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CBSE LIVING SCIENCE CHEMISTRY

CLASS 9

Chapter 3

Atoms and Molecules

Learning Objectives

- ❖ Particle nature of matter
- ❖ Laws of chemical combination – law of conservation of mass, law of constant proportions, law of Multiple proportions
- ❖ Atoms
- ❖ Dalton's atomic theory
- ❖ Molecules
- ❖ Ions
- ❖ Atomic mass
- ❖ Molecular mass
- ❖ Mole concept

INTRODUCTION

It is essential to know about atoms and molecules in order to understand matter within and outside our body. Chemistry deals with the composition, structure and properties of matter. These aspects are best described and understood in terms of the basic constituents of matter, i.e. atoms and molecules.

LAWS OF CHEMICAL COMBINATION

An Indian philosopher Maharishi Kanad (400 BC) postulated that if matter keeps on dividing then a stage will come when the particles obtained by division would no longer be divided further, i.e. further division will not be possible. These particles were called *parmanu* by him. The word atom (meaning indivisible) was first used by Democritus for these indivisible particles. Three important laws of chemical combination given by Antoine L Lavoisier and Joseph L Proust are discussed further.

Law of conservation of mass

According to the law of conservation of mass (Lavoisier, 1774), during a chemical reaction, matter is neither created nor destroyed but remains conserved. This means that in a chemical reaction, the total mass of the products is equal to the total mass of the reactants.
For example, in the reaction of hydrogen (H_2) and chlorine (Cl_2) represented below, 2 g of H_2 reacts with 71 g of Cl_2 to give 73 g of HCl.

$$\begin{array}{lccccc} & H_2 & + & Cl_2 & \longrightarrow & 2HCl \\ \text{Molecular mass} & 2 & & 71 & & 2 \times 36.5 = 73 \end{array}$$

The above molecular mass data indicates that the mass of the reactants is equal to the mass of the products, i.e. the total mass is conserved in the reaction. The law of conservation of mass can be utilised to find out the mass of any reactant or product.

Law of constant proportions

The law of constant proportions is also known as the law of definite proportions. This law, proposed by Proust in 1797, states that whatever be the method of its formation or whatever may be its source, a chemical compound always consists of the same elements combined together in the same proportions by mass. This law means that a chemical compound will always have its elements combined in a fixed ratio by mass.
Let us take a real example. Carbon dioxide gas (CO_2) is always found to contain only carbon and oxygen. The ratio in which carbon and oxygen are present in carbon dioxide is always found to be fixed at 3:8 (actual is 12:32) by mass, independent of the source of carbon dioxide. Thus, if 3.0 g of carbon is burnt, it will always combine with 8.0 g of oxygen to form 11.0 g of carbon dioxide.

Also, the compound calcium oxide (CaO) can be prepared by the following three independent methods:

a. By heating calcium carbonate strongly:

$$CaCO_3 \longrightarrow CaO + CO_2$$

b. By heating calcium hydroxide strongly:

$$Ca(OH)_2 \longrightarrow CaO + H_2O$$

c. By heating calcium nitrate strongly:

$$2Ca(NO_3)_2 \longrightarrow 2CaO + 4NO_2(\uparrow) + O_2(\uparrow)$$

The analysis of calcium oxide thus prepared by different methods and different sources shows that it contains the elements calcium and oxygen only, and the ratio by mass of calcium to oxygen is always fixed, that is 5:2.

Law of multiple proportions

This law, proposed by Dalton in 1803, states that when two elements combine to form two or more compounds, then the different masses of one element which combine with a fixed mass of the other, bear a simple ratio to one another.

Oxygen combines with hydrogen to form two oxides. The oxygen content in one oxide (water) is 88.89% while that in the other oxide (hydrogen peroxide) is 94.12%.

For H_2O:

Mass % of oxygen = 88.89

Mass % of hydrogen = 100 – 88.89 = 11.11

Thus,

11.11 g of hydrogen reacts with 88.89 g of oxygen

$$1 \text{ g of hydrogen reacts with } \frac{88.89}{11.11} \text{ g of oxygen} = 8.00 \text{ g of oxygen}$$

For H_2O_2:

Mass % of oxygen = 94.12

Mass % of hydrogen = 100 – 94.12 = 5.88

Thus,

5.88 g of hydrogen reacts with 94.12 g of oxygen

1 g of hydrogen reacts with $\frac{94.12}{5.88}$ g of oxygen

= 16.00 g of oxygen

The ratio of oxygen masses which combine with 1 g of hydrogen = 8.00 : 16.00 = 1 : 2

Since 1 : 2 is a simple ratio, the law of multiple proportions is supported by the above data.

Example 1 12.0 g of calcium carbonate was heated till no further loss in weight and 5.28 g of carbon dioxide was produced. The mass of the residue (calcium oxide) was found to be 6.72 g. Show that these observations are in accordance with the law of conservation of mass.

Solution

$$\underset{12.0\text{ g}}{\text{Calcium carbonate}} = \underset{6.72\text{ g}}{\text{Calcium oxide}} + \underset{5.28\text{ g}}{\text{Carbon dioxide}}$$

Mass of reactant = 12.0 g

Mass of products = 6.72 g + 5.28 g = 12.0 g

Since the total mass of the products is equal to the total mass of the reactant, the observations are in agreement with the law of conservation of mass.

Example 2 In one experiment, 6.54 g of iron(III) oxide was reduced with hydrogen to give 4.73 g of iron. In another experiment, 2.87 g of iron was heated and treated with steam when 3.97 g of iron(III) oxide was obtained. Show that these data illustrate the law of constant proportions

Solution

Expt. No.	Mass of iron	Mass of iron(III) oxide	Mass of oxygen
1.	4.73 g	6.54 g	6.54 – 4.73 = 1.81 g
2.	2.87 g	3.97 g	3.97 – 2.87 = 1.10 g

In the experiment no. 1, the ratio of Fe : O

$$= 4.73 : 1.81 = 1 : \frac{1.81}{4.73} = 1 : 0.38$$

In the experiment no. 2, the ratio of Fe : O

$$= 2.87 : 1.10 = 1 : \frac{1.10}{2.87} = 1 : 0.38$$

Thus, the ratio of the masses of Fe : O in each experiment is constant at 1 : 0.38. Hence, the given data is in agreement with the law of constant proportions.

