Introduction :-

Chemistry is defined as the branch of science which deals with composition properties & interaction of matter.

Importance of chemistry :-

NO acts as a messenger in ~~transmiss~~ transmission of brain waves & in immune system it slows down the growth of tumor.

Chemical fertilizers like urea & $(NH_4)_2SO_4$

Preservatives like sodium Benzoate & sodium metabisulphite.

Insecticide & Pesticide → DDT & gammexane

Life saving drugs cisplatin & taxol both for treatment of Cancer.

AZT for AIDS victims.

Matter

anything which occupies space and has a volume.

- Rigid → Solid
- fluid
  - liquid
  - gas

These states can be easily interconverted by changing temp. & Pressure.

Based on the purity matter can also be divide as.

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Matter

- Pure substance
  - element
  - Compounds.
- Impure substance (mixture)
  - Homo
    - true Solution
  - Hetero
    - colloid
    - Suspension

Element :- Purest form of matter made up of only one kind of particle called atoms or molecules.

atoms	molecule
they are basic unit consisting of only subatomic particle & no further interaction with other ~~ra~~ atoms	One ~~are~~ or more atoms of same element or different & held together by strong interaction (covalent)
They may or may not exist independently.	They can exist independently.
eg. H, Cl, He	eg. H_2, Cl_2, He, CO_2

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Molecule :- Can be ~~de~~ defined for both element & ~~mo~~ ~~compo~~ compound.

If by one type of atom => elemental molecule.
H_2, He, S_8, O_3

If by different type of atoms => molecule of a compound.

~~Elemental molecule~~
Based on the no. of atoms.

-> Monoatomic => one atom
He, Ne, Ar ...

-> Diatomic => 2 atoms.
CO, O_2, HCl

-> Polyatomic => more than 2 atoms.
O_3, SO_2, CH_4, S_8

Law of chemical combinations:

1. Law of conservation of Mass
2. Law of Definite proportion.
3. Law of Multiple proportion.
4. Gay-Lussac's law of gases
5. Avogadro's Law

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1. Law of conservation of mass

No loss of mass will occur during a chemical reaction.

eg. $CaCO_3 \longrightarrow CaO + CO_2$
50g 28g 22g.

exception Nuclear Reactions.

2. Law of constant or Definite proportion

Percentage of an element in any given compound is constant irrespective of the source from which it is obtained.

3. Law of Multiple proportion

Two parts of the Law

Elements can combine with each other to form more than one compound.

eg. 1. CO_2, CO
2. NO, NO_2, N_2O, N_2O_5, N_2O_3
3. C_2H_3Br, C_2H_5Br

On ~~examining~~ analysising a compound formed from 2 ~~atoms~~ elements.

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then on fixing one of the element another can show a a simple whole no. ratio.

4. **Guy lussac's Law of gaseous volume:-**

Applicable for gases which are involved in a chemical reaction

It's compulsory to have atleast 2 gaseous terms ~~in it. the~~ involved in the chemical Reaction.

$A_{(g)} + B_{(s)} \longrightarrow C_{(l)}$ ✗

$A_{(g)} + B_{(s)} \longrightarrow C_{(g)}$ ✓

$3A_{(g)} + 2B_{(s)} \longrightarrow 4C_{(g)}$

⇒ 3vol A forms 4vol C

$3A_{(g)} + 4B_{(g)} \longrightarrow 2C_{(g)}$

⇒ 3vol A reacts with 4 vol B to form 2 vol C

$5A_{(g)} + 3B_{(l)} \longrightarrow 6C_{(l)} + 4D_{(g)}$

?

1. When **8.4g** of $NaHCO_3$ is added to a solution of CH_3COOH weighing **20g**. It is Observed that **4.4g** of CO_2 is released into atmosphere and residue is left behind. Calculate the mass of residue by applying **law of conservation of mass.**

2. **2.75g** of **cupric oxide** was reduced by heating in a current of **hydrogen** and weight of **copper** that remained was **2.196g**. Another experiment , **2.358g** of **copper** was dissolved in **nitric acid** and resulting **copper nitrate** converted into **cupric oxide** by ignition. The weight of **cupric oxide** formed was **2.952g**. Show that these results illustrate **law of constant composition**.

3. **Sulphur** and **oxygen** are known to form two compounds. The sulphur content in one of these is **51%** while in the other is **41%.** Show that this data is in agreement with the **law of multiple proportions**.

