Matter
Solid, Liquid and Gas

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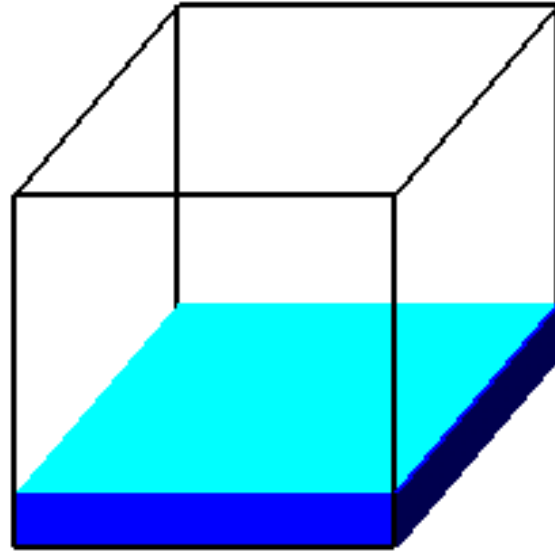**GMSSS-23A Chandigarh**
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States of Matter
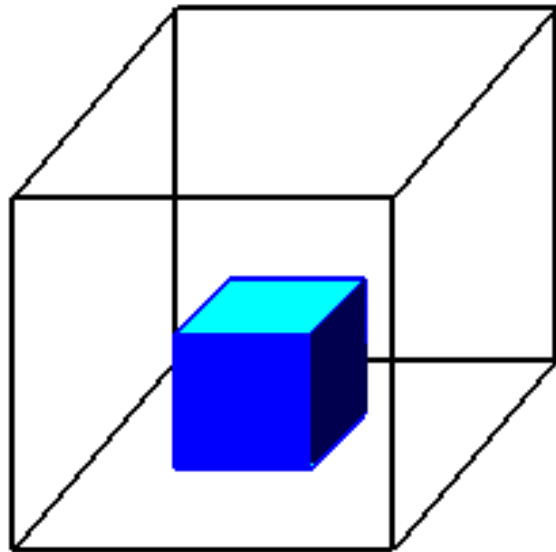
Solid

Holds Shape

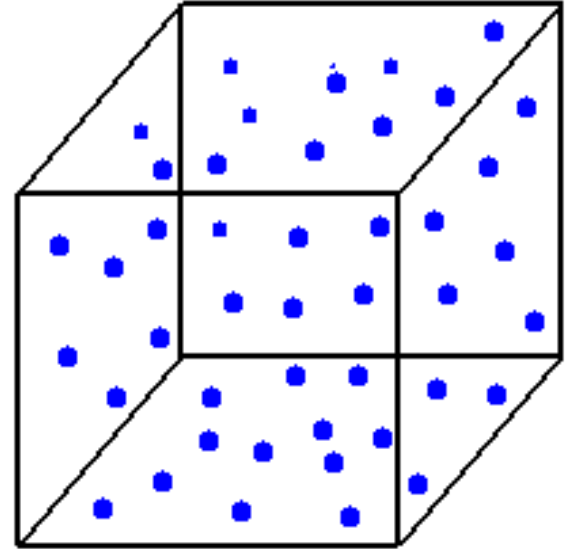Fixed Volume

Liquid

Shape of Container

Free Surface

Fixed Volume

Gas

Shape of Container

Volume of Container

• *INTERMOLECULAR FORCES*

Intramolecular forces are attractive forces between molecules.
They hold atoms together in a molecule.

INTERMOLECULAR Vs INTRAMOLECULAR
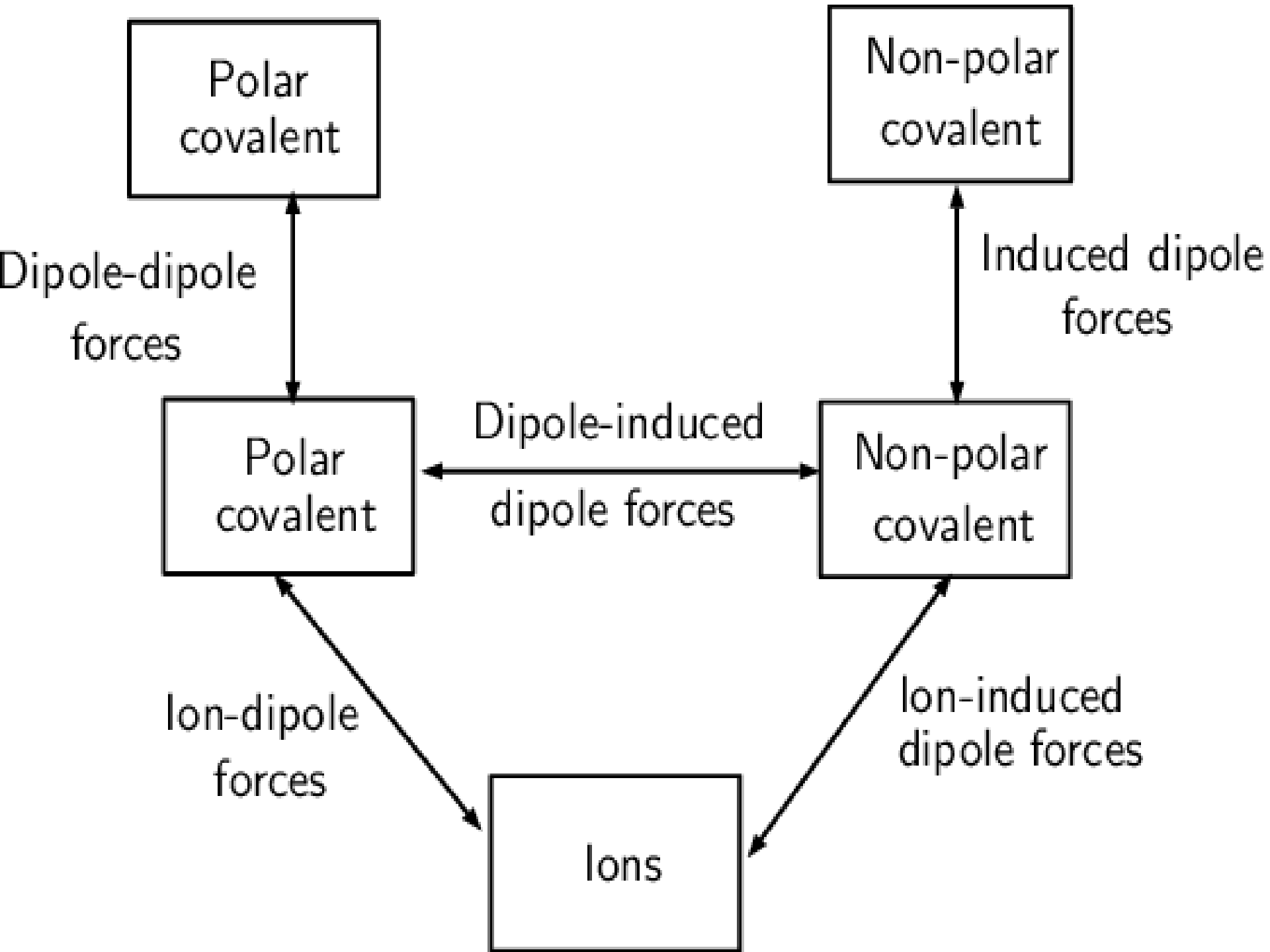- 41 KJ to vaporize 1 mole of water (inter)
- 930 KJ to break all O-H bonds in 1 mole of water (intra)

Measure of intermolecular force
boiling point
melting point
ΔHvap
ΔHsub
ΔHfus

Generally, Intermolecular forces are much weaker than Intramolecular forces.