Example 3 Carbon combines with oxygen to form two oxides. The carbon content in one of the oxides is 42.9% while in the other oxide it is 27.3%. Show that these data are in agreement with the law of multiple proportion

Solution For the first oxide, mass % of carbon = 42.9 and mass % of oxygen = 100 – 42.9 = 57.1

42.9 g of carbon reacts with 57.1 g of oxygen

1 g of carbon reacts with $\frac{57.1}{42.9}$ g of oxygen

= 1.33 g of oxygen

For the second oxide, mass % of carbon = 27.3 and mass % of oxygen = 100 – 27.3 = 72.7.

27.3 g of carbon reacts with 72.7 g of oxygen

1 g of carbon reacts with $\frac{72.7}{27.3}$ g of oxygen

= 2.66 g of oxygen

Ratio of oxygen masses which combine with 1 g of carbon = 1.33 : 2.66 = 1 : 2

Since 1 : 2 is a simple ratio, the law of multiple proportion is supported by the data.

ATOMS AND ATOMIC THEORY OF MATTER

The word 'atom' is derived from the Greek word '*atomos*' meaning '*indivisible*'. This idea about atom was only speculative and it was not verified experimentally. It was in 1805, John Dalton, a British school teacher and now known as the father of modern atomic theory, put the atomic model of matter on a firm scientific basis and considered atom as the ultimate particle of matter incapable of being created or destroyed.

Postulates of Dalton's atomic theory

The main postulates of Dalton's atomic theory are as follows:

1. Matter is composed of very tiny particles called atoms.
2. Atoms are incapable of being destroyed or created.
3. Atoms of any pure substance can neither be subdivided nor changed into atoms of another element.
4. All atoms of an element are identical in terms of mass, size and other properties.
5. Atoms of one element differ in mass and other characteristics from those of other elements.
6. Chemical combination results in the union of atoms in numerical proportions to form a large number of new substances.
7. The relative number and kinds of atoms are constant in a given compound.

The modern atomic theory correlates Dalton's atomic theory in the fact that the atoms are the smallest particles of matter which take part in chemical reactions, and in a given compound the relative number and kind of atoms are constant.

The modern atomic theory contradicts some of the postulates (e.g., indivisibility of atom) of Dalton’s atomic theory.

Drawbacks of Dalton’s atomic theory

1. Atom is no longer considered as the smallest indivisible particle. It was found that an atom was composed of smaller subatomic particles like electron, proton and neutron.
2. Atoms of the same element have different masses. For example, chlorine has two types of atoms with masses 35 u and 37 u. Such atoms of an element with different masses are called isotopes.
3. Atoms of some different elements have same masses. The atoms of different elements with same mass numbers are called isobars. For example, the atoms of potassium and calcium have the same mass number (40).
4. Substances made up of same kind of elements have different properties. For example, diamond, graphite and charcoal are made up of carbon atoms but they possess different physical properties.
5. The ratio in which the different atoms combine to form a compound is fixed and integral but may not be simple. For example, a molecule of sugar having the formula $C_{12}H_{22}O_{11}$ contains C, H and O in the ratio 12 : 22 : 11. The ratio is fixed and integral but is not simple

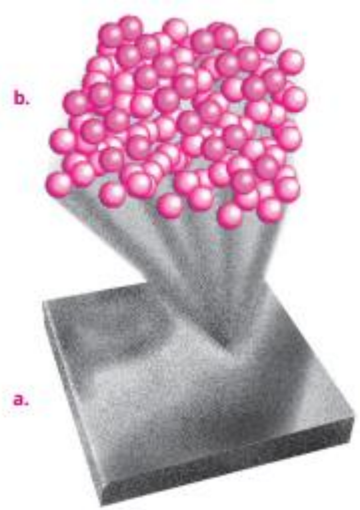
What is an atom?

An atom is the smallest particle of an element which takes part in a chemical reaction and maintains its chemical identity throughout all physical and chemical changes. Atoms are the building blocks of all matter. It has now been proved that atoms are made up of even smaller particles called electrons, protons and neutrons. Under normal conditions, the atoms of elements except those of noble gases such as helium, neon, argon, etc. are not capable of independent existence.

How big are atoms?

The size of atoms is very small. The idea of size of an atom is obtained from the radius of the atom. The radius of an atom is measured in nanometer (nm) [$1 \text{ nm} = 10^{-9}$ m]. The smallest atom is hydrogen atom, whose radius is of the order of 10^{-10} m.

We cannot see atoms and molecules even with the help of the most powerful microscope available on the earth. Modern techniques such as Scanning Tunneling Electron Microscopy (STEM) have been used to record the magnified images.

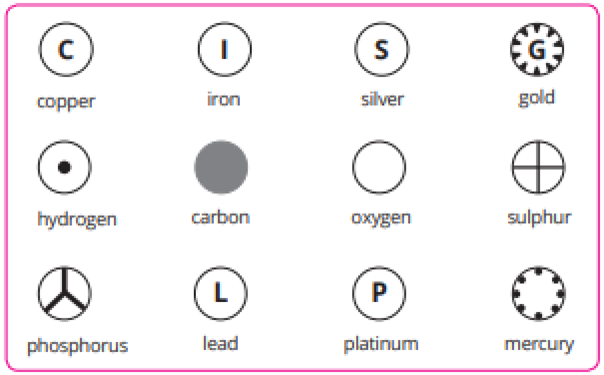


a. An image of the surface of silicon

b. The arrangement of atoms of silicon on the surface of silicon

Symbol of atoms

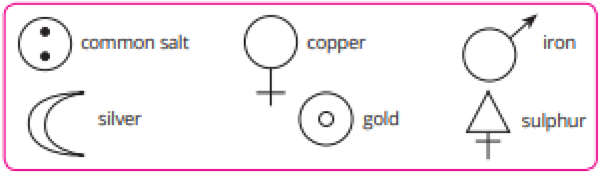
Elements are represented by symbols. Dalton used the circular symbols for representing the atoms of various elements. He differentiated the atoms of different elements by putting different markings inside the circles.



Dalton’s symbols for atoms of various elements

Alchemist's symbols

Alchemists or the chemists of the old time also used some symbols for elements. The symbols used by them were of strange type as they wanted to hide their knowledge from others. These symbols are shown here.



Most of the universally accepted symbols for elements at present were first introduced by the Swedish chemist J J Berzelius (1779–1848). He suggested that the symbols of elements be made from one or two letters of the name of the element. So generally, the symbol of an element is the first letter of the name of the element (either the old Latin or Greek name or the modern name). If there are many elements having the same first letter, the first and the second (or the first and third) letters are used as the symbol of an element. The first letter is always a capital letter and the second letter is a lowercase letter (if a symbol consists of two letters). For example, the symbol for aluminium is Al and not AL and for cobalt is Co and not CO.

