SOME BASIC CONCEPTS OF CHEMISTRY

Chemistry is the science of molecules and their transformations. It is the science not so much of the one hundred elements but of the infinite variety of molecules that may be built from them. By **Roald Hoffmann**

DEVELOPMENT OF CHEMISTRY

- Chemistry came up as a result of search for two interesting things:

  i. Philosopher's stone (Paras) which would convert all baser metals e.g., iron and copper into gold.
  ii. 'Elixir' of life' which would grant immortality.

IMPORTANCE OF CHEMISTRY

- Principles of chemistry are applicable in diverse areas, such as weather patterns, functioning of brain and operation of a computer, production in chemical industries, manufacturing fertilisers, alkalis, acids, salts, dyes, polymers, drugs, soaps, detergents, metals, alloys, etc., including new material.

- Chemistry contributes in a big way to the national economy.

- It also plays an important role in meeting human needs for food, healthcare products and other material aimed at improving the quality of life.

- Chemistry provides methods for the isolation of life- saving drugs from natural sources and makes possible synthesis of such drugs.

- Some of these drugs are cisplatin and taxol, which are effective in cancer therapy. The drug AZT (Azidothymidine) is used for helping AIDS patients.

- With a better understanding of chemical principles it has now become possible to design and synthesise new material having specific magnetic, electric and optical properties.

- This has lead to the production of superconducting ceramics, conducting polymers, optical fibres, ete.

- Chemistry has helped in establishing industries which manufacture utility goods, like acids, alkalies, dyes, polymers metals, etc. These industries contribute in a big way to the economy of a nation and generate employment.

- Safer alternatives to environmentally hazardous refrigerants, like CFCs (chlorofluorocarbons), responsible for ozone depletion in the stratosphere, have been successfully synthesised.

- One problem is the management of the Green House gases, like methane, carbon dioxide, etc.

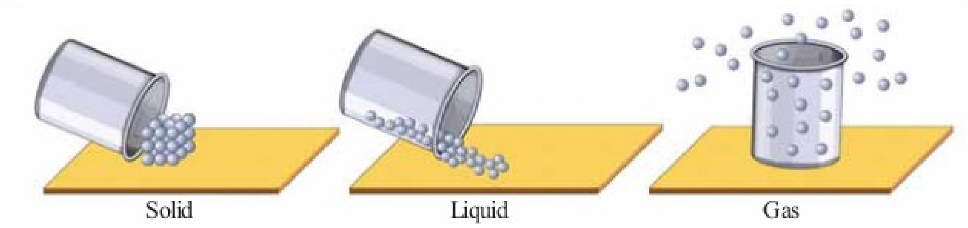
- Understanding of biochemical processes, use of enzymes for large-scale production of chemicals and synthesis of new exotic material are some of the intellectual challenges for the future generation of chemists.

NATURE OF MATTER

- Anything which has mass and occupies space is called matter.

- Everything around us, for example, book, pen, pencil, water, air, all living beings, etc., are composed of matter.

- Matter can exist mainly in three physical states viz. solid, liquid and gas.



- Particles are held very close to each other in solids in an orderly fashion and there is not much freedom of movement.

- In liquids, the particles are close to each other but they can move around.

- In gases, the particles are far apart as compared to those present in solid or liquid states and their movement is easy and fast.

- Based on the arrangement of particles, different states of matter exhibit the following characteristics:
- Solids have definite volume and definite shape.
- Liquids have definite volume but do not have definite shape. They take the shape of the container in which they are placed.

- Gases have neither definite volume nor definite shape. They take the volume of the container in which they are taken.

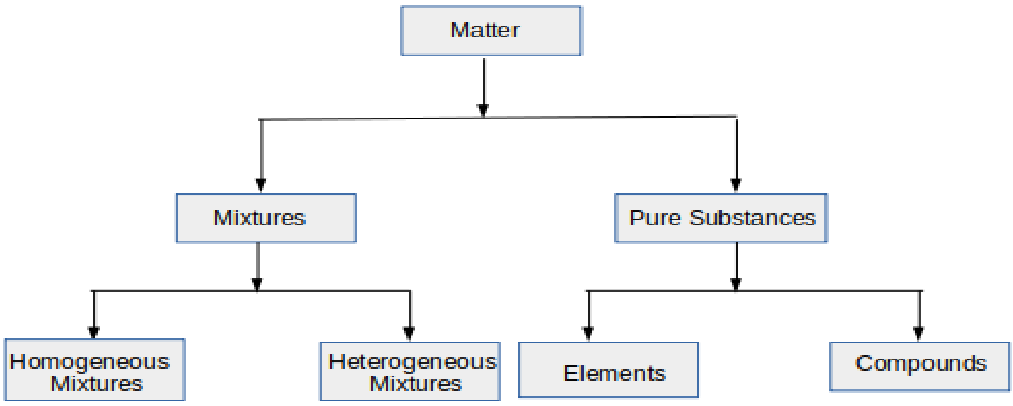
- These three states of matter are interconvertible by changing the conditions of temperature and pressure.

- Solid ==> liquid == Gas

- On heating, a solid usually changes to a liquid, and the liquid on further heating changes to gas (or vapour).
- In the reverse process, a gas on cooling liquifies to the liquid and the liquid on further cooling freezes to the solid.

Classification of Matter

Generally matter can be classified as mixture or pure substance. These can be further sub-divided as shown.



✔ When all constituent particles of a substance are same in chemical nature, it is said to be a pure substance.

✔ A mixture contains particles of two or more pure substances which may be present in it in any ratio. Hence their composition is variable.

✔ Many of the substances present around you are mixtures. For example, sugar solution in water, air, tea, etc., are all mixtures.

✔ A mixture may be homogeneous or heterogeneous.

✔ In a homogeneous mixture, the components completely mix with each other.The particles of components of the mixture are uniformly distributed throughout the bulk of the mixture and its composition is uniform throughout.

✔ Sugar solution and air are the examples of homogeneous mixtures.

✔ In a heterogeneous mixture, the composition is not uniform throughout and sometimes different components are visible.

✔ For example, mixtures of salt and sugar, grains and pulses along with some dirt (often stone pieces) are heterogeneous mixtures.

✔ The components of a mixture can be separated by using physical methods, such as simple hand-picking, filtration, crystallisation, distillation, etc.

✔ Copper, silver, gold, water and glucose are some examples of pure substances.

- Glucose contains carbon, hydrogen and oxygen in a fixed ratio and its particles are of same composition. Also its constituents—carbon, hydrogen and oxygen—cannot be separated by simple physical methods.

- Pure substances can further be classified into elements and compounds.

