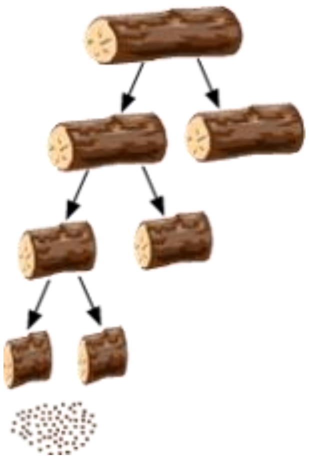
# Atoms and Molecules

## Dalton's Atomic Theory

### Dividing Matter



**Matter cannot be divided infinite number of times.**

For example, if we keep chopping a log of wood into smaller and smaller pieces, then we will reach a point when the wood will not be divisible any further. Minute particles of wood will remain and these will not be visible to the naked eye. This is true for all forms of matter.The same was believed by the early Indian and Greek philosophers. In India, around 500 BC, an Indian philosopher named Maharishi Kanad called matter as padarth and these smallest particles (atoms) as 'parmanu'. The word 'atom' is derived from the Greek word 'atomos' which means 'indivisible'. It was the Greek philosopher Democritus who coined the term. However, for these ancient thinkers, the idea of the minute indivisible particle was a purely philosophical consideration.

By the end of the eighteenth century, scientists had begun to distinguish between elements and compounds. Two French chemists named Antoine Lavoisier and Joseph Proust observed that elements combine in definite proportions to form compounds. On the basis of this observation, each of them proposed an important law of chemical combination. The laws proposed by them helped Dalton formulate his atomic theory.

### Dalton's Atomic Theory

In the early nineteenth century, an English chemist named John Dalton proposed a theory about **atoms**. Known as 'Dalton's atomic theory', it proved to be one of the most important theories of science. The various laws of chemical combination also supported Dalton's theory. Dalton asserted that 'atoms are the smallest particles of matter, which cannot be divided further'. He published his atomic theory in 1808 in his book *A New System of Chemical Philosophy*. The postulates of Dalton's atomic theory are as follows:

- All matter is made up of very tiny particles. These particles are called atoms.
- An atom cannot be divided further, i.e., atoms are indivisible.
- Atoms can be neither created nor destroyed in a chemical reaction.
- All atoms of an **element** are identical in all respects, e.g. in terms of mass, chemical properties, etc.
- Atoms of different elements have different masses and chemical properties.
- Atoms of different elements combine in small whole-number ratios to form **compounds**.
- In a given compound, the relative numbers and types of atoms are constant.

**Know Your Scientist**

**John Dalton (1766–1844)** was born into the poor family of a weaver in Eaglesfield, England. He was colour-blind from childhood. He became a teacher when he was barely twelve years old. By the time he was nineteen, he had become the principal of a school. In 1793, Dalton left for Manchester to teach physics, chemistry and mathematics at a college. Elected a member of the Manchester Literary and Philosophy Society in 1794, he became its president in 1817 and remained in that position until his death. During his early career, he identified the hereditary nature of red–green colour blindness. In 1803, he postulated the law of partial pressures (known as Dalton's law of partial pressures). He was the first scientist to explain the behaviour of atoms in terms of relative atomic weight. He also proposed symbolic notations for various elements.

**Law of Conservation of Mass**

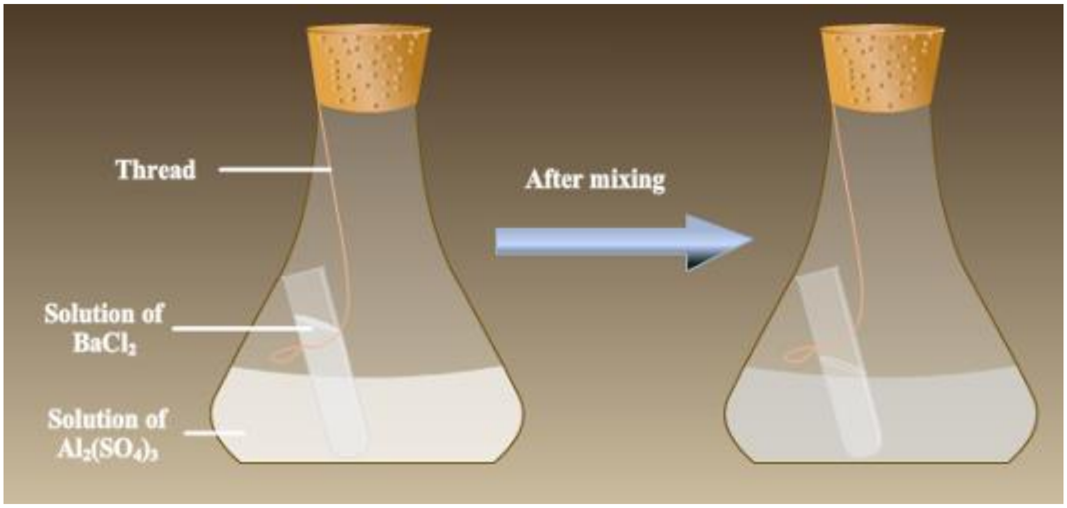
**Law of Chemical Combination**

Lavoisier and Proust proposed two laws that sought to explain the chemical combinations of elements. These laws are the law of conservation of mass and the law of constant proportions, respectively.

Let us first study about the law of conservation of mass.

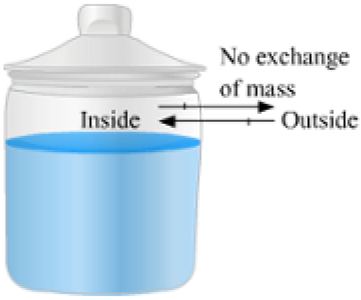
According to this law, '**mass can be neither created nor destroyed.**' In other words, for a chemical reaction taking place in a closed system, the total mass of the reactants is the same as that of the products.

For example, an unmixed solution of barium chloride and aluminium sulphate weighs the same as that of the mixed solution of the two.




**Know More**

**Closed System**




A system is said to be closed if there is no exchange of mass across the boundaries of the system and the surroundings. However, it can exchange energy with its surroundings.

**Solved Examples**

**Medium**

**Example 1**:

**In a decomposition reaction**

**, 100 g of mercuric oxide—when heated in a closed test tube—decomposes to produce mercury and oxygen gas. If the mass of the produced mercury is 92.6 g, then what is the mass of the produced oxygen?**

**Solution**:

According to the law of conservation of mass:

Total mass of the reactants = Total mass of the products

It is given that:

Mass of the decomposing mercuric oxide = 100 g

Mass of the produced mercury = 92.6 g

Let the mass of the produced oxygen be $x$.

So, we have:

$100\ g = 92.6\ g + x$

$\Rightarrow x = (100 - 92.6)\ g$

$\Rightarrow \therefore x = 7.4\ g$

**Laws of Chemical Combination and Dalton's Atomic Theory**

**The law of conservation of mass can be explained using the first and third postulates of Dalton's atomic theory.**

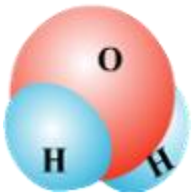
The law of conservation of mass states that mass can be neither created nor destroyed. According to this law, in case of a chemical reaction taking place in a closed system, the total mass of the reactants equals that of the products.

Mass is the amount of matter in something. As per the first postulate of Dalton's atomic theory, all matter is made up of atoms. The third postulate of the same theory asserts that atoms can be neither created nor destroyed in a chemical reaction, i.e., the total number of atoms and their mass should remain the same before and after the reaction. This is the same as the law of conservation of mass.