Nonbonding (Intermolecular)

Ion-dipole		Ion charge– dipole charge	40–600	$Na^+ \cdots O \overset{H}{\underset{H}{\big<}}$
H bond	$\overset{\delta^-}{-A} - \overset{\delta^+}{H} \cdots \overset{\delta^-}{:B} -$	Polar bond to H– dipole charge (high EN of N, O, F)	10–40	$:\overset{..}{O} - H \cdots :\overset{..}{O} - H$ with H below each O
Dipole-dipole		Dipole charges	5–25	$I - Cl \cdots I - Cl$
Ion–induced dipole		Ion charge– polarizable e^- cloud	3–15	$Fe^{2+} \cdots O_2$
Dipole–induced dipole		Dipole charge– polarizable e^- cloud	2–10	$H - Cl \cdots Cl - Cl$
Dispersion (London)		Polarizable e^- clouds	0.05–40	$F - F \cdots F - F$
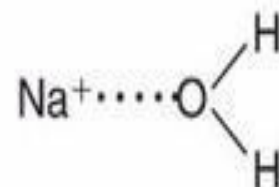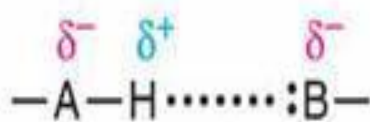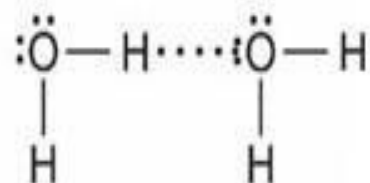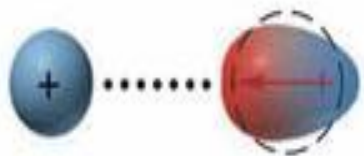
BOYLE'S LAW
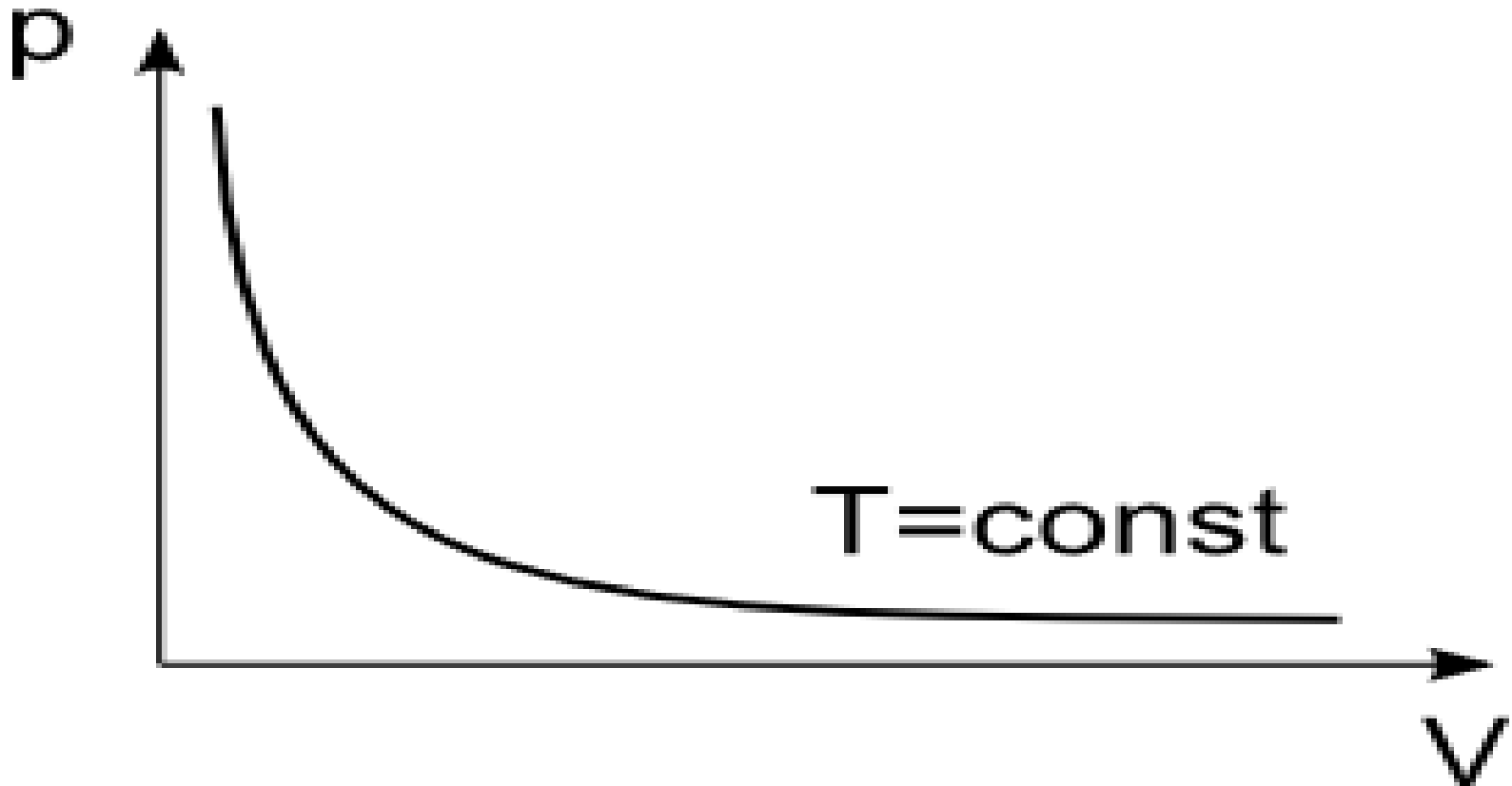
$$PV = \text{constant}$$

CHARLE'S LAW (Temperature – Volume Relationship)

Charle's Law states that pressure remaining constant, the volume of a fixed mass of a gas is directly proportional to it's absolute temperature.

