

Atomic Mass and Molecular Mass

You fairly have a rough idea of what atoms and molecules are. But, did you ever wonder that atomic mass and molecular mass are important concepts as well? What do we do by knowing the mass of such tiny atoms and molecules? Well, the significance of these atomic and molecular masses is huge! In this chapter, we will cover the concepts and also see how important they are.

Atomic Mass

An atom is such small a particle that it cannot be seen or isolated. Therefore, it is impossible to determine the actual mass of a single atom by weighing it. However, Avogadro's hypothesis finally solved the problem. He took equal volumes of two different gases under similar conditions of temperature and pressure. He then weighed them.

Surprisingly, the ratio of their masses was equal to the ratio of their single molecules. Thus, though the actual masses of the atoms could not be determined, their relative masses could be determined. If the atomic mass of hydrogen is taken as 1, the relative atomic mass of oxygen is 16.

Initially, scientists obtained the atomic masses of all the elements by comparing with the mass of hydrogen taken as 1. However, the problem was that the atomic masses of most of the elements came out to be fractional in this method. Therefore, carbon is taken as the reference for the determination of atomic masses.

Browse more Topics under Some Basic Concepts Of Chemistry

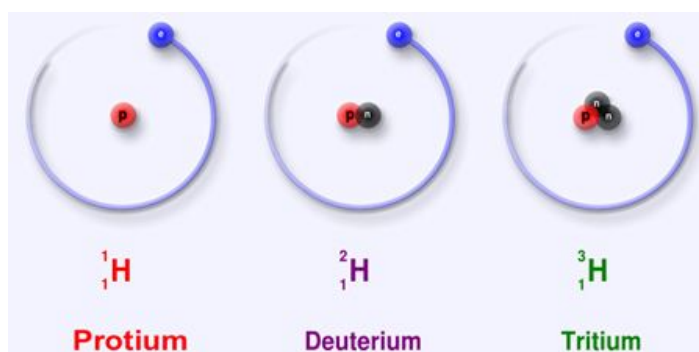
- [Concentrations](#)
- [Dalton's Atomic Theory](#)
- [Importance and Scope of Chemistry](#)
- [Laws of Chemical Combination](#)
- [Mole and Equivalent Weight](#)
- [Nature of Matter](#)
- [Percentage Composition](#)
- [Properties of Matter and Their Measurement](#)
- [Stoichiometry and Stoichiometric Calculations](#)
- [Uncertainty in Measurement](#)

Learn about [Stoichiometry and Stoichiometric Calculations](#) here.

What is Atomic Mass?

Atomic mass of an element is the number of times an atom of that element is heavier than an atom of carbon taken as 12. One atomic mass unit is equal to one-twelfth of the mass of an atom of carbon 12 isotope. The atomic mass of an element is the average relative mass of its atoms as compared to an atom of carbon 12 taken as 12.

Fractional abundance of an isotope is the fraction of the total number of atoms that is comprised of that particular isotope. Atomic mass of an element = (Fractional abundance of isotope 1 \times mass of isotope 1) + (Fractional abundance of isotope 2 \times mass of isotope 2)



Gram Atomic Masses

The atomic masses of elements expressed in grams is their gram atomic masses. For eg: the atomic mass of oxygen is 16 amu.

\therefore Gram atomic mass of oxygen = 16 g.

Molecular Mass

The molecular mass of a substance is the number of times the molecule of the substance is heavier than one-twelfth the mass of an atom of carbon -12. Or, the molecular mass is equal to the sum of its atomic masses of all the atoms present in one molecule of a substance.

For eg: Water (H_2O)

Atomic mass of H = 1 unit

Atomic mass of O = 16 units

The molecular mass of water = $2 \times \text{atomic mass of H} + 1 \times \text{atomic mass of O}$
 $= 2 \times 1 + 16 \times 1$
 $= 18 \text{ units}$

Gram Molecular Mass

The molecular mass of a substance expressed in grams is the gram molecular mass. For eg: Molecular mass of oxygen = 32u

\therefore Gram molecular mass of oxygen = 32 g

Learn how to calculate [the number of mole and equivalent weight](#).

Solved Example For You

Q: The specific heat of a metal of atomic mass 32 is likely to be:

- a. 0.25
- b. 0.24
- c. 0.20
- d. 0.15

Solution: Specific heat = $6.4/\text{atomic mass}$

Therefore, Specific Heat = $6.4/32 = 0.2$

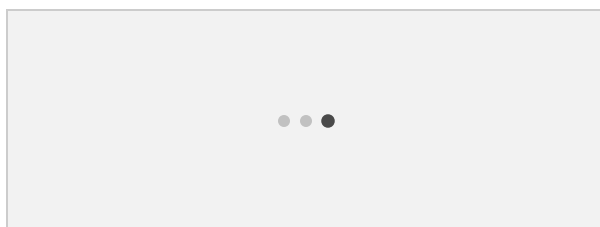
Hence, the answer is (c).

Concentrations

Do you like your coffee thick or more ‘watery’? What do we mean by this? Maybe the ratio of **water** to milk? Or, maybe the concentration of milk in water! We always discuss a solution being diluted or concentrated. However, this is a qualitative way of expressing the concentration of the **solution**.

Dilute and Concentrated Solutions

A dilute solution means the quantity of solute is relatively very small. On the other hand, a concentrated solution implies that the solution has a large amount of solute. But, unfortunately, these are relative terms. They do not give us an idea of the quantitative concentration of the solution.



So, to quantitatively describe the concentrations of various solutions around us, we commonly express levels of concentration in different ways. Let us look at a few of them below.

Quantitative Measurement of Concentration

1) Mass Percentage (w/w)

When we express the concentration of a solution as the percent of one component in the solution, we call it the mass percentage (w/w).

Suppose, we have a solution containing component A as the solute and B as the solvent, then its mass **percentage** is expressed as:

Mass % of A = $\frac{\text{Mass of component A in the solution}}{\text{Total mass of the solution}} \times 100$

10% solution of sugar by mass means that 10 grams of sugar is present in 100 grams of the solution, i.e., 10 grams of sugar has been dissolved in 90 grams of water.

2) Volume percentage (V/V)

Sometimes, we express the concentration as a percent of one component in the solution by volume. In such cases, it is the volume percentage. It is given as:

Volume % of A = $\frac{\text{Volume of component A in the solution}}{\text{Total volume of the solution}} \times 100$

For example, if a solution of NaCl in water is said to be 10 % by volume that means a 100 ml solution will contain 10 ml NaCl.

3) Mass by Volume Percentage (w/V)

This unit is majorly used in pharmaceutical industry. It is the mass of a solute dissolved per 100 mL of the solution.