Now the names of elements are approved by the International Union of Pure and Applied Chemistry (IUPAC). For example, the symbol of the yet to be discovered element with the atomic number 119 is Uue (Ununennium).

The symbol of an element is thus an abbreviation (short name) for the full name of the element, or the symbol of an element is a shorthand notation for its name.

Significance of the symbol of an element

1. Symbol represents name of the element.
2. It represents one atom of the element.
3. It represents a definite mass of the element (equal to atomic mass expressed in gram).
4. It represents mass of the element which contains one Avogadro's number (6.022×10^{23}) of atoms of that element.

Symbols of some common elements

Modern name	Symbol	Latin or Greek name
Aluminium	Al	
Antimony	Sb	Stibium (*Latin*)
Argon	Ar	Argon (*Greek*)
Arsenic	As	
Barium	Ba	Barys (*Greek*)
Boron	B	
Bromine	Br	Bromos (*Greek*)
Calcium	Ca	Calx (*Latin*)
Carbon	C	Carbonium (*Latin*)
Chlorine	Cl	Chloros (*Greek*)
Chromium	Cr	Chrom (*Greek*)
Cobalt	Co	Cobold (*German*)
Copper	Cu	Cuprum (*Latin*)
Fluorine	F	Fluo (*Latin*)
Gold	Au	Aurum (*Latin*)
Hydrogen	H	Hydrogenium (*Latin*)
Iodine	I	Iodes (*Greek*)
Iron	Fe	Ferrum (*Latin*)
Krypton	Kr	Kryptos (*Greek*)
Lead	Pb	Plumbum (*Latin*)
Magnesium	Mg	
Molybdenum	Mo	Molybdos (*Greek*)

Modern name	Symbol	Latin or Greek name
Mercury	Hg	Hydrargyrum (*Latin*)
Neon	Ne	Neos (*Greek*)
Nickel	Ni	
Nitrogen	N	Nitrogenium (*Latin*)
Oxygen	O	Oxygenium (*Latin*)
Phosphorus	P	Phosphoros (*Greek*)
Polonium	Po	
Potassium	K	Kalium (*Latin*)
Platinum	Pt	
Selenium	Se	Selene (*Greek*)
Silicon	Si	Silex (*Latin*)
Silver	Ag	Argentum (*Latin*)
Sodium	Na	Natrium (*Latin*)
Strontium	Sr	
Sulphur	S	Sulfur (*Latin*)
Tantalum	Ta	Tantalos (*Greek*)
Tin	Sn	Stannum (*Latin*)
Titanium	Ti	Titan (*Latin*)
Tungsten	W	Wolfram (*Latin*)
Uranium	U	
Vanadium	V	
Xenon	Xe	Xenon (*Greek*)
Zinc	Zn	Zink (*Greek*)

Relative atomic mass (A_r)

The mass of an atom is the sum of the masses of protons ($m_p = 1.67262 \times 10^{-24}$ g) and neutron ($m_n = 1.67493 \times 10^{-24}$ g). The mass of electron ($m_e = 9.10939 \times 10^{-28}$ g) is neglected since it is 1/1837 times the mass of a proton. Since atoms are extremely small particles, it is very difficult to measure their actual masses. For example, the mass of one hydrogen atom is 1.672×10^{-24} g. Such a small mass cannot be measured accurately even with the help of a very sensitive balance.

In order to overcome such problems, the atomic masses are expressed as relative atomic masses with reference to the mass of a standard reference atom. The reference atom is ^{12}C which has been assigned an atomic mass of 12.000 atomic mass unit. The mass equal to $1/12^{th}$ of the mass of a ^{12}C atom is called one atomic mass unit. The atomic mass unit is abbreviated as amu and is denoted by the symbol ‘u’ (unified mass). Hence,

$$1 \text{ atomic mass unit} = 1\text{ u} = \frac{\text{Mass of one } ^{12}\text{C atom}}{12}$$

The absolute mass of one ^{12}C atom

$$= 1.9924 \times 10^{-23}\text{ g}$$

$$\text{Hence,}\quad 1\text{ u} = \frac{1.9924 \times 10^{-23}}{12}\text{ g} = 1.66 \times 10^{-24}\text{ g}$$

Thus, the mass of one ^{12}C atom is 12 u.

The atomic mass (A) of an element is defined as the average mass of an atom of an element in atomic mass unit. Due to the presence of isotopes, the elements have fractional atomic masses. Actually the atomic mass of an element is the weighted average of the atomic masses of all the isotopes of an element.

Since atomic mass is the actual mass, it has the unit of mass, viz. gram, kilogram, atomic mass unit, etc.

The relative atomic mass (A_r) of an element is the average mass of an atom of an element as compared to 1/12th the mass of one carbon-12 atom. Hence,

$$\text{Relative atomic mass of an element } (A_r) = \frac{\text{Average mass of one atom of the element}}{\frac{1}{12} \times (\text{Mass of one } ^{12}\text{C atom})}$$

The relative atomic mass (A_r) of an element is a pure number and it does not have a unit. The relative atomic masses of some elements taking ^{12}C (m = 12.000 u) as reference are given.

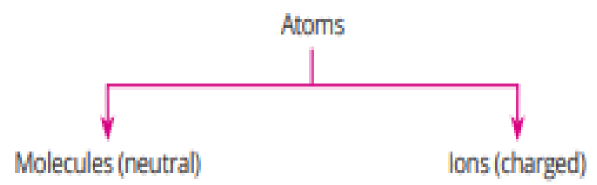
Element name	Symbol	Relative atomic mass (A_r)	
		Exact value	Common value*
Aluminium	Al	26.982	27
Antimony	Sb	121.750	121
Argon	Ar	39.948	40
Arsenic	As	74.922	75
Barium	Ba	137.330	137
Beryllium	Be	9.012	9
Bismuth	Bi	208.980	209
Boron	B	10.810	11
Bromine	Br	79.904	80
Cadmium	Cd	112.410	112
Calcium	Ca	40.080	40
Carbon	C	12.011	12
Caesium	Cs	132.906	133
Chlorine	Cl	35.453	35.5
Chromium	Cr	51.996	52
Cobalt	Co	58.933	59
Copper	Cu	63.546	63.5
Fluorine	F	18.918	19
Gallium	Ga	69.720	70
Germanium	Ge	72.590	72.5
Gold	Au	196.966	197

