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Ch - 1 - Some Basic Concepts Of Chemistry

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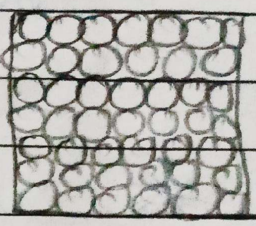
* **Matter** :- A substance that having mass and occupy space.
  → solid , liquid, gas

* **Characteristics properties of matter** :-

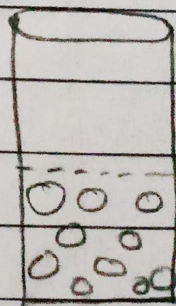
① **Solid** :-



→ They have fix shape and volume.
→ Intermolecular force is very high.
→ They are rigid.
→ They have no space.
→ Atom can't move from it's position.
→ It can only vibrate on their fix position
→ Highly incompressible
→ density is very high.

exception - sponge

② **Liquid** :-



→ They have fix volume only.
→ Intermolecular force is less compared to solid.
→ They are not rigid.
→ They have space between them.
→ It can move from its position.
→ Density is less compare to solid.
→ They are less compressible.

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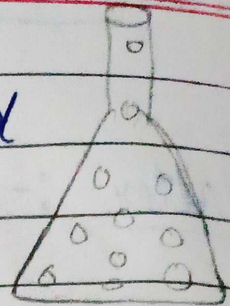
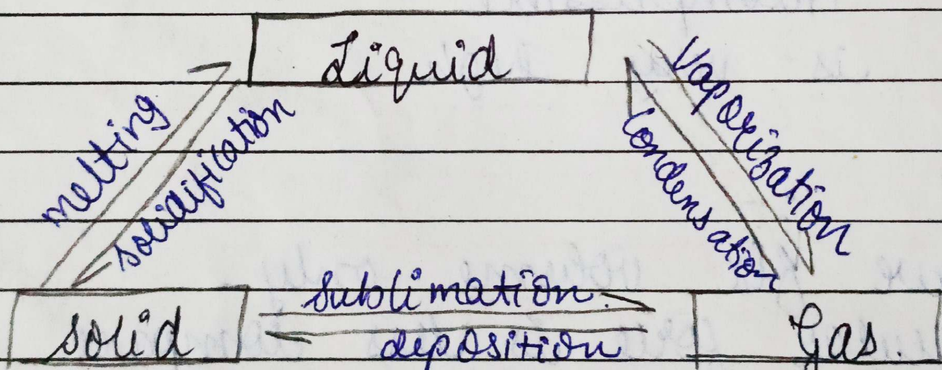
(3) Gas :-



- They don't have fix shape and volume.
- Intermolecular force is very low.
- Not rigid
- They have large space between them.
- It can move from its position.
- Density is very low.
- They are highly compressible

* Conversion :-

Solid $\underset{\text{cool}}{\overset{\text{Heat}}{\rightleftharpoons}}$ Liquid $\underset{\text{cool}}{\overset{\text{Heat}}{\rightleftharpoons}}$ Gas



→ $H_2O + O \rightarrow H_2O_2$
(Hydrogen peroxide)

* Classification of Matter

- Matter
  - Mixtures
    - Homogeneous Mixture
    - Heterogeneous Mixture
  - Pure Substances
    - Elements
      - Atom
      - Molecules
    - Compounds

* Matter :-

- Mixtures :- when 2 or more pure substances are mixed together in any definit ratio is called mixtures.

(1) Homogeneous mixture :-

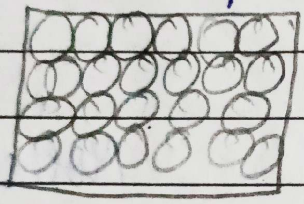
→ When 2 or more ~~mixture~~ substance are completely mixed with each other.
eg. :- sugar solution.

(2) Heterogeneous mixture :-

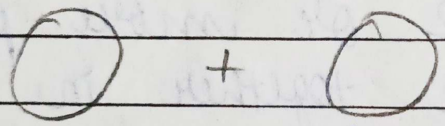
→ When 2 or more substance are not completely mixed with each other.