**Law of Constant Proportions**

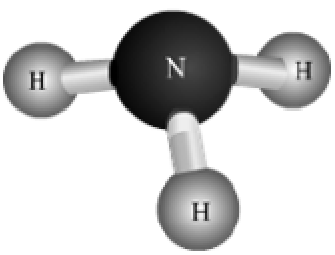
This law is also known as the **law of definite proportions**. According to this law, '**in a compound, elements are always present in definite proportions by masses.**' We know that compounds are composed of two or more elements. So, according to this law, the proportions in which elements are present in a compound remain the same, irrespective of its method of preparation. Let us understand this law with the help of a couple of examples.

Water Molecule




Pure water obtained from any source (well, river, lake or sea) and from any country (India, Russia or America) will always contain two hydrogen atoms and one oxygen atom. The atoms of hydrogen combine with the atom of oxygen in the ratio of 1 : 8 by mass to form water. The ratio by the number of atoms for water will always be 2 : 1. The mass of a water molecule is 18 g. So, a molecule of water contains 2 g of hydrogen and 16 g of oxygen.

Ammonia Molecule




Ammonia contains one nitrogen atom and three hydrogen atoms. Irrespective of the source from which ammonia is obtained, it will always contain nitrogen and hydrogen in the ratio of 14 : 3 by mass. The mass of an ammonia molecule is 17 g. So, a molecule of ammonia contains 14 g of nitrogen and 3 g of hydrogen. Similarly, 34 g of ammonia contains 28 g of nitrogen and 6 g of hydrogen.

**Solved Examples**

**Medium**

**Example 2**:

**Hydrogen and oxygen combine in the ratio of 1 : 8 by mass to form water. How much mass of hydrogen gas will be required to react completely with 24 g of oxygen gas?**

**Solution**:

We know that hydrogen and oxygen combine in the fixed ratio of 1 : 8 by mass to form water.

Thus, it can be said that 8 g of oxygen gas reacts completely with 1 g of hydrogen gas.

Therefore, 24 g of oxygen will react completely with $\frac{1}{8} \times 24 = 3\text{ g}$ of hydrogen gas.

**Laws of Chemical Combination and Dalton's Atomic Theory**

**The law of constant proportions follows directly from the sixth and seventh postulates of Dalton's atomic theory.**

The sixth postulate of Dalton's atomic theory states that atoms of different elements combine in small whole-number ratios to form compounds. The seventh postulate states that the relative numbers and types of atoms in a compound are constant. This has the same meaning as the law of constant proportions, which states that the elements in a compound are always present in definite proportions by masses.

Now, we know that a sample of carbon dioxide (no matter how it is prepared) is made up of carbon and oxygen. A molecule of carbon dioxide comprises 1 carbon atom and 2 oxygen atoms. The mass of a carbon dioxide molecule is 44 g. The relative atomic masses of carbon and oxygen are 12 u and 16 u respectively. As per the law of constant proportions, in carbon dioxide, carbon and oxygen always combine in the ratio of 3 : 8 by mass.

**Atoms**

**Atoms: An Overview**

When we talk about atoms, two questions usually strike our mind...


Let us go through this lesson to find the answers to these questions. We will also learn how to represent different atoms in symbolic forms. So, in short, we are going to study:

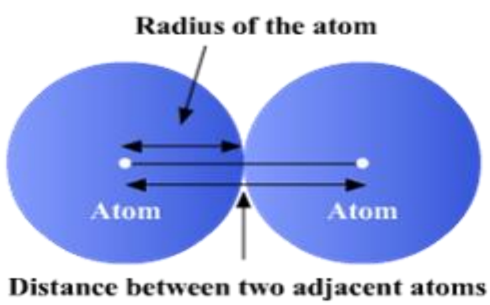
- **Size of an atom**
- **Representation of atoms**
- **Atomic mass**

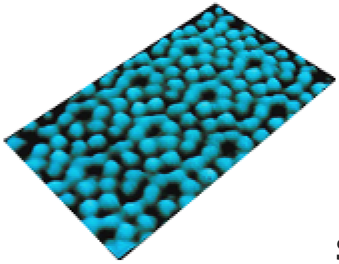
**Size of an Atom**

- The size of an isolated atom cannot be measured; however, we can estimate its size by assuming that its radius is half the distance between adjacent atoms in a solid.
- Atoms are very small in size. They are so small that it is not possible to see them even under a powerful optical microscope.

.The size of an atom is indicated by its radius, called the **atomic radius**
.Since an atom is very small, we need a very small unit for reporting the atomic radius; thus, the radius of an atom is often expressed in **nanometre**

.



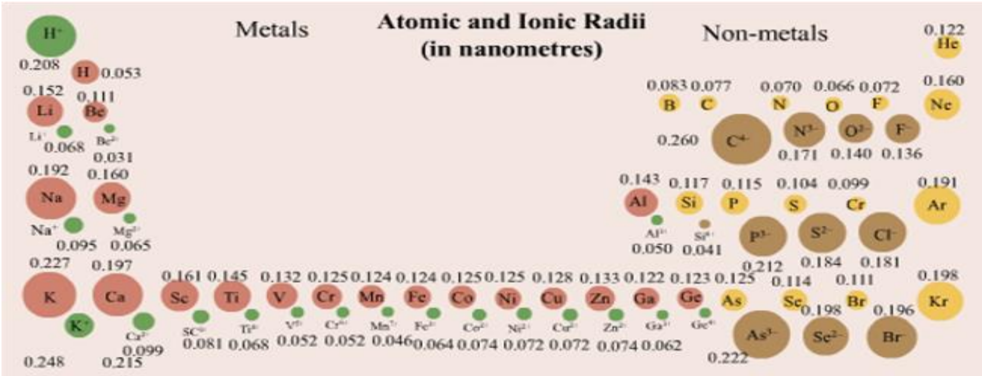


**Surfaces of Silicon Atoms**

Atoms cannot be seen with the naked eye, but the use of modern techniques has enabled us to see the surfaces of atoms. The magnified image of the surfaces of silicon atoms is shown in the following figure.

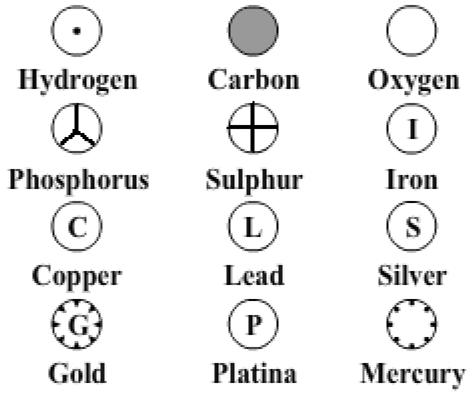
### Size of an Atom

Hydrogen atom is the smallest of all atoms. The given figure shows the atomic radii of some elements.

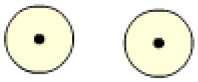



### Classical Representation of Atoms

- A large number of elements are known to us today. It would be cumbersome to refer to them by their names all the time in our studies. For the sake of convenience, we need symbols that represent these elements. Toward the end of the nineteenth century, scientists felt this need to assign standard characteristic symbols to the elements.
- John Dalton was the first scientist to use symbols to represent different elements. Dalton's proposed symbols for some elements are shown in given figure.




- Each symbol proposed by Dalton represents an atom of the respective element. For example, if someone wanted to represent two hydrogen atoms, then he would have to draw the symbol of hydrogen atom twice as shown.