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5. ## Avogadro's law:-

Equal vol. of all gases under similar conditions of temp & pressure contain equal no. of molecules.

Dalton's Atomic Theory:-

Matter is made up of atom.

Atom is extremely small, indivisible particle.

Atoms of an element are identical

Atoms of different elements are different

Atom combine with each other in a fixed, simple whole no. ratio to form a compound.

Atoms can neither be created nor be destroyed.

Atomic mass:-

Standard is ^{12}C isotope.
all the atomic mass are obtained

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in relation to this C-isotope.

Atomic mass ⇒ a mass exactly equal to $\frac{1}{12}^{th}$ of mass of ^{12}C-atom

seen for one entity.

Unit is amu or u or unified mass.

As atoms has many isotopes the atomic mass is usually taken as the average of all the isotopes.

⇓

Average atomic mass.

$$AAM = \frac{\sum R.A \times \text{Isotopic no.}}{\sum R.A.}$$

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4. Given that the abundances of isotopes ^{54}Fe, ^{56}Fe, ^{57}Fe are **5%**, **90%** and **5%** respectively. The atomic mass of ***Fe*** is:

(a) 55.85 u (b) 55.95 u (c) 55.75 u (d) 56.05 u

5. Boron has two isotopes **boron-10** and **boron-11** whose percentage abundances are **19.6%** and **80.4%** respectively. The average atomic mass of **boron** is :

(a) 10.5 *amu* (b) 10.804 *amu* (c) 11.01 *amu* (d) 10.79 *amu*

6. Carbon occurs in nature as a mixture of **carbon-12** and **carbon-13**. The average atomic mass of carbon is **12.011 *amu*** . The percentage abundance of **carbon-12** is :

Base Physical Quantity	Symbol for Quantity	Name of 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

The definition of SI base units are given in table below.

Definitions of SI Base Units

Unit of length	metre	The metre is the length of the path travelled by light vacuum during a time interval of 1/299 792 458 of a secon
Unit of mass	kilogram	The kilogram is the unit of mass; it is equal to the mass the international prototype of the kilogram.
Unit of time	second	The second is the duration of 9 192 631 770 periods of th radiation corresponding to the transition between the tw hyperfine levels of the ground state of the caesium-13 atom.
Unit of electric current	ampere	The ampere is that constant current which, if maintained i two straight parallel conductors of infinite length, c negligible circular cross-section, and placed 1 metre apar in vacuum, would produce between these conductors force equal to 2×10^{-7} newton per metre of length.
Unit of thermodynamic temperature	kelvin	The kelvin, unit of thermodynamic temperature, is the fraction 1/273.16 of the thermodynamic temperature of the triple point of water
Unit of amount of substance	mole	1. The mole is the amount of substance of a system which contains as many elementary entities as there are atoms in 0.012 kilogram of carbon-12; its symbol is "mol". 2. When the mole is used, the elementary entities must be specified any may be atoms, molecules, ions, electrons, other particles, or specified groups of such particles.
Unit of luminous intensity	candela	The candela is the luminous intensity, in a given direction, of a source that emits monochromatic radiation of frequency 540×10^{12} hertz and that has a radiant intensity in that direction of 1/683 watt per steradian.

Prefixes used in the SI system

Multiple	Prefix	Symbol
10^{-24}	yocto	y
10^{-21}	zepto	z
10^{-18}	atto	a
10^{-15}	femto	f
10^{-12}	pico	p
10^{-9}	nano	n
10^{-6}	micro	μ
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 given below are some relations between various units and some useful conversion factor which can be useful in solving numerical problems.

Table : Relations between various units

Units of Length	Units of Volume
1 ft = 12 in	1 m^3 = 10^3 L
1 yd = 3 ft	1 dm^3 = 1 L
1 mile = 5280 ft	1 cm^3 = 10^{-3} L
1 in = 2.54 cm	1 ft^3 = 28.32 L
1 m = 39.37 in	**Units of Mass**
1 mile = 1.609 km	1 kg = 10^3 g
1 Å = 10^{-10} m	1 mg = 10^{-6} kg
1 nm = 10^{-9} m	1 metric tonne = 10^3 kg
1 pm = 10^{-12} m	1 lb = 453.6 g

