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STATES OF MATTER

Intermolecular Forces

Intermolecular forces are the forces of attraction and repulsion between interacting particles (atoms and molecules). This term does not include the electrostatic forces that exist between the two oppositely charged ions and the forces that hold atoms of a molecule together i.e., covalent bonds.

Attractive intermolecular forces are known as van der Waals forces.

However, the attractive forces between an ion and a dipole (i.e. ion-dipole forces) are not van der Waals forces.

van der Waals forces

It is one of the weakest physical forces of nature which operates between non-bonded atoms and molecules. It arises due to following three types of interactions-

(1) Dipole - Dipole interaction
(2) Dipole – Induced dipole interaction
(3) Induced dipole – Induced dipole interaction (London Dispersion)

(1) Dipole - Dipole Interactions

Dipole-dipole interactions operates between polar molecules (i.e. between the molecules possessing permanent dipole such as HCl, NH_3, H_2O etc.). Here, positively charged pole of one molecule attracts the negatively charged pole of another molecule.

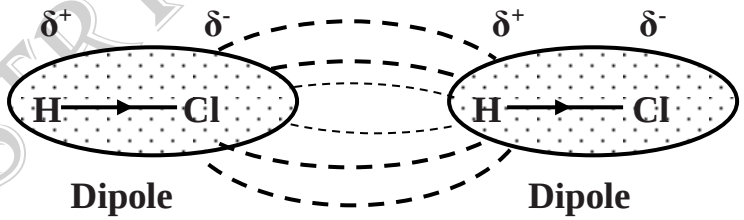


Fig- Dipole-dipole interaction between two HCl molecules

Here, the attractive force decreases with the increase of distance between the dipoles.

Dipole-dipole interaction energy between stationary polar molecules (as in solids) is proportional to $1/r^3$ and that between rotating polar molecules is proportional to $1/r^6$, where r is the distance between polar molecules.

- This interaction is stronger than the London forces but is weaker than ion-ion interaction as it involves only partial charges.

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- Besides dipole-dipole interaction, polar molecules can also interact by London forces. Thus, due to cumulative effect this type of intermolecular forces is strongest amongst three.
- Hydrogen bonding is a special case of dipole-dipole interaction between highly polar molecules.

2. Dipole – Induced Dipole Interactions

This type of attractive forces operate between the polar molecules having permanent dipole and the non-polar molecules lacking permanent dipole. For example, between HCl and O_2 molecules.

Here, permanent dipole of the polar molecule induces dipole in non-polar molecule by deforming its electron cloud and therefore attraction starts operating between them.

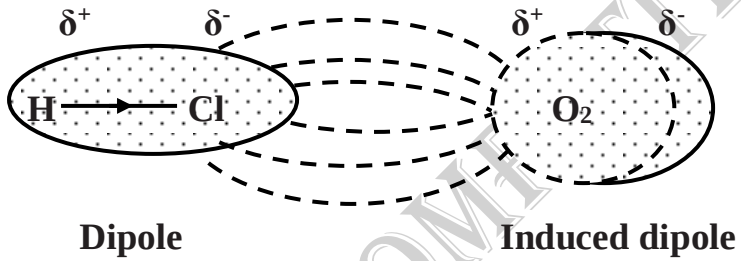


Fig.- Dipole - Induced dipole interaction

Here, the interaction energy is proportional to $1/r^6$, where r is the distance between two molecules.

Induced dipole moment depends upon the dipole moment present in the permanent dipole and the polarizability of the non-polar molecule.

- Here, in fact, a cumulative effect of dispersion forces (London dispersion) and dipole-induced dipole interactions operates.

(3) Induced Dipole–Induced Dipole Interactions (London Dispersion Forces or London Forces)

It operates between non-bonded atoms (e.g. between atoms of Inert gases) and non-polar molecules (e.g. O_2, N_2, CH_4 etc.), where the electron clouds are symmetrically distributed and therefore they have no dipole moment. However, in these atoms and molecules a dipole may develop momentarily due to disturbance of their electron clouds either by heating or jerk.

Suppose two atoms He are in the close vicinity of each other. Let, the electronic charge distribution of one atom becomes unsymmetrical momentarily *i.e.*, the charge cloud becomes more on one side than the other. This results in the

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development of instantaneous dipole in the atom for a very short time. This instantaneous or transient dipole distorts the electron density of the other He atom and therefore, a dipole also gets induced in this atom. The temporary dipoles of these atoms start attracting each other.

Similarly, temporary dipoles get induced in non-polar molecules and they attract each other.

This force of attraction between two temporary dipoles was first proposed by Fritz London, and therefore it is termed as London force or London dispersion.

