Course Name: Chemistry
Course NO: CHE1203

State of Matter

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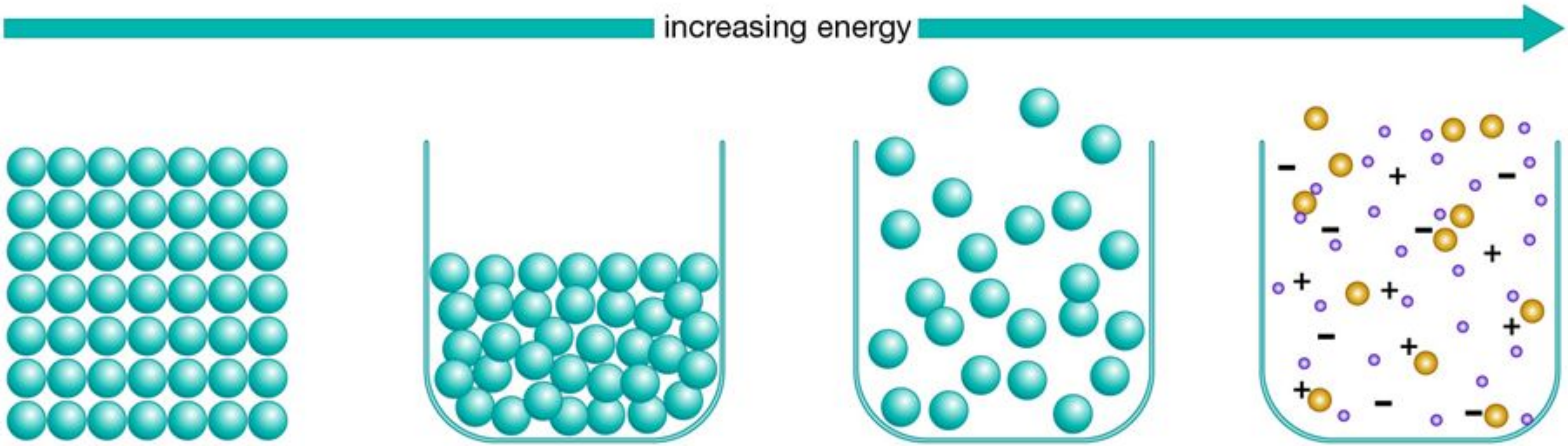
State of Matter

Physical states



Solid

The molecules that make up a solid are arranged in regular, repeating patterns. They are held firmly in place but can vibrate within a limited area.

Liquid

The molecules that make up a liquid flow easily around one another. They are kept from flying apart by attractive forces between them. Liquids assume the shape of their containers.

Gas

The molecules that make up a gas fly in all directions at great speeds. They are so far apart that the attractive forces between them are insignificant.

Plasma

At the very high temperatures of stars, atoms lose their electrons. The mixture of electrons and nuclei that results is the plasma state of matter.

Gas Laws

The three fundamental gas laws discover the relationship of pressure, temperature, volume and amount of gas. Boyle’s Law tells us that the volume of gas increases as the pressure decreases. Charles’ Law tells us that the volume of gas increases as the temperature increases. And Avogadro’s Law tell us that the volume of gas increases as the amount of gas increases. The ideal gas law is the combination of the three simple gas laws.

