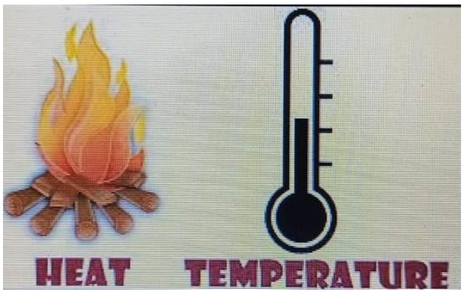
5- CONCEPT OF THERMAL PHYSICS

CONTENT: Heat , modes of heat transfer, Temperature, Scales, Absolute Zero Temperature, Pressure, volume of a gas, Gas Laws, Ideal as equation, problems.

Heat: (path dependent)



" It is an energy which can transfer from one body (Hot, high temperature) to another body (cold, low temperature) due to temperature difference between them."

(or) "it is total kinetic energy of the molecules in a gas (body)."

Unit for Heat energy is Calorie (cal) or Joule.

1 Cal = 4.2 Joule -------> chemical equivalent of heat J = 4.2 joule/cal

1 Kcal = 4.2×10^3 joule = 4200 Joule

----->Explain the modes of Heat Transfer?

Conduction:

1)It is the process of heat transfer without actual movement of particles in the medium.

2)It is medium dependent.

3)Here the particles in the medium are in vibrating motion.

4)Conduction of heat occurred in solids.

5)Conduction of heat is slow process.

6)Examples: heating metal rods, cooking vessels, ironing of cloth, heat from other person by touch.

Convection:

1)It is the process of heat transfer with actual movement of particles in the medium.

2)It is medium dependent.

3)Here the particles in the medium are in translational motion.

4)Convection of heat occurred in liquids and gases.

5)It is faster than conduction (or) it is fast process.

6)example : Boiling of water or milk, room heaters , air coolers, sea breeze, land breeze.

Radiation:

1)It is the process of heat transfer through EM waves.

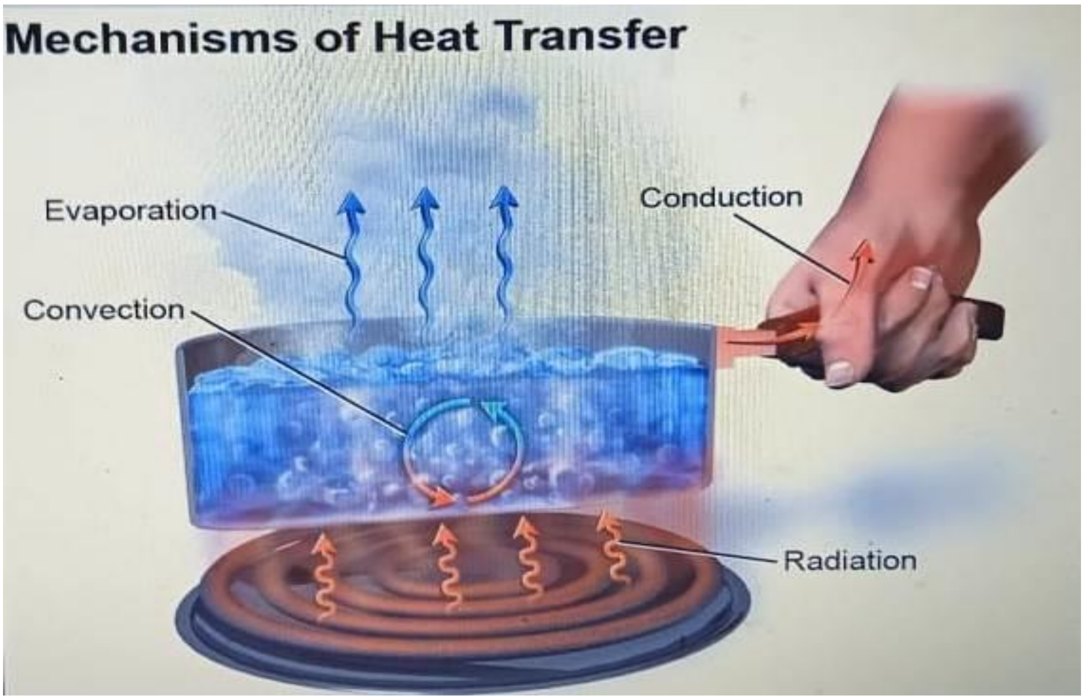
2)It may (or) may not require medium.

3)Here the particles in the medium are in translator motion , vibratory and oscillatory motion.

4)Radiation may takes place with (or) without matter(particles).

5)It is the fastest process.

6)examples: heat transfer from SUN to earth, heat received near by fire, Black body radiation.



Temperature:

It is measure of degree of hotness (or) coldness of a gas (or)body(or)the intensity of heat is called as temperature.

((OR))

It is a physical property that tells whether a body is in thermal equilibrium (or) not with another.

Using melting and boiling points of water at normal atmospheric pressure we can define 3 temperature scales.

(1)centigrade (or) Celsius scale of temperature: (C)

In this temperature scale 0^0C (ice point)is the lower temperature point, 100^0C (steam point) is the upper temperature point. Total temperature scale has 100 equal parts, each part is equal to 1^0C.

(2)Fahrenheit scale of temperature: (F)

In this temperature scale lower fixed point is 32^0F (ice point) and upper fixed point is 212^0F(steam point). Total scale is divide into 180 equal parts, each par is equal to 1°F.

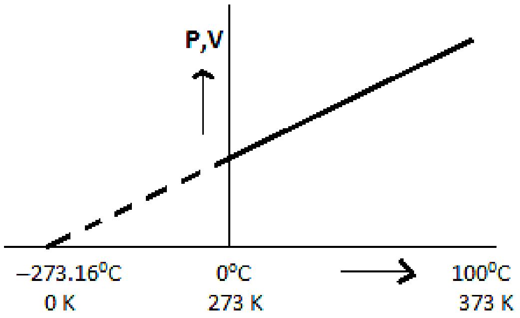
(3)Kelvin scale in temperature: (T)

In this temperature scale lowest fixed point is 273 K (ice point) and upper fixed point is 373 K (steam point). Total scale has 100 equal parts, each part is equal to 1K.