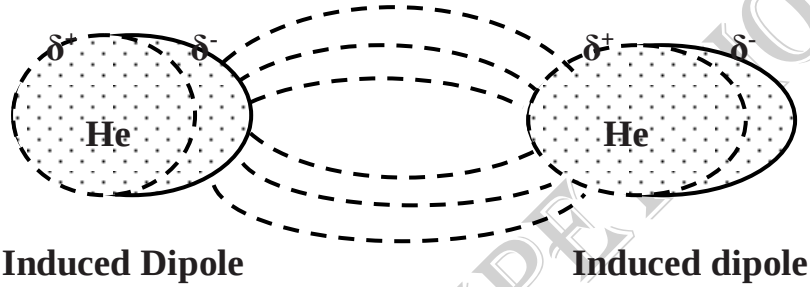


Fig.- Dispersion forces or London forces between He atoms

These forces are always attractive and interaction energy is inversely proportional to the sixth power of the distance between two interacting particles (i.e., $1/r^6$ where r is the distance between two particles). These forces are important only at short distances (~500 pm) and their magnitude depends on the polarisability of the particle.

In addition to attractive forces, molecules also exert repulsive forces on one another. When two molecules are brought into close contact with each other, then a repulsion starts operating between their electron clouds and their nuclei. The magnitude of this repulsion rises rapidly with decrease of the distance separating the molecules. This is the reason that solids and liquids are hard to compress. In fact, in solids and liquids, the molecules are already in close contact and therefore they resist further compression as it would result in the increase of repulsive interactions.

Thermal Energy

It is the energy of a body arising from motion of its atoms or molecules. It is directly proportional to the temperature of the substance. It is the measure of average kinetic energy of the particles of the matter and is thus responsible for movement of particles. This movement of particles is called thermal motion.

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Itermolecular Forces *vs* Thermal Interactions

We know that, intermolecular forces tend to keep the molecules together but thermal energy of the molecules tends to keep them apart. In fact, the three states of matter are the result of balance between intermolecular forces and the thermal energy of the molecules.

When molecular interactions are very weak, molecules do not cling together to make liquid or solid unless thermal energy is reduced by lowering the temperature. Gases do not liquify on compression only, although molecules come very close to each other and intermolecular forces operate to the maximum. However, when thermal energy of molecules is reduced by lowering the temperature, the gases can be liquified very easily.

Gaseous Laws

The laws which govern the behavior of gases are called gaseous laws. The important gaseous laws are as under –

(1) Boyles Law (Pressure-Volume Relationship)

This law governs the effect of pressure on the volume of a definite amount of a gas at a given temperature. According to this law-

"At constant temperature, the volume of a definite amount of a gas is inversely proportional to the pressure applied on the gas."

Mathematically,

$$V \quad \alpha \quad \frac{1}{P} \qquad (\text{At constant T})$$

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

or $$\boxed{PV = K} \qquad \text{---------- (1)}$$

Where, K is a proportionality constant called Boyle's constant

This is mathematical form of Boyle's law.

Let, at a given temperature, the volume of a definite amount of a gas is V_1 at pressure P_1 and is V_2 at pressure P_2.

Then, following Boyle's law,

$$\boxed{P_1V_1 = P_2V_2 = K} \qquad \text{---------- (2)}$$

Thus, "At constant temperature, the product of the pressure and volume of a definite amount of a gas is always constant, no matter how large the pressure is.

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Graphical Representation of Boyle's Law

(1) If volume of a definite amount of a gas is plotted against pressure applied on the gas at constant temperature, a rectangular hyperbola curve is obtained.

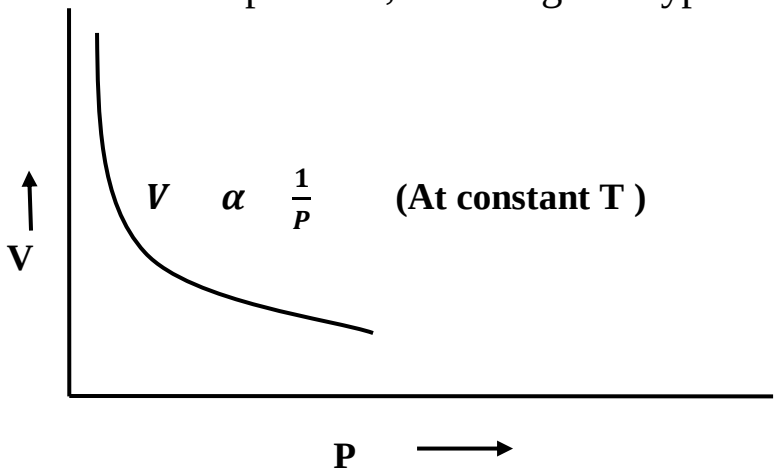


Fig. – Plot of V against P at constant T (Isotherm)

From the curve it is clear that, at constant temperature, the volume of a definite amount of a gas goes on decreasing with increase of pressure applied on the gas.

(2) If volume of a definite amount of a gas is plotted against reciprocal of pressure applied on the gas (1/P) at constant temperature, then a straight line passing through origin is obtained.