SOME USEFUL CONVERSION FACTORS

Common Units of Mass and Weight

1 pound = 453.59 grams

1 pound = 453.59 grams = 0.45359 kilogram

1 kilogram = 1000 grams = 2.205 pounds

1 gram = 10 decigrams = 100 centigrams = 1000 milligrams

1 gram = 6.022×10^{23} atomic mass units or u,

1 atomic mass unit = 1.6606×10^{-24} gram

1 metric tonne = 1000 kilograms = 2205 pounds

Common Units of Volume

1 quart = 0.9463 litre, 1 litre = 1.056 quarts

1 litre = 1 cubic decimetre = 1000 cubic centimetres = 0.001 cubic metre

1 millilitre = 1 cubic centimetre = 0.001 litre = 1.056×10^{-3} quart

1 cubic foot = 28.316 litres = 29.902 quarts = 7.475 gallons

Common Units of Energy

1 joule = 1×10^{7} ergs

1 calorie* = 4.184 joules

= 4.184×10^{7} ergs = 4.129×10^{-2} litre-atmospheres

= 2.612×10^{19} electron volts

1 ergs = 1×10^{-7} joule = 2.3901×10^{-8} calorie

1 electron volt = 1.6022×10^{-19} joule = 1.6022×10^{-12} erg = 96.487 kJ/mol

1 litre-atmosphere = 24.217 calories = 101.32 joules = 1.0132×10^{9} ergs

1 British thermal unit = 1055.06 joules = 1.05506×10^{10} ergs = 252.2 calories

Common Units of Length

1 inch = 2.54 centimetres (exactly)

1 mile = 5280 feet = 1.609 kilometers

1 yard = 36 inches = 0.9144 metre

1 metre = 100 centimetres = 39.37 inches = 3.281 feet = 1.094 yards

1 kilometer = 1000 metres = 1094 yards = 0.6215 mile

1 Angstrom = 1.0×10^{-8} centimetres = 0.10 nanometre

= 1.0×10^{-10} metre = 3.937×10^{-9} inch

Common Units of Force** and Pressure

1 atmosphere = 760 millimetres of mercury

= 1.013×10^{5} Pascals

= 14.70 pounds per square inch

1 bar = 10^{5} Pascals

1 torr = 1 millimetre of mercury

1 pascal = 1 $kg/m.s^2$ = 1 N/m^2

Temperature

SI Base Unit : Kelvin (K)

0 K = –273.15°C

K = °C + 273.15°

°F = 1.8(°C) + 32°

$$°C = \frac{F - 32°}{1.8}$$

Scientific Notation.

$N \times 10^{x}$

=> 600000000000

6.0×10^{11}

0.00000000632

=> 6.32×10^{-9}

Addition & subtraction

10 power should be same

$3.425 \times 10^{4} + 3.425 \times 10^{3}$

$34.25 \times 10^{3} + 3.425 \times 10^{3}$

37.675×10^{3}

3.7675×10^{3}

Multiplication & division.

$10^x \times 10^y \Rightarrow 10^{x+y}$

$\frac{10^x}{10^y} \Rightarrow 10^{x-y}$

Significant figures

Precision: Closeness of various measurements.

Accuracy: agreement with the true value.

Rules of Significant figures

1. All non zero digits are significant.
   75, 68, 786
2. Zero on left of non zero digit is not significant
   0.068   0.00003
3. Zero inbetween is always significant.
   608, 305
4. Zero after non zero is significant
   600   6.00
5. Exact no. like 5 pens => infinite SF but 5 gram => 1 S.F.
6. Scientific notation only 'N' is considered.
   4.5×10^3 4.50×10^3 4.500×10^4

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Rounding off :-

1. Right most digit more than 5
   -> Preceding no. is ↑ with 1.
2. Right most digit less than 5
   => Preceding no. is not changed.
3. Right most digit is 5

if preceding no. is even then it is not changed.

If preceding no. is odd then it is ↑ by 1.

Significant figures with Addition & subtraction.

~~Highest~~ minimum no. after decimal ~~pa~~ point

eg. 33.6342
0.003
0.02
0.1   ?

For Multiplication & Division.
Same no. of significant digits as the no. with least no of significant digit

20 × 3.33 => 66.6

⇓ infinite ⇓ 3 SF ⇓ 3 SF.

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Mole concept:-

What is mole?

It is chemistry dozen.

1 dozen => 12

similarily 1 mole => 6.022×10^{23} => N_A or N_0

So if I have

1 mol pens => N_A pens.

deals with atoms, molecules & ions.

Practical difficulties

→ Very small can't see them

→ large no. to count.

=> It should be equated with a constant value for easier studies.