Note: Relation between above three temperature scales are $\frac{C-0}{100} = \frac{F-32}{180} = \frac{K-273}{100}$

Every gas has a definite volume and pressure at 0°C. As the temperature of a gas increases then volume and pressure are proportionally increase. This shows the dark line (straight line) in a graph. The extension of this line will touch at -273.16^{0}C. In this temperature the pressure and volume of a gas becomes Zero. This temperature is known as 0K (or) Absolute Zero.

-----→ What is Absolute Zero temperature?



The temperature at which the essential properties of a gas are found to be Zero is defined as "absolute Zero". It is denoted by 0K (or) –273^{0}C.

A temperature change in 1 Kelvin is equal to a temperature change in 1^{0}C.

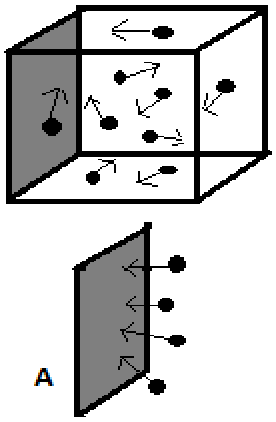
That is 0°C is equal in Kelvin scale is 0+273=273K

And 100^{0}C is equal in Kelvin scale is 100+273=373K

The relation between the Kelvin and Celsius scale of temperature is **T = 273 + t**

Pressure by a Gas:

Let certain amount of mass of gas is enclosed in a cube (or) cylinder of fixed volume. The particles in a gas are continuously moving and hitting the walls of a cube.



Let the area of one of the wall is A. Particles applying force (hitting) on this area. This force on area is known as pressure.

"So force per unit area is known as pressure of the gas".

P = $\frac{F}{A}$ --------> Newton/m^2 (or) Pascal

And P = $\frac{F}{A} = \frac{F.l}{A.l} = \frac{W}{V} = \frac{Joule}{Volume}$ -------> joule/m^3

So pressure is also a energy density. Ex:- whistling, Balloon, Pump, Cooker, Cycle tube, Water pipe.

Atmospheric Pressure:

It is the weight (force) applied by the air on unit area of the ground surface of earth. That is

$$P = \frac{F}{A} = \frac{F.l}{A.l} = \frac{\mathbf{mg}.l}{\mathbf{V}} = \frac{\boldsymbol{mg}\ .l}{(\frac{\boldsymbol{m}}{\boldsymbol{d}})} = \boldsymbol{\ell}\mathbf{dg}$$

$\boxed{P = \ell dg}$, d---->density of Mercury, g = 9.8 m/sec^2

So pressure can expressed in terms of height (L) of the mercury in barometer.

Height L = 760mm = 76cm = 0.76m,, of Hg at Sea level (low level of earth)

L = 253mm = 25.3 cm = 0.253m,, of Hg at Mount Everest (high level of earth)

At sea level, Pressure $P = 0.76 \times 9.8 \times 13.6 \times 10^3 = 101292.8$ Pascal ------> 1 atm pressure.

At Mount Everest, Pressure $P = 0.253 \times 9.8 \times 13.6 \times 10^3 = 33719.8$ Pascal -----> $\frac{1}{3}$ atm Pressure.

So $P_{sea} > P_{Everest}$ -----→ due to this reason boiling of food will take more time on mountains.

Volume of a gas:

The space occupied by a gas in the universe is known as volume of the gas. It is measured in litre (or) m^3 (or) CC that is cm^3 .

Volume occupied by a 1 gram mole of gas at room temperature (NTP) and at 1 atm pressure is V =22.4 litre.

$V = 22.4$ litre $= 22.4 \times 1/1000\ m^3 = 0.0224\ m^3 = 22400\ cm^3 = 22400$ CC $= (28.19\ cm)^3$

Thermodynamic system and surroundings:

A cylinder containing gas with movable piston is known as System. (OR) a definite quantity of matter bounded by some closed surface is refers to a thermodynamic system. The walls of a cylinder is either conducting or insulating. The outer part of the system which influences on it is known as surroundings.

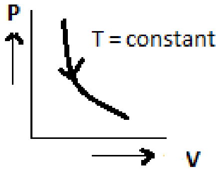
Gas laws: These are the laws to study the behaviour of gases with the relations between pressure(P),volume(V),temperature(T). These are (1) Boyle's law

(2)Charles 1st law (or) constant pressure law

(3)Charles 2nd law (or) constant volume law

(1)Boyle's law with graphical explanation:

"At constant temperature, the volume of a given mass of gas is (varies) inversely proportional to its pressure." Ex: compression of rubber ball (or) sponge. If P and V are the pressure and volume of a gas at constant temperature,



then $\mathbf{P} \propto \frac{1}{V}$ ---> PV=constant ----> energy =constant

if volume varies as V_1, V_2, V_3.....then corresponding pressures are P_1, P_2, P_3 ......

then $P_1 V_1 = P_2 V_2 = P_3 V_3 = \ldots\ldots\ldots =$ constant

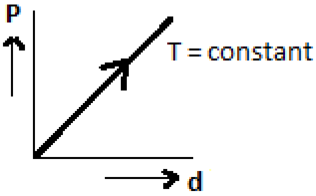
This graph shows the rectangular parabola.