$$\frac{V_1}{T_1} = \frac{V_2}{T_2}$$
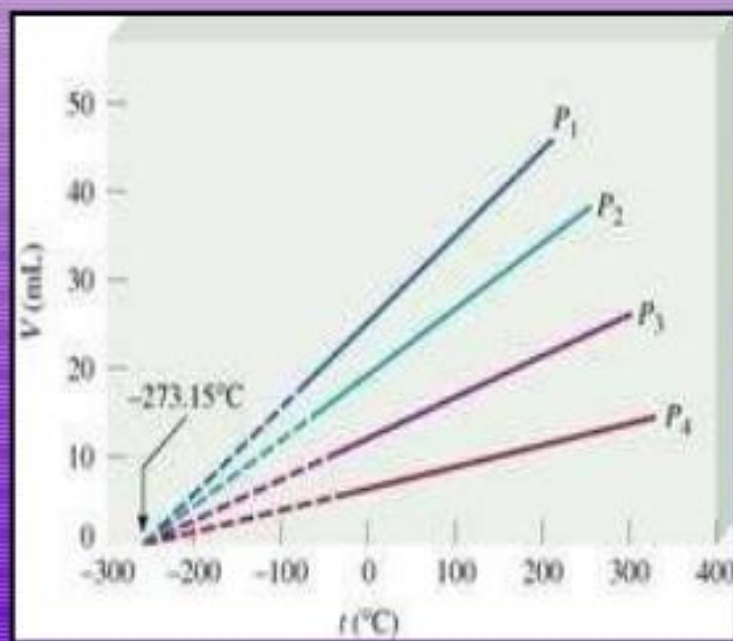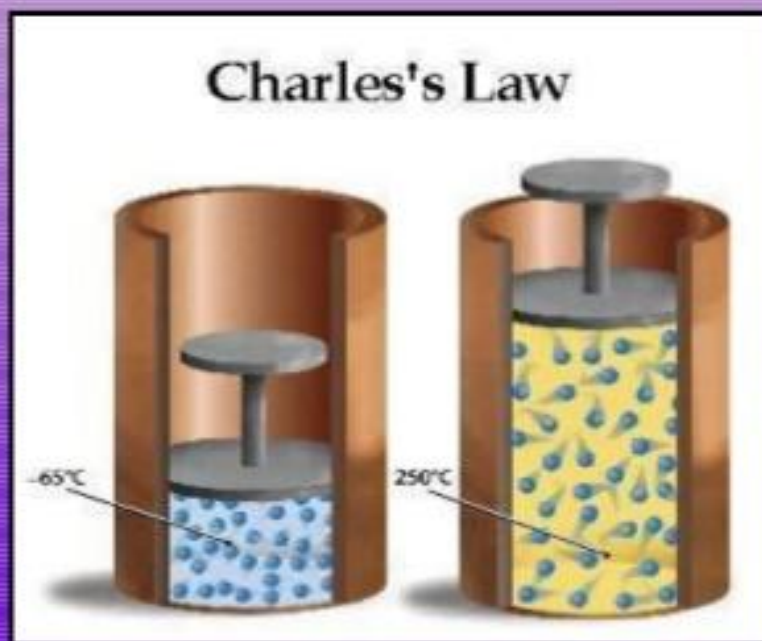
AVAGADRO LAW (Volume – Amount Relationship)

It states that equal volumes of all gases under the same conditions of temperature and pressure contain equal number of molecules.

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

$$\frac{V_1}{n_1} = \frac{V_2}{n_2}$$
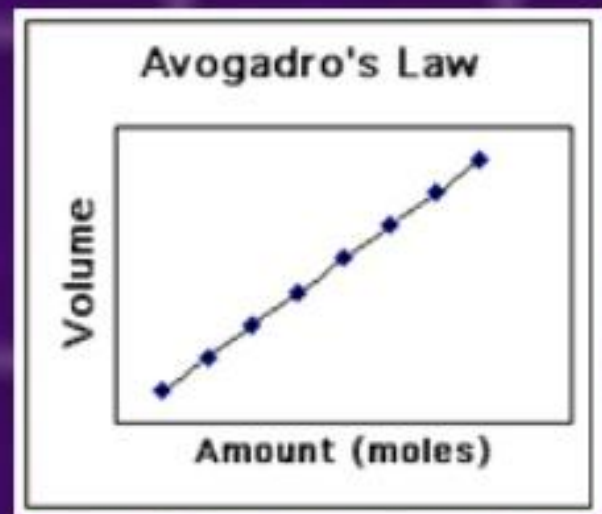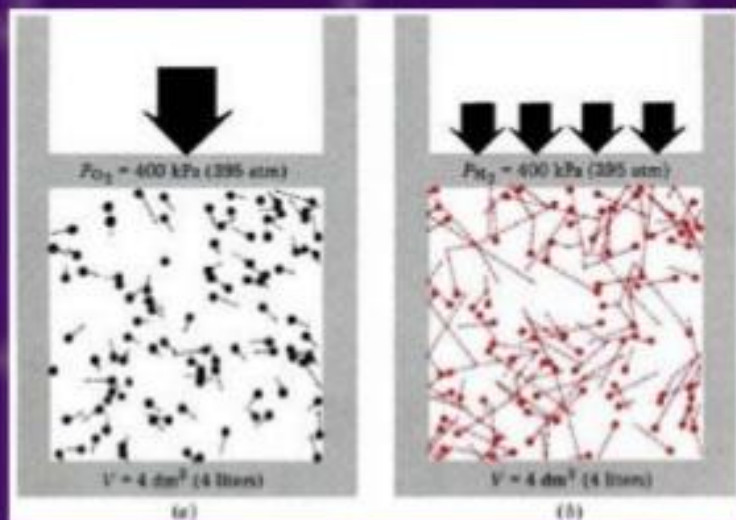
MOLAR VOLUME OF GAS UNDER DIFF. CONDITIONS:

1. When STP is taken as 273K and 1atm the molar volume is 22.4L

2. When STP is taken as 273K and 1bar the molar volume is 22.7L (1 Bar=0.987atm)

3. When SATP conditions are used i.e. 298K and 1bar the molar volume is 24.8L

• DALTON'S LAW OF PARTIAL PRESSURE

This law states that the total pressure exerted by the mixture of non-reactive gases is equal to the sum of the partial pressures of individual gases.

$$P_{total} = P_1 + P_2 + \cdots = \sum_i P_i$$

$$(P_1) = \frac{\text{Number of moles of the gas } (n_1) \times P_{\text{Total}}}{\text{Total number of moles } (n) \text{ in the mixture}} = \text{Mole fraction } (X_1) \times P_{\text{Total}}$$

GAS A   GAS B   GASES A & B

P = P_A + B_B

P = 300 mm Hg   P = 400 mm Hg   P = 300 + 400
                                   =700 mm Hg

O_2 N_2 Ar H_2O CO_2 =   AIR

20,9   78,1   0,97   1,28   0,05   101,3
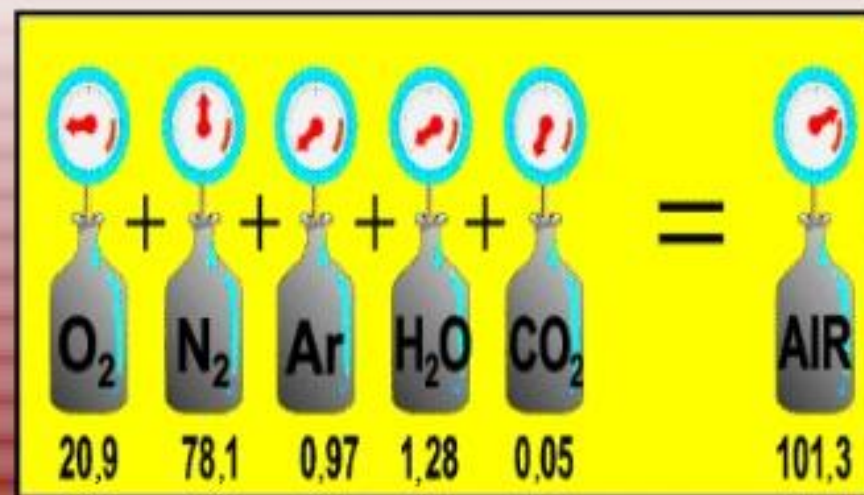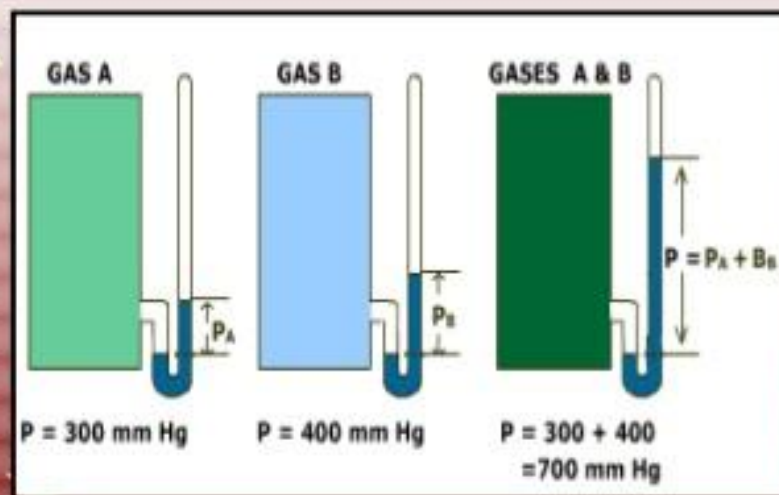
• IDEAL GAS EQUATION

Ideal gas equation is a relation between four variables and it describes the state of any gas, therefore, it is also called equation of state.

