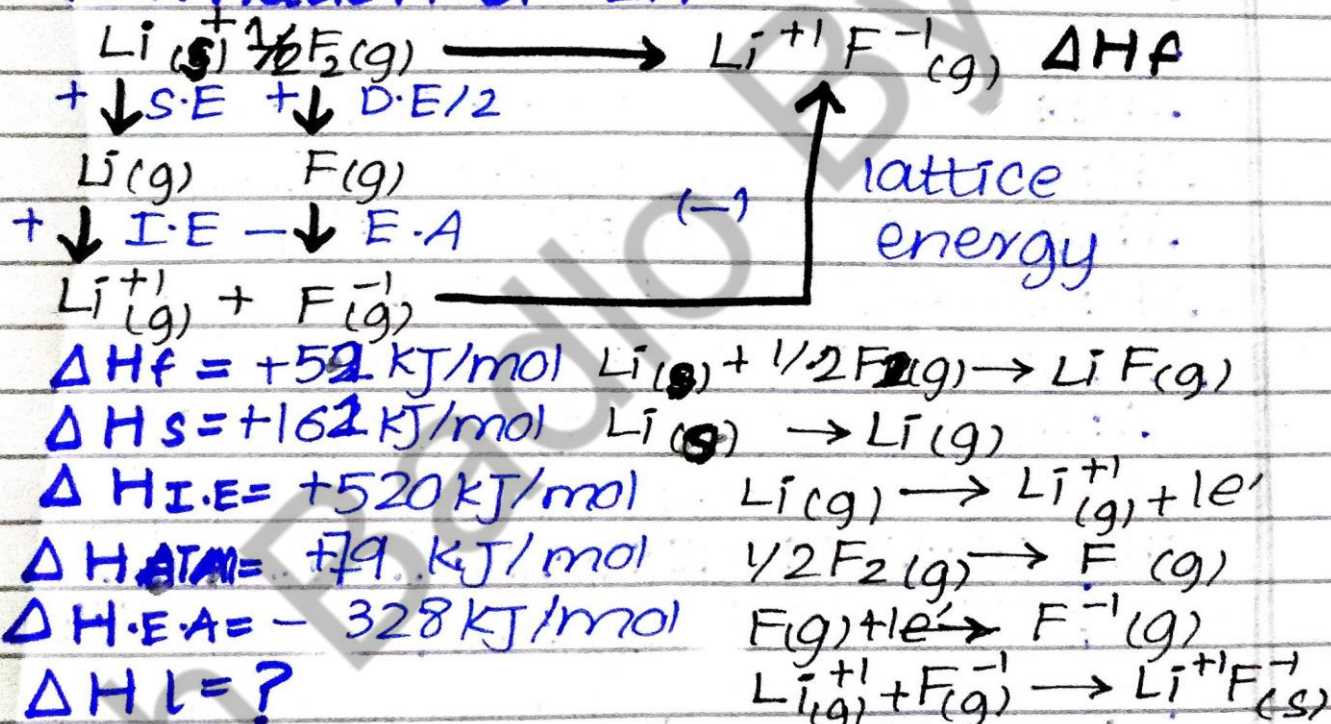
→ The first form/ state of atoms used in reaction must be in it's natural allotropic forms.

→ Firstly make gaseous state of atoms by sublimation / D.E.

→ Also check the charge on atoms.



Formation of LiF :