$P \propto \frac{1}{V}$ ------> $P = \mathbf{k} \left(\frac{1}{V}\right)$

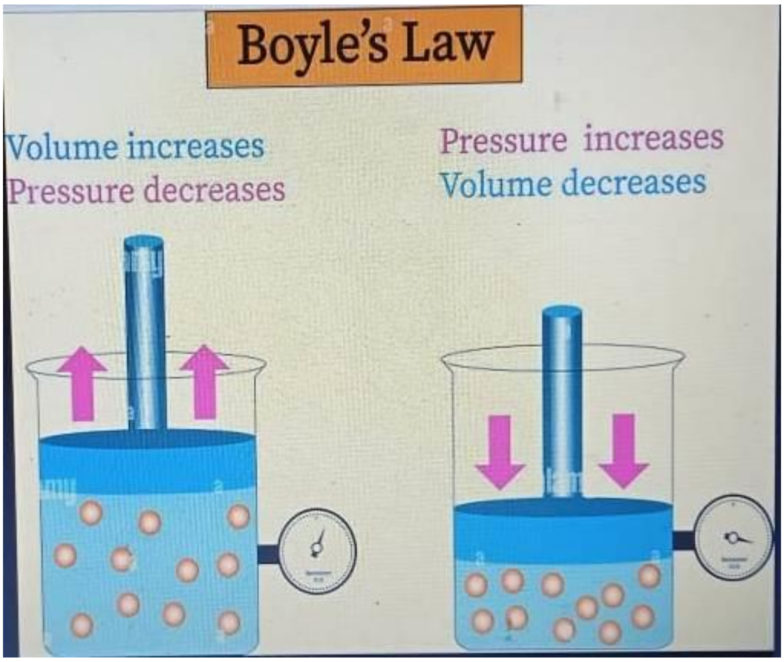
Density of a gas is inversely proportional to volume then $d = \frac{m}{V}$---->$V = \frac{m}{d}$---->$V_1 = \frac{m}{d_1}$, $V_2 = \frac{m}{d_2}$

Then $P_1 V_1 = P_2 V_2$ -------> $\frac{P_1 \cancel{m}}{d_1} = \frac{P_2 \cancel{m}}{d_2}$ ----> $\frac{P_1}{d_1} = \frac{P_2}{d_2}$ = constant(K)

$\frac{P}{d}$ = constant -----> $P \propto d$



"Hence at constant temperature, pressure of a given mass of gas is directly proportional to density." Example: Real gases

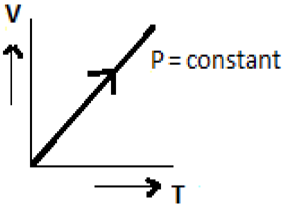


Charles 1st law with graphical explanation:

At constant pressure, the volume of a fixed mass of a gas at 0°C is V_o , at 1°C it is increased by $\frac{1}{273}$ of V_o. That is at 1°C , volume of gas $V_1 = V_o + \frac{Vo}{273}$

at 2°C , volume of gas $V_2 = V_o + \frac{2Vo}{273}$

at 3°C , volume of gas $V_3 = V_o + \frac{3Vo}{273}$



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at t°C , volume of gas $V_t = V_o + \frac{tVo}{273} = V_o[1 + \frac{t}{273}] = V_o \frac{[273+t]}{273}$

$V_t = V_o[\frac{T}{T_0}]$ -------> $\frac{V_t}{T_t} = \frac{V_0}{T_0}$-----> $\frac{V}{T}$ =constant -----> $V \propto T$ at constant pressure

Hence Charles 1st law stated as "**at constant pressure, the volume of a given mass of a gas is directly proportional to its absolute temperature.**"

That is V α T -------> $\frac{V_1}{T_1} = \frac{V_2}{T_2}$

Ex: liquid(mercury) thermometer. P=1atm=constant------> V α T ---> T↑ then V ↑

Mercury freeze at –38.8°C and boils at 356°C. So mercury thermometer can't be used at very low & very high temperature. Alcohol thermometer up to –112°C.

Density of a gas is inversely proportional to volume then $d = \frac{m}{V}$---->$V = \frac{m}{d}$ ----> $V_1 = \frac{m}{d_1}$, $V_2 = \frac{m}{d_2}$

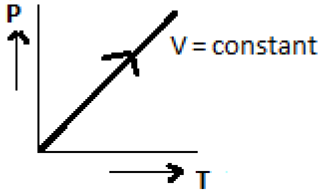
V α T -----> Then $\frac{V_1}{T_1} = \frac{V_2}{T_2}$ -------> $\frac{\cancel{m}}{T_1 d_1} = \frac{\cancel{m}}{T_2 d_2}$ ----> $T_1 d_1 = T_2 d_2$ = constant(K)

Td = constant -----> $d \propto \frac{1}{P}$ ----->density of gas is inversely proportional to pressure.

And at absolute Zero T_t = 0 K (or) t = –273°C then $V_t = V_o = \frac{[273-273]}{273} = 0$ ---->Zero volume.

Charles 2nd law with graphical explanation:

At constant volume, the pressure of a fixed mass of a gas at 0°C is P_o, at 1°C it is increased by 1/273 of P_o.



That is at 1°C , pressure of gas $P_1 = P_o + \frac{Po}{273}$

at 2°C , pressure of gas $P_2 = P_o + \frac{2Po}{273}$

at 3°C , pressure of gas $P_3 = P_o + \frac{3Po}{273}$

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at t°C , pressure of gas $P_t = P_o + \frac{t\,Po}{273} = P_o[1 + \frac{t}{273}] = P_o\frac{[273+t]}{273}$

$P_t = P_o[\frac{T_t}{T_0}]$ -------> $\frac{P_t}{T_t} = \frac{P_0}{T_0}$-----> $\frac{P}{T}$ = constant

-----> **P α T** at constant volume

Hence Charles 2nd law stated as **"at constant volume, the pressure of a given mass of a gas is directly proportional to its absolute temperature."** That is P α T -------> $\frac{P_1}{T_1} = \frac{P_2}{T_2}$

Ex: Gas thermometer. V=constant------> P α T ----> T↑ then P ↑

Hydrogen thermometer -----> –200°C to 500°C
Helium thermometer ----------> below –200 °C
Nitrogen thermometer ---------> up to 1500 °C

and at absolute Zero that is T_t = 0 K (Kelvin) (or) t = –273°C then $P_t = P_o\frac{[273-273]}{273} = 0$ ---->Zero pressure.