4) Molarity (M)

One of the most commonly used methods for expressing the concentrations is molarity. It is the number of moles of solute dissolved in one litre of a solution. Thus, if one gram molecule of a solute is present in 1 litre of the solution, the concentration of the solution is said to be one molar. Unit of molarity: mol L^{-1}

Suppose a solution of ethanol is labelled as 0.25 M. By this, we mean that in one litre of the given solution 0.25 moles of ethanol is dissolved.

5) Molality (m)

Molality represents the concentration regarding moles of solute and the mass of solvent. It is given by moles of solute dissolved per kg of the solvent. The molality formula is given as:

$$\text{Molality}(m) = \text{Moles of solute} / \text{Mass of solvent in kg}$$

Relationship Between Molality and Molarity
--

$$\frac{\text{Molarity}}{\text{Molality}} = \frac{\text{No. of moles of solute}}{\text{Volume of solution in L}} \times \frac{\text{Mass of solvent in Kg}}{\text{Number of moles of solute}}$$
$$= \frac{\text{Mass of solvent in Kg}}{\text{Volume of solution in L}} = \frac{W}{V}$$

Let the density of the solution be 'd'. Unit= g mL⁻¹

$$\text{Mass of solution} = V \times d$$

The mass of solute = number of moles \times molecular mass of solute = $n \times m_A$

$$\begin{aligned}\text{Mass of solvent, } W &= \text{mass of solution} - \text{mass of solute} \\ &= V \times d - n \times m_A\end{aligned}$$

Thus,

$$\frac{\text{Molality}}{\text{Molarity}}$$

$$\frac{M}{m} = V \times d - n \times \frac{m_A}{V}$$

Where m_A is the molecular mass of solvent.

6) Mole Fraction

If the solution has a solvent and the a solute, mole fraction gives a concentration as the ratio of moles of one component to the total moles present in the solution. It is denoted by x .

The above-mentioned methods are commonly used ways of expressing the concentration of solutions. All the methods describe the same thing that is, the concentration of a solution, each of them has their own advantages and disadvantages. Molarity depends on temperature while mole fraction and molality are independent of

temperature. All these methods are used on the basis of requirement of expressing the concentrations.

Solved Example for You

Q: What do you mean by the formality of a solution?

Answer: It is the number of formula mass in grams present per litre of solution. In case, formula mass is equal to molecular mass, formality is equal to molarity. Like molarity and normality, the formality also depends on temperature.

Dalton's Atomic Theory

It's amazing to think about how life evolved on this planet. Ever since our existence on this planet, the [human brain](#) has been curious to understand nature. This curiosity has been the mother of all inventions. There have been a number of theories to predict the same. One such theory is the Dalton's Atomic Theory. We will learn about this theory and its postulates in this chapter. We will also look at its limitations. Let's begin.

Dalton's Atomic Theory

Matter has been one of the most important subjects of research for the science enthusiasts. Scientists and philosophers have always tried to simplify things. They wanted to know about the fundamental particles that make matter, their properties, structure etc. This led to the formulation of a number of atomic theories.

Democritus was the first man who proposed that matter is made up of particles. He named these particles, ‘atomos’ meaning indivisible. This was the Democritus’ Atomic Theory. Due to the lack of technological setup back then, scientists had very limited information on this theory.

Almost after two thousand years, the works on the simplifying matter was materialized by scientist, John Dalton. In 1808, John Dalton postulated the famous Dalton’s Atomic Theory. He published this theory in a paper titled “A New Chemical Philosophy”; indeed the philosophy was new for that era. Let us now look at the postulates of this theory.

Browse more Topics under Some Basic Concepts Of Chemistry

- [Atomic Mass and Molecular Mass](#)
- [Concentrations](#)

- [Importance and Scope of Chemistry](#)
- [Laws of Chemical Combination](#)
- [Mole and Equivalent Weight](#)
- [Nature of Matter](#)
- [Percentage Composition](#)
- [Properties of Matter and Their Measurement](#)
- [Stoichiometry and Stoichiometric Calculations](#)
- [Uncertainty in Measurement](#)

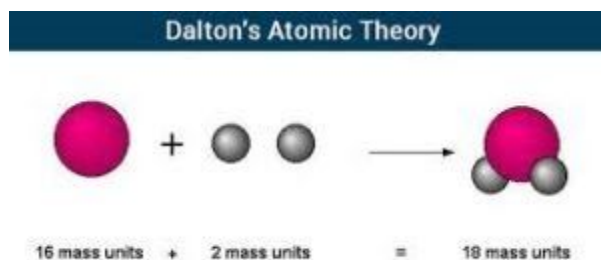
Learn Dalton's Atomic Theory from this Video Lecture

Postulates of Dalton's Atomic Theory

- The matter is made up of indivisible particles known as atoms.
- The properties of all the atoms of a given element are the same including mass. This can also be stated as all the atoms of an element have identical mass while the atoms of different elements have different masses.
- Atoms of different elements combine in fixed ratios to form compounds.
- Atoms are neither created nor destroyed. This implies that during chemical reactions, no atoms are created nor destroyed.

- The formation of new products (compounds) results from the rearrangement of existing atoms (reactants).
- Atoms of an element are identical in mass, size and many other chemical or physical properties, but atoms of two-different elements differ in mass, size, and many other chemical or physical properties.
- Learn about [Atomic Mass and Molecular Mass](#) here.

However, this theory is not entirely free of limitations. Let us now look at the drawbacks of this theory.



(Source: Sutori)

How do Scientists calculate the exact number of Atoms and Molecules? Find by learning [Mole and Equivalent Weight](#) here.

Drawbacks of Dalton's Atomic Theory of Matter

- It was proved that an atom is not indivisible. As an atom can be subdivided into electrons, protons and neutrons. But remember that atom is the tiniest particle that takes part in a chemical reaction.
- According to Dalton Atomic Theory, atoms of an element are identical in mass, size and many other chemical or physical properties. But, practically we observe that atoms of several elements differ in their densities and masses. These atoms with the different masses are known as isotopes. For example, Chlorine (Cl) has 2 isotopes with the mass numbers of 35 and 37.
- Also, according to Dalton Atomic Theory, atoms of two-different elements differ in mass, size and many other chemical or physical properties. However, this is not correct for all situations. For example, Argon (Ar) and Calcium (Ca) atoms, each have an atomic mass of 40 amu. These atoms with similar atomic masses are isobars.
- According to Dalton Atomic Theory, when atoms of different elements (atoms of two or more elements) combine in simple whole number ratios, we get chemical compounds. But this is not true in case of complex organic compounds.