Two Hydrogen atoms

**Modern Representation of Atoms**

- Many of the symbols proposed by Dalton were difficult to draw and remember. Therefore, an alternative method of representing elements was required.
- Another scientist named Jöns Jacob Berzelius suggested that letters of the alphabet can be used as symbols to represent the elements. The modern symbols of elements are based on this idea.
- The International Union of Pure and Applied Chemistry (**IUPAC**) approves the names and symbols for the elements.
- The modern symbol of an element is made up of one or two letters of the English or Latin name of that element.
- As a rule, the first letter of a symbol is always written as a capital letter and the second as a small letter.
- The modern representation of atoms is more convenient and meaningful than the classical representation.
- To conclude, the symbols of the various elements are significant as:
- They represent distinct elements.
- They represent single atoms of the elements.

**Modern Representation of Atoms**

Table mentioned below shows the modern representation of atoms:

| **Elements** | **Symbols** | **Elements** | **Symbols** |
|---|---|---|---|
| Aluminium | Al | Iron (from Latin: ferrum) | Fe |
| Argon | Ar | Lead (from Latin: plumbum) | Pb |
| Calcium | Ca | Magnesium | Mg |
| Carbon | C | Nitrogen | N |
| Chlorine | Cl | Oxygen | O |
| Copper (from Latin: cuprum) | Cu | Potassium (from Latin: kalium) | K |
| Fluorine | F | Silicon | Si |
| Gold (from Latin: aurum) | Au | Silver (from Latin: argentum) | Ag |
| Hydrogen | H | Sodium (from Latin: natrium) | Na |
| Iodine | I | Zinc | Zn |

**Did You Know?**

**IUPAC**
**The International Union of Pure and Applied Chemistry (IUPAC)** is an international non-governmental organisation established in 1919.
IUPAC nomenclature, in general, is a system of naming chemical compounds and describing the science of chemistry.
Some man-made elements are given temporary three-letter symbols.

Atomic Mass

- Every atom has some characteristic mass of its own and this **is known as atomic mass.**
- All the atoms of an element have the same atomic mass. Atoms of different elements have different atomic masses.
- **Determination of atomic mass:**

It is difficult to determine the mass of an individual atom; so, the mass of an atom is ascertained in relation to the mass of C-12 **isotope**

. Thus, atomic mass is in fact relative atomic mass.

- C-12 is an isotope of carbon and its mass is used as a standard reference to calculate the relative atomic masses of all elements.

**The IUPAC** adopted one-twelfth $\left(\frac{1}{12}\right)^{th}$ of the mass of a C-12 isotope as the standard unit to measure relative atomic masses. It named this unit as **atomic mass unit**

(amu) or **unified atomic mass unit** (u).

So,

$$\text{Atomic mass unit} = \frac{1}{12} \times \text{Mass of C-12 atom}$$

In simple words, Atomic mass is a term which gives the total mass of protons and neutrons in an atom. Also, the atomic mass is measured with respect to mass of 1/12 the mass of one carbon atom.

**Atomic Mass**

The atomic masses of some common elements are given in the following table.

| Elements | Atomic masses (u) | Elements | Atomic masses (u) |
|---|---|---|---|
| Hydrogen | 1 | Chlorine | 35.5 |
| Helium | 4 | Potassium | 39 |
| Carbon | 12 | Calcium | 40 |
| Nitrogen | 14 | Argon | 40 |
| Oxygen | 16 | Iron | 56 |
| Fluorine | 19 | Copper | 63.5 |
| Neon | 20 | Zinc | 65 |
| Sodium | 23 | Bromine | 80 |

| | | | |
|---|---|---|---|
| Magnesium | 24 | Silver | 108 |
| Sulphur | 32 | Gold | 197 |

**Atomic Mass Unit**

Atomic mass unit (1 u) is defined as exactly one twelfth the mass of an atom of carbon-12. Atomic mass unit is only number and it has no units. On the basis of above unit, the atomic mass of carbon atom is 12 amu.

**Relative atomic mass or atomic weight**

Definition with respect to hydrogen:

It is the ratio of mass of one atom of an element to the mass of an atom of hydrogen taken as unity.

Definition with respect to carbon:

It is the ratio of mass of one atom of an element to 1/12th mass of an atom of carbon.

**Relative molecular mass or molecular weight**

Definition with respect to hydrogen:

It is the ratio of mass of one molecule of a substance to the mass of an atom of hydrogen taken as unity.

Definition with respect to carbon:

It is the ratio of mass of one atom of an element to 1/12th mass of an atom of carbon.

**Gram Molecular Volume**

The volume occupied by 1 gram molecule of a dry gas at S.T.P is called gram molecular volume. The experimental value of 1 gram molecular volume of a gas is 22.4 litres at S.T.P.

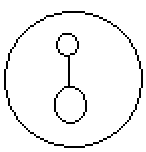
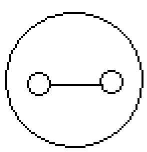
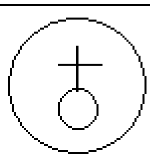
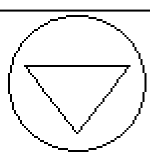
The relationship between the mole, Avogadro's number, and mass is summarised as follows:

1 mole of any particle (atom, molecule, ion) → $6.022 \times 10^{23}$ number of atoms/molecules/ions → Relative mass of those particles in grams

1 mole of molecules → $6.022 \times 10^{23}$ number of molecules → Molecular mass in grams

1 mole of atoms → $6.022 \times 10^{23}$ number of atoms → Atomic mass in grams

Symbols are very important in chemistry as they are used to represent different elements. The usage of symbols to represent elements has been in trend from the ancient Greek time. Greeks used symbols to represent the four elements, earth, air, fire and water.
In the era of alchemists, different materials were represented using pictorial symbols.
For example,

| **Element** | **Symbol** |
|---|---|
| Nickel |  |
| Arsenic |  |
| Antimony |  |
| Water |  |

**Do You Know?**
The process of converting a less valuable metal into a more valuable metal like gold is called **alchemy**, and the men who started this process are called **alchemists**.

**System for determining symbols for different elements:**
Certain rules have been framed to determine symbols of different elements. They are as follows:
1. The symbols of common elements, mainly non-metals, are the first letter of their respective names.
For example, the symbols of oxygen and fluorine are O and F respectively.

2. If the name of an element shares the initial letter with another element, then the first and the second letter of its name are used as the symbol.
For example, symbols used for **B**arium and **B**eryllium are Ba and Be respectively.

3. If the first two letters of the names are the same, then the first and third lettes are used as the symbol.
For example, symbols used for **M**agnesium and **M**anganese are Mg and Mn respectively.

4. Symbols of some elements are based on their old or Latin names. There are 11 elements whose symbols are derived from their Latin names.
For example, symbol of sodium is Na, which is derived from its Latin name Natrium.

5. If the symbol of any element is a single letter, it should be written in capital.

6. If the symbol of any element has two letters then the first one should be in capital followed by small letter.

**Significance of symbols:**
1. Symbol of an element signifies the name of the element.
2. It also signifies that one atom of that element is present.

**Molecules and Ions**

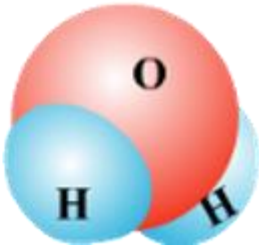
**A Brief Introduction to Molecules and Ions**

Most atoms are not stable in free state. So, they combine with other atoms to form molecules.

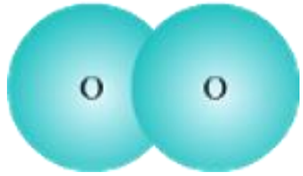
For example:

A water molecule is formed when two hydrogen atoms combine with one oxygen atom. An oxygen molecule is formed when two oxygen atoms combine with each other.