Ideal gas: A gas which obeys Boyle's and Charles laws at all temperatures and pressures is called an ideal gas. In general there is no ideal gases. It is assumption only. It is comparing purpose only.

Characteristics of ideal gas:

1)Ideal gas has very high number of molecules are in a container.

2)sizes of the gas molecules are negligibly small when compared to volume of container.

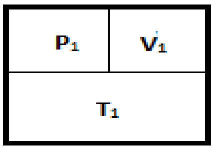
3)There are no intermolecular forces of attraction (or) repulsion among the molecules. ----->potential energy=0

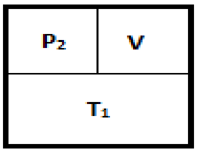
4)collisions among the molecules are perfectly elastic. That is total internal energy(U) = sum of all kinetic energies of the molecules.

------->Derivation of Ideal gas Equation:

The relationship among pressure, volume and temperature is called as gas equation. It is the combined relation of Boyle's and Charles laws.

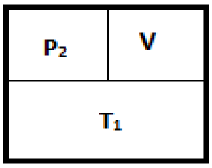
Consider a fixed mass of a gas at pressure P_1, volume V_1 and absolute temperature T_1. If the pressure and volume of a gas will change to P_2 and V at constant temperature T_1.

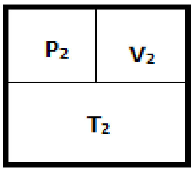




then from Boyle's law

$P_1 V_1 = P_2 V$ ------> $V = \frac{P_1V_1}{P_2}$ --------(1)





Now the gas is heated to temperature T_2 at constant pressure P_2 then it's volume changes to V_2.

then from Charles 1st law

$\frac{V}{T_1} = \frac{V_2}{T_2}$ --------> $V = \frac{V_2T_1}{T_2}$ --------->(2)

from equations (1) and (2) $\frac{P_1V_1}{P_2} = \frac{V_2T_1}{T_2}$ ------- $\boxed{\frac{P_1V_1}{T_1} = \frac{P_2V_2}{T_2}}$

----> $\frac{PV}{T}$ = constant (**R**) for 1 gram mole of gas.

----> PV=**n**RT for **n** mole of a gas. Here **R**=universal gas constant

Since density $d = \frac{m}{V}$ ----> volume $V = \frac{m}{d}$, mass m=constant ----- $V \propto \frac{1}{d}$

Then ideal gas equation $\frac{P}{dT}$ = constant --- $\boxed{\frac{P_1}{d_1T_1} = \frac{P_2}{d_2T_2}}$

This equation represents the gas equation in terms of density.

Universal gas constant: (R)

Gas equation for 1 gram mole of gas is $\frac{PV}{T}$ = constant(R)

Here in SI units pressure of a gas in STP is **P = hdg,**

here h = 76cm = 0.76m,, of Hg at NTP,
d = mercury density = 13.6×10^3 kg/m^3
g = 9.8 m/sec^2
therefore pressure P = $0.76 \times 13.6 \times 10^3 \times 9.8$ = 101292.8 Pascal ------> **1 atm**

volume occupied by 1 gram molecular gas is V = 22.4 litre = 0.0224 m^3

temperature T = 273 K that is t = 0^{o}C

therefore gas constant **R** = $\frac{PV}{T}$ = 101292.8 x 0.0224 / 273

R = 8.31 Joule/mole/Kelvin ------> SI system

= 8.31×10^7 erg/mole/Kelvin -------> C.G.S system

R = 8.31/4.2 = 1.99 Cal/mole/Kelvin ------> heat terms

This is same for all gases. So for n gram mole of a gas, the gas equation is PV=**nR**T

Where n = $\frac{m}{M}$ = mass of a gas in grams/molecular weight.

Ex: 64 grams of O_2 gas ----> n = 64/32 = 2 moles---> $2 \times 6.023 \times 10^{23}$ molecules ------> PV=**2R**T

Note: 1)Water can boil (change to vapour) at 100^{o}C temperature, 1 atm pressure. Water can also boil below 100^{o}C, due to decrease in pressure. Ex: At Himalayas water boiled at 98^{o}C because there is 1/3 atm atmosphere pressure at constant volume.

2)Ice can melt (change to liquid water) at 0^{o}C, 1 atm pressure. Ice can also melt below 0^{o}C, due to increase in pressure at constant volume.

General gas constant: (r)