Starting with the three gas laws...

$P \propto 1/V$ $V \propto T$ $V \propto n$

Therefore, using the Laws of Proportionality
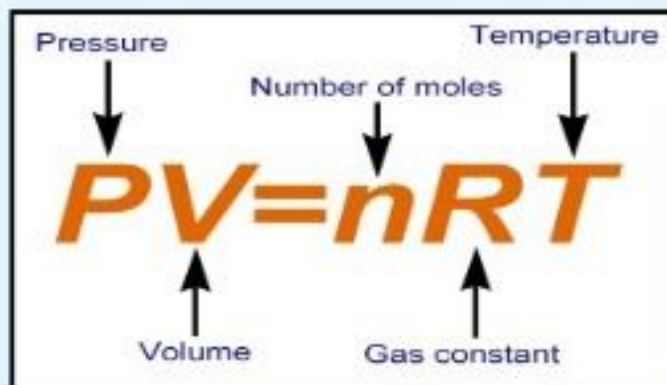
$PV = C_B$ $V/T = C_C$ $V/n = C_A$ (C is some constant)

Using Algebraic Substitution

$PV/nT = R$ (R is combination of 3 constants)

$\therefore PV = nRT$

Pressure   Temperature

Number of moles

$$PV = nRT$$

Volume   Gas constant

$$\frac{p_1 V_1}{T_1} = \frac{p_2 V_2}{T_2}$$

Combined gas law

KINETIC MOLECULAR THEORY ON GASES

The Assumptions of Kinetic Molecular Theory:

- A gas is composed of particles in constant motion.

- The average kinetic energy depends on temperature, the higher the temperature, the higher the kinetic energy and the faster the particles are moving.

- Compared to the space through which they travel, the particles that make up the gas are so small that their volume can be ignored.

- The individual particles are neither attracted to one another nor do they repel one another.

- When particles collide with one another (or the walls of the container) they bounce rather than stick. These collisions are elastic; if one particle gains kinetic energy another loses kinetic energy so that the average remains constant.

Root mean square speed at lower T

Root mean square speed at higher T

Fraction of particles at that speed

Particle speed

$$\frac{N_P(v)}{N}$$

$$\frac{N_P(v)}{N} = 4\pi\left(\frac{m}{2\pi kT}\right)^{\frac{3}{2}} v^2 \, e^{-\frac{KE}{kT}}$$
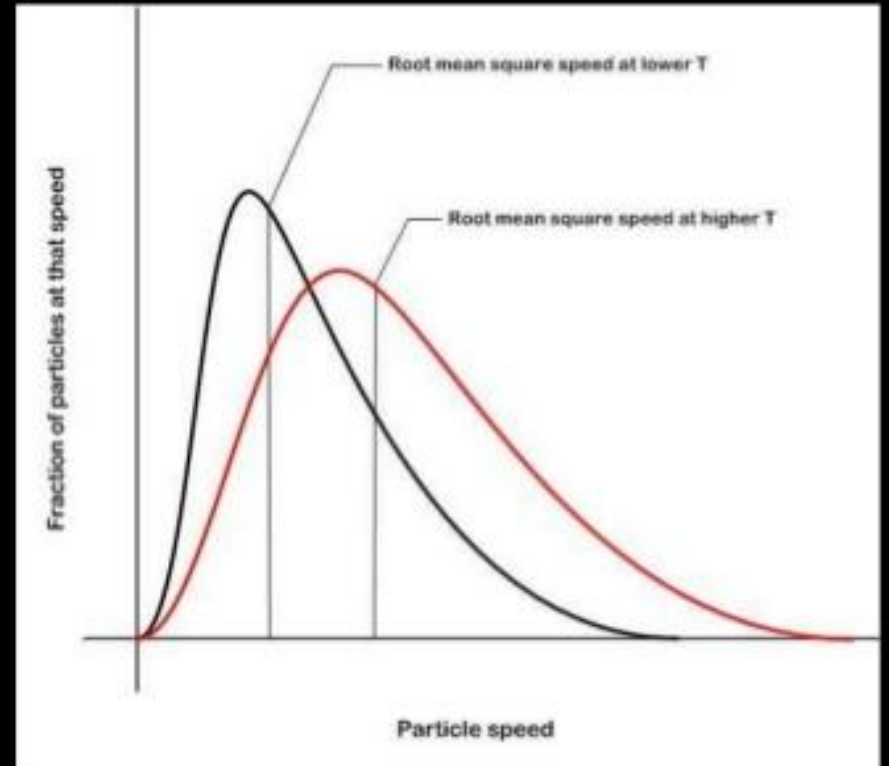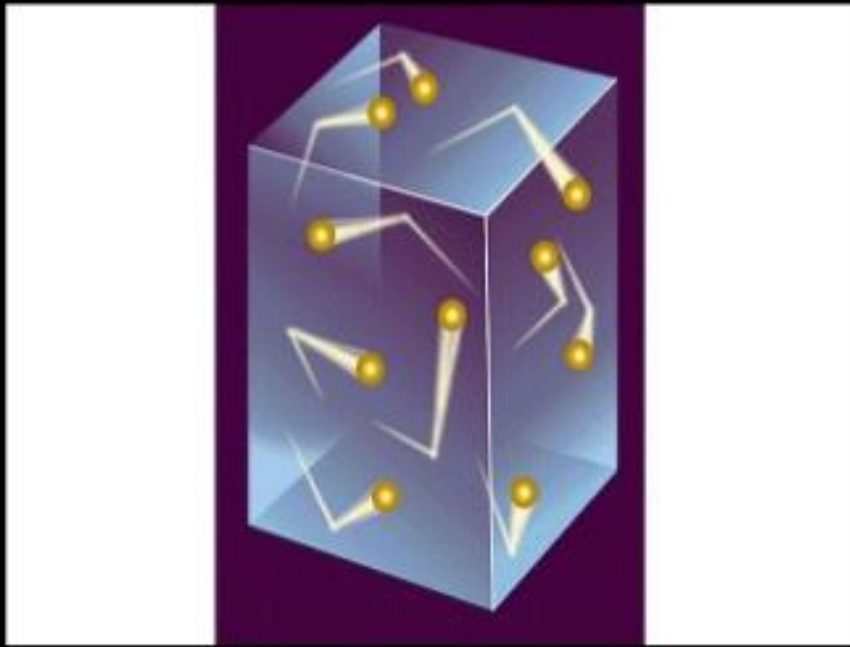
Fraction of Molecules moving at speed v

