

Ch - 5 -

Thermodynamics

System
Surrounding
Boundary

* Types of System:-

→ Open / Closed / Isolated

* Thermodynamic Properties:-

1) Extensive Properties:-

→ "The properties of system which depends on quantity of matter or depends on size or mass of the system is called ..."

eg:-	weight	Internal energy	Enthalpy
	Volume	potential energy	Entropy
	mole	kinetic energy	Gibb's Free energy
		Heat capacity	

2) Intensive Properties:-

→ The properties of system which does not depend on mass or size are called ...

Melting point
 Boiling point
 concentration
 density
 viscosity
 pH
 specific heat capacity
 molar heat capacity
 Refractive index (glass)

Characteristics of extensive & Intensive properties :-

1) A → Extensive
 B → Extensive

→ The ratio of extensive properties is always intensive properties.

$e.g. : \frac{\text{mass}}{\text{volume}} = \text{density}$	$\frac{\text{No. of moles}}{\text{Volume}} = \text{Molarity}$
---	---

2) When extensive properties are in the form of
 per unit mass
 per unit volume
 per unit mole } → Intensive properties

3) Extensive properties are additive &
 Intensive properties are not additive

Thermodynamic Process :-

- 1) Isothermal process :- $\Delta T = 0, \Delta U = 0$
- 2) Isobaric process :- $\Delta P = 0$
- 3) Isochoric process :- $\Delta V = 0$
- 4) Adiabatic process :- $\Delta q = 0$

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Cyclic process :-

$$\Delta U = 0$$

$$\Delta S = 0$$

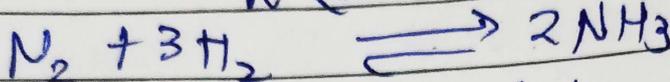
$$\Delta T = 0$$

$$q \neq 0$$

$$w \neq 0$$

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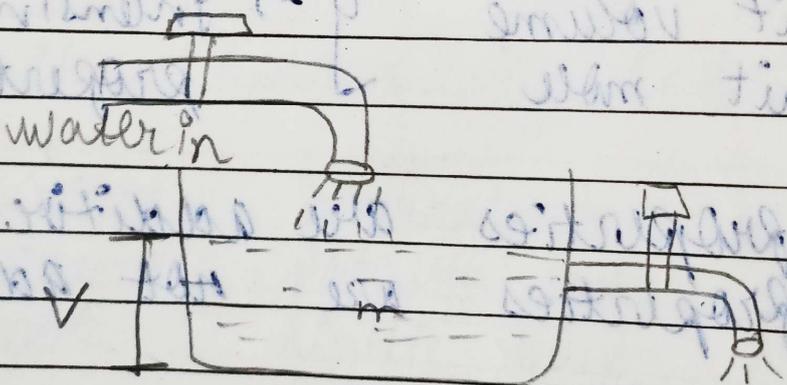
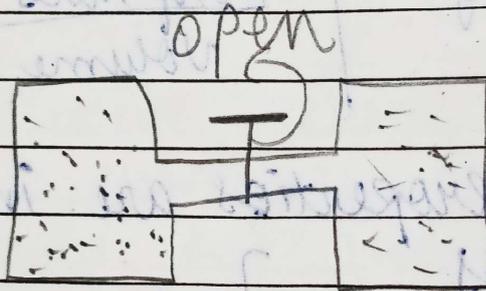
Reversible process



→ This reaction take infinite time to complete.

→ work done (w) is maximum

→ Reaction is in equilibrium (dynamic)



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Irreversible Reaction :-



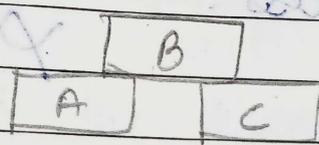
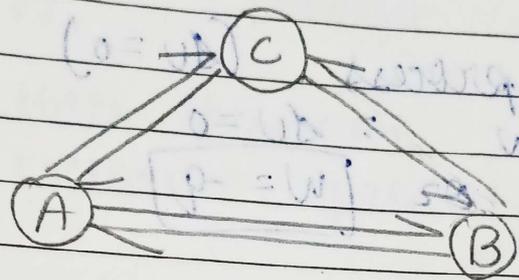
→ easily completed

→ work done - minimum

* Irreversible / Reaction

* Zeroth law of Thermodynamic

→ "When Body A is equilibrium with Body B & Body B is equilibrium with Body C, then Body A & C are also in equilibrium".



* 1st law of Thermodynamic :-

- "Energy can neither be created nor be destroyed.
- It is transferred from one phase to another phase.
- Total energy of universe remains constant."

→ Mathematically,

$$\Delta U = q + w \quad \text{or} \quad \Delta U = \Delta q + \Delta w \quad \text{or} \quad dU = dq + dw$$

→ Change in internal energy is the sum of heat given to the system & work done on the system.

Heat is given
to the system
(+q)

work done
on system
(+w)

System

work done by
the system
(-w)

Heat released by
the system
(-q)

Application of 1st law :-

1) Isothermal process ($\Delta U = 0$)

$$\Delta U = q + w = \Delta U = 0$$

$$q = -w \quad \text{or} \quad w = -q$$

2) Isobaric process
 $\Delta p = 0$

γ

3) Isochoric process

$$\Delta U = 0, w = 0$$

$$\Delta U = q + w$$

$$\Delta U = q$$

3) Adiabatic process

$$\Delta q = 0$$

$$\Delta U = q + w$$

$$\Delta U = w$$

4) Cyclic Process

$$\Delta U = 0$$

$$\Delta T = 0$$

$$\Delta S = 0$$

$$\Delta U = q + w$$

$$q = -w$$

or

$$w = -q$$

* Internal Energy:

→ Internal energy is total energy of system i.e. electric energy, chemical energy, mechanical energy, here :-

$$\text{Internal Energy} = \text{Kinetic Energy} + \text{Potential Energy}$$

- It is state function.
- It is extensive property.

• How can one change the internal energy of system?

- Heat given to the system or heat is given by
- Work done on the system / by the system
- mass in / mass out
- Absolute value of I.E. can't be determined.
- For isothermal process & cyclic process, $\Delta U = 0$

* Work Done: (IMP)

- Here work done is Mechanical work done = $F \times d$
- gravitational work = mgh } out of box
- electrical work done = $q \times V$ } box

Here $W = \text{Force} \times \text{Disp.}$
 $W = F \times d$ — (1)

1) Compression ($V_1 > V_2$)

$$W = F \times d \quad \text{--- (1)}$$

$$P = \frac{F}{a} \quad \text{--- (2)}$$

$$F = P \times a$$

$$W = P \times a \times d$$

here $a \times d = \text{volume}$

$$\Delta V = a \times d$$

$$\Delta V = V_2 - V_1$$

$$W = P \times a \times d$$

$$W = P \times \Delta V$$

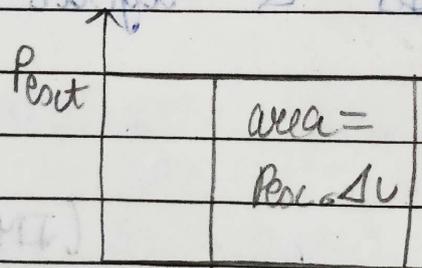
$$W = P_{\text{ext}} \times V_2 - V_1$$

$$W = -P_{\text{ext}} \times \Delta V$$

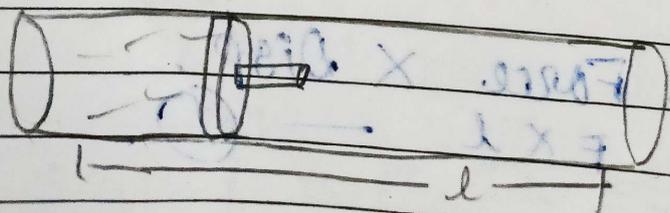
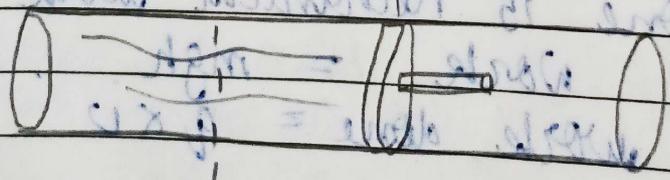
$$W = +P_{\text{ext}} \times \Delta V$$