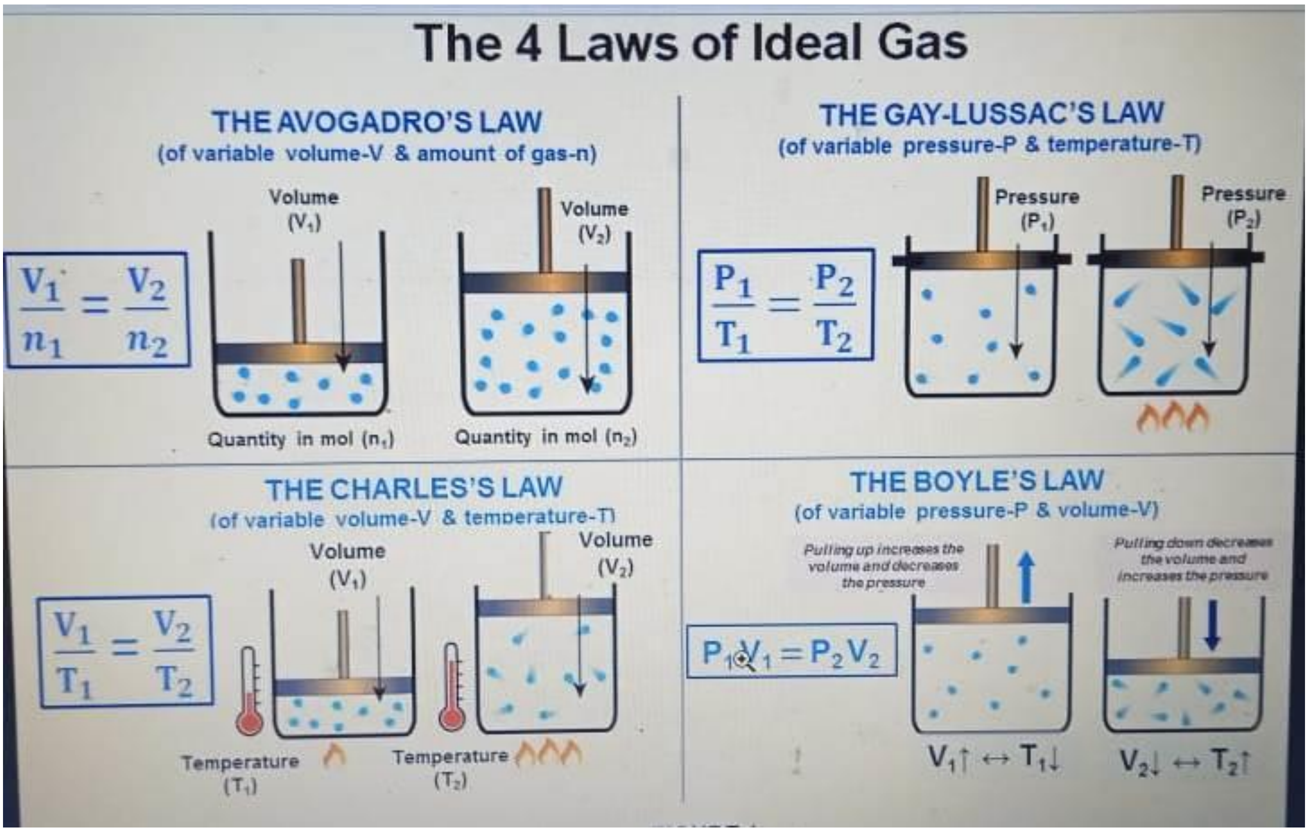
When 1 gram of gas at NTP (t =27^{o}C=300K, P=1bar = 0.987 atm) is taken then the gas constant is denoted by '**r**', then the gas equation is PV = **r**T. 'r' value varies from gas to gas.

and **r** = $\frac{universal\ gas\ constant}{molecular\ weight\ of\ a\ gas}$ = $\frac{\mathbf{R}}{\mathbf{M}}$

For '**m**' grams, gas equation is PV = **mr**T = **m** $\frac{\boldsymbol{R}}{\boldsymbol{M}}$ T

------->Distinction between gas constant (r) and universal gas constant(R)?

General gas constant (r)	Universal gas constant(R)
1)It is the value of PV/T when 1 gram of gas is considered. 2)It varies from gas to gas. 3) It's SI unit is Joule/Kg/Kelvin 4)dimensional formula of R is $M^0 L^2 T^{-2} K^{-1}$	1)It is the value of PV/T when 1 gram mole of gas is considered. 2)It is constant for all gases. 3)It's SI unit is Joule/mole/Kelvin 4)dimensional formula of R is $M^1 L^2 T^{-2} K^{-1}$



The 4 Laws of Ideal Gas

THE AVOGADRO'S LAW

(of variable volume-V & amount of gas-n)

$$\frac{V_1}{n_1} = \frac{V_2}{n_2}$$

THE GAY-LUSSAC'S LAW

(of variable pressure-P & temperature-T)

$$\frac{P_1}{T_1} = \frac{P_2}{T_2}$$

THE CHARLES'S LAW

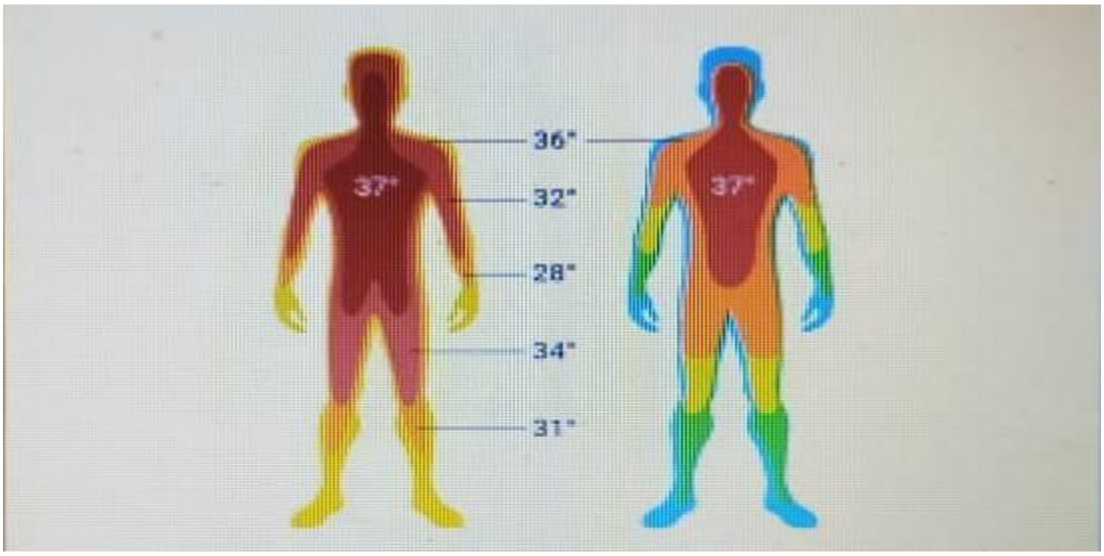
(of variable volume-V & temperature-T)