Boyle's Law

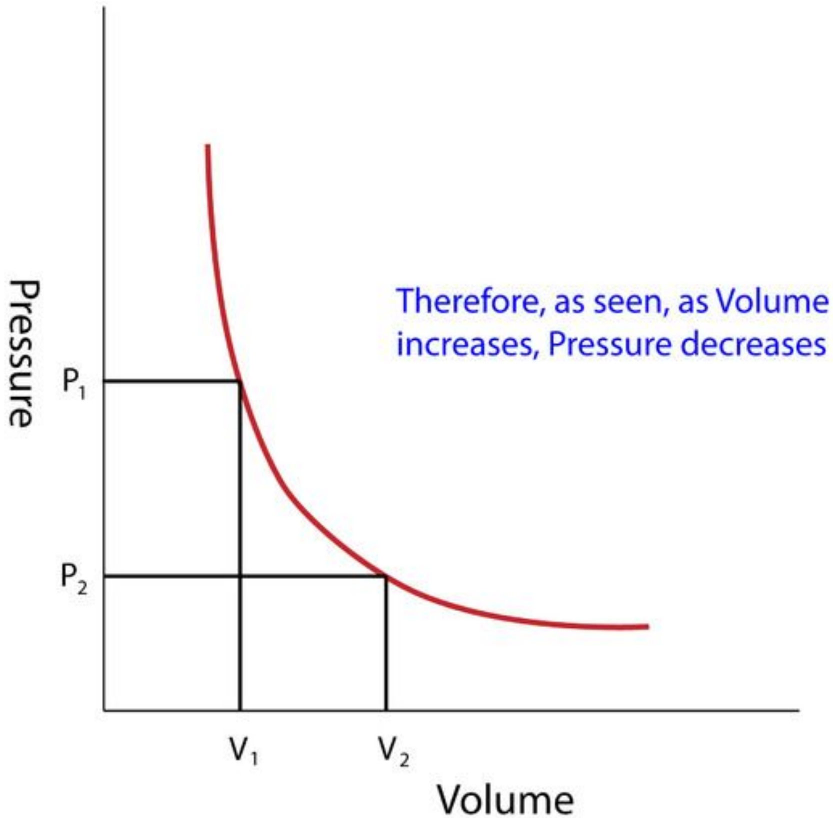
In 1662, Robert Boyle discovered the correlation between pressure (P) and Volume (V) [assuming the Temperature (T) and amount of gas(n) remained constant:

$P \propto 1/V \rightarrow PV = x$ (Gas Laws.2)

$V \propto 1/P \rightarrow PV = x$

where x is a constant depending on amount of gas at a given temperature.

Pressure is inversely proportional to Volume



Pressure (P) vs Volume (V) graph at constant temperature

Another form of the equation (assuming there are 2 sets of conditions, and setting both constants to each other) that might help solve problems is:

$$P_1V_1=x=P_2V_2$$

Charles' Law

In 1787, French physicists Jacques Charles, discovered the correlation between Temperature (T) and Volume (V) [assuming pressure (P) and amount of gas (n) remain constant]:

It states that, for a given mass of an ideal gas at constant pressure, the volume is directly proportional to its absolute temperature, assuming in a closed system.

$V \propto T \rightarrow V = yT$

where y is a constant depending on amount of gas and pressure.

