(c) **Internal energy :** *The internal energy of a gas is sum of internal energy due to moleculer motion (called internal kinetic energy $U_K$) and internal energy due to molecular configuration (called internal potential energy $U_{P.E.}$)*

*i.e.,* $U = U_K + U_{P.E.}$ .....(1)

(i) In ideal gas, as there is no intermolecular attraction, hence

$$U = U_K = \frac{3n}{2}RT \qquad .....(2)$$

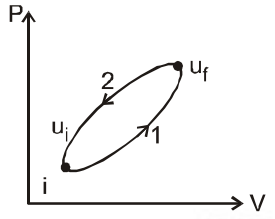
(for *n* mole of ideal gas)

(ii) Internal energy is path independent i.e., point function.

(iii) In cyclic process, there is no change in internal energy (shown in fig.)

*i.e.,* $dU = U_f - U_i = 0$

$\Rightarrow$ $U_f = U_i$




(iv) Internal energy of an ideal gas depends only on temperature eq.(2).

**First law of thermodynamics** is a generalisation of the law of conservation of energy that includes possible change in internal energy.

**First law of thermodynamics** *"If certain quantity of heat dQ is added to a system, a part of it is used in increasing the internal energy by dU and the rest is used in performing external work dW"*

*i.e.,* $dQ = dU + dW \Rightarrow dU = dQ - dW$

The quantity dU (i.e., $dQ - dW$) is path independent but dQ and dW individually are not path independent.

### Applications of First Law of Thermodynamics

(i) In **isobaric process** P is constant

so $dW = \int_{V_1}^{V_2} PdV = P(V_2 - V_1)$

so $dQ = dU + dW = n\,C_P\,dT$

(ii) In **cyclic process** heat given to the system is equal to work done (area of cycle).

(iii) In **isothermal process** temperature T is constant and work done is

$$dW = \int_{V_1}^{V_2} PdV = nRT\,\text{Log}_e\frac{V_2}{V_1}$$

Since, T = constant so for ideal gas dU = 0

Hence, $dQ = dW = nRT\ \text{Log}_e\frac{V_2}{V_1}$ (for ideal gas)

(iv) In **isochoric process** W = 0 as V = constant

It means that heat given to system is used in increasing internal energy of the gas.

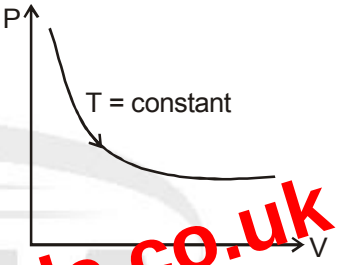
(v) In **adiabatic process** heat given or taken by system from surrounding is zero i.e., dQ = 0

$$dU = -dW = -\left[\frac{nR}{\gamma - 1}(T_1 - T_2)\right] = \left[\frac{(P_1V_1 - P_2V_2)}{\gamma - 1}\right]$$

It means that if system expands dW is +ive and dU is –ive (i.e., temperature decrease) and if system contracts dW is –ive and dU is +ive (i.e., temperature increase).

### THERMODYNAMIC PROCESSES

(i) ***Isothermal process :*** *If a thermodynamic system is perfectly conducting to surroundings and undergoes a physical change in such a way that temperature remains constant throughout, then process is said to be isothermal process.*



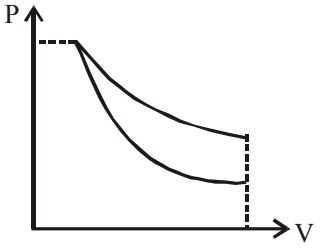

For isothermal process, the equation of state is $PV = nRT = constant$, where n is no. of moles.

For ideal gas, since internal energy depends only on temperature

$$dU = 0 \Rightarrow dQ = dW = \int_{V_1}^{V_2} PdV = nRT\int_{V_1}^{V_2}\frac{dV}{V}$$

or $dQ = nRT\,\log_e\frac{V_2}{V_1} = 2.303nRT\,\log_{10}\frac{V_2}{V_1}$

(ii) ***Adiabatic process :*** *If system is completely isolated from the surroundings so that no heat flows in or out of it, then any change that the system undergoes is called an adiabatic process.*




For ideal gas, dQ = 0

$dU = \mu C_V dT$ (for any process)

$$dW = \int_{V_1}^{V_2} PdV = \int_{V_1}^{V_2}\frac{K}{V^\gamma}dV$$

(where $PV^\gamma = K$ = constant)

$$= \frac{K}{1-\gamma}\left(\frac{1}{V_2^{\gamma-1}} - \frac{1}{V_1^{\gamma-1}}\right) = \frac{(P_2V_2 - P_1V_1)}{1-\gamma}$$