Water Molecule




Oxygen Molecule




Some atoms are charged. Such charged atoms and molecules are called ions.

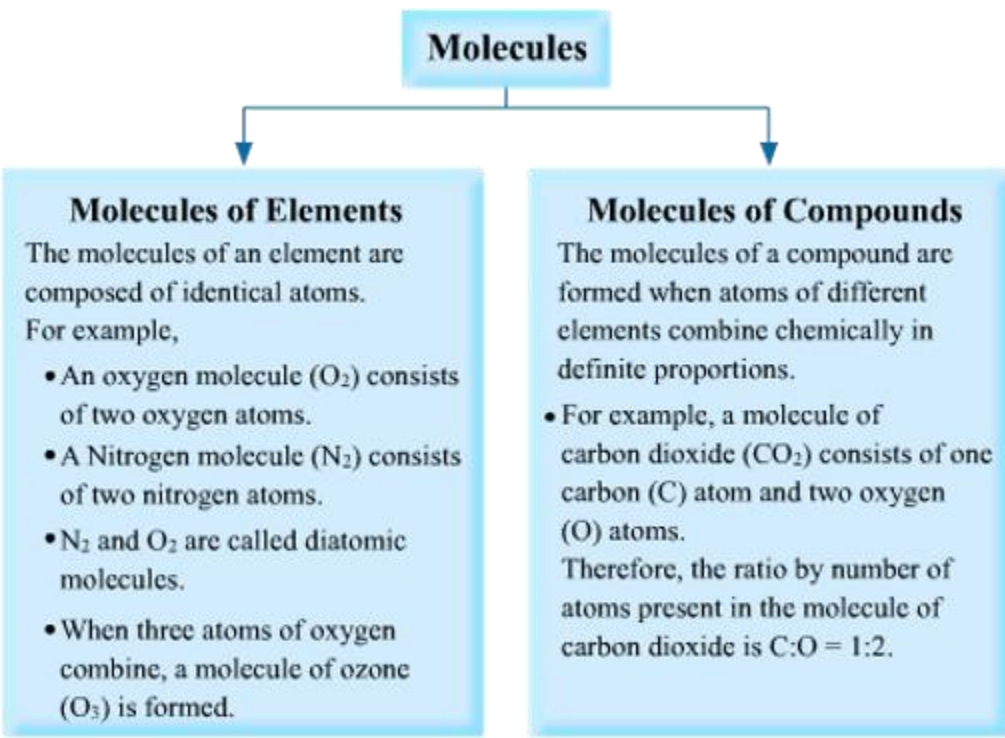
A positively charged ion is called **cation**.

A negatively charged ion is called **anion**.

In this lesson, we are going to study about:

- Molecules and molecular compounds
- Ions and ionic compounds

**Molecules**



**Did You Know?**

The term '**molecule**' originates from the French word '*molécule*', which means 'extremely minute particle'. It was coined by the French philosopher and mathematician Rene Descartes in the early seventeenth century.

In view of John Dalton's laws of definite and multiple proportions, the existence of molecules was accepted by many chemists since the early nineteenth century. However, it is the work of Jean Baptiste Perrin on the Brownian motion (1911) of particles of liquids and gases which is considered to be the final proof of the existence of molecules.

**Atomicity of Molecules**

**The number of atoms constituting a molecule is known as its atomicity. The given table lists the atomicity of some common elements.**

| **Elements** | **Atomicity** |
|---|---|
| Helium (He), Neon (Ne), Argon (Ar) | Monoatomic (1 atom per molecule) |
| Oxygen (O2), Hydrogen (H2), Nitrogen (N2) Chlorine (Cl2), Fluorine (F2) | Diatomic (2 atoms per molecule) |
| Phosphorus (P4) | Tetratomic (4 atoms per molecule) |
| Sulphur (S8) | Polyatomic (8 atoms per molecule) |

**Did You Know?**

**Buckminsterfullerene** is an allotrope of carbon in which sixty carbon atoms are bonded together.

**Ions**

An ion is a charged atom or molecule. This charge arises because the number of electrons do not equal the number of protons in the atom or molecule. An ion is also known as **radical**. A positively charged ion is called **cation**; they are also called basic radicals. While a negatively charged ion is called **anion**. Such ions are called acid radicals.

There are many ions which are **polyatomic ions**

.

The given table lists the symbols and atomicity of some common ions.

| Cations | Symbols | Atomicity | Anions | Symbols | Atomicity |
|---|---|---|---|---|---|
| Aluminium | $Al^{3+}$ | Monoatomic | Bromide | $Br^-$ | Monoatomic |
| Ammonium | $NH_4^+$ | Polyatomic | Carbonate | $CO_3^{2-}$ | Tetra-atomic |
| Calcium | $Ca^{2+}$ | Monoatomic | Chloride | $Cl^-$ | Monoatomic |
| Cuprous ion | $Cu^+$ | Monoatomic | Fluoride | $F^-$ | Monoatomic |
| Cupric ion | $Cu^{2+}$ | Monoatomic | Hydride | $H^-$ | Monoatomic |
| Hydrogen | $H^+$ | Monoatomic | Hydroxide | $OH^-$ | Diatomic |

**Ions**

The given table lists the symbols and atomicity of some other common ions.

| Cations | Symbols | Atomicity | Anions | Symbols | Atomicity |
|---|---|---|---|---|---|
| Ferric ion | $Fe^{3+}$ | Monoatomic | Iodide | $I^-$ | Monoatomic |
| Magnesium | $Mg^{2+}$ | Monoatomic | Nitrate | $NO_3^-$ | Tetra-atomic |
| Nickel | $Ni^{2+}$ | Monoatomic | Nitride | $N^{3-}$ | Monoatomic |
| Potassium | $K^+$ | Monoatomic | Nitrite | NO−2NO2- | Tetra-atomic |
| Silver | $Ag^+$ | Monoatomic | Oxide | $O^{2-}$ | Monoatomic |
| Sodium | $Na^+$ | Monoatomic | Phosphate | $PO_4^{3-}$ | Polyatomic |
| Zinc | $Zn^{2+}$ | Monoatomic | Sulphate | $SO_4^{2-}$ | Polyatomic |

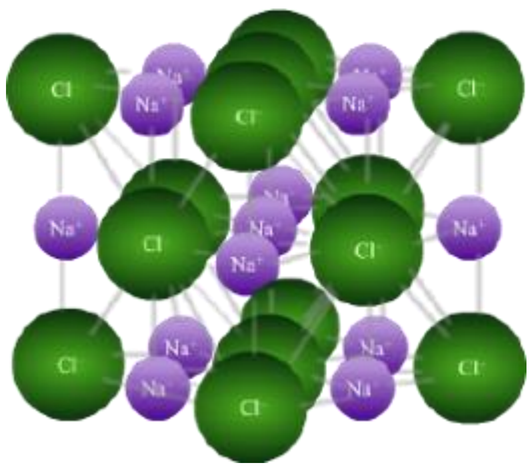
| | | | | | |
|---|---|---|---|---|---|
| Hydrogen carbonate | HCO−3HCO3- | Polyatomic | Sulphite | $SO_3^{2-}$ | Tetra-atomic |

**Ionic Compounds**

The compounds which are formed by the combination of cations and anions are known as **ionic compounds**.