Volume is directly proportional to Temperature

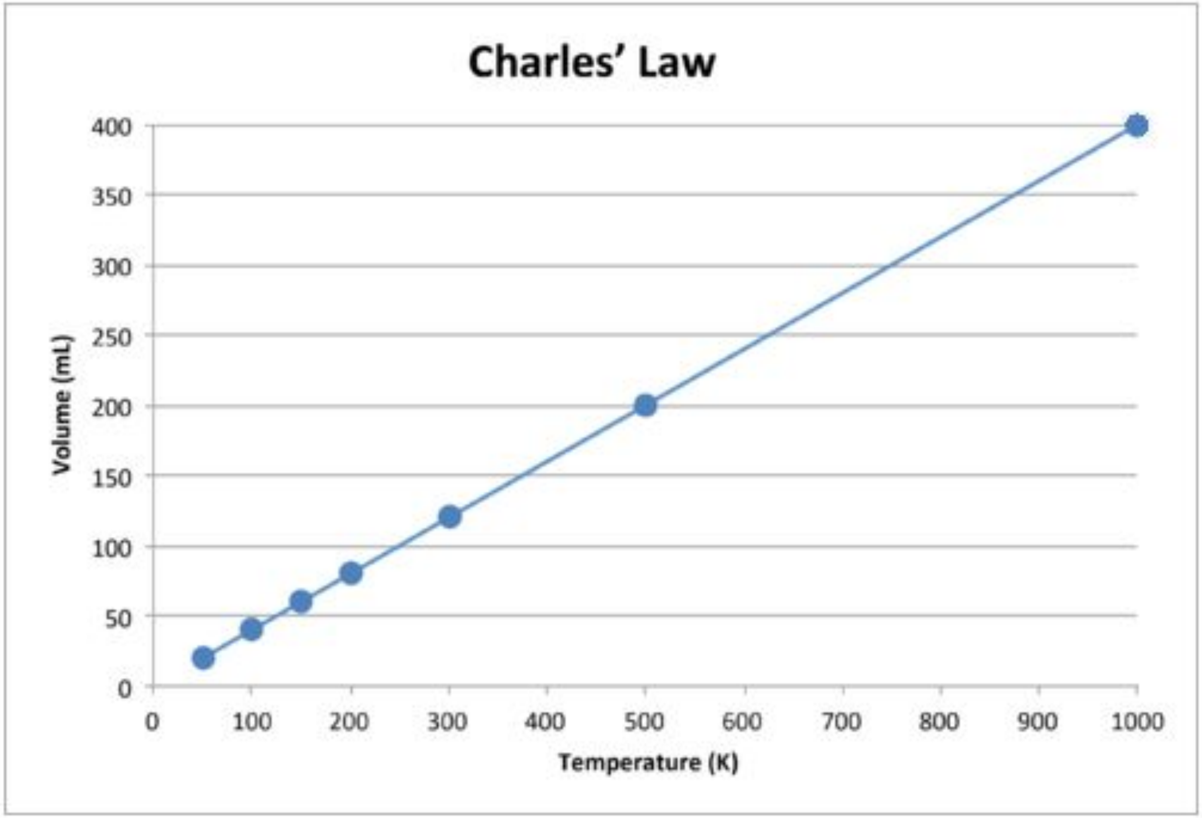


Fig. The volume of a gas increases as the Kelvin temperature increases.

Another form of the equation (assuming there are 2 sets of conditions, and setting both constants to each other) that might help solve problems is:

$$V_1/T_1=y=V_2/T_2$$

Avogadro’s Law

In 1811, Amedeo Avogadro fixed Gay-Lussac's issue in finding the correlation between the Amount of gas(n) and Volume(V) (assuming Temperature(T) and Pressure(P) remain constant):