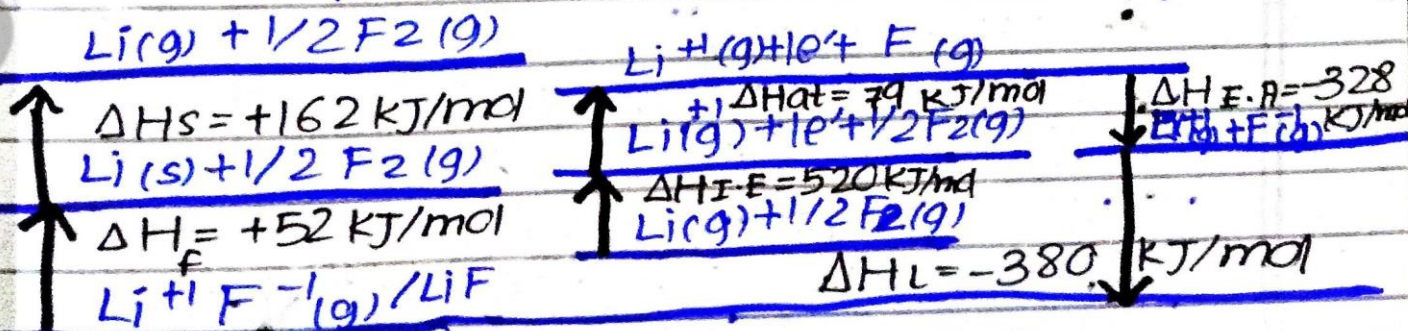


According to the haber's process.

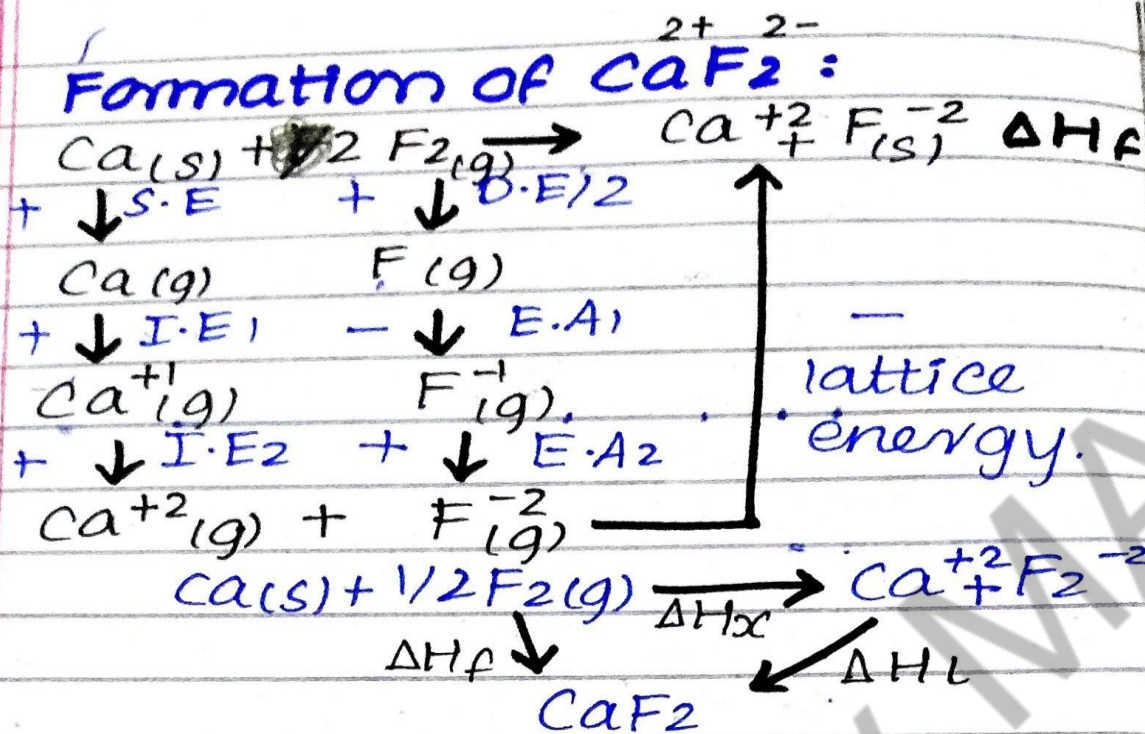
$$\Delta H_f = \Delta H_s + \Delta H_{I.E} + \Delta H_{at} + \Delta H_{E.A.} + \Delta H_L$$

$$+ 592 = 162 + 520 + 79 + 328 + \Delta H_L$$

$$|\Delta H_L = -380 \text{ kJ/mol}|$$



Formation of CaF_2 :



$$\Delta H_f = -1220 \text{ kJ/mol} \quad \text{Ca(s)} + 1/2 \text{F}_2(\text{g}) \rightarrow \text{CaF}_2(\text{s})$$