- Particles of an element consist of only one type of atoms. These particles may exist as atoms or molecules.

- Elements such as sodium or copper, contain atoms as their constituent particles. whereas, in some others, the constituent particles are molecules which are formed by two or more atoms.

- For example, hydrogen, nitrogen and oxygen gases consist of molecules.

- When two or more atoms of different elements combine together in a definite ratio, the molecule of a compound is obtained.

- The constituents of a compound cannot be separated into simpler substances by physical methods.
- They can be separated by chemical methods.

- Examples of some compounds are water, ammonia, carbon dioxide, sugar, ete.

- The properties of a compound are different from those of its constituent elements.

- For example, hydrogen and oxygen are gases, whereas, the compound formed by their combination i.e., water is a liquid.

- Note that hydrogen burns with a pop sound and oxygen is a supporter of combustion, but water is used as a fire extinguisher.

PROPERTIES OF MATTER

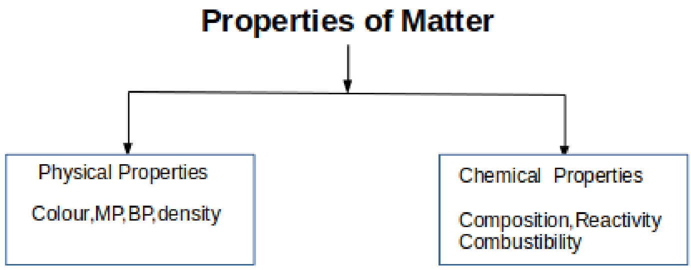
- The properties of matter can be classified into two categories — physical properties, and chemical properties

- Physical properties such as colour, odour, melting point, boiling point, density, etc.,,

- Chemical properties like composition, combustibility, reactivity with acids and bases, etc.

- Physical properties can be measured or observed without changing the identity or the composition of the substance.

- The measurement or observation of chemical properties requires a chemical change to occur.



The International System of Units (SI)

- The SI system has seven base units . These units pertain to the seven fundamental scientific quantities.

- The other physical quantities, such as speed, volume, density, etc., can be derived from these quantities.

Physical Quantity	Symbol for Quantity	SI Unit	Symbol for 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

Prefixes used in the SI System

Multiple	Prefix	Symbol
10^{-24}	yocto	y
10^{-21}	zepto	z
10^{-18}	atto	z
10^{-15}	femto	f
10^{-12}	pico	p
10^{-9}	nano	n
10^{-6}	micro	
10^{-3}	milli	m

10^{-2}	centi	c
10^{-1}	deci	d
10	deca	da
10^{2}	hecto	h
10^{3}	Kilo	k
10^{6}	mega	M
10^{9}	giga	G
10^{12}	tera	T
10^{15}	peta	P
10^{18}	exa	E
10^{21}	zeta	Z
10^{24}	yotta	Y

The temperatures on two scales are related to each other by the following relationship:

- **^{o}F = 9/5 (^{0}C) +32** **(**Relation between Fahrenheit and degree Celsius**)**
- The kelvin scale is related to Celsius scale. **K=°C + 273.15**

Scientific Notation

- A chemist has to deal with numbers as large as 602, 200,000,000,000,000,000,000 for the molecules of 2 g of hydrogen gas or as small as 0.00000000000000000000000166 g mass of a H atom.
- Other constants such as Planck's constant, speed of light, charges on particles, etc., involve numbers of the above magnitude.
- Scientific notation is an exponential notation in which any number can be ex-presented in the form **N * 10^{n}, where n is an exponent having positive or negative values and N is a number (called digit term) which varies between 1.000... and 9.999....**

Examples.

- **We can write 232.508 as 2.32508 *10^{2} in scientific notation**. While writing it, the decimal had to be moved to the left by two places and same is the exponent (2) of 10 in the scientific notation.
- **Similarly, 0.00016 can be written as 1.6 * 10^{-4}**. Here, the decimal has to be moved four places to the right and (-4) is the exponent in the scientific notation.

Significant Figures

- Every experimental measurement has some amount of uncertainty associated with it because of limitation of measuring instrument and the skill of the person making the measurement.

For example,

- Mass of an object is obtained using a platform balance and it comes out to be **9.4g.**
- On measuring the mass of this object on an analytical balance, the mass obtained is **9.4213g.**
- The mass obtained by an analytical balance is slightly higher than the mass obtained by using a platform balance. Therefore, digit 4 placed after decimal in the measurement by platform balance is uncertain.

➢ The uncertainty in the experimental or the calculated values is indicated by mentioning the number of significant figures.

➢ Significant figures are meaningful digits which are known with certainty

Rules for determining the number of significant figures.

✔ (1) **All non-zero digits are significant.**

For example

- In 285 cm, there are **three significant figures** and in 0.25 mL, there are **two significant** figures.

✔ (2) **Zeros between two non-zero digits are significant.**

For example

- Thus, **2.005 has four significant figures.**

✔ (3) **Zeros preceding to first non-zero digit are not significant.** Such zero indicates the position of decimal point.

For example

- Thus, 0.03 has **one significant figure** and 0.0052 has t**wo significant figures.**

✔ (4) Zeros at the end or right of a number are significant, provided they are on the right side of the decimal point.

For example,

- 0.200 g has three significant figures. But, if otherwise, 200, the terminal zeros are not significant if there is no decimal point.

- 100 has only one significant figure, but 100. has three significant figures and 100.0 has four significant figures.

Such numbers are better represented in scientific notation.

- We can express the number 100 as $1x10^2$ for **one significant figure.**

- we can express the number 100 as $1.0x10^2$ for **two significant figures.**

- We can express the number 100 as $1.00 * 10^2$ for t**hree significant figures.**

✔ I**n numbers written in scientific notation, all digits are significant**

Example

- $4.01X10^2$ has **three significant figures**, and 8.256×10^{-3} has **four significant figures.**

Addition and Multiplication of Significant Figures

- Here, 12.11 has 4 significant figures, 18.0 has 3 significant figures and 1.012 has 4 significant figures. Sum of these numbers is 31.122 which contains 5 significant figures.

- The result should contain minimum number significant . That is 3 significant figures.

a) 12.11
18.0
1.012
=======
31.122 Result = 31.1

b) **2.5**X1.25 = 3.125 Result = 3.1

- Sum is **31.122** which has **5 significant figures**. Hence result is adjusted to 3 significant figures, that is , **31.1**

- Similarly, in multiplication of 2.5 and 1.25 result is 3.125.Where in 2.5 contains 2 significant figures and 1.25 contains 3 significant figures. So the result should contain minimum number of 2 significant figures. Hence the result is 3.1

- **Similarly for Subtraction and Division.**

c) 12.258 / **2.32** = 5.28362069
= 5.28

d) 12.231–**6.98** = 5.251
= 5.25

While limiting the result to the required number of significant figures as done in the above mathematical operation, one has to keep in mind the following points for rounding off the numbers.