$V \propto n \rightarrow V = zn$

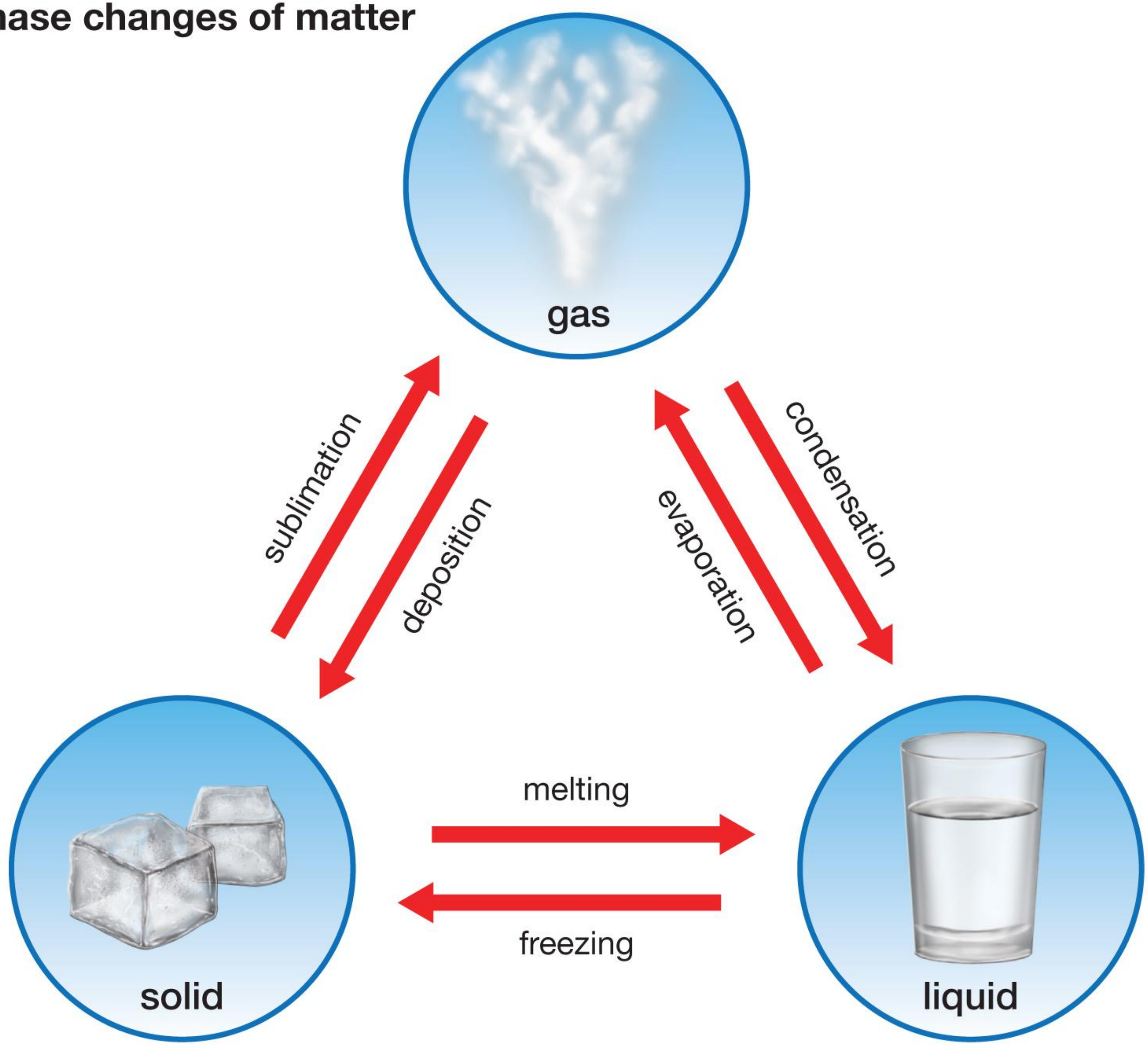
where z is a constant depending on Pressure and Temperature.