Element name	Symbol	Relative atomic mass (A_r)	
		Exact value	Common value*
Helium	He	4.003	4
Hydrogen	H	1.008	1
Indium	In	114.820	115
Iodine	I	126.904	127
Iridium	Ir	192.220	192
Iron	Fe	55.847	56
Krypton	Kr	83.800	84
Lanthanum	La	138.906	139
Lead	Pb	207.200	207
Magnesium	Mg	24.305	24
Manganese	Mn	54.938	55
Mercury	Hg	200.59	200
Molybdenum	Mo	95.94	96

Element name	Symbol	Relative atomic mass (A_r)	
		Exact value	Common value*
Neon	Ne	20.179	20
Nickel	Ni	58.700	59
Niobium	Nb	92.906	93
Nitrogen	N	14.007	14
Osmium	Os	190.200	190
Oxygen	O	15.999	16
Phosphorus	P	30.974	31
Platinum	Pt	195.090	195
Potassium	K	39.098	39
Radium	Ra	226.025	226
Rhenium	Re	186.207	186
Rhodium	Rh	102.906	103
Rubidium	Rb	85.468	85.5
Ruthenium	Ru	101.070	101
Scandium	Sc	44.956	45
Selenium	Se	78.960	79
Silicon	Si	28.086	28
Silver	Ag	107.868	108
Sodium	Na	22.980	23

How do atoms exist?

At present 118 elements are known to us. Each element is composed of atoms only. The atoms of each element are different from the atoms of other elements. Also the atom is the smallest unit of an element which can take part in a chemical reaction. Actually the atoms of these 118 elements combine in so many different ways that give rise to the substances around us. The atoms of most elements, except the noble gases such as helium, neon, argon, krypton and xenon, do not exist independently. Atoms combine together to form molecules and ions. Molecules and ions differ in terms of the charge they are carrying – while the molecules are neutral compounds, the ions possess either positive or negative charge.



Large numbers of molecules or ions aggregate together to form various types of substances we see around us. The compounds containing molecules are called molecular compounds. The compounds containing ions are called ionic compounds. Some molecules also interact with some ions to give new substances.

MOLECULES AND THEIR CHEMICAL FORMULAE

Molecules

A molecule is the smallest particle of an element or a compound which can exist in the free state independently under ordinary conditions and exhibits all the properties of the substance (element or compound).

Chemical formula

A chemical formula is the symbolic representation of a molecule of the compound. It denotes the number of atoms of different elements present in one molecule of the compound. The formula of a compound indicates the fixed proportion in which, by mass, the atoms combine.

Chemical formulae are of two types – molecular formulae and empirical formulae.

Molecular formula

The symbolic representation of a molecule of a substance representing the actual number of various atoms present in it is called the molecular formula. For example, the molecular formula of carbon dioxide is CO_2, containing one atom of carbon and two atoms of oxygen. The number of atoms of all the elements present in a molecule of a substance is known as atomicity of that molecule.

- Molecules containing two atoms are called diatomic molecules.
- Molecules containing three atoms are called triatomic molecules.
- Molecules containing four atoms are called tetratomic molecules.
- Molecules containing more than four atoms are called polyatomic molecules.

Types of molecule	No. of atoms present	Examples
Monoatomic	1	He, Ne, Ar, Kr, Xe
Diatomic	2	H_2, N_2, O_2, F_2, Cl_2, Br_2, I_2, HF, HCl, HBr, HI, CO
Triatomic	3	O_3, CO_2, H_2O, NO_2, SO_2
Tetratomic	4	P_4, NH_3, PH_3, SO_3, H_2O_2
Polyatomic	> 4	S_8, Se_8, CH_4

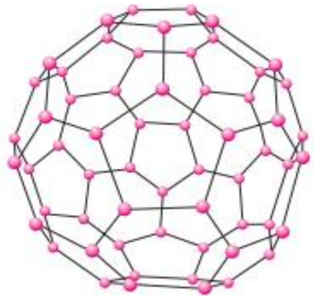
Different types of molecules

Significance of molecular formula

1. It indicates the names of various elements present in the compound.
2. It indicates the number of various atoms present in one molecule of the compound.
3. Relative molecular mass of the compound can be calculated from the molecular formula.
4. Mass of each element present in one mole of the compound can be found out from the molecular formula. Thus, the percentage composition of each element in the compound can be found out.
5. Molecular formula gives the number of gram-atoms of each element present in one mole of the compound.

Molecules of elements

Molecules of elements consist of the same type of atoms. For example, hydrogen (H_2), oxygen (O_2), nitrogen (N_2), etc. The atomicity of noble gases such as helium (He), neon (Ne), argon (Ar), krypton (Kr) and xenon (Xe) is 1 each and these elements are monoatomic. The atomicity of oxygen (O_2), nitrogen (N_2), fluorine (F_2), chlorine (Cl_2), bromine (Br_2) and iodine (I_2) is 2 each and these elements are diatomic. The atomicity of ozone (O_3) is 3 and it is triatomic. The atomicity of phosphorus (P_4) and sulphur (S_8) is 4 and 8, respectively and they are tetratomic and polyatomic (octatomic), respectively. The molecules of carbon (diamond, graphite, buckminsterfullerene) do not possess simple structures. Their molecules contain a very large number of atoms chemically bonded together. The simplest of these molecules is buckminsterfullerene which contains sixty carbon atoms (C_{60}) bonded together to produce a structure similar to a football. The properties of such solid element are not the properties of its single atom but the properties of the cluster of atoms.



Structure of buckminsterfullerene

Molecules of compounds

Molecules of compounds consist of different types of atoms. These different atoms are bonded together in definite fashions and in definite proportions. For example, carbon dioxide has a linear structure (O=C=O), water has a bent structure , ammonia has a trigonal pyramidal structure and methane has a tetrahedral structure.

Ions and ionic compounds

The charged species formed when an atom gains or loses electrons is called ion. An ion can be positively or negatively charged. A positively charged ion is called cation and a negatively charged ion is called anion.

A cation is formed when an atom loses one or more electrons. For example, when the potassium atom (K) loses one electron, the potassium ion (K^+) is formed.

$$\underset{\text{Potassium atom}}{K(s)} \longrightarrow \underset{\text{Potassium ion}}{K^+} + \underset{\text{Electron}}{e^-}$$

When the chlorine atom (Cl) gains one electron, the chloride ion (Cl^-) is formed.

$$\underset{\text{Chlorine atom}}{Cl(g)} + \underset{\text{Electron}}{e^-} \longrightarrow \underset{\text{Chloride ion}}{Cl^-}$$

Examples of cations are Na^+, K^+, Ca^{2+}, Al^{3+}, etc. Examples of anions are Cl^-, Br^-, I^-, O^{2-}, S^{2-}, etc.

An ion exists as a part of a compound and not independently. This is because of the presence of electrostatic force of attraction between the cations and the anions.