- If the rightmost digit to be removed is more than 5, the preceding number is increased by one.

For example,

- 1.386. If we have to remove 6, we have to round it to 1.39.

- If the rightmost digit to be removed is less than 5, the preceding number is not changed.

For example,

- 4.334 if 4 is to be removed, then the result is rounded up to 4.33.

- If the rightmost digit to be removed is 5, then the preceding number is not changed if it is an even number but it is increased by one if it is an odd number.

For example,

- If 6.35 is to be rounded by removing 5, we have to increase 3 to 4 giving 6.4 as the result. However, if 6.25 is to be rounded off it is rounded off to 6.2.

Precision and accuracy are often referred to while we talk about the measurement.

- ✗ Precision refers to the closeness of various measurements for the same quantity.
- ✗ However, accuracy is the agreement of a particular value to the true value of the result.

For example,

- ➢ If the true value for a result is 2.00 g and student ‘A’ takes two measurements and reports the results as 1.95 g and 1.93 g. These values are precise as they are close to each other but are not accurate.

- ➢ Another student ‘B’ repeats the experiment and obtains 1.94 g and 2.05 g as the results for two measurements. These observations are neither precise nor accurate.

- ➢ When the third student ‘C’ repeats these measurements and reports 2.01 g and 1.99 g as the result, these values are both precise and accurate.

- ➢ **This can be more clearly understood from the data given in Table .**

Student	1	2	Average (g)
Student A	1.95	1.93	1.940
Student B	1.94	2.05	1.995
Student C	2.01	1.99	2.000

LAWS OF CHEMICAL COMBINATIONS

The combination of elements to form compounds is governed by the following **five fundamental laws.**

1. Law of Conservation of Mass

[Antoine Lavoisier after the French Revolution, was a French nobleman and chemist who was central to the 18th-century chemical revolution and who had a large influence on both the history of chemistry and the history of biology.]

This law was put forth by **Antoine Lavoisier in 1789**. He performed careful experimental studies for combustion reactions and reached to the conclusion that **in all physical and chemical changes, there is no net change in mass during the process.**
That is, **matter can neither be created nor destroyed in any chemical reaction**. This is called **'Law of Conservation of Mass'**.In fact, this was the result of exact measurement of masses of reactants and products.

For Example,

- Reaction between Hydrogen and Oxygen to give Water.
  $2H_2$ (g) + O_2 (g) ....> $2H_2O$ (l)
  4 g H_2 + 32 g O_2 ....> 36 g H_2O

  Mass of reactants = Mass of products = 36 g

2. Law of Definite Proportions

[Joseph Louis Proust was a French chemist. He was best known for his discovery of the law of constant composition in 1794, stating that chemical compounds always combine in constant proportions.]

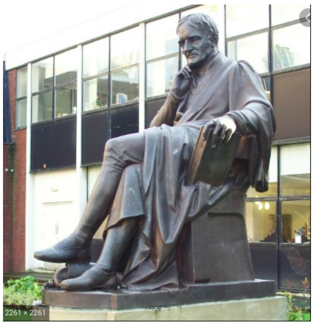
This law was given by, a French chemist, J**oseph Proust.**

He stated that **a given compound always contains same elements combined together in the same proportion by mass**. This law has been confirmed by various experiments. It is known as **Law of Definite Composition.**

Proust worked with two samples of cupric carbonate — one of which was of natural origin and the other was synthetic. He found that the composition of elements present in it was same for both the samples as shown below:

	copper %	**carbon %**	**oxygen%**
Natural Sample	51.35	9.74	38.91
Synthetic Sample	51.35	9.74	38.91

3. Law of Multiple Proportions.



[John Dalton was an English chemist, physicist, and meteorologist. He is best known for introducing the atomic theory into chemistry, and for his research into colour blindness, sometimes referred to as Dalton ism in his honour.]

This law was proposed by Dalton in 1803.

According to this law, if **two elements can combine to form more than one compound, the masses of one element that combine with a fixed mass of the other element, are in the ratio of small whole numbers.**

For example.

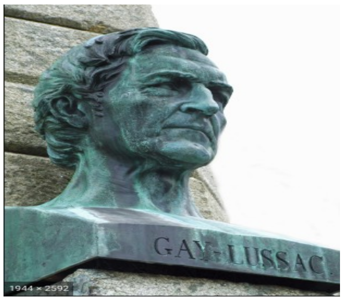
Hydrogen combines with Oxygen to form two compounds, namely, water and hydrogen peroxide.

Hydrogen	+	Oxygen	>	Water
2g		16g		18g

Hydrogen	+	Oxygen	>	Hydrogen Peroxide
2g		32g		34g

Here, **the masses of oxygen (i.e., 16 g and 32 g), which combine with a fixed mass of hydrogen (2g) bear a simple ratio, i.e., 16:32 or 1: 2.**

4. Gay Lussac's Law of Gaseous Volumes



[Joseph Louis Gay-Lussac was a French chemist and physicist. He is known mostly for his discovery that water is made of two parts hydrogen and one part oxygen, for two laws related to gases.]

This law was given by Gay Lussac in 1808.

He observed that when gases combine or are produced in a chemical reaction they do so in a simple ratio by volume, provided all gases are at the same temperature and pressure.

Thus, 100 mL of hydrogen 5 combine with 50 mL of oxygen Gay Lussac to give 100 mL of water vapour.

Hydrogen	+	Oxygen	>	Water
100 mL		50 mL		100 mL

Thus, the volumes of hydrogen and oxygen which combine (i.e., 100 mL and 50 mL) bear a simple ratio of 2:1.

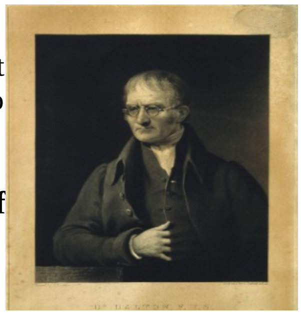
Gay Lussac's discovery of integer ratio in volume relationship is actually the law of definite proportions by volume.

