

Branches of chemistry

- Organic chemistry : This branch deals with study of compounds of carbon and its compounds with any number of other elements, including hydrogen (most compounds contain at least one carbon–hydrogen bond), nitrogen, oxygen, halogens, phosphorus, silicon, and sulfur. Except carbonates, bicarbonates, cyanides, isocyanides, carbides and oxides
- Inorganic chemistry: This branch deals with study of known elements and their compounds except organic compounds. It is concerned with materials obtained from minerals, air, sea and soil.
- Physical chemistry: This branch deals with study of physical properties and constitution of matter, the laws of chemical combination and theories governing reactions. The effect of temperature, pressure, light, concentration on reaction.
- Analytical chemistry: This branch deals with various methods of analysis of chemical substances both qualitative and quantitative.
- Industrial chemistry : Chemistry involved in industrial process is studied in this branch.
- Biochemistry : It comprise the studies of substances for the prevention and cure of various diseases in living beings
- Nuclear chemistry : This branch deals with study of nuclear reaction, the production of radioactive isotopes and their application in various field.

Measurement in chemistry

A physical quantity is expressed in terms of pure number and unit
Physical quantity = (a pure number) × unit

For example when we say 5 kg. It means 5 times of 1kg

'A unit is defined as the standard of reference chosen in order to measure a definite physical quantity'.

Standard weight and measure are given by International system of Unit (SI)

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Basic physical quantity	Symbol of quantity	Name of SI unit	Symbol of SI unit
Length	l	metre	m
Mass	m	kilogram	kg
Time	t	second	s
Electric current	I	ampere	A
Thermodynamic temperature	T	kelvin	K
Amount of substance	n	mole	mol
Luminous Intensity	I _v	candela	Cd

Some time Sub-multiple and Multiples are used to reduce and enlarge the size of different units

Sub - multiples			Multiples		
Prefix	Symbol	Sub-multiple	Prefix	Symbol	Multiple
deci	d	10 ⁻¹	deca	da	10
centi	c	10 ⁻²	hecto	h	10 ²
milli	m	10 ⁻³	kilo	k	10 ³
micro	μ	10 ⁻⁶	mega	M	10 ⁶
namo	n	10 ⁻⁹	giga	G	10 ⁹
pico	p	10 ⁻¹²	tera	T	10 ¹²
femto	f	10 ⁻¹⁵	peta	P	10 ¹⁵
atto	a	10 ⁻¹⁸	exa	E	10 ¹⁸

Following convention is followed in writing a unit or its symbol

1. The symbols of the units are never expressed in plural form. For example, we write 15 cm and not 15 cms.
2. If a unit is named after the name of a person, it is not written with capital letter. For example, we write newton and not Newton
3. The symbol of the unit named after the name of a person is expressed in capital letter. For example, the symbol for unit newton is N and not n. Symbols of the other units are not written in capital letters. For example, symbol for unit meter is m and not M
4. Not more than one solidus is used. For example 1 poise is expressed as

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1g/s cm or 1 g s⁻¹ cm⁻¹ and not 1g/s/cm.

Some derived units

- a) Area = length × breadth
= m × m = m²
- b) Volume = length × breadth × height
= m × m × m = m³
- c) Density = mass/ volume
= kg/ m³ = kg m⁻³
- d) Speed = distance / time
= m/s = m s⁻¹
- e) Acceleration = change in velocity / time
= m s⁻¹ /s = m s⁻²
- f) Force = mass × acceleration
= kg m s⁻²
- g) Pressure = force per unit area
= kg m s⁻² / m²
= kg m⁻¹ s⁻² or N m⁻² (pascal – Pa)
- h) Energy = force × distance travelled
= kg m s⁻² × m
= kg m² s⁻² (joule – J)

Some conversions

- a) Temperature
Degree Centigrade = Kelvin – 273.15

$$^{\circ}\text{C} = \text{K} - 273.15$$

$$^{\circ}\text{F} = \frac{9}{5} \times (^{\circ}\text{C}) + 32$$

- b) Volume

$$1 \text{ litre (lit or L)} = 1000 \text{ cm}^3$$

$$1 \text{ L} = (10 \text{ cm})^3$$

$$1 \text{ L} = (10 \times 10^{-2} \text{ m})^3$$

$$1 \text{ L} = 10^3 \times 10^{-6} \text{ m}^3$$

$$1 \text{ L} = 10^{-3} \text{ m}^3$$

$$\text{OR } 1 \text{ m}^3 = 1000 \text{ L}$$

Also

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$$1 \text{ L} = 1000 \text{ mL} = 10^{-1} \text{ m}^3 = 1 \text{ dm}^3$$

c) Energy

$$1 \text{ cal} = 4.184 \text{ J}$$

$$1 \text{ eV} = 1.6 \times 10^{-19} \text{ J}$$

d) Pressure

$$1 \text{ atm} = 760 \text{ torr} = 760 \text{ mmHg} = 76 \text{ cmHg} = 1.013 \times 10^5 \text{ Pa}$$

Significant figures

To indicate the precision of measurement, scientists use the term significant figure.

The following rules are observed in counting the number of significant figures in a given measured quantity.