- Dalton Atomic Theory fails to explain the existence of allotropes. This implies that the Dalton atomic theory fails to explain the differences in properties of charcoal, graphite, and diamond (allotropes of carbon).
- Dalton's Atomic Theory also suggested that an atom is the smallest part of an atom that can take part in a chemical reaction. Some postulates of this theory remain valid even in today's modern chemical thoughts. The atomic structure model proposed by indeed proves to be a significant, stepping stone in chemistry. It forms the base for modern atomic theories and quantum mechanics.

Learn more [Basic Concepts of Chemistry here](#).

Solved Examples for You

Q: Give the supporting laws for Dalton's Atomic Theory.

Ans: John Dalton based his theory on two laws. They are explained below:

- Law of Conservation of Mass

According to the law of conservation of mass, the matter is neither created nor destroyed. This means, in a chemical reaction, amount of elements remains same in starting when only reactants there and at the completion of the reaction when product formed. We always use the “Law of conservation of mass” when we balance chemical equations.

- Law of Constant Composition

According to the law of constant composition, a pure compound will always have the same proportion of the same elements. For example, table salt with the molecular formula of NaCl holds the same proportions of the elements Na (sodium) and Cl (chlorine). This composition doesn't depend on where the salt came from and how much salt one should have.

Importance and Scope of Chemistry

Have you ever asked yourself this question as to why do we study chemistry? What do we get from learning all these elements and

compounds and their applications? Does it seem vague to you? Well, it is not! The importance and scope of chemistry are huge! In this chapter, we will look at some of the practical applications of chemistry. However, before we start, can you tell what exactly is chemistry?

What is Chemistry?

Chemistry is that branch of science dealing with the study of [composition](#), structure, and properties of matter. It deals with the study of the changes which different forms of matter undergo under different conditions. Chemistry also had branches that look at the laws governing these changes. Now that we know what chemistry is, let us move on and read about the importance of the same.

Importance and Scope of Chemistry

There are many instances in your day-to-day life that involves chemistry, its applications, and its rules. Let us look at them one by one.

1) Supply of Food

The study of chemistry provided the world with chemical fertilizers such as urea, calcium superphosphate, sodium nitrate, and Ammonium Sulphate. These chemicals have helped greatly in increasing the yield of fruits, vegetables, and other crops. Thus, we can cater to the ever-growing demand for food. It has helped to protect the crops from insects and harmful bacteria by the use of certain effective insecticides, fungicides, and pesticides.

Chemistry also led to the discovery of preservatives. These chemicals have greatly helped to preserve food products for a longer period. It has given methods to test the presence of adulterants. This ensures the supply of pure foodstuff.



2) Contribution to Improved Health and Sanitation Facilities

Chemistry provided mankind with a large number of life-saving drugs. We could find a cure for **dysentery** and pneumonia due to the discovery of sulphur drugs and penicillin. Besides this, life-saving drugs like cisplatin and taxol are effective for cancer therapy and AZT is used for AIDS victim. Some of the common drugs that chemistry has blessed us with include:

- Analgesics: To reduce the pain of different types.
- Antibiotics: To curb infection and cure diseases.
- Tranquillisers: To reduce tension and bring about calm and peace to patients suffering from mental diseases.
- Antiseptics: To stop infection of the wounds.
- Disinfectants: To kill the **microorganism** present in toilets, floor, and drains.
- Discovery of anaesthetics has made surgical operations more and more successful.
- The use of insecticides such as DDT and Gammexane has greatly reduced the dangers of diseases caused by the rats, mosquitoes, and flies.

3) The Scope of Chemistry in Saving the Environment

Thanks to science, now we have environment-friendly chemicals that help us conserve the nature. One such example is the replacement of CFCs in the refrigerators.

4) Increase in Comfort, Pleasure, and Luxuries

Because of the advancements in science and the discoveries of chemistry, we lead a more comfortable life today. You may ask how? Let us see below.

- Synthetic fibres: These are more attractive, comfortable and durable. They include terylene, nylon, and rayon. They are easy to wash, dry quickly and do not need ironing. Chemistry provides a large number of synthetic dyes which impart bright and fast colour to the clothes.
- Building materials: By supplying steel and cement, chemistry helps in the construction of safer homes and multi-storage buildings. It also helps in the construction of long-lasting and durable dams and bridges.
- Supply of metals: Metals like gold, silver, copper, iron, aluminium, zinc and a large number of the alloys are used for making various objects. These include ornaments, utensils, coins, and many Industrial and agricultural implements.

- Articles of domestic use: Chemistry has made our homes more comfortable by supplying a large number of articles of domestic uses. Examples include detergents, oils, and fats, sugar, paper, glass, plastic, paints, cosmetics, perfumes, cooking gas etc. We are able to beat the heat in summers by using refrigerants like ammonia, liquid sulphur dioxide, and freon.
- Entertainment: Cinema, video cameras, simple cameras make use of films which are made of Celluloid and coated with suitable chemicals. Fireworks which amuses us on festival and marriages are chemical products. Can you imagine how boring life would have been if you wouldn't have been able to take all those cute selfies?
- Transport and communication: All means of transport use either petrol or diesel or coal which are all chemical products.
- Nuclear atomic energy: Chemistry has come to the rescue by providing an alternative source of energy which is nuclear energy.

5) The Scope of Chemistry in Industry

Chemistry plays an important and useful role towards the development and growth of a number of industries. This includes industries like glass, cement, paper, textile, leather, dye etc. We also see huge applications of chemistry in industries like paints, pigments, petroleum, sugar, plastics, Pharmaceuticals.

It has also helped in the greater production of [sulphuric acid](#), [nitric acid](#), and [Ammonia](#), hydrogenated oils by providing suitable catalysts.

6) The Scope of Chemistry in War

Chemistry plays an important role in the discovery of highly explosive substances such as TNT, nitroglycerine, and dynamite. It also plays a role in finding poisonous gases like mustard gas, Phosgene etc.

Solved Example for You

Q: List down some problems caused by Chemistry.

Ans:

- Nuclear energy is useful but the disposal of nuclear waste poses a serious problem to humanity.

- Phonograph records have added to our pleasure for listening to music but they are made of polyvinyl chloride. This is produced from vinyl chloride which can cause liver cancer in industrial workers.
- Antibiotics have eliminated infectious diseases but the overuse is very harmful. Chemistry has given drugs like LSD, cocaine, brown sugar. These prove to be a curse to the society.

Laws of Chemical Combination

When you write the various [chemical equations](#), do you balance them?

What happens when you don't balance them? Your marks get cut.

That is okay. But, what exactly happens? There are various laws of Chemical Combinations that govern these facts. So, apart from getting your marks cut, these laws are important because they help to keep things in place. What do we mean by this? Let us read about it below!

Laws of Chemical Combinations

Chemistry is the study of the change of matter from one form to the other. These changes often occur as a result of the combination of two different types of matter. There are certain rules that govern the

combination of different elements to form different compounds. These rules are, what we call, the laws of chemical combinations.

