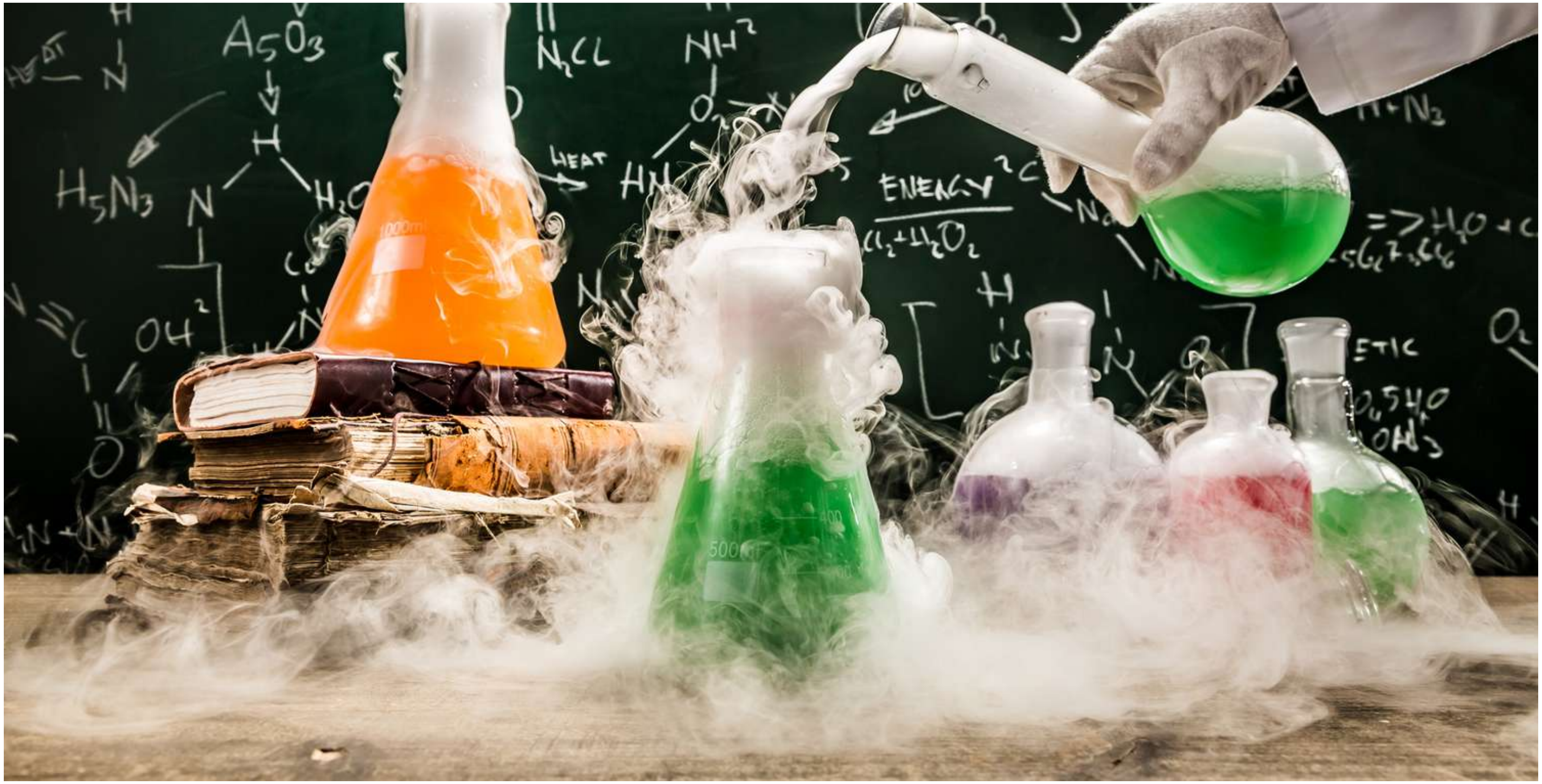
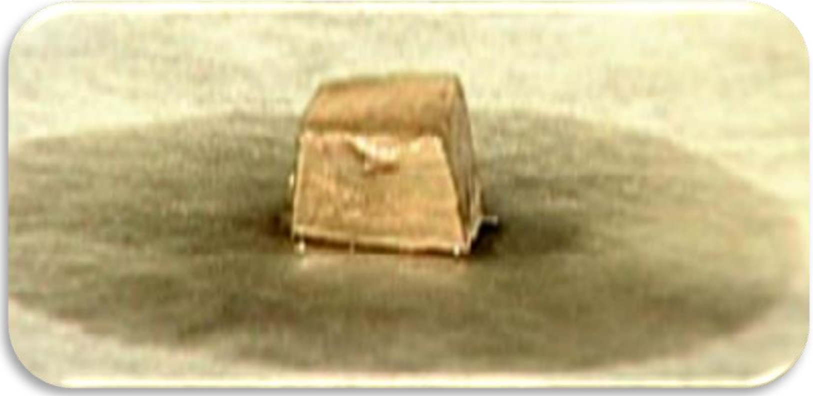
1.O MATTER