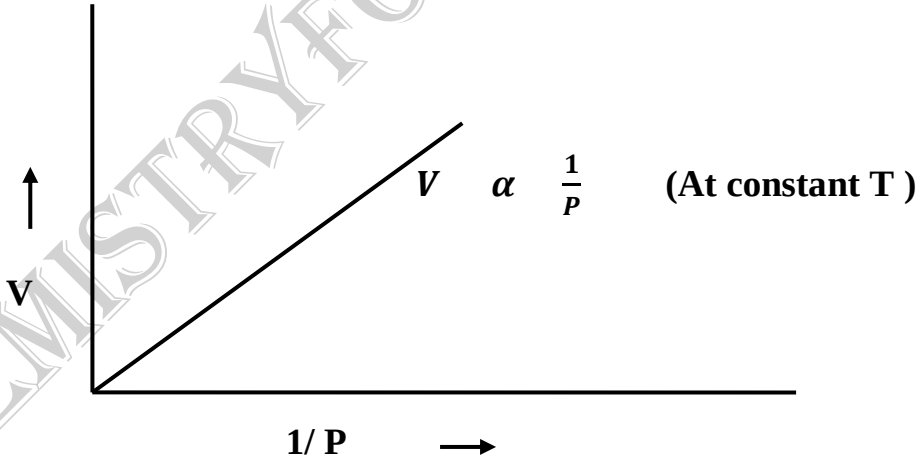


Fig. – Plot of V against 1/P at constant T

From the plot it is clear that, at constant temperature, the volume of a definite amount of a gas goes on increasing with increase of 1/P.

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(3) If product of pressure and volume (PV) of a definite amount of a gas is plotted against pressure applied on the gas (P) at constant temperature, then a straight line parallel to pressure axis is obtained.

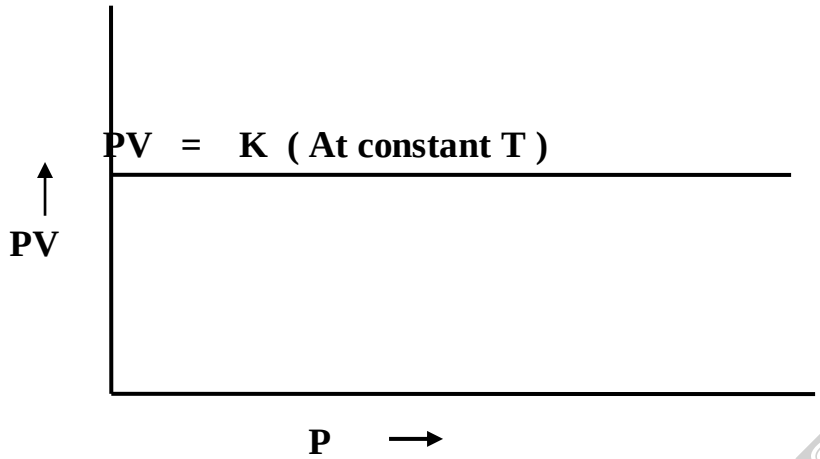


Fig. – Plot of PV against P at constant T

From the plot it is clear that, the product of pressure and volume (PV) of a definite amount of a gas is always constant at constant temperature, no matter how large the pressure is.

Numericals

Q.1. A balloon is filled with hydrogen at room temperature. It will burst if pressure exceeds 0.2 bar. If at 1 bar pressure the gas occupies 2.27 L volume, up to what volume can the balloon be expanded ?

Solution-

According to Boyle's Law, $P_1V_1 = P_2V_2$

If P_1 = 1bar, V_1 = 2.27 L
If P_2 = 0.2 bar, V_2 = ?

$\therefore$ 1 bar × 2.27 L = 0.2 bar × V_2

V_2 = 11.35 L

Since, the balloon would bursts at 0.2 bar pressure, the volume of balloon should be less than 11.35 L.

(2) Charle's Law (Temperature-Volume Relationship)

This law governs the effect of temperature on the volume of a definite amount of a gas at a given pressure. According to this law-

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"At constant pressure, if temperature of a definite amount of a gas is increased or decreased by 1^0C , then its volume gets increased or decreased by 1/273rd part of its volume at 0^0C temperature."

Let, at 0^0C temperature, the volume of a definite amount of a gas is V_0, then following Charle's law its volume at 1^0C temperature,

$$V_1{}^0C = V_0 + \frac{1 \times V_0}{273} \quad (\text{at constant P})$$

Similarly, at t^0C temperature,

$$V_t{}^0C = V_0 + \frac{t \times V_0}{273} \quad (\text{at constant P})$$

$$V_t{}^0C = V_0\left(1 + \frac{t}{273}\right) \quad (\text{at constant P}) \quad \text{---------- (1)}$$

This is mathematical form of Charle's law.