Browse more Topics under Some Basic Concepts Of Chemistry

- [Atomic Mass and Molecular Mass](#)
- [Concentrations](#)
- [Dalton's Atomic Theory](#)
- [Importance and Scope of Chemistry](#)
- [Mole and Equivalent Weight](#)
- [Nature of Matter](#)
- [Percentage Composition](#)
- [Properties of Matter and Their Measurement](#)
- [Stoichiometry and Stoichiometric Calculations](#)
- [Uncertainty in Measurement](#)

Watch Videos on Laws of Chemical Combination

There are five basic laws of chemical combinations that govern the chemical combinations of elements. What are they and what do they signify? Let us read more about these laws in the section below.

1) Law of Conservation of Mass

French chemist, Antoine Lavoisier in 1789, studied this [law](#). This law states that “In all physical and chemical changes, the total mass of the reactants is equal to that of the [products](#)” or “Mass can neither be created nor destroyed.”

Suggested Videos on Law of Conservation of Mass –

We also refer to this law as the law of indestructibility of matter. The mass and energy are interconvertible but the total sum of the mass and [energy](#) during any physical or chemical change remains constant. That is why your teacher cuts your marks when you don't balance the equations! Understand now?

Learn what [quantity](#) of elements are perfect for reaction by studying [Stoichiometric Calculations](#).



2) Law of Constant Composition or Definite Proportions

French chemist, J.L. Proust in 1799, discovered this law. It states that “A chemical compound is always found to be made up of the same elements combined together in the same fixed proportion by mass”.

For example, a sample of pure water from various sources or any country is always made up of only [hydrogen](#) and [oxygen](#). These elements are always in the same fixed ratio of 1:8 by mass. We can prepare a sample of carbon dioxide in the laboratory. We can do this in various ways like:

- Heating limestone
- Burning [coal](#) in air
- The action of dilute hydrochloric acid on marble
- Heating sodium carbonate

However, we will always find that it contains the same elements, carbon, and oxygen, in the same fixed ratio of 3:8 by mass.

Suggested Videos on Law of Constant Proportion

Limitation of this Law

It is not applicable if an element exists in different [isotopes](#) which may be involved in the formation of the compound. The elements may combine in the same ratio but the compounds formed may be different.

Learn [how to determine the Percentage composition of the solution](#).

3) Law of Multiple Proportions

When two elements combine to form two or more chemical compounds, then the masses of one of the elements which combined with a fixed mass of the other, bear a simple ratio to one another. For eg, Carbon combines with oxygen to form two compounds namely carbon dioxide and carbon monoxide.

In carbon dioxide, 12 parts by mass of carbon combine with 32 parts by mass of oxygen while in carbon monoxide, 12 parts by mass of carbon combine with 16 parts by mass of oxygen. The masses of oxygen which combined with a fixed mass of carbon in carbon monoxide and carbon dioxide are 16 and 32. These masses of oxygen bear a simple ratio of 16:32 or 1:2 to each other.

Taking another example of the Compounds of Sulphur and oxygen, we see something similar. The element sulphur also forms two oxides Sulphur dioxide and sulphur trioxide. In sulphur dioxide, 32 parts by mass of Sulphur combine with 32 parts by mass of oxygen. On the other hand, in sulphur trioxide, 32 parts by mass of Sulphur combines with 48 parts by mass of oxygen.

The masses of oxygen which combined with the fixed mass of sulphur in the two oxides are 32 and 48. These bear a simple ratio of 32:48 or 2:3 to each other.

Suggested Videos on Law of Multiple Proportion

4) Law of Reciprocal Proportion

This law was put forward by Richter in 1792. It states that “The ratio of masses of 2 elements, A and B which combines separately with a fixed mass of the third element C is either the same or some multiple of the ratio of the masses in which A and B combine directly with each other”.

For example, the elements carbon and oxygen combine separately with the third element hydrogen to form methane and water. However, they combine directly with each other to form carbon dioxide. In

methane, 12 parts by mass of carbon combine with 4 parts by mass of hydrogen. In water, 2 parts by mass of hydrogen combine with 16 parts by mass of oxygen.

Watch Videos on Law of Reciprocal Proportion

The masses of carbon and oxygen which combine with a fixed mass of hydrogen are 12 and 32 ie. they are in the ratio 12:32 or 3:8. In carbon dioxide, 12 parts by mass of carbon combine directly with 32 parts by mass of oxygen ie they combined directly in the ratio of 12:32 or 3:8. This is exactly the same as the first ratio.

5) Gay Lussac's Law of Gaseous Volume

When gases react together they always do so in volumes which bear a simple ratio to one another and to the volume of the products, if these are also gases. This holds true provided all measurements of volumes are done under similar conditions of temperature and pressure.

Read the [Importance and Scope of Chemistry here](#).

Solved Example for You

Q: Give an example of Gay Lussac's Law of Gaseous Volume.

Ans: A simple example to prove the Gay Lussac's law is that of hydrogen and chlorine. 1 volume of hydrogen and 1 volume of chlorine always combine to form two volumes of hydrochloric acid gas. The ratio between the volumes of the reactants and the product in this reaction is simple, i.e., 1:1:2.

Mole and Equivalent Weight

If we ask you to count the total number of stars in the sky, can you do that? No. Similarly, scientists can not count the exact number of **atoms** and molecules as they are very tiny and present in a large number. Therefore, we have the concepts of mole and equivalent weight to make this calculation easy. In this chapter, we will study these two concepts and look at a few examples for the same.

What Do You Mean By A Mole?

As we already mentioned, atoms and molecules are very tiny. That is why they are very difficult to be counted in exact figures. Well, this problem is solved by using the **Avogadro's number**. This number is expressed as $N_A = 6.023 \times 10^{23}$. We define a mole as the number

equal to Avogadro's number. This is just like you say dozen is equivalent to 12, a score is 20 and a century is 100.

In other **words**, a mole is a unit that we use to represent 6.023×10^{23} particles of the same matter. Describing further, a mole is the total amount of substance that contains as many atoms, molecules, ions, electrons or any other elementary entities as there are carbon atoms in exactly 12 gm of it. There are various kinds of moles of a substance that we can calculate. They are as follows:

- The number of moles of molecules
- Number of moles of atoms
- The number of moles of gases (Standard molar **volume** at STP = 22.4 lit)
- Number of moles of particles e.g. atoms, molecules ions etc

We calculate the mole fraction of a substance as the **fraction** of the substance in the **mixture** expressed in terms of mol. It is represented as X.

