GAS LAWS-CLASS 9

INTRODUCTION

We live under a thick layer of gases – the atmosphere. The air that we breathe is a mixture of gases containing nitrogen (78% by volume), oxygen (21% by volume) and a few others including carbon dioxide, argon, etc. Oxygen is essential for the survival of animal life on earth and also for combustion. Green plants require carbon dioxide for photosynthesis. Ozone in the upper layers of the atmosphere protects life on earth from the ultraviolet rays of the sun.

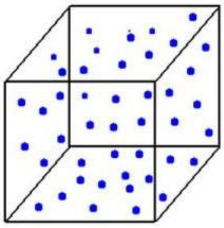
Characteristic Properties of Gases

Gases Have Very Low Density

Irrespective of the substance, the gaseous state is characterized by very low density in comparison with the solid and liquid states.

Gases Fill the Container Completely

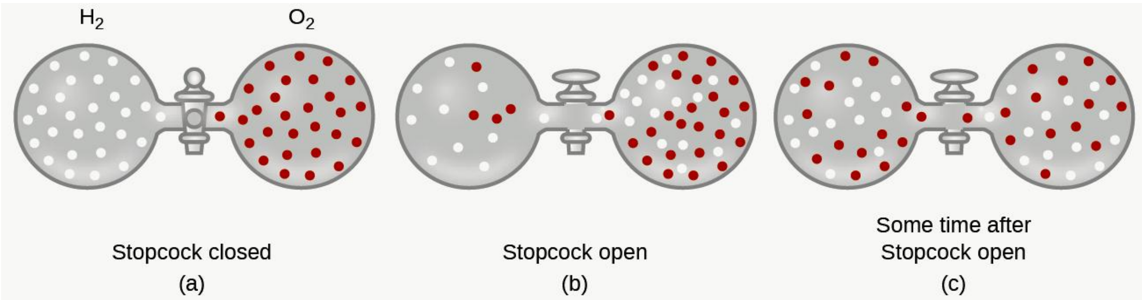
A gas expands freely and so it completely fills the container, whatever be the size of the container.



Even if you allow some amount of gas to escape from a gas jar, the jar remains completely filled. In other words, the volume occupied by the gas remains equal to that of the container. Gases can also be compressed; that is, their volumes can be reduced by applying pressure. The same amount of gas may be made to occupy containers of different capacities and, in all cases, the container is completely filed. However, the pressure of the gas in all the cases will be different.

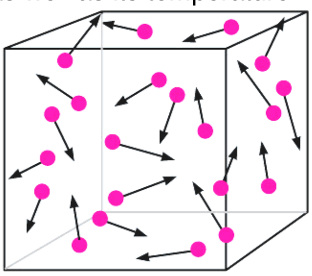
Gases Diffuse into One Another and Effuse

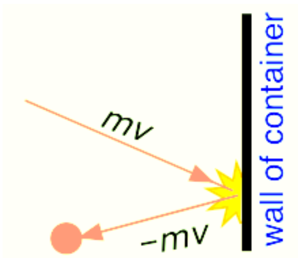
Gases diffuse into each other. If two bulbs filled with different gases are connected, the gases diffuse and mix with each other, forming a homogenous mixture. That is, the composition of the mixture is the same in both the bulbs. The components of the mixture do not separate on standing. Gases also effuse through a hole in the container. Gas-filled rubber balloons gradually shrink due to the effusion through the pores of the balloon.



Gases Exert Pressure in All Directions

A gas exerts pressure on any surface in contract with it. Hence, a gas exerts pressure on the entire surface area of the walls of the container. The pressure exerted by a given mass of a gas depends on the volume occupied by the gas as well as its temperature.





Measurable Properties of Gases

A gaseous system can have following measurable properties:

(a) Number of moles (b) Pressure (c) Temperature (d) Volume

(a) Number of Moles: As in case of gases the mass of gas is insignificant and better choice is to represent number of moles.

For a gas its molar volume is constant at S.T.P. i.e. 22.4 litre irrespective of type of gas.

No. of moles of a gas at S.T.P.

$$= \frac{\text{vol. of gas at STP}}{22.4}$$

(b) Pressure: The force exerted per unit area on the walls of container by the gaseous molecules is measured as pressure of a gas.

$$\text{Pressure} = \frac{\text{force}}{\text{area}}$$

Units of Pressure

Pressure can be expressed in several different units.