work done on the system is always +ve

for compression



$$\Delta V = 0$$



Q1) Work done for expansion :-

$$W = F \times d$$

$$W = F \times a \times d.$$

$$P = \frac{F}{a}$$

$$F = P \cdot a$$

$$a \times d = \Delta V$$

$$\Delta V = V_2 - V_1$$

+ve

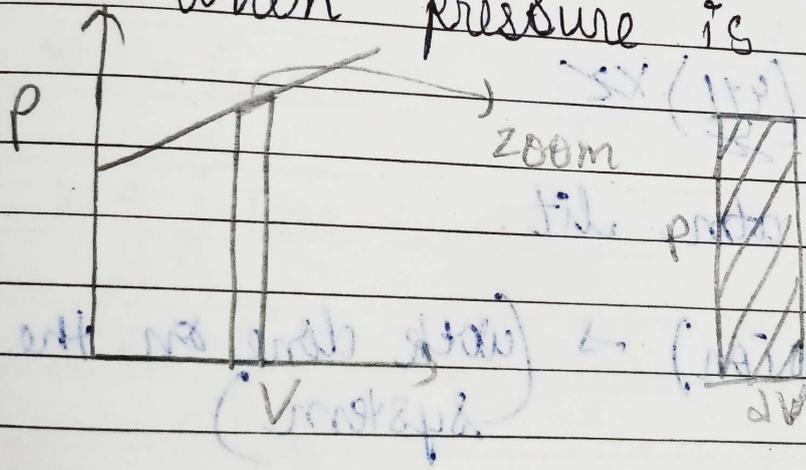
$$W = P_{ext} \times \Delta V$$

$$W = -P_{ext} \times \Delta V$$

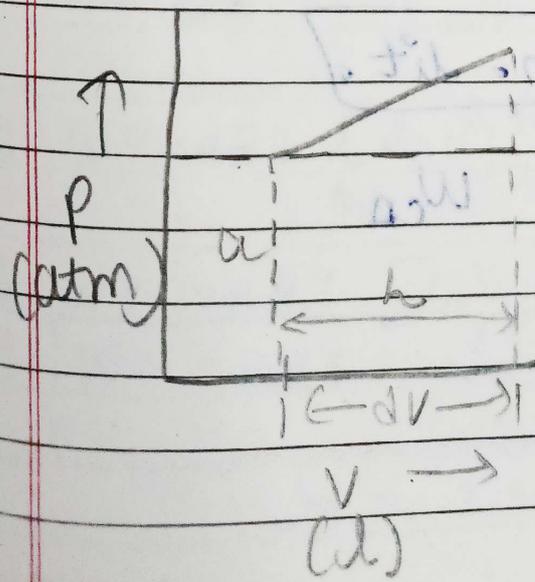
Q2) $-W = P_{ext} \times \Delta V$

→ Above process is considered at constant pressure.

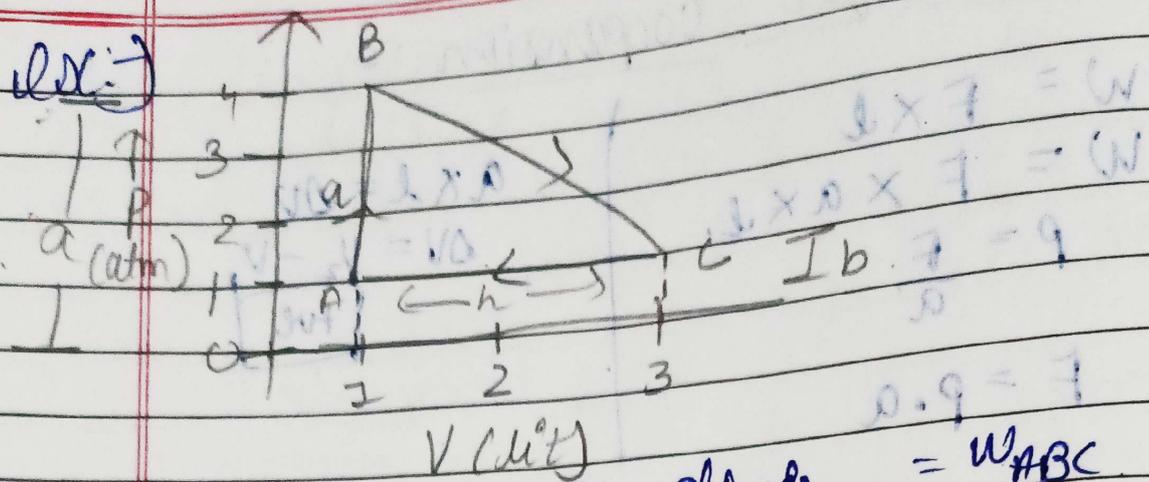
* When pressure is variable:-



$\Delta V =$ small change in volume
work done = Area of $P \times \Delta V$ diag.



Area of trapezium = $(a+b) \times h$.



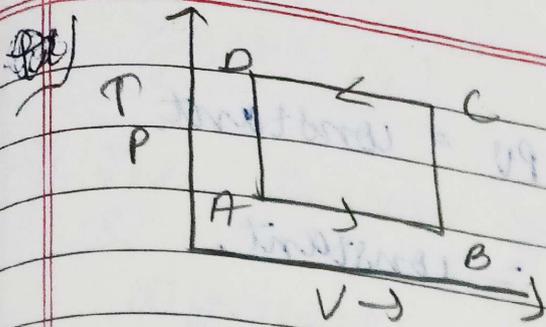
① $W_{AB} = P_{ext} \times \Delta V$ | volume \Rightarrow constant
 $= P_{ext} \times 0$
 $\boxed{W_{AB} = 0}$

② W_{BC} (work done by the system)
 $W = \text{Area of trapezium}$
 $= \frac{(a+b) \times h}{2}$
 $= \frac{(1+2) \times 3}{2}$

$W_{BC} = 1.5 = \boxed{5}$ atm. lit.

③ W_{CA} (compression) \rightarrow (work done on the system)
 $W_{CA} = 2 \times 1 = \boxed{2}$ atm. lit.

④ $W_{ABC} = W_{AB} + W_{BC} + W_{CA}$
 $= 0 - 5 + 2$
 $W_{ABC} = -3$ atm. lit.

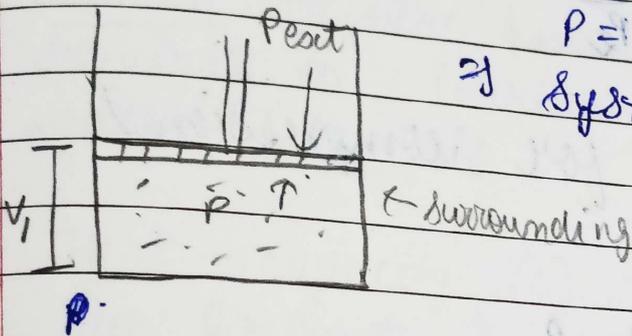


$W = +ve$

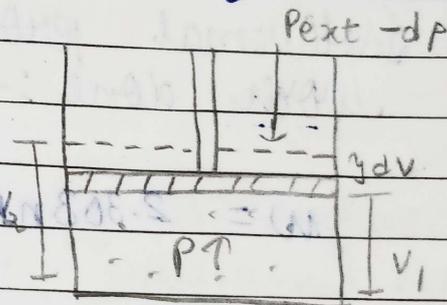
∴ anticlockwise = +ve

clockwise = -ve

* Work done for Isothermal process / Isothermal expansion.



$P = P_{ext}$
∴ System is in equilibrium



∴ System not in equilibrium

∴ $W = -P_{ext} \cdot dV$ — (1)

$W = \int_{V_1}^{V_2} -P_{ext} \cdot dV$

∴ $W = - \int_{V_1}^{V_2} nRT \cdot \frac{dV}{V}$

$W = -nRT \cdot \int_{V_1}^{V_2} \frac{1}{V} \cdot dV$

∴ $W = -nRT \left[\ln V \right]_{V_1}^{V_2}$

∴ $W = -nRT \left[\ln V_2 - \ln V_1 \right]$

∴ $W = -nRT \ln \frac{V_2}{V_1}$

$\ln = 2.303 \log$

∴ $W = -2.303 nRT \log \frac{V_2}{V_1}$

∴ work done by the system.

$PV = nRT$
ideal gas eqⁿ

P = Pressure

V = Volume

n = no. of moles

R = gas const.

= 8.314 J

T = temp. (K)

∴ $P_{ext} \cdot V = nRT$

$P_{ext} = \frac{nRT}{V}$

→ Boyle's law:- $P \propto \frac{1}{V}$ $\Rightarrow P = \frac{k}{V}$ $\Rightarrow PV = \text{constant}$

$\Rightarrow P_1 V_1 = \text{constant}$ | $P_2 V_2 = \text{constant}$,
 $\therefore P_1 V_1 = P_2 V_2$

$$\boxed{\frac{V_2}{V_1} = \frac{P_1}{P_2}}$$

$$\therefore W = -2.303 nRT \log \frac{P_1}{P_2}$$

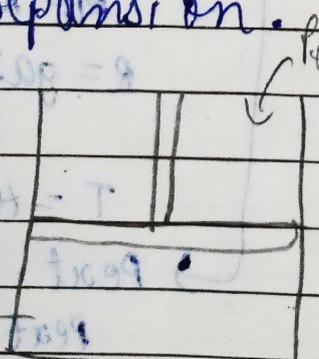
* Isothermal process for compression / work done :-

$$\Rightarrow W = 2.303 nRT \log \frac{P_1}{P_2}$$

$$W = 2.303 nRT \log \frac{V_2}{V_1}$$

* Free expansion :-

→ expansion in vacuum is called free expansion.