Browse more Topics under Some Basic Concepts Of Chemistry

- [Atomic Mass and Molecular Mass](#)

- Concentrations
- Dalton's Atomic Theory
- Importance and Scope of Chemistry
- Laws of Chemical Combination
- Nature of Matter
- Percentage Composition
- Properties of Matter and Their Measurement
- Stoichiometry and Stoichiometric Calculations
- Uncertainty in Measurement

Equivalent Weight

We define the equivalent weight of a substance (element or compound) as:

*“The number of parts by weight of it, that will combine with or displace directly or indirectly 1.008 parts by weight of **hydrogen**, 8 parts by weight of oxygen, 35.5 parts by weight **chlorine** or the equivalent parts by weight of another element”.*

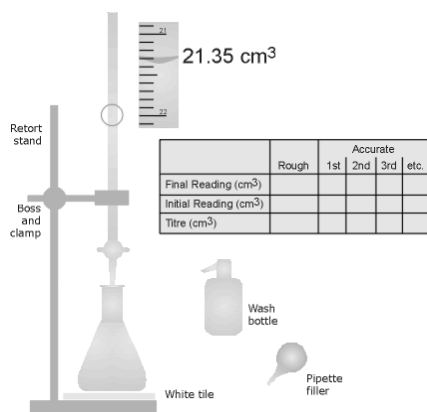
It is important to know that the equivalent weight of any substance is dependent on the reaction in which it takes part. Equivalent weight is a

relative quantity so it does not have any unit. When we express the equivalent weight of a substance in grams, we call it Gram Equivalent Weight (GWE).

How to Calculate Equivalent Weight?

$$\text{Equivalent weight} = \text{Molar mass} / \text{Valence factor}$$

(The Valence factor for a base = acidity of the base, the Valence factor for an acid = basicity of the acid and the Valence factor for an element = **valency**)



Titration

Titration is a procedure in which we can determine the concentration of an unknown solution with the help of solution, the concentration of which we know. In this procedure, we determine the concentration of solution A by adding carefully measured volumes of a solution of

known concentration B . We continue adding these solutions until the reaction of A with B is just complete.

Law of Equivalence

Titration is based on the Law of Equivalence. This law states that,

“At the endpoint of a titration, volumes of the two titrants reacted have the same number of equivalents or milliequivalents”

Acid-Base Titration

One gram equivalent of acid is **neutralized** by one gm equivalent of base. It means,

One equivalent of Acid = One Equivalent of Base

$$\text{Acid } [N_1V_1] = \text{Base } [N_2V_2]$$

$$[\text{Gram equivalent} = \text{Normality} \times \text{Volume}]$$

Let us understand this with the help of a simple example. Let us calculate the number of milliequivalents of H_2SO_4 present in 10 mL of

N/2 H_2SO_4 solution. Milliequivalents = Normality \times Volume (mL) =
 $\frac{1}{2} \times 10 = 5$ milliequivalent of H_2SO_4

Limiting Reactant

The reactant that is completely consumed in the course of the reaction is the limiting reactant. When it is fully consumed, the reaction itself stops. However, what we must note is that the concept of limiting reactant is applicable to reactions other than monomolecular reactions. These are the reactions where more than one type of reactant is involved.

To determine the limiting reagent, first, we must be aware of the amount of all reactants and mole ratio of reactants. If the ratio of moles of reactant A with respect to reactant B is greater than the ratio of the moles of A to moles of B for a balanced chemical equation then B is the limiting reactant.

We can use the concept of stoichiometry to determine the other terms like left (unused) mass of other reactants, amount of formed product once we know the limiting reactant.

Solved Example for You

Q: Calculate the number of moles in each of the following:

1. 392 grams of sulphuric acid
2. 44.8 litres of carbon dioxide at STP
3. 6.022×10^{23} molecules of oxygen

Solution:

1) 1 mole of $\text{H}_2\text{SO}_4 = 98 \text{ g}$

Thus, $98 \text{ g of H}_2\text{SO}_4 = 1 \text{ mole of H}_2\text{SO}_4$

$392 \text{ g of H}_2\text{SO}_4 = 4 \text{ moles of H}_2\text{SO}_4$

2) 1 mole of $\text{CO}_2 = 22.4 \text{ litres at STP}$

i.e. $22.4 \text{ litres of CO}_2 \text{ at STP} = 1 \text{ mole}$

$44.8 \text{ litres of CO}_2 \text{ at STP} = 2 \text{ moles CO}_2$

3) 1 mole of O_2 molecules = 6.022×10^{23} molecules.

$6.022 \times 10^{23} \text{ molecules} = 1 \text{ mole of oxygen molecules.}$

Nature of Matter

Do you matter? Okay, that was silly on our part to ask! Of course, you do. You matter and you ARE matter! Do you know what it means? What is matter? And what is the nature of matter? Well, that is what we are going to study in this chapter. We will look at the different **aspects** of matter and study about their **nature**. But, first tell us, what is matter?

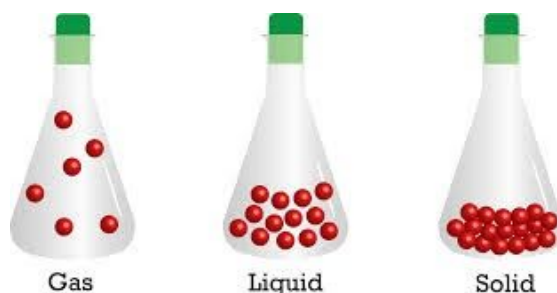
What is Matter?

We all know what matter is. Isn't it? Everything we see around is matter. But, then air doesn't classify as matter. NO! Air IS matter! So, a matter is anything that occupies space and has mass.

All the buildings, the bridges, the **atomic particles** are all matter. Even our **DNA**, the air, the molecules inside our bodies, everything is matter.

We know that matter comprises of particles. These particles are atoms and **molecules**. In this section, we will cover the nature of matter.

Based on its physical state, we can divide the nature of matter into three major categories.



Nature of Matter

- **Solids:** Solids are all those substances having their particles very close to each other. There exist strong intermolecular forces between these particles. The particles are firmly held in their positions. These particles have only vibratory [motion](#). Solids have a definite shape and definite volume. Examples include Wood, [iron](#), aluminium etc.
- **Liquids:** Liquids comprise of all those substances with weak intermolecular forces. The particles are capable of minimum movement. They have a definite volume. However, they do not have a definite shape. They usually take the shape of the container in which we place them. Examples include water, milk, etc.

- Gases: Gases are those forms of matter having very weak forces between their molecules. Hence, in gases, the molecules are free to move. The distance between molecules is large as compared to solids and liquids. Gases have neither fixed shape nor a definite volume. They tend to completely occupy the container in which they are placed. E.g. air, oxygen, hydrogen, methane, etc.