From equation (1),

$$V_t{}^0C = V_0\left(\frac{273 + t}{273}\right) \quad (\text{at constant P})$$

$\because$ $(273 + t)^0C = T$ (Where T = Absolute temperature)

$\therefore$ $V_t{}^0C = V_0\left(\frac{T}{273}\right)$ (at constant P)

or $V = KT$ (at constant P) -------- (2)

$\therefore$ $V \alpha T$ (At constant P)

Thus, "At constant pressure, the volume of a definite amount of a gas is directly proportional to the absolute temperature."

This is simplified form of Charle's law.

From eq. (2), $\frac{V}{T} = K$

$$\therefore \frac{V_1}{T_1} = \frac{V_2}{T_2} = K \quad (\text{at constant P}) \quad \text{---------- (3)}$$

Thus, from Charle's law, "The ratio of the volume of a definite amount of a gas and absolute temperature is always constant at a given pressure, no matter how large the temperature is."

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Graphical Representation of Charle's Law

If volume of a definite amount of a gas is plotted against absolute temperature, at constant pressure, then a straight line passing through origin is obtained.

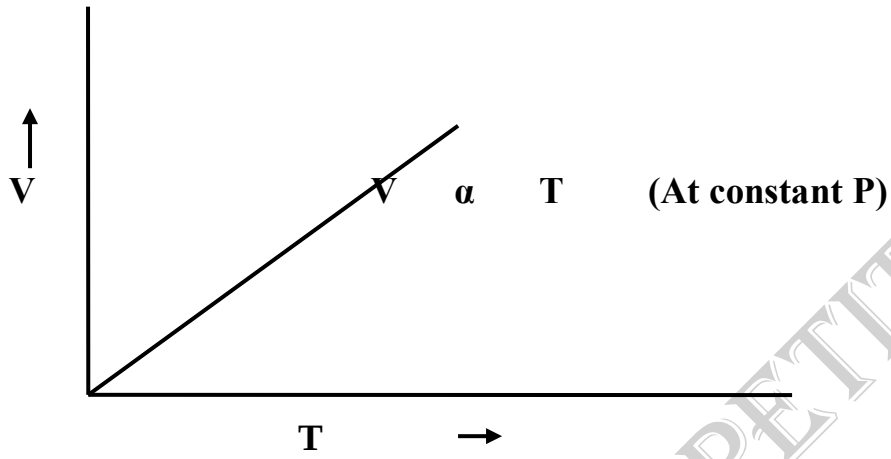


Fig. – Plot of V against T at constant P (Isobar)

From the plot it is clear that, the volume of a definite amount of a gas goes on increasing with increase of absolute temperature at constant pressure.

Absolute Zero Temperature

From Charle's law,

$$V_t{}^0C = V_0\left(\frac{T}{273}\right) \text{ (at constant P)}$$

If t = -273^0C,

Then,

$$V_{-273}{}^0C = V_0\left(1+\frac{-273}{273}\right) \text{ (at constant P)}$$

or $\boxed{V_{-273}{}^0C = 0}$

Thus, at -273^0C temperature, the volume of all the gases theoretically becomes zero. Hence, this temperature is called Absolute zero temperature.

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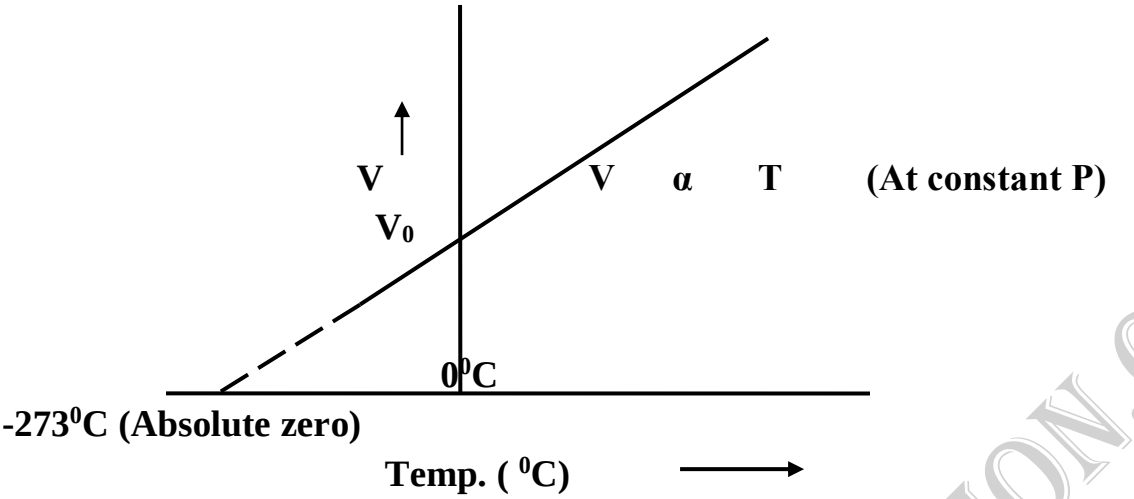


Fig.– Plot of Volume against Temp. (^{0}C) at constant Pressure

From the plot it is clear that, the volume of a definite amount of a gas goes on decreasing with decrease of temperature at constant pressure and at -273^{0}C temperature (Absolute zero temperature), it theoretically becomes zero .