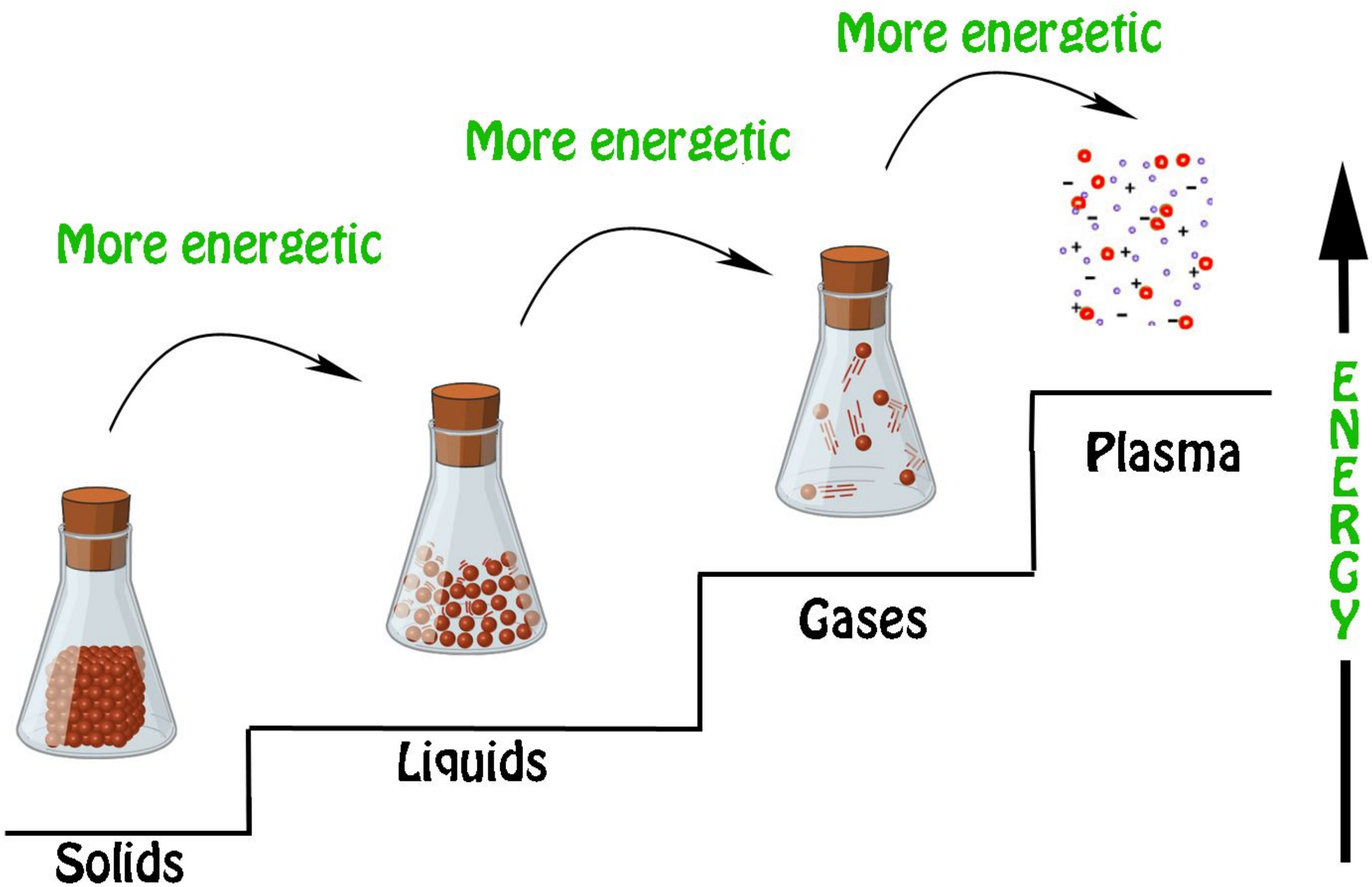
Volume(V) is directly proportional to the Amount of gas(n)

Another form of the equation (assuming there are 2 sets of conditions, and setting both constants to eachother) that might help solve problems is:

$$P_1/n_1=z=P_2/n_2$$

Phase changes of matter





As you go from solids, to liquids, to gases, and finally to plasma, the energy levels of the particles are increasing

Ideal Gas Law

The laws which deal with ideal gases are naturally called ideal gas laws and the laws are determined by the observational work of Boyle in the seventeenth century and Charles in the eighteenth century.

Ideal Gas Equation

The Ideal gas law is the equation of state of a hypothetical ideal gas. It is a good approximation to the behavior of many gases under many conditions, although it has several limitations. The ideal gas equation can be written as

$$\boldsymbol{PV = nRT}$$

Where,

P is the pressure of the ideal gas.

V is the volume of the ideal gas.

n is the amount of ideal gas measured in terms of moles.

R is the **universal gas constant**.

T is the temperature.

Derivation of the Ideal Gas Equation

Let us consider the pressure exerted by the gas to be '**p,**'

The volume of the gas be – '**v**'

Temperature be – **T**

n – be the number of moles of gas

Universal gas constant – **R**

According to **Boyle's Law**,

At constant **n** & **T,** the volume bears an inverse relation with the pressure exerted by a gas.

i.e. $v \propto 1/p$ ………………………………(i)

According to Charles' Law,

When **p** & **n** are constant, the volume of a gas bears a direct relation with the Temperature.

i.e. $v \propto T$ ……………………………….(ii)

According to Avogadro's Law,

When **p** & **T** are constant, then the volume of a gas bears a direct relation with the number of moles of gas.

i.e. $v \propto n$ ……………………………….(iii)

Combining all the three equations, we have-

$v \propto nT/p$ or $pv=nRT$

where **R** is the **Universal gas constant**

Ideal Gases

Ideal gas, or perfect gas, is the theoretical substance that helps establish the relationship of four gas variables, pressure (P), volume(V), the amount of gas(n) and temperature(T). It has characters described as follow:

(i) The particles in the gas are extremely small, so the gas does not occupy any spaces.