The properties of an ion are different from those of the parent element(s). For example, chlorine gas is poisonous but the chloride ion (Cl^-) is non-poisonous.

An ionic compound contains cations and anions. The examples of ionic compound are NaCl, KCl, Na_2SO_4, K_2SO_4, $MgSO_4$, $AlPO_4$, etc.

An ionic compound is formed when a metallic element reacts with a non-metallic element. For example, when potassium metal reacts with chlorine gas (a nonmetal) the ionic compound, potassium chloride is formed.

The formation of potassium chloride is illustrated below:

$$\underset{\text{Potassium atom}}{K(s)} \longrightarrow \underset{\text{Potassium ion}}{K^{+}} + \underset{\text{Electron}}{e^{-}} \quad \ldots (1)$$

$$\underset{\text{Chlorine atom}}{Cl(g)} + \underset{\text{Electron}}{e^{-}} \longrightarrow \underset{\text{Chloride ion}}{Cl^{-}} \quad \ldots (2)$$

$$\underset{\substack{\text{Potassium}\\ \text{ion}}}{K^{+}} + \underset{\substack{\text{Chloride}\\ \text{ion}}}{Cl^{-}} \longrightarrow \underset{\substack{\text{Potassium chloride}\\ \text{(an ionic compound)}}}{K^{+}Cl^{-}}$$

The electron generated by potassium atom [Eq. (1)] is transferred to the chlorine atom [Eq. (2)].

In the formation of ionic compounds, the combining elements react in a definite ratio by mass. The ratio by mass in some ionic compounds are as follows:

Ratio by mass	Ionic compound	Elements present
23 : 35.5	Sodium chloride ($NaCl$)	Sodium, Chlorine
39 : 35.5	Potassium chloride (KCl)	Potassium, Chlorine
24 : 16 = 3 : 2	Magnesium oxide (MgO)	Magnesium, Oxygen
40 : 32 = 5 : 4	Calcium sulphide (CaS)	Calcium, Sulphur

Ratio by mass in some ionic compounds

Most of the ionic compounds are soluble in water and insoluble in organic solvents such as ethanol, chloroform, benzene, etc. They have high melting and boiling points.

Empirical formula of a compound

Empirical formula of a compound is the simplest formula which gives the simplest ratio in whole numbers between the number of atoms of different elements present in one molecule of the compound. For example, the empirical formula of benzene is CH. It indicates that the simplest ratio between the carbon and hydrogen atoms in its molecule is 1 : 1 whereas its actual formula is C_6H_6. Therefore, the empirical formula of the benzene having molecular formula of C_6H_6 is CH.

Empirical formula of a compound does not indicate the actual number of atoms of the elements present in the compound. It only gives the simplest whole number ratio between the number of atoms of all the elements present in the compound.
Empirical formula mass is the sum of the atomic masses of various elements representing the empirical formula. Thus, for benzene it is 13 (12 + 1).
The relation between empirical formula and molecular formula of a compound is given by:

Molecular formula = n × Empirical formula

where n is an integer

The value of n is obtained by the following relationship:

$$n = \frac{\text{Molecular mass}}{\text{Empirical formula mass}}$$

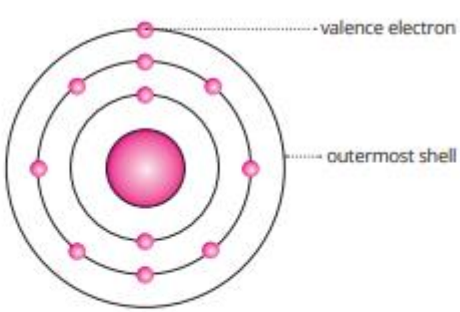
Writing of chemical formulae

The following procedure is followed for writing the formulae of compounds:

1. The symbols of the constituent elements are written so that the symbol of the positive atom or radical is placed on the left-hand side and the symbol of the negative atom or radical is placed on the right-hand side.
2. The valencies of the symbols are written below them and shifted crosswise to the lower right corner of the symbols and the valency numbers become subscripts of the corresponding symbols.
3. The valency numbers are reduced to a simple ratio by dividing with a common factor, if required.

Valence electrons and valency

Valency of an atom or radical is the number of electrons a particle loses or gains or shares during the course of a chemical reaction. The particles may be atoms or group of atoms. The electrons present in the outermost shell of an atom are called **valence electrons**. The valence electrons dictate the valency and chemical reactions of an element. The outermost shell of an atom is called its valence shell.



Electronic configuration of sodium atom showing valence electrons

Metals such as sodium, magnesium and aluminium lose one, two or three valence electrons, respectively to form cations. Thus, the valency of sodium, magnesium and aluminium is 1, 2 and 3, respectively.

$$\underset{\text{Sodium atom}}{\underset{(2,\,8,\,1)}{Na}} \xrightarrow{-\,1\text{ electron}} \underset{\text{Sodium ion}}{\underset{(2,\,8)}{Na^{+}}}$$

$$\underset{\text{Magnesium atom}}{\underset{(2,\,8,\,2)}{Mg}} \xrightarrow{-\,2\text{ electrons}} \underset{\text{Magnesium ion}}{\underset{(2,\,8)}{Mg^{2+}}}$$

$$\underset{\text{Aluminium atom}}{\underset{(2,\,8,\,3)}{Al}} \xrightarrow{-\,3\text{ electrons}} \underset{\text{Aluminium ion}}{\underset{(2,\,8)}{Al^{3+}}}$$

Non-metals such as nitrogen, oxygen and fluorine gain three, two and one electrons, respectively in their valence shell to form anions. Thus, the valency of nitrogen, oxygen and fluorine is 3, 2 and 1, respectively.

$$\underset{\text{Nitrogen atom}}{\underset{(2,\,5)}{N}} \xrightarrow{+\,3\text{ electrons}} \underset{\text{Nitride ion}}{\underset{(2,\,8)}{N^{3-}}}$$

$$\underset{\text{Oxygen atom}}{\underset{(2,\,6)}{O}} \xrightarrow{+\,2\text{ electrons}} \underset{\text{Oxide ion}}{\underset{(2,\,8)}{O^{2-}}}$$

$$\underset{\text{Fluorine atom}}{\underset{(2,\,7)}{F}} \xrightarrow{+\,1\text{ electron}} \underset{\text{Fluoride ion}}{\underset{(2,\,8)}{F^{-}}}$$

Valency is the combining capacity of an element. It can be defined as:

The number of hydrogen atoms or chlorine atoms or twice the number of oxygen atoms that combine with one atom of an element to form a compound is called valency of that element.