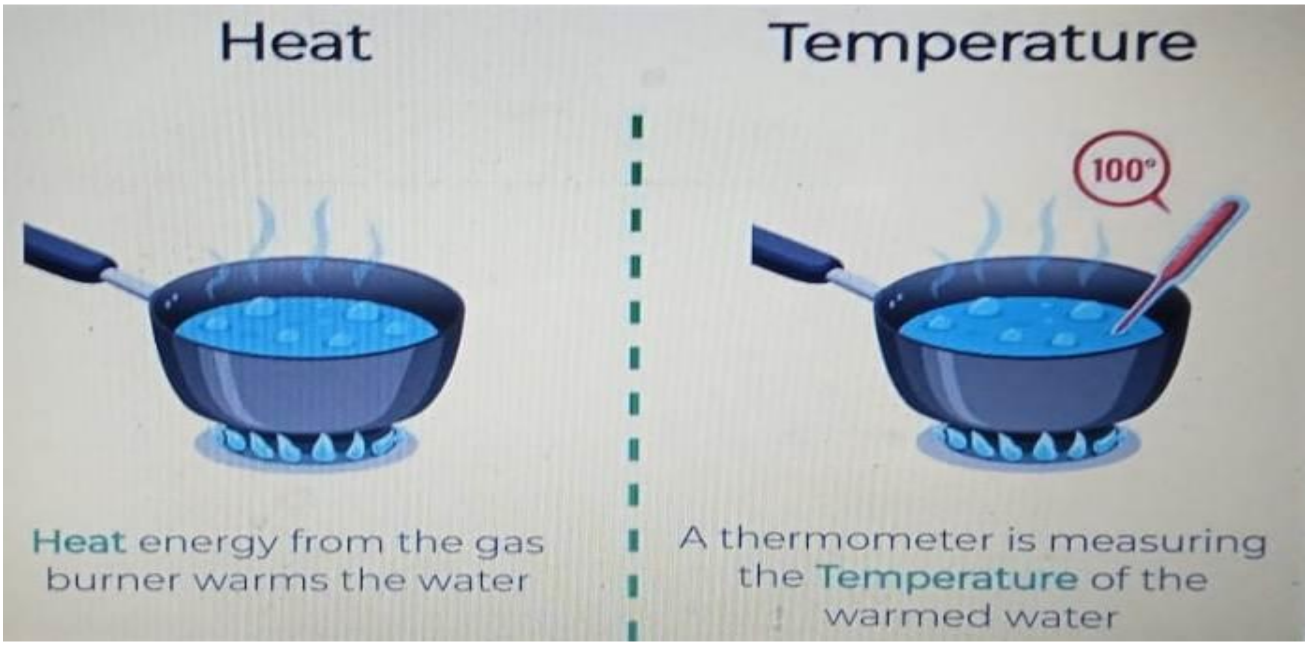
$$\frac{V_1}{T_1} = \frac{V_2}{T_2}$$

THE BOYLE'S LAW

(of variable pressure-P & volume-V)

$$P_1V_1 = P_2V_2$$





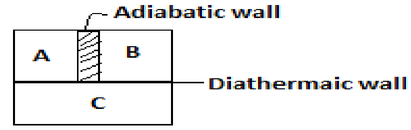
Heat

Heat energy from the gas burner warms the water

Temperature

A thermometer is measuring the **Temperature** of the warmed water

Zeroth law of thermodynamics:



"If two bodies (or) systems A and B are separately in thermal equilibrium with a third body C, then A and B are in thermal equilibrium with each other."

First law of thermodynamics:

The heat energy supplied to the system is equal to the sum of increase in internal energy and work done by a gas.

$$\boxed{dQ = dU + dW}$$

Ex: working of diesel engine, petrol, steam engines. The explosions of a fire cracker in an insulated steel drum.

Internal energy of a system:

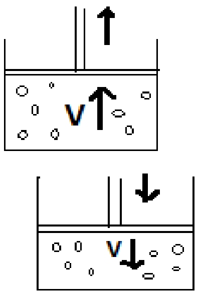
In the gas , molecules (or) atoms are in a state of constant rapid motion and it has internal kinetic energy.

The energy arising due to the inter molecular attraction is called the internal potential energy. Gas molecules generally have rotational, vibration energies.

"The sum of internal kinetic energy, internal potential energy, rotational and vibration energies is known as total internal energy of a gas." $\Delta U = U_f - U_i$

Ex: energy is released when water changes to ice. This stored energy is called the internal energy. If we heat a system it's internal energy increases, If we push the piston down (or) when it cooled then it's internal energy decreases.

External Work: (path dependent)

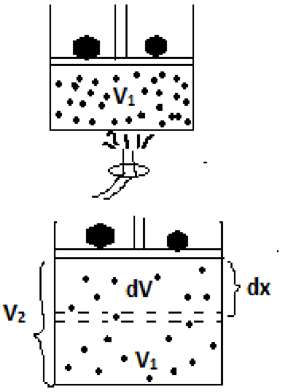


If the system does work on surroundings (or) surrounding does some work on system then the work done is called as external work.

If the piston goes up by increase in volume as temperature increases, it is known as system done work on surroundings(work done by the system).

If the piston comes down (pressure increases) then volume will decrease, it is known as work done on the system by the surroundings.

------->Derivation of work done in expanding a gas at constant pressure:



Consider a gas in a cylindrical vessel with a movable piston. Initial volume of the gas is V_1 , pressure on a gas is fixed by taking fixed weights on piston.

