4

GAS LAWS AND MOLE CONCEPT

As in this picture, we can see huge balloons flying in the air in some tourist centres. Do you feel like rising up and flying in such balloons?

Have you ever thought about the reasons why these balloons rise up in the air, float and come down? Helium gas is filled in such balloons. Why do balloons filled with helium gas rise up in the air?

Is the density of helium gas filled in these balloons greater or lesser than that of the atmospheric air?

....

Tabulate the substances around us on the basis of their physical states.

Solid	Liquid	Gas

Table 4.1

Do you know, which of these states has the least density?

....

Find out the elements which exist in gaseous state at ordinary temperature (20°C-25°C) and write them down in the science diary.

....

From the following, tabulate the compounds which exist in gaseous state at ordinary temperature.

- Carbon dioxide, sodium chloride, ammonia, glucose, ammonium chloride, methane, sulphur dioxide, carbon monoxide, nitric acid, nitrogen dioxide.

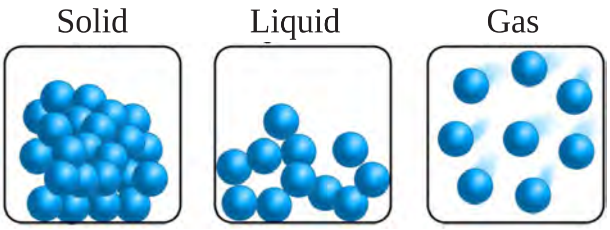


Figure 4.1

You have learnt about the arrangement of particles in different physical states.

Compare the distance between the molecules, force of attraction between them, freedom of movement of the molecules and their energy in gaseous state, with those in the other states of matter and write down.

Distance between the molecules	Very high
Force of attraction between the molecules	
Freedom of movement of the molecules	
Energy of the molecules	

Table 4.2

Scientists have conducted numerous studies on the general properties of gases. The postulates of kinetic molecular theory proposed by James Clerk Maxwell and Ludwig Boltzmann to explain these properties are given below.

- Gases are made up of minute particles (atoms/molecules).
- The attractive force between gaseous molecules is very low.
- As the molecules are so far apart, the volume of gaseous molecules is negligible in comparison with the total volume of the gas.
- The volume can be reduced by reducing the distance between the gaseous molecules.
- Gaseous molecules are in constant motion in all directions. As a result, the molecules collide with one another and with the walls of the container. The force produced due to the collision of molecules with the walls of the container results in gaseous pressure.

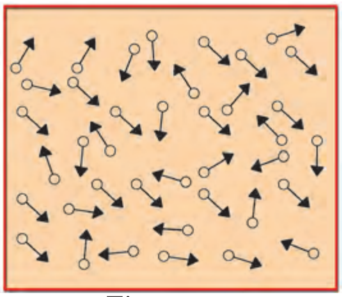


Figure 4.2

- Collisions of gaseous molecules are elastic in nature. That means, the kinetic energy of the molecules before and after the collision will be same.
- The average kinetic energy of the gaseous molecules is directly proportional to its temperature.

General properties of gases

The measurable properties of gases such as volume, pressure and temperature are explained below.

Volume

The space occupied by a substance is taken as its volume.

If a liquid filled in a bottle of volume 1L is transferred to a bottle of volume 2 L, what will be its volume?

....

If oxygen gas filled in a bottle of volume 1L is transferred to a bottle of volume 2 L, what will be its volume?

....

1 L Oxygen 2 L Oxygen

Figure 4.3

What if a gas of volume 2 L is transferred to a bottle of 10 L volume?

....

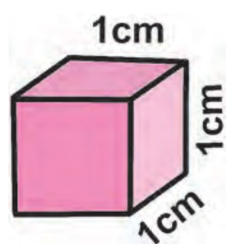
> The volume of a gas is the volume of the container in which it is occupied.

Let us examine the units used to state the volume of a gas.

Generally, the unit used is litre (L).

1 cm^3 is the volume of a container having length 1cm, breadth 1 cm and height 1 cm.

This is equal to 1mL.



$1\ cm^3 = 1\ cc$
(cc=cubic centimeter)

Figure 4.4

$$1000\ cm^3 = 1000\ mL = 1\ L$$

The SI unit of volume is m^3.

$1m^3$ is the volume of a container having length 1m, breadth 1m and height 1m.