5. Avogadro's Law

[**Lorenzo Romano Amedeo Carlo Avogadro**, Count of Quaregna and Cerreto was an Italian scientist, most noted for his contribution to molecular theory now known as Avogadro's law, which states that equal volumes of gases under the same conditions of temperature and pressure will contain equal numbers of molecules]

- In 1811, Avogadro proposed that equal volumes of all gases at the same temperature and pressure should contain equal number of molecules.
- Avogadro made a distinction between atoms and molecules which is quite understandable in present times.
- If we consider again the reaction of hydrogen and oxygen to produce water, we see that two volumes of hydrogen combine with one volume of oxygen to give two volumes of water without leaving any unreacted oxygen.

DALTON'S ATOMIC THEORY



[John Daltonwas an English chemist, physicist, and meteorologist. He is best known for introducing the atomic theory into chemistry, and for his research into colour blindness, sometimes referred to as Daltonism in his honour.]

- In 1808, Dalton John Dalton published '**A New System of Chemical Philosophy**', in which he proposed the following :

  i) Matter consists of indivisible atoms.

  ii) All atoms of a given element have identical properties, including identical mass. Atoms of different elements differ in mass.

  Iii) Compounds are formed when atoms of different elements combine in a fixed ratio.

  vi) Chemical reactions involve reorganisation of atoms. These are neither created nor destroyed in a chemical reaction.

- ✔ Dalton's theory could explain the laws of chemical combination. However, it could not explain the laws of gaseous volumes. It could not provide the reason for combining of atoms

The atomic mass

- The mass of an atom is actually very-very small because atoms are extremely small, but we have mass spectrometry for determining the atomic masses fairly accurately.
- But in the nineteenth century, scientists could determine the mass of one atom relative to another by experimental means, Hydrogen, being the lightest atom was arbitrarily assigned a mass of 1 and other elements were assigned masses relative to it.
- The present system of atomic masses is based on carbon-12 as the standard, Carbon-12 is one of the isotopes of carbon and can be represented as ^{12}C.
- In this system, ^{12}C is assigned a mass of exactly 12 atomic mass unit (amu) and masses of all other atoms are given relative to this standard.
- One atomic mass unit is defined as a mass exactly equal to one-twelfth of the mass of one carbon -12
- ➢ 1 amu = 1.66056×10^{-24} g

- ✔ Mass of hydrogen atom = 1.0080 amu

- ✔ Mass of oxygen atom= 15.995 amu.
- ✔ At present, ‘amu’ has been replaced by ‘u’, which is known as unified mass.

Average Atomic Mass

- ◆ Many naturally occurring elements exist as more than one isotope. When we take into account the existence of these isotopes and their relative abundance (per cent occurrence), the average atomic mass of that element can be computed.

For example,

- ➔ Carbon has the following three isotopes with relative abundances and masses.

Isotope	Relative Abundance (%)	Atomic mass (amu)
^{12}C	98.892	12
^{133}C	1.108	13.00335
^{14}C	$2 * 10^{-12}$	14.00317

- ➢ **From the above data, the average atomic mass of carbon will come out to be**:
  (0.98892) (12 u) + (0.01108) (13.00335 u) + (2 x 10^{-12}) (14.00317 u) = **12.011 u**

- ➔ Similarly, average atomic masses for other elements can be calculated.
- ➔ In the periodic table of elements, the **atomic masses** mentioned for different elements actually represent their **average atomic masses**.

Molecular Mass

- ◆ Molecular mass is the sum of atomic masses of the elements present in a molecule. It is obtained by multiplying the atomic mass of each element by the number of its atoms and adding them together.

  For example,

- ◆ molecular mass of methane, which contains one carbon atom and four hydrogen atoms, can be obtained as follows:
- ➔ Molecular mass of methane, (CH_4) = 1*(12.011 w) +4 *(1.008 u) = 12.011 + 4.032 = **16.043** u
- ➔ Similarly, molecular mass of water (H_2O) = 2X atomic mass of hydrogen + 1 x atomic mass of oxygen = 2 * (1.008 u) + 1 * (16.00 u) = **18.02** u

Formula Mass

- ➔ Some substances, such as sodium chloride, do not contain discrete molecules as their constituent units.
- ➔ The formula, such as NaCl, is used to calculate the formula mass instead of molecular mass as in the solid state sodium chloride.

MOLE CONCEPT AND MOLAR MASSES

- ➔ we use the idea of mole to count entities at the microscopic level (i.e., atoms, molecules, particles, electrons, ions, etc).
- ➔ 1 mole is known as 'Avogadro constant', or Avogadro number denoted by N, in honour of Amedeo Avogadro.
- ➔ We can write it with all zeroes without using any powers of ten. 6022 1367000000000000000000 (**$6.022 * 10^{23}$**) Hence, so many entities (atoms, molecules or any other particle) constitute one mole of a particular substance.

Examples

- ➔ We can say that 1 mol of hydrogen atoms = $6.022x10^{23}$ atoms
- ➔ 1 mol of water molecules = $6.022x10^{23}$ water molecules
- ➔ 1 mol of sodium chloride = 6.022×10^{23} formula units of sodium chloride
- ➔ **The mass of one mole of a substance in grams is called its molar mass.**
- ➔ **The molar mass in grams is numerically equal to atomic/molecular/ formula mass in u.**
- ◆ Molar mass of water = 18.02 g/ mol ; Molar mass of sodium chloride = 58.5 g /mol

PERCENTAGE COMPOSITION

- ◆ Information regarding the percentage of a particular element present in a compound is required. The purity of a given sample by analysing this data.
- ➢ Let us understand it by taking the example of water (H_2O). Since water contains hydrogen and oxygen, the percentage composition of both these elements can be calculated as follows:

- Generally,

Mas % of any element

$$= \frac{\text{mass of that element in the compound}}{\text{Molar mass of the compound}} * 100$$

Molar mass of water = 18.02 g

$$\text{Mass \% of hydrogen} = \frac{2 * 1.008}{18.02} * 100 = 11.18\%$$

$$\text{Mass \% of Oxygen} = \frac{1 * 16.00}{18.02} * 100 = 88.79\%$$

- ✔ Let us take one more example. What is the percentage of carbon, hydrogen and oxygen in ethanol?

Molecular formula of ethanol = **C_2H_6O**

Molar mass of ethanol **=** (2x12.01 + 6x1.008 + 16.00) g = 46.068 g

Mass per cent of carbon = $\frac{(2x12.01)}{46.068} * 100$ = **52.14 %**

Mass per cent of hydrogen = $\frac{6x1.008}{46.068} * 100$ = **13.13 %**

Mass per cent of oxygen = $\frac{16.00}{46.068} * 100$ = **34.73 %**

Empirical Formula for Molecular Formula

- An empirical formula represents the simplest whole number ratio of various atoms present in a compound .
- The molecular formula shows the exact number of different types of atoms present in a molecule of a compound.
- If the mass per cent of various elements present in a compound is known, its empirical formula can be determined. Molecular formula can further be obtained if the molar mass is known.