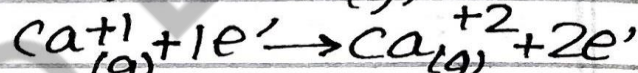
$$\Delta H_s = +178$$



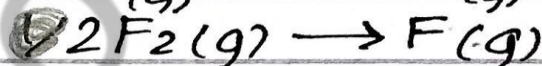
$$\Delta H_{\text{I.E}_1} = +590$$



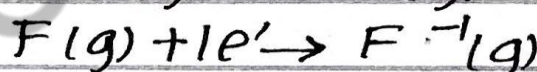
$$\Delta H_{\text{I.E}_2} = +1145$$



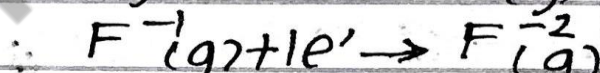
$$\Delta H_{\text{A.E}} = +79$$



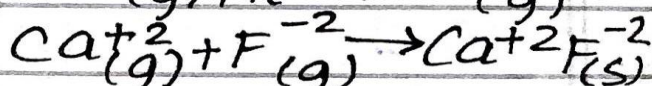
$$\Delta H_{\text{E.A}_1} = -328$$



$$\Delta H_{\text{E.A}_2} = +65.6$$



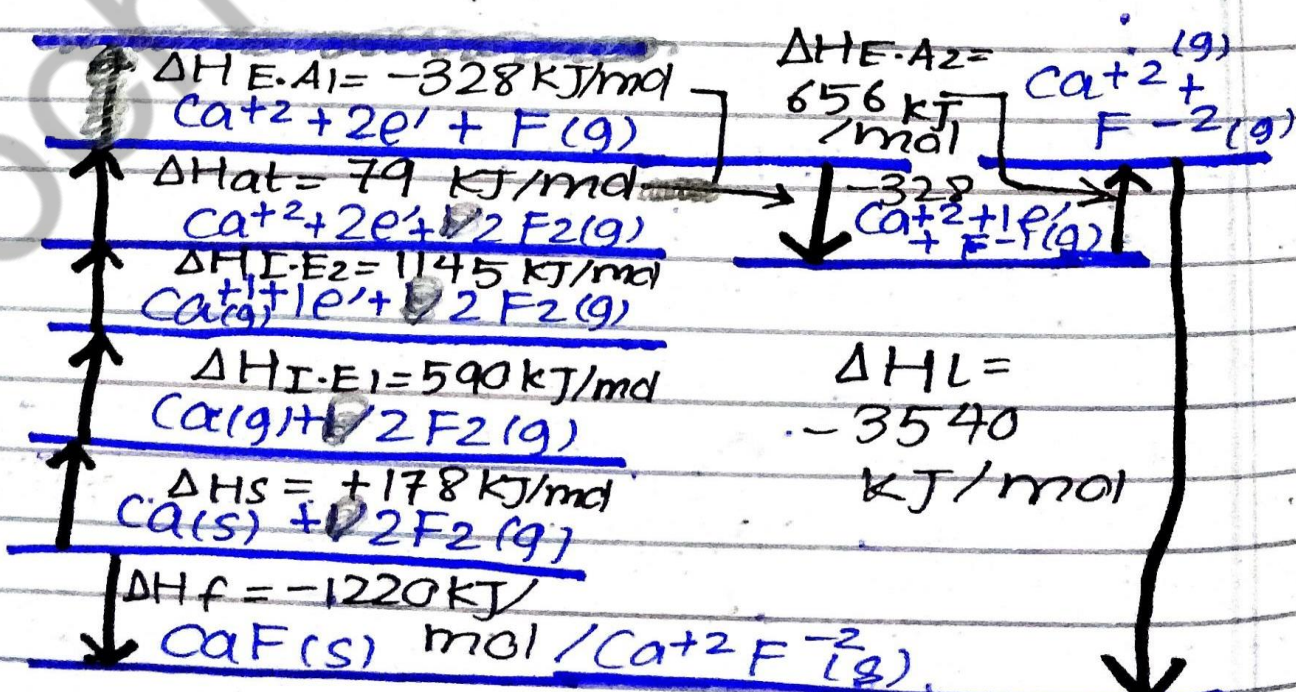
$$\Delta H_L = ?$$

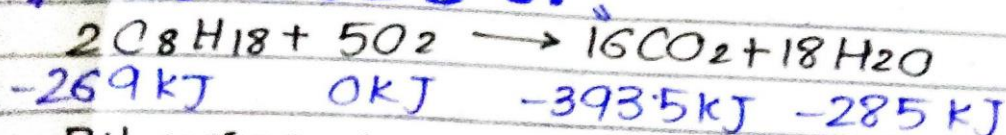


According to Born Haber's cycle:

$$\Delta H_f = \Delta H_s + \Delta H_{\text{I.E}_1} + \Delta H_{\text{I.E}_2} + \Delta H_{\text{A.E}} + \Delta H_{\text{E.A}_1} + \Delta H_{\text{E.A}_2} + \Delta H_L$$

$$\Delta H_L = -3540 \text{ kJ/mol}$$



EXAMPLE 6-8:-

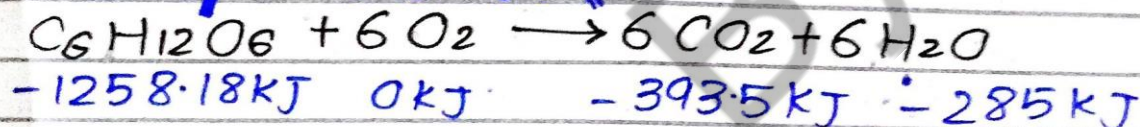
By using the formulae:

$$\Delta H^\circ_{\text{reaction}} = \sum \text{Coeff}_p \Delta H_f^\circ(\text{product}) - \sum \text{Coeff}_r \Delta H_f^\circ(\text{reactant})$$

$$\begin{aligned} \Delta H_r^\circ &= (16 \times CO_2 + 18 \times H_2O) - (2 \times C_8H_{18} + 5 \times O_2) \\ &= (16 \times -393.5 + 18 \times -285) - (2 \times -269 + 5 \times 0) \\ &= -11426 + 538 \end{aligned}$$