1) All non-zero digits are significant : for example

24.3 has three significant figure

243.2 has four significant figure

2) A zero becomes significant figure if it appears between two non-zero digits: for example

3.04 has three significant figure

3.506 has four significant number

3) Leading zeros or the zero placed to the left of the numbers are never significant : For example

0.542 has three significant number

0.054 has two significant number

0.006 has one significant number

4) Trailing zero or the zeros placed to the right of the decimal are significant: For example

431.0 has four significant figures

432.00 has five significant number

5) If a number ends in zero but these zeros are not to the right of decimal point, these zeros may or may not be significant

For example , 11400 g may have three, four or five significant figures. Such a number is first written in exponential form. The above mass may be written in the following exponential forms

$1.14 \times 10^4 \text{ g}$ which has three significant figures

$1.140 \times 10^4 \text{ g}$ has four significant figures

$1.1400 \times 10^4 \text{ g}$ has five significant figures

6) The non-significant figures in the measurements are round off

a) If the figure following the last number to be retained is less than 5, all the unwanted figures are discarded and the last number is left unchanged e.g.

5,6724 is 5.67 to three significant figure

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b) If the figure following the last number to be retained is greater than 5, the last figure to be retained is increased by 1 unit and the unwanted figures are discarded e.g 8.6526 is 8.653 to four significant figures

c) If the figure following the last number to be retained is 5, the last figure is increased by 1 only in case it happened to be odd. In case of even number the last figure remains unchanged. For example 2.3524 is 2.4 to two significant figures (3 is odd hence increased by 1)

7.4511 is 7.4 to two significant figure (4 is even hence no increase)

7) Exact number have an infinite number of significant figures because infinite number of zero can be placed after decimal point
Example $20 = 20.00000 \dots$

Calculations Involving Significant Figures

Rule 1 : The resultant of an addition or subtraction in the numbers having different precisions should be reported to the same number of decimal places as are present in the number having the least number of decimal places. The rule is illustrated by the following examples

$$\begin{array}{r} \text{a) } 33.3 \quad (\text{ only one decimal place}) \\ 03.11 \\ +00.313 \\ \hline \end{array}$$

Sum 36.723 (Answer should be reported to one decimal place)

Correct answer 36.7

$$\begin{array}{r} \text{b) } 3.1421 \\ 0.241 \\ \hline 0.09 \quad (\text{ two decimal places}) \end{array}$$

Sum 3.4721 (answer should be reported to 2 decimal place)

Correct answer = 3.47

$$\begin{array}{r} \text{c) } 62.831 \\ - 24.5495 \end{array}$$

Difference = 38.2815 (Answer should be reported to 3 decimal place after round off)

Correct answer = 38.282 (as 1 is odd number hence increased by one)

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Rule 2: The answer to a multiplication or division is rounded off to the same number of significant figures as is possessed by the least precise term used in the calculation. Example

a) 142.06

$\times 0.23$ (two significant figures)

32.6738 (answer should have two significant figures, after rounding off)

Correct answer = 33

b) 51.028

$\times 1.31$ (three significant figures)

66.84668

Correct answer = 66.8

c) $\frac{0.90}{4.26} = 0.2112676$

Correct answer = 0.21

d) $\frac{3.24 \times 0.0866}{5.046} = 0.055653$

Correct answer = 0.0556

e) $\frac{4.28 \times 0.146 \times 3}{0.0418} = 44.84784$

Correct answer = 44.8

Elements and compounds

Elements are pure substances that cannot be decomposed into simpler substances by chemical change. The smallest particles of an element possess the same properties as the bigger particles. An element can also be defined as a pure substance which consists of only one type of atoms (smallest particle)

Elements are classified into Metals, non-metals and metalloids

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Compounds are also pure substances that are composed of two or more different elements in a fixed proportion by mass. The property of compound is altogether different than the constituent elements. For example water is made up of hydrogen which can burn and oxygen which helps to burn but Water extinguishes fire. Components of the compound cannot be separated by physical process but only by chemical process. If we just boil water Hydrogen and oxygen cannot be separated for that we have to go for electrolysis process

Compound are classified into two types:

- i) Organic compounds : The compounds obtained from living sources are termed organic compounds. The term organic is now applied to hydrocarbons (compounds of carbon and hydrogen) and their derivatives
- ii) Inorganic compounds: The compounds obtained from non-living sources such as rock and minerals are termed inorganic compounds. The compounds of all elements except hydrocarbons and their derivatives are included in this category.

Some specific properties of substances:

- i) Deliquescence: The property of certain compounds of taking up the moisture present in atmosphere and becomes wet when exposed is known as deliquescence. For example Sodium hydroxide, potassium hydroxide, calcium chloride, magnesium chloride anhydrous ferric chloride
- ii) Hygroscopicity : Certain compounds combine with the moisture and are converted into hydroxides or hydrates. Such substances are called hygroscopic. Anhydrous copper sulphate, quick lime (CaO), anhydrous sodium carbonate etc are of hygroscopic nature
- iii) Efflorescence: The property of some crystalline substances of losing their water of crystallization on exposure and becoming powdery on the surface is called efflorescence. Such salts are known as efflorescent. The examples are Ferrous sulphate ($\text{FeSO}_4 \cdot 7\text{H}_2\text{O}$), sodium carbonate (

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$\text{Na}_2\text{CO}_3 \cdot 10\text{H}_2\text{O}$), Sodium sulphate ($\text{Na}_2\text{SO}_4 \cdot 10\text{H}_2\text{O}$) potash alum [$\text{K}_2\text{SO}_4 \cdot \text{Al}_2(\text{SO}_4)_3 \cdot 24\text{H}_2\text{O}$] etc