For example, two atoms of hydrogen combine with one atom of oxygen to form a molecule of water (H_2O). Hence, the valency of oxygen is 2

Element	Name of anion	Anion	Valency
Chlorine	Chloride	Cl^-	1
Fluorine	Fluoride	F^-	1
Sulphur	Sulphide	S^{2-}	2
Bromine	Bromide	Br^-	1
Oxygen	Oxide	O^{2-}	2
Iodine	Iodide	I^-	1
Nitrogen	Nitride	N^{3-}	3

Ions of common non-metals and their valency

Element	Cation	Valency
Sodium	Na^+	1
Potassium	K^+	1
Cobalt	Co^{2+}	2
Nickel	Ni^{2+}	2
Silver	Ag^+	1
Zinc	Zn^{2+}	2
Magnesium	Mg^{2+}	2
Tin (stannous)	Sn^{2+}	2
Tin (stannic)	Sn^{4+}	4
Barium	Ba^{2+}	2
Calcium	Ca^{2+}	2
Lead (plumbous)	Pb^{2+}	2
Lead (plumbic)	Pb^{4+}	4
Copper (cuprous)	Cu^+	1
Copper (cupric)	Cu^{2+}	2
Mercury (mercurous)	Hg^+	1
Mercury (mercuric)	Hg^{2+}	2
Manganese	Mn^{2+}	2
Chromium	Cr^{3+}	3
Arsenic	As^{3+}	3

Ions of some common metals and their valencies

The elements exhibiting the valency of 0, 1, 2, 3, 4, 5, 6, 7 and 8 are called zerovalent, monovalent, divalent, trivalent, tetravalent, pentavalent, hexavalent, heptavalent and octavalent, respectively.

When two or more atoms combine by sharing of electrons to form covalent compounds, the valencies of elements do not have a charge, i.e. neither positive nor negative. For example, in carbon dioxide (CO_2), the valency of carbon and oxygen is 4 and 2, respectively. Valencies of some common anions are given here.

Anion	Symbol	Valency
Hydrogencarbonate	HCO_3^-	1
Manganate	MnO_4^{2-}	2
Hydrogensulphate	HSO_4^-	1
Permanganate	MnO_4^-	1
Sulphite	SO_3^{2-}	2
Acetate	CH_3COO^-	1
Sulphate	SO_4^{2-}	2
Nitrate	NO_3^-	1
Carbonate	CO_3^{2-}	2
Nitrite	NO_2^-	1
Peroxide	O_2^{2-}	2
Phosphate	PO_4^{3-}	3

Valencies of some common anions

Chemical formulae of some compounds

1. The symbols of the constituent elements are written so that the symbol of the positive atom or radical is placed on the left-hand side and the symbol of the negative atom or radical is placed on the right-hand side.
2. The valencies of the symbols are written below them and shifted crosswise to the lower right corner of the symbols and the valency numbers become subscripts of the corresponding symbols.
3. The valency numbers are reduced to a simple ratio by dividing with a common factor, if required.

The procedure discussed is illustrated as follows by writing the formulae of few compounds.

1.												
Element	H	Cl	H	O	N	H	C	H	C	O	Si	O
Valency	1	1	1	2	3	1	4	1	4	2	4	2
Formula	HCl		H_2O		NH_3		CH_4		C_2O_4 or CO_2		Si_2O_4 or SiO_2	

2.												
Element	Mg^{2+}	NO_3^-	K^+	SO_4^{2-}	Ca^{2+}	Cl^-	Fe^{3+}	CO_3^{2-}	Ca^{2+}	SO_4^{2-}	K^+	PO_4^{3-}
Valency	2	1	1	2	2	1	3	2	2	2	1	3
Formula	$Mg(NO_3)_2$		K_2SO_4		$CaCl_2$		$Fe_2(CO_3)_3$		$CaSO_4$		K_3PO_4	

MOLECULAR MASS

Relative molecular mass (M_r)

The molecular masses of compounds are also expressed as relative molecular masses (M_r). The relative molecular mass of a compound is the average mass of its one molecule as compared to 1/12th the mass of one carbon-12 atom.

$$\text{Relative molecular mass } (M_r) = \frac{\text{Average mass of one molecule of the compound}}{\frac{1}{12} \times (\text{Mass of an atom of } ^{12}\text{C})}$$

The relative molecular mass (M_r) of a compound is a pure number and it does not have a unit.

Molecular mass (M)

The average mass of one molecule of a compound in atomic mass unit is called molecular mass (M). Hence,

Molecular mass (M) = $M_r \times 1\ \text{u} = M_\text{r}\ \text{u}$

The molecular mass (M) has the unit of mass, i.e. g, kg or u. Note that the magnitudes of molecular mass (M) and relative molecular mass (M_r) are equal. They differ only in their units.

Calculation of molecular mass from atomic masses

The molecular mass of a compound is calculated by adding the atomic masses of all the atoms present in one molecule of the compound. This is illustrated below.

1. Molecular mass of water: The molecular formula of water is H_2O. Hence, Molecular mass of water = (2 × Atomic mass of hydrogen) + (1 × Atomic mass of oxygen)
= (2 × 1) + (1 × 16) = 18 u.

2. Molecular mass of nitric acid (HNO_3): The molecular formula of nitric acid is HNO_3. Hence, Molecular mass of nitric acid = (1 × Atomic mass of hydrogen) + (1 × Atomic mass of nitrogen) + (3 × Atomic mass of oxygen) = (1 × 1) + (1 × 14) + (3 × 16) = 1 + 4 + 48 = 63 u.

Formula unit mass of compounds

The sum of the atomic masses of all the atoms in a formula unit of a compound is called the formula unit mass of the compound. The concept of the formula unit mass is used in case of ionic compounds. This is because in case of ionic compounds there are no separate individual molecules but they exist as aggregates, i.e. as cluster of ions.