1. Centimeters or millimeters of Hg: The length of the mercury column expressed in centimeters or millimeters of Hg.
2. Torr: Named in honour of Torricelli, 1 Torr = 1 mm of Hg
3. Standard atmosphere (atm.): The atmospheric pressure at sea level, i.e., 76cm or 760 mm of Hg, is said to be one standard atmosphere (atm.)
4. Pascal(Pa): Named in honour of Blaise Pascal, 1 Pascal (Pa) is equal to the pressure due to a force of 1 Newton (N) on a surface area of 1 m^2. 1 N is the fore that will give an acceleration of 1 m/s^2 to a body of mass of 1 kg. The Pascal is a very small unit – even low pressures are generally expressed in kilopascal (kPa).
5. Bar: Because the Pascal is a very small unit, the more convenient bar (1 bar = 10^5 Pa) is commonly used for expressing pressure near normal atmospheric pressure. The below table gives the relationships among the different units of pressure.

Conversion Relationships Among Pressure Units

	Atm.	Torr	Pascal	Bar
1 Atm.	1	760	1.01325×10^5	1.01325
1 Torr	1.31579×10^{-3}	1	133.322	1.33322×10^{-5}
1 Pascal	9.8692×10^{-6}	7.50064×10^{-3}	1	1×10^{-5}
1 Bar	0.98692	750.06168	1×10^5	1

(c) Temperature: The degree of hotness of substance is measured as temperature.

Units: There are four common temperature scales are used in which minimum and maximum temperature is decided by freezing and boiling point of water.

Temperature scale	Freezing point of H_2O	Boiling point of H_2O	Range
Celsius scale	0^oC	100^oC	100
Fahrenheit scale	32^oF	212^oF	180
Kelvin Scale	273 K	373 K	100

The above mentioned scales can be interrelated by using following formula.

$$\frac{C}{5} = \frac{F-32}{g} = \frac{K-237}{5}$$

Volume: The space occupied by any matter is measured as its volume.

Units

1 $\text{Metre}^3 = 10^6 cm^3$ 1 Litre = 10^3 ml = $10^3 cm^3$ 1 Metre^3 = 1000 litre.

Illustration: ***Convert (a) 114 cmHg pressure into the atmosphere unit, and (b) 0.5 atm into torr.***

Solution: (a) 114 cmHg = $\frac{114 \text{cmHg} \times 1 \text{atm}}{76 \text{cmHg}} = \frac{114}{76}$ atm = 1.5 atm.

(b) 0.5 atm = $\frac{0.5\,atm \times 760\,torr}{1\,atm}$ = 380 torr.

GAS LAWS

BOYLE' S LAW

On the basis of his experimental data Boyle gave the following law, called Boyle' s law.
The volume of a fixed mass of gas at constant temperature is inversely proportional to its pressure.

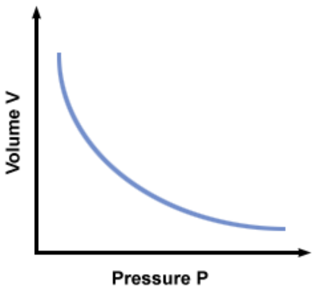
Mathematically, for a given mass of a gas at constant temperature. If the pressure be p and the volume V,

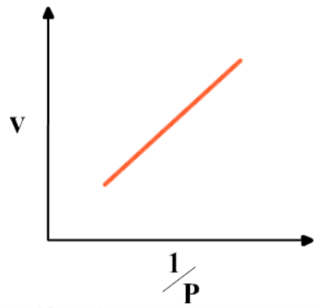
$V \propto \frac{1}{p}$

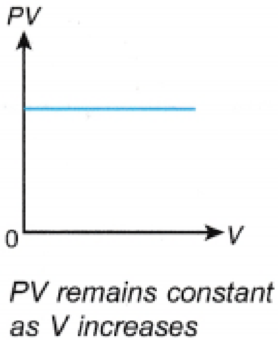
$\therefore \quad V = k\,\frac{1}{p}$,

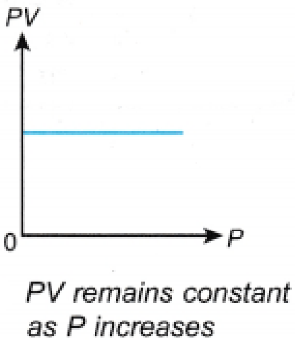
where k is the proportionality constant.