=> 1mol linked with its GMM/GAM/GFM

1mol = GMM / GAM / GFM

GMM ⇓ For molecules / GAM ⇓ Atoms / GFM ⇓ ionic compound

For gases
1 mole = 22.4 L at NTP
(1 bar press + 273 K)

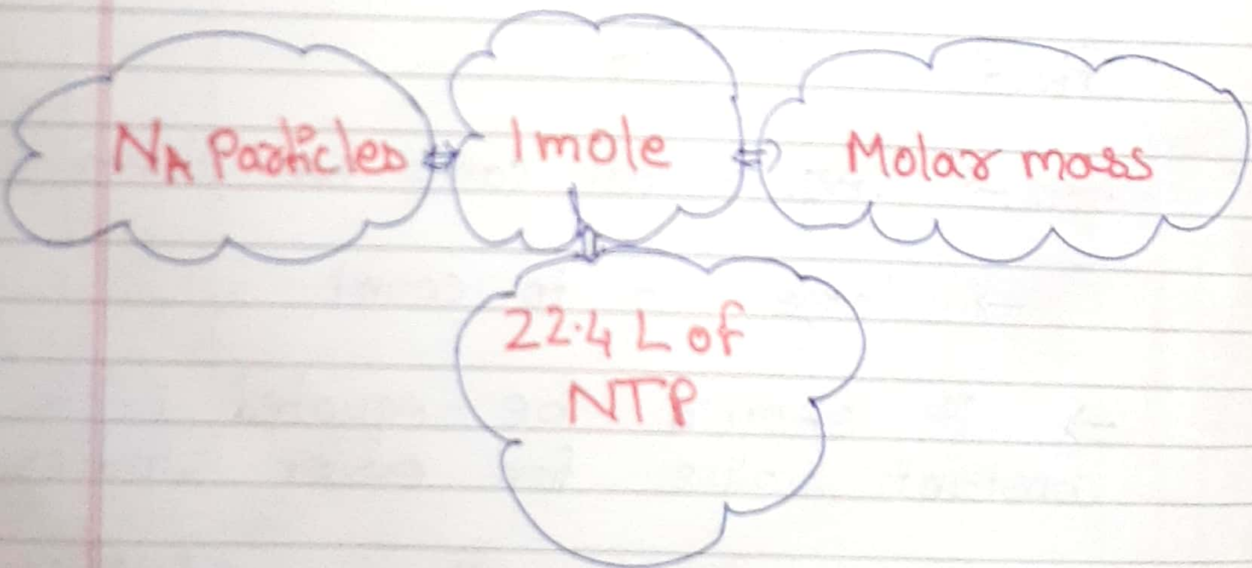
1 mole is also called as:

1 gram molecule — For molecules

1 gram atom — For atoms

1 gram Formual — For ionic compounds

From above the final conclusion can be drawn as.



Type of Questions which can be asked from mole concept.

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Type I :- Direct relations

$$n = \frac{wt}{mol \cdot wt} \qquad n = \frac{\text{no. of particles}}{6.022 \times 10^{23}}$$

$$n = \frac{\text{Vol. of NTP}_{(g)}}{22.4L}$$

Type II :- Interconversions

$$\frac{wt}{\text{molecular wt}} = \frac{\text{no. of particles}}{6.022 \times 10^{23}}$$

$$\frac{wt}{mol \cdot wt} = \frac{\text{vol. of gas at NTP}}{22.4(L)}$$

$$\frac{\text{no. of particle}}{6.022 \times 10^{23}} = \frac{\text{vol. of gas at NTP}}{22.4\ (L)}$$

Type III :- (a) Finding no. of atoms in a molecule or ionic radicals

Step 1
no. of moles of molecule or ionic radicals is calculated.

Step 2 no. of corresponding atoms or total atoms are calculated.

(b) Subatomic particles

Step 1
no. of moles of molecules or ionic

radicals are calculated.

step 2

Finding total no. of subatomic particles.

For protons = Z_{Total}

For neutrons = $(A-Z)_{Total}$

For electron

cations = Z_{Total} - charge.

anions = Z_{Total} + charge

uncharged = Z_{Total}.

Type IV :- Relation between gram & amu (avogram)

~~1 gram = 1.67 × 10^-24~~

$1 \text{ amu} = 1.67 \times 10^{-24}$ gram

$= \frac{1}{6} \times 10^{-23}$ gram

Type V comparision

step I Find no. of mol of given value

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Step II ~~we can~~ Find the no. of moles of given atom in the particle

Step III Find the no. of mol. of another particle w.r.t. the atoms obtained in step II

Type VI :- Finding charge (only for charged particles)

1 mol charge = 96500 C

Try Yourself

41. Calculate mass of one atom of calcium in gram.
42. Calculate mass of one molecule of sulphur dioxide (SO_2) in gram.
43. Calculate number of atoms in 0.5 mole of oxygen atoms.
44. Calculate number of atoms in 0.2 mole of oxygen atoms.