When the gas is heated at this constant pressure it will moves the piston up with distance dx, then volume increases (gas expand) to V_2 , with the amount dV. That is $V_2 = V_1 + dV$ --------> $dV = V_2 - V_1$

Then the work done by the gas on piston is dW = force on piston x distance = F x dx

But pressure $P = \frac{F}{A}$ -------> F = PA , here A----> area of piston

therefore dW = PA x dx = P x (A dx) ,here dV = A dx =increase in volume.

------> $\mathbf{dW = P (V_2 - V_1)}$

$$\boxed{\mathbf{dw = P\ dV}}$$

------> $dw = nr (T_2 - T_1)$

Thermodynamics:

It is the branch of physics which can deals the transformation of heat (thermal) energy into mechanical work (or) work to heat.

Ex: 1)Train engine(old), Piston moving ------>Heat energy into Mechanical work conversion

2)Rubbing of iron --------> Mechanical work into Heat energy

It is divided into two parts: 1)deals the relations between different energies. That is first law of thermodynamics.

2)deals the direction of heat changing. That is second law of thermodynamics.

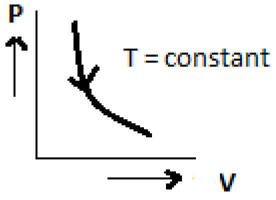
Isothermal Process (changes):

It is the process at which the pressure and volume are change at constant temperature with the exchange of heat energy between the system and surrounding by conducting walls.

Ex: when a gas compressed slowly, the heat generated is gives up (conducted) to the surroundings and if a gas expand slowly, slight cooling will takes place, here heat is taken from surroundings to makes temperature constant.

So isothermal process is a slow process. And T = constant ------- > dT = 0 change in temperature is Zero

------> dU = 0 ----> dQ = dW



From gas equation PV/T = constant -----> PV = constant ----->

So isothermal process obeys Boyle's law

Ex: 1)Conversion of water at 100°C into steam of 100°C.

2)Conversion of ice at 0°C into water at 0°C.

Adiabatic Process: (Changes)

It is the process at which the pressure, volume and temperature are change at constant heat. Here there is no heat energy between the system and surroundings.

Ex: 1)When a gas compressed suddenly, then it's temperature increases, volume decreases but there is no time for heat to exchange with surrounding.

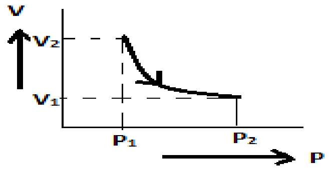
2)Similarly when a gas expands suddenly, then it's temperature decreases, volume increases but there is no time for heat to exchange with surroundings.

So adiabatic process is a sudden(quick) process, and Q=constant----->dQ = 0 ----> dU = –dW

This process can be represented by PV^{γ} =constant

-----> $TV^{\gamma-1}$ = constant

------> $T^{\gamma}P^{1-\gamma}$ =constant



Here $\gamma = C_p/C_v$ ---->Ratio between two specific heats.

For ideal gas γ =5/3 , for diatomic gas γ = 7/5

3)Cycle tube (or) football bursts suddenly, then air in it expands and cools.

Isochoric process:

A process can takes place at constant volume is known as isochoric process.

That is V = constant-----> dV = 0------>work done dW = P dV = P (0) = 0-------> therefore dQ = dU

So total heat supplied (or) rejected is equal to increase (or) decrease in internal energy of the system.

Here Charles 2nd law holds good ------> V = constant -----> P α T ----> $\frac{P_1}{T_1} = \frac{P_2}{T_2}$

Ex: 1)Atmosphere change 2)Explosion in gas

Isobaric process:

A process that occurs at constant pressure is known as isobaric process.

Here Charles 1st law holds good ------> P = constant -----> V α T ----> $\frac{V_1}{T_1} = \frac{V_2}{T_2}$

Reversible process:

A process which can be returned in the opposite direction from final state to initial state is called reversible process. Ex: 1) ice + heat = water ---->vaporisation of water

2) water – heat = ice ------>fusion of ice

3)mass attached to spring

Irreversible process:

A process which can not be returned to initial state is called irreversible process. Almost all processes of nature are irreversible.

Ex:1) Heat produced due to friction (or) work done against friction.

2)Sugar dissolved in water.

3)Rusting of iron in the environment

Cyclic process:

A process which can be taken from initial to final by one or more processes and returned from final to initial state by one or more processes, then it is known as cyclic process.

That is change in internal energy dU = 0 -------->Hence dQ = dW

Thus heat given to a system is equals to work done by the system. Ex:1)water (yes) 2)petrol (no)

Limitations of first law of thermodynamics:

1)First law does not tell about the direction of heat flow, that is it does not specify the condition under which a body can use its heat energy to produce work.

2)It does not tells us about the efficiency with which heat can be converted into work.

Second law of thermodynamics:

"Heat can not flow from low temperature body to high temperature body without any external work done on the system." --------> it tells direction of flow.-------> Scientist Clausius statement

Ex:1)Refrigerator, water coolers, cold storage units can works on this law.

------->Differences between Isothermal and Adiabatic processes:

Isothermal	Adiabatic
1)In this pressure, volume, heat are changes at constant temperature.	1)In this pressure, volume, temperature are changes a constant heat.
2)Here conducting walls are needed.	2)Here non conducting walls are needed.
3)Internal energy is constant	3)Internal energy changes.
4)It is slow process	4)It is quick process.
5)PV = constant holds good	5) PV^{γ} =constant holds good. Here $\gamma = \frac{C_p}{c_v}$
6)Entropy changes	6)Entropy is constant

Specific Heat:

The increase in temperature of the body due to heat supplied will depends on nature of the body. That nature is called specific heat (constant).