The formula unit mass of ionic compounds is calculated in the same manner as in the case of the molecular mass described above. For example, the formula unit mass of NaCl = Atomic mass of Na + Atomic mass of Cl = 23 u + 35.5 u = 58.5 u

Percentage composition of a compound

The percentage composition of a compound is the mass of each element of the compound, present in 100 g of that compound, i.e. the mass percentage of each element present in the compound. The mass percentage of each element in a compound can be calculated using either of the following two equations:

When the masses of compound and each element are given:

$$\text{Mass percentage of an element X} = \frac{\text{Mass of element in the given mass of the compound}}{\text{Total mass of the compound}} \times 100$$

When the formula of the compound and the atomic masses of the elements are given:

$$\text{Mass percentage of an element} = \frac{\text{Total mass of the element in one molecule of the compound}}{\text{Molecular mass of the compound}} \times 100$$

MOLE CONCEPT

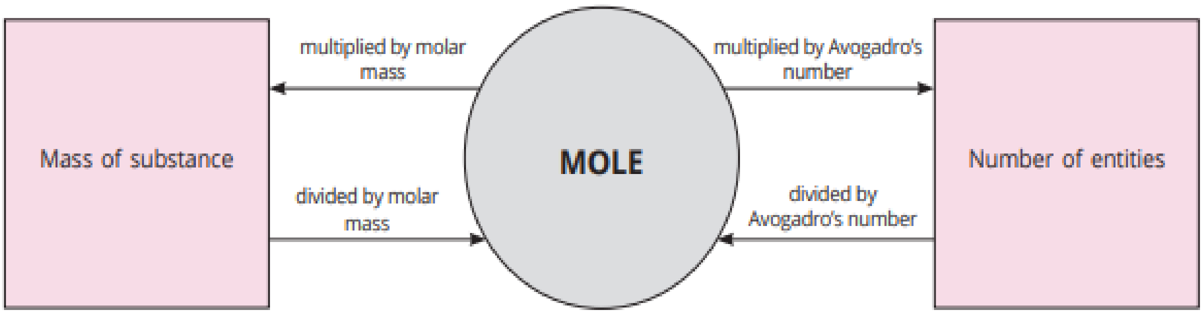
A mole (Latin: *moles* = pile or heap) of any substance (atoms, molecules or ions) represents 6.022×10^{23} particles of that substance. This number is called the Avogadro's constant or Avogadro's number (N_0).

The term mole is applicable to a wide variety of items such as atoms, molecules, ions, electrons, protons, photons or chemical bonds. The SI symbol of mole is mol.

A mole of electrons: The Avogadro's number of electrons is called one mole of electrons. The charge on one electron = -1.6022×10^{-19} C. The charge on Avogadro's number of electrons (i.e., one mole of electrons) = $6.022 \times 10^{23} \times 1.6022 \times 10^{-19}$C ~96500 C. This quantity of charge is called one faraday (F). The charge possessed by the Avogadro's number of electrons is called one faraday.

One mole of hydrogen atoms
= 6.022×10^{23} atoms of hydrogen

One mole of hydrogen molecules
= 6.022×10^{23} molecules of hydrogen



Mole concept illustrated schematically

Gram atomic mass of an element

The quantity of an element equal to the relative atomic mass in gram is termed as the gram-atomic mass (or simply gram-atom) of the element. The mass of 1 mole of atoms of an element is equal to its relative atomic mass in grams. Therefore, we can say that gram-atomic mass or gram-atom is the mass of one Avogadro's number of atoms expressed in grams, i.e.

1 gram-atomic mass = 1 gram-atom= Mass of 6.022×10^{23} atoms of the substance
One gram-atom of all the elements contains 6.022×10^{23} atoms.

Gram-molecular mass of a compound

The quantity of a compound equal to its relative molecular mass in gram is termed as the gram-molecular mass or gram-molecule or simply mole of the compound. The mass of 1 mole of molecules of a compound is equal to its relative molecular mass in grams. Therefore, we can say that gram-molecular mass or gram-molecule is the mass of one Avogadro's number of molecules expressed in grams, i.e.

1 gram-molecular mass = 1 gram-molecule= Mass of 6.022×10^{23} molecules of the substance
One gram-molecule of all compounds contains 6.022×10^{23} molecules.

Molar mass

Mass of one mole of any substance is called its molar mass. Alternatively, the average mass of one mole of any substance is called its molar mass. It is given by,

$$\text{Molar mass } (M) = \frac{\text{Mass of the substance in grams}}{\text{Amount of the substance in moles}} = \frac{m}{n}$$

The unit of molar mass is g mol^{-1}.

How many moles are there in a certain mass of a substance?

The number of moles in a certain mass of a substance is calculated using the following formula:

$$\text{No. of moles of a substance, X}$$
$$= \frac{\text{Mass of the substance in grams}}{\text{Molar mass of the substance in grams per mole}}$$
$$= \frac{m \text{ g}}{M \text{ g mol}^{-1}} = \frac{m}{M} \text{ mol}$$

where *m* is the mass of the substance and *M* is the molar mass of the substance.

How many molecules are there in a certain mass of a substance?

$$\text{No. of molecules of a substance}$$
$$= \frac{\text{Mass of the substance}}{\text{Molar mass of the substance}} \times \begin{array}{c}\text{Avogadro's}\\ \text{number}\end{array}$$
$$= \frac{m}{M} \times 6.022 \times 10^{23} \text{ molecules}$$

where *m* is the mass of the substance and *M* is the molar mass of the substance.

SUMMARY

1. **Law of conservation of mass:** The law of conservation of mass states that during a chemical reaction matter is neither created nor destroyed.
2. **Law of constant proportions:** The law of constant proportions states that whatever be the method of its formation, a chemical compound always consists of the same elements combined together in the same proportion by mass.
3. **Law of multiple proportions:** When two elements combine to form two or more compounds, the different masses of one element which combine with a fixed mass of the other, bear a simple ratio to one another.
4. **Dalton's atomic theory:** According to the Dalton's atomic theory, matter is composed of very tiny indivisible particles called atoms. Atoms are incapable of being destroyed or created.
5. **Atom:** An atom is the smallest particle of an element, which takes part in a chemical reaction and maintains its chemical identity throughout all physical and chemical changes.
6. **Atomic symbols:** Modern atomic symbols were introduced by J J Berzelius. Symbol of an element is the abbreviation of the full name of that element. Symbol of an element is the first letter of the name of the element (Latin/Greek/modern name). For elements having the same first letter, the symbol consists of two letters – first letter in capital followed by second letter not in common, in lower case.
7. **Atomic mass unit:** The mass equal to $\frac{1}{12}$th of the mass of a ^{12}C atom is called one **atomic mass unit**. The mass of $\frac{1}{12}$th of a ^{12}C atom (1.66×10^{-24} g) is taken as one atomic mass unit (u).
8. **Relative atomic mass (A_r):** It is the average atomic mass of an atom of an element as compared to 1/12th the mass of one carbon-12 atom.
9. **Gram-atomic mass:** The quantity of an element equal to the relative atomic mass in gram is termed as the **gram-atomic mass** of the element.
10. **Molecule:** A molecule is the smallest particle of a substance which can exist in a free state independently and exhibits all the properties of the substance.