We can change the state of matter from one form to another by changing the conditions of pressure and [temperature](#). We must also note that the nature of matter depends on its composition as well. If the matter consists of more than one type of particles then it is a mixture. On the other hand, if it consists of a single type of particles then it is a pure substance.

We can further classify [mixtures](#) into homogeneous and heterogeneous mixtures. Pure substances are also categorised further as elements and compounds. The nature of matter continues to be a vast subject of research and recent advancements have revealed some other states of matter. The two other states of matter that scientists

have found recently include the Boson-Einstein condensate and plasma

Solved Example for You

Q: What is a compound?

Answer: A compound is a matter consisting of two or more different elements bonded chemically. A molecule forms when two or more atoms or elements (may be the same type or different types) join together chemically. For example, O_2 is a molecule (diatomic) as it contains 2 oxygen atoms that are chemically bonded. However, CO_2 is a compound containing two different atoms, Carbon and Oxygen that are bonded chemically.

Percentage Composition

Percentage composition of a compound is a ratio of an amount of each element to the total amount of individual elements in a compound, which is then multiplied with 100. Let us take an example of H_2O i.e. Water. Water has 2 molecules of hydrogen, and one mole of water is of 18.0152 grams. And one mole of a hydrogen atom is of 1.008 grams. So, 2 hydrogen moles weighs 2.016 grams. Hence, one mole of

water has 2.016 grams of hydrogen mole. Therefore, the percentage composition of hydrogen would be $2.016/18.0152 = 11.19\%$. Now let us study in detail about Percentage composition and its formula.

What is Percentage Composition?

The percentage composition of any given compound is nothing but the ratio of the amount of each element present in the compound to the total amount of individual elements present in the compound multiplied by 100. Here, we measure the quantity in terms of grams of the elements present in the solution.

The percent composition of any compound is an expression of its composition in terms of all the elements present. The significance of this composition calculation is found in the chemical analysis.



Percentage Composition Formula

We can express the percentage composition of a given element using the formula below:

$$\% C_E = \frac{\text{mass of element E}}{\text{total mass of compound}} \times 100$$

Here, % C_E is the percentage composition of the element E. This is the value that we are going to calculate. The numerator on the right side indicates the total amount of element E present in the compound. On the other hand, the denominator is an expression for the total amount of all the elements present in the compound.

We multiply this ratio by 100 to get the percentage form of the composition. Let us now look at the mass percentage of composition in more details. We will also look at its importance.

Learn more about [Atomic mass and molecular mass](#) here.

Browse more Topics Under Some Basic Concepts Of Chemistry

- [Atomic Mass and Molecular Mass](#)
- [Concentrations](#)
- [Dalton's Atomic Theory](#)
- [Importance and Scope of Chemistry](#)
- [Laws of Chemical Combination](#)
- [Mole and Equivalent Weight](#)
- [Nature of Matter](#)
- [Percentage Composition](#)
- [Properties of Matter and Their Measurement](#)
- [Stoichiometry and Stoichiometric Calculations](#)
- [Uncertainty in Measurement](#)

Mass Percentage Composition

We use the concept of mass percentage composition to denote the concentration of an element in a compound or a component in a mixture. We use this term to signify the total percent by mass of each element that is present in a compound.

It is important to note that we can calculate the mass percentage composition by dividing the mass of a component by the total mass of the mixture. This ratio is then to be multiplied by 100. We also call it the mass percent(w/w) %.

Importance of Mass Percentage

For many scientific and practical reasons, we ought to know what quantity of an element is present in a given compound. This is indispensable for the chemists to get to the empirical formula of many compounds.

This formula helps in showing the lowest whole number of moles and the relative number of atoms of each element in a compound. With the help of the empirical formula, chemists can also calculate the actual molecular formula. This formula shows the exact number of atoms in the compound.

Determining the Mass Percent from a Chemical Formula

We will explain this section with the help of an example. It will clear your fundamentals on how to calculate the mass percent. Let us consider glucose. The formula for glucose is $C_6H_{12}O_6$. Let us calculate the mass percent of each of the elements in glucose.

From the formula, we can see the number of moles of each of the elements. Glucose has 6 carbon atoms, 12 hydrogen atoms and also 6 oxygen atoms. Let us now multiply each of the atoms by its molar mass. Thus, we arrive at the mass of each of the element in glucose.

We now divide each mass by mass of 1 mole of glucose to arrive at the mass fraction of each element. Multiplying this value by 100 gives the mass percentage of the elements.

The Example in Detail

Each mole of carbon has a mass of 12.01g/mol of carbon. This, we know, from the periodic table. So, 6 moles of carbon will have $12.01\text{g/mol} \times 6 = 72.06\text{g}$ of Carbon. Similarly, 1 mole of Hydrogen

has a mass of 1.008g/mol of Hydrogen. Therefore, 12 moles of Hydrogen will have the mass of $12 \times 1.008 = 12.096\text{g}$ of Hydrogen.

Going by the same logic for Oxygen, 1 mole of oxygen has a mass of 16.00g/mol. Therefore, 6 moles of oxygen will have $16.00 \times 6 = 96\text{ g}$ of Oxygen. Thus, 1 mole of Glucose ($\text{C}_6\text{H}_{12}\text{O}_6$) has a total mass of $72.06 + 12.096 + 96 = 180.16\text{ g/mol}$

Calculating the Mass Percentage

To find out the mass % of the three elements of glucose, let us first calculate the mass fraction of each element in 1 mole of glucose. So, what do we mean by the mass fraction? It is the mass that each element contributes to the total mass of glucose. By multiplying this by 100, we get the mass percentage of each element.

$$\text{Mass fraction of Carbon} = 72.06\text{g}/180.16\text{g} = 0.4000$$

$$\text{Therefore, mass \% of Carbon} = 0.4000 \times 100 = 40.00\%$$

$$\text{Mass fraction of Hydrogen} = 12.096/180.16 = 0.06714$$

$$\text{Therefore, mass \% of Hydrogen} = 0.06714 \times 100 = 6.714\%$$

Mass fraction of Oxygen = $96/180.16 = 0.53286$

Therefore, mass% of Oxygen = $0.53286 \times 100 = 53.286\%$

It is interesting to note that even if carbon and oxygen have an equal number of moles of the compound, their mass percentages are different. Oxygen has a higher value in this as its molar mass is higher than that of carbon.

Learn how to calculate the [number of Moles and Equivalent weight](#) here.

Solved Example For You

Q: Find the percent composition of each element in water.