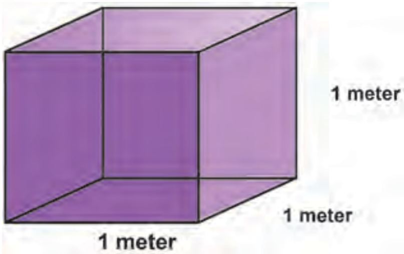


Figure 4.5

$$1\ m^3 = 1\ m \times 1\ m \times 1\ m$$
$$= 100\ cm \times 100\ cm \times 100\ cm = 1000000\ cm^3$$
$$= 1000000\ mL$$
$$= \frac{1000000}{1000}\ L\ [1000\ mL = 1L]$$
$$= 1000\ L$$

Now it is clear that $1\ m^3 = 1000\ L$.

Pressure

We have seen that gaseous molecules are in constant random motion. As a result, they collide with one another and with the walls of the container. Due to these collisions, a force is experienced on the surface. Gaseous pressure is the force experienced per unit surface area.

The pressure of gases is measured using manometer.

Different types of manometers.

$$\text{Pressure} = \frac{\text{Force}}{\text{Surface area}}$$

$$P = \frac{F}{A}$$

Generally, pressure is expressed in terms of atmospheric pressure (atm). You have learnt about atmospheric pressure in lower classes. The SI unit of pressure is Pascal (Pa).

(Pa = N/m^2 or Newton per square meter).

$$1\ atm = 1.01325 \times 10^5\ Pa.$$

Temperature

Temperature is another measurable property of a gas.

Which is the energy acquired by molecules due to its movement? (Kinetic energy / potential energy)

What happens to the kinetic energy of the molecules if a gas is heated? (Increases / decreases)

When gases are heated, the energy of molecules increases and hence temperature also increases.

The SI unit of temperature is Kelvin(K). To convert the common unit °C to Kelvin, add 273 to it.

If temperature is t°C,

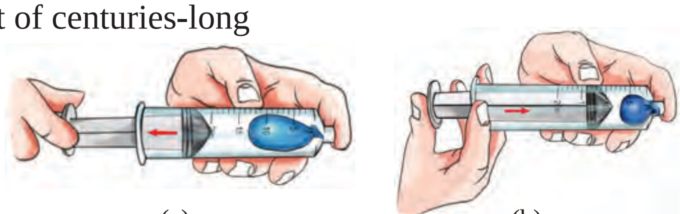
$$t°C = (t + 273)\ K$$

We can learn more about Kelvin scale in the following parts of this unit.

Gas Laws

The gas laws were established as a result of centuries-long experiments and observations on the physical properties of gases.

a) The relation between pressure and volume of a gas - Boyle's law



(a) (b)

Figure 4.6

Take a large syringe and remove its piston. Put into the syringe a small inflated balloon with its open end tied up. Now, refix the piston in its position (Fig. 4.6). Close the other end of the syringe with your finger. Record the changes happening to the balloon when the piston is pushed in and pulled back.

Activity	Observation
Piston is pushed in.	
Piston is pulled back.	

Table 4.3

What happens to the pressure inside the syringe when the piston is pushed in? (Increases / decreases)

What happens when the piston is pulled back?

....

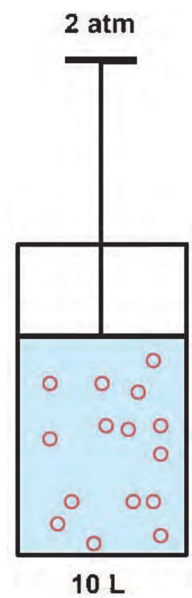


Figure 4.7 (a)

Now, let us look at another situation.

A definite mass of a gas taken in a cylinder with a freely movable piston is illustrated in the figure (Fig. 4.7(a))

Let us increase the pressure applied on this without changing the temperature (Fig. 4.7 b & c).

P (atm)	V (L)
2	10
4	5
10	2

Table 4.4

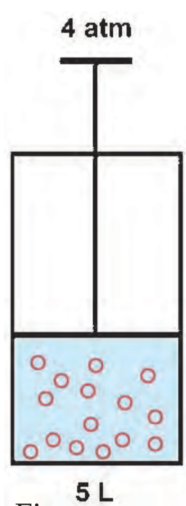


Figure 4.7 (b)

Figure 4.7 (c)

What happens to the volume when pressure is increased in each situation?

....

Robert Boyle
(1627-1691)

As pressure increases, volume decreases.

This relation was proposed by the scientist, Robert Boyle in 1662.