(ii) The ideal gas has constant, random and straight-line motion.

(iii) No forces between the particles of the gas. Particles only collide elastically with each other and with the walls of container.

Real Gas

Real gas, in contrast, has real volume and the collision of the particles is not elastic, because there are attractive forces between particles. As a result, the volume of real gas is much larger than of the ideal gas, and the pressure of real gas is lower than of ideal gas. All real gases tend to perform ideal gas behavior at low pressure and relatively high temperature.

Limitation of Ideal Gas Equation

Ideal Gas equation makes two assumptions:

(a) Gas particles have mass but zero volume

(b) There are no internal forces between particles

- The difference between the an ideal gas and a real gas are especially noticeable at very, very high pressure and very, very low temperature under these conditions.
- Molecules are close to each other
- Volume of molecules is not negligible compared with volume of container
- There are Van der Waals or dipole-dipole force of attraction between molecules

- Attractive forces pull the molecules towards each other and away from the walls of the container
- The pressure lower than expected for an ideal gas
- The effective volume of the gas is smaller than expected for an ideal gas

Van der Waals Equation

An equation due to van der Waals extends the ideal gas equation in a straightforward way. Van der Waals' equation is

$$(P + n^2a/V^2)(V\text{-}nb)=nRT$$

Derivation of Van der Waals Equation

The ideal gas law treats gas molecules as point particles that interact with their containers but not each other, meaning they neither take up space nor change kinetic energy during collisions. The ideal gas law states that volume (V) occupies by n moles of any gas has a pressure (P) at temperature (T) given by the following relationship, where R is the gas constant:

$$PV=nRT$$

To account for the volume that a real gas molecule takes up, the Van der Waals equation replaces V in the ideal gas law with (V_m-b), where V_m is the molar volume of the gas and b is the volume that is occupied by one mole of the molecules. This leads to:

$$P(V_m-b)=RT$$

The second modification made to the ideal gas law accounts for the fact that gas molecules do in fact interact with each other (they usually experience attraction at low pressures and repulsion at high pressures) and that real gases therefore show different compressibility than ideal gases.

Van der Waals provided for intermolecular interaction by adding to the observed pressure P in the equation of state a term a/V_m^2, where a is a constant whose value depends on the gas. The Van der Waals equation is therefore written as:

$$(P + a/V_m^2)(V_m - b) = RT$$

and, for n moles of gas, can also be written as the equation below:

$$(P + n^2 a/V^2)(V - nb) = nRT$$

where V_m is the molar volume of the gas, R is the universal gas constant, T is temperature, P is pressure, and V is volume.

When the molar volume V_m is large, b becomes negligible in comparison with V_m, a/V_m^2 becomes negligible with respect to P, and the Van der Waals equation reduces to the ideal gas law, $PV_m=RT$.

Kinetic Theory of Gases

The five basic tenets of the kinetic-molecular theory are as follows:

(i) A gas is composed of molecules that are separated by average distances that are much greater than the sizes of the molecules themselves. The volume occupied by the molecules of the gas is negligible compared to the volume of the gas itself.

(ii) The molecules of an ideal gas exert no attractive forces on each other, or on the walls of the container.

(iii) The molecules are in constant random motion, and as material bodies, they obey Newton's laws of motion. This means that the molecules move in straight lines until they collide with each other or with the walls of the container.

(iv) Collisions are perfectly elastic; when two molecules collide, they change their directions and kinetic energies, but the total kinetic energy is conserved. Collisions are not “sticky".

(v)The average kinetic energy of the gas molecules is directly proportional to the absolute temperature.