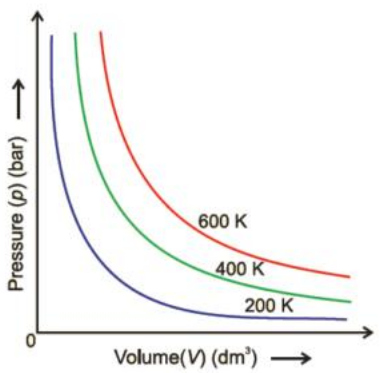
$P = \frac{k}{V}$

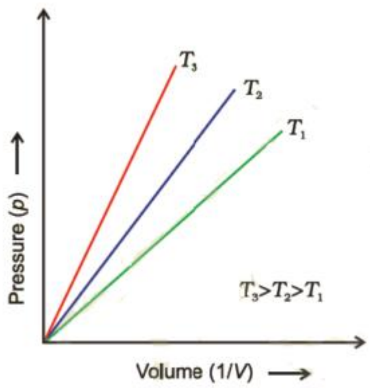


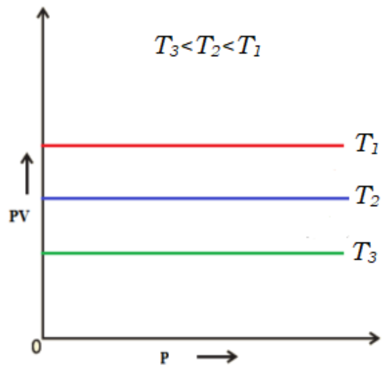












Thus Boyle' s law may also be stated as follows.
The pressure of a fixed mass of a gas, at constant temperature, is inversely proportional to its volume.
Equations lead to the relationship

$pV = k$

We can use this relationship to calculate the new volume of a sample of gas as the same temperature. Let the pressure be changed to p_2, resulting in volume V_2 at the temperature. Then,

$p_1V_1 = k_1$ and $p_2V_2 = k_2$ when $k_1 = k_2$

Then, $p_1V_1 = p_2V_2$

Thus, Boyle' s law may also be stated as follows…

The product of the pressure and the volume of a fixed mass of gas, at constant temperature, is a constant.

Pressure-Volume Calculations Using Boyle' s Law

If any three of the four quantities in above equation are known, the fourth can be calculated. For example, if we know the initial pressure p_1, the initial volume V_1 and the final pressure p_2, then we can easily calculate the final volume V_2.

$$V_2 = \frac{p_1V_1}{P_2}$$

The pressure p_2 required to change the volume of a gas without mentioning the pressure because the volume is dependent on the pressure. The same amount of gas occupies different volumes under different pressures. For example, if a fixed mass of a gas occupies a volume of 100 mL under a pressure of 1 atm, then the volume will be 10 mL under a pressure of 10 atm at the same temperature. As we will see later, it is essential to specify the temperature as well because the volume of a gas depends upon the temperature too.

Illustration : ***A sample of a gas has a volume of 80 mL at a pressure of 432 mmHg at a certain temperature. What will be the volume if the pressure is changed to 720mmHg, keeping the temperature constant?***

Solution: Here, p_1 = 432 mmHg, V_1 = 80 mL, p_2 = 720 mmHg, V_2 =?

From Boyle' s law $p_1V_1 = p_2V_2$, we get

$$V_2 = \frac{p_1V_1}{p_2}$$

or $$V_2 = \frac{432\,mmHg \times 80\,mL}{720\,mmHg} = \frac{432 \times 80}{720}\,mL = 48\,mL.$$

The new volume of the gas is 48 mL.

Illustration : ***The volume of a sample of a gas is 25 mL at a pressure of 76.0 cmHg. At what pressure will the volume 15 mL, keeping the temperature constant?***

Solution: Let the required pressure be p cmHg. Then,

p cmHg x 15 mL = 76.0 cmHg x 25 mL

$$\therefore p = \frac{76.0\,cmHg \times 25\,mL}{15\,mL} = \frac{76.0 \times 25}{15}\,cmHg = 126.7\,cmHg.$$

CHARLES' LAW

Charles discovered that air expands on heating and contracts on cooling, but the change in volume is not proportional to the temperature on the Celsius scale.

Thus, a doubling of temperature from 20^oC to 40^oC does not double the volume of gas. In fact, Charles stated his finding in the following form, known as Charle' s law.

The change in the volume of a sample of a gas, at constant pressure, due to a change in temperature by 1^oC is 1/273 part of the volume of the sample at 0^oC.