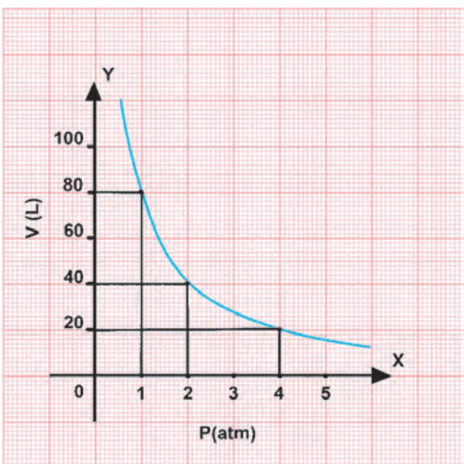
At constant temperature, the volume of a fixed mass of gas is inversely proportional to its pressure.

That is, $V \propto \frac{1}{P}$ (temperature, mass constant)

$V = \text{a constant} \times \frac{1}{P}$

$PV = k$ (k = a constant)

The value of k depends on the gas and its mass.



Pressure-volume (P - V) graph
Figure 4.8

If the volume of a fixed mass of gas is V_1 at pressure P_1 and V_2 at pressure P_2, Then

$$P_1V_1 = P_2V_2$$

To substantiate Boyle's law, we can illustrate the relation between volume and pressure of a fixed mass of gas, using a graph (Fig 4.8). From this graph, we can see that PV is a constant at constant temperature, for a definite mass of gas.

It can be seen that PV is a constant at all points in the graph.

Explain the following situations on the basis of Boyle's law.

a. The size of weather balloons goes on increasing as they rise above the sea level.

b. The size of the air bubbles rising up from the bottom of an aquarium increases as they reach the surface of the water.

- The volume of a definite mass of gas at a pressure 1 atm is 44 L. If the pressure is increased to 4 atm, what will be its volume?

According to Boyle's law,

$$P_1V_1 = P_2V_2$$
$$1 \times 44 = 4 \times V_2$$
$$V_2 = 44 /4= 11 \text{ L}$$

The volume of a definite mass of gas at 1 atm pressure is 1200 L. How much pressure is to be applied to change its volume to 30 L? (Temperature constant)

?

b) The relation between volume and temperature - Charles's Law

Let us do an experiment (Fig. 4.9). Make a small hole on the lid of a glass bottle and fix an empty refill into it. Add a small drop of ink into the refill, then blow gently to move the ink to the centre of the refill. Cover the bottle with your palms. What do you observe?

Figure 4.9

....

Upto which point of the arrangement does air occupy?

It is from the bottom of the bottle to that part of the refill that contains the ink.

What change occurs to the temperature of the air inside the bottle when covered the bottle with palms?

....

Does the volume of the gas increase or decrease during this period? ....

It is now clear why the ink rises up in the refill.

Place the bottle in water and cool it. What do you observe?

....

What happens to the volume of the gas when the bottle is cooled?

....

From this, can you find out the relation between the temperature of a gas and its volume?

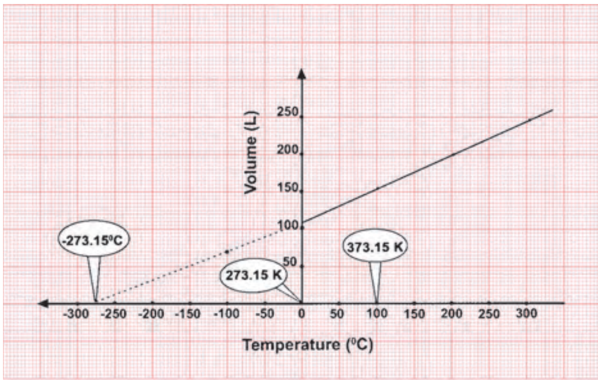
....

Here, you can see that the ink in the refill acts like a freely moving piston. When temperature of the gas increases, ink rises up in the refill and when temperature decreases, it comes down.

Jacques Alexandre Ceasre Charles (1746-1823)

Towards the end of the 18th century, it was Jacques Alexandre Ceasre Charles, a French scientist, who studied and recorded the changes in the volume of gases in accordance with temperature.

Plot the relation between temperature and volume of gases at constant pressure in a graph. If this graph is extrapolated backwards, you can see that it will meet the temperature axis (X axis) at –273.15°C (Fig 4.10).