PERCENTAGE COMPOSITION

Percentage composition of the compound is the relative mass of the each of the constituent element in 100 parts of it.

$$\text{Mass \% of an element} = \frac{\text{Mass of that element in one mole of the compound}}{\text{Molar mass of the compound}} \times 100$$

Let us take an example of water (H_2O), it contains hydrogen and oxygen, the percentage composition of both these elements can be calculated as follows :

Molar mass of water = 18.02 g

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

$$\text{Mass \% of oxygen} = \frac{16.00}{18.02} \times 100 = 88.79\%$$

One can check the purity of a given sample by analysing percentage composition.

Example 30 : *Find the percentage of calcium in calcium carbonate ($CaCO_3$).*

Solution : Molecular mass of $CaCO_3$ = 40 + 12 + (16 × 3) = 100 u

Atomic mass of Ca = 40

$\therefore$ Percentage of Ca $= \frac{40}{100} \times 100 = 40\%$

Try Yourself

45. Calculate percentage of sulphur in sulphuric acid (H_2SO_4).
46. Calculate percentage of carbon in ethanol (C_2H_5OH).

Empirical Formula and Molecular Formula

An **empirical formula** of a compound may be defined as the formula which gives the simplest whole number ratio of atoms of the various elements present in the molecule of the compound.

The **molecular formula** of a compound may be defined as the formula which gives the actual number of atoms of various elements present in the molecule of the compound.

For example, the molecular formula of glucose is $C_6H_{12}O_6$ and empirical formula is CH_2O (which shows that C, H and O are present in the simplest ratio of 1 : 2 : 1. Relation between the two formulae :

Molecular formula = Empirical formula × n

Where, n = 1, 2, 3 ... $n = \frac{\text{Molecular formula mass}}{\text{Empirical formula mass}}$

Steps for Writing Empirical and Molecular Formula

If the mass percent of various elements present in a compound is known, its empirical formula can be determined. Molecular formula can be obtained if the molar mass is also known. This will become clear with following example :

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?

Step 1. Conversion of mass percent to grams.

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

Step 2. Convert into number moles of each element

Divide the masses obtained above by respective atomic masses of various elements.

Moles of hydrogen $= \frac{4.07 \text{ g}}{1.008 \text{ g}} = 4.04$

Moles of carbon $= \frac{24.27 \text{ g}}{12.01 \text{ g}} = 2.021$

Moles of chlorine $= \frac{71.65 \text{ g}}{35.453 \text{ g}} = 2.021$

Step 3. Divide the mole value obtained above by the smallest number

Since, 2.021 is smallest value, division by it gives a ratio of 2 : 1 : 1 for H : C : Cl .

In case the ratios are not whole numbers, then they may be converted into whole number by multipl by the suitable coefficient.

Step 4. Write empirical formula by mentioning the number after writing the symbols of respec elements.

CH_2Cl is, thus, the empirical formula of the above compound.

Step 5. Writing molecular formula

(a) Determine empirical formula mass.

Add the atomic masses of various atoms present in the empirical formula.

For CH_2Cl, empirical formula mass is

$12.01 + 2 \times 1.008 + 35.453 = 49.48$ g

(b) Divide molar mass by empirical formula mass.

$$\frac{\text{Molar mass}}{\text{Empirical formula mass}} = \frac{98.96 \text{ g}}{49.48 \text{ g}} = 2 = (n)$$

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

Empirical formula = CH_2Cl, n = 2.

Hence, molecular formula is $C_2H_4Cl_2$.