$$\boxed{\Delta H_r^\circ = -1.09 \times 10^4 \text{ kJ}}$$

Enthalpy of reaction is exothermic.

Example # 6.9:-

By using the formulae:

$$\begin{aligned} \Delta H_r^\circ &= \sum \text{Coeff}_p \Delta H_f^\circ(P) - \sum \text{Coeff}_r \Delta H_f^\circ(r) \\ &= (1 \times C_6H_{12}O_6 + 6 \times O_2) - (6 \times CO_2 + 6 \times H_2O) \\ \Delta H_r^\circ &= (1 \times -1258.18 \text{ kJ} + 6 \times 0) - (6 \times -393.5 + 6 \times -285) \end{aligned}$$

$$\begin{aligned} \Delta H_r^\circ &= (-1258.18) - (-2361 - 1710) \\ &= -1258.18 + 4071 \end{aligned}$$

$$\boxed{\Delta H_r^\circ = 2812.82 \text{ kJ}}$$

For 1g of glucose we can:

$$1 \text{ mol } C_6H_{12}O_6 = 2812.82 / 180$$

$$\boxed{1 C_6H_{12}O_6 = 15.62 \text{ kJ}}$$

FACTORS affecting lattice energy:-

* Na & Cl (Amount of energy required to break the ionic bond) lattice energy.

lattice energy $\rightarrow$ more $\rightarrow$ bond strong

lattice energy $\rightarrow$ less $\rightarrow$ bond weak

" Magnitude of (Amount) lattice energy is influenced by charge density of

an ion in compound:

$$\text{charge density} = \frac{\text{Ionic charge}}{\text{Ionic radius}}$$

(i) **Ionic charge** $\propto$ Lattice energy.