Volume-temperature (V - T) graph

Figure 4.10

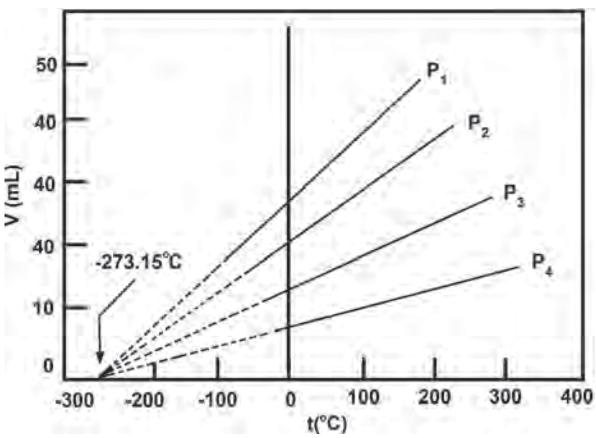
This means that at –273.15°C, the volume of a gas becomes zero.

Even if the pressure is changed, it is found that the temperature-volume graph is a straight line and all the straight lines meet at –273.15°C (Fig. 4.11).

The analysis of the graph (Fig. 4.11) is given in the table below.

	Kelvin scale	Celsius scale	Volume
Absolute zero	0 K	–273°C	0
Freezing point of water	273 K	0°C	115 L
Boiling point of water	373 K	100°C	150 L

Table 4.5



Volume-temperature (V - T) graph at different pressures

Figure 4.11

It was Lord Kelvin who identified that the lowest temperature that can be attained by a gas is –273.15°C and named it Absolute zero. For practical purposes, –273°C is taken as the value of Absolute zero. Using this, he developed the Kelvin scale of temperature.

According to Table 4.5, it is evident that as temperature increases, the volume of gas increases.

That is,

> At constant pressure, the volume of a definite mass of a gas is directly proportional to its temperature in Kelvin Scale. This is Charles's law.
>
> $V \alpha T$ (pressure, mass constant)
>
> $V = k \times T$ (k - a constant)
>
> $V/T = k$, a constant

If the volumes of a definite mass of gas are V_1 and V_2 at temperatures T_1 and T_2 respectively, then

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

- If the volume of a gas is 150 L at 27°C, what will be its volume at 0°C?

  27°C = 27 +273 = 300 K

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

  $$\frac{150}{300} = \frac{V_2}{273}$$

  $$V_2 = \frac{150 \times 273}{300} = 136.5 \text{ L}$$

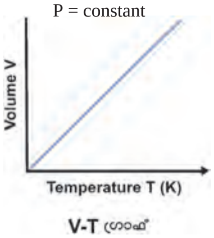
- The volume of a definite mass of hydrogen gas at 300 K temperature is 60 mL. At what temperature the volume of this gas becomes 20 mL?
- The relation between volume and temperature of a definite mass of a gas at constant pressure is illustrated in the two graphs given below.

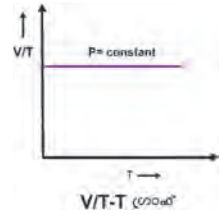
a) What is the relation between temperature and volume?

b) To which gas law do these graphs relate?

c) State the law.

d) What is the peculiarity of the value of V / T ?





?

Practical applications of Charles's law in everyday life.

- During summer, vehicle tyres are filled with air at lower pressure.

- Liquid ammonia is a substance that quickly changes from liquid to gaseous state. The containers of liquid ammonia are submerged in cold water for some time, before opening.

Avogadro's Law

Avogadro's law shows the relation between the number of particles in gases (N) and their volumes (V).

That is,

At constant temperature and pressure, equal volumes of all gases contain an equal number of molecules. In other words, at constant temperature and pressure, equal number of molecules of different gases occupy equal volumes.

At constant temperature and pressure, the volume of a gas is directly proportional to the number of molecules. This is Avogadro's Law.

$$V \propto N \text{ (Temperature, pressure constant)}$$

Applications of Avogadro's law in daily life.

- Inflating balloons.
- Filling air in footballs.

Combined gas equation

Let us consider Boyle's law and Charles's law which we have already discussed.

Boyle's law

$V \propto \frac{1}{P}$ (mass, temperature constant)

Charles's law

$V \propto T$ (pressure, mass constant)

Considering both laws together,

$$V \propto \frac{1}{P} \times T$$

$$V = \text{a constant} \times \frac{1}{P} \times T$$

$$\frac{PV}{T} = k \text{ (a constant)}$$

If the pressure, volume and temperature of a definite mass of a gas are changed from P_1, V_1 and T_1 to P_2, V_2 and T_2 respectively, then

$\frac{P_1 V_1}{T_1} = \frac{P_2 V_2}{T_2}$. This is combined gas equation.

The volume of a definite mass of gas at 1 atm pressure and 300 K temperature is 30 L. What will be its volume at 273 K temperature and 0.5 atm pressure?