The expansion of gas against zero pressure

$$W = P_{\text{ext}} \times \Delta V$$

$$\boxed{W=0}$$

→ 1st law of thermo.

$$\Delta U = q + w$$

$$\Rightarrow W=0.$$

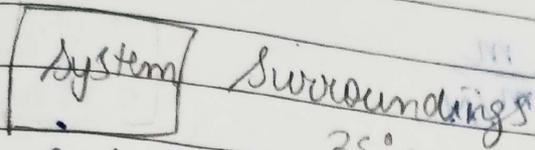
$$\therefore \boxed{\Delta U = q}$$

* Heat (q)

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→ The amount of heat is flow between system & surrounding on account of temperature difference.



Unit = calorie
 $1 \text{ cal} = 4.184 \text{ J}$

• Calorie :- The amount of heat energy is required to increase temperature of 1g of water by 1°C .

* Heat capacity (C) :-

→ The amount of heat energy is required to increase temperature of substance by 1°C .

$$C = \frac{q}{\Delta t}$$

→ It is extensive property.

* Specific Heat Capacity :-

→ The amount of heat energy is required to increase temperature of 1g substance by 1°C .

→ Intensive

$$c = \frac{q}{m \Delta t}$$

* Molar Heat capacity :- (C_M)

→ The amount of heat energy is required to increase temperature of 1mol substance by 1°C .

→ Intensive

$$C_M = \frac{q}{n \Delta t}$$

- Relation b/w specific heat capacity & molar heat capacity :-

$$\frac{C_M}{C} = \frac{q/m\Delta T}{q/m\Delta T}$$

$$\therefore \frac{C_M}{C} = \frac{m}{n} = \frac{m}{m/M}$$

\therefore

$$C_M = M \times C$$

$C_M = \text{molar mass} \times \text{specific heat capacity.}$

M-IMP

- * Relation b/w C_p & C_v or Prove $C_p - C_v = R$.

$\Rightarrow C_p = \text{Heat capacity at const. pressure.}$

$C_v = \text{Heat capacity at const. volume.}$

\rightarrow Heat capacity is given by.

$$C = \frac{q}{\Delta T}$$

$$\boxed{q = C \times \Delta T} \quad \text{--- (1)}$$

- Write above eqⁿ (1) for const. volume

$$\boxed{q_v = C_v \times \Delta T.} \quad \text{--- (2)}$$

- Write eqⁿ (1) for const. pressure

$$\boxed{q_p = C_p \times \Delta T} \quad \text{--- (3)}$$

\rightarrow For constant volume, $\Delta U = q + w$ But $w=0$

$$\therefore \boxed{\Delta U = q_v}$$

Put in (2)

ideal gas \rightarrow no attraction b/w
 $\Delta V = 0$ when mixed
 \rightarrow no change in energy when mixed

• Enthalpy :- Heat at const. pressure is denoted as

Enthalpy (ΔH)
 $\Delta H = q_p$ put in (3)

$\Delta U = C_v \times \Delta T$ ($\because q_v = \Delta U$)
 $C_v = \frac{\Delta U}{\Delta T}$

$\Delta H = C_p \times \Delta T$ ($\because \Delta H = q_p$)
 $C_p = \frac{\Delta H}{\Delta T}$

• Enthalpy is given by $H = U + PV$

$\Delta H = \Delta U + \Delta PV$

$dH = dU + d(PV)$

• $PV = nRT$ (\in ideal gas)

$\Delta H = \Delta U + \Delta(PV)$

$\Delta H = \Delta U + \Delta(nRT)$

($n=1$) ($R = \text{const.}$)

For 1 mol ideal gas value of $n=1$

$\Delta H = \Delta U + R\Delta T$

Put the value of ΔH & ΔU

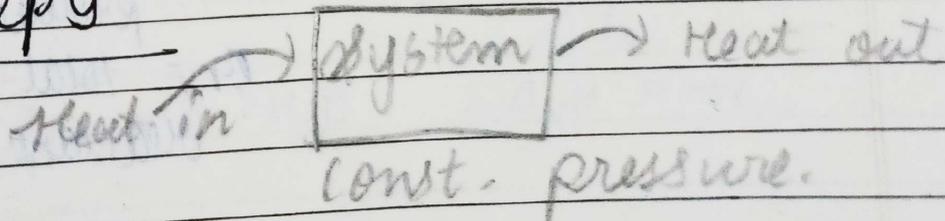
$C_p \times \Delta T = C_v \Delta T + R\Delta T$

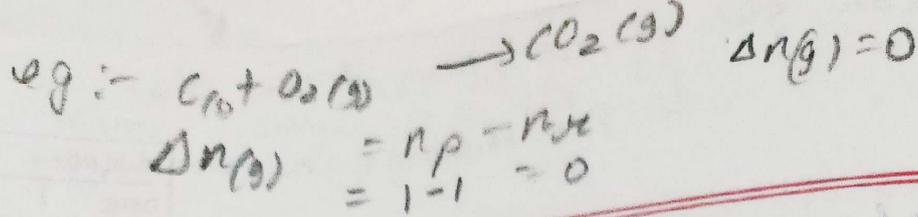
$C_p \Delta T - C_v \Delta T = R\Delta T$

$C_p - C_v = R$

gas const.

* Enthalpy





→ Heat is added to the system or removed from the system at constant pressure is called enthalpy (ΔH).

→ Heat at constant pressure is called enthalpy (ΔH).

→ The absolute value of enthalpy cannot be determined.

$$\Delta H = \Delta U + \Delta PV$$

Case: 1

at constant pressure

$$\Delta H = \Delta U + p \Delta V$$

$$\Delta H = q_p$$

$$q_p = \Delta U + p \Delta V$$

Case: 2

at constant volume

$$\Delta H = \Delta U + V \Delta P$$

$$q_p = q_v + V \Delta P$$

$$q_p - q_v = V \Delta P$$

$$\Delta H = \Delta U + \Delta PV$$

For ideal gas $PV = nRT$

$$\Delta H = \Delta U + \Delta n(g) RT$$

$$\Delta n(g) = n_p - n_r$$

$\Delta n(g)$ = change of total moles in gaseous phase.

n_p = total no. of moles of product in gaseous phase

n_r = total no. of moles of reactant in gaseous phase

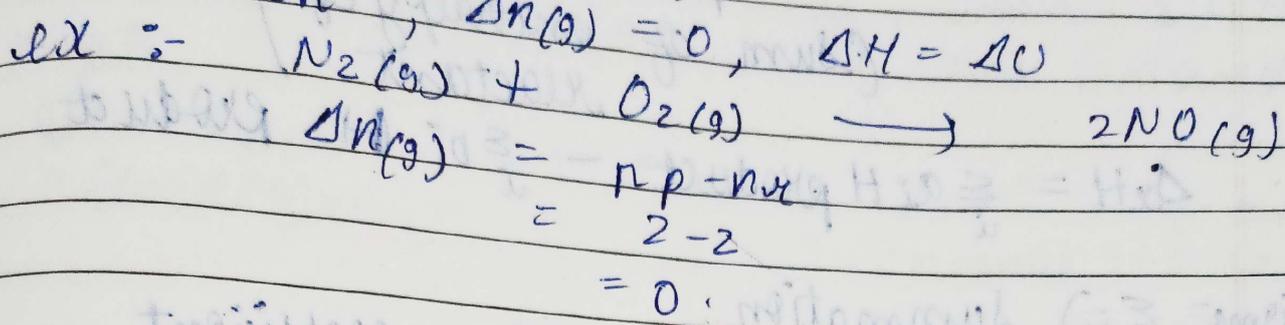
⇒ Case :- 1

$$\Delta H = \Delta U + \Delta PV$$

$$\Delta H = \Delta U + \Delta n(g)RT$$

When

$$\Delta n(g) = 0, \Delta H = \Delta U$$

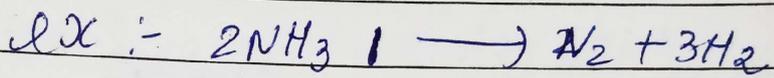


⇒ Case :- 2

$$\Delta H > 0$$

$$\Delta H = \Delta U + \Delta n(g)RT$$

$$\Delta n(g) > 1$$



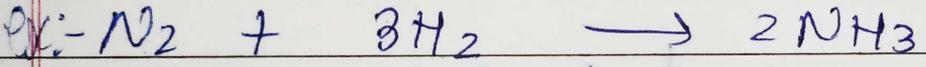
$$\Delta n(g) = n_p - n_r$$

$$= (1 + 3) - 2 = 2$$

⇒ Case :- 3

$$\Delta H < \Delta U$$

$$\Delta n(g) < 1$$

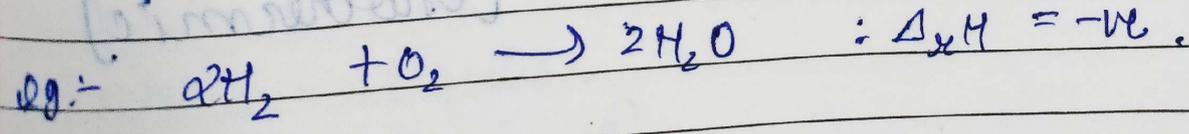
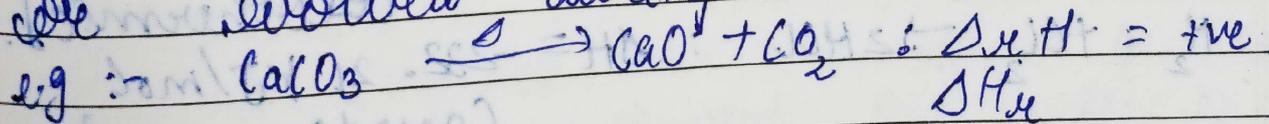