$\text{Mg}^{+2}\text{Cl}_2^{-2}$ has more lattice energy than $\text{Na}^{+1}\text{Cl}^{-1}$. $\text{MgO} > \text{MgCl}_2$
 $\text{Na}^{+1}\text{Cl}^{-1} < \text{Mg}^{+2}\text{Cl}_2^{-2} < \text{Al}^{+3}\text{Cl}_3^{-3}$

As they have more magnitude of charge in ionic molecule.

(ii) **Ionic radius** $\propto 1/\text{lattice energy}$.

KF has large size between LiF, so LiF has more lattice energy than KF. $\text{K}^{+1}\text{F}^{-1} < \text{Li}^{+1}\text{F}^{-1}$

$\text{K}^{+1}\text{Cl}^{-1} < \text{K}^{+1}\text{F}^{-1}$ (more lattice energy).

ROLES OF hydration in dissolving process:-

Hydration: When a compound is dissolved in water process is called hydration. $\text{Na}^{+}\text{Cl}^{-}(\text{aq})$

Dissolution: When a solid is dissolved in water at a given temperature, its process is called dissolution.

Solubility: Grams of solute dissolved in 1 litre / dm^3 of solvent. such as dissolving 58.5g of NaCl in $1\text{dm}^3 \text{H}_2\text{O}$.

→ Example of dissolution: dissolving NaCl in water at room temperature

solute-solute interaction **solvent-solvent interactions**

$\text{NaCl} - \text{NaCl}$

$\text{H}_2\text{O} - \text{H}_2\text{O}$

solute-solvent

interactions

$\text{NaCl} - \text{H}_2\text{O}$

* when a solution is formed: interactions between solute-solute particles are broken. (endothermic).

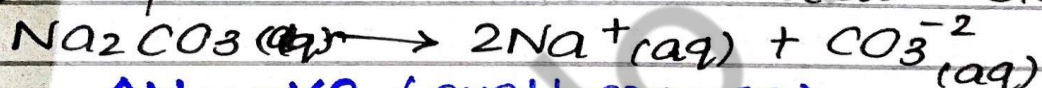
→ solvent molecules also move apart: and interactions are broken. (Endothermic)

→ Attractions between solute-solvent will occur. (hydration).

* This solute-solvent interactions or the solution may be endothermic / exothermic.

Heat of solution: when a compound is dissolved in water, the amount of heat released or absorbed is called as the heat of solution.

(i) compound + water → heat released



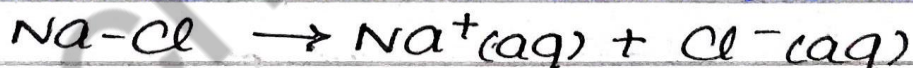
$\Delta H = -ve$ (exothermic).

(ii) compound + water → heat absorbed.



$\Delta H = +ve$ (endothermic).

* when a solute dissolves in H_2O the attractive forces of solute particles must be overcome to break the solute particles out of solid lattice (ion).



strong attractions → difficult to dissolve.

weak attractions → easily dissolves

→ strong attractions: more hydration energy

→ weak attractions: less hydration energy

strong imf = High boiling point = harder to break in water (ionize) = can be boiled to dissolve in water.

Lattice energy: $\Delta H_{\text{lattice}} / \Delta H_L$
Energy required to break the solute particles. $\text{Na} \nrightarrow \text{Cl}$

Enthalpy of hydration: ΔH_{hyd}
Energy [required] released when solute binds with water molecules.

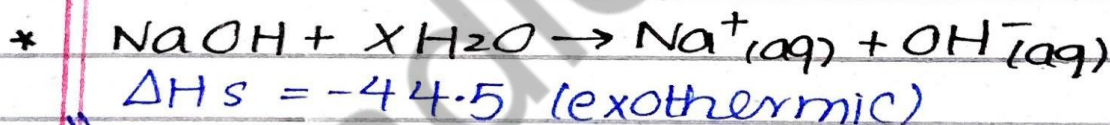
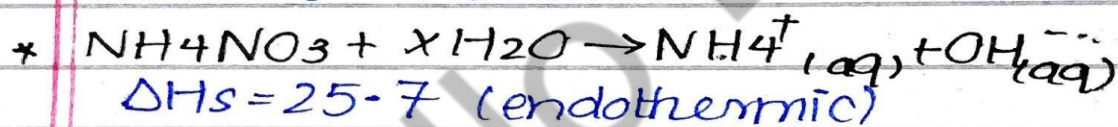
$$\Delta H_{\text{solution}} = \Delta H_L - \Delta H_{\text{hyd}}$$

~~ex~~othermic $\leftarrow$ $\rightarrow$ exothermic

* If ΔH_s is +ve (endothermic), then solution is -ve (exothermic).