For example:

- **Zinc oxide (ZnO)**: It is formed when a zinc ion ($Zn^{2+}$) combines with an oxide ion ($O^{2-}$).
- **Magnesium chloride ($MgCl_2$)**: It is formed when a magnesium ion ($Mg^{2+}$) combines with two chloride ions ($Cl^-$).
- **Potassium bromide (KBr)**: It is formed when a potassium ion ($K^+$) combines with a bromide ion ($Br^-$).
- **Sodium chloride (NaCl)**: It is formed when a sodium ion ($Na^+$) combines with a chloride ion ($Cl^-$). The structure of NaCl crystals is shown in the given figure. You can see that there is a group of $Na^+$ and $Cl^-$ ions combined with each other.



**Solved Examples**

**Easy**

**Example 1**:

**Find the atomicity of each of the following ions.**

i) $S^{2-}$

ii) $SO_4^{2-}$

iii) $NH_4^+$

iv) **$OH^-$**

**Solution**:

| Ions | Atomicity |
|---|---|
| $S^{2-}$ | 1 |
| $SO_4^{2-}$ | 5 |
| $NH_4^+$ | 5 |
| $OH^-$ | 2 |

**Medium**

**Example 2**:

**Identify the anions and cations present in the following compounds.**

| Compounds | Anions | Cations |
|---|---|---|
| NaCl | | |
| $KMnO_4$ | | |
| NaOH | | |
| KBr | | |
| $NH_4OH$ | | |

**Solution**:

| Compounds | Anions | Cations |
|---|---|---|
| NaCl | $Cl^-$ | $Na^+$ |
| KMnO4 | $MnO_4^-$ | $K^+$ |
| NaOH | $OH^-$ | $Na^+$ |
| KBr | $Br^-$ | $K^+$ |

| | | |
|---|---|---|
| NH4OH | OH− | $NH_4^+$ |

**Hard**

**Example 3**:

**Give the symbols and valence numbers for the following ions.**

| Ions | Symbols | Valence numbers |
|---|---|---|
| Ammonium | | |
| Carbonate | | |
| Sulphate | | |
| Chloride | | |
| Phosphate | | |

**Solution**:

| Ions | Symbols | Valence numbers |
|---|---|---|
| Ammonium | $NH_4^+$ | +1 |
| Carbonate | $CO_3^{2-}$ | −2 |
| Sulphate | $SO_4^{2-}$ | −2 |

| | | |
|---|---|---|
| Chloride | $Cl^-$ | −1 |
| Phosphate | $PO_4^{3-}$ | −3 |

**Writing Chemical Formulae of Compounds**

**Molecular Formula: A Brief Overview**

Just like each atom has a unique symbol, each compound has a unique molecular formula.

The molecular formula of a compound provides information about the names and numbers of atoms of the different elements present in a molecule of that compound.

**Molecular formula** is a **chemical formula** that indicates the kinds of atoms and the numbers of each kind of atom in a molecule of a compound.

**Examples**

- The molecular formula of glucose is $C_6H_{12}O_6$. One molecule of glucose contains 6 atoms of carbon, 12 atoms of hydrogen and 6 atoms of oxygen.
- The molecular formula of water is $H_2O$. One molecule of water contains 2 atoms of hydrogen and 1 atom of oxygen.

**Salient features of chemical formula:**

- Compounds are formed when two or more elements combine chemically. Hence, compounds can also be represented using symbols.
- The notation used for representing any compound is called chemical formula of that compound.
- Each compound has a unique chemical formula.
- The chemical formula of any compound tells us about : The different elements which combine to form the compound and the number of atoms of each element present in a molecule of the compound
- For example, $H_2O$ is the chemical formula of water. This denotes that there are two atoms of hydrogen and one atom of oxygen present in one molecule of water.

**Chemical Formulae**

Let us understand the information derived from chemical formulae by taking the example of carbon dioxide. The chemical formula of carbon dioxide is $CO_2$. Using this formula, we can derive the following information about carbon dioxide.

- Two elements are present in carbon dioxide: carbon(C) and oxygen (O).
- $CO_2$ represents one molecule of carbon dioxide.

- Since one atom of carbon combines with two atoms of oxygen, the **valency** of carbon is twice that of oxygen.
- $CO_2$ is a neutral molecule. It has no charge.
- The relative atomic masses of carbon and oxygen are 12 u and 16 u respectively. So, the ratio by mass between carbon and oxygen is 12 : 32, i.e., 3 : 8.

**Writing Chemical Formulae**

To write the chemical formula of a compound, one should have prior knowledge of two things.

- **The symbols of the constituent elements.**
- **The combining capacity of the atom of each element constituting the compound.**

The number of atoms of other elements with which one atom of an element combines is decided by the valency of that element.

For example, both hydrogen (H) and chlorine (Cl) have a valency of 1. Therefore, one atom of hydrogen reacts with one atom of chlorine to form one molecule of hydrogen chloride (HCl).

The valency of an ion is equal to the charge on it.

**Chemical Formulae**

The valencies of some common ions are given in the following table.

| Names of ions | Symbols | Valencies | Names of ions | Symbols | Valencies |
|---|---|---|---|---|---|
| Aluminium | $Al^{3+}$ | 3 | Sulphite | $SO_3^{2-}$ | 2 |
| Ammonium | $NH_4^+$ | 1 | Bromide | $Br^-$ | 1 |
| Calcium | $Ca^{2+}$ | 2 | Carbonate | $CO_3^{2-}$ | 2 |
| Copper(II) | $Cu^{2+}$ | 2 | Chloride | $Cl^-$ | 1 |
| Hydrogen | $H^+$ | 1 | Hydride | $H^-$ | 1 |

**Chemical Formulae**

The valencies of some common ions are given in the following table.

| Names of ions | Symbols | Valencies | Names of ions | Symbols | Valencies |
|---|---|---|---|---|---|

| | | | | | |
|---|---|---|---|---|---|
| Iron(II) | Fe2+ | 2 | Hydrogen carbonate | $HCO_3^-$ | 1 |
| Iron(III) | Fe3+ | 3 | Hydroxide | OH− | 1 |
| Magnesium | Mg2+ | 2 | Nitrate | $NO_3^-$ | 1 |
| Nickel | Ni2+ | 2 | Nitrite | $NO_2^-$ | 1 |
| Potassium | K+ | 1 | Oxide | O2− | 2 |
| Silver | Ag+ | 1 | Phosphate | $PO_4^{3-}$ | 3 |
| Sodium | Na+ | 1 | Sulphate | $SO_4^{2-}$ | 2 |
| Zinc | Zn2+ | 2 | Sulphide | S2− | 2 |

**Chemical Formulae**

The following rules need to be kept in mind while writing the chemical formulae of compounds.

•The valencies or charges on the ions must be balanced. The charge on a cation must be equal in magnitude to the charge on an anion so that the opposite charges cancel each other out and the net charge of the molecule becomes zero.

**Examples**

- In case of CaO, the valency of Ca is +2 and that of O is −2. These are then crossed over and the compound formed is CaO.

**Formula of calcium oxide**

**Symbols Ca O**

**Charges 2+ 2−**

- The charge on $Mg^{2+}$ is +2 and that on $Cl^-$ is −1. Thus, one $Mg^{2+}$ ion combines with two $Cl^-$ ions to form a molecule with the formula $MgCl_2$.

- In case of a compound consisting of a metal and a non-metal, the symbol of the metal is written first.

**Chemical Formulae**

**Example**

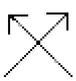
- In calcium chloride ($CaCl_2$) and zinc sulphide ($ZnS$), calcium and zinc are metals, so they are written first; chlorine and sulphur are non-metals, so they are written after the metals.
- In case of compounds consisting of polyatomic ions, the polyatomic ions are enclosed in brackets before writing the number to indicate the ratio.