Matter
- Pure Substance
  - Elements
    - Atom
    - Molecules
  - Compound

1) Pure Substance :- A substance contain same constituent particle in which present.

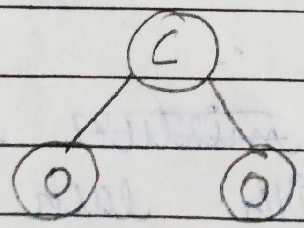


2) Molecules :- ex. :- O_2



Oxygen + Oxygen Atom

3) CO_2



* Compounds :-

→ When two or more different atom of elements are combined together in definite ratio is called compound.

eg :- glucose = $C_6H_{12}O_6$

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$C = 12$, $H = 1$, $O = 16$

$= (6 \times 12) + (12 \times 1) + (6 \times 16)$

$\boxed{180}$

Na
$p^+ = 11$
$e^- = 11$
$n^\circ = 12$

* Properties of Matter
→ 1) Physical properties
2) chemical properties

(1) Physical properties :- A properties of matter that can be observed without changing their identity or composition are called physical properties.

eg :- Melting point, boiling point, colour, odour, density.

(2) Chemical properties :- A properties of substance that can be observed by chemical change is called chemical properties.

eg :- Acidity, Basicity, toxicity, combustion, pH, etc.

* SI Unit System :-

Quantity	Units
1) Mass (m)	Kilogram (kg)
2) Length (l)	metre (m)
3) time (t)	Second (s)
4) temperature (T)	kelvin (K)
5) Current (I)	ampere (A)
6) Amount of substance (n)	mole (mol)
7) Luminous intensity (Iv)	candela (cd)

(d) deci	→ 10^{-1}		deca	→ 10^{1}	(da)
(c) centi	→ 10^{-2}		hecto	→ 10^{2}	(h)
(m) milli	→ 10^{-3}		kilo	→ 10^{3}	(K)
(μ) micro	→ 10^{-6}		mega	→ 10^{6}	(M)
(n) nano	→ 10^{-9}		giga	→ 10^{9}	(G)
(P) pico	→ 10^{-12}		tera	→ 10^{12}	(T)
(f) Femto	→ 10^{-15}		peta	→ 10^{15}	(P)
(a) atto	→ 10^{-18}		esca	→ 10^{18}	(E)
(z) zepto	→ 10^{-21}		zeta	→ 10^{21}	(Z)
(y) yocto	→ 10^{-24}		yotta	→ 10^{24}	(Y)

* Mass (m)
→ The amount of matter present in it is called mass.
→ Mass is constant everywhere.
→ SI Unit → Kg or gm.

* Weight :- The external force, gravity is acting on the body.

* Volume :- space occupied by substance is called volume.
→ Volume → m^3 or cm^3

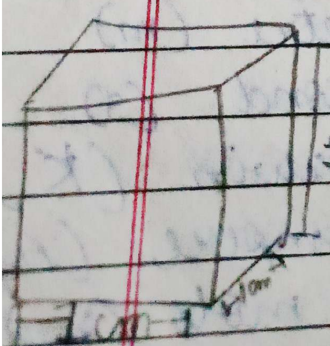
ex :- $1 ml = 1 cm^3$

$10 cm = 1 dm$

$1000 cm^3 = 1 dm^3$

$1 l = 1000 ml$

$1 l = 1 dm^3$



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Temperature (T)

→ It is degree of hotness or coldness.

→ Unit → degree celsius (°C)

→ degree Fahreinheit (°F)

→ kelvin (K)

$$\boxed{\begin{aligned} K &= °C + 273.15 \\ °F &= \frac{9}{5}(°C) + 32 \end{aligned}}$$

1) ex :- 40°C = ______ °F

$°F = \frac{9}{\cancel{5}}(\overset{8}{\cancel{40}}) + 32$

$\boxed{°F = 104}$

2) 30°C = $\underline{86}$ °F

3) 103°F = ______ °C

$°F = \frac{9}{5}(°C) + 32$

$°C = (°F - 32) \times \frac{5}{9}$

$= (103 - 32) \times \frac{5}{9}$

$= 71 \times \frac{5}{9}$

$= 39.4\ °C$

4) 230°F = $\underline{383.15}$ K.