Mathematically, if V_t be the volume of a sample of a gas at a temperature t^oC and V_0 be the volume at 0^oC, V_t and V_0 are related, according to Charles' law, by the relationship.

$$V_t - V_0 = V_0 \times \frac{t}{273}$$

or $$V_t = V_0 = \frac{V_0 \times t}{273} = V_0 \times \frac{273+t}{273}$$

or $$V_t = V_0\left(1+\frac{t}{273}\right)$$

or $$V_0 = \frac{V_t \times 273}{273+t}$$

Volume-Temperature Calculations Using Charles' Law

Given the volume of a gas sample at temperature $t_1{}^oC$, we can calculate the volume at any other temperature $t_2\,{}^0C$ from Charles law if we know its volume at 0^0C. The volume of the gas sample at 0^oC can be calculated by using the above equation.

The Absolute Temperature Scale (Kelvin Scale)

From Charles law, it is clear that the volume of gas at a temperature lower than 0^oC, say $-t^oC$, may be inferred from Equation

$$V_{-t} = V_0\left(1-\frac{t}{273}\right)$$

And, at -273^oC,

$$V_{-273} = V_0\left(1-\frac{273}{273}\right) = V_0(1-1) = V_0 \times 0 = 0.$$

This leads to the very important conclusion that the volume of gas at -273^oC will be zero. It may be further concluded that at any temperature below -273^oC, the volume of the gas will be less than zero, i.e., negative, which is impossible.

Illustration : ***(a) Convert the following temperatures on the Celsius scale to the Kelvin scale.***

(i) 25°C (ii) 77°C (iii) -25°C

(b) Convert the following temperatures on the Kelvin scale to the Celsius scale.

(i) 353 K (ii) 159 K (iii) 254 K

Solution: (a) (i) $25^oC = (273 + 25)\ K = 298\ K$

(ii) $77^oC = (273 + 77)\ K = 350\ K$

(iii) $254\ K = (254-273)^oC = -19^oC$

Charles Law in Terms of the Kelvin Scale of Temperature

If V_1 and V_2 be the volumes of a sample of gas at t_1^oC respectively then from Equation, we can state that

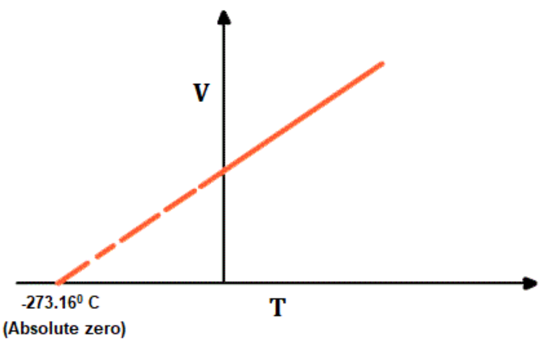
$$V_1 = V_0 \times \frac{273+t_1}{273} \text{ and } V_2 = V_0 \times \frac{273+t_2}{273}$$

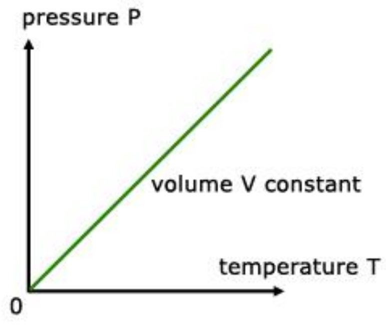
$$\therefore \frac{V_2}{V_1} = \frac{V_0 \times (273+t_2)}{V_0 \times (273+t_1)} = \frac{273+t_2}{273+t_1} = \frac{T_2}{T_1} \quad \text{or} \quad \frac{V_2}{V_1} = \frac{T_2}{T_1}$$

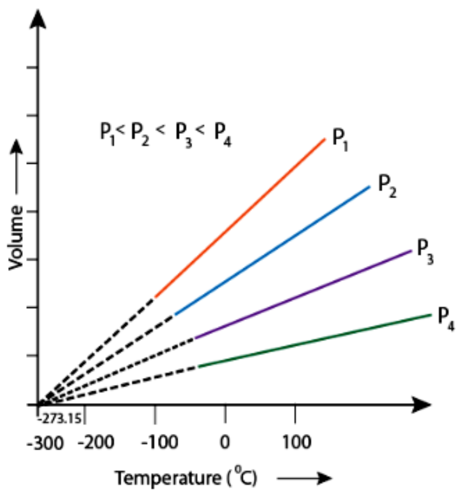
where T_1 and T_2 are the Kelvin temperatures corresponding to t_1^oC respectively. From the above equation, we see that the volume of a sample of a gas, at a constant pressure, is directly proportional to its temperature on the Kelvin scale.