- iv) Malleability : this property show by metals. When solid is beaten and does not break but converted into a thin sheet, it is said to posses the property of malleability. Example Gold , silver, copper
- v) Ductility : The property of metal to be drawn into wires is termed ductility. Example Gold , silver, copper. Platinum is most ductile.
- vi) Elasticity : When the stress (force per unit area normal to cross-section) is small, the solid completely regain its original shape, size or volume after deforming force is removed. The solid is then said to be elastic. Steel, glass, ivory etc are elastic bodies
- vii) Plasticity : When stress is increased on a metal, a limit is reached beyond which, if the stress is removed, the solid does not come back to its original shape or size. It acquires a permanent deformation. Such material can be given any shape without any difficulty
- viii) Brittleness: The solid materials which break into small pieces on hammering are called brittle. The solids of non-metals are generally brittle in nature
- ix) Hardness: A material is said to be harder than the other if it can scratch it. The hardness is measured on Mho's scale. For this purpose, ten minerals have been selected which have been assigned hardness from 1 to 10
On Mho's scale, hardness of diamond is maximum and that of talc is minimum.

Law of Chemical Combination

- (i) Law of conservation of mass : The law was first stated by Lavoisier in 1774 .According to this law, in all chemical changes, the total mass of the system remains constant or in chemical change, mass is neither created nor destroyed. Total mass of reactant = Total mass of products + masses of the un-reacted reactants

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- (ii) Law of definite or constant proportions: This law was presented by Proust in 1799 . According to this law " A chemical compound always contains the same element combined together in fixed proportion by mass. i.e. a chemical compound has fixed composition and it does not depends on the method of its preparation or the source from which it has been obtained.

For example carbon dioxide can be obtained by any of the following methods

- by heating calcium carbonates
- By burning carbon in oxygen
- by reacting calcium carbonate with hydrochloric acid

Whatever sample of carbon dioxide is taken it is observed that carbon and oxygen are always combined in the ratio of 12:32 or 3:8

The converse of this law that when same elements combines in the same proportion, the same compound will be formed, is always not true.

For example. Carbon, hydrogen and oxygen when combined in the ratio of 12:3:8 may form either ethyl alcohol (C_2H_5OH) or dimethyl ether (CH_3OCH_3) under different experimental condition

- (iii) Law of multiple proportions : this law was put forward by Dalton in 1808 . According to this law, if two elements combine to form more than one compound, then the different masses of one element which combines with a fixed mass of the other element, bear a simple ratio to one another.

Nitrogen forms five stable oxides

N_2O	Nitrogen 28 parts	Oxygen 16 parts
N_2O_2	Nitrogen 28 parts	Oxygen 32 parts
N_2O_3	Nitrogen 28 parts	Oxygen 48 parts
N_2O_4	Nitrogen 28 parts	Oxygen 64 parts
N_2O_5	Nitrogen 28 parts	Oxygen 80 parts

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Thus for same mass of nitrogen combines with different masses of oxygen and the ratio of oxygen with same mass of Nitrogen is in ratio

16:32:48:64:80 or 1:2:3:4:5

(iv) Law of reciprocal proportion : This law was given by Richter in 1794. The law states that when definite mass of an element A combines with two other elements B and C to form two compounds and if B and C also combine to form a compound, their combining masses are in same proportion or bear a simple ratio to the masses of B and C which combine with a constant mass of A.

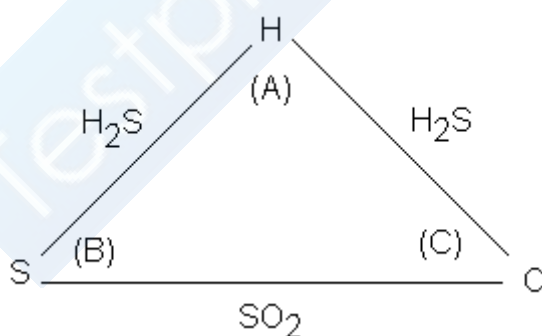
(v) For example , hydrogen combines with sodium and chlorine to form compounds NaH and HCl respectively

In NaH Sodium 23 parts Hydrogen one part

In HCl Chlorine 35.5 parts Hydrogen one part

Thus according to law when Sodium combines with Chlorine to form NaCl is 23 parts of sodium and 35.5 parts of chlorine

However in following example it do not hold good



In H₂S Hydrogen two parts Sulphur 32 parts

In H₂O Hydrogen two parts Oxygen 16 parts

Thus expected that 32 part of Sulphur should combine with 16 parts of oxygen but actually both combine to form SO₂ in the ratio of 32:32 or 1:1

The law of reciprocal proportions is a special case of a more general law, the law of equivalent masses, which can be stated as " In all chemical reactions, substances always react in the ratio of their equivalent masses"

DALTON'S ATOMIC THEORY

Main postulates of this atomic theory are

- Matter is discrete (i.e. discontinuous) and is made up of atoms. An atom is the smallest (chemically) indivisible part of an element, which can take part in a chemical change
- Atoms of the same element are identical in all aspect, size, shape, structure etc and especially weight.
- Atoms of different elements have different properties and weights
- Atoms cannot be created or destroyed. So a chemical reaction is nothing but a rearrangement of atoms and the same number of atoms must be present before and after reaction
- A compound is formed by the union of atoms of one element with atoms of another in a fixed ratio of small whole numbers (1:, 1:2, 1:3 etc)

All the postulates of Dalton's atomic theory have been proved to be incorrect.

Atom is divisible in the sense that it has structure. Sub-atomic particles are known.