However, in practice, much more before this temperature, all the real gases either get liquefied or solidified and therefore show deviation from Charle's law.

Numericals

Q.1. On a ship sailing in pacific ocean where temperature is 23.4 °C , a balloon is filled with 2 L air. What will be the volume of the balloon when the ship reaches Indian ocean, where temperature is 26.1°C ?

Solution-

Given,

$V_1 = 2$ L $\quad T_1 = (23.4 + 273)$ K $= 296.4$ K

$V_2 =$? $\quad T2 = (26.1 + 273)$ K $= 299.1$ K

From Charles law,

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

$$\therefore \quad \frac{2}{299.1} = \frac{V_2}{296.4}$$

or $\quad V_2 = 2.018$L

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(3) Avogadro Law (Volume – Amount Relationship)

According to this law,"Under same conditions of temperature and pressure, equal volumes of all gases contain equal number of molecules.

This means that, as long as the temperature and pressure remain constant, the volume of a gas depends upon number of molecules of the gas i.e. amount of the gas.

Mathematically,

$V \alpha n$ Where *n* is the number of moles of the gas

or $V = k\,n$ ---------- (1)

The number of molecules in one mole of a gas has been determined to be 6.022×10^{23} and is known as Avogadro constant.

Since volume of a gas is directly proportional to the number of moles, one mole of all the gases will have same volume at STP (standard temperature and pressure).

∵ Number of moles of a gas,

$$n = \frac{m}{M}$$

Where, *m* = mass of the gas under investigation and M = molar mass of the gas

Thus,

$$V = \frac{k\ m}{M} \quad \text{---------- (2)}$$

From eq. (1) and (2) we get,

$$M = \frac{k\ m}{V}$$

or $M = k\,.\,d$ ---------- (3)

Where '*d*' is the density of the gas.

Thus, the density of a gas is directly proportional to its molar mass.

- Standard temperature and pressure means 273.15 K (0°C) temperature and 1 bar (i.e. exactly 105 pascal) pressure. These values approximate freezing temperature of water and atmospheric pressure at sea level.
- At STP molar volume of an ideal gas or a combination of ideal gases is 22.71098 L.
- Molar volume of some real gases in litres per mole at 273.15 K temperature and 1 bar pressure (STP) is as under-

Argon	22.37
Carbon dioxide	22.54
Dinitrogen	22.69
Dioxygen	22.69
Dihydrogen	22.72
Ideal gas	22.71

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- The previous standard is still often used, and applies to all chemistry data more than decade old. In this definition STP denotes the same temperature of 0°C (273.15 K), but a slightly higher pressure of 1 atm (101.325 kPa). One mole of any gas of a combination of gases occupies 22.413996 L of volume at STP.
- Standard ambient temperature and pressure (SATP), conditions are also used in some scientific works. SATP conditions means 298.15 K and 1 bar (i.e., exactly 105 Pa). At SATP (1 bar and 298.15 K), the molar volume of an ideal gas is 24.789 $Lmol^{-1}$.

(4) Gay Lussac's Law (Pressure - Temperature Relationship)

This law gives the mathematical relationship between pressure and Temperature. According to this law,

"At constant volume, pressure of a fixed amount of a gas varies directly with the temperature."

Mathematically,

$$P \quad \alpha \quad T$$

$$P \quad = \quad k\,T \qquad \text{-------} \quad (1)$$

This is mathematical form of Gay Lussac's law.

If pressure of a definite amount of a gas is plotted against absolute temperature at constant volume (Isochore) then a straight line passing through origin is obtained.

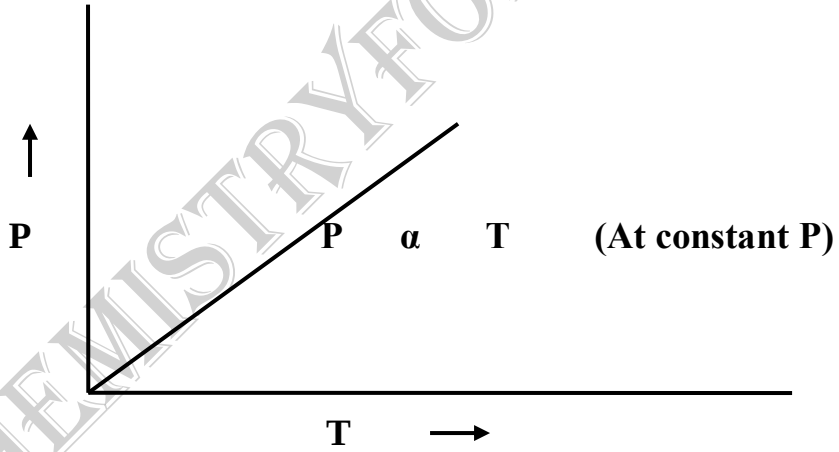


Fig. – Plot of P against T at constant V (Isochore)

- Pressure in well inflated tyres of automobiles is almost constant, but on a hot summer day this increases considerably and tyre may burst if pressure is not adjusted properly.
- During winters, on a cold morning one may find the pressure in the tyres of a vehicle decreased considerably.

