STATES OF MATTER

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❖ **STATES OF MATTER**

INTRODUCTION

Matter is made up of atoms or molecules. The arrangement of these molecules determines the state of matter. There are three recognized states of matter: Solid, Liquid and Gas. Matter can change between states when the temperature or pressure is changed. State changes of matter are physical rather than chemical.

DIFFERENCE BETWEEN STATES OF MATTER

S. NO.	GAS	LIQUID	SOLID
i	Assumes the Shape & volume of its container.	Assumes the shape of the part of a contaiener which occupies.	Retains a fixed volume and shape.
ii	Particles can move past one another.	Particles can move/slide past one another.	Rigid particles locked into one place
iii	Compressible, lots of free space between particles.	Not easily compressible little space between particles.	Not easily compressible, little free space between particles.
iv	Flows easily, particles can move past one another	Flow easily, particles can move past one another.	Doesn't flow easily, rigid particles cannot move past one another.
v	Low density	Intermediate density	High density
vi	Very small intermolecular attraction but high kinetic energies.	Considerable intermolecular attraction, kinetic energy is less.	Intermolecular forces are high, vibrational motion only.

GASEOUS STATE

1. **Important Properties of Gases**

(i) **Mass:** Mass in gm = Moles × Molecular mass.

(ii) **Volume :** Volume of the gas is the volume of container in which they are filled in.

(iii) **Temperature:** Temperature of a gas is the measure of kinetic energy of gas.

Kinetic energy $\propto$ Temperature

(iv) **Pressure :** Pressure of gas is defined as the force exerted by the gas on the walls of its container. It is often assumed that pressure is isotropic, i.e. it is the same in all the three directions.

Pressure $= \frac{Force}{Area}$,

$$P = \frac{Mg}{A} = \frac{v \times d \times g}{A}$$

$$= \frac{A \times h \times d \times g}{A}$$

$$P = hdg$$

Where.... h = height of the mercury column supported by the barometer.

d = density of mercury.

g = acceleration due to gravity.

2. Units and Determination of Pressure of Gas

(a) In SI unit the unit of pressure is the pascal (N/m^2) instead, the unit bar, kPa or MPa is used.

$$1 \text{ bar} = 10^5 \text{ N/m}^2 = 100 \text{ kN/m}^2 = 100 \text{ kPa}$$

(b) Pressure are also stated in mm or cm of mercury.

$$1 \text{ atm} = 760 \text{ mm Hg} = 1.01325 \text{ bar} = 1.01325 \times 10^5 \text{ Pa}$$

$$= 101.325 \text{ KN/m}^2 = 1.0332 \text{ Kgf/cm}^2$$

$$1 \text{ Pa} = 1 \text{ Nm}^{-2} = 1 \text{ Kgm}^{-1} \text{ S}^{-1}$$

$$1\text{L} = 1\text{dm}^3 = 10^{-3} \text{ m}^3 \text{ (SI unit)}$$

$$1\text{L atm} = 101.325 \text{ J}$$

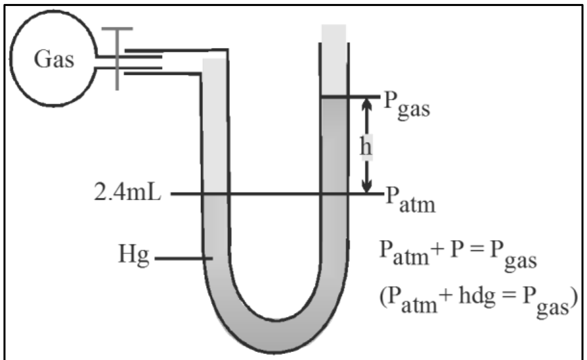
$$1 \frac{\text{KN}}{\text{m}^2} = 1 \times 10^3 \frac{\text{N}}{\text{m}^2} = \frac{1 \times 10^3 \times \text{kg}}{9.8 \times 10^4 \text{ cm}^2} = \frac{1}{98} \text{kgf/cm}^2$$

$$1\text{Torr} = \frac{101325}{760} \text{ Pa} = 133.322 \text{ Pa}$$

(d) The pressure relative to the atmosphere is called gauge pressure. The pressure relative to the perfect vacuum is called absolute pressure.

Absolute pressure = Gauge pressure + Atmosphere pressure

(e) When the pressure in a system is less than atmospheric pressure, the gauge pressure becomes negative, but is frequently designated and called vacuum.



An open arm manometer

For example, 16 cm vacuum will be -

$$\frac{76-16}{76}\times 1.013 = 0.80 \text{ bar}$$

(f) Pressure is measured using a manometer which is a simple device with a horizontal arm and a U-tube carrying mercury or any other suitable liquid.

