# THE MOLE

## Atomic Weight and Molecular Weight

In the early stages of scientific inquiry, the determination of atomic and molecular weights involved disparate measurement scales. However, the current universally accepted standard was established by the International Union of Pure and Applied Chemistry (IUPAC) in 1961. This standardized system relies on the most stable isotope of carbon, specifically the (C-12) isotope, as the designated standard for comparing the atomic masses of different elements.

To elaborate on this standard, an element's atomic weight is defined as "the number of times an atom of that element is heavier than $\frac{1}{12}$th of the mass of a single atom of $C^{12}$. In simpler terms, the atomic weight of an element represents the relative mass of its atoms compared to the standardized mass of a $C^{12}$ atom. This standardized approach, set by IUPAC, establishes a consistent and universally accepted framework for measuring atomic and molecular weights across various chemical elements.

$$\therefore \quad \text{Atomic weight} = \frac{\text{Mass of one atom}}{\frac{1}{12}\text{th of mass of one atom of } C^{12}}$$

$$\text{Molecular weight} = \frac{\text{Mass of one molecule}}{\frac{1}{12}\text{th of mass of one atom of } C^{12}}$$

Since both of these values lack units, they are commonly referred to as relative atomic mass and relative molecular mass. These quantities, devoid of specific units, serve as measures that describe the relative mass of atoms or molecules without being tied to a particular unit of measurement. In scientific contexts, the terms "relative atomic mass" and "relative molecular mass" are employed to underscore this characteristic, emphasizing the comparative nature of these mass values across different elements or compounds.

## Atomic And Molecular Mass

One of the key principles derived from Dalton's atomic theory is the concept of atomic mass, stating that each element possesses a distinct atomic mass. Due to the minuscule size of atoms, measuring their absolute masses proves challenging. Nonetheless, it becomes feasible to establish the relative masses of various atoms by adopting a small unit of mass as a standard.

In this context, the mass of one hydrogen atom was chosen as the unit and accepted as the standard. The atomic mass of an element can be defined as the numerical value indicating how many times the mass of one atom of that element surpasses the mass of one hydrogen atom.

$$A = \text{Atomic mass of an element}$$
$$= \frac{\text{Mass of an element}}{\text{Mass of one atom of hydrogen}}$$

In 1858, the oxygen atom was established as the standard for several reasons:

(i) Obtaining compounds of elements with oxygen is more convenient than with hydrogen due to the higher reactivity of oxygen compared to hydrogen.

(ii) Using oxygen as the standard results in the atomic masses of most elements approximating whole numbers. In contrast, if hydrogen were the standard, the atomic masses of many elements would be fractional. The mass of one atom of natural oxygen was set at 16.0, thereby defining the atomic mass of an element

$$= \frac{\text{Mass of one atom of the element}}{\frac{1}{16}\text{ th part of the mass of one atom of oxygen}}$$

$$= \frac{\text{Mas of one atom of the element}}{\text{mass of one atom of oxygen}}$$

By adopting oxygen as the standard, the atomic masses are determined as follows: hydrogen has an atomic mass of 1.008, sodium 22.991, and sulfur 32.066.
In 1961, the International Union of Chemists introduced a new unit for expressing atomic masses, choosing the stable isotope of carbon ($^{12}C$) with a mass number of 12 as the standard. The atomic mass of an element is defined as the numerical value indicating how many times the mass of one atom of that element is greater compared to $1/12^{th}$ of the mass of one atom of carbon-12 ($^{12}C$).

$$A = \text{Atomic mass of an element}$$

$$= \frac{\text{Mass of one atom of the element}}{\frac{1}{12}\text{th part of the mass of one atom of carbon}-12}$$

$$= \frac{\text{Mass of one atom of the element}}{\text{Mass of one atom of carbon}-12} \times 12$$

[The term 'A' was previously referred to as atomic weight. However, this terminology is now obsolete due to the ambiguity associated with the word 'weight,' which commonly denotes gravitational force.]

