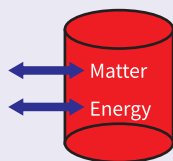


TYPES OF SYSTEM

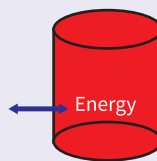
Open System



Isolated System



Closed System



THERMODYNAMICS PROCESSES

Isothermal Process

$$\Delta T = 0$$

Isochoric Process

$$\Delta V = 0$$

Adiabatic Process

$$q = 0$$

Cyclic Process

$$\Delta U_{\text{cyclic}} = 0$$

HEAT (q)

Exchange of energy due to temperature difference.

THERMODYNAMICS PROPERTIES

Intensive Properties

P, T, N

Properties of the system which are independent of amount of system.

State Function

Values of such functions do not depend on the path of the system.
Example: ΔU , ΔH , ΔS etc.

Extensive Properties

V, U, H

Properties of the system which depend on the amount of the system.

Path Function

Values of such functions depend on the path of the system.
Example: W and q.

INTERNAL ENERGY (U)

Total energy within the substance

WORK (W)

Work is a mode of energy transfer when temperature difference is not involved.

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

5. THERMODYNAMICS

FIRST LAW OF THERMODYNAMICS

Law of conservation of energy total energy of an isolated system is constant.
Mathematically, $\Delta u = q + w$

SIGN CONVENTION

Heat absorbed by the system = +ve
Heat evolved by the system = -ve
Work done by the system = -ve
Work done on the system = +ve

FREE EXPANSION

When an ideal gas expands in vacuum then, $p_{\text{ext}} = 0$.
 $W = 0$

Isothermal Irreversible Work

$$W_{\text{irr}} = -P_{\text{ext}}\Delta V$$

Isothermal Reversible

$$W_{\text{rev}} = -2.303nRT \log\left(\frac{V_2}{V_1}\right)$$

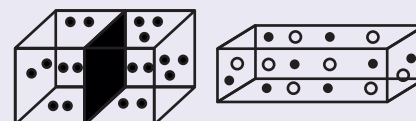
$$W_{\text{rev}} = -2.303nRT \log\left(\frac{P_1}{P_2}\right)$$

Adiabatic Reversible

$$W_{\text{rev}} = \frac{nR}{\gamma - 1} (T_2 - T_1)$$

SPONTANEITY

It is natural direction of a process.



ENTHALPY

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

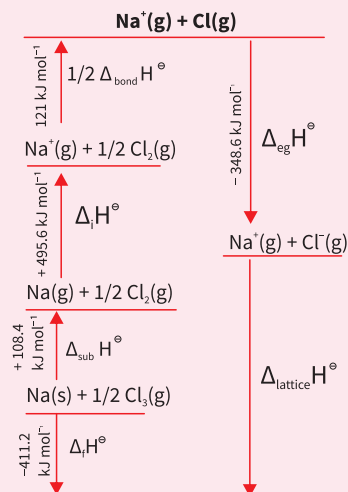
$$\Delta H = \Delta U + \Delta n_g RT$$

Enthalpy of reaction (ΔH)

$$\Delta H = \sum \Delta H_f(\text{Products}) - \sum \Delta H_f(\text{Reactants})$$

$$\Delta H = \sum B.E._{\text{Reactants}} - \sum B.E._{\text{Products}}$$

EORN HABER CYCLE



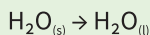
Molar HEAT CAPACITY

Amount of heat required to raise the temperature of a 1 mole substance.
Heat capacity = $q = C \times m \times \Delta T = C \Delta T$

MAYER'S FORMULA

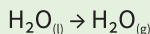
$$C_p - C_v = R$$

Enthalpy of fusion (ΔH°_f)



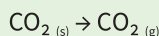
Enthalpy change when 1 mole of compound undergoes melting at constant temperature and atmospheric pressure.

Enthalpy of vapourization (ΔH°_{vap})



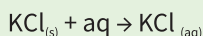
Enthalpy change when 1 mole of compound undergoes boiling at constant temperature and atmospheric pressure.

Enthalpy of Sublimation (ΔH°_{sub})



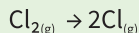
Enthalpy change when 1 mole of solid substance is directly converted into gaseous state at a constant temperature and standard pressure.

Enthalpy of Solution (ΔH°_{sol})



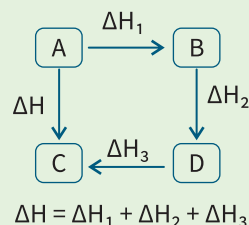
Enthalpy change when 1 mole of substance is dissolved in fixed quantity of solvent.

Bond dissociation enthalpy (B.E.)



It is the enthalpy change to break 1 mole bonds of a particular kind.

HESS' LAW CONSTANT HEAT SUMMATION



ENTROPY

This is the measure of the degree of randomness or disorder of the system.

$$\Delta S = \frac{q_{rev}}{T}$$

$$\Delta S_{Total} = \Delta S_{system} + \Delta S_{surroundings}$$

Entropy changes during phase transformation

$$\Delta S_{fusion} = \frac{\Delta H_{fusion}}{T}, \Delta S_{vap} = \frac{\Delta H_{vap}}{T}$$

$$\text{Entropy change of a reaction. } \frac{\Delta_{subs} H}{T}$$

$$\Delta_r S^\circ = \sum S^\circ_{products} - \sum S^\circ_{reactants}$$

SECOND LAW OF THERMODYNAMICS

The Total entropy of the universe is always increasing in the course of every spontaneous or natural change.

$$\Delta S_{Total} > 0$$

THIRD LAW OF THERMODYNAMICS

The Entropy of a perfectly crystalline substance at 0 K or absolute zero is zero.

GIBB'S FREE ENERGY

The net energy available to do useful work and it is a measure of spontaneity.

Standard free energy of a reaction:

$$\Delta G^\circ = \sum \Delta G^\circ_{Products} - \sum \Delta G^\circ_{Reactants}$$

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

Sign of ΔH

Negative
Positive
Positive

Negative

Sign of ΔS

Positive
Negative
Positive

Negative

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

Always Negative
Always Positive
+ve at low temp
-ve at low temp.
-ve at low temp
+ve at low temp.

Spontaneity

Spontaneous at all temperature.
Non-spontaneous at all temperature.
Non-spontaneous at all temperature.
spontaneous at high temperature.
spontaneous at low temperature.
spontaneous at low temperature.

GIBB'S ENERGY CHANGE & EQUILIBRIUM

$$\Delta G^\circ = -2.303RT \log K$$

$\Delta G^\circ = 0$ $\log K = 0$ Equilibrium reached.
 $\Delta G^\circ < 0$ $\log K > 1$ Mixture contains mostly products.
 $\Delta G^\circ > 0$ $\log K < 1$ Mixture contains mostly reactants.

$\Delta G < 0$, Process is spontaneous

$\Delta G = 0$, Process is at equilibrium

$\Delta G > 0$, Process is non-spontaneous