The existence of isotopes for more elements shows that atoms of the same element need not have same mass. The atomic weight of an element is , in fact, a mean of the atomic masses of the different isotopes of the element.

Part of atomic mass can be destroyed and an equivalent amount of energy is released during nuclear fission.

Atoms combine in fixed integral ratios; however there are instances where atoms combine in non-integral ratios. E.g. iron(II) with a composition $\text{Fe}_{(1-x)}\text{S}$ ($x = 0$ to 0.2).

In zinc oxide, zinc and oxygen have not combined in exactly an integral ratio. The atomic ratio of Zn : = $(1+x) : 1$, where x is a very small fraction. Compounds of this kind are called non-stoichiometric compounds or Berthollide compounds as against

compounds whose formulae are accordance with atomic theory and law of definite proportion.

Atomic Weight or Atomic mass

The atomic mass of an element can be defined as the number which indicates how many times the mass of one atom of the element is heavier in comparison to the mass of one atom of hydrogen. If Hydrogen atom selected as standard

Thus Atomic weight of element

$$= \frac{\text{Weight of 1 atom of the element}}{\text{Weight of 1 atom of hydrogen}}$$

When we state atomic mass of chlorine is 35.5, we mean that an atom of chlorine is 35.5 times heavier than an atom of hydrogen.

It was later felt that the standard for reference for atomic weight may be oxygen, the advantage being that the atomic weights of most other elements become close to whole number

Thus Atomic weight of element

$$= \frac{\text{Weight of 1 atom of the element}}{\frac{1}{16} \times \text{Weight of 1 atom of oxygen}}$$

The modern reference standard for atomic weight is carbon isotope of mass number 12

Thus Atomic weight of element

$$= \frac{\text{Weight of 1 atom of the element}}{\frac{1}{12} \times \text{Weight of 1 atom of carbon} - 12}$$

On this basis, atomic weight of oxygen 16 was changed to 15.9994

Nowadays atomic weight is called relative atomic mass and denoted by amu or u (atomic mass unit). The standard for atomic mass is C^{12}

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Atomic mass unit : The quantity of one twelfth mass of an atom of carbon-12 (C^{12}) is known as the atomic mass unit (amu). The actual mass of one atom of carbon -12 is $1.9924 \times 10^{-23}g$ or $1.9924 \times 10^{-26}kg$

Atomic mass of Hydrogen is 1.008 amu. In terms of gram it will be
 $= 1.008 \times 1.9924 \times 10^{-23}g = 1.6736 \times 10^{-24} g$

Atomic mass of oxygen is 16.00 amu. In terms of gram it will be
 $= 16 \times 1.9924 \times 10^{-23}g = 2.656 \times 10^{-23} g$

It is clear from above that

- Atomic weight or mass is not a weight but a number
- Atomic weight is not absolute but relative to the weight of the standard reference element C^{12}
- Gram atomic weight is atomic weight expressed in gram, but it has a special significance with reference to a mole

Most of the elements occur in nature as a mixture of isotopes. Thus average atomic masses of various elements are determined by multiplying the atomic mass of each isotope by its fractional abundance and adding the values thus obtained. The fractional abundance is determined by dividing percentage abundance by hundred

If a, b are the atomic masses of isotopes in the ratio m:n then average isotopic mass

$$= \frac{m \times a + n \times b}{m + n}$$

If x and y are percentage abundance of the two isotopes ($y = 100 - x$) then

$$= \frac{x}{100} \times a + \frac{y}{100} \times b$$

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Illustration

Q) Given that the abundance of isotopes ^{54}Fe , ^{56}Fe , ^{57}Fe are 5%, 90% and 5% respectively. The atomic mass of Fe is

Solution

From the formula

$$= \frac{x}{100} \times a + \frac{y}{100} \times b + \frac{z}{100} \times c$$

X = 5, y = 90, z = 5 and a = 54, b = 56, c = 57

On substitution of above data in Formula we get

$$= \frac{5}{100} \times 54 + \frac{90}{100} \times 56 + \frac{5}{100} \times 57 = 55.95 \text{ amu}$$

Q) Carbon occurs in nature as a mixture of carbon-12 and carbon-13. The average atomic mass of carbon is 12.011amu. the percentage abundance of Carbon-13 is

Solution

Let x be the percentage abundance of carbon-12 then (100-x) will be the percentage abundance of carbon-13. From formula

$$12.011 = \frac{x}{100} \times 12 + \frac{(100 - x)}{100} \times 13$$

On simplification we get x = 98.9 or abundance of C-12 is 98.9% and C-13 is 1.1%

Avogadro's number

It is the number of atoms in exactly 12 grams of carbon -12. This experimentally determined value is approximately 6.0221×10^{23} particles.

Gram-atomic mass or gram atom

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Gram-atomic mass can be defined as the absolute mass in gram of 6.02×10^{23} atoms of any element.

$$\text{Mass of Oxygen} = 16 \text{ u} = 16 \times 1.66 \times 10^{-24} \text{ g}$$

$$\text{Thus Absolute mass of } 6.02 \times 10^{23} \text{ atoms of oxygen} = 16 \times 1.66 \times 10^{-24} \times 6.02 \times 10^{23} \text{ g} = 16 \text{ g}$$

Thus mass of 1 atom of Oxygen is 16 amu or u and Gram-atomic mass of Oxygen is 16g

The Molecule

Avogadro suggested that the fundamental; chemical unit is not an atom but a molecule, which may be a cluster of atoms held together in some manner causing them to exist as unit. The term molecule means the smallest particle of an element or a compound that can exist free and retain all its properties.