**Example**

- In case of aluminium sulphate, to balance the charges, two $SO_4^{2-}$ ions combine with one $Al^{3+}$ ion. Thus the formula for aluminium sulphate is $Al_2(SO_4)_3$. Here, the brackets with the subscript 3 indicate that three sulphate ions are joined to two aluminium ions.

**Formula of aluminium sulphate**

**Symbols Al $SO_4$**



**Charges 3+ 2−**

## Chemical Formulae

### Naming Certain Compounds

| Compound | Rule | Example |
|---|---|---|
| A metal and a non-metal | Metal is written first<br>Non-metal is written last with suffix −ide | Calcium nitride (Ca3N2) |
| Two non-metals | Less electronegative non-metal is written first<br>In case, more than one atom of a non-metal is present then prefix like −di, −tri, −tetra etc. is added | Phosphorous pentachloride (PCl5) |

| | | |
|---|---|---|
| Two elements and oxygen | Oxygen is placed at end of the formula<br>Following prefixes or suffixes are used depending on the number of oxygen atoms present:<br>Less than two oxygen atom: hypo (prefix)<br>Two oxygen atoms: −ite (suffix)<br>Three oxygen atoms: −ate (suffix)<br>More than three oxygen atoms:<br>−per (prefix) | Sodium hypochlorite ($NaClO$)<br><br>Sodium chlorite ($NaClO2$)<br><br>Sodium chlorate ($NaClO3$)<br><br>Sodium perchlorate ($NaClO4$) |
| Acids | Binary acids<br>Prefix: hydro<br>Suffix: −ic with the name of second element<br>Polyatomic radicals<br>Suffix: −ic on the basis of second element<br>Prefix not used | Hydrochloric acid ($HCl$)<br><br>Sulphuric acid ($H2SO4$) |
| Trivial names | Used for specific compounds<br>No systemic rule followed | Ammonia ($NH3$)<br><br>Water ($H2O$) |

**Solved Examples**

**Easy**

**Example 1**:

**Give two examples each of molecules having one atom, two atoms and three atoms.**

**Solution**:

**Molecules having one atom (/monatomic molecules)**: Argon (Ar) and Neon (Ne)

**Molecules having two atoms (/diatomic molecules)**: Nitrogen ($N_2$) and Oxygen ($O_2$)

**Molecules having three atoms (/triatomic molecules)**: Nitrogen dioxide ($NO_2$) and carbon dioxide ($CO_2$)

**Medium**

**Example 2**:

**The valencies of a few ions are provided below.**

**$H^+ = 1$, $SO_4^{2-} = 2$, $Br^- = 1$, $Mg^{2+} = 2$ and $K^+ = 1$**

**Write the formulae for magnesium bromide, magnesium sulphate, hydrogen bromide and potassium sulphate.**

**Solution**:

•Magnesium bromide: $MgBr_2$

•Magnesium sulphate: $MgSO_4$

•Hydrogen bromide: HBr

•Potassium sulphate: $K_2SO_4$

**Hard**

**Example 3**:

**Write the names of the following compounds.**

**i)$H_2CO_3$**

**ii)$KNO_3$**

**iii)$(NH_4)_3PO_4$**

**iv)$Na_2CO_3$**

**v)$Al(NO_3)_3$**

**vi)$NaHCO_3$**

**Solution**:

i)$H_2CO_3$: Hydrogen carbonate

ii)$KNO_3$: Potassium nitrate

iii)$(NH_4)_3PO_4$: Ammonium phosphate

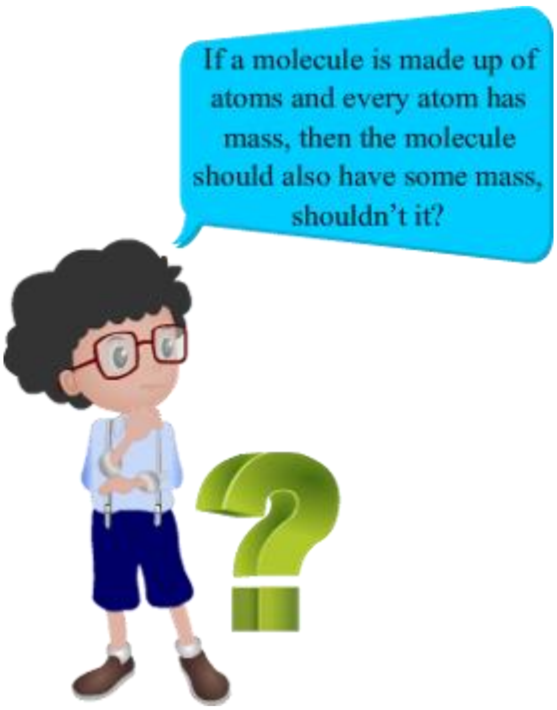
iv)$Na_2CO_3$: Sodium carbonate

v)$Al(NO_3)_3$: Aluminium nitrate

vi)$NaHCO_3$: Sodium hydrogen carbonate

**Molecular Mass and Formula Unit Mass**

**Mass of a Molecule-An Overview**




The **molecular mass** of a substance is the sum of the **atomic masses** of all the atoms present in a molecule of that substance. Molecular mass is expressed in atomic mass unit (u).

To calculate the molecular mass of a substance, the masses of all its constituent atoms are added. For example:

Molecular mass of glucose, $C_6H_{12}O_6$

= 6 (Atomic mass of C) + 12 (Atomic mass of H)
+ 6 (Atomic mass of O)

= 180 u

**Molecular Mass and Its Calculation**

In the early twentieth century, scientists performed a number of experiments to conclude that an atom can be further divided into three types of subatomic particles— **electrons** Mass of one molecule of any compound , **protons** and **neutrons** Mass of one molecule of any compound **. The neutrons and protons of an atom provide it its mass. This is called atomic mass.**

**As mentioned before, the atoms constituting a molecule provide it its characteristic mass. The molecular mass of a substance is the sum of the atomic masses of all the atoms present in a molecule of that substance.** This is also called 'relative molecular mass' and its unit is atomic mass unit (u).

**So, to calculate the mass of a molecule, we need to know the masses of the individual atoms present in that molecule.**

Let us consider a molecule $A_aB_bC_cD$

Here,

- $A$, $B$, $C$ and $D$ are the atoms of four arbitrary elements.
- $a$, $b$ and $c$ are the respective numbers of atoms of elements $A$, $B$ and $C$ present in the molecule.
- There is only one atom of element $D$ in the molecule.

The molecular mass of $A_aB_bC_cD$ is calculated as follows:

($a$ × Atomic mass of $A$) + ($b$ × Atomic mass of $B$) + ($c$ × Atomic mass of $C$)
+ (1 × Atomic mass of $D$)

**Solved Examples**

**Easy**

**Example 1**:

**Calculate the formula unit mass of sodium hydroxide. The atomic masses of sodium, hydrogen and oxygen are 23 u, 1 u and 16 u respectively.**

**Solution**:

The chemical formula of sodium hydroxide is NaOH.

It is given that:

Atomic mass of sodium (Na) = 23 u

Atomic mass of oxygen (O) = 16 u

Atomic mass of hydrogen (H) = 1 u

A formula unit of sodium hydroxide contains one atom each of sodium, oxygen and hydrogen.

So, formula unit mass of NaOH = (23 + 16 + 1) u = 40 u

**Medium**

**Example 2**:

**Calculate the formula unit mass of potassium sulphate. The atomic masses of potassium, oxygen and sulphur are 39 u, 16 u and 32 u respectively.**

**Solution**:

The chemical formula of potassium sulphate is $K_2SO_4$.