mole = $\frac{\text{given mass of subs.}}{\text{molar mass of subs.}}$

$= \frac{g}{mol}$

* Derived Unit :-

1) Density :- It is given mass of substance per unit volume is density.

$$\text{Density} = \frac{\text{mass}}{\text{volume}} = \frac{Kg}{m^3} = \frac{gm}{cm^3}$$

2) $$\text{velocity} = \frac{\text{distance}}{\text{time}} = \frac{m}{s}$$

3) $$\text{Acceleration} = \frac{\text{velocity}}{\text{time}} = \frac{m/s}{s} = \frac{m}{s^2}$$

4) $F = ma$

$= Kg \times \frac{m}{s^2}$ Or N

* Scientific Notation :-

3×10^8 m/s

⇒ Avagadro no. :- 6.022×10^{23}

→ 322.342×10

→ 32234.2×10^{-2}

→ 322.342×10

→ 0.322342×10^3

$\xleftarrow{10^+}$ $\xrightarrow{10^-}$

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ex:- 45.32×10^{-2}
$= 0.04532 \times 10^{-2+3}$
$= 0.04532 \times 10^{1}$
$= 0.4532 \times 10^{1-1}$
$= 0.4532 \times 1$

ex :- $2.323 \times 10^{-3} + 3.22 \times 10^{-2}$
$= 2.323 \times 10^{-3+1} + 3.22 \times 10^{-2}$
$= 0.232 \times 10^{-2} + 3.22 \times 10^{-2}$
$= (0.232 + 3.22) \times 10^{-2}$
$= 3.452 \times 10^{-2}$

* Significant Figures :-

1) All non-zero digits are significant.
eg. 285 → 3 significant figure.

2) Zeros preceding to first non-zero digit are not significant.
eg. → 0.03 → one significant figure.

3) Zeros between two non-zero digits are significant.
eg. :- 2.005 → 4 significant figure.

4) Zeros at the end or right of a number are significant, provided they are on the right side of the decimal point.
eg. :- 0.200 → 3 significant fig.

5) Counting the numbers of object, for eg, 2 balls have infinite significant figures.

* Rounding off numbers :-

① Addition and Subtraction.

→ In addition or subtraction, select minimum decimal point as a significant value.

② Multiplication and Division.

(i) If the rightmost digit is more than 5, then remove it and increase its preceding value by one.

eg. :- 23.568 = 23.57

(ii) If the rightmost digit is less than 5, then remove it and don't change it's preceding value.

eg. :- 5.444 = 5.44

(iii) If the rightmost digit is 5, remove it, if it is an even number does not change it but if it is an odd number, increase by one.

eg. :- 6.35 → 6.4

6.25 → 6.2

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⇒ In the case of multiplication or division select minimum significant value as final significant value.

* **Dimensional Analysis :-**

⇒ To convert given unit into another unit.

(1) Unit factor method or Factor label method.

eg. (i) 6 ft = ____ cm.

1 ft = 12 inch.

$\frac{1\,ft}{12\,inch} = \frac{12\,inch}{1\,ft}$

1 inch = 2.54 cm

$\frac{1\,inch}{2.54\,cm} = \frac{2.54\,cm}{1\,inch}$

$6\,\cancel{ft} \times \frac{12\,\cancel{inch}}{1\,\cancel{ft}} \times \frac{2.54\,cm}{1\,\cancel{inch}}$

$= 6 \times 12 \times 2.54\ cm$

$= 182.88\ cm.$

(ii) 2 days = ____ sec

1 day = 24 hr

$\frac{1\,day}{24\,hr} = \frac{24\,hr}{1\,day}$

1 hr = 3600 sec

$\frac{1\,hr}{3600\,sec} = \frac{3600\,sec}{1\,hr}$

$2\,\cancel{days} \times \frac{24\,\cancel{hr}}{1\,\cancel{day}} \times \frac{3600\,sec}{1\,\cancel{hr}}$

$= 2 \times 24 \times 3600\ sec$

$= 172800\ sec.$

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(3) $5 \text{ lit} = \_\_\_\_ m^3$