Consider a molecule of carbon dioxide. It has been established that it contains one atom of carbon and two atoms of oxygen. This molecule can be split into atoms of carbon and oxygen. So the smallest particle of carbon dioxide that can exist free and retain all its properties is the molecule of carbon dioxide. A compound molecule should contain at least two different atoms.

The term molecule is also applied to describe the smallest particle of an element which can exist free. Thus a hydrogen molecule is proved to contain two atoms; when it is split into atoms there will be observed a change in property.

Molecules of elementary gases like hydrogen, oxygen, nitrogen, chlorine etc contain two atoms in molecules; they are diatomic. Molecules of noble gas like helium, neon argon, krypton and xenon are monoatomic. Molecule of Sulphur contains eight atoms while molecule of phosphorus contains four atoms.

The number of atoms of an element in a molecule of the element is called its atomicity

Molecular weight

It is the number of times a molecule is heavier than one twelve of an atom of carbon-12

Molecular weight

$$= \frac{\text{weight of 1 molecule}}{\frac{1}{12} \times \text{weight of 1 carbon - 12 atom}}$$

- Molecular weight is not a weight but a number
- Molecular weight is relative and not absolute.
- Molecular weight expressed in grams is called gram-molecular weight
- Molecular weight is calculated by adding all the atomic weight of all the atoms in a molecule

The molecular weight of carbon dioxide CO_2 is calculated as Atomic mass of carbon 12 u and atomic mass of oxygen is 16 u
Thus molecular mass of $\text{CO}_2 = 1 \times 12 + 2 \times 16 = 44 \text{ u}$

- Molecular weight is now called relative molecular mass

Gram-molecular Mass or Gram molecule

A quantity of substance whose mass in gram is numerically equal to its molecular mass is called gram-molecular mass. In other words, molecular mass of a substance expressed in gram is called gram-molecular mass or gram molecule.

For example the molecular mass of chlorine is 71u and therefore its gram-molecular mass or gram molecule is 71g which is the mass of 6.02×10^{23} molecules of chlorine

Gram molecular mass of oxygen is 32 and Nitric acid $\text{HNO}_3 = 1 + 14 + 3 \times 16 = 63 \text{ g}$

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

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

Avogadro's Hypothesis

It states that equal volumes of gases at the same temperature and pressure contain equal number of molecules. It means that 1ml of hydrogen, oxygen, carbon dioxide or mixture of gases taken at the same temperature and pressure contains the same number of molecules

Application of Avogadro's hypothesis

- a) To prove that simple elementary gas molecules are diatomic

Consider the experimental result

1 volume of hydrogen + 1 volume of chlorine → 2 volume of hydrogen chloride at the same temperature

1 volume of contains 'n' molecules.

Then n molecules of hydrogen + n molecules of chlorine → 2n molecules of hydrogen chloride

A molecule of hydrogen chloride should contain at least 1 atom of hydrogen and 1 atom of chlorine. Two molecules of hydrogen chloride should contain at least 2 atoms of hydrogen and 2 atoms of chlorine and these should have come from 1 molecule of hydrogen and 1 molecule of chlorine respectively. Thus Avogadro's hypothesis enables us to establish that hydrogen and chlorine molecules must contain at least 2 atoms.

- b) To establish the relationship between molecular weight and vapour density of a gas

The absolute density of gas is the weight of 1 L of gas at STP (Standard temperature 0°C and pressure 1 atm)

The relative density or vapour density of gas (V.D.)

$$= \frac{\text{Density of the gas}}{\text{Density of hydrogen}}$$

hydrogen

$$V.D. = \frac{\text{Weight of 1 L of gas at STP}}{\text{Weight of 1 L of Hydrogen gas at STP}}$$

$$V.D. = \frac{\text{Weight of a certain volume of the gas}}{\text{Weight of the same volume of hydrogen at the same temperature and pressure}}$$

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So Vapour density of gas is defined as the ratio of the weight of a certain volume of the gas to the weight of the same volume of hydrogen at the same temperature and pressure.

Let V litres of the gas contains 'n' molecules

$$V.D. = \frac{\text{Weight of 'n' molecules of the gas}}{\text{weight of 'n' molecules of hydrogen}}$$

$$V.D. = \frac{1}{2} \times \frac{\text{Weight of 1 molecule of gas}}{\text{Weight of 1 atom of hydrogen}}$$

$$V.D. = \frac{1}{2} \times \text{Molecular weight of the gas}$$

Molecular weight of gas = 2 × Vapour density of the gas

c) Gram-Molecular volume or Molar Volume

Molecular weight of gas = 2 × Vapour density of the gas

$$\text{Molecular weight of gas} = 2 \times \frac{\text{Weight of 1L of gas at STP}}{\text{Weight of 1L of hydrogen at STP}}$$

$$\text{Gram - Molecular weight of gas} = 2 \times \frac{\text{Weight of 1L of gas at STP}}{0.089\text{g}}$$

Gram-Molecular weight of gas = 22.4 × Weight of 1L of the gas at STP

Gram-Molecular weight of gas = Weight of 22.4L of gas at STP

This establish that gram -molecular weight of any gas (or vapour) occupies the same volume of 22.4L at STP. The volume occupied by a gram-molecular weight of any gas is called a molar volume and its value is 22.4L at STP

MOLE

The term 'mole' was introduced by Ostwald in 1896. This is the Latin word 'moles' means heap or pile.