v_p

v_{av}

v_{rms}

v_p = **Most Probable Speed**

v_{av} = **Average Speed**

v_{rms} = **Root-Mean-Square Speed**

Velocity

v
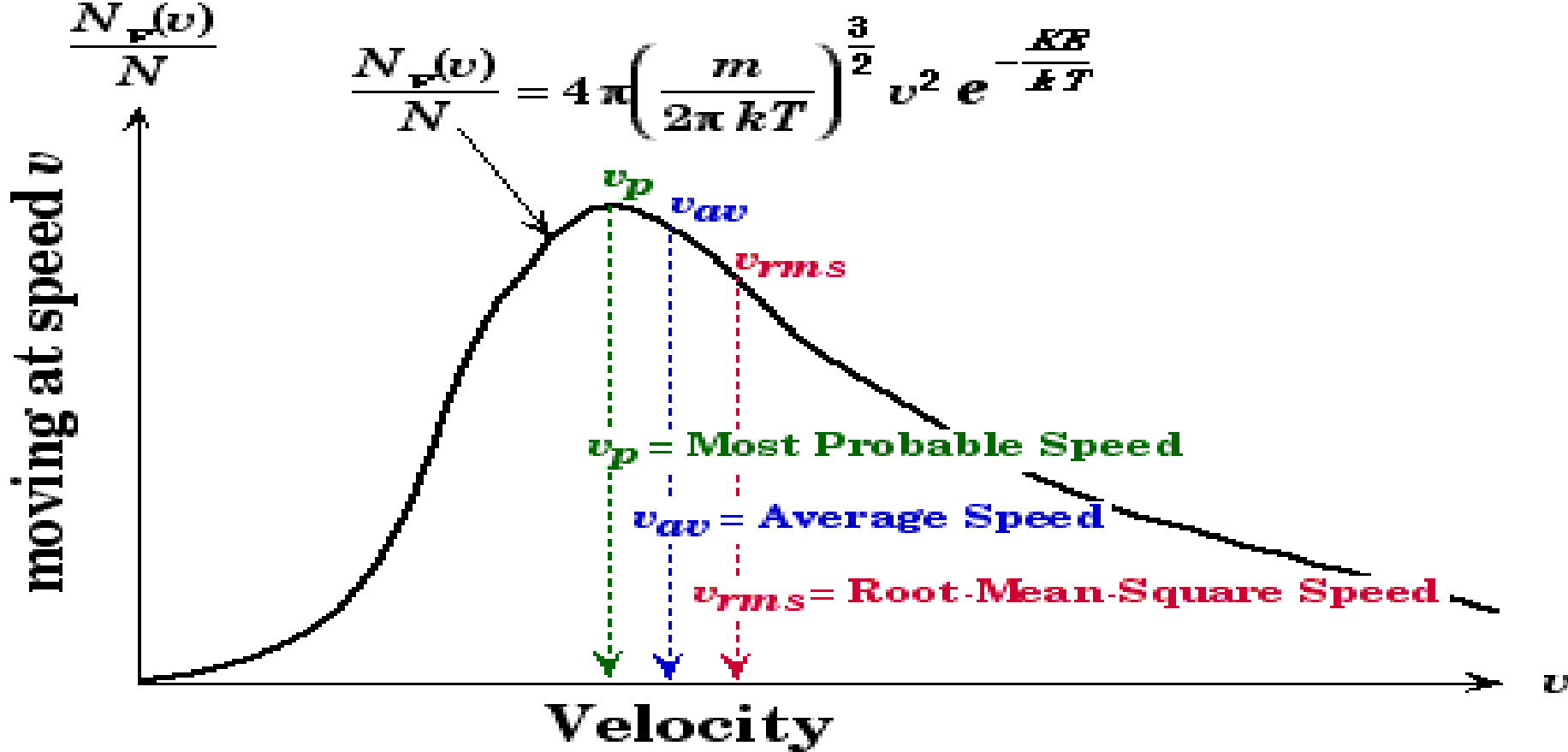
$$v_p = \sqrt{2\frac{kT}{m}} \qquad v_{av} = \sqrt{\frac{8}{\pi}\frac{kT}{m}} \qquad v_{rms} = \sqrt{3\frac{kT}{m}}$$

RATIO OF MOLECULAR SPEEDS

1   :   1.128   :   1.224

PLEASE NOTE
*R=K when m=M

BEHAVIOUR OF REAL GASES : DEVIATION FROM IDEAL GAS BEHAVIOUR

A gas which obeys the gas laws and the gas equation PV = nRT strictly at all temperatures and pressures is said to be an ideal gas. The molecules of ideal gases are assumed to be volume less points with no attractive forces between one another. But no real gas strictly obeys the gas equation at all temperatures and pressures. Deviations from ideal behaviour are observed particularly at high pressures or low temperatures. The deviation from ideal behaviour is expressed by introducing a factor Z known as compressibility factor in the ideal gas equation. Z may be expressed as Z = PV / nRT

- In case of ideal gas, PV = nRT so, Z = 1

- In case of real gas, PV ≠ nRt so,Z ≠ 1
  Thus in case of real gases Z can be < 1 or > 1

(i) When Z < 1, it is a negative deviation. It shows that the gas is more compressible than expected from ideal behaviour.

(ii) When Z > 1, it is a positive deviation. It shows that the gas is less compressible than expected from ideal behaviour.

Causes of deviation from ideal behaviour

The causes of deviations from ideal behaviour may be due to the following two assumptions of kinetic theory of gases. There are

- *The volume occupied by gas molecules is negligibly small as compared to the volume occupied by the gas.*
- *The forces of attraction between gas molecules are negligible.*
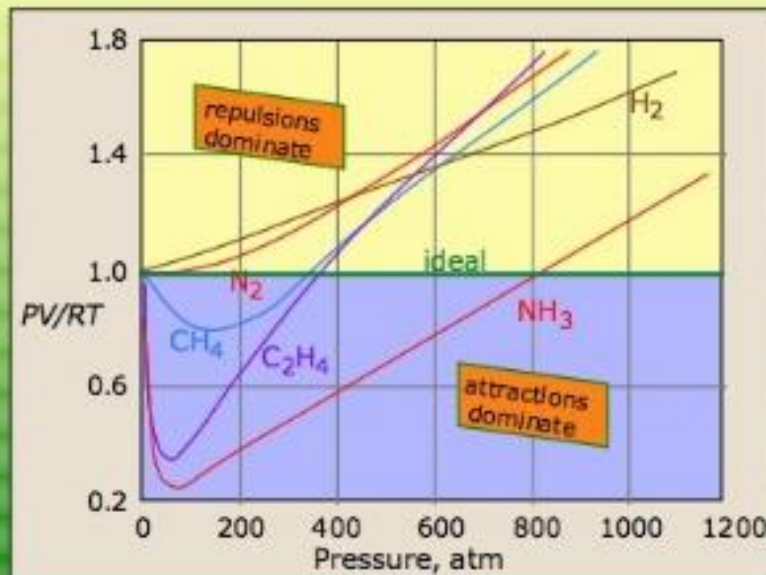
VAN DER WAALS EQUATION