Heat energy supplied is directly proportional to mass and temperature changes (irrespective of shape of the body). Water **C** = 1 cal/gm °c = 4200 J / kg Kelvin

Copper **C** = 0.093 Cal /gm °c

Iron **C** = 0.11 Cal /gm °c

Glass **C** = 0.2 Cal /gm °c

$dQ \propto m$, $dQ \propto dT$ ---------> $dQ \propto m\, dT$ -------> $dQ = \mathbf{C}\, m\, dT$, C------> proportionality constant

$\boxed{\mathbf{C} = \frac{dQ}{m\, dT}}$ units ----> Joule / kg Kelvin (or) Cal/gm kelvin

Dimensions -----> $M^o L^2 T^{-2} K^{-1}$

Definition:

"It is the amount of heat required to raise the temperature of unit mass of the substance through 1°C (or) 1 Kelvin."

In the case of solids and liquids there is only one specific heat because only volume changes with temperature.

In the case of gases both pressure and volume changes with temperature. So gas has two specific heats.

Specific heat of gas at constant volume (c_v):

"It is the amount of heat required to raise the temperature of unit mass of the substance through 1°C (or) 1 Kelvin at constant volume." $c_v = (\frac{dQ}{m\, dT})_v$ ------>Joule/(kg. Kelvin)

Specific heat of gas at constant pressure (C_p):

"It is the amount of heat required to raise the temperature of unit mass of the substance through 1°C (or) 1 Kelvin at constant pressure." $c_p = (\frac{dQ}{m\, dT})_p$ ------>Joule/(kg .Kelvin)

-------> Why $c_p > c_v$? Explain it :

When a gas is heated at constant volume, the supplied heat is utilised only to increase the internal energy of the gas. Because V =constant

$$dV = 0 \text{ ------> } dQ = dU + \cancel{P\,dV} \text{ -------> } dQ = dU$$

but when a gas is heated at constant pressure, the heat supplied is utilised for two purposes.

1)For increasing the internal energy of the gas.
2)To do external work in expansion to keep the pressure constant.

So to raise the temperature, more heat is required at constant pressure that at constant volume.

That is $\mathbf{c_p > c_v}$ ------> $\frac{Cp}{Cv} > 1$ -------> $\Upsilon > 1$, $\gamma = \frac{Cp}{Cv}$ = Specific heat ratio

For ideal gas $\gamma = 5/3$, for monoatomic gas $\gamma = 1.67$,
for Diatomic gas $\gamma = 7/5 = 1.4$, for Triatomic gas $\gamma = 1.33$

Molar Specific heat (or) Heat capacity:

It is the amount of heat required to raise the temperature of 1 gram mole of gas through 1°C (or) 1K

$dQ \,\alpha\, n$, $dQ \,\alpha\, dT$ ---------> $dQ \;\alpha\; n\,dT$ -------> $dQ = \mathbf{C}\,n\,dT$, here **C** ------> proportionality constant

Molar Specific heat = Heat energy supplied/ temperature difference -----> $\mathbf{C} = \frac{dQ}{n\,dT}$ Joule / Kelvin.mole

Molar specific heat = molecular weight x specific heat ---->$\mathbf{C} = M \times c$, Dimensions -----> $M^1 L^2 T^{-2} K^{-1}$

Molar specific heat of a gas at constant volume (C_v):

It is the amount of heat required to raise the temperature of 1 gram mole of gas through 1°C (or) 1K at constant volume. $\mathbf{C_v} = M \times \mathbf{c_v}$ -----> $\mathbf{C_v} = \left(\frac{dQ}{n\,dT}\right)_v$

Molar specific heat of a gas at constant Pressure (C_p):

It is the amount of heat required to raise the temperature of 1 gram mole of gas through 1°C (or) 1K at constant pressure. $\mathbf{C_p} = M \times \mathbf{c_p}$ ------> $\mathbf{C_p} = \left(\frac{dQ}{n\,dT}\right)_p$

------->Show that $C_p - C_v = R$ (or) $c_p - c_v = r$ ----Mayer's formula .

Consider '**n**' gram mole of ideal gas at pressure P, volume V, temperature T. Suppose that gas is heated at constant volume(dV=0) then the molar specific heat of the gas is $\mathbf{C_v} = \left(\frac{dQ}{n\,dT}\right)_v$

Now from 1st law of thermodynamics $dQ = dU + \cancel{P\,dV}$ ----> $dQ = dU$

Therefore $\mathbf{C_v} = \left(\frac{dU}{n\,dT}\right)_v$ -----> $dU = \mathbf{n}\,C_v\,dT$ ----->(1) it represents change in internal energy.

Now the gas is heated again at constant pressure(dP = 0) then the molar specific heat of the gas is

$\mathbf{C_p} = \left(\frac{dQ}{n\,dT}\right)_p$ ------> $dQ = \mathbf{n}\,C_p\,dT$ ------> (2) it represents amount of heat supplied.

Now from 1st law of thermodynamics $dQ = dU + P\,dV$

-----> $\mathbf{n\,C_p}\,dT = \mathbf{n\,C_v}\,dT + P\,dV$

$\mathbf{n}[\mathbf{C_p} - \mathbf{C_v}]\,dT = P\,dV$ ------->(3)

but from ideal gas equation $PV = \mathbf{nR}T$ for n mole gas.

$P\,dV + \cancel{V\,dP} = \mathbf{n\,R}\,dT$ (since now P= constant)

$P\,dV = \mathbf{n\,R}\,dT$

Now equation (3) becomes $\mathbf{n}[\mathbf{C_p} - \mathbf{C_v}]\,dT = \mathbf{n\,R}\,dT$ -------> $\mathbf{C_p} - \mathbf{C_v} = R$ for ideal gas.

That is difference in molar specific heats is equal to universal gas constant (**R=8.31**)

In general for any other gases, it will becomes $\mathbf{Mc_p - Mc_v = R}$ -----> $\mathbf{c_p - c_v} = \frac{R}{M}$ ----> $\mathbf{c_p - c_v = r}$

That is difference in specific heats is equal to general gas constant.