It is given that:

Atomic mass of potassium (K) = 39 u

Atomic mass of sulphur (S) = 32 u

Atomic mass of oxygen (O) = 16 u

A formula unit of potassium sulphate contains two atoms of potassium, one atom of sulphur and four atoms of oxygen.

So, molecular mass of $K_2SO_4$ = [(2 × 39) + (1 × 32) + (4 × 16)] u

= (78 + 32 + 64) u

= 174 u

**Hard**

**Example 3**:

**Calculate the formula unit mass of $[Co(NH_3)_5Br]SO_4$.**

**Solution**:

Atomic mass of cobalt (Co) = 59 u

Atomic mass of nitrogen (N) = 14 u

Atomic mass of hydrogen (H) = 1 u

Atomic mass of bromine (Br) = 80 u

Atomic mass of sulphur (S) = 32 u

Atomic mass of oxygen (O) =16 u

The formula unit mass of $[Co(NH_3)_5Br]SO_4$ is calculated as follows:

(Atomic mass of Co) + {5 × [(Atomic mass of N) + (3 × Atomic mass of H)]}

+ (Atomic mass of Br) + (Atomic mass of S) + (4 × Atomic mass of O)

= 59 u + {5 × [14 u + (3 × 1) u]} + 80 u + [32 u + (4 × 16) u]

= (59 + 85 + 80 + 96) u

= 320 u

**Formula Unit and Formula Unit Mass**

- Molecular mass and formula unit mass are basically the same. Molecular mass is the mass of a molecule, while formula unit mass is the mass of an ionic compound.
- The term **formula unit** is used for substances whose constituent particles are ions. For example, the formula unit of calcium oxide is CaO.
- The **formula unit mass** of a substance is the sum of the atomic masses of all the atoms in the formula unit of a compound.

Let us try to calculate the formula unit mass of nitric acid.

The chemical formula of nitric acid is $HNO_3$.

Atomic mass of hydrogen (H) = 1 u

Atomic mass of nitrogen (N) = 14 u

Atomic mass of oxygen (O) = 16 u

In a formula unit of $HNO_3$, there are one hydrogen (H) atom, one nitrogen (N) atom and three oxygen (O) atoms.

Thus, formula unit mass of $HNO_3$ = (1 × 1) u + (1 × 14) u + (3 × 16) u

= (1 + 14 + 48) u

= 63 u

**The Mole Concept**

**The Mole Concept: A Brief Overview**

Mole defines the quantity of a substance.

One mole of any substance will always contain **$6.022 \times 10^{23}$** particles, no matter what that substance is.

Therefore, we can say:

- 1 mole of sodium atoms (Na) contains $6.022 \times 10^{23}$ sodium atoms.
- 1 mole of sodium ions ($Na^+$) contains $6.022 \times 10^{23}$ sodium ions.
- 1 mole of hydrogen atoms (H) contains $6.022 \times 10^{23}$ hydrogen atoms.
- 1 mole of hydrogen molecules ($H_2$) contains $6.022 \times 10^{23}$ hydrogen molecules.

The word 'mole' is derived from the Latin word 'moles' which means 'heap' or 'pile'. It was first used by the German chemist Wilhelm Ostwald in 1896. It was accepted universally much later, in 1967, as a way of indicating the number of atoms or molecules in a sample.

Thus, mole can be defined as **a unit of measurement used for determining the number of atoms or molecules or ions in a given sample. It is also used to express the number of reactants and products in a chemical reaction**.

**The Mole Concept: A Brief Overview**

Consider the formation of water by the combination of hydrogen and oxygen.

$2H_2 + O_2 \rightarrow 2H_2O$

This reaction implies that 2 moles of hydrogen molecules combine with 1 mole oxygen molecules to form 2 moles of water molecules.

When carbon (C) reacts with oxygen (O), carbon dioxide is produced. **Can you write the chemical equation for the same?**

The chemical equation for the reaction is:

$$\underset{\text{Carbon}}{C} + \underset{\text{Oxygen}}{O_2} \rightarrow \underset{\text{Carbon dioxide}}{CO_2}$$

In this reaction, one atom of carbon combines with one molecule (or two atoms) of oxygen to form one molecule of carbon dioxide. We can also say that in this chemical reaction, 12 u of carbon combines with 32 u of oxygen to give 44 u of carbon dioxide. Clearly, we can represent the quantities of substances in terms of their masses. However, a chemical equation only indicates the numbers of atoms or molecules taking part in the chemical reaction. Therefore, it is easier to represent the quantities of substances involved in a chemical reaction by the numbers of their atoms or molecules rather than their masses. In order to do the same, the concept of mole is used.

**Mole Concept**

In 1909, the French physicist Jean Perrin found that one gram atom of any element contains the same number of atoms and one gram molecule of any substance contains the same number of molecules, which is equal to $\mathbf{6.022 \times 10^{23}}$.

He proposed naming this number in honour of the Italian physicist Amedeo Avogadro. Hence, $\mathbf{6.022 \times 10^{23}}$ is known as **Avogadro's number** (or Avogadro's constant) and the amount of a substance containing $\mathbf{6.022 \times 10^{23}}$ atoms/molecules/ions is called a **mole**.

Mole is a counting unit in chemistry as it is used to express large numbers of atoms or molecules.One mole of any substance can be defined as the amount of a substance that contains as many particles (atoms, molecules or ions) as there are atoms in 12 g of carbon-12 isotope. So,

1 mole of oxygen atoms (O) = $6.022 \times 10^{23}$ oxygen atoms

1 mole of oxygen molecules ($O_2$) = $6.022 \times 10^{23}$ oxygen molecules

**Know Your Scientist**

Jean Perrin (1870-1942) was a French physicist. He was awarded the Nobel Prize in Physics in 1962, for his contribution to the establishment of the atomic nature of matter, while conducting research on Brownian motion. In 1895, he showed that cathode rays are made up of negatively charged particles. He is also known for explaining the origin of solar energy through thermonuclear reaction of hydrogen (nuclear fusion) in the sun. In 1908, he studied Brownian motion using an ultramicroscope and gave experimental confirmation to the hypothesis that the random motion of suspended particles is due to the particulate nature of matter and the inter-particle interactions. He is also credited with estimating the size of a water molecule and the number of molecules of water present in a given amount of water.

Amedeo Avogadro (1776-1856) was an Italian lawyer; however, his interest in the natural sciences led him to study physics and mathematics privately. In 1809, while teaching the natural sciences in Vercelli, he hypothesized that under the same conditions of temperature and pressure, equal volumes of gases contain the same number of particles. This hypothesis later came to be known as Avogadro's law.

**Mole Concept**

The molar mass of a substance can be defined as the mass of one mole of a substance in grams. It is numerically equal to atomic/molecular/formula unit mass in u.

$$\text{Molar mass} = \frac{\text{Mass of substance}}{\text{Number of moles}}$$

The mass of one atom is called **atomic mass** and its unit is unified mass (u), while the mass of one mole of atoms is called **molar mass of atoms** and its unit is gram (g). Molar mass of atoms is also called **gram atomic mass**.

For example, the atomic mass of nitrogen (N) is 14 u, while its gram atomic mass is 14 g. So, while 14 u of nitrogen contains only 1 atom of nitrogen, 14 g of nitrogen contains 1 mole of nitrogen atoms, i.e., $6.022 \times 10^{23}$ nitrogen atoms.

The mass of one molecule is called **molecular mass** and its unit is unified mass (u), while the mass of one mole of molecules is called **molecular mass** and its unit is gram (g). When molecular mass is expressed in grams, it is called **gram molecular mass** or **gram molecule**.