Solution: We know that the chemical formula for water is H_2O . Let us now calculate the molar mass of water. The molar mass of Oxygen = $16.00 \times 1 = 16 \text{ g/mole}$ and of Hydrogen = $1.01 \times 2 = 2.02 \text{ g/mole}$

Now, using the molar mass of each of the given elements, we find out the percentage composition of each element in H_2O . It is given as the ratio of the grams of the element to the grams of the total element in

the compound, multiplied by 100. Calculating the percentage composition of Hydrogen,

$$\% \text{ H} = 2.02/18.02 \times 100$$

Therefore, % H= 11.21 %

Calculating the percentage composition of Oxygen,

$$\text{Therefore, } \% \text{ O} = 16/18.02 \times 100 = 88.79 \%$$

Properties of Matter and Their Measurement

Can you see air? No! Can you eat it? No! Then why do we call it as matter? Well, that is interesting. To understand that, you have to first know what matter is. You would also have to know the various properties of matter. Only then you can truly know why we classify air as matter. Are you ready? Let's begin the journey through the concept of the matter!

What is Matter?

Matter is anything that has mass and occupies some space. So, even if you can't see air, you can feel it. You know it occupies space. That is why a balloon bursts when you pump extra air into it! Also, you are aware of the fact that air exerts pressure. If it weren't matter, how would it exert pressure?

Browse more Topics under Some Basic Concepts Of Chemistry

- [Atomic Mass and Molecular Mass](#)
- [Concentrations](#)
- [Dalton's Atomic Theory](#)
- [Importance and Scope of Chemistry](#)
- [Laws of Chemical Combination](#)
- [Mole and Equivalent Weight](#)
- [Nature of Matter](#)
- [Percentage Composition](#)
- [Stoichiometry and Stoichiometric Calculations](#)
- [Uncertainty in Measurement](#)

Classification of Matter

We can classify matter at both micro and macroscopic levels. At the microscopic level, we can classify it into solid, liquid and gas. These

are the three physical states of matter. We cannot convert these three states within themselves by applying temperature or pressure.

At the macroscopic level, we can classify matter into pure substances and mixtures. This classification depends on the chemical composition of various substances. Pure substances are further of two types: elements and compounds.



Elements and Compounds

Element refers to the primary stuff present in all the substance. The smaller unit of an element is the atom. Examples of elements include Nitrogen, Hydrogen, Oxygen, Carbon etc. A compound is a pure substance having more than one atom of elements linked by chemical bonds. Chemical reactions result in these bonds. Examples include water.

Mixture

Mixtures refer to the aggregate of more than one type of pure substance. The chemical identity of a mixture remains maintained. Their constituent ratio could vary unlike in compounds. For example, sugar syrup is a mixture of water and sugar.

Physical and Chemical Properties of Matter

Let us look at the physical and chemical properties of matter in a nutshell.

- **Physical Properties of matter:** These are the properties that we can measure without changing the chemical composition of the substance. The physical properties include mass, volume, density, refractive index etc.
- **Chemical Properties of matter:** These are the properties that we can evaluate at the cost of matter itself. As an example, we can test the sweet taste of sugar by eating it only.
- **Physical Quantities:** Many physical properties of matter are quantitative in nature. Such properties of matter are physical properties. The measurement of any physical quantity consists of two parts: the number and the unit. We express the accuracy of the number by using the concepts of significant figures. We

express the units of measurement in S.I. units. Usually, we use the concept of dimensional analysis to derive the units.

Units for Measurement

We need to measure all physical quantities. We can express the value of a physical quantity as the product of the numerical value and the unit in which it is expressed.

Fundamental Units

Fundamental units are those units which can neither be derived from one another nor they can be further resolved into any other units.

Below, we have listed the seven fundamental units of measurement in S.I. system.

Quantity	Name of unit	Abbreviation
Mass	Kilogram	Kg
Length	Meter	m

Temperature	Kelvin	K
Amount of substance	Mole	Mol
Time	Second	S
Electric current	Ampere	A
Luminous intensity	Candela	Cd

Derived units

Some quantities are expressed as a function of more than one fundamental units known as derived units. For example velocity, acceleration, work, energy etc.

Quantity with Symbol	Unit (S.I.)	Symbol
Velocity (v)	Metre per sec	ms^{-1}
Area (A)	Square metre	m^2

Volume (V)	Cubic metre	m^3
Density (ρ)	Kilogram m^{-3}	Kg m^{-3}
Energy (E)	Joule (J)	$\text{Kg m}^2\text{s}^{-2}$
Force (F)	Newton (N)	Kg ms^{-2}
Frequency (ν)	Hertz	Cycle per sec
Pressure (P)	Pascal (Pa)	Nm^{-2}
Electrical charge	Coulomb (C)	A-s (ampere – second)

Units and Dimensional Analysis: Conversion of Units

Dimensional Analysis is one of the simplest ways to carry out the calculations involving different units. Here, we convert a quantity expressed in one unit into an equivalent quantity with a different unit.

We do this by the use of a conversion factor which expresses the relationship between units.

Original quantity \times conversion factor = equivalent quantity

(in former unit)

(in other units)

This is based on the fact that ratio of each fundamental quantity in one unit with their equivalent quantity in other unit is equal to one.

Solved Example for You

Q: What is the difference between homogeneous and heterogeneous mixtures?

Ans: A homogeneous mixture has a uniform composition throughout the entire mixture. The common examples include the mixture of sugar or salt in water, iodine in alcohol or any alloy. These have a uniform composition throughout their entirety.

A heterogeneous mixture is one having a non-uniform composition throughout the entire mixture. Examples of such mixtures include the

mixture of sodium chloride and iron filings and a mixture of oil and water.

Stoichiometry and Stoichiometric Calculations

You may ask why do we need to measure the elements in **chemical reactions**? What happens when your mom adds extra salt to your favourite pasta? Do you eat it? No! Similarly, without a proper study of Stoichiometry, we can't know what **quantity** of elements are perfect for any suitable reaction. In this chapter, we will understand the concept of Stoichiometry and look how we can make the stoichiometric calculations from those reactions.

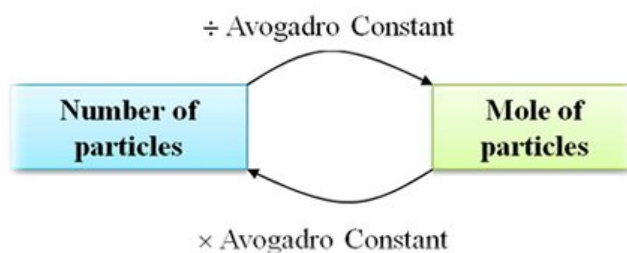
What is Stoichiometry?

The branch of stoichiometry deals with the calculation of various quantities of reactants or products of a chemical reaction. The word “stoichiometry” itself is derived from two Greek words “stoichion” that means element and “metry” means to measure. We have the following two sub-sections in this concept of stoichiometry.