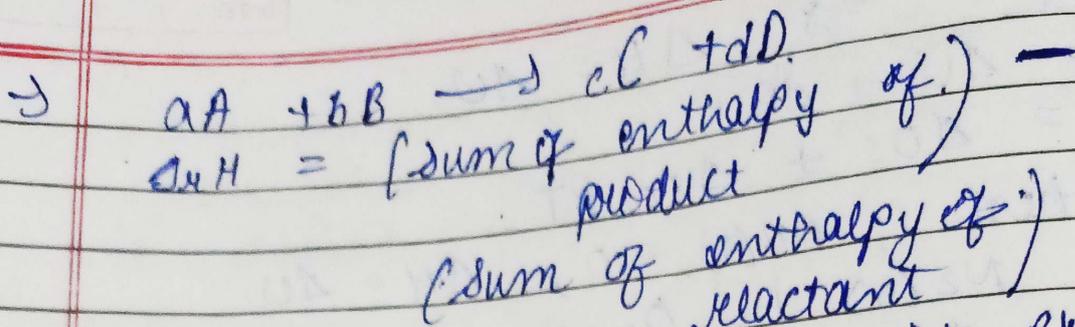
$$\Delta n(g) = n_p - n_r$$

$$= 2 - 4 = -2$$

* Heat of Reaction / Enthalpy of Reaction

→ It is defined as amount of heat absorbed or evolved during the chemical reaction.



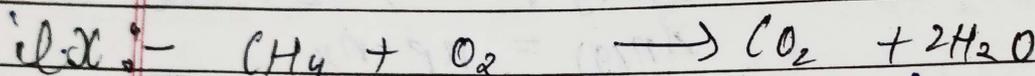


$$\Delta_r H = \sum_i a_i H_{\text{product}} - \sum_i b_i H_{\text{reactant}}$$

sigma = \sum = summation.

a_i & b_i = stoichiometric coefficient.

$$\rightarrow \Delta_r H = \left[\sum_i (c \Delta_f H_c + d \Delta_f H_d) \right] - \left[\sum_i (a \Delta_f H_a + b \Delta_f H_b) \right]$$



$$\Delta_r H = \sum_i (\Delta_f H_{\text{CO}_2} + 2 \Delta_f H_{\text{H}_2\text{O}}) -$$

$$\sum_i (\Delta_f H_{\text{CH}_4} + 2 \Delta_f H_{\text{O}_2})$$

Endothermic

$$\Delta H = +ve.$$

$$\Delta U = +ve.$$

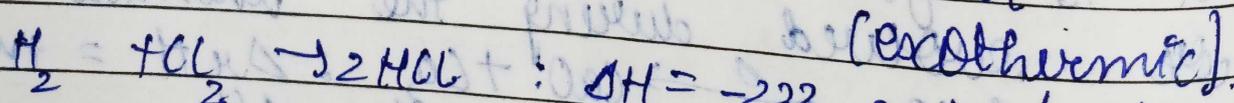
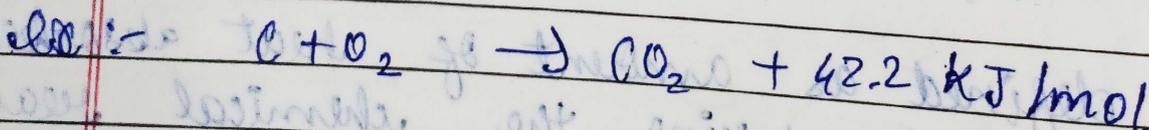
$$q = +ve$$

Exothermic

$$\Delta H = -ve.$$

$$\Delta U = -ve.$$

$$q = -ve.$$



(exothermic)

* Standard enthalpy of reaction ($\Delta_r H^\circ$) -

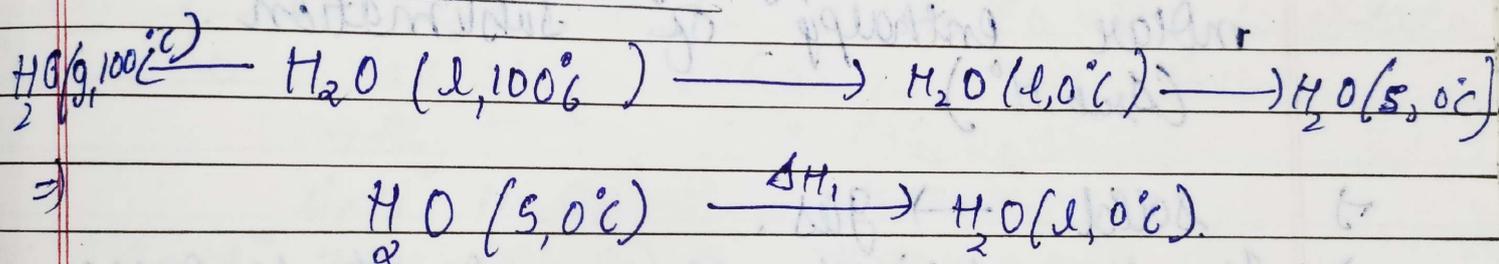
→ It is the enthalpy change when all the reacting species are in their standard state.

→ Standard Condition → STP → Tempⁿ → 25°C = 298K.
Pressure → 1atm
con^c → 1mol.

ex:- ethanol → 25°C → liquid.
iron rod → 500K → most stable state

$$\Delta_r H^\circ = \sum a_i \Delta_f H^\circ_{\text{product}} - \sum b_i \Delta_f H^\circ_{\text{reactant}}$$

* Enthalpy change during phase transformation :-



→ Temperature is constant at phase conversion (Volume nearly constant).

① Standard enthalpy of fusion or molar enthalpy of fusion ($\Delta_{\text{fus}} H^\circ$)

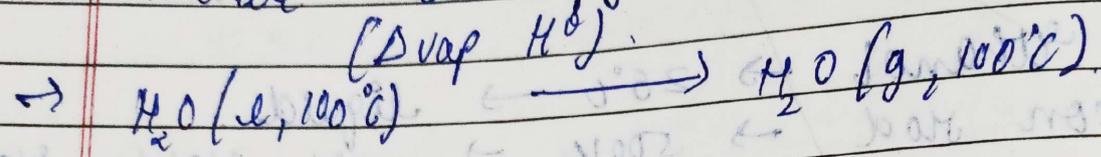
→ It is defined as amount of heat is required to convert one mole of solid into liquid.

Ques

It is the enthalpy change, when melting of 1 mol solid into liquid state.

→ It is endothermic process.
 $\Delta H = +ve.$

② Standard enthalpy of vaporization or molar enthalpy of vaporization.



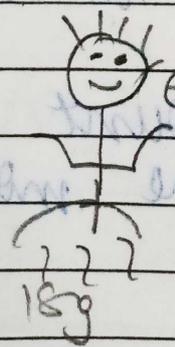
→ It is defined as amount of energy is required to convert 1 mol of liquid into gas.

③ Standard enthalpy of sublimation / molar enthalpy of sublimation.
 $(\Delta_{sub} H^\circ)$

→ solid \rightarrow gas.

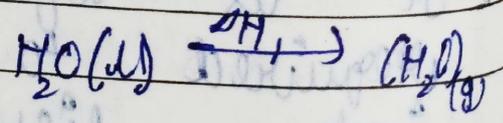
→ It is defined as amount of energy required to convert 1 mol of solid into gas.