For example, the molecular mass of oxygen ($O_2$) is 32 u, while its gram molecular mass is 32 g. So, while 32 u of oxygen contains only 1 molecule of oxygen, 32 g of oxygen contains 1 mole of oxygen molecules, i.e., $6.022 \times 10^{23}$ oxygen molecules.

The volume of one mole of any substance is called its **molar volume**.

The molar volume of a gas at **STP** is numerically equal to 22.4 L.

**Solved Examples**

**Easy**

**Example 1**:

**Calculate the mass of 3.3 moles of ammonia molecule.**

**Solution**:

Molar mass of ammonia molecule ($NH_3$) = 17 g

Number of moles of ammonia molecule = 3.3

We know that:

$$\text{Molar mass} = \frac{\text{Mass of substance}}{\text{Number of moles}}$$

So,

$$17\text{ g} = \frac{\text{Mass of ammonia molecule}}{3.3}$$

$$\Rightarrow \text{Mass of ammonia molecule} = (17 \times 3.3)\text{ g} = 56.1\text{ g}$$

**Medium**

**Example 2:**

**Calculate the volume of 14 g of nitrogen gas at STP.**

**Solution**:

Mass of nitrogen gas ($N_2$) = 14 g

Molar mass of nitrogen gas = 28 g

We know that:

$$\text{Molar mass} = \frac{\text{Mass of substance}}{\text{Number of moles}}$$

So,

$$28\,\text{g} = \frac{14\ \text{g}}{\text{Number of moles}}$$

$$\Rightarrow \text{Number of moles} = \frac{14\ \text{g}}{28\ \text{g}} = 0.5$$

The volume of 1 mole of a gas at STP is 22.4 L.

Therefore,

Volume of 0.5 mole of nitrogen gas at STP = $\frac{22.4\,\text{L}}{1} \times 0.5 = 11.2\,\text{L}$

**Avogadro's Law**

In 1811, Avogadro hypothesized that under the same conditions of temperature and pressure, equal volumes of all gases contain an equal number of moles. For example, at the same temperature and pressure, the two gases, oxygen and nitrogen possessing the same volume contain the same number of molecules. This hypothesis is named Avogadro's law. The mole concept provides the following information.

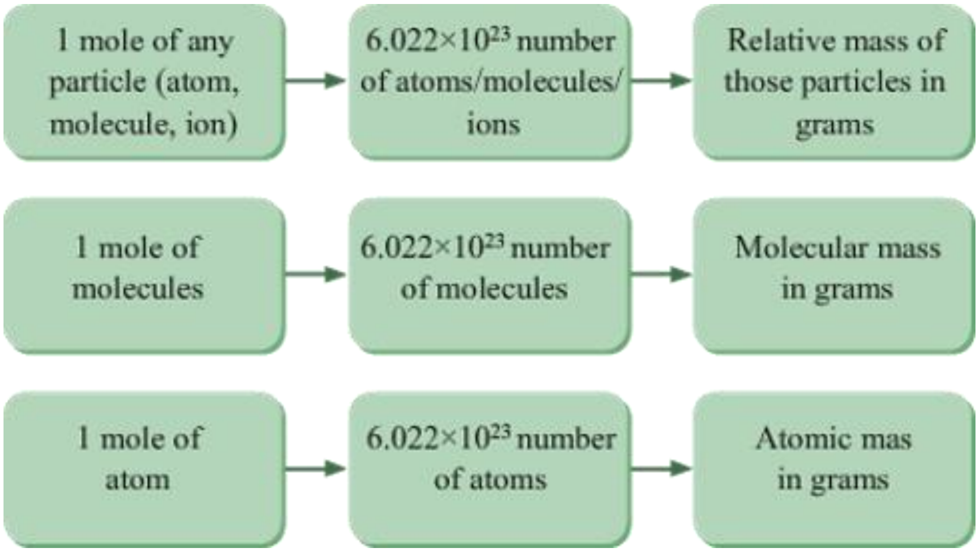
- If one mole of a substance (atoms, molecules or ions) is present, then the number of elementary particles present in that substance is equal to $6.022 \times 10^{23}$.
- The mass of one mole of a substance (atoms, molecules or ions) is equal to its molar mass.
- While carrying out reactions, scientists require the number of atoms and molecules. This requirement is fulfilled by the use of the mole concept as follows:

1 mole = $6.022 \times 10^{23}$ = Relative mass in grams.

**Avogadro's Law**

**Relationship between Mole, Avogadro's Number and Mass**

The relationship between mole, Avogadro's number and mass is summarized in the given figure.




**Applications of Avogadro's Law**

- It provides an explanation for Gay-Lussac's law.

  The volumes of different combining gases bear a simple ratio to one another because according to Avogadro's law at constant temperature and pressure equal volumes of gases contain the same number of molecules.

- It helps in determination of atomicity of gases.

  Consider the formation of hydrogen chloride gas by the direct combination of hydrogen and chlorine gases:
  $H_2 + Cl_2 \rightarrow 2\ HCl$
  According to Avogadro's law:
  1 molecule of $H_2$ + 1 molecule of $Cl_2 \rightarrow$ 2 molecules of HCl
  Or, 1/2 molecule of $H_2$ + 1/2 molecule of $Cl_2 \rightarrow$ 1 molecule of HCl
  As atoms are indivisible, therefore, half a molecule of $H_2$ and $Cl_2$ indicate they both contain two atoms per molecule.

- It helps in determination of molecular formula of a gas.

  Consider the formation of hydrogen chloride gas by the direct combination of hydrogen and chlorine gases:
  $H_2 + Cl_2 \rightarrow 2\ HCl$

According to Avogadro's law:
1 molecule of $H_2$ + 1 molecule of $Cl_2$ → 2 molecules of HCl
Or, 1/2 molecule of $H_2$ + 1/2 molecule of $Cl_2$ → 1 molecule of HCl
Or, 1 atom of H + 1 atom of Cl → 1 molecule of HCl
Hydrogen chloride gas has one atom of hydrogen and one atom of chlorine. Therefore, its molecular formula is HCl.

- It helps in the establishment of the relationship between molecular mass and vapour density (VD).

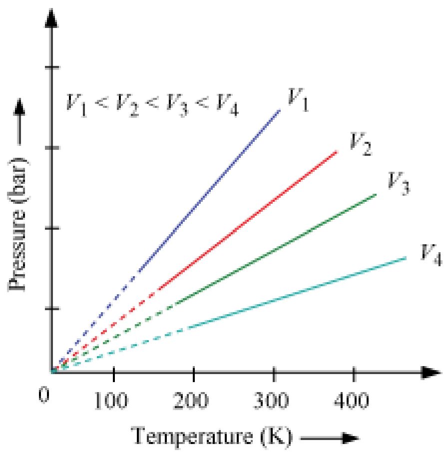
$$2 \times \text{VD} = \frac{\text{Mass of 1 molecule of a gas}}{\text{Mass of 1 atom of hydrogen}}$$

2 × VD = Mass of 1 molecule of a gasMass of 1 atom of hydrogen

- It provides the relation between gram molecular mass and gram molecular volume.

Molar volume of a gas = 22.4 L
Gram molecular mass of a gas occupies 22.4 L and contains $6.02 \times 10^{23}$ molecules/atoms of the gas.

**Gay-Lussac's Law**




This law gives the relationship between pressure and temperature. According to this law, *at constant volume, the pressure of a fixed amount of a gas is directly proportional to the temperature*. It can be represented mathematically as,

Mathematically,

$p \propto T$

$\Rightarrow \frac{p}{T}$ = constant = $k_3$

If at constant volume,

$p_1$ = Pressure of a gas at $T_1$

$p_2$ = Pressure of the same gas at $T_2$

Then,

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