11. **Chemical formula:** Chemical formula is the symbolic representation of a molecule of the compound.
12. **Molecular formula:** The symbolic representation of a molecule of a substance representing the actual number of various atoms present is called **molecular formula**.
13. **Empirical formula:** The simplest formula which gives the simplest ratio in whole numbers between the number of atoms of different elements present in one molecule of the compound is called **empirical formula**.

    Molecular formula of a substance = $n \times$ Empirical formula
14. **Valency:** Valency is the combining capacity of an element.
15. **Relative molecular mass (M_r):** It is the average mass of the molecules of a compound as compared to 1/12th the mass of one carbon-12 atom.
16. **Molecular mass (M):** The average mass of one molecule of a compound in atomic mass unit is called its **molecular mass**.
17. **Gram-molecular mass:** The quantity of an element or compound equal to the relative molecular mass in gram is termed as the **gram-molecular mass** of the element or compound.
18. **Mole:** The quantity of a substance which contains Avogadro's number of chemical units (atoms, molecules or ions) of the substance is called **mole**. The SI symbol of mole is mol.
19. **Molar mass:** The mass of one mole of any substance is called its **molar mass**.

MIND MAP

ATOMS AND MOLECULES

- **LAW OF CONSERVATION OF MASS**
  - mass can neither be created nor destroyed in a chemical reaction
- **LAW OF CONSTANT PROPORTIONS**
  - a chemical compound always consists of same elements combined in the same proportion by mass

ATOMS

- the building blocks of matter
- atomic radius can be measured in (nm)
- 1 amu = 1 u = mass of one ^{12}C atom/12

MOLECULES

- atoms of same or different kinds
- two or more atoms chemically bonded together

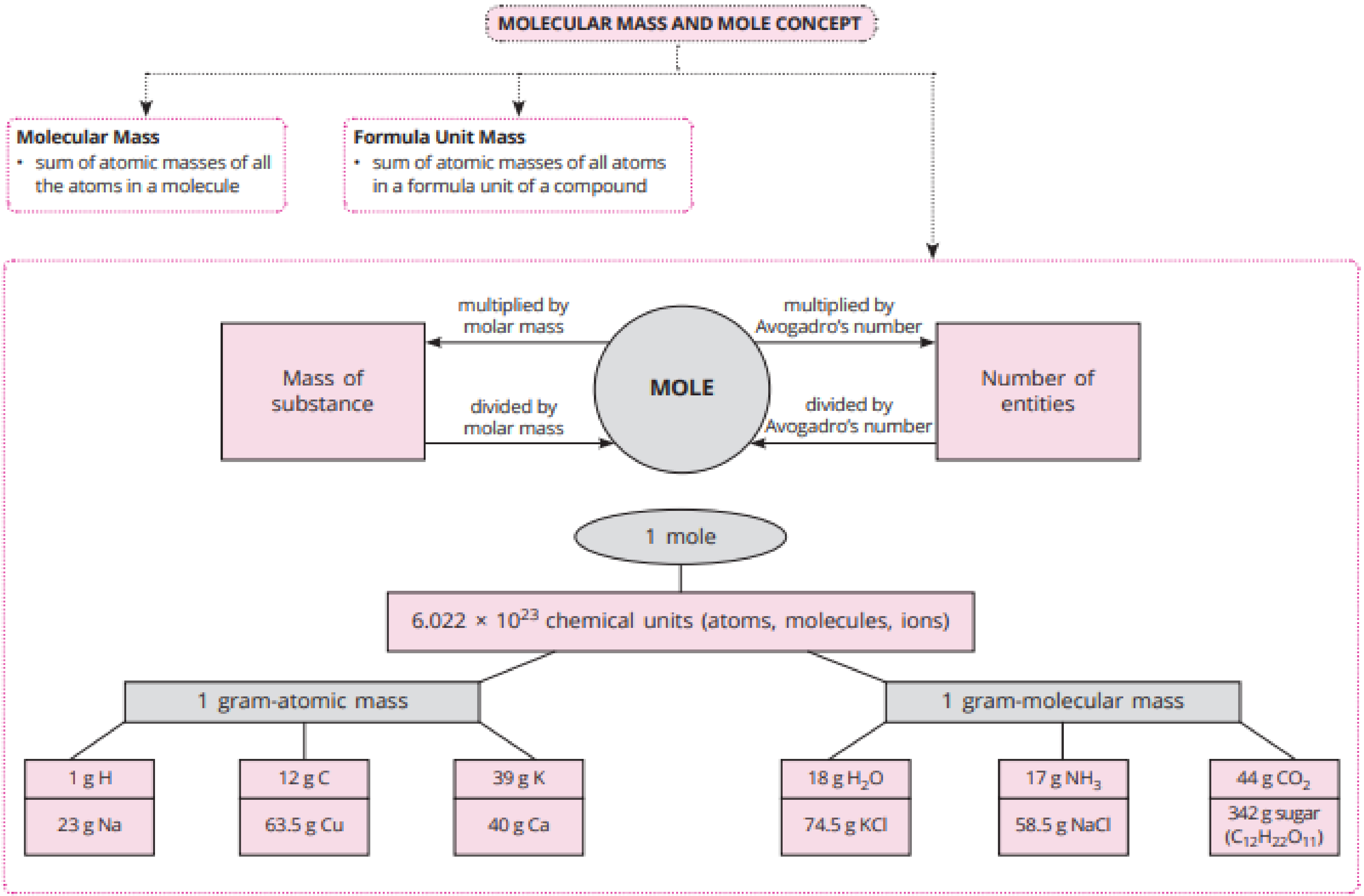
Molecules of Elements

- **Monoatomic:** He, Ne
- **Diatomic:** O_2, H_2
- **Triatomic:** O_3
- **Tetratomic:** P_4
- **Polyatomic:** S_8

Molecules of Compounds

- different elements join together in definite proportions

MIND MAP


MOLECULAR MASS AND MOLE CONCEPT
Molecular Mass
• sum of atomic masses of all the atoms in a molecule
Formula Unit Mass
• sum of atomic masses of all atoms in a formula unit of a compound
Mass of substance
multiplied by molar mass
divided by molar mass
MOLE
multiplied by Avogadro's number
divided by Avogadro's number
Number of entities
1 mole
6.022×10^{23} chemical units (atoms, molecules, ions)
1 gram-atomic mass
1 g H
23 g Na
12 g C
63.5 g Cu
39 g K
40 g Ca
1 gram-molecular mass
18 g H_2O
74.5 g KCl
17 g NH_3
58.5 g NaCl
44 g CO_2
342 g sugar ($C_{12}H_{22}O_{11}$)