## Atomic Mass:

The quantity denoted as 'A' represents the mass of an atom of carbon-12 ($^{12}$ C) and is termed the atomic mass unit, abbreviated as amu. The precise mass of one atom of carbon-12 is measured at $1.9924 \times 10^{-23}$g or $1.99.24 \times 10^{-26}$ kg.
Thus,

$$1\text{ amu} = \frac{1.9924 \times 10^{-23}}{12} = 1.66 \times 10^{-24}\text{ g or } 1.66 \times 10^{-2}\text{ kg}$$

$$A = \text{atomic mass of an element}$$

$$= \frac{\text{Mass of one atom of the element}}{1\text{ amu}}$$

The atomic masses of some elements on the basis ofcarbon-l2 are given below:

| | | | |
|---|---|---|---|
| Hydrogen | 1.008 amu | Iron | 55.847 amu |
| Oxygen | 16.00 amu | Sodium | 22.989 amu |
| Chlorine | 35.453 amu | Zinc | 65.38 amu |
| Magnesium | 24.305 amu | Silver | 107.868 amu |
| Copper | 63.546 amu | | |

The true mass of an atom of an element, measured in grams, is obtained by multiplying the atomic mass of the element in atomic mass units (amu) by $1.66 \times 10^{-24}$ g. Consequently, the actual mass of a hydrogen atom is calculated as $1.008 \times 1.66 \times 10^{-24} = 1.6736 \times 10^{-24}$ g. Similarly, the actual mass of an oxygen atom is determined by the formula $16 \times 1.66 \times 10^{-24} = 2.656 \times 10^{-23}$ g.
The atomic masses listed above reveal that the atomic masses of several elements are not whole numbers. These values represent average relative masses, as most elements exist in nature as a combination of isotopes (atoms of the same element with different atomic masses). Although elements typically exhibit constant mixtures of isotopes, there are exceptions.
For instance, chlorine occurs naturally as a mixture with two isotopes: Cl-35 (34.969 amu) and Cl-37 (36.966 amu), present in the ratio of 75.53% (Cl-35) to 24.47% (Cl-37).
Therefore, the average relative mass of chlorine is calculated as:

(34.969 x 0.7553) + (3'6.966 x 0.2447) = 35.46 amu

Derived from the average mass, the atomic mass of chlorine is often expressed as 35.46 or 35.5 amu. However, it's crucial to note that it is impossible for an atom to possess a precise relative mass of 35.5 amu. Instead, it may have a relative mass of approximately 35.0 or 37.0 amu, depending on the specific isotope.

Consequently, the average relative mass of any naturally occurring sample of chlorine is denoted as 35.46 or 35.5 amu, given that it comprises a mixture of two isotopes in defined proportions. This reasoning is applicable to all other elements as well. The average atomic masses of various elements are determined by multiplying the atomic mass of each isotope by its fractional abundance and subsequently summing the obtained values.
The fractional abundance is determined by· dividing percentage abundance by hundred.

**Example.** Boron has two isotopes boron-1O and boron-II whose percentage abundances are 19.6% and 80.4% respectively. What is the average atomic mass of boron?

**Solution:** Contribution of boron-1O 10.0 $X$ 0.196 = 1.96amu

Contribution of boron-II =ll. Ox 0.804 = 8.844amu

Adding both =1.96+8.844=10.804amu

Hence, the average atomic mass of boron is 10.804 amu. When the numerical value of an element's atomic mass is expressed in grams, it is referred to as gram-atomic mass or gram atom. For instance, the atomic mass of oxygen is 16, whereas the gram-atomic mass or gram atom of oxygen is 16 g. Correspondingly, the gram-atomic masses of hydrogen, chlorine, and nitrogen are 1.008 g, 35.5 g, and 14.0 g, respectively. It is noteworthy that the gram atomic mass or gram atom of every element comprises the same number of atoms, a quantity designated as Avogadro's number. The value of Avogadro's number is 6.02 x 1023. Absolute mass of one oxygen atom

= 16amu =16x 1.66 x 10-24 g

Therefore, the mass of 6.02 x 1023 atoms of oxygen will be

= *16x* 1.66 x 10-24 x 6.02 X 1023

16g (gram-: atomic mass)

Thus, gram-atomic mass can be defined as the absolute mass in grams of 6.02 x 1023 atoms of any element. Number of gram atoms of any element can be calculated with the help of the following formula:

$$\text{No. of gram atoms} = \frac{\text{Mass of the element in grams}}{\text{Atomiv mass of the element in grams}}$$

## Molecular Mass

Similar to an atom, a molecule of a substance is also a minute particle with a mass typically ranging from $10^{-24}$ to $10^{-22}$ g. Much like atomic mass, molecular mass is expressed as a relative mass concerning the mass of a standard substance, which can be an atom of hydrogen, oxygen, or carbon-12. The molecular mass of a substance is defined as the mass of a molecule relative to the mass of an atom of hydrogen (taken as 1.008), oxygen (taken as 16.00), or one atom of carbon (taken as 12). The molecular mass is a numerical value indicating how many times one molecule of a substance is heavier compared to $\frac{1}{16}$ th of the mass of an oxygen atom or $\frac{1}{12}$ th of the mass of one atom of carbon-12.