Ex 5.7



298K
(25°C)

$\Delta H_{vap}^\circ = 44.01 \text{ kJ/mol}$



→ moles of water = $\frac{18g}{18}$

= 1 mol

→ Amount of energy required to vaporize the water = $n \times \Delta_{\text{vap}} H^\circ$
 $= 1 \times 44.01 \text{ kJ/mol}$

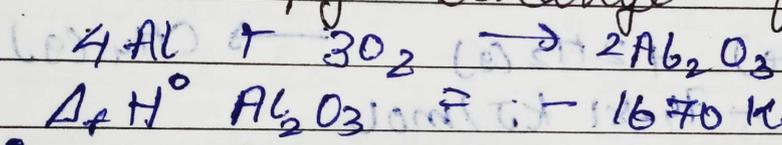
→ $\Delta H = \Delta U + \Delta PV$
 $\Delta U = \Delta H - \Delta PV$
 $\Delta U = \Delta H - \Delta n_{\text{gas}} RT$
 $\Delta n_{\text{gas}} = n_{\text{p}} - n_{\text{r}}$
 $= 1 - 0 = 1$

$\text{H}_2\text{O (l)} \rightarrow \text{H}_2\text{O (g)}$
 $\Delta U = 44.01 \text{ kJ/mol} - 1 \times 8.314 \text{ J/mol} \cdot \text{K} \times 298 \text{ K}$
 $= 44010 \text{ J/mol} - 1 \times 8.314 \text{ mol} \cdot \text{K} \times 298 \text{ J/mol}$

$\Delta U = 41.53 \text{ kJ/mol}$

Calc:-

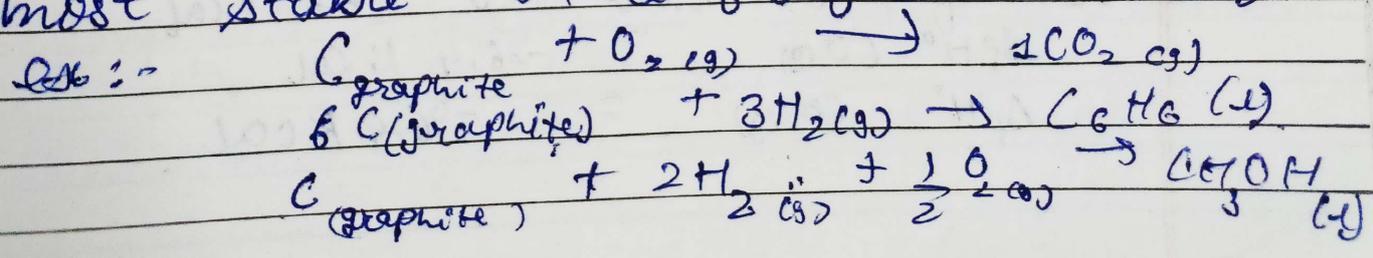
$\Delta_f H^\circ$ for Al_2O_3 is -1670 kJ . Calculate the enthalpy change for given

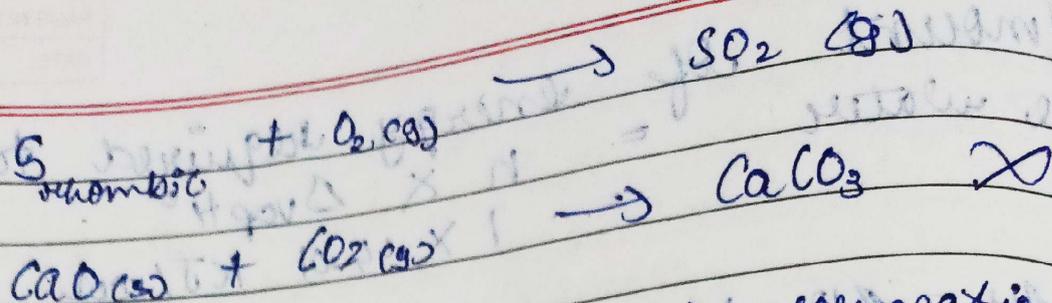


$\Delta_f H^\circ 2 \times (\text{Al}_2\text{O}_3) = -1670 \times 2 \text{ kJ}$

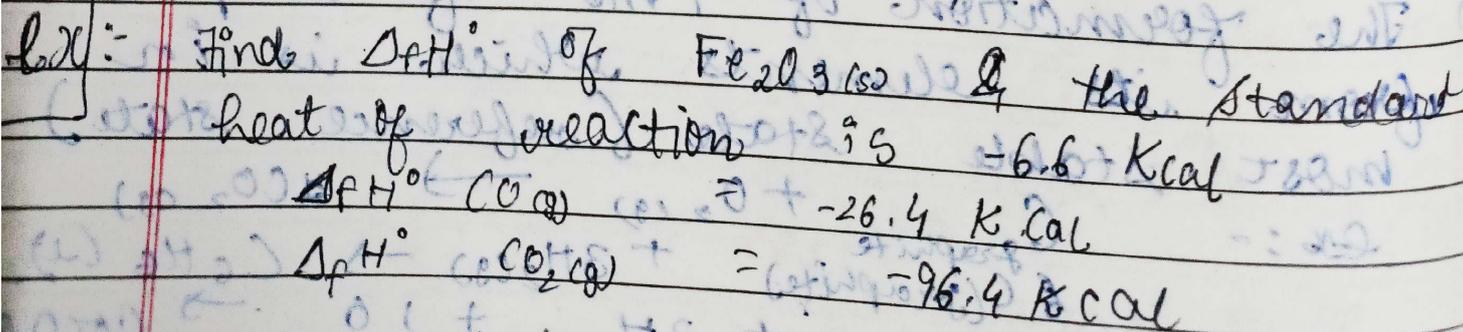
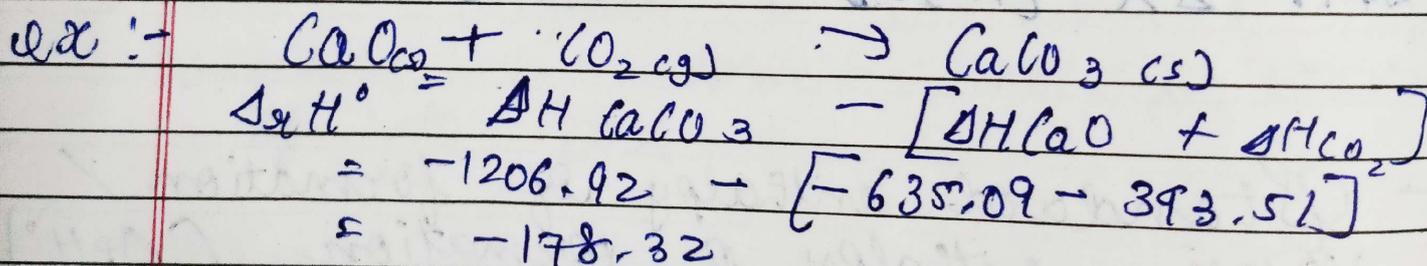
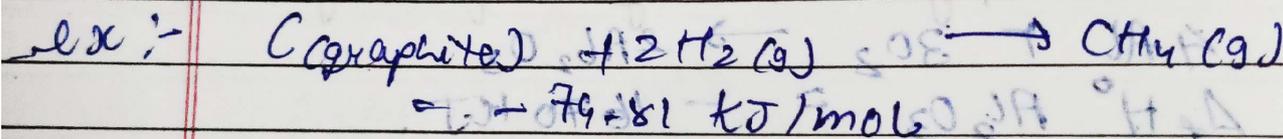
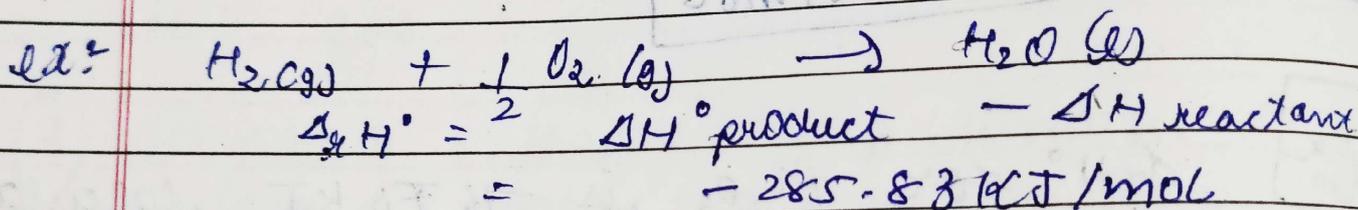
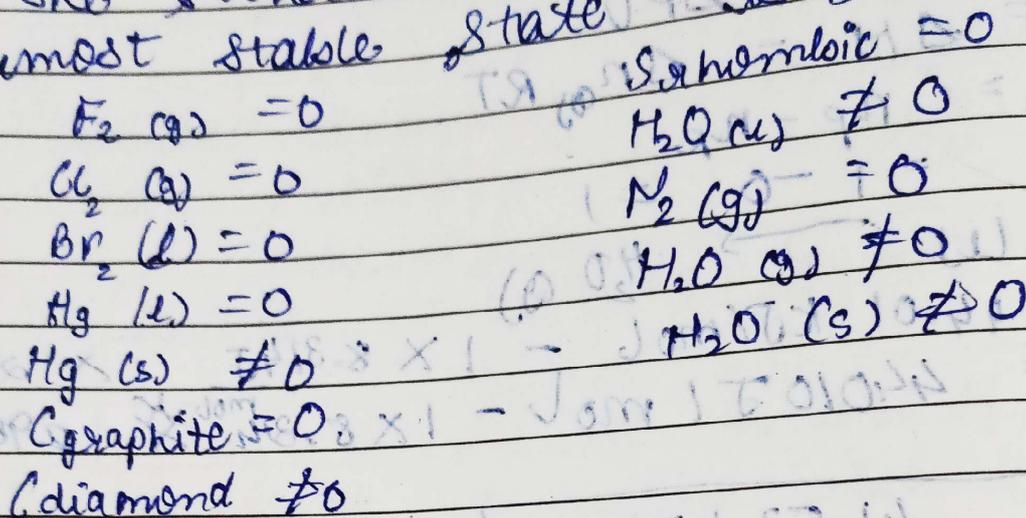
* Standard enthalpy of formation / molar enthalpy of formation ($\Delta_f H^\circ$)