- Gravimetric analysis

- Volumetric analysis

We can solve the problems on gravimetric and volumetric analysis by using the two well-known concepts of mole concept and the concept of equivalence. To understand Stoichiometric calculations, we must be able to understand the **relationship** between the various reactants and **products** in a chemical reaction.



For a reaction to be balanced, both sides of the **equation** should have an equal number of elements. We use various stoichiometry coefficients to adjust the number of each element on both sides of the reaction. These are the number of atoms that we write to balance the reaction.

Let us now discuss the various conversion factors to solve stoichiometric problems. We have to follow a few steps for the same.

They are:

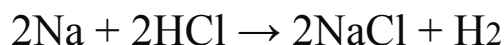
1. Balancing the given equation.
2. Giving the substance the unit as a mole.
3. Calculating the number of moles.

Stoichiometric Calculations are Based on Chemical Formulas

Let's learn some terms used in Stoichiometry first.

- **Formula Mass:** It is the sum of the atomic weights of the various atoms present in the molecule of the substance. For example, we can calculate the **formula** mass of Na_2S as $2(23) + 1(32) = 78$
- **Avogadro number:** Avogadro's number is the total number of **particles** present in one mole of a **substance**. It is the number of atoms present in exactly 12g of C-12. Avogadro number is valued as 6.022×10^{23} .
- **Molar Mass:** It is the sum of the total mass of all the atoms that constitute a molecule per mole.

We can explain the mole [ratio](#) of reactants and products with the help of the following reaction:



From the above reaction, we can see that 2 moles of Na react with 2 moles of HCl. This reaction forms 2 moles of NaCl (Sodium Chloride) and 1 mole of H₂. Hence, for a given amount of sodium, let us say x mole, we get the x mole of NaCl in reaction with x mole of HCl. The hydrogen gas produced will be x/2 moles.

Learn more about [the Percentage Composition here](#).

Watch Numericals on Stoichiometry

Solved Example for You

Q: What is a limiting reagent?

Ans: Limiting reagent is an important concept in a chemical calculation. It refers to a reactant which is present in a minimum stoichiometric quantity for a given chemical reaction. This particular reactant is fully consumed in the chemical reaction. So, we make all

the calculations related to various products or in a sequence of reactions on the basis of limiting reagent.

The limiting reagent is necessary to find out in a chemical reaction that involves more than two reactants. In that case, we need to understand which one is the limiting reagent for analysing the chemical reaction.

Uncertainty in Measurement

If we ask you what your **weight** is, you can easily say it. However, if we ask you what the value of π exactly is, there is uncertainty in **measurement**. Isn't it? In the subject of **chemistry**, a lot of times, we have to deal with both experimental and theoretical calculations. Therefore, we have to follow more than one methods to measure or calculate these number with minimum errors and uncertainty. In this chapter, we will deal with the concept of uncertainty in measurement.

Uncertainty in Measurement

Too often, we come across values that are close to each other and their average values. In such cases, we can say that the measurement is absolutely correct or precise. However, at times you may not

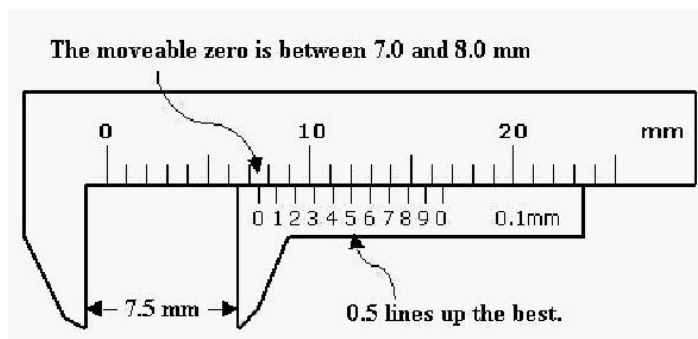
experience this. At all those times, you will have to mention the uncertainty in measurement.

Browse more Topics under Some Basic Concepts Of Chemistry

- [Atomic Mass and Molecular Mass](#)
- [Concentrations](#)
- [Dalton's Atomic Theory](#)
- [Importance and Scope of Chemistry](#)
- [Laws of Chemical Combination](#)
- [Mole and Equivalent Weight](#)
- [Nature of Matter](#)
- [Percentage Composition](#)
- [Properties of Matter and Their Measurement](#)
- [Stoichiometry and Stoichiometric Calculations](#)

Specifying this uncertainty is important as it will help you study the overall effect on [output](#). We indicate uncertainty through significant [figures](#). So, what are significant figures? A significant figure is the total number of digits in a number. It includes the last digit whose

value is uncertain. Let us now look at the practical and scientific notation of uncertainty in measurement.



Scientific Notation

We know that **atoms** and molecules have extremely low masses. However, we must not forget that they are present in massive numbers. Scientists have to deal with numbers that are as large as 123,456,789,101,110,000,000,987,000,870 and more.

Sometimes, they also have to deal with numbers as small as 0.000000000000000000000000166 g. Do you know what this is? Yes! It is the mass of Hydrogen atom. In science, there are other constants that have difficult figures. They include speed of light and charges on particles. So, how do we handle these numbers?

Handling These Numbers

We use notation like $m \times 10^n$ while handling these numbers. Here, we signify m times ten raised to the power of n . In this, we can also see that n is an exponent that has positive and negative values and m is that number which varies from 1.000... and 9.999...

Similarly, we can write the scientific notion of 578.677 as 5.78677×10^2 . In this, we move the decimal to the left by two spaces. If we move it three spaces to the left, the power of 10 becomes 3. In the same way, we can also write 0.000089 as 8.9×10^{-5} . In this, we move the decimal five places towards the right, (-5) is the **exponent**.

This method eases the handling for us and gives better precision. We arrive at results that are more accurate when we are dealing with high magnitude numbers.

Uncertainty in Multiplication and Division

We can apply the same rules as above, for the methods of **multiplication and division** too. For e.g: $(3.9 \times 10^6) \times (2.1 \times 10^5) = (3.9 \times 2.1)(10^{6+5}) = (3.9 \times 2.1) \times (10^{11}) = 11.31 \times 10^{11}$

Solved Example for You

Q: Explain the uncertainty in addition and subtraction with an example.

Solution: In addition and subtraction, we need to place these numbers in such a way that they have same exponents. This will be the first step of our [solution](#). Therefore, when we add 5.43×10^4 and 3.45×10^3 , we make the powers equal first. After that, we add and subtract the coefficients.

For example: $5.43 \times 10^4 + 0.345 \times 10^5 = (5.43 + (0.345 \times 10)) \times 10^4 = 8.88 \times 10^4$.

In the case of subtraction, $5.43 \times 10^4 - 0.345 \times 10^5 = (5.43 - (0.345 \times 10)) \times 10^4 = (5.43 - 3.45) \times 10^4 = 1.98 \times 10^4$