$$\text{M} = \text{Molecular mass}$$
$$= \frac{\text{Mass of one molecule of the substance}}{\frac{1}{12}\text{ th mass of one atom of carbon}-12}$$

The mass of a molecule is the aggregate of the masses of the atoms within it.
For example, a water molecule comprises two hydrogen atoms and one oxygen atom, resulting in a molecular mass of water calculated as (2 × 1.008) + 16.06, which equals 18.016 amu. Similarly, a molecule of sulfuric acid (H2SO4) is composed of two hydrogen atoms, one sulfur atom, and four oxygen atoms.
Thus, the molecular mass of sulphuric acid is

= (2 x 1.008) + 32.00 + (4 x 16.00)

= 98.0160r 98.016amu

Gram-molecular Mass or Gram Molecule is a term used to denote a quantity of a substance whose mass in grams corresponds numerically to its molecular mass. To put it simply, when the molecular mass of a substance is expressed in grams, it is referred to as gram-molecular mass or gram molecule.
For instance, considering chlorine with a molecular mass of 71, its gram-molecular mass or gram molecule would be 71 g.
Similarly, molecular mass of oxygen (02) is 32, i. e.,

$$2x\ 16 = 32amu.$$

Gram-molecular mass of oxygen 32g Molecular mass of nitric acid (HN03) is 63, i. *e.,*

$$= 1 + 14 + 3\ x\ 16 = 63\ amu$$

Gram-molecular mass of nitric acid = 63 g
Gram-molecular mass should not be confused with the mass of one molecule of the substance in grams. The mass of one molecule of a substance is known as its actual mass.
For example, the actual mass of one molecule of oxygen is equal to.

$$32x\ 1.66\ x\ 10\text{-}24\ g,\ i.e., 5.32\ x\ 10\text{-}23\ g.$$

The number of gram molecules of a substance present in a given mass of a substance can be determined by the application of following formula:

$$\text{No. of gram molecules} = \frac{\text{Mass of a Substance in grams}}{\text{Molecular mass of the substance in grams}}$$

$$\text{Mass of single molecule} = \frac{\text{Molar mass in grams}}{6.023 \times 10^{23}}$$

$$= \text{Molar mass in amu} \times 1.66 \times 10^{-2}\ \text{ grams}$$

Calculate the mass of 2.5-gram atoms of oxygen.

**Solution:** We know that,

$$\text{No. of gram atoms} = \frac{\text{Mass of the element in grams}}{\text{Atomic mass of the element in grams}}$$

So, Mass of oxygen = 2.5 x 32 80.0g

Calculate the mass of 1.5-gram molecule of sulphuric acid.

**Solution:** Molecular mass of H

$$2S0_4 = 2x\ 1 + 32 + 4\ x\ 16 = 98.0amu$$

Gram-molecular mass of $H_2S0_4$ = 98.0g
Mass of 1.5-gram molecule of $H_2S0_4$ = 98.0 x 1.5 = 147.0g

## Mole Concept:

The counting of items often involves the use of units like dozen or gross, regardless of the nature of the items. For instance, one dozen pencils equal 12 pencils, one dozen apples equal 12 apples, and one gross of books equals 144 books, or one gross of oranges equals 144 oranges. Similarly, in the realm of chemistry, when counting atoms, molecules, ions, etc., chemists employ the unit mole. Coined by Ostwald in 1896, the term "mole" originates from the Latin word 'moles,' signifying a heap or pile. A mole (mol) is defined as the number of atoms in 12.00 g of carbon-12. Experimentally determined to be $6.02 \times 10^{23}$, this number is known as Avogadro's number, named in honor of Amedeo Avogadro (1776 - 1856).
Thus, one mole encompasses $6.02 \times 10^{23}$ units, which can be atoms, molecules, ions, electrons, or other entities.
For example, 1 mole of hydrogen atoms corresponds to $6.02 \times 10^{23}$ hydrogen atoms, 1 mole of hydrogen molecules equates to $6.02 \times 10^{23}$ hydrogen molecules, and so on for different entities.

It's crucial to specify the type of entity when using the mole designation. Therefore, a mole of oxygen atoms comprises $6.02 \times 10^{23}$ oxygen atoms, while a mole of oxygen molecules contains the same number of molecules. Consequently, a mole of oxygen molecules is equivalent to two moles of oxygen atoms, i.e., $2 \times 6.02 \times 10^{23}$ oxygen atoms.

The weight of one mole depends on the nature of the particles or units. The mass of one mole of atoms of any element is precisely equal to the atomic mass in grams (gram-atomic mass or gram atom) of that element.