Atoms and molecules are never truly ideal because they all interact with other gas particles; weak attractions between separate gas particles are known as intermolecular attractions or van der Waals forces after the chemist who proposed a correction to the ideal gas law to calculate pressure of a real gas. Van der Waals proposed that the ideal gas equation could be corrected for real gas behavior by subtracting the effective gas particle volume from the volume of the container and by correcting for intermolecular attractions:

$$PV = nRT$$

$$(P + attractions)(V - gasVolume) = nRT$$

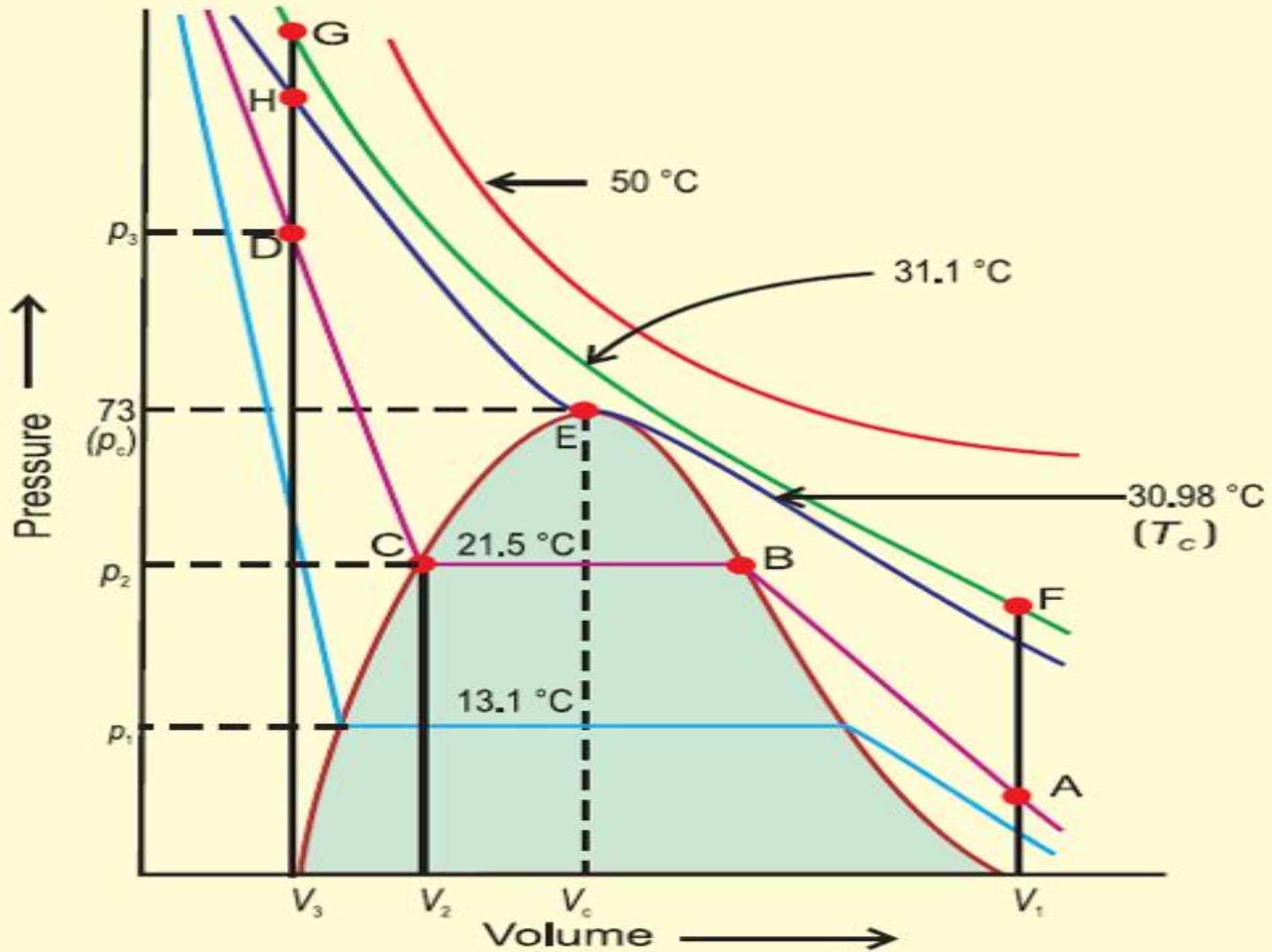$$\left(P + \frac{n^2 a}{V^2}\right)(V - nb) = nRT$$

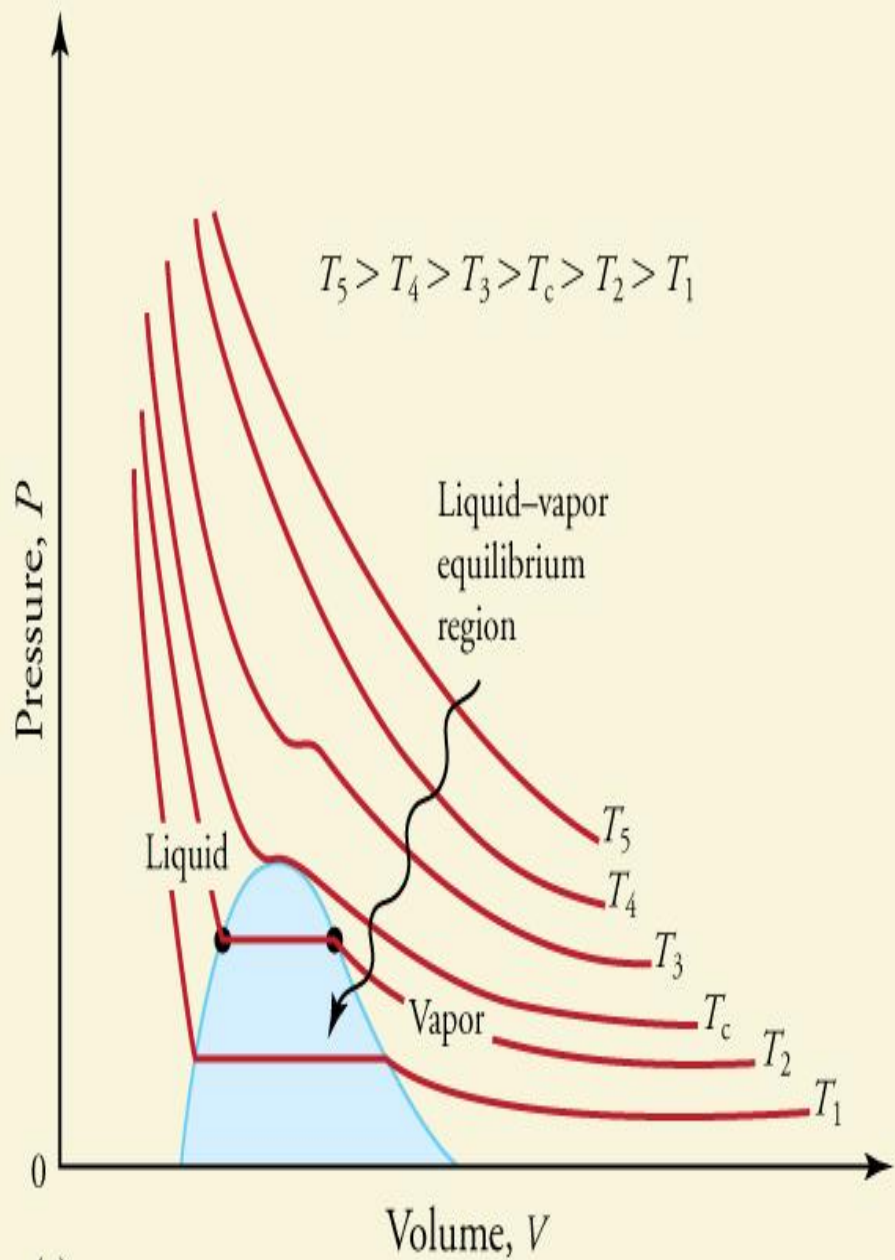# SIGNIFICANCE OF VAN DER WAALS CONSTANTS

- The value of constant "a" is measure of magnitude of attractive forces among the molecules of the gas. Its units are atm litre^2/mole^2

- The value of constant "b" is measure of the effective size of the molecules of the gas. Its units are litre/mole.