Analyse the question to understand the concept of Empirical formula & Molecular formula.

- A compound contains 4.07% hydrogen, 24.27% carbon and 71.65% chlorine. Its molar mass is 98.96 g. What are its empirical and molecular formulas?

Solution

Step 1. Conversion of mass per cent to grams

✔ Since we are having **mass per cent**, it is convenient to use **100 g of the compound** as the starting material. Thus, in the **100 g sample of the above compound, 4.07g hydrogen, 24.27g carbon and 71.65g chlorine are present.**

Step 2. Convert into number moles of each element

✔ Divide the masses obtained above by respective atomic masses (GAM) of various elements. This gives the number of moles of constituent elements in the compound.

$$\text{Moles of hydrogen} = \frac{\text{Mass \%}}{\text{GAM}} = \frac{4.07\text{ g}}{1.008\text{ g}} * 100 = 4.04$$

$$\text{Moles of carbon} = \frac{\text{Mass \%}}{\text{GAM}} = \frac{24.27\text{ g}}{12.01\text{ g}} * 100 = 2.021$$

$$\text{Moles of chlorine} = \frac{\text{Mass \%}}{\text{GAM}} = \frac{71.65\text{ g}}{35.453\text{ g}} * 100 = 2.021$$

Step 3. Divide each of the mole values obtained above by the smallest number amongst them.

✔ Since 2.021 is smallest value, division of each above value by it (**4.04/2.021 : 2.021/2.021 : 2.021/2.021**) **gives a ratio of 2:1:1 for H:C:Cl.**

(Note : In case the ratios are not whole numbers, then they may be converted into whole number by multiplying by the suitable coefficient.)

Step 4. Write down the empirical formula by mentioning the numbers after writing the symbols of respective elements.

CH_2CI is the empirical formula of the above compound.

Step 5. Writing molecular formula.

- (a) **Determine empirical formula mass by adding the atomic masses of various atoms present in the empirical formula.**

  For CH_2Cl, empirical formula mass = 12.01 + (2 x 1.008) + 35.453 = **49.48 g**

(b) Divide Molar mass by empirical formula mass. $n = \frac{\text{Molecular Formula Mass}}{\text{Empirical Formula mass}}$

$$n = \frac{98.96}{49.48} = 2$$

(c) Multiply empirical formula by n obtained above to get the molecular formula.

Molecular Formula = (Empirical Formula)$_n$

Molecular Formula = $(CH2Cl)_2$ = $C_2H_4Cl_2$

STOICHIOMETRY AND STOICHIOMETRIC CALCULATIONS

- Stoichiometry deals with the calculation of masses (sometimes volumes also) of the reactants and the products involved in a chemical reaction.

- Let us study what information is available from the balanced chemical equation of a given reaction.

- Let us consider the combustion of methane.

  CH_4 (g) + $2O_2$ (g) .....> CO_2(g) + 2 H_2O (g)

- ✔ Here, methane and dioxygen are called reactants and carbon dioxide and water are called products.

- ✔ The coefficients 2 for O_2 and H_2O are called stoichiometric coefficients. Similarly the coefficient for CH_4 and CO_2 is one in each case.

Different interpretation of the balanced chemical equation.

Thus, according to the above chemical reaction,

- One mole of CH_4(g) + two moles of O_2(g) to give one mole of CO_2(g) .....> two moles of H_2O(g)

- One molecule of CH_4(g) + 2 molecules of O_2(g).....> one molecule of CO_2(g) + 2 molecules of H_2O(g)

- 22.4 L ofCH_4(g) + 44.8 L of O_2 (g) ......> 22.4 L of CO_2 (g) + 44.8 L of H_2O(g) (STP)

- 16 gofCH_4 (g) + 64 g of O_2 (g) ......> 44 g of CO_2 (g) + 36 g of H_2O (g).

Limiting Reagent

- The reactant which is present in the least amount gets consumed after sometime and after that further reaction does not take place whatever be the amount of the other reactant. Hence, the reactant, which gets consumed first, limits the amount of product formed and is, therefore, called the limiting reagent.

Reactions in Solutions

- It is important to understand as how the amount of substance is expressed when it is present in the solution.

- The concentration of a solution or the amount of substance present in its given volume can be expressed in any of the following ways.

  1. Mass per cent or weight per cent (w/w %)
  2. Mole fraction
  3. Molarity
  4. Molality

Mass per cent

It is obtained by using the following relation:

$$\text{Mass per cent} = \frac{\text{Mass of solute}}{\text{Mass of solution}} * 100$$

Mole Fraction (x_A)

- It is the ratio of number of moles of a particular component to the total number of moles of the solution,

- If a substance 'A' dissolves in substance 'B' and their number of moles are n_A and n_B respectively, then the mole fractions of A and B are given as:

- Mole fraction of A = $\frac{\text{No.of molesof A}}{\text{No. of moles of solutions}} = \frac{n_A}{n_A + n_B}$

- Mole fraction of B = $\frac{\text{No.of moles of B}}{\text{No. of moles of solutions}} = \frac{n_B}{n_A + n_B}$

Molarity (M)

It is defined as the number of moles of the solute in 1 litre of the solution.

Thus,

$$\text{Molarity} = \frac{\text{No.of moles of solute}}{\text{Volume of solution in litres}}$$

- For the calculation of a particular concentration , a general formula, $\mathbf{M_1 x\ V_1 = M_2\ x\ _2}$, where M and V are molarity and volume, respectively, can be used.

Molality

- It is defined as the number of moles of solute present in 1 kg of solvent. It is denoted by m.

$$\text{Molality (m)} = \frac{\text{No. of moles of solute}}{\text{Mass of solvent in kg}}$$

- **Note :** Molarity of a solution depends upon temperature because volume of a solution is temperature dependent.

Problem 1.