For instance, the atomic mass of aluminum is 27 amu (atomic mass units). Since one amu is equal to $1.66 \times 10^{-24}$ g, one mole of aluminum contains $6.02 \times 10^{23}$ aluminum atoms.

Mass of one atom aluminum = $27 \text{ x } 1.66 \text{ x } 10^{-24}$ g

Mass of one mole aluminum =$27 \text{ x } 1.66 \text{ x } 10^{-24} \text{ X } 6.02\text{x}10^{23}$

=27g

This corresponds to the atomic mass of aluminum expressed in grams, often referred to as one-gram atomic mass or one gram atom of aluminum. Similarly, the mass of $6.02 \times 10^{23}$ molecules (1 mole) of a substance equals its molecular mass in grams, also known as gram-molecular mass or gram molecule. For instance, the molecular mass of water is 18 amu. Therefore, the mass of one mole of water is calculated as $18 \times 1.66 \times 10^{-24} \times 6.02 \times 10^{23}$, resulting in 18 grams.

This represents the molecular mass of water in grams, denoted as one gram-molecular mass or one gram molecule. The concept of moles is applicable to ionic compounds that lack molecules. In such instances, the formula of an ionic compound signifies the ratio between constituent ions. The mass of $6.02 \times 10^{23}$ formula units represents one mole of an ionic compound.

One mole of BaC12

= $6.02 \text{ x } 10^{23}$BaCl2 units

= 208.2g BaCl2

Molecular mass (formula mass) of BaCl2

= $6.02 \text{ x } 10^{23}$ $Ba^{2}$ + ions + $2 \text{ x } 6.02 \text{ x } 10^{23}$CI- ions

= 137.2 + 71.0 = 208.2g

One mole of a substance will have mass equal to formula mass of that substance expressed in grams.

## Mole:

A mole is defined as the quantity of a substance containing the same number of entities (atoms, molecules, or other particles) as there are atoms precisely in 0.012 kg (or 12 g) of the carbon-12 isotope.

Through mass spectrometry, it has been determined that $6.023 \times 10^{23}$ atoms are present in 12 g of the C-12 isotope. This specific number of entities in one mole is so significant that it is assigned a distinct name and symbol, known as Avogadro's constant, denoted by NA. In essence, we can express 1 mole as the assembly of $6.02 \times 10^{23}$ entities. These entities may encompass atoms, ions, molecules, or even everyday objects like pens, chairs, or paper.

1 mole of atom is also termed as 1 g – atom

1 mole of ions is also termed as 1 g – ion

1 mole of molecule is also termed as 1 g – molecule

Methods of Calculations of Mole

(a) If no. of some species is given,

$$\text{then no. of moles} = \frac{\text{Given no.}}{N_A}$$

(b) If weight of a given species is given,

then no. of moles $= \frac{\text{Given wt.}}{\text{Atomic wt.}}$ (for atoms),

or $\frac{\text{Given wt.}}{\text{Molecular wt.}}$ = (for molecules)

(c) If volume of a gas is given along with its temperature (T) and pressure (P).

use n $\frac{PV}{RT}$ =

where R = 0.0821 lit-atm/mol-K

(When P is in atmosphere and V is in liters)

1 mole of any gas at STP occupies 22.4 litter.

**Ex.** Chlorophyll the green colouring material of plants contains 3.68 % of magnesium by mass. Calculate the number of magnesium atom in 5.00 g of the complex.

**Sol.** Mass of magnesium in 5.0 g of complex $= \frac{3.68}{100} \times 5.00 = 0.184$ g

Atomic mass of magnesium = 24

24 g of magnesium contain $= 6.023 \times 10^{23}$ atoms

$\therefore$ 0.184 g of magnesium would contain $= \frac{6.023 \times 10^{23}}{24} \times 0.184 = 4.617 \times 10^{21}$ atom Therefore, 5.00 g of the given complex would contain $4.617 \times 10^{21}$ atoms of magnesium.

## Gravimetric Analysis

Once we get a balanced chemical equation then we can interpret a chemical equation by following ways.