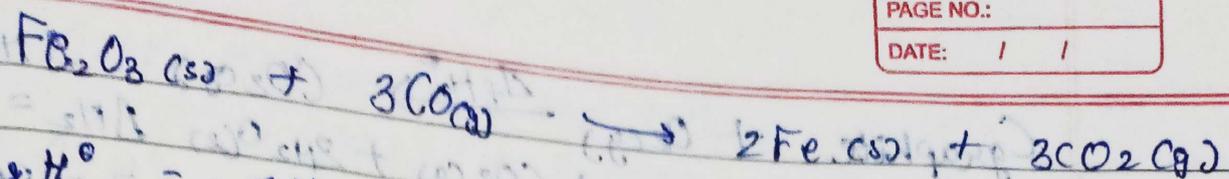
→ The formation of 1 mol of compound from its elements which is in most stable state (reference state)





The standard enthalpy of formation of most stable state is zero. (Element)





$$\Delta_r H^\circ = \Delta H_{\text{product}} - \Delta H_{\text{reactant}}$$

$$= [2\Delta_f H^\circ (\text{Fe} (\text{s})) + 3\Delta_f H^\circ (\text{CO}_2 (\text{g}))] - [\Delta_f H^\circ (\text{Fe}_2\text{O}_3 (\text{s})) + 3\Delta_f H^\circ (\text{CO} (\text{g}))]$$

$$-6.6 \text{ kcal} = [2(0) + 3(-94)] - [x + 3(-26.4)]$$

$$-6.6 = [0 - 282] - [x - 79.2]$$

$$-6.6 = -282 - x + 79.2$$

$$-6.6 + 202.8 = -x$$

$$196.2 \text{ kcal} = x$$

* Enthalpy of combustion / Standard enthalpy of combustion.

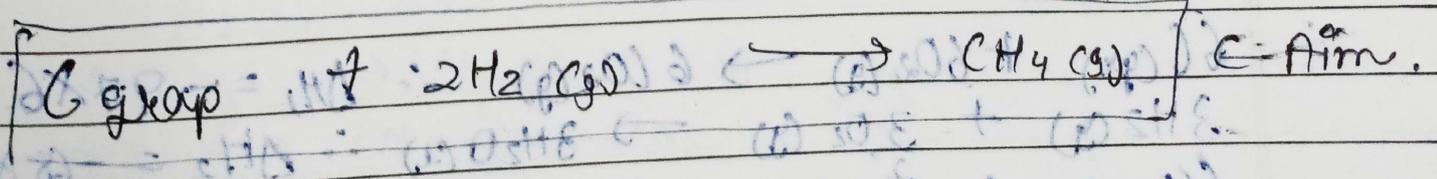
→ The amount of heat evolved when the combustion of 1 mol substance.

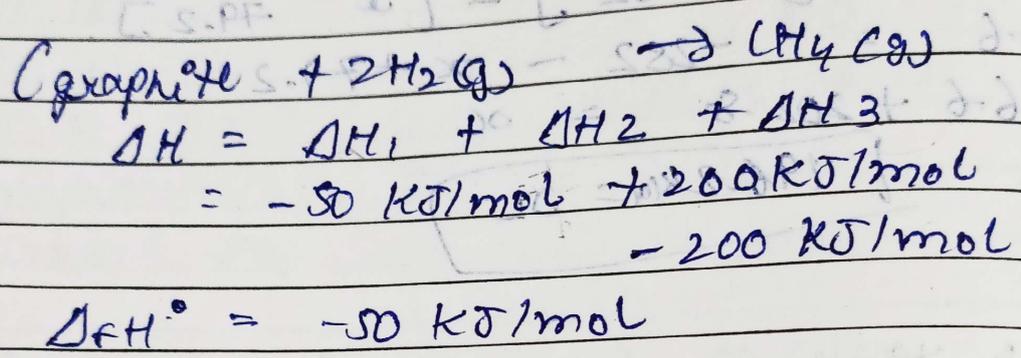
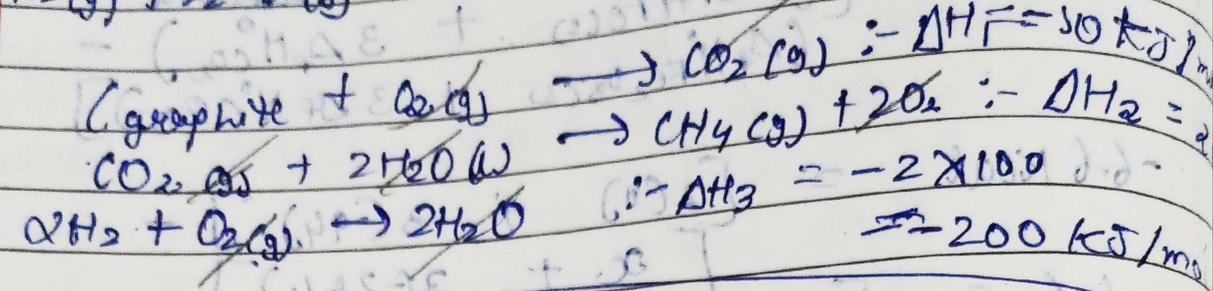
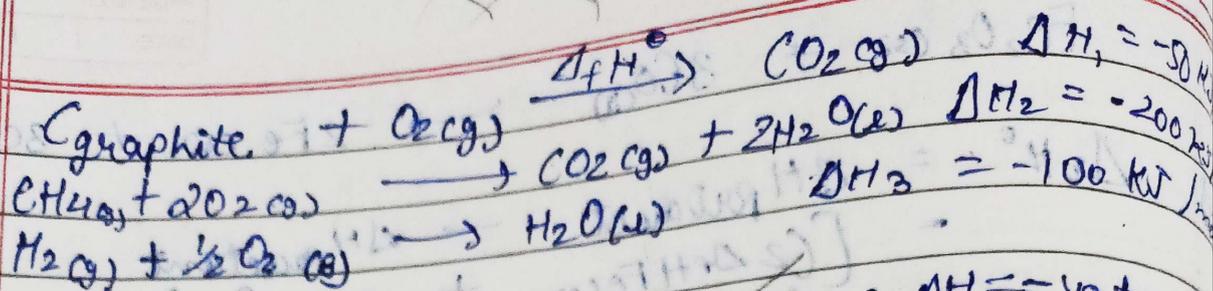
$$\text{C} + \text{O}_2 \rightarrow \text{CO}_2 \quad \therefore \Delta_c H^\circ = -ve$$

⇒ This is exothermic process.