Charles law, therefore, may be stated as follows.

The volume of a fixed mass of gas, at constant pressure, is directly proportional to its Kelvin temperature (T).







Equation can be used to calculate the volume V_2 of a sample of a gas at any given temperature T_2 if the volume V_1 at any other temperature T_1 is known. In fact, if any three of the four quantities in Equation are known, the fourth can be calculated.

Illustration : ***A 452-mL of nitrogen gas is heated from 37°C to 192°C at constant pressure. What is its new volume?***

Solution: Let V_1 and T_1 represent the initial volume and temperature (Kelvin scale), and V_2 and T_2 the final volume and temperature (Kelvin scale). Then

V_1 = 452 mL, T_1 = (273 + 37) = 310 V_2 = ? T_2 =(273 + 192) = 465

From the equation, $\frac{V_2}{V_1} = \frac{T_2}{T_1}$

$$\Rightarrow \quad V_2 = \frac{V_1 \times T_2}{T_1} = \frac{452\text{mL} \times 465}{310} = 678 \text{ mL.}$$

The volume of the sample of the gas at 192°C is 678 mL.

The Combined Gas Law Equation

By now we are fully conversant with the fact that volume, pressure and temperature are the important variables of a given mass of a gas. We have also seen that Boyle' s and Charle' s laws connect only two of these variables, keeping the third constant. In practical conditions, however, all three vary. Boyle' s law or Charles law alone are, therefore, not helpful in calculating the final value of a variable (pressure, volume or temperature) in cases where all the three change. Combining Boyle' s law and Charles' law, however, we get an expression which connects the three variables, and this equation is called the combined gas law equation. Let us see how we can obtain such an equation. For a fixed mass of gas, the gas laws are

Boyle' s law: $V \propto \frac{1}{p}$, when T is constant, and

Charles' law: $V \propto T$, when p is constant.

Combining the two laws, we get

$$V \propto \frac{1}{p}T$$

When T and P vary.

$$\therefore V = k\frac{T}{P} \quad \text{or} \quad \frac{pV}{T} = k,$$

where k is the proportionality constant.

We should remember that the value of k is different for different masses of a gas.

Let us suppose that, for a given mass of a gas, the initial pressure, volume and temperature are p_1, V_1 and T_1, which change to p_2, V_2 and T_2 respectively. Then,

$$\frac{p_1V_1}{T_1} = k \text{ and } \frac{p_2V_2}{T_2} = k.$$

Therefore, $\frac{p_1V_1}{T_1} = \frac{p_2V_2}{T_2}$

The above equation is called the combined gas law equation. This equation is applicable to a fixed mass of a gas. If any five of the six quantities are known, the sixth can be calculated.

Illustration : ***A sample of a gas occupies 10.0 L at 240°C under a pressure of 80.0 kPa. At what temperature would the gas occupy 20.0 L, if we increase the pressure to 100.0 kPa?***

Solution: p_1 = 80.0 kPa, V_1 = 10.0 L, T_1 = (240 + 273) = 513.

Since $\frac{p_1V_1}{T_1} = \frac{p_2V_2}{T_2}$

$$T_2 = \frac{p_2V_2T_1}{p_1V_1}$$

Substituting the values,

$$T_2 = \frac{100.0\,kPa \times 20.0\,L \times 513}{80.0\,kPa \times 10.0\,L} = 1282.5\ K$$

On the Celsius scale, the required temperature = $(1282.5 - 273)^0C$
= 1009.5^0C

Standard Temperature and Pressure

The standard temperature and pressure (s.t.p.), by general convention, are 0^oC (273 K) and 1 atm (=760 mmHg or 760 Torr).

When we say that the volume of a certain sample of a gas at s.t.p. is 1 L, we simultaneously define the temperature as 273 K and the pressure as 1 atm. Given the volume of a gas at any arbitrary temperature and pressure, we can calculate its volume at s.t.p. using the combined gas equation.