A mole(mol) is defined as the number of atoms in 12.0 g of carbon-12 which is equal to 6.02×10^{23} . This number is also known as Avogadro's number

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Thus mole is a unit which counts, similar to dozen or gross.

Thus heap of 6.02×10^{23} particles, atoms, ions, molecules or object is called one mole

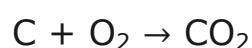
Now if mole is heap then it must have weight and occupy volume

Mass of One mole of substance is equal to gram-molecular mass for example

One mole of water (H_2O) contains 6.02×10^{23} particles (molecules of water) and weight of this molecules = 18 g (2+16)

One mole of Chlorine (Cl_2) contains 6.02×10^{23} particles (molecules of chlorine) and weight of this molecules = 71 g (2×35.5)

So one mole of substance will have mass equal to formula mass of that substance.



In above reaction one mole of carbon (C) reacted with one mole of Oxygen gas (O_2) to produce one mole of carbon dioxide (CO_2) or

CO_2 is made up of one mole of carbon and one mole of Oxygen gas

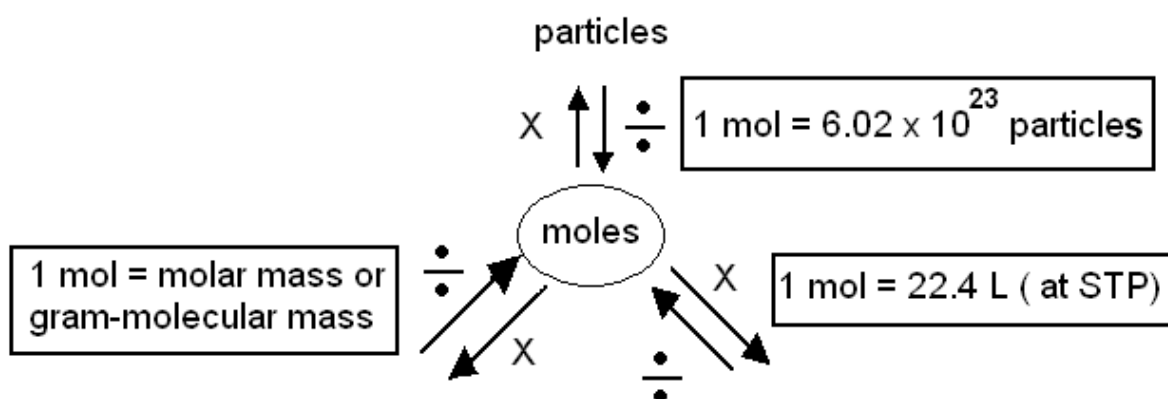
It is established by Avogadro's hypothesis that one gram-molecular mass which is mass of one mole of gaseous substance occupy volume of 22.4 L at NTP.

1 mole = 22.4 L at STP

1mole = 6.02×10^{23} number of particles, atom, molecules, ions

1mole = gram-molecular weight

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Illustrations

Q) What is the mass of one molecule of water

Solution

Molecular weight of water = 18 g

One mole contains 6.02×10^{23} number of molecules of water

Thus Wt of one molecule of water = $18 / 6.02 \times 10^{23} = 3 \times 10^{-23} \text{ g}$

Q) At STP, 5.6 L of gas weight 60g. The vapour density of gas will be

Solution

Molecular mass = $2 \times$ Vapour density --- eq(1)

Calculate molecular mass

5.6 L = 60 g

22.4 L = ?

Thus 22.4 L have mass = molecular mass = $\frac{60 \times 22.4}{5.6} = 240 \text{ g/mol}$

Now from equation(1)

$240 = 2 \times$ Vapour density

\therefore Vapour density = 120

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Q) x gram of CaCO_3 was completely burnt in air. The mass of the solid residue formed is 28. What is the value of 'x' in gram. [Molecular weight of Ca=40, O =16, C =12]



Solution

From given data molecular weight of $\text{CaCO}_3 = 40 + 12 + 48 = 100$

$\text{CaO} = 40 + 16 = 56\text{g}$

From given chemical equation 100 g of CaCO_3 produces 56 g of CaO

Or $56\text{g} = 100\text{g} \therefore 28 = x?$ On simplification we get 50g

Q) If an organic compound contain 4% sulphur. Then, what will the minimum molecular mass of the organic compound.

Solution

Here we will assume organic compound contain one atom of sulphur, so that we can estimate minimum molecular weight

Now molecular weight of sulphur = 32 g/mol

From given

5 g of S = 100 g of compound

32 = ? of compound

On simplification we get 740 g/mole

Or 740 amu is the weight of one molecule of Compound

Q) How many atoms are present in one mole of water

Solution

Molecular formula of water is H_2O

One molecule of water contains 3 atom

Thus One mole of water contain 3 mole of atom

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Now 1 mole of atom = 6.02×10^{23} atoms

\therefore 3 mole atom = $3 \times 6.02 \times 10^{23}$ atoms = 1.806×10^{24} atoms

Q) Which of the following have maximum number of molecules present

a) 15 L of H_2 gas at STP b) 5 L of N_2 gas at STP

c) 0.5g of H_2 gas d) 10g of O_2 gas

Solution

22.4 L of gas at STP = 1 mol

Thus 15L of H_2 gas at STP = $15/22.4 = 0.669$ mol

5 L of N_2 gas at STP = $5/22.4 = 0.223$ mol

Molecular weight of H_2 gas = 2 g or

2g of H_2 gas = 1 mol

\therefore 0.5g = 0.25 mol

Molecular weight of O_2 gas = 32 g or

32 g = 1 mole

\therefore 10 g of $O_2 = 10/32 = 0.312$ mol

Thus

15 L of H_2 gas at STP = 0.669 mol

5 L of N_2 gas at STP = 0.223 mol

0.5g of H_2 gas = 0.25 mol

10 g of $O_2 = 0.312$ mol

From above 15 L of H_2 gas at STP contain more number of moles thus more number of molecules

Percentage Composition of a Compound

Percentage composition of a compound is the relative mass of the each of the constituent element in 100 parts of it. It is readily calculated from the formula of the compound. Molecular mass of a compound is obtained from its formula by adding up the masses of all the atoms of the constituent elements present in the molecule.