1. Calculate the molarity of NaOH in the solution prepared by dissolving its 4 g in enough water to form 250mL of the solution. (Atomic Mass Na = 23,O = 16, H =1)

$$\text{No.of moles of solute (NaOH)} = \frac{\text{Mass in gram}}{\text{GMM}} = 4/40 = 0.1 \text{ moles}$$

$$\text{Molarity (M)} = \frac{\text{No.of moles of solute}}{\text{Volume of solution in mL}} * 1000 = \frac{0.1}{250} * 1000 = \mathbf{0.4\ M}$$

=====

*** GMM = Gram Molecular Mass**

Problem 2

- The density of 3 M solution of NaCl is 1.25 g /mL Calculate the molality of the solution.(GMM of NaCl = 58.5 g)

  M = 3 mol/L
  Mass of solute (NaCl) in 1L solution = 3 * 58.5 = 175.5 g
  Mass of 1L solution = 1000 * 1.25 = 1250 g
  Mass of water in 1L solution = 1250 - 175.5 = 1074.5 g

  $$\text{Molality (m)} = \frac{\text{No. of moles of solute}}{\text{Mass of solvent in gram}} * 1000 = \frac{3}{\mathbf{1074.5}} * 1000 = \frac{3000}{\mathbf{1074.5}} = \mathbf{2.79\ m}$$

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NCERT Questions & Answers

Question 1

- Calculate the molar mass of the following: (i) H_2O (ii) CO_2 (iii) CH_4
  [Atomic mass ; C=12 , H = 1, O = 16]

Answer

Molar mass of H_2O = (2*1) + 16 = 18 g ; CO_2 = 44 g ; CH_4 = 16 g

Question 2

- Calculate the mass per cent of different elements present in sodium sulphate (Na_2SO_4).
  [Atomic mass ; Na = 23 , S = 32 , O = 16]

Answer

- a) Mass % of Na = $\frac{\text{Atomic mass of Na}}{\text{Molecular Mass of }(Na_2SO_4)} * 100 = \frac{46}{142} * 100 = 32.39\ \%$
  =======
- b) Mass % of S = $\frac{\text{Atomic mass of S}}{\text{Molecular Mass of }(Na_2SO_4)} * 100 = \frac{32}{142} * 100 = 22.54\ \%$
  =======
- c) Mass % of O = $\frac{\text{Atomic mass of O}}{\text{Molecular Mass of }(Na_2SO_4)} * 100 = \frac{64}{142} * 100 = 45.07\ \%$
  =======

Question 3

- Determine the empirical formula of an oxide of iron, which has 69.9% iron and 30.1% dioxygen by mass.

Answer

Element	Fe	O
% by mass	69.9%	30.1%
Atomic mass	55.85	16
% / Atomic mass	69.9/55.85 = **1.25 (Least Value)**	30.1/16 = 1.88
Value / Least Value	1.25/1.25 = 1	1.88/1.25 = 1.5 (decimal value)
Make decimal value whole number value	1*2 = **2**	1.5 * 2 = **3**
Empirical Formula		**Fe_2O_3**

Question 4

- Calculate the amount of carbon dioxide that could be produced when

  (i) 1 mole of carbon is burnt in air.

  (ii) 1 mole of carbon is burnt in 16 g of dioxygen.

  (iii) 2 moles of carbon are burnt in 16 g of dioxygen.

Answer

- Chemical equation of the formation of Carbon Dioxide.
  $C + O_2$ ....> CO_2
- (i) Amount of Carbon Dioxide = 1 mole
- (ii) Amount of Carbon Dioxide = 1/2 mole
- (iii)Amount of Carbon Dioxide = 1/2 mole

Question 5

- Calculate the mass of sodium acetate (CH_3COONa) required to make 500 mL of 0.375 molar aqueous solution. Molar mass of sodium acetate is 82.0245 g mol.

Answer

- *Molarity* $= \dfrac{\text{No. of moles of solute}}{\text{Volume of Solution (mL)}} * 1000 = \dfrac{\text{Mass of solute/GMM of solute}}{\text{Volume of Solution (mL)}} *1000$

$$M = \frac{w/GMM}{V} * 1000 = \frac{w}{V * GMM} * 1000$$

w = mass of solute
V= Volume of solution
GMM = Gram Molecular Mass

$$0.375 = \frac{w}{500 * 82.0245} \mathbf{1000} \quad ; \quad \mathbf{w = 0.375} * 500 * 82.0245 / 1000 = 15.38\ g$$

Mass of sodium acetate = 15.38 g

Question 6

- *Calculate the concentration of nitric acid in moles per litre in a sample which has a density, 1.41 g mL and the mass per cent of nitric acid in it being 69%.*

Answer

- *Mass % =* $\frac{\text{Mass of solute}}{\text{Mass of solution}} * 100$ *; mass% is given as 69% .*

  Therefore , mass of solute = 69 g ; GMM of Nitric Acid = 63 g

  No. of moles of Nitric Acid = Mass / GMM = 69 / 63 = 1.095

  Volume of solution = Mass / Density = 100 g / 1.41 g mL = 70.92 mL

  $$\text{Molarity} = \frac{\text{No. of moles of solute}}{\text{Volume of Solution (mL)}} * 1000 = \frac{1.095}{70.92} * 1000 = 15.44\ \text{moles / L}$$

Question 7

- *How much copper can be obtained from 100 g of copper sulphate ($CuSO_4$)?*

Answer

- *Mass of copper obtained from 100 g $CuSO_4$* $= \frac{\text{Atomic mass of Cu}}{\text{GMM of } CuSO_4} * \text{Given mass of } CuSO_4$

  $$= \frac{63.54}{159.54} * 100 = 39.83\ g$$

Question 8

Calculate the atomic mass (average) of chlorine using the following data:

Isotope of Chlorine	**% natural Abundance**	**Molar Mass**
^{35}Cl	75.77	34.9689
^{37}Cl	24.23	36.3659

Answer

- *Average atomic mass = (34.9689 * 75.77%) + (36.3659 * 24.23%) = 35.45 g*
  =====

Question 9

- *In three moles of ethane (C_2H_6), calculate the following:*

I. *Number of moles of carbon atoms.*
II. *Number of moles of hydrogen atoms.*
III. *Number of molecules of ethane.*

Answer

- *i) No. Of moles of carbon atoms in 3 moles of C_2H_6 = 3*2 = 6 moles Carbon atoms*
  *ii) No. Of moles of hydrogen atoms in 3 moles of C_2H_6 = 3*6 = 18 moles Carbon atoms*
  iii) No. Of moles of molecules C_2H_6 in 3 moles = $3*6.022*10^{23}$ molecules

Question 10

- *What is the concentration of sugar ($C_{12}H_{22}O_{11}$,) in mol/ L if its 20 g are dissolved in enough water to make a final volume up to 2L?*

Answer

- *Molarity* $= \dfrac{\text{No. of moles of solute}}{\text{Volume of Solution (L)}} = \dfrac{\text{Mass of solute/GMM of solute}}{\text{Volume of Solution (L)}}$

  $= \dfrac{20/342}{2} = \dfrac{20}{2*342} = 0.02924\ M$
  =======

Question 11

- *If the density of methanol is 0.793 kg L', what is its volume needed for making 2.5 L of its 0.25 M solution?*