Illustration : ***A gas sample occupies 55mL at 182^oC and 3 atm. What will be the volume of the sample at s.t.p.?***

Solution: $p_1 = 3$ atm., $V_1 = 55$ mL, At s.t.p.,
$T_1 = (273 + 182) = 455$ $p_2 = 1$ atm., $V_2 = ?$ $T_2 = 273$.
From the combined gas equation,

$$\frac{p_1V_1}{T_1} = \frac{p_2V_2}{T_2} \Rightarrow V_2 = \frac{p_1V_1T_2}{p_2T_1}.$$

Substituting the values, we get

$$V_2 = \frac{3\,atm \times 55\,mL \times 273}{1\,atm \times 455\,K} = 99\ mL$$

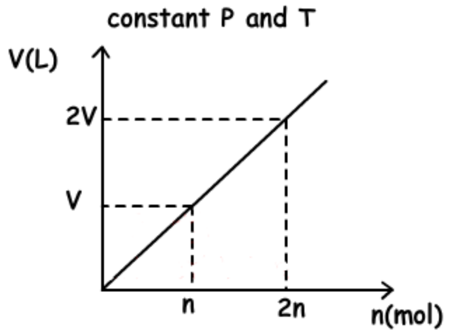
The volume of the gas sample at s.t.p. will be 99 mL.

Avogadro' s Hypothesis, Mole and Molar Volume

Under the same conditions of temperature and pressure, equal volumes of all gases contain an equal number of molecules or moles.

Mathematically, $V \propto n$, at constant T and p (n is the number of molecules or moles)

The mole, a package that contains 6.022×10^{23} molecules, is the connection between the microscopic world of molecules and the real world of macroscopic gas samples. The number of molecules contained in 1 mol of gas (i.e., 6.022×10^{23}) is also known as Avogadro constant. The volume of a gas measured at s.t.p. that contains 1 mol of it is called the s.t.p. molar volume or standard molar volume. Experiments have shown that for every gas (or a mixture of gases) the s.t.p. molar volume is 22.4 L. Thus, if we know the volume of gas sample at s.t.p., we can calculate the number of moles and thereby the number of molecules in the sample.



1 mol of a substance is also equal to the mass in grams numerically equal to the molecular mass. For example, the molecular mass of CO_2 is 44, and 1 mol of CO_2 has a mass of 44 g. Thus, we can also calculate the number of moles (hence the number of molecules) contained in a gas by knowing its mass. Thus, while mass of 22.4 L of hydrogen at s.t.p. is equal to 2g that of 22.4 L of N_2 is 28 g.

Illustration : ***Calculate the number of molecules contained in (a) 2.24 L of a gas at s.t.p., and (b) 4.4 g of hydrogen gas at s.t.p..***

Solution: (a) We must first find the number of moles equivalent to 2.24 L of the gas at s.t.p.
Since 22.4 L of a gas at s.t.p. = 1 mole,

2.24 L of the gas at s.t.p. = 2.24 L x $\frac{1mol}{22.4L}$

= 0.1 mol

Since 1 mol of a gas contains 6.022 x 10^{23} molecules,

0.1 mol of the gas contains 0.1 x 6.022 x 10^{23} molecules

= 6.022 x 10^{22} molecules.

(b) We must find the number of moles equivalent to 4.4g of H_2 gas.

Since the molecular mass of H_2 = 2.2 mol.

2.2 mol of H_2 gas contains 2.2 x 6.022 x 10^{23} molecules

= 1.32 x 10^{24} molecules.

Ideal Gas Equation

Boyle' s law: $V \propto \frac{1}{p}$ (T, n constant).

Charles' law: $V \propto T$;(p, n constant).

Avogadro' s law: $V \propto n$; (p, T constant).

Combining the three relations,

$V \alpha \frac{nT}{p}$, when n, p and T all vary.

Hence, we can say that

$V = \frac{RnT}{p}$,

Where R is the constant of proportionality.

pV = nRT.

We can rearrange the equation as follows.

The above equation is called the gas equation. The value of the constant of proportionality R does not change with a change in the values of the mass, pressure, volume or temperature of the gas, and is the same for all gases. It is, therefore, called the universal gas constant. Its value may be found by employing the knowledge that 1 mol of a gas (n =1) at s.t.p. occupies 22.4 L. The numerical value of R depends upon the units used for pressure and volume. Solving Equation for R and substituting the known values,

$$R = \frac{1atm \times 22.4L}{1mol \times 273K} = 0.0821 \text{ atm/mol K}$$

If the pressure is expressed in pascals and the volume in cubic meters, then

$$R = \frac{1.01325 \times 10^5 Pa \times 22.4 \times 10^{-3} m^3}{1mol \times 273K} = 8.314 \frac{Pa m^3}{mol K} = 8.314 J/mol\ K$$

(1 Pa = 1 N/m^2 ; $\therefore$ 1 Pa x m^3 = 1 N x m = 1J.)