For example : formula of methyl alcohol is CH_3OH

Molecular weight of $\text{CH}_3\text{OH} = 12 + 3(1) + 16 + 1 = 32\text{g}$

Amount of carbon = 12 g there for % of carbon

32 g of $\text{CH}_3\text{OH} = 12\text{ g}$

100 g = ?

On simplification we get % of Carbon (C) = 37.5%

Similarly hydrogen = 12.5%

And Oxygen = 50%

Empirical formula and molecular formula

Empirical formula represents the simplest relative whole number ratio of atoms of each element present in the molecule of the substance. For example CH is the empirical formula of benzene in which ratio of the atoms of carbon and hydrogen is 1:1. It also indicate ratio of carbon and hydrogen is 12:1 by mass.

Molecular formula of a compound is one which expresses as the actual number of atoms of each element present in one molecule. C_6H_6 is the molecular formula of benzene indicating that six carbon atoms and six hydrogen atoms are present in a molecule of benzene. Thus

Molecular formula = $n \times$ Empirical formula

Determination of Empirical and molecular formula.

The following steps are followed to determine the empirical formula of the compound

- i) The percentage composition of the compound is determined by quantitative analysis
- ii) The percentage of each element is divided by its atomic mass. It gives atomic ratio of the elements present in the compound
- iii) The atomic ratio of each element is divided by the minimum value of atomic ratio as to get the simplest ratio of the atoms of element present in the compound.
- iv) If the simplest ratio is fractional, then values of simplest ratio of each element is multiplied by a smallest integer to get a simplest whole number for each of the element
- v) To get the empirical formula, symbols of various elements present are written side by side with their respective whole ratio as a subscript to lower right hand corner of the symbol

Illustration

Q) Caffeine has the following percent composition: carbon 49.48%, hydrogen 5.19%, oxygen 16.48% and nitrogen 28.85%. Its molecular weight is 194.19 g/mol. What is its molecular formula?

Solution

1) Calculate the empirical formula:

carbon: $49.98 \text{ g} \div 12.011 \text{ g/mol} = 4.16$
 hydrogen: $5.19 \text{ g} \div 1.008 \text{ g/mol} = 5.15$
 nitrogen: $28.85 \text{ g} \div 14.007 \text{ g/mol} = 2.06$
 oxygen: $16.48 \text{ g} \div 15.999 \text{ g/mol} = 1.03$

carbon: $4.16 \div 1.03 = 4.04 = 4$
 hydrogen: $5.15 \div 1.03 = 5$

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nitrogen: $2.06 \div 1.03 = 2$

oxygen: $1.03 \div 1.03 = 1$

2) Empirical formula is $C_4H_5N_2O$. The "empirical formula weight" is about 97.1, which gives a scaling factor of two.

3) The molecular formula is $C_8H_{10}N_4O_2$.

Method 2:

Number of carbon atoms in molecule

$$= \frac{\%}{100} \times \frac{\text{Molecular mass}}{\text{Atomic mass}}$$

$$= \frac{49.48}{100} \times \frac{194.19}{12} = 8.00$$

Number of oxygen atoms in molecule

$$= \frac{\%}{100} \times \frac{\text{Molecular mass}}{\text{Atomic mass}}$$

$$= \frac{16.48}{100} \times \frac{194.19}{16} = 2.00$$

Number of nitrogen atoms in molecule

$$= \frac{\%}{100} \times \frac{\text{Molecular mass}}{\text{Atomic mass}}$$

$$= \frac{28.85}{100} \times \frac{194.19}{14} = 4.00$$

Number of hydrogen atoms in molecule

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$$= \frac{\%}{100} \times \frac{\text{Molecular mass}}{\text{Atomic mass}}$$

$$= \frac{5.15}{100} \times \frac{194.19}{1} = 10.00$$

Thus formula is $C_8H_{10}N_4O_2$

Chemical Equation and Stoichiometry

Stoichiometry is the calculation of the quantities of reactant and products involved in a chemical reaction.