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

$$\frac{1 \times 30}{300} = \frac{0.5 \times V_2}{273}$$

$$V_2 = \frac{30 \times 273}{300 \times 0.5} = 54.6 \text{ L}$$

The volume of a gas at 27°C and 1atm pressure is 100 mL. What will be its volume at 273 K temperature and 2 atm pressure?

Mole concept

In everyday life various units are used to count objects.

Eg.

2 numbers : pair	12 numbers : dozen
20 numbers : score	144 numbers : gross

Are these units sufficient to count extremely small particles like atoms, molecules and ions?

Even the presence of such tiny particles can be detected only with the help of powerful modern microscopes. Counting such minute particles is impossible.

Can you believe that a single drop of water contains approximately 10^{19} water molecules?

Can you imagine how huge this number is?

How many years it will take to finish counting this number?

Mole is the unit used to indicate such large numbers.

A mole is the quantity of a substance containing 6.022×10^{23} particles (atoms /molecules/ ions). This number came to be known as Avogadro number (N_A) named in honour of the Italian scientist Amedeo Avogadro.

Amedeo Avogadro (1776-1856)

Mole is the SI unit of quantity of matter. This represents the number of particles in that substance.

One mole of water contains 6.022×10^{23} water molecules.

The concept of mole is highly significant in Chemistry. It enables accurate measurement of the quantity of reactants and products that are to be used in chemical reactions.

Avogadro number

In 1908, Jean Perrin, a physicist, discovered that the number of particles in one mole is 6.7×10^{23}. This number came to be known as Avogadro number. The value of Avogadro number changes from time to time. A more accurate value of 6.022×10^{23}, which was found by De Bievre in 2001, is being considered now. The value of Avogadro's number continues to change based on advanced methods of accurate quantifications, new findings on atomic structure and the modified scientific definition of the concept of a mole.

Year	Scientist who discovered	Avogadro number
1908	Perrin	6.7×10^{23}
1917	Mullikan	6.064×10^{23}
1929	Birge	6.0644×10^{23}
1931	Bearden	6.019×10^{23}
1945	Birge	6.02338×10^{23}
1951	DuMond	6.02544×10^{23}
1965	Bearden	6.022088×10^{23}
1973	Cohen	6.022045×10^{23}
1987	Deslattes	6.022134×10^{23}
1994	Basile	6.0221379×10^{23}
2001	De Bievre	6.0221339×10^{23}

The changed values of Avogadro's number

Relative atomic mass and mole

Relative atomic mass expresses the mass of one atom relative to that of the another. It indicates how many times one atom is heavier as compared to another. The atomic mass of an element is expressed as, how many times is the mass of the atom, when compared to the $1/12^{th}$ mass of a carbon-12 atom, which is considered as a single unit. This mass is known as unified mass.

The atomic masses have decimal values because the average atomic mass is calculated by taking into account the presence of isotopes.

However, for practical purposes, they are often used as whole numbers.

Element	Average atomic mass	Relative atomic mass
Hydrogen	1.0079	1
Oxygen	15.9994	16
Sodium	22.989	23
Carbon	12.011	12
Nitrogen	14.0067	14

Table 4.6

Why atomic mass is a fraction.

The atomic mass of elements is calculated by considering the average mass of their isotopes based on their natural abundance.

For example, the natural abundance of neon is as, $^{20}Ne = 90.48\%$, $^{21}Ne = 0.27\%$, $^{22}Ne = 9.25\%$.

$$\text{Average atomic mass} = \frac{(20 \times 90.48) + (21 \times 0.27) + (22 \times 9.25)}{100}$$

$$= 20.18\ \text{u}$$

The natural abundance of Cl-35 isotope is 75% and that of Cl-37 is 25%.

$$\text{The average atomic mass of this element} = \frac{(35 \times 75) + (37 \times 25)}{100}$$

$$= 3550/100 = 35.5\ \text{u}$$

The atomic mass of most elements is a fraction, since average atomic mass is calculated in this way.

Analyse the table given below.

Element	Relative atomic mass	Gram atomic mass	Number of mole atoms	Number of atoms
Copper	63.5	63.5 g	1	6.022×10^{23}
Iron	55.8	55.8 g	1	6.022×10^{23}
Zinc	65.3	65.3 g	1	6.022×10^{23}
Aluminium	27	27 g	1	6.022×10^{23}
Nitrogen	14	14 g	1	6.022×10^{23}

Table 4.7

An element that weighs as much as its relative atomic mass in grams contains 6.022×10^{23} atoms.