$1 \text{ lit} = 1000 \, cm^3$

$\frac{1 \, l}{1000 cm^3} = \frac{1000 \, cm^3}{1 \, l}$

$1 m = 100 cm$

$1 m^3 = (100 \, cm)^3$

$1 m^3 = 10^6 cm^3$

$\frac{1 m^3}{10^6 cm^3} = \frac{10^6 cm^3}{1 m^3}$

$\Rightarrow 5 \cancel{l} \times \frac{1000 \cancel{cm^3}}{\cancel{1 \, l}} \times \frac{1 m^3}{\cancel{10^6}^{\,3} \cancel{cm^3}}$

$\Rightarrow 5 \times 10^{-3} m^3.$

(4) $3.2 \frac{kg}{m^3} = \_\_\_\_ g/cm^3$

$1 kg = 1000 g$

$\frac{1 Kg}{1000 g} = \frac{1000 g}{1 Kg}$

$1 m = 100 cm$

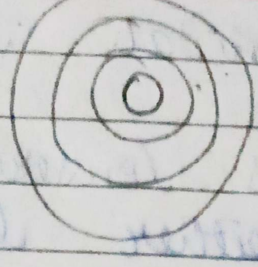
$1 m^3 = 10^6 cm^3$

$\frac{1 m^3}{10^6 cm^3} = \frac{10^6 cm^3}{1 m^3}$

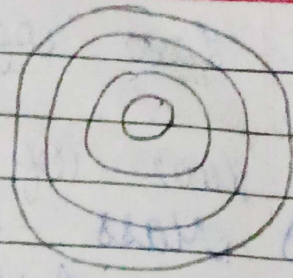
$3.2 \frac{\cancel{kg}}{\cancel{m^3}} \times \frac{1000 g}{\cancel{1 Kg}} \times \frac{\cancel{1 m^3}}{\cancel{10^6}^{\,3} cm^3}$

$= 3.2 \times 10^{-3} \, g/cm^3$

(A)
Accurate ×
Precise ×



(B)
not accurate
precise



(C)
Accurate and
precise

(5) 100 miles per hour = ____ $\frac{m}{s}$

1 miles = 1.6 Km
1 Km = 1000 m
1 hr = 60 min
1 min = 60 sec

$\Rightarrow 100 \frac{\cancel{miles}}{\cancel{hr}} \times \frac{1.6 \cancel{Km}}{1 \cancel{mile}} \times \frac{1000 m}{1 \cancel{Km}} \times \frac{1 \cancel{hour}}{60 \cancel{min}} \times \frac{1 \cancel{min}}{60 sec}$

$= \frac{100 \times 1.6 \times 1000}{60 \times 60} \frac{m}{s}$

$= 0.044 \times 10^3$

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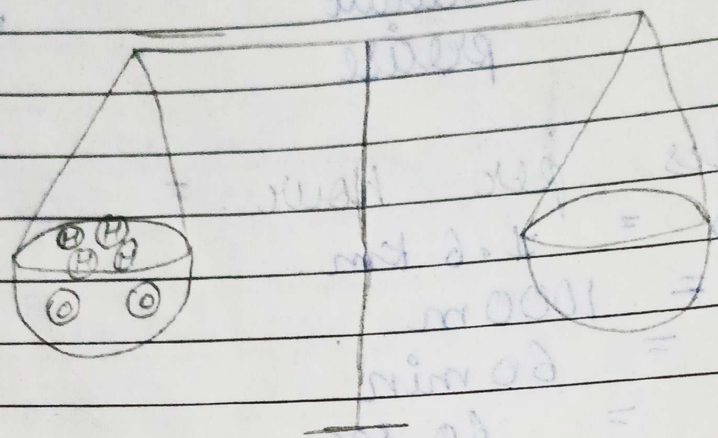
* Law of chemical combination

1) Law of mass conservation

⟹ Mass can neither be created nor be destroyed in chemical reaction.



$$2H_2 + O_2 \rightarrow 2H_2O$$

⟹ Mass of Reactant (Before reaction) = Mass of Product (After reaction)

2) Law of definite proportion

⟹ 'A given compound always contains exactly the same no. of elements in simple ratio.'

→ wood (C) + O_2 → CO_2.