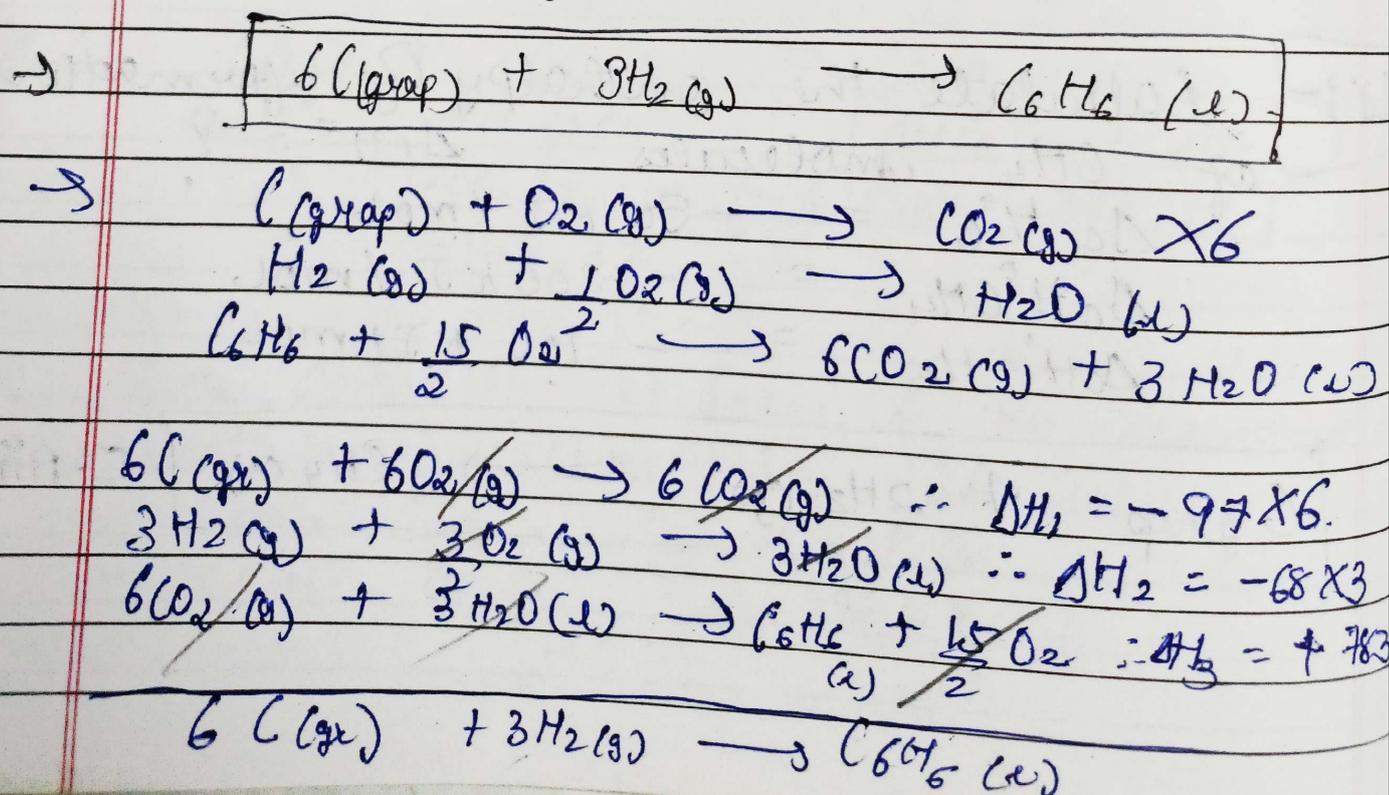
Ex: - Calculate the enthalpy of formation of CH_4 molecules $\Delta_f H^\circ = ?$

- $\Delta_c H^\circ (\text{C}) = -50 \text{ kJ/mol}$
- $\Delta_c H^\circ (\text{CH}_4) = -200 \text{ kJ/mol}$
- $\Delta_c H^\circ (\text{H}_2) = -100 \text{ kJ/mol}$





Q. The heat of formation of CO_2 & H_2O is -97 kcal & -68 kcal . The heat of combustion of benzene is -783 kcal . Calculate the heat of formation of benzene.

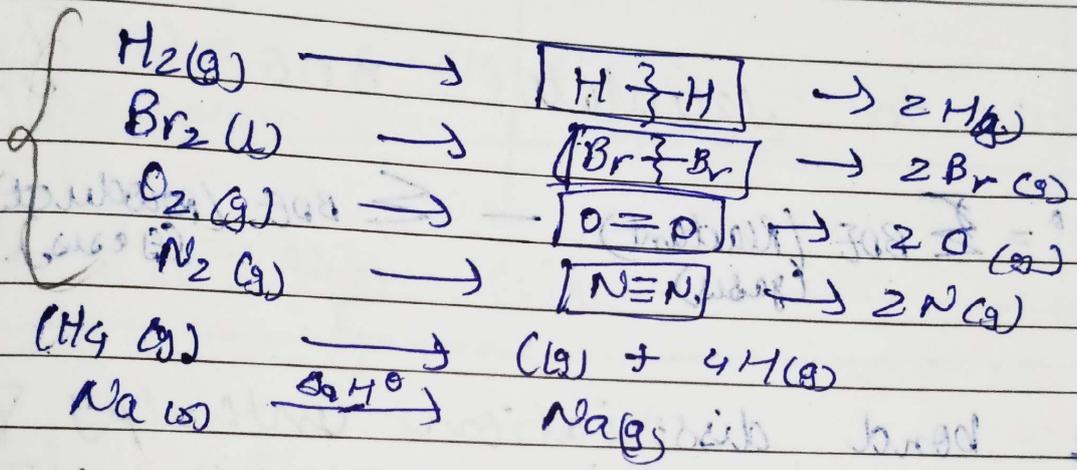


$$\Delta H = \Delta H_1 + \Delta H_2 + \Delta H_3$$

$$\Delta H^\circ = (-97 \times 6) + (-68 \times 3) + (783) \text{ kcal}$$

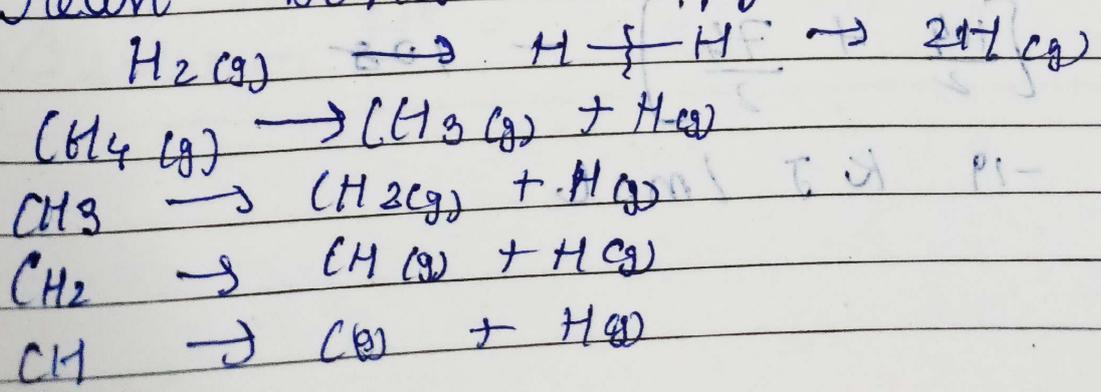
$$= -3 \text{ kcal}$$

* Standard enthalpy of atomization ($\Delta_a H^\circ$)
 → It is the amount of energy required to dissociate 1 mole of given substance into gaseous atom.



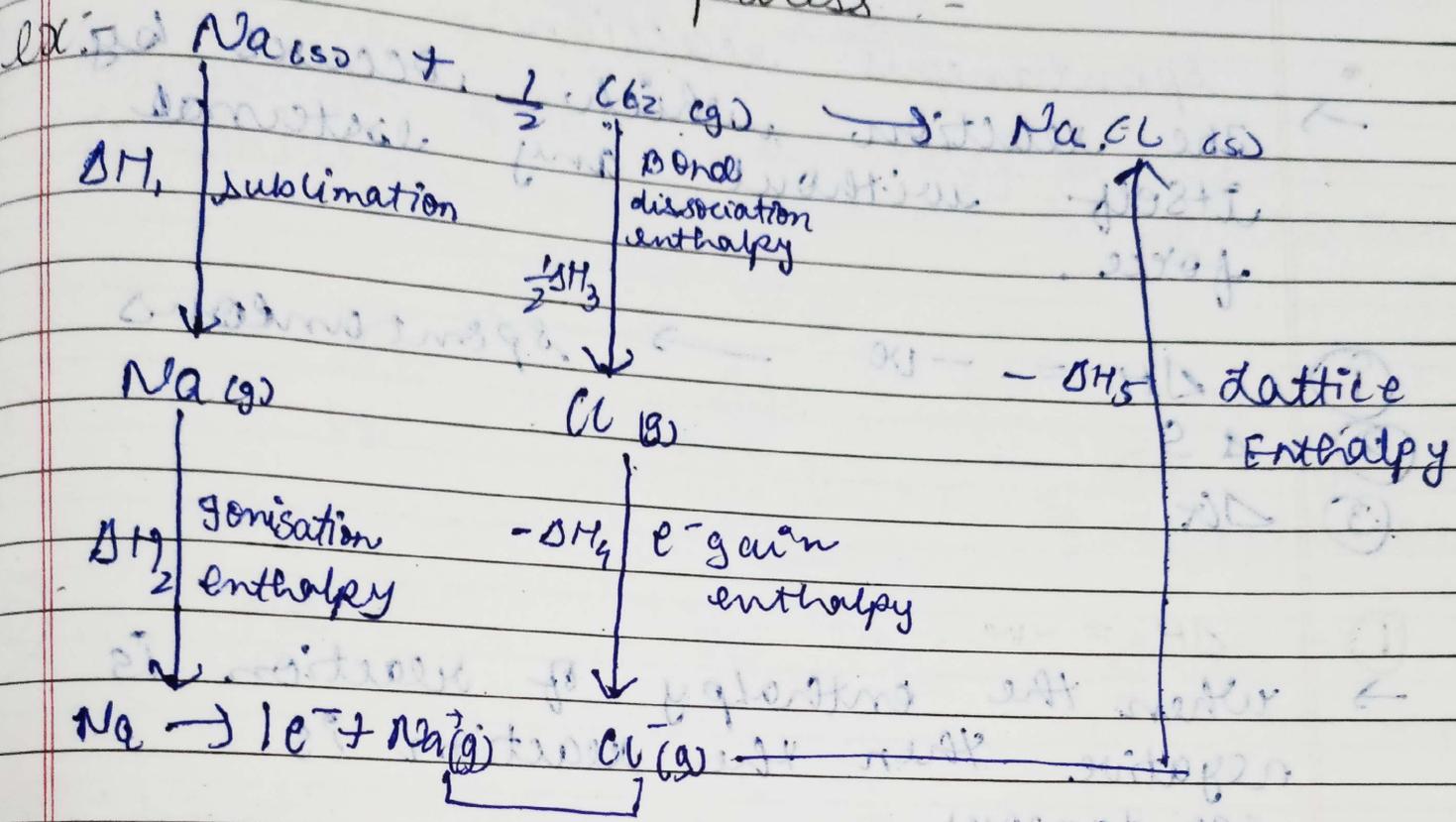
→ In the case of diatomic molecules it is also known as bond dissociation enthalpy.