Consider a reaction $N_2(g) + 3H_2(g) \rightarrow 2NH_3(g)$

The calculation based on mole-mole relationship

In such calculations, number of moles of reactants are given and those of products required. Conversely if the number of moles of products are given, the number of moles of reactants are required. Above reaction can be interpreted as

1 mole of N_2 + 3 mole of H_2 \rightarrow 2 mol of NH_3

Calculation based on mass- mass relationship

It making necessary calculations, following steps are followed

- Write down the balanced chemical equation
- Write down the theoretical amount of reactants and products involved in the reaction
- Calculate the unknown amount of substance using unitary method

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

Calculation based on volume-volume relationship

In such calculation after balancing reaction quantity of gases are expressed in volume based on STP conditions

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Limiting reagent:

Limiting reactant or reagent is the reactant that is entirely consumed when a reaction goes to completion. Other reactants which are not completely consumed in the reaction are called excess reactant

Calculation involving percent yield

The amount of the product actually obtained is called the actual yield. Theoretical yield is expected quantity of product to achieve.

$$\text{Percent yield} = \frac{\text{Actual yield}}{\text{Theoretical yield}} \times 100$$

CONCENTRATION OF SOLUTION

1. Molarity

$$M = \frac{\text{Number of moles of solute}}{\text{Volume of solution (in litre)}}$$
$$M = \frac{W \times 1000}{M_0 \times V}$$

W : Weight of solute in grams

M₀ : Molecular mass of solute

V: Volume of solution in ml

2. Normality

$$N = \frac{\text{Number of gram equivalents of solute}}{\text{Volume of solution (in litre)}}$$
$$N = \frac{W \times 1000}{Eq \times V}$$

W : Weight of solute in grams

Eq : Equivalent weight of solute

V: Volume of solution in ml

Equivalent weight of solute

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$$Eq = \frac{\text{molecular mass of solute}}{n}$$

n factor : is acidity, basicity, Charge on ion, change in oxidation state

Acidity : Number replaceable -OH group in basic compound

Example: NaOH : number of -OH groups 1, Thus Acidity = 1

Ca(OH)₂ : Number of -OH groups 2, Acidity = 2

Al(OH)₃ : Number of -OH groups 3, Acidity = 3

Basicity of acid is defined as the no of ionizable hydrogen (H⁺) ions present in one molecule of an acid.

Example: HCl, HNO₃, CH₃COOH : H⁺ ions are 1, thus Basicity is 1,

H₂SO₃ : H⁺ ions are 2, thus Basicity is 2,

Oxalic acid : H⁺ ions are 2, thus Basicity is 2

Charge: Na⁺ thus take charge as 1

Ca⁺⁺ thus take charge as 2

SO₄⁻² thus take charge as 2

Change is oxidation state:

Example KMnO₄

Basic medium : Change in oxidation state of Mn= +1

Acidic medium : Change in oxidation state of Mn= +5

Neutral medium : Change in oxidation state of Mn= +3

From equation

$$N = \frac{W \times 1000}{Eq \times V}$$
$$NV = \frac{W \times 1000}{Eq}$$

Here NV is called milliequivalence

The law of equivalence

The law of equivalence provide, us the molar ratio of reactants and products without knowing the complete balanced reaction, which is as good as having a balanced chemical reaction. The molar ratio of reactants and products can be known by knowing the n-factor of relevant species. According to the law of equivalence, whenever two

BASIC PP

substances react, the equivalents of one will be equal to the equivalents of other and the equivalents of any product will also be equal to that of the reactant.

In general, whenever two substance react with their n-factor in the ratio of a:b, then their molar ratio in a balanced chemical reaction would be b:a

Also for dilution of solution and titration

$$N_1V_1 = N_2V_2$$

Illustration

Q) 28 g of metal requires 24.5 g of sulphuric acid to get dissolved. Calculate the equivalent weight of metal and the volume of hydrogen liberated at STP

Solution

Molecular weight of Sulphuric acid = 98 and basicity = 2 (because of two H⁺)

Thus Equivalent weight = $98/2 = 49$

Let E be the equivalent weight of metal

According to law of chemical equivalence,

Number of gram equivalent of metal = Number of gram equivalents of sulphuric acid

$$\frac{28}{E} = \frac{24.5}{49}$$

$$E = 56$$

Now one equivalent of a metal liberate one equivalent volume of hydrogen

Hydrogen is a diatomic gas with valency 1

Equivalent volume of hydrogen = 11.2L at STP

It means 56 g of metal liberate 11.2 L hydrogen

∴ Volume of hydrogen liberated by 28g metal

$$= \frac{11.2 \times 28}{56} = 5.6L$$

BASIC PP

Q) 6.78 g of copper is displaced from copper sulphate solution by 7 g zinc. Find the equivalent weight of copper, if equivalent weight of zinc is 32.5

Solution

According to law of chemical equivalence

Let E be the equivalent weight of copper

Number of gram equivalents of copper = number of gram equivalents of zinc

$$\frac{6.78}{E} = \frac{7}{31.5}$$

$$E = 31.5$$

Q) 100 mL of 0.1N NaOH solution is required for neutralisation of 0.49 g of dibasic acid. What is the molecular weight of acid.

Solution:

From the formula of milliequivalence

$$NV = \frac{W \times 1000}{Eq}$$

we will calculate milliequivalence of acid = $NV = 0.1 \times 100 = 10$

Milliequivalence of base = $\frac{W \times 1000}{Eq}$

$$\frac{0.49 \times 1000}{Eq}$$

Milliequivalence of base = milliequivalence of acid

$$10 = \frac{0.49 \times 1000}{Eq}$$

$$Eq = 49$$

Molecular weight = Equivalent weight \times basicity

Given acid is dibasic thus Molecular weight = $49 \times 2 = 98$

Q) Calculate the volume of each of the two hydrochloric acids of strength 12N and 4N respectively that are mixed to make one litre solution of strength 5N

Solution

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Let Volume of 12N solution be = x

Then volume of 4N solution will be = 1000 -x

Now $N_1V_1 + N_2V_2 = NV$

$$12x + 4(1000 -x) = 5(1000)$$

On simplification we get x = 125 mL

That is if 125 mL of 12 N solution is mixed with 875mL of 4 N solution we will get 1000mL of 5N solution.

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