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Ideal Gas

A gas that follows Boyle's law, Charles' law and Avogadro law strictly is called an ideal gas. Such a gas is hypothetical. It is assumed that, intermolecular forces are not present between the molecules of an ideal gas.

Real gases follow these laws only under certain specific conditions when forces of interaction are practically negligible. In all other situations, they deviate from ideal behaviour.

Ideal Gas Equaeion

In order to give combined effect of change of temperature and pressure on the volume of a definite amount of an ideal gas an equation has been derived by combining Boyle's, Charles' and Avogadro's laws, which is called Ideal gas equation.

From Boyle's law,

$$V \quad \alpha \quad \frac{1}{P} \qquad \text{(At constant T)} \qquad \text{--------} \quad (1)$$

From Charles's law,

$$V \quad \alpha \quad T \qquad \text{(At constant P)} \qquad \text{---------} \quad (2)$$

From Avogadro's law,

$$V \quad \alpha \quad n \qquad \text{(At constant P and T)} \qquad \text{---------} \quad (3)$$

Combining these three laws, we get-

$$V \quad \alpha \quad \frac{nT}{P} \qquad \text{(at constant P and T)}$$

or $$V = \frac{nRT}{P}$$

or $$PV = nRT \qquad \text{----------} \quad (4)$$

Where, R is a proportionality constant called gas constant. It is same for all gases and therefore it is also called Universal Gas Constant.

Equation (4) is called ideal gas equation. If three variables in this equation are known, then fourth can be calculated using this equation.

This equation is applicable to any gas, under the conditions when behaviour of the gas approaches ideal behaviour.

Ideal gas equation describes the state of an ideal gas and therefore, it is also called equation of state.

If temperature, volume and pressure of a fixed amount of gas changes from T_1, V_1 and P_1 to T_2, V_2 and P_2 then,

In initial state

$$P_1V_1 = nRT_1$$

or $$\frac{P_1V_1}{T_1} = nR \qquad \text{--------} \quad (5)$$

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Similarly, in final state

$$P_2V_2 = nRT_2$$

$$\text{or} \quad \frac{P_2V_2}{T_{21}} = nR \qquad \text{-------- (6)}$$

From eq. (5) and (6) we get,

$$\frac{P_1V_1}{T_1} = \frac{P_2V_2}{T_2}$$

This equation is known as Combined gas law. If out of six, values of five variables are known, the value of unknown variable can be calculated from the equation.

Numericals

Q.1. At 25°C and 760 mm of Hg pressure a gas occupies 600 mL volume. What will be its pressure at a height where temperature is 10°C and volume of the gas is 640 mL.

Solution- Given that,

$P_1 = 760$ mm Hg, $P_2 = ?$

$V_1 = 600$ mL $V_2 = 640$ mL

$T_1 = 25 + 273 = 298$ K $T_2 = 10 + 273 = 283$K

According to Combined gas law,

$$\frac{P_1V_1}{T_1} = \frac{P_2V_2}{T_2}$$

$$\therefore \quad \frac{760 \times 600}{298} = \frac{P_2 \times 640}{283}$$

$$\text{or} \quad P_2 = 676.6 \text{ mm Hg}$$

Relation Between Density and Molar Mass of a Gaseous Substance-

From Ideal gas equation,

$$PV = nRT$$

$$Here, \quad n = \frac{m}{M}$$

$$\therefore \quad PV = \frac{m}{M}RT$$

$$\text{or} \quad \frac{m}{V} = \frac{P \times M}{RT}$$

$$\therefore \text{ Density, } d = \frac{P \times M}{RT}$$

This expression gives the relation between density and molar mass of a gaseous substance.

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On rearranging this equation we get,

$$M = \frac{dRT}{P}$$

Using this expression the molar mass of a gas can be calculated.

(5) Dalton's Law of Partial Pressures

According to this law, "the total pressure exerted by a mixture of non-reactive gases is equal to the sum of the partial pressures exerted by the individual gases".

Mathematically,

$$P_{Total} = P_1+P_2+P_3+..... \text{ (at constant T, V) --------} \quad (1)$$

Where, P_{Total} is the total pressure exerted by the mixture of gases and P_1, P_2, P_3 +.....are partial pressures of individual gases.

In a mixture of gases, the pressure exerted by the individual gas is called partial pressure. It is the pressures which a gas would exert, if it is enclosed separately in the vessel of same volume under the same conditions of temperature.

The partial pressure of a gas A is related to its mole fraction X_A as per the following equation-

$$P_A = X_A \times P_{total}$$

Where, P_{total} = Total pressure of gaseous mixture

Gases are generally collected over water and therefore, are moist. Pressure of dry gas is calculated by subtracting vapour pressure of water from the total pressure of the moist gas. Pressure exerted by saturated water vapour is called aqueous tension.