LIQUEFACTION OF GASES

Andrew's Experiments

$T_5 > T_4 > T_3 > T_c > T_2 > T_1$

Liquid–vapor equilibrium region

Pressure, P

Liquid

Vapor

T_5
T_4
T_3
T_c
T_2
T_1

0

Volume, V

(a)

Pressure, P

True but not ideal gas

Critical point

Liquid region

Saturation curve

Critical isotherm

Liquid–vapor region

Vapor region

T_c

0

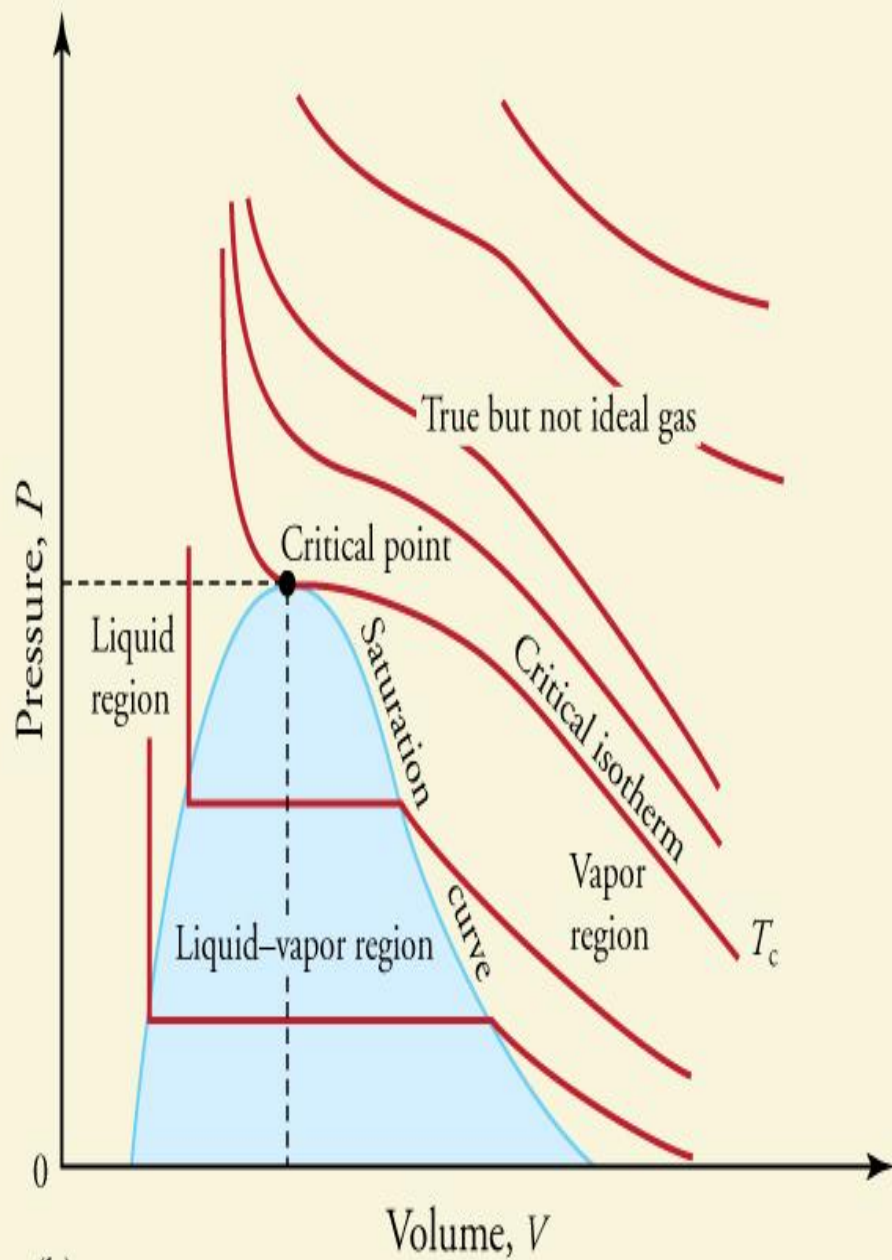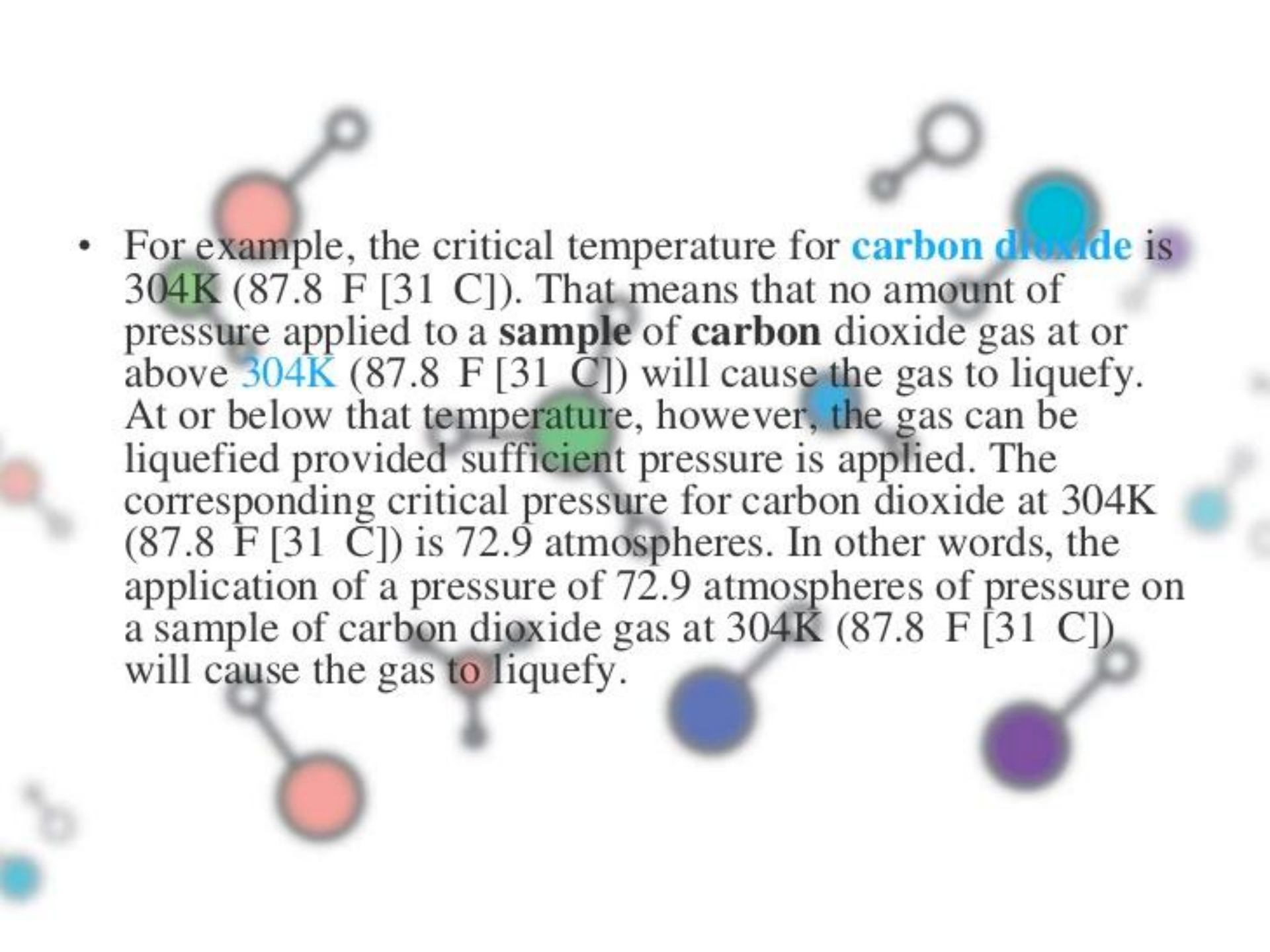Volume, V