Example 31 : *What is the simplest formula of the compound which has the following percentage compositior Carbon 80%, Hydrogen 20%? If the molecular mass is 30, calculate its molecular formula.*

Solution : Calculation of empirical formula

Element	% age	Atomic mass	Constent in sample	No of moles	Sample ratio	Simplest whole no. ratio
C	80	12	80 g	$\frac{80}{12} = 6.66$	$\frac{6.66}{6.66} = 1$	1
H	20	1	20 g	$\frac{20}{1} = 20$	$\frac{20}{6.66} = 3$	3

∴ Empirical formula is CH_3

Calculation of molecular formula

Empirical formula mass = $12 \times 1 + 1 \times 3 = 15$

$$n = \frac{\text{Molecular mass}}{\text{Empirical formula mass}} = \frac{30}{15} = 2$$

Molecular formula = Empirical formula × 2 = $CH_3 \times 2 = C_2H_6$

Try Yourself

47. A substance, on analysis, gave the following percentage composition : Na = 43.4%, C = 11.3%, O = 45.3%. Calculate the empirical formula. (Na = 23, C = 12, O = 16)

48. An organic compound on analysis gave the following data : C = 57.82%, H = 3.6% and 38.58% is oxygen. Find its empirical formula.

EXERCISE 1.2

[Molecular Mass, Empirical and Molecular formula, Mole Concept]

16. Choose the correct option regarding atomic mass unit

(1) $1\,u = \frac{1}{12}$ mass of 1 atom of C^{12}
(2) $1\,u = \frac{1}{16}$ mass of 1 atom of O^{16}
(3) $1\,u = \frac{1}{44}$ mass of 1 molecule of CO_2
(4) $1\,u = \frac{1}{6}$ mass of 1 atom of C^{12}

17. The simplest formula of a compound containing 50% of element X (Atomic weight = 10) and 50% of element Y (atomic weight = 20) is

(1) XY
(2) X_2Y
(3) XY_2
(4) X_2Y_3

18. Number of carbon atoms present in 22 g CO_2 is

(1) 6.02×10^{23}
(2) 3.01×10^{23}
(3) 3.01×10^{-23}
(4) 6.02×10^{-23}

19. Volume of oxygen at NTP, required to completely burn 1 kg of coal (100% carbon) is

(1) 22.4 L
(2) 1.86×10^3 L
(3) 22.4×10^3 L
(4) 1000 L

20. Which of the following possess highest mass?

(1) 0.2 mol of CO_2 gas
(2) 2.24 L SO_2
(3) 3.01×10^{23} molecules of water
(4) 20 g calcium

21. How many water molecules are present in one mL of water (d = 1 gm/ml)?

(1) 5.55×10^{21}
(2) 6.023×10^{23}
(3) 3.346×10^{22}
(4) 6.023×10^{22}

22. Number of electrons present in 1.6 g of methane is

(1) 6.02×10^{24}
(2) 6.02×10^{23}
(3) 6.02×10^{22}
(4) 6.02×10^{21}

23. How many times an atom of sulphur is heavier than an atom of carbon?

(1) 32
(2) 12
(3) $\frac{8}{3}$
(4) $\frac{12}{32}$

24. The total mass of 1 mole of electrons is? (Mass of $1e^- = 9.1 \times 10^{-28}$ g)

(1) 9.1×10^{-28} g (2) 1.008 mg

(3) 0.55 mg (4) 9.1×10^{-27} g

25. How many moles of $Al_2(SO_4)_3$ would be present in 50 g of the substance? (Given at volume of Al = 27, S = 32, O = 16)

(1) 0.083 mole (2) 0.952 mole

(3) 0.481 mole (4) 0.14 mole

26. How many moles of electrons weigh one kilogram?

(1) 6.022×10^{23} (2) $\frac{1}{9.1} \times 10^{31}$

(3) $\frac{6.022 \times 10^{54}}{9.1}$ (4) $\frac{10^8}{9.1 \times 6.022}$

27. 0.0833 mol of carbohydrate of empirical formula CH_2O contain 1 g of hydrogen. The molecular formula of carbohydrate is

(1) $C_5H_{10}O_5$ (2) $C_6H_{12}O_6$

(3) $C_{12}H_{22}O_{11}$ (4) $C_3H_4O_3$

28. How many moles of magnesium phosphate, $Mg_3(PO_4)_2$ will contain 0.25 mole of oxygen atoms?

(1) 3.125×10^{-2} (2) 1.25×10^{-2}

(3) 2.5×10^{-2} (4) 0.125

29. Which has maximum number of atoms?

(1) 10.8 g of Ag(108) (2) 2.4 g of C(12)

(3) 5.6 g of Fe(56) (4) 54g of Al(27)

30. In a compound C, H and N are present in the ratio of 9 : 1 : 3.5 by weight. Molecular mass of compound is 108. Molecular formula of the compound is

(1) $C_2H_6N_2$ (2) C_3H_4N

(3) $C_6H_8N_2$ (4) $C_9H_{12}N_3$