* Bond dissociation enthalpy / energy
 Bond dissociation enthalpy
 Mean Bond enthalpy

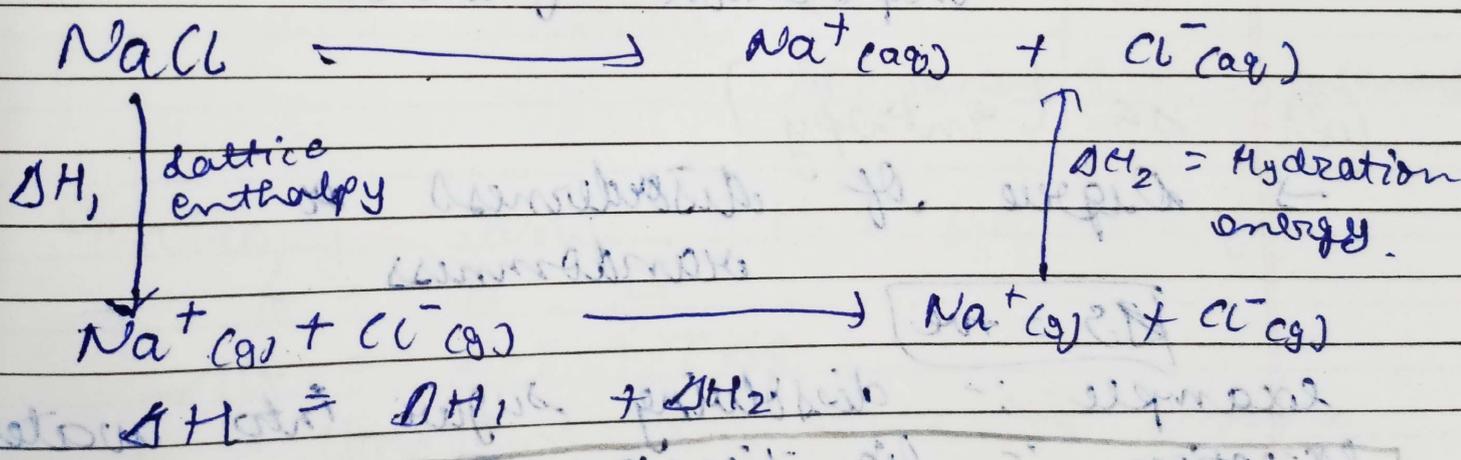


* Lattice enthalpy

→ Born Haber process :-



* $\Delta_{\text{sol}} H^\circ$



$$\Delta_{\text{solution}} H^\circ = \Delta_{\text{lattice}} H^\circ + \Delta_{\text{Hydration}} H^\circ$$

* Spontaneous Reaction or Spontaneity & Gibb's free energy

→ Spontaneous reaction:-
The reaction which occurs by itself without any external force.

- ① $\Delta H = -ve$ → Spontaneous
- ② $\Delta S = +ve$
- ③ $\Delta G = -ve$ always

① $\Delta H = -ve$
→ When the enthalpy of reaction is negative then the reaction is spontaneous.

② exception :-

- melting of ice.
- evaporation of water

Ice

② ΔS (Entropy)

→ Degree of disorderness or randomness

$\Delta S = +ve$

example :- dissolving sugar into water

exception :- liquification of gas

③ ΔG (Gibb's free energy)

$$G = H - TS$$

G = Gibb's free energy

H = Enthalpy

T = temperature

S = Entropy

→ at constant temperature & pressure
 $\Delta G = \Delta H - T\Delta S$

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

→ when ΔG is negative, then reaction is spontaneous (exothermic)

$$\Delta H$$

-ve

$$\Delta S$$

+ve

$$\Delta G$$

$$\Delta G = -ve$$

→ spontaneous

+ve

-ve

$$\Delta G = +ve$$

→ Non spontaneous

-ve

-ve

$$\Delta G < 0$$

$$\Delta G > 0$$

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

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

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

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

$$\Delta S$$

$$\Delta S$$

Spontaneous

Non-spontaneous

* Entropy and gibb's free energy

$$G = H - TS$$

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

→ constant temperature & pressure
 → Helmholtz's factor

$$\Delta S_{total} = \Delta S_{system} + \Delta S_{surrounding}$$

→ when system is in equilibrium

$$\Delta S = 0 \quad \Delta S > 0 \quad \therefore \text{spontaneous}$$

$$0 = \Delta S_{\text{system}} + \Delta S_{\text{surroundings}}$$

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

Reaction is in equilibrium

$$\Delta G = 0$$

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

$$T\Delta S = \Delta H$$

$$\Delta S = \frac{\Delta H}{T}$$

$$\Delta S_{\text{surroundings}} = \frac{\Delta H_{\text{surroundings}}}{T} = -\frac{\Delta H_{\text{sys}}}{T}$$

$$T\Delta S_{\text{sys}} = \Delta H_{\text{sys}}$$

$$\Delta S_{\text{total}} = \Delta S_{\text{system}} + \Delta S_{\text{surroundings}}$$

$$\Delta S_{\text{total}} = \Delta S_{\text{sys}} - \frac{\Delta H_{\text{sys}}}{T}$$

$$T\Delta S_{\text{sys}} = \Delta H_{\text{sys}} - \Delta H_{\text{sys}}$$

$$T\Delta S_{\text{sys}} = -\Delta H_{\text{sys}} + T\Delta S_{\text{system}}$$

$$= -(\Delta H_{\text{sys}} - T\Delta S_{\text{sys}})$$

$$= -\Delta G$$

$$= \boxed{-ve}$$

* Gibbs free energy & equilibrium constant.

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

$$\Delta G = 0, \text{ equilibrium}$$

$$\Delta G < 0, \text{ spontaneous}$$

$$\Delta G > 0, \text{ non-spontaneous}$$

→ STP: conditions

$$T = 298 \text{ K}$$

$$P = 1 \text{ atm}$$

$$C = 1 \text{ mol}$$

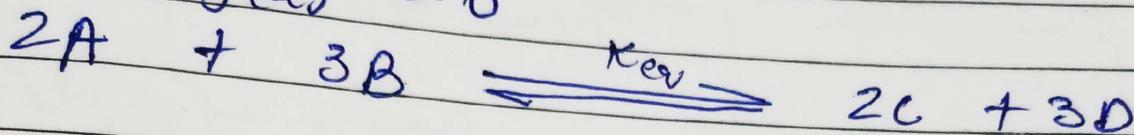
$$\Delta G_r^\circ = \sum a_i \Delta G^\circ_{\text{product}} - \sum b_i \Delta G^\circ_{\text{reactant}}$$

Standard stable glibb's free energy of most states of element is zero.

$$\text{Ca(s)} \neq 0$$

$$\text{Br}_2(\text{g}) \neq 0$$

$$\text{Hg(l)} = 0$$



$$K_{eq} = \frac{[C]^2 [D]^3}{[A]^2 [B]^3}$$

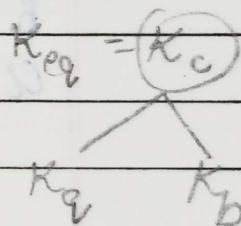
$\Delta G = \Delta G^\circ + RT \ln K_{eq}$ concⁿ change is zero.

$$\Delta G = 0$$

$$0 = \Delta G^\circ + RT \ln K_{eq}$$

$$\Delta G^\circ = -RT \ln K_c$$

$$\Delta G^\circ = -2.303 RT \log_{10} K_c$$



$$\ln = 2.303 \log_{10}$$