(b)

CRITICAL TEMPRATURE AND PRESSURE
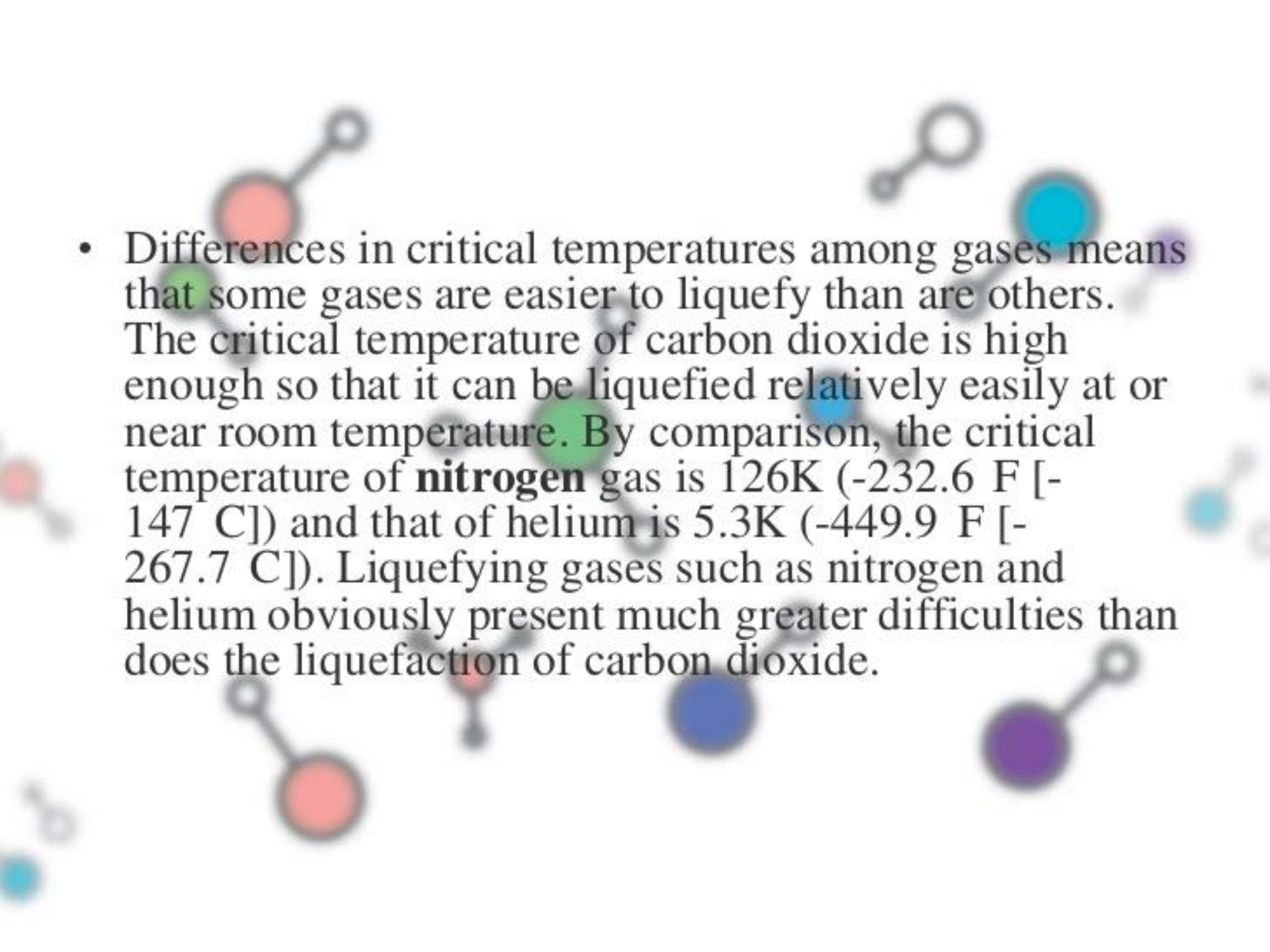
- Two important properties of gases are important in developing methods for their liquefaction: **critical temperature** and **critical pressure**.

- The highest temperature at which a substance can exist as a liquid, is called its critical temperature(Tc).There is corresponding pressure which is requires to bring about liquification at this critical temprature called critical pressure.

- For example, the critical temperature for **carbon dioxide** is 304K (87.8 F [31 C]). That means that no amount of pressure applied to a **sample** of **carbon** dioxide gas at or above 304K (87.8 F [31 C]) will cause the gas to liquefy. At or below that temperature, however, the gas can be liquefied provided sufficient pressure is applied. The corresponding critical pressure for carbon dioxide at 304K (87.8 F [31 C]) is 72.9 atmospheres. In other words, the application of a pressure of 72.9 atmospheres of pressure on a sample of carbon dioxide gas at 304K (87.8 F [31 C]) will cause the gas to liquefy.

- Differences in critical temperatures among gases means that some gases are easier to liquefy than are others. The critical temperature of carbon dioxide is high enough so that it can be liquefied relatively easily at or near room temperature. By comparison, the critical temperature of **nitrogen** gas is 126K (-232.6 F [-147 C]) and that of helium is 5.3K (-449.9 F [-267.7 C]). Liquefying gases such as nitrogen and helium obviously present much greater difficulties than does the liquefaction of carbon dioxide.

Gas	P_c, atm	V_c, $cm^3\ mol^{-1}$	T_c, K	$Z_c = \dfrac{P_c V_c}{nRT_c}$
He	2.3	57.8	5.3	0.306
H_2	12.8	65.0	33.2	0.304
Ne	26.9	41.7	44.4	0.302
N_2	33.6	90.1	126.1	0.291
O_2	50.3	74.4	154.5	0.302
CO_2	72.7	95.0	304.2	0.275
H_2O	218.0	55.6	647.3	0.227
NH_3	112.0	72.0	405.5	0.243
CH_4	45.8	99.0	191.0	0.290
C_2H_6	48.2	139.0	305.5	0.267
C_2H_4	50.5	124.0	417.2	0.275

Q1) 28gm of each of the following gases are taken at 300K and 600mm pressure .Which of these will have least volume?

- A)HBr
- B)HCl
- C)HF
- D)HI

Q2) By what factor does the avg. velocity of a gaseous molecule inc. when absolute temperature is doubled.

- A)1.4
- B)2
- C)2.8
- D)4

Q3) Dipole-induced dipole interactions are present in which of the following pairs?

- A)Water and Alcohol
- B)Chlorine and Carbon Tetrachloride
- C)Hydrogen Chloride and Helium
- D)Silicon Tetra fluoride and Helium.

THANK YOU

For any queries mail at:

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