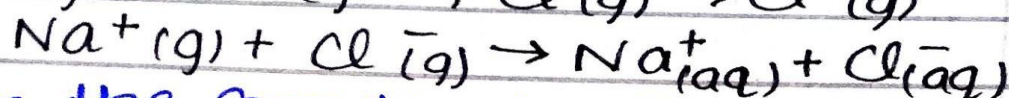
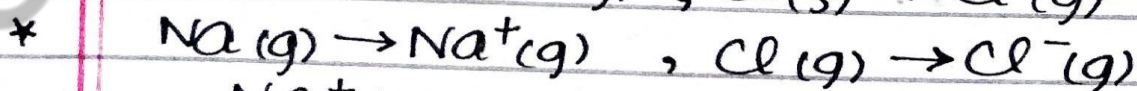
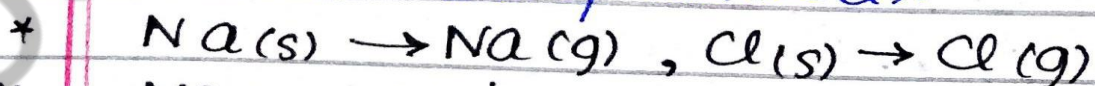
→ ΔH_s is very large: the difficult to dissolve the solute.

→ ΔH_s is very small: it's easy to dissolve the solute.



"Energy cycle involving enthalpy of solution:-"

After making the ionic gaseous state of the metal and non-metal atoms, now make them (ionic gases) into the aqueous states (1 step added).

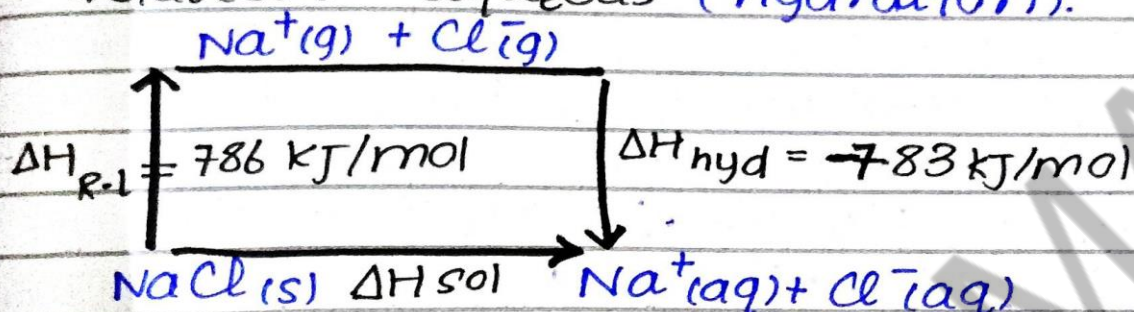


→ So, the formulae $\text{NaCl(s)} \rightarrow \text{Na}^+(\text{aq}) + \text{Cl}^-(\text{aq})$ is that:

$$\Delta H_s = \Delta H_{R.L} + \Delta H_{\text{hyd}}$$

* Energy formed = lattice energy.

- Energy breaking = reverse lattice energy.
- solid → gaseous state (Reverse lattice)
- Gaseous → aqueous (hydration).