Ex. Calculate the mass of mercury in a uniform column 760 mm high and 1.00 cm^2 in cross-sectional area. Is there any change in.

(a) mass and

(b) pressure of column of same height but with 2.00cm^2 cross sectional area is taken?

(density of Hg = 13.6 g/cm^3)

Sol. $760 \text{ mm} = 76 \text{ cm}$

$\therefore \quad V = 76 \times 1 = 76 \text{ cm}^3$

$\therefore \quad \text{Mass} = V \times d = 76 \times 13.6 = 1033.6 \text{ g}$

(a) If area of cross-section is 2cm^2 (doubled) then volume is also doubled hence,

$\therefore \quad \text{Mass} = 76 \times 2 \times 13.6 = 2067.2 \text{ g}$

(b) This mass would rest on twice the area and so exert the same pressure.

GAS LAWS

The gas laws are a set of laws that describe the relationship between thermodynamic temperature (T), pressure (P) and volume (V) of gases.

1. Boyle's Law

It states "at a constant temperature (T), the pressure (P) of a given mass (or moles, n) of any gas varies inversely with the volume (V)". Pressure (P) and volume (V) of gases.

i.e. $P \propto \frac{1}{V}$ (for given n and T)

or, $PV = K = \text{constant}$

For two or more gases at constant temperature

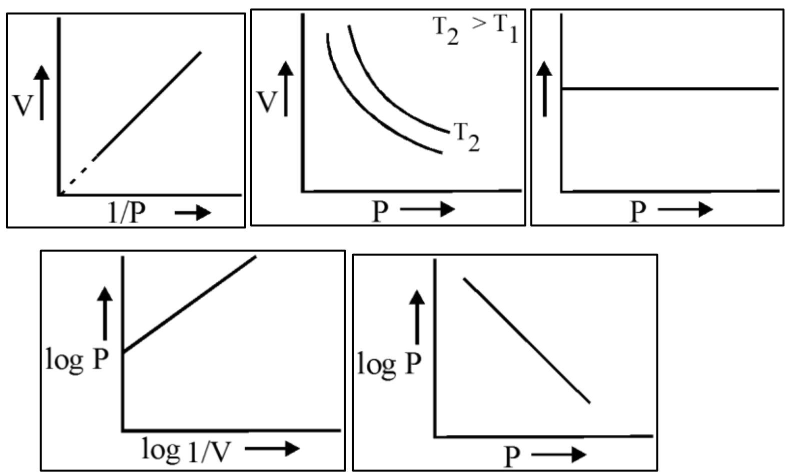
$$P_1V_1 = P_2V_2 = ... = K$$

Also, density $d \propto \frac{1}{V}$

Hence $P \propto d$

or $\frac{P_1}{d_1} = \frac{P_2}{d_2} = ... = K$

◆ **Graphical Representation**



These plots drawn at constant temperature for a gas are called Isotherms.

Ex. The diameters of a bubble at the surface of a lake is 4 mm and at the bottom of the lake is 1 mm. If atmospheric pressure is 1 atm and the temperature of the lake-water and the atmosphere are equal, what is the depth of the lake? (The density of lake-water and Hg are 1 gm/ml and 13.6 gm/ml respectively. Also neglect the contribution of pressure due to surface tension).

Sol. The pressure on the bubble = 1 atm (when it is at the surface)

The pressure on the bubble = P atm (say) (when it is at the bottom)

The volume of the bubble $= \frac{1}{6}\pi\,(0.1)^3$

$$P \times \frac{1}{6}\,\pi \times (0.1)^3 = 1 \times \frac{1}{6}\,\pi\,(0.4)^3$$

or $P = 64$ atm. Thus, the pressure due to water is 63 atm

($\because$ atmospheric pressure = 1 atm)

Now, consider the depth of lake is h cm

$\therefore \quad 63 \times 76 \times 13.6 \times g = h \times 1 \times g$

or $\quad h = 63 \times 76 \times 13.6$ cm $= 65116.8$ cm.