Here, $P_{Dry\ gas} = P_{Total}$ – Aqueous tension

Numerical-

Q. A neon-dioxygen mixture contains 70.6 g dioxygen and 167.5 g neon. If pressure of the mixture of gases in the cylinder is 25 bar. What is the partial pressure of dioxygen and neon in the mixture ?

Solution- Number of moles of dioxygen,

$$n_{O_2} = \frac{70.6\ g}{32\ g}$$

$$= 2.21 \text{ moles}$$

Number of moles of neon,

$$n_{Ne} = \frac{167.5\ g}{20\ g}$$

$$= 8.375 \text{ moles}$$

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Mole fraction of dioxygen,

$$X_{O_2} = \frac{2.21}{2.21+8.375} = 0.21$$

Mole fraction of neon,

$$X_{Ne} = \frac{8.375}{2.21+8.375} = 0.79$$

Partial pressure = mole fraction of a gas × total pressure

$\therefore$ Partial pressure of O_2 = 0.21 × 25 bar
= 5.25 bar

and Partial pressure of Ne = 0.79 × 25 bar
= 19.75 bar

Kinetic Molecular Theory of Gases

It is microscopic model of gases and has been proposed to explain the behavior of gases. This theory is the result of combined effort of Maxwell, Boltzman, Clausious, Claperon etc. The main postulates of this theory are as under-

(1) All the gases are consisting of very minute, individual and invisible spherical particles called molecules.

(2) The inter-molecular space between these molecules at ordinary temperature is very large. Hence, they are highly compressible at ordinary temperature.

(3) The gaseous molecules are highly rigid and incompressible in nature.

(4) These molecules are always in state of random motion in all the possible directions. Hence, they constantly collide to each other as well as to the wall of the container.

(5) The collision of gaseous molecules to the wall of the container is responsible for gaseous pressure.

(6) The collisions of gaseous molecules with each other as well as to the wall of the container is perfectly elastic i.e. there occurs no energy loss during these collisions.

(7) Gaseous molecules are mere massive particles having negligible volume. Hence, the actual volume of the gaseous molecules is negligible as compared to the volume occupied by the gas.

(8) There operates no force of attraction or repulsion between the gaseous molecules.

(9) The average kinetic energy of gaseous molecules depends only upon absolute temperature and is directly proportional to the absolute temperature.

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Deviation of Real Gases from Ideal Behaviour

Real gases almost obey all the gaseous laws and follow ideal gas equation (PV = nRT) at high temperature and low pressure but show deviation at low temperature and high pressure.

For example, if PV of real gases is plotted against P at constant temperature T (Boyle's law), then instead of a straight line parallel to pressure axis, following two types of graphs are obtained-

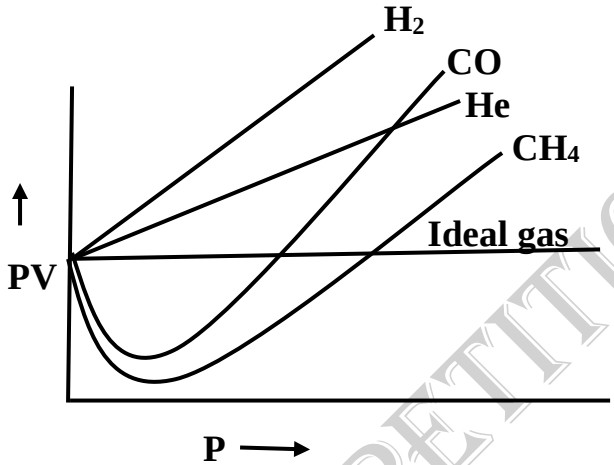


Fig.- Plot of PV vs P for real gas and ideal gas

(1) For hydrogen and helium, the value of PV increases with increase of the pressure.Thus, these gases show positive deviation from ideal behavior at all the pressures.

(2) For gases like carbon monoxide and methane the value of PV first decreases with increase of the pressure and reaches to a minimum value (which is the characteristic of the respective gas) i.e. they first show negative deviation from ideal behavior. After this the value of PV starts increasing and crosses the line for ideal gas. Thus, the gase start showing positive deviation from ideal behaviour.

Similarly, according to Charle's law at -273^0C temperature (Absolute zero temperature), the volume of all the real gases should become theoratically zero. However, in practice, much more before this temperature, all the real gases either get liquefied or solidified and therefore show deviation from Charle's law.

Thus, real gases do not follow Boyle's law, Charles law and Avogadro's law perfectly under all conditions of temperature and pressure and therefore, do not follow ideal gas equation perfectly under all conditions.

Compressibility Factor (Z)

The deviation of real gases from ideal behaviour is measured in terms of compressibility factor Z. By definition,

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"The ratio of product PV and nRT of a gas in a given conditions of temperature and pressure is called compressibility factor, Z of the respective gas."