* $\Delta H_{\text{lattice}} \rightarrow \Delta H_{\text{reverse lattice}}$.

$$\Delta H_{\text{L}} = -786, \Delta H_{\text{r.l}} = 786 \text{ kJ/mol.}$$

$$\Delta H_{\text{hyd}} = -783 \text{ kJ/mol}$$

According to this formulae:

$$\Delta H_{\text{sol}} = \Delta H_{\text{r.l}} + \Delta H_{\text{hyd}}$$

$$= 786 + (-783)$$

$$\boxed{\Delta H_{\text{sol}} = +3 \text{ kJ/mol}}$$

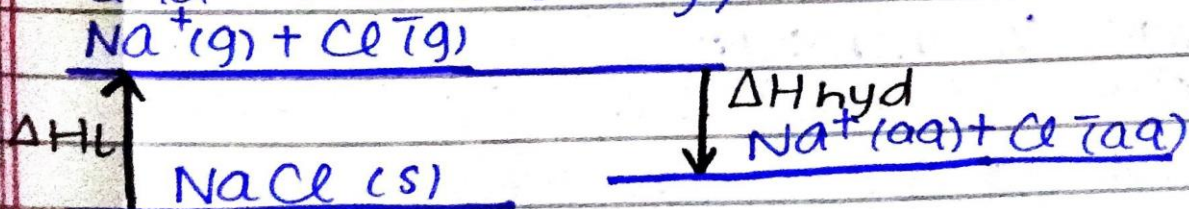
So, this process is endothermic process.

lattice energy: Energy released when 1 mole of ionic compound is formed from gaseous ion. $\text{Na}^+(\text{g}) + \text{Cl}^-(\text{g}) \rightarrow \text{NaCl(s)}$

Reverse lattice energy: Energy absorbed when 1 mol of ionic compound breaks down into gaseous ion. $\text{NaCl} \rightarrow \text{Na}^+(\text{g}) + \text{Cl}^-(\text{g})$

$\Delta H_{\text{solution}}$: Energy released or absorbed when 1 mol compound is formed.

- If ΔH_{sol} is +ve, the solution process is endothermic (requires energy).
- If ΔH_{sol} is -ve, it's exothermic process and releases energy.



"Enthalpy affecting ΔH_{hyd} :-"

Following are its types:

- (i) **Ionic charge**: The greater the ionic charge on the compound, the greater the hydration energy. $I \cdot C \propto \Delta H$
such as $\text{Na}^+ \propto -1 < \text{Mg}^{+2} \propto -2$
- (ii) **Ionic radius**: The bigger the size of an atom, the lesser the hydration energy. $I \cdot R \propto 1/\Delta H$
such as $\text{Ca}^{+2} \text{SO}_4^{-2} < \text{Mg}^{+2} \text{SO}_4^{-2}$

"ENTROPY:-"

"The degree of disorder and the randomness of a system is called entropy."

ice (s)	$\xrightarrow{\text{melt}}$	liquid water (l)
fixed arrangement		random arrangement
less entropy		more entropy

Symbolic representation:

It is represented by symbol S .

Factors for entropy:

- * As a system moves from more ordered state to less ordered state, entropy of a system increases and vice versa.

solid $\rightarrow$ liquid $\rightarrow$ gases

$S < L < g$ (order decreasing).

Changes in entropy:

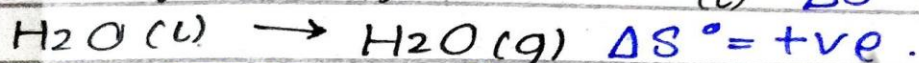
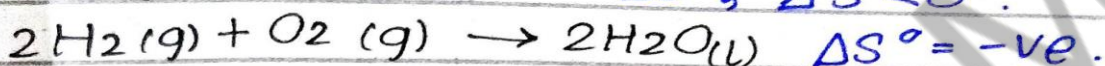
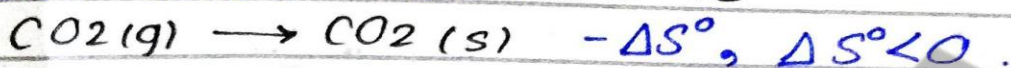
It's represented by ΔS .