In solving any problem, it is essential that all units of volume, pressure and temperature be consistent with the value used for R. For example, if the pressure is expressed is expressed in millimetres of Hg (or torr), and volume milliliters, you can express the value of R in litre atmospheres per mole Kelvin only after changing the unit of pressure to atmospheres and that of volume to litres.

Illustration : ***At what temperature will 0.005 mole of a gas occupy 600 mL at a pressure of 750 mmHg?***

Solution: From the gas equation,

T = pV/nR.

Let us choose to express the value of the gas constant R in litre atmospheres per mole Kelvin. Then we must express the volume in litres and the pressure in atmospheres.

600 mL = 600 mL x $\frac{1L}{1000mL}$ = 0.6 L.

750 mmHg = 750 mmHg x $\frac{1atm}{760mmHg} = \frac{75}{76}$atm

Substituting the values in the expression for T,

$$T = \frac{(75/76)atm \times 0.6L}{0.005mol \times 0.0821 L\,atm/K\,mol} = \frac{75 \times 0.6}{76 \times 0.005 \times 0.0821} K = 1442.4K.$$

The required temperature on the Celsius scale

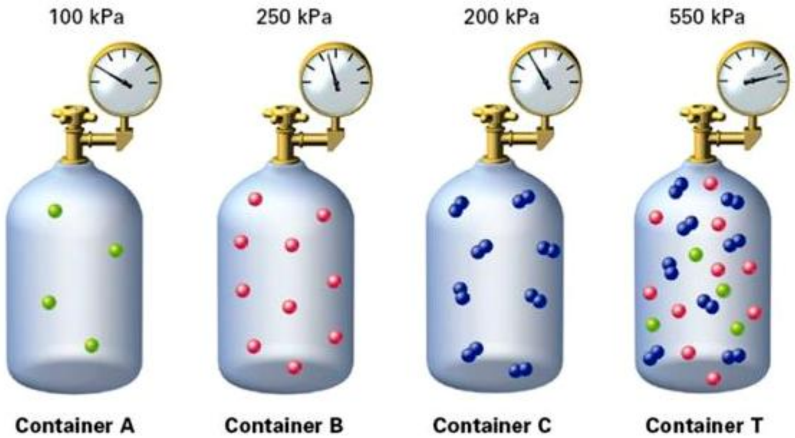
$$= \frac{75 x 0.6}{0.005 mol x 0.0821 L atm/K mol} = \frac{75 x 0.6}{76 x 0.005 x 0.0821} \text{ K} = 1442.4\text{K}$$

The required temperature on the Celsius scale = $(1442.4 - 273)^{o}C$

$= 1169.4^{o}C$.

Hence, at a temperature of $1169.4^{0}C$ and under a pressure of 750 mmHg, 0.005 mol of a gas will occupy 600 mL.

Dalton' s Law: John Dalton visualized that in a mixture of gases, each component of gas exerted a pressure as if it were alone in the container. The individual pressure of each gas in the mixture is defined as its partial pressure. It states that: the total pressure of a mixture of non - reacting gases is equal to the sum of the partial pressures of the individual gases present.



Mathematically the law can be expressed as $P = P_1 + P_2 + P_3$.... (V and T are constant) where P_1, P_2 and P_3 are partial pressures of the three gases 1, 2 and 3; and so on:

Dalton' s Law of partial pressures follows by application of the ideal gas equation PV = nRT separately to each gas of the mixture. Thus we can write the partial pressures P_1, P_2 and P_3 of the three gases as:

$$P_1 = n_1\left(RT/V\right), P_2 = n_2\left(RT/V\right), P_3 = n_3\left(RT/V\right)$$

Where n_1, n_2 and n_3 are moles of gases 1, 2 and 3. The total pressure, P_t, of the mixture is

$$P_t = (n_1 + n_2 + n_3) RT/V, P_t = n_t RT/V$$

In other words the total pressure of the mixture is determined by the total number of moles present whether of just one gas or mixture of gases.

$$P_1 = \frac{n_1}{n_1 + n_2} P_t \quad ; P_2 = \frac{n_2}{n_1 + n_2} P_t$$

$P_1 = X_1 . P_t$ and $P_2 = X_2 . P_t$

Partial pressure of gas = its mole fraction $\times$ Total pressure

$$\% \text{ of gas in mixture} = \frac{\text{Partial pressure of gas}}{\text{Total Pressure}} \times 100$$

$$P_{\text{Dry Gas}} = P_{\text{Moist Gas}} - P_{\text{Water Vapour}} = P_{\text{Observed}} - \text{Aqueous Tension}$$