Answer

- $$\text{Molarity} = \frac{\text{No. of moles of solute}}{\text{Volume of Solution (L)}} =$$

$$0.25 = \frac{w}{V(L) * GMM} = \frac{w}{2.5 * 32} = w / 80$$

$w = 0.25 * 80 = 20\ g = 0.02\ kg$

Volume = Mass / Density = 0.02/ 0.793 = 0.02522 L = 25.22 mL

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Question 12

- *Pressure is determined as force per unit area of the surface. The SI unit of pressure, pascal is as shown below:*
  IPa = IN m^{-2}
  If mass of air at sea level is 1034 g cm^{-2}*, calculate the pressure in pascal.*

Answer

$$\text{Mass of air} = 1034\ g/cm^2 = \frac{1034 * 10^{-3}\ kg}{10^{-4}\ m^2} = 10340\ kg / m^2$$

- $$\text{Pressure} = \text{Force Area} = mg / \text{Area} = \frac{10340 * 9.8}{\text{Area}} = 101232\ N/m^2 = 1.01232 * 10^5\ N/m^2$$

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Question 13

- *A sample of drinking water was found to be severely contaminated with chloroform,* $CHCI_3$ *supposed to be carcinogenic in nature. The level of contamination was 15 ppm (by mass).*

I. *Express this in per cent by mass.*
II. *Determine the molality of chloroform in the water sample.*

Answer

- $$\text{i) ppm} = \frac{\text{No. of parts of the component } CHCI_3}{\text{Total No. Of parts of all the components in the solution}} * 10^6$$

$$= \frac{\text{No. of parts by mass of the component } CHCI_3}{\text{Total No. Of parts of all the components by mass in the solution}} = 15*10^{-6} * 10^2$$

$= 15 * 10^{-4}\ g$

=======

ii) Molality of solution = $\frac{\text{No. Of moles of solute } CHCI_3}{\text{Mass of the solution (g)}} * 1000$

$= \frac{w}{GMM * \text{Mass of solution (g)}} 1000$

$= \frac{15*10^{-4}}{119.5 * 100} *1000 = 0.0001255\ m = 1.255 *10^{-4}\ m$

========

Question 14

Express the following in the scientific notation:

a) 0.0048
b) 234,000
c) 8008
d) 500.0
e) 6.0012

Answer

- $0.0048 = 4.8 * 10^{-3}$
- $234,000 = 2.34 * 10^{5}$
- $8008 = 8.008 * 10_{3}$
- $500.0 = 5.000 * 10_{2}$
- $6.0012 = 6.0012$

Question 15

- *How many significant figures are present in the following?*

(a) 0.0025
(b) 208
(c) 5005
(d) 126,000
(e) 500.0
(f) 2.0034

Answer

(a) $0.0025 = 2.5 * 10^{-3}$ = 2 Significant Figures
(b) $208 = 2.08 * 10^{2}$ = 3 Significant Figures
(c) $5005 = 5.005 * 10^{3}$ = 4 Significant Figures
(d) $126,000 = 1.26 * 10^{3}$ = 3 Significant Figures

*(e) 500.0 = 5.000 * 10^2 = 4 Significant Figures*
*(f) 2.0034 = 2.0034 * 10^0 = 5 Significant Figures*

Question 16

- *Round up the following up to three significant figures:*

a) *34.216*
b) *10.4107*
c) *0.04597*
d) *2808*

Answer

a) *34.216 = 3.42*
b) *10.4107 = 10.4*
c) *0.04597 = 0.046*
d) *2808 = 281*

Question 17

- *The following data are obtained when dinitrogen and dioxygen react together to form different compounds:*

Mass of dinitrogen	***Mass of dioxygen***
4g	*16g*
4g	*32g*
28 g	*32g*
28g	*80g*

a) *Which law of chemical combination is obeyed by the above experimental data? Give its statement.*

Answer

a) *i) In the first case 4 g dinitrogen combines with 16 g and 32 g dioxygen separately. Here , the masses of dioxygen combines (16 g and 32 g) combines with fixed mass of dinitrogen bear a simple whole number ratio..That is , 16:32 = 1:2.So it follows the **Law of Multiple proportions.***

*ii) In the second case 28 g dinitrogen combines with 32 g and 80 g dioxygen separately. Here , the masses of dioxygen combines (32 g and 80 g) combines with fixed mass of dinitrogen bear a simple whole number ratio..That is , 32:80 = 2:5.So it follows the **Law of Multiple proportions.***

Question 18

- *If the speed of light is 3.0×10^8 m /s, calculate the distance covered by light in 2.00 ns.*

Answer

- *Distance covered = speed * time = $(3*10^8) * (2.00 * 10^{-9})$ m = $6 * 10^{-1}$ m = 0.6m*
  =====   ===

Question 19

- *In a reaction*
  $A+B_2$ ...... > AB_2
  Identify the limiting reagent, if any, in the following reaction mixtures.

i. *300 atoms of A + 200 molecules of B*
ii. *2 mol A+3 mol B*
iii. *100 atoms of A + 100 molecules of B*
iv. *5mol A+2.5mol B*
v. *2.5mol A+5mol B*

Answer

I. *Ratio of reactants is 1:1.So 200 molecules of B is the limiting reactant.*
II. *Ratio of reactants is 1:1.So 2 moles A is the limiting reactant.*
III. *No limitating reactant.*
IV. *Ratio of reactants is 1:1.So 2.5 moles of B is the limiting reactant.*
V. *Ratio of reactants is 1:1.2.5 moles A is the limiting reactant.*

Question 20

- *Dinitrogen and dihydrogen react with each other to produce ammonia according to the following chemical equation:*
  N_2 (g) + $3H_2$ (g) .....> $2NH_3$ (g)

i. *Calculate the mass of ammonia produced if 2.00×10^3 g dinitrogen reacts with 1.00×10^3 g of dihydrogen.*
ii. *Will any of the two reactants remain unreacted?*
iii. *If yes, which one and what would be its mass?*

Answer

- *No. of moles of nitrogen = 2000/28 = 71.4286 moles*
  The ratio of reactants and products of the reaction is 1:3:2
  *Therefore, the moles of ammonia produced = 2*71.4286 moles = 142.8571 moles*
  *= 142.8571 * 17 = 2429 g*
  =====
- *Nitrogen will remain unreacted.*

- *Mass of nitrogen remain unreacted = 500 – 214.2857 = 285.7143 g*

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Question 21

- *How are 0.50 mol Na_2CO_3 and 0.50 M Na_2CO_3 different?*

Answer

- *0.50 mol Na_2CO_3 contains 0.5 moles in a given sample and 0.50 M Na_2CO_3 contains 0.5 moles in 1 Litre solution.*

Question 22

- *If 10 volumes of dihydrogen gas reacts with five volumes of dioxygen gas, how many volumes of water vapour would be produced?*

Answer

- *$2H_2 + O_2$ ......> $2H_2O$*
  Volume ratio of reactants and products are 2:1:2
  If 10 volumes of dihydrogen gas reacts with five volumes of dioxygen gas volumes of water vapour produced would be 10 volumes.