Mathematically,

$$\text{Compressibility factor, } Z = \frac{PV}{nRT}$$

For ideal gas, the value of compressibility factor, Z is always unity (equal to 1). But for real gases, it may be lesser or greater than 1.

If compressibility factor, Z of different gases are plotted against pressure, P at constant temperature, T then following curves are obtained-

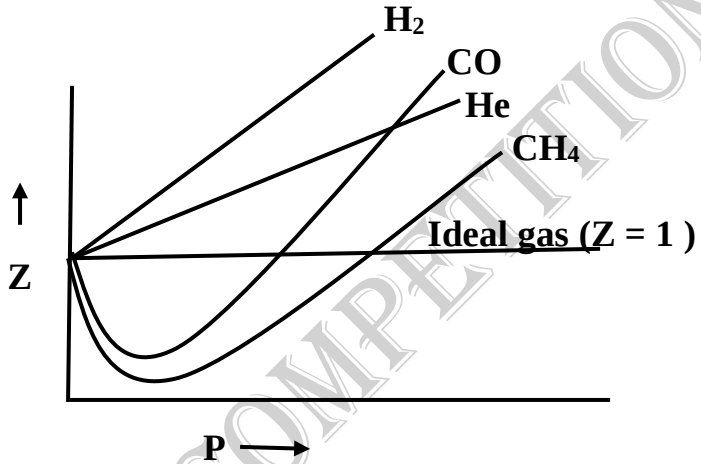


Fig.- Plot of compressibility factor, Z vs P for real gas and ideal gas

From the graph, it is clear that-

(1) For an ideal gas the graph of Z vs P is always a straight line parallel to pressure axis.

(2) For real gases, following two types of graphs are obtained-

(a) For hydrogen and helium, the value of Z increases continuously with increase of the pressure. Thus, for these gases Z is greater than 1 at all the pressures and these gases show positive deviation from ideal behavior at all the pressures.

(2) For gases like N_2, O_2, CH_4, CO_2 etc. the value of Z first decreases with increase of the pressure , reaches to a minimum value (which is the characteristic of the respective gas) and then it starts increasing and crosses the line for ideal gas. Thus, at low pressure, these gases show negative deviation from ideal behavior ($Z<1$) while at high pressure they show positive deviation from ideal behaviour ($Z > 1$).

- At very low pressures all the real gases have $Z = 1$ and behave like ideal gas.
- At very high pressure these gases have $Z > 1$. These are more difficult to compress.

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- At intermediate pressures, for most of the real gases $Z < 1$.
- Since, for a real gas, Compressibility factor,

$$Z = \frac{PV_{real}}{nRT} \quad \text{-------} \quad (1)$$

If gas is behaving like ideal gas, then-

$$Z = 1$$

Hence,
$$V_{ideal} = \frac{nRT}{P} \quad \text{-------} \quad (2)$$

From eq. (1) and (2), we get-

$$Z = \frac{V_{real}}{V_{ideal}}$$

Thus, compressibility factor is the ratio of actual molar volume of a gas to its molar volume when it behaves like ideal gas at same temperature and pressure.

Boyle's Temperature

For all the real gases, there is a definite temperature at which it behaves like ideal gas over an appreciable range of pressure. This temperature is called Boyle temperature or Boyle point.

Boyle point of a gas depends upon its nature. Above Boyle point, real gases show positive deviations from ideality and the value of Z for the gas is greater than one. Here, the force of attraction between the molecules is very feeble.

Below Boyle temperature real gases first show decrease in the value of Z with increase of pressure, reaches to a minimum value and there after it starts increasesing continuously.

Causes of Deviation of Real Gases from Ideal Behaviour

Real gases almost obey all the gaseous laws and follow ideal gas equation (PV = nRT) at high temperature and low pressure when the volume occupied by the gas is very large but show deviation at low temperature and high pressure. The principal cause of this deviation is that, following two assumptions of the kinetic theory of gases do not hold good for real gases at low temperature and high pressure-

(1) There is no force of attraction or repulsion between the molecules of a gas.
(2) The actual volume of the molecules of a gas is negligible as compared to the volume occupied by the gas.

If assumption (1) is correct, the gas will never liquefy, but in practice real gases get liquified when cooled and compressed. Furthermore, the liquids formed are very difficult to compress. This means that, the forces of repulsion are powerful enough and prevent squashing of molecules in tiny volume.

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If assumption (2) is correct, the pressure vs volume graph of experimental data (real gas) and that theoretically calculated from Boyles law (ideal gas) should coincide. But in practice, that is not so.

Real gases show deviations from ideal behaviour at low temperature and high pressure because, in these conditions the gaseous molecules come very close together, so that volume occupied by the gas is very small. Hence, appreciable attraction operates amongst the gaseous molecules and the actual volume of the gaseous molecules cannot be neglected as compared to the volume occupied by the gas.

The real gases show ideal behaviour when the temperature is very high and the pressure approaches zero.

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