1. Mass - mass analysis
2. Mass - volume analysis
3. Mole - mole analysis
4. Vol - Vol analysis (separately discussed as eudiometry or gas analysis)

Now you can understand the above analysis by following example

### 1. Mass - Mass Analysis

Consider the reaction $2\,KClO_3 \longrightarrow 2KCl + 3O_2$ According to stoichiometry of the reaction

mass-mass ratio: $2 \times 122.5$: $2 \times 74.5$: $3 \times 32$

or $\frac{\text{Mass of } KClO_3}{\text{Mass of KCl}} = \frac{2\times122.5}{2\times74.5}$ $\frac{\text{Mass of } KClO_3}{\text{Mass of } O_2} = \frac{2\times122.5}{3\times32}$

**Ex.** Calculate the weight of iron which will be converted into its oxide by the action of 36 g of steam.

(Given: $3Fe + 4H_2O \longrightarrow Fe_3O_4 + H_2$)

**Sol.** Mole ratio of reaction suggests,

$$\frac{\text{Mole of Fe}}{\text{Mole of } H_2O} = \frac{3}{4}$$

$\therefore$ Mole of Fe $= \frac{3}{4} \times$ mol of $H_2O$

$$= \frac{3}{4} \times \frac{36}{18} = \frac{3}{2}$$

wt. of Fe $= \frac{3}{2} \times 56 = 84$ g

**Ex.** In a gravimetric determination of P of an aqueous solution of dihydrogen phosphate in $H_2PO^-_4$ is treated with a mixture of ammonium and magnesium ions to precipitate magnesium ammonium phosphate, Mg $(NH_4)$ $PO_4.6H_2O$. This is heated and decomposed to magnesium pyrophosphate, $Mg_2P_2O_7$. A solution of $H_2PO^-_4$ yielded 2.054 g of $(Mg_2P_2O_7)$ which is weighed. What weight of $NaH_2PO_4$ was present originally?

**Sol.** $NaH_2PO_4 + Mg^{2+} + NH_4^+ \longrightarrow Mg(NH_4)PO_4.6H_2O \xrightarrow{\Delta} Mg_2P_2O_7$

As P atoms are conserved, applying POAC for P atoms, moles of P in $NaH_2PO_4$ = Moles of P in $Mg_2P_2O_7$

$\Rightarrow$ $1 \times \text{Moles of } NaH_2PO_4 = 2 \times \text{Moles of } Mg_2P_2O_7$

$\therefore$ $$\frac{W_{SH_2PO_4}}{M_{NH_2PO_4}} = 2 \times \frac{W_{Mg_2P_2O_7}}{M_{Mg_2P_2O_7}}$$

$\Longrightarrow$ $$\frac{W_{NaH_2PO_4}}{120} = 2 \times \frac{2.054}{222}$$

$\therefore$ $$W_{NaH_2PO_4} = 2.22 \text{ g}$$

## 2. Mass - Volume Analysis

Now again consider decomposition of $KClO_3$

$$2\ KClO_3 \longrightarrow 2KCl + 3\ O_2$$

mass volume ratio: $2 \times 122.5$ g: $2 \times 74.5$ g: $3 \times 22.4$ L at STP

we can use two relations for volume of oxygen

$$\frac{\text{Mass of } KClO_3}{\text{volume of } O_2 \text{ at STP}} = \frac{2\times122.5\ \text{g}}{3\times22.4\ \text{L}} \quad \text{..........(i)}$$

And $$\frac{\text{Mass of KCl}}{\text{volume of } O_2 \text{ at STP}} = \frac{2\times74.5\ \text{g}}{3\times22.4\ \text{L}} \quad \text{.........(ii)}$$

**Ex.** How much marble of 90.5 % purity would be required to prepare 10 liters of $CO_2$ at STP when the marble is acted upon by dilute HCl?

**Sol.** $$CaCO_3 + 2HCl \longrightarrow CaCl_2 + H_2O + CO_2$$

100 g              22.4litre

22.4 L of $CO_2$ at STP will be obtained from 100 g of $CaCO_3$

$\therefore$ 10 L of $CO_2$ at STP will be obtained from pure $CaCO_3 = \frac{100}{22.4} \times 10 = 44.64$ g

$\therefore$ Impure marble required $= \frac{100}{90.5} \times 44.64 = 49.326$ g

## 3. Mole - Mole Analysis

This analysis is very much important for quantitative analysis point of view.

Now consider again the decomposition of $KClO_3$.

$$2\ KClO_3 \longrightarrow 2\ KCl + 3\ O_2$$

In very first step of mole-mole analysis you should read the balanced chemical equation like 2 moles $KClO_3$ on decomposition gives you 2 moles KCl and 3 moles $O_2$ and from the stoichiometry of reaction we can write

$$\frac{\text{Moles of KClO}_3}{2} = \frac{\text{Moles of KCl}}{2} = \frac{\text{Moles of O}_2}{3}$$

Now for any general balance chemical equation like

$$aA + bB \longrightarrow cC + dD$$

you can write.

$$\frac{\text{Moles of A reacted}}{a} = \frac{\text{Moles of B reacted}}{b} = \frac{\text{Moles of C reacted}}{c} = \frac{\text{Moles of D reacted}}{d}$$