Dalton

Dalton is a unit used to express the atomic mass. In 1993, IUPAC put forward a non-SI unit, Dalton which was equivalent to unified mass unit. This is represented by the symbol Da.

Eg. The atomic mass of Hydrogen - 1.008 Da

This mass is known as gram atomic mass. One gram atomic mass contains 1 mole atoms.

Why is the quantity of one mole atom of different substances different, even though they contain the same number of atoms?

The difference in quantity is due to the different sizes of the atoms, even though they are same in number.

Molar Mass

The total mass of atoms in a molecule is known as molecular mass.

Eg. Calculate the molecular mass of carbon dioxide (CO_2, atomic mass O = 16, C = 12).

Molecular mass $= 1 \times C + 2 \times O$

$= 1 \times 12 + 2 \times 16 = 12 + 32 = 44$

Calculate the molecular mass of molecules given in the table. Atomic masses of constituent elements are given in the bracket.

(Atomic mass - H = 1, O = 16, N = 14, C = 12, S = 32)

Element / Compound	Chemical formula	Molecular mass
Oxygen	O_2	$2 \times 16 = 32$
Ammonia	NH_3	$14 + 3 \times 1 = 17$
Water	H_2O	....
Glucose	$C_6H_{12}O_6$	....
Sulphuric acid	H_2SO_4	....
Nitrogen	N_2	....

Table 4.8

If 6.022×10^{23} oxygen molecules are taken, the mass will be 32 g. This is the molecular mass of oxygen expressed in grams. This is known as gram molecular mass.

What will be the mass of 6.022×10^{23} CO_2 molecules?......... g

This is the molar mass of the compound.

One molar mass of a compound contains one mole molecules.

Complete the following table.

(Atomic mass - H = 1, O = 16, N = 14, C = 12, S = 32, Ca = 40)

Compound	Molecular mass	Molar mass	Number of moles	Number of Molecules
NH_3	17	17g	1	6.022×10^{23}
CO_2	....	....	....	6.022×10^{23}
H_2O	....	....	1	....
NO_2	....	....	....	6.022×10^{23}
$CaCO_3$	....	....	....	....

Table 4.9

How many moles are there in 44g of CO_2?

....

Can you find out the number of moles in 88 g CO_2 ?

44 g = 1 mole

88 g = $\frac{88}{44}$ = 2 mole

That is, number of moles = $\frac{\text{Given mass}}{\text{Molar mass}}$

- What is the number of molecules in this sample of CO_2?

Number of molecules in 1 mole of CO_2 = 6.022×10^{23}

Number of molecules in 2 moles of CO_2 = $2 \times 6.022 \times 10^{23}$

Complete the following table.

(Atomic mass - H = 1, O = 16, N = 14, C = 12, S = 32, Ca = 40, Na = 23, Cl = 35.5)

Substance	Molar Mass	Given mass	Number of moles	Number of molecules
NH_3	....	85 g	....	.... $\times 6.022 \times 10^{23}$
CO_2	....	220 g	....	....
H_2O	....	....	10	$10 \times 6.022 \times 10^{23}$
NO_2	....	92 g	....	....
$C_6H_{12}O_6$	....	360 g	....	....
NaCl	58.5 g	1170 g	....	....

Table 4.10

Volume of gases and mole

According to Avogadro's law, at constant temperature and pressure equal volumes of all gases contain equal number of molecules.

What is the peculiarity of the volume occupied by an equal number of molecules of different gases kept at same temperature and pressure?

....

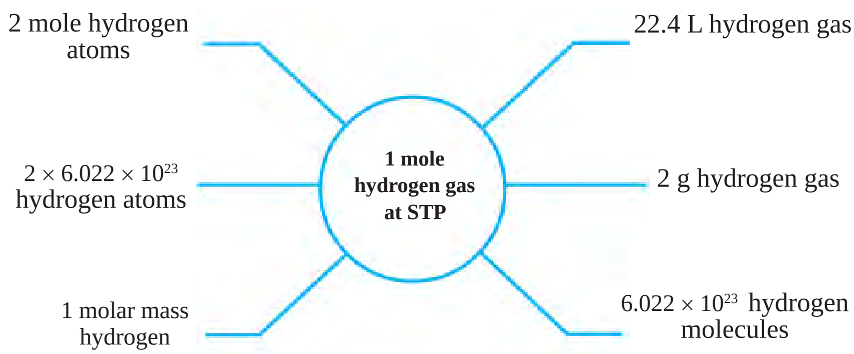
If temperature and pressure are fixed at 273 K and 1 atm respectively, it is known as STP (Standard Temperature and Pressure). At STP, one mole of any gas occupies a volume of 22.4 L. This is known as molar volume at STP.

$$\text{Number of moles of gases at STP} = \frac{\text{Given volume (in litres)}}{\text{Molar volume at STP}}$$

Complete the table given below.