12g + 32g → 44g

→ $\text{mass \%} = \dfrac{\text{mass of substance}}{\text{mass of solution or total mass}} \times 100$

$\% C = \frac{12}{44} \times 100$

$= 27.27\%$

$\% O_2 = \frac{32}{44} \times 100$

$= 72.72\%$.

Eg. :- $CuCO_3$

Synthetically, $CuO + CO_2 \rightarrow CuCO_3$

Naturally, $CuCO_3$

③ Law Of Multiple proportions :-

If 2 elements combine to form more than two compound, mass of one element combine with fix mass of other element. are in ratio of small whole number.

fix $H_2 + \frac{1}{2}O_2 \rightarrow H_2O$.

$2g \quad 16g \rightarrow 18g$.

$H_2 + O_2 \rightarrow H_2O$.

$2g \quad 32g$ Hydrogen peroxide

$34\ g$.

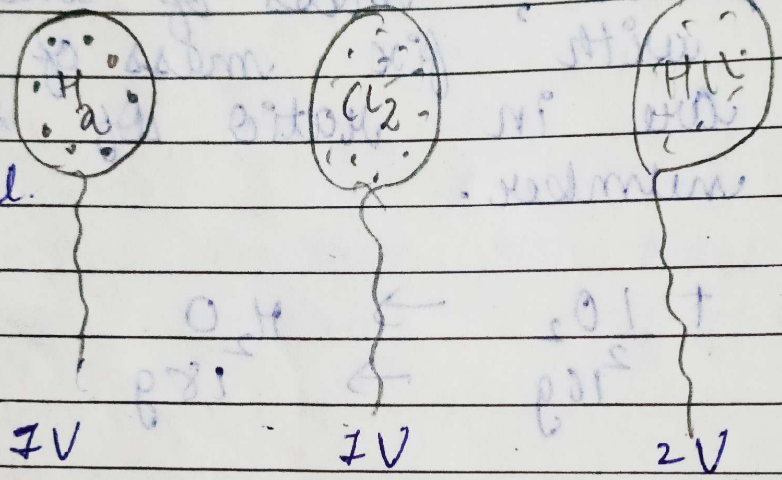
$16 : 32$.

$\boxed{1:2}$

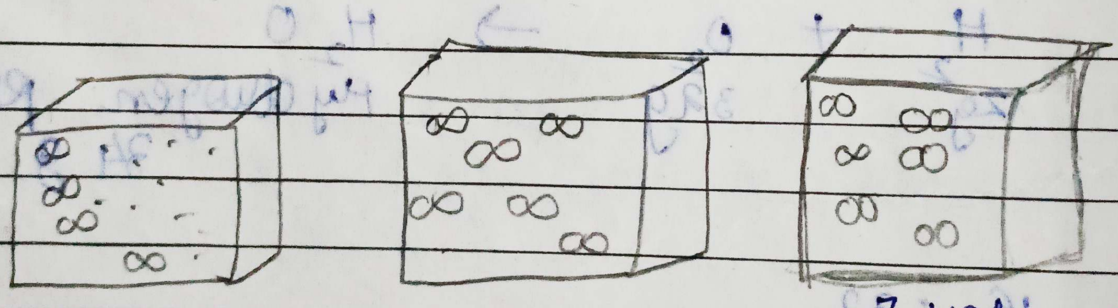
(4) Gay Lussac's law of gaseous volume
→ When gases are combine or are produced in chemical reaction, they do so in simple ratio of volume provided by all the gases at same temperature and pressure.

(5) Avogadro's law:-
→ Law state that, equal volume of all gases at same temperature and pressure should contain equal no. of molecules.

25°C
1 atm.
Volume equal.



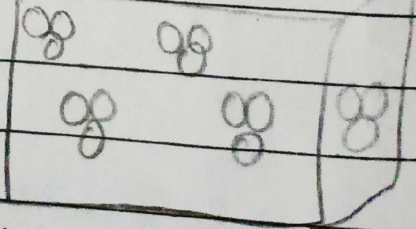
⇒



1 Volume H_2

1 Volume H_2 gas.

1 volume O_2 gas.



Diatomic molecules

$2H_2 + O_2 \rightarrow 2H_2O$

2 volume of water vapour.