2. Charles' Law

It states "at constant pressure, the volume of a given mass of a gas, increases or decreases by $\frac{1}{273.15}$ th of its volume at 0ºC for every rise or fall of one degree in temperature".

$$\frac{V_t}{V_0} = 1 + \frac{1}{273.15}t \qquad \text{(at constant n and P)}$$

or $$V_t = V_0\left(1 + \frac{t}{273.15}\right)$$

or $$V_t = \frac{V_0(273.15 + t)}{273.15}$$

0ºC on the Celsius scale is equal to 273.15 K at the Kelvin or absolute scale.

i.e., T_t (Temperature in Kelvin scale) $= 273.15 + t$

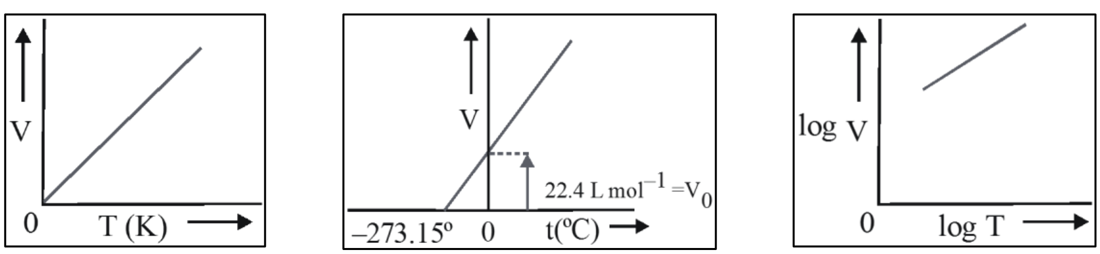
$\therefore$ From the above equation we get $\frac{V_t}{V_0} = \frac{T_t}{T_0}$

or $$\frac{V_t}{T_t} = \frac{V_0}{T_0}$$

i.e. The volume of a given gas is proportional to the absolute temperature.

$$\frac{V_1}{T_1} = \frac{V_2}{T_2} \text{ (at constant P)}$$

◆ **Graphical Representation**



Graphs between V and T at constant pressure are called Isobars.

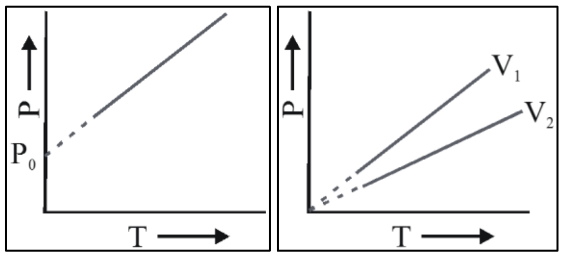
3. Gay-Lussac's Law : Dependence of Pressure on Temperature

It states "at constant volume, the pressure of a given mass of a gas is directly proportional to the absolute temperature of the gas".

$P \propto T$ or $P = KT$

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

◆ **Graphical Representation**



The plots drawn at constant volume for a gas is called as Isochore.

◆ **The Combined Gas Law**

It states "for a fixed mass of gas, the volume is directly proportional to absolute temperature and inversely proportional to the pressure".

Boyle's Law, $V \propto \frac{1}{P}$ (at constant n, T)

Charle's Law, $V \propto T$ (at constant n, P)

Therefore, $V \propto \frac{T}{P}$

or $V = K \frac{T}{P}$

or $\frac{PV}{T} = K$

or $\frac{P_1 V_1}{T_1} = \frac{P_2 V_2}{T_2}$

◆ **Volume Coefficient (αv) of a Gas**

The ratio of increase in volume of a gas at constant pressure per degree rise of temperature to its volume at 0ºC is the volume coefficient of the gas.

$\alpha_V = \frac{V_t - V_0}{V_0 \times t}$ or $V_t = V_0 (1 + \alpha_V t)$

For all gases, $\alpha_V = \frac{1}{273}$

- **Pressure Coefficient (αp) of a Gas**

The ratio of increase in pressure of the gas at constant volume per degree rise of temperature to its pressure at 0ºC is the pressure coefficient of the gas.

$$\alpha_p = \frac{P_t - P_0}{P_0 \times t} \qquad \text{or} \qquad P_t = P_0 (1 + \alpha_p t)$$

For all gases, $\alpha_p = \frac{1}{273}$

4. Avogadro's Law

It states "equal volumes of any two gases at the same temperature and pressure contain the same number of molecules".

$V \propto n$ (At constant P and T)

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

Ex. The temperature of a certain mass of a gas is doubled. If the initially the gas is at 1 atm pressure. Find the % increase in pressure?

Sol. $\frac{P_1}{T_1} = \frac{P_2}{T_2}; \frac{1}{T} = \frac{P_2}{2\ T}$

% increase $= \frac{2-1}{1}$ x $100 = 100\%$

Ex. The temperature of a certain mass of a gas was increased from 27°C to 37°C at constant volume. What will be the pressure of the gas.

Sol. $\frac{P_1}{T_1} = \frac{P_2}{T_2}; \frac{P}{300} = \frac{P_2}{310}; P_2 = \frac{31}{30} P$