Gas	Molar mass	Given Mass	Number of moles	Volume (L)
O_2	......	32 g	....	22.4
NH_3	....	170 g	....	....
CO_2	....	....	3	....
NO_2	....	....	2	....

Table 4.11



Ideal gas equation

Considering Boyle's law, Charles's law and Avogadro's law,

$V \ \alpha \ \frac{1}{P}$ (T, n constant)

$V \ \alpha \ T$ (P, n constant)

$V \propto n$ (n = number of moles), (P, T constant)

$V \propto \frac{1}{P} \times T \times n$

PV = a constant × nT

This constant is known as Universal Gas Constant. This is represented by the letter R.

PV = nRT

This is ideal gas equation.

The gases which obey ideal gas equation at all temperature and pressure are known as ideal gases.

1. 224 L O_2 (oxygen) gas is taken at STP.
   a) How many moles of oxygen are there in it?
   b) How many molecules are there in it?
   c) Calculate the mass of this oxygen sample. (Hint: atomic mass O = 16)
2. Find the samples with equal volumes from the ones given below at STP.
   a) 64 g O_2
   b) 44 g CO_2
   c) $2 \times 6.022 \times 10^{23}$ NH_3 molecules

Mole concept and chemical equations

You have already learnt to balance equations.

Have a look at the given equation.

$$2H_2 + O_2 \rightarrow 2H_2O$$

Here, it can be seen that two hydrogen molecules combine with one oxygen molecule to form two water molecules.

If we take two moles of hydrogen molecules instead of two hydrogen molecules?

$$2 \text{ mol } H_2 + 1 \text{ mol } O_2 \rightarrow 2 \text{ mol } H_2O$$

How can the equation be written in terms of mass?

What is the mass of 2 mol hydrogen?

$2 \times 2 = 4$ g

What is the mass of 1 mol oxygen? ....

You have already found that the mass of 2 mol H_2O is 36 g.

$$4\text{ g } H_2 + 32\text{ g } O_2 \rightarrow 36\text{ g } H_2O$$

According to this equation, when 4 g hydrogen completely reacts with oxygen, 36 g water is obtained.

How many grams of water will be obtained if 40 g hydrogen reacts completely with oxygen?

Water obtained when 4 g H_2 reacts completely = 36 g

Water obtained when 1 g H_2 reacts completely = 36/4= 9 g

Water obtained when 40 g H_2 reacts completely = 9×40 =360 g

How much oxygen is required for this much hydrogen to react completely and form water?

$$4\text{ g } H_2 + 32\text{ g } O_2 \rightarrow 36\text{ g } H_2O$$

The oxygen required for 4 g hydrogen to react completely = 32 g

The oxygen required for 1 g hydrogen to react completely = 32/4 = 8 g

The oxygen required for 40 g hydrogen to react completely
= 8 x 40 = 320 g

Let us analyse another situation.

The reaction in which nitrogen and hydrogen react to give ammonia is of high significance in the field of agriculture.

$$N_2 + 3H_2 \rightarrow 2NH_3$$

If this equation is analysed on the basis of mole concept,

1 mol N_2 + 3 mol H_2 $\rightarrow$ 2 mol NH_3

On the basis of mass, it is

28 g N_2 + 6 g H_2 $\rightarrow$ 34 g NH_3

If we know how much ammonia is to be produced, we can determine the amount of reactants to be used?

- How many moles of nitrogen and hydrogen are required to produce 6 mol ammonia?

  1 mol N_2 + 3 mol H_2 $\rightarrow$ 2 mol NH_3

What is the ratio of the reactants?

....

The amount of nitrogen required to produce 2 mole ammonia - 1 mol

The amount of nitrogen required to produce 1 mole ammonia - ½ mol

The amount of nitrogen required to produce 6 mole ammonia - ½ × 6 = 3 mol

The amount of hydrogen required to produce 2 mole ammonia - 3 mol

The amount of hydrogen to be reacted for producing 1 mole ammonia - 3/2 mol

The amount of hydrogen to be reacted for producing 6 mole ammonia - 3/2 × 6 = 9 mol.