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* Dalton's Atomic Theory

→ Matter consist indivisible atom.
→ Given atom of element having identical properties and mass.

ex :- C → solid
→ 12g
O → gas
→ 16 g

→ Different atom of element having different mass.
→ A compound is formed when atom of different element are combined together in definite ratio. eg. $C_6H_{12}O_6$

→ Atom can't be created or destroyed during chemical or physical change.

• Limitations :-

→ It can't explain law of chemical combination
→ It can't explain law of gaseous volume.

* Atomic Mass and molecular mass
→ 1 atomic mass unit (amu).

* **Isotopes :-**

→ Same element having different mass is called isotopes.

eg. :- $^{12}C_6$, $^{13}C_6$, $^{14}C_6$

* **Isobars :-**

→ Different element having same mass.

eg :- $^{40}_{18}Ar$ $^{40}_{20}Ca$

$$\boxed{\Rightarrow \; 1 \text{ amu} = 1.66056 \times 10^{-24} \text{ g}}$$

→ mass of H atom := 1.6736×10^{-24} g.

→ Mass of H-atom := $\dfrac{1.6736 \times \cancel{10^{-24}}}{1.66056 \times \cancel{10^{-24}}}$

= 1.0078

= 1.0080 amu.

• 1 amu defined as mass exactly equal to $1/12^{th}$ mass of one C-12 atom.

• amu → u → unified mass.

eg :- 1.0080 u.

→ avg. atomic mass.

* **Molecules. :-**

H_2, Cl_2, H_2O, CO_2

* **Mole concept and Molar Mass**

→ 1 mole is amount of substance containing particles or entities that are in fix mass of C-12 or

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isotopes ^{12}C.

$$\boxed{\Rightarrow \text{mole} = \frac{\text{given mass}}{\text{molar mass}}}$$

- Spectrometry → 1.992648×10^{-23} g/atom
  mass of C-12

$$= \frac{12 \text{ g/mol}}{1.992648 \times 10^{-23} \text{ g/atom}}$$

$$= \boxed{6.022 \times 10^{23} \frac{\text{atom}}{\text{mol}}} \rightarrow \text{Avagadro No. } (N_A)$$

---

* Mole :-
(i) In terms of particle
(ii) In terms of volume
(iii) In terms of mass.

(i) In terms of particle.
$\Rightarrow 6.022 \times 10^{23} \frac{\text{atom}}{\text{mol}}$.

- 1 mol of oxygen atom $= 6.022 \times 10^{23} \frac{\text{atom}}{\text{mol}}$
  or
  1 mol of oxygen molecules
  $= 6.022 \times 10^{23} \frac{\text{atom}}{\text{mol}}$

(ii) In terms of volume.
for gases → 22.4 lit. at NTP.

- Pressure → STP → Standard temperature and pressure.

$T = 298 K$
$P = 1 atm / 1 bar$
$C = 1 mol / lit$

(iii) In terms of mass.
(a) Mass of atom
eg :- $C = 12 g/mol$, $O = 16 g/mol$.

(b) Mass of molecules
eg :- $CO_2 = 44 g/mol$
$H_2O = 18 g/mol$

(c) Compound.
eg :- $C_6H_{12}O_6 = 180 g/mol$

(d) Formula mass
$NaCl = 58.5$, $CaCO_3 = 100$.

- 1 mol CO_2 $= 6.022 \times 10^{23}$ $\frac{atom}{mol}$
or
44 g of CO_2.
or
22.4 l volume.

* Percentage Composition.

⇒ Mass % of elements $= \frac{\text{Mass of that element in 1 mole compound}}{\text{Molecular Mass}} \times 100$

→ eg1 :- $CO_2 = 44$ elements : C, O. ⇒ C=12, O=32	Mass % of C $= \frac{12}{44} \times 100$ $= 27.27\%$	Mass % of O $= \frac{32}{44} \times 100$ $= 72.72\%$

Ex → 2 :- CH_3CH_2OH.
→ Molecular Mass $= 24+6+16$
$= 46$ g/mol

(i) Mass % of C $= \frac{24}{46} \times 100 = 52.17\%$

(ii) Mass % of H $= \frac{6}{46} \times 100 = 13.04\%$

(iii) Mass % of O $= \frac{16}{46} \times 100 = 34.78\%$.