Question 23

- *Which one of the following will have the largest number of atoms?*

1) *1gAu (s)*
2) *1gNal (s)*
3) *1gLi (s)*
4) *1 g of Cl_2 (g)*

Answer

1) *No.of mole atoms 1 g Au = $\frac{\text{Mass in gram}}{\text{GAM}}$ = 1/196.97 = 0.005077 mole atoms*

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2) *No.of mole atoms 1 g NaI = $\frac{\text{Mass in gram}}{\text{GAM}}$ = 1/149.9 = 0.00667 mole atoms*

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3) *No.of mole atoms 1 g Li= $\frac{\text{Mass in gram}}{\text{GAM}}$ = 1/6.94 = 0.14409 mole atoms*

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4) *No.of mole atoms 1 g Cl_2 = $\frac{\text{Mass in gram}}{\text{GAM}}$ = 1/70.9 = 0.0141 mole atoms*

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i) ***1g Li has the largest number of atoms*** ***ii) 1g Au has the least number of atoms***

Question 24

- *Calculate the molarity of a solution of ethanol in water, in which the mole fraction of ethanol is 0.040 (assume the density of water to be one).*

Answer

$$\text{Mole fraction of ethanol} = \frac{\text{No.of moles of ethanol } (n_E)}{\text{No.of moles ethanol } (n_E) + \text{No.of moles water } (n_W)}$$

$$0.04 = \frac{n_E}{n_E + n_W}$$

$$0.04 = \frac{n_E}{n_E + 1000/18}$$ *(Density of water is 1g.1 mL water= 1g)*

$$0.04 = \frac{n_E}{n_E + 55.55}$$

$$0.04(n_E + 55.55) = n_E$$

$$0.04*n_E + 2.22 = n_E$$

$$n_E(1 – 0.04) = 2.22$$

$$n_E = 2.22/0.96 = 2.31 \text{ moles}$$

Molarity = No.of moles ethanol / volume of solution(L)

= 2.31/1 = 2.31 M

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Question 25

- *What will be the mass of one ^{12}C atom in g?*

Answer

$$\text{the mass of one } ^{12}C \text{ atom in g} = \frac{GAM}{6.022x10^{23}} = \frac{12}{6.022x10^{23}} g$$

Question 26

- *How many significant figures should be present in the answer of the following calculations (i) (02856 x 298.15 x 0.112) / 0.5785 (ii) 5x 5.364 (iii) 0.0125 + 0.7864 + 0.0215*

Answer

(i) (0.2856 x 298.15 x 0.112) / 0.5785 = 9.5369836 / 0.5785 = 16.48571 = 16.5
=====

Since 0.112 contains the least number no of significant figures which contains 3 significant figures. So the answer should contain 3 significant figures.

(ii) 5x 5.364 = 26.82 Significant figures 4.

(iii) 0.0125 + 0.7864 + 0.0215 = 0.8204 = 0.820 significant figures 3.

Question 27

- *Calculate the number of atoms in each of the following (i) 52 moles of Ar (ii) 52 u of He (iii) 52 g of He.*

Answer

(i) The number of atoms in 52 moles of Ar = 52 x $6.022x10^{23}$
(ii) The number of atoms in 52 u of He = 52/4 = 13 atoms
(iii) The number of atoms in 52 g of He = (52/4) x $6.022x10^{23}$ = 13 x $6.022x10^{23}$ atoms

Question 28

- *A welding fuel gas contains carbon and hydrogen only. Burning a small sample of it in oxygen gives 3.38 g carbon dioxide, 0.690 g of water and no other products. A volume of 10.0 L (measured at STP) of this welding gas is found to weigh 11.6 g. Calculate* ***(i) empirical formula, (ii) molar mass of the gas, and (iii) molecular formula.***

Answer

(i)

% of Carbon = 12X3.38
----------- x 100 =92.18 %
44

% of Hydrogen = 0.690 X 2
----------- x 100 = 7.67 %
18

Element	Atomic mass	% composition	% composition/Atomic mass	Value /Least Value
C	12	92.18	92.18/12 = 7.68	7.68/7.67 = 1
H	1	7.67	767 /1 = 7.67	7.67/7.67 = 1

Empirical formula = CH

ii) Mass of 22.4 L at STP = (11.6/10) x 22.4 = 25.98 g = molar mass *n = Molecular mass / Empirical formula mass = 25.98 / 13 = 2*

iii) Molecular formula = C_2H_2

Question 29

- *Calcium carbonate reacts with aqueous HCI to give* $CaCl_2$ *and* CO_2 *according to the reaction,* $CaCO_3 (s) + 2 HCI (aq) \dashrightarrow CaCl_2 (aq) + CO_2(g) + H_2O(1)$

  What mass of $CaCO_3$ *is required to react completely with 25 mL of 0.75 M HCI?*

Answer

$$0.75 = \frac{\textbf{No of moles of Hcl}}{\text{Volume in ml}} \times 1000 = \frac{\textbf{mass in gram}}{35.5 \times 25} \times 1000 = \frac{\textbf{Mass in gram x 40}}{35.5}$$

$$\text{Mass in gram} = \frac{0.75 \times 35.5}{40} = 0.6656 \text{ g}$$

$$\text{Mass of CaCO3 needed to react with 0.6656 g Hcl} = \frac{100 \times 0.6656}{71} = 0.937 \text{ g}$$

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Question 30

- *Chlorine is prepared in the laboratory by treating manganese dioxide* (MnO_2) *with aqueous hydrochloric acid according to the reaction*
  $4 HCl (aq) + MnO_2(s) > 2H_20 (1) + MnCI_2(aq) + CI_2 (g)$

  How many grams of HCI react with 5.0 g of manganese dioxide?

Answer

Molecular mass of MnO_2 *= 54.94 + 32 = 86.94 g*
Molecular mass of HCl = 36.5 g
According to the chemical equation 4HCl reacts with 86.94 g MnO_2
That is, 146 g HCl reacts with 86.94 g MnO_2

$$\text{Mass of HCl reacts with 86.94 g } MnO_2 = \frac{146 \times 5}{86.94} = 8.396 \text{ g}$$

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