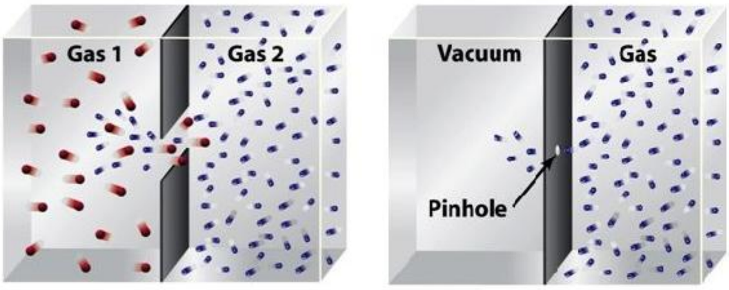
Graham' s Law

When two gases are placed in contact, they mix spontaneously. This is due to the movement of molecules of gases by random motion of the molecules and is called Diffusion. Thomas Graham observed that molecules with smaller masses diffused faster than heavy molecules. In 1829 Graham formulated what is now known as Graham' s Law of Diffusion.

It states that: under the same conditions of temperature and pressure, the rates of diffusion of different gases are inversely proportional to the square roots of their molecular masses or densities.

Mathematically the law can be expressed as

$$\frac{r_1}{r_2} = \sqrt{\frac{M_2}{M_1}} = \sqrt{\frac{d_2}{d_1}} = \frac{V_1 T_2}{V_2 T_1}$$



Diffusion is the mixing of gas molecules by random motion under conditions where molecular collisions occur.

Effusion is the escape of a gas through a pinhole into a vacuum without molecular collisions.

Where r_1 and r_2 are the rates of diffusion of gases 1and 2, while M_1 and M_2 are their molecular masses and V_1 and V_2 are the volumes of the gases diffused, T_1 and T_2 are the time taken to diffuse of gases1 and 2

When a gas escapes through a pin-hole into a region of low pressure or vacuum, the process is called Effusion. The rate of effusion of gas also depends on the molecular mass of the gas. Mathematically,

$$\frac{E_1}{E_2} = \sqrt{\frac{M_2}{M_1}}$$

Illustration: ***350 cm^3 of oxygen and 275 cm^3 of another gas ' A' diffused in same time under similar conditions. Find the molecular mass of the gas ' A' .***

Solutions: When two different volumes of gases have same time of diffusion, according to Graham' s law of diffusion.

$$\sqrt{\frac{M_{O_2}}{M_A}} = \frac{V_A}{V_{O_2}}$$

$$\sqrt{\frac{32}{M_A}} = \frac{275cm^3}{350cm^3}$$

Squaring both sides we get

(Or) $M_A = \frac{32 \times 350cm^3 \times 350cm^3}{275cm^3 \times 275cm^3} = 51.86$

(Or) Molecular mass of gas ' A' = 51.86 g. mol^{-1}.

Illustration: ***If 500 cm^3 of hydrogen diffused in 16 minutes through a fine hole, how much time does the same volume of ozone (O_3) take for diffusion?***

Solution: According to Graham' s law of diffusion, when same volume of different gases diffused in different times.

$$\sqrt{\frac{M_{O_3}}{M_{H_2}}} = \frac{t_{O_3}}{t_{H_2}}$$

$$\sqrt{\frac{48}{2}} = \frac{t_{O_3}}{t_{H_2}}$$

$$\sqrt{\frac{48}{2}} = \frac{t_{O_3}}{16\ min}$$

Or $t_{O_3} = \sqrt{24} \times 16\ min = 78.4\ min$

$\therefore$ Time of diffusion of Ozone = 78.4 min.

Kinetic Theory

In order to explain the experimental observations that have been summarized by the different gas laws, Rudolf Clausius, in 1857, put forward his kinetic theory of gases. The basic assumptions of the kinetic theory are as follows.

1. Gases consist of particles (molecules or atoms) that are in constant random motion.
2. Gas particles are constantly colliding with each other and the walls of their container. These collisions are elastic; that is, there is no net loss of energy from the collisions.
3. Gas particles are small and the total volume occupied by gas molecules is negligible relative to the total volume of their container.
4. There are no interactive forces (i.e., attraction or repulsion) between the particles of a gas.
5. The average kinetic energy of gas particles is proportional to the absolute temperature of the gas, and all gases at the same temperature have the same average kinetic energy.
6. The average kinetic energy of the molecule is proportional to the Kelvin temperature.