* Mass of Cu = 63.5.
Fe = 56.
S = 32
P = 31

* Empirical Formula & Molecular Formu

(i) Empirical formula :- It is represented as the ratio of fix atom of different elements.

→ It is simplest ratio of atom of different element

eg :- $C_6H_{12}O_6$

→ $C:H_2:O$ → empirical formula

eg :- C_2H_2 or $CH\equiv CH$

$C:H$ → empirical formula.

(ii) Molecular formula :- It is represented as the exact no. of atom of different element.

eg :- 4.07% → Hydrogen

24.27% → C

71.65% → Cl

Molar Mass = 98.96 g.

Method 1

Step 1 :- Convert given mass % in mass.

Base = 100g.

H = 4.07g.

C = 24.27g

Cl = 71.65g

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Step : 2 :- Mass of atom
$H = 1g$, $C = 12g$, $Cl = 35.5\ g$

Step 3 :- Find moles.
$$\text{moles} = \frac{\text{given mass}}{\text{molar mass}}$$

$$\text{Moles of } H = \frac{4.07}{1} = 4.07$$

$$\text{Moles of } C = \frac{24.27}{12} = 2.0225$$

$$\text{Moles of } Cl = \frac{71.65}{35.5} = 2.0183$$

Step 4 :- Find the minimum ratio of moles.

$$\frac{4.07}{2.0183} : \frac{2.0225}{2.0183} : \frac{2.0183}{2.0183}$$

$$2 : 1 : 1$$

$$H_2 : C : Cl$$

$CH_2Cl \rightarrow$ empirical formula

Step 5 :- Find Molecular formula.

$$\boxed{n = \frac{\text{Molar mass of compound}}{\text{empirical molar mass}}}$$

Empirical mass $= CH_2Cl = 12 + 2 + 35.5 = 49.5$

$$n = \frac{98.96}{49.5} = \boxed{2}$$

$(CH_2Cl) \times n$ $= C_2H_4Cl_2 \rightarrow$ molecular formula
$(CH_2Cl) \times 2$

Method 2

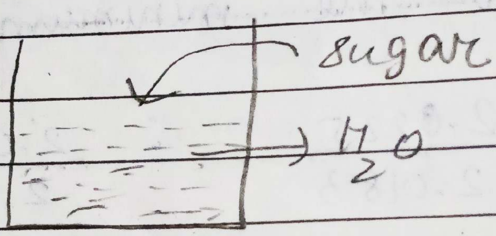
elements	mass of elements	atomic Mass	moles	ratio of moles.	simplest ratio
H	4.07	1	4.07	$\frac{4.07}{2.0183} = 2$	2 : 1 : 1
C	24.27	12	2.0225	$\frac{2.0225}{2.0183} = 1$	H_2 : C : Cl
Cl	71.65	35.5	2.0183	$\frac{2.0183}{2.0183} = 1$	

After that continue with step 5.

* Concentration of solution :-

→ solution = solute + solvent

ex :-



→ solvent :- In the solution which is present in large amount.

→ solute :- In the solution which is present in less amount.

→ $\boxed{H_2O \text{ is a universal solvent.}}$

→ solvent : 1 or A
solute : 2 or B

Concentration

- Normality
- formality
- w/w %
- v/v %
- w/v %
- ppm
- mole (n)
- mole - fraction
- Molarity
- Molality

① w/w % or mass / mass %

$$= \frac{\text{mass of solute}}{\text{mass of solution}} \times 100$$

→ w = mass , M = molar mass , n = moles

$$\boxed{w/w\ \% = \frac{w_2}{w_1 + w_2} \times 100}$$

Ex.:- Calculate mass percentage of 100g of glucose in to 300g of water.