- * If entropy of a system increases represented by $+\Delta S$ $\Delta S > 0$

→ If entropy of system decreases, it's represented by $+\Delta S$. $\Delta S < 0$

→ Temperature increases, then the entropy increases, then it's $+\Delta S$ $\Delta S > 0$

→ Temperature decreases = Entropy decreases. $-\Delta S$. $\Delta S < 0$



calculations of Entropy changes:

Following is it's formulae.

$$\Delta S^\circ = \sum S^\circ_{\text{product}} - \sum S^\circ_{\text{reactant}}$$

"Gibb's Free Energy:-"

Gibb's free energy is a measure of useable energy of a system at the constant pressure and temperature.

Also known as Available energy.

It's represented by symbol G .

$$G = H - TS \quad (H = \text{Enthalpy}, T = \text{temp}, S =$$

It's defined as: Entropy).

"Enthalpy of a system minus product of temperature and entropy."

Gibb's free energy is a state function it doesnot depend on path followed.

by the reaction.

Change in Gibb's free energy:

$$\Delta G^\circ = \Delta H^\circ - \Delta(TS)$$

At constant temperature, $\Delta T = 0$

So by putting this value in above.

$$\Delta G^\circ = \Delta H^\circ - \Delta S^\circ$$

It's the .Gibb's helmholtz equation.

* The sign and magnitude of ΔG° indicate process is spontaneous at equilibrium.

(i) $\Delta G^\circ < 0$ (-) spontaneous and in the forward direction.

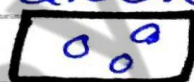
→ exothermic reaction.

(ii) $\Delta G^\circ > 0$ (+) non-spontaneous and forward direction.

→ Endothermic reaction.

(iii) $\Delta G^\circ = 0$ (same) The reaction is at equilibrium.

* **Entropy**: degree of disorderness.



(spontaneous)

less disorder

more disorder

small entropy, -S

Large S, +S

Entropy = +S (It's spontaneous).

* **Enthalpy**:  releases heat to hot water get stability.

$\Delta H = -ve$ (exothermic, spontaneous)

Enthalpy = $-\Delta H \rightarrow$ spontaneous

* **Temperature**:  Melting ice (melt fast)

spontaneous process

$\Delta T \rightarrow$ change spontaneous.

Gibb's Free energy:

$$\Delta G = \Delta H - T\Delta S$$

"It tells us whether the reaction is spontaneous or not".

So, enthalpy and entropy can't explain spontaneous processes

$\Delta G < 0$ spontaneous $\Delta G > 0$ non-spontaneous

$\Delta G = 0$ equilibrium

#	ΔH	ΔS	It's a
	-	+	spontaneous process.

$$\Delta G = \Delta H - T\Delta S$$

$$= (-) - (+)$$

$$= (-)(-)$$

$$\boxed{\Delta G = -ve} \text{ Exothermic process}$$

#	ΔH	ΔS	It's a
	+	-	non-spontaneous process

$$\Delta G = \Delta H - T\Delta S$$

$$= (+) - (-)$$

$$= +(+)$$

$$\boxed{\Delta G = +ve} \text{ Endothermic}$$

A process will be spontaneous at all temperatures if $\Delta H < 0$ and $\Delta S > 0$

#	ΔH	ΔS	It's a spontaneous
	+	+	and non spontaneous
	-	-	process.

$$\Delta G < 0$$

spontaneous

$$\Delta H - T\Delta S < 0$$

$$\Delta H < T\Delta S$$

$$\boxed{T > \frac{\Delta H}{\Delta S}}$$

This reaction is spontaneous above a specific temperature.

$$\Delta G > 0$$

non-spontaneous

$$\Delta H - T\Delta S > 0$$

$$\Delta H > T\Delta S$$

$$\boxed{T < \frac{\Delta H}{\Delta S}}$$

This reaction is non spontaneous below a specific temperature.

"Enthalpy and Calorie content:"

Enthalpy of food is the total heat energy stored in the chemical bonds of the food.

calorie content tells us how much of that energy of the body can use

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when it metabolizes the food.
* The calorie content of food is essentially a measure of its enthalpy change when metabolized. The enthalpy of combustion of food macronutrients directly corresponds to the calories they provide.

- > carbohydrates 4 kcal/g
- > proteins 4 kcal/g
- > fats 9 kcal/g

The higher the enthalpy of a food item, the more calories it provides body with.

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