- 3 moles of nitrogen should react with 9 moles of hydrogen to form 6 moles of ammonia . If the amount of any of these reactants is less, the product will be formed proportionately.
- What is the quantity of ammonia produced when 6 moles of nitrogen react with 6 moles of hydrogen?

Only 2 moles of nitrogen react with 6 moles of hydrogen. (N : H = 1 : 3)

2 moles of nitrogen + 6 moles of hydrogen $\rightarrow$ 4 moles of ammonia

How many moles of nitrogen will be left after the reaction?

.... mol.

The main component of biogas is methane (CH_4). Note the equation of its combustion.

$CH_4 + 2O_2 \rightarrow CO_2 + 2H_2O$

On the basis of moles,

.... mol CH_4 + ... mol O_2 $\rightarrow$ .... mol CO_2 + ... mol H_2O

On the basis of mass,

.... g CH_4 + ... g O_2 $\rightarrow$ .... g CO_2 + ... g H_2O

What is the mass of oxygen required to react with 16 g of methane?

....

Find the volume of CO_2 produced by the complete combustion of 16 g of methane?

....

Thus, it can be seen that the amount of CO_2 released when each gram of methane is burned is about three times the mass of methane. Butane (C_4H_{10}) present in cooking gas also emits about three times CO_2 when burned.

All of these are fossil fuels. Now, do you understand why it is advised to control the use of fossil fuels?

Compounds such as lime and baking soda are used to neutralise acid wastes from industries. The above examples will help you to find out the amount of compounds to be used for this purpose.

Eg. $H_2SO_4 + Ca(OH)_2 \rightarrow CaSO_4 + 2H_2O$

$H_2SO_4 + 2NaHCO_3 \rightarrow Na_2SO_4 + 2H_2O + 2CO_2$

The amount of lime and baking soda used by factories can be found out from these equations.

- Calculate the amount of lime in kg required to neutralise 980 kg of sulphuric acid.

Let us assess

1. Convert the units of the given temperatures.

°C	K
0	273
100	373
30	......
....	300
40	....

2. Complete the table. (Atomic Mass - H = 1, C = 12, O = 16, N = 14)

Substance	Molecular Mass	Given mass	Number of moles	Volume at STP
H_2	....	10 g	....	112 L
CO_2	....	440 g	....	....
NH_3	....	340 g	....	....

3. A balanced equation of the reaction to produce ammonia is given.

$$N_2 + 3H_2 \rightarrow 2NH_3$$

a) How many moles of hydrogen is required for 10 moles of nitrogen to react completely?

b) How much ammonia will be produced if 10 moles of nitrogen react completely?

4. 448 L of gas is stored in a cylinder at 0°C and 1 atm pressure.

a) How many moles of molecules does this gas contain?

b) What is the number of molecules in this sample?

5. 400 L of gas is stored in a cylinder at 27 °C and constant pressure.

a) What will be the temperature if the volume of this gas is reduced to 200 L at the same pressure?

b) Which gas law is relevant to this context?

c) The boiling point of a substance is 3°C. Above what temperature in Kelvin does this substance obey the gas laws?

6. a) What is the number of Cl_2 molecules in 710 g chlorine gas?

b) What is the total number of atoms in this sample?

7. How many grams of hydrogen is required to react with 700 g nitrogen in the production of ammonia?
8. Calculate the following.
   a) How many moles of $CaCO_3$ are present in its 1 kg?
   b) Mass of the same number of NH_3 molecules as contained in 88 g of CO_2.
   c) Mass of 22.4 L CO_2 kept at STP.
   (Hint: Atomic mass - H = 1, C = 12, O = 16, N = 14, Ca = 40)
9. The volume of a cylinder containing NH_3 at STP is 4480 mL.
   a) Find the number of moles of NH_3 in the cylinder.
   b) Find the mass of this NH_3 gas.
   c) How many molecules of NH_3 are there in this gas cylinder??
10. Consider the chemical reaction

$$2H_2 + O_2 \rightarrow 2H_2O$$

   a) How many moles of oxygen (O_2) should react to obtain 10 mole H_2O?
   b) Find the mass of Oxygen gas (O_2) required to obtain 10 mole H_2O?
   (Atomic mass - H = 1, O = 16)

Extended activities

1. 448 L HCl gas kept at STP completely reacts with ammonia gas at STP to form NH_4Cl.

$$NH_3 + HCl \rightarrow NH_4Cl$$

   Find the mass of ammonia gas used in this reaction.

2. 48 g C burns in air and carbon monoxide gas is formed. If the mass of CO gas formed is 56 g, what is the volume of oxygen gas at STP used in this reaction?