$w_1 = 300g$ | $w_2 = 100g$

→ w/w % of glucose $= \frac{w_2}{w_1 + w_2} \times 100$

$= \frac{100}{400} \times 100 = \boxed{25\%}$

② v/v % $= \frac{\text{volume of solute}}{\text{volume of solution}} \times 100$

$$\boxed{v/v\ \% = \frac{v_2}{v_1 + v_2} \times 100}$$

③ w/v % = $\frac{\text{mass of solute}}{\text{volume of solution}} \times 100$

④ mole (n) = $\frac{\text{given mass of substance}}{\text{molar mass of substance}}$

→ $\boxed{n_2 = \frac{W_2}{M_2}}$ $\boxed{n_1 = \frac{W_1}{M_1}}$

ex) → Calculate moles of 70 ml of $CHCl_3$ (chloroform) into 140 ml of C_6H_6 (benzene)

$$n_2 = \frac{W_2}{M_2}$$

$$= \frac{70}{119.5} = 0.58$$

⑤ Mole ~~Formation~~ Fraction (x):-

→ Mole fraction of solute (x_2) = $\frac{\text{moles of solute } (n_2)}{\text{moles of solution } (n_1+n_2)}$

$$\boxed{x_2 = \frac{n_2}{n_1+n_2}}$$

→ Mole fraction of solvent (x_1) = $\frac{\text{moles of solvent}}{\text{moles of solution}}$

$$\boxed{x_1 = \frac{n_1}{n_1+n_2}}$$

→ $x = x_1 + x_2 + x_3 + \ldots\ldots + x_n$

$\boxed{x = 1 \text{ (unity)}}$

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(6) ppm :- parts per million

$$ppm = \frac{\text{No. of parts of component}}{\text{Total no. of parts of component}} \times 10^6$$

eg.:- I l sea water ⟶ 1030 g

1030 g of water → 6×10^{-3} g O_2

10^6 g " " → ?

$$= \frac{10^{6 3} \times 6 \times 10^{-3}}{1030} = 5.82 \text{ ppm or}$$

5.82 μg/ml

* Molarity (M): The no. of moles of solute present in I l of solⁿ.

→ Unit = $\boxed{\frac{mol}{lit}}$ or $\frac{mol}{dm^3}$

→ Molarity or Molar solution or M $= \frac{\text{No. of moles of solute}}{\text{Volume of sol}^n \text{ in lit.}}$

$$M = \frac{n_2}{V_2 \text{ (lit)}} \qquad n_2 = \frac{w_2}{M_2}$$

$$\boxed{M = \frac{w_2}{M_2 \times V(\text{lit})} \text{ or } M = \frac{W_2 \times 1000}{M_2 \times V(\text{ml})}}$$

w_2 = mass of solute   V = volume

M_2 = Molar mass of solute

* Molality (m) = No. of moles of solute present in 1 Kg of solvent

Unit → $m = \frac{mol}{Kg}$

Molality or molal solⁿ or m $= \frac{\text{No. of moles of solute}}{\text{mass of solvent (Kg)}}$

$$\boxed{\begin{aligned} m &= \frac{n_2}{w_1 (Kg)} \\ m &= \frac{w_2}{M_2 \times w_1 (Kg)} \quad \text{or} \\ m &= \frac{w_2 \times 1000}{M_2 \times w_1 (g)} \end{aligned}}$$

* Normality :-

$$M = \frac{n_2}{V(mL)} \times 1000$$

$$M = \frac{w_2 \times 1000}{M_2 \times V(mL)}$$

molecular mass $= M_2$

→ in Normality :-

GM_2 = gram equimolar mass.

$$\boxed{N = \frac{w_2 \times 1000}{GM_2 \times V(mL)}}$$

$$GM_2 = \frac{\text{molecular mass}}{n\text{-Factor}}.$$

→ eg. NaCl → $Na^+ Cl^-$ → n-Factor = 1

$$GM_2 = \frac{\text{Molecular mass}}{n\text{-Factor}} = \frac{58.5}{1} = 58.5.$$

eg. H_2SO_4 = 98 g/mol.

$H_2SO_4 \rightarrow 2H^+ + SO_4^{-2}$.

n- Factor = 2.

$$GM_2 = \frac{\text{Molecular mass}}{n\text{-Factor}} = \frac{98}{2} = 49 \text{ g/mol}$$

* Molarity of Mixtures :-

$$M = \frac{M_1V_1 + M_2V_2}{V_1 + V_2}.$$

* Dilution Formula :-

$$M_1V_1 = M_2V_2.$$

* Relation between Molarity & Molality

M & m

$$\frac{1}{m} = \frac{d}{M} - \frac{\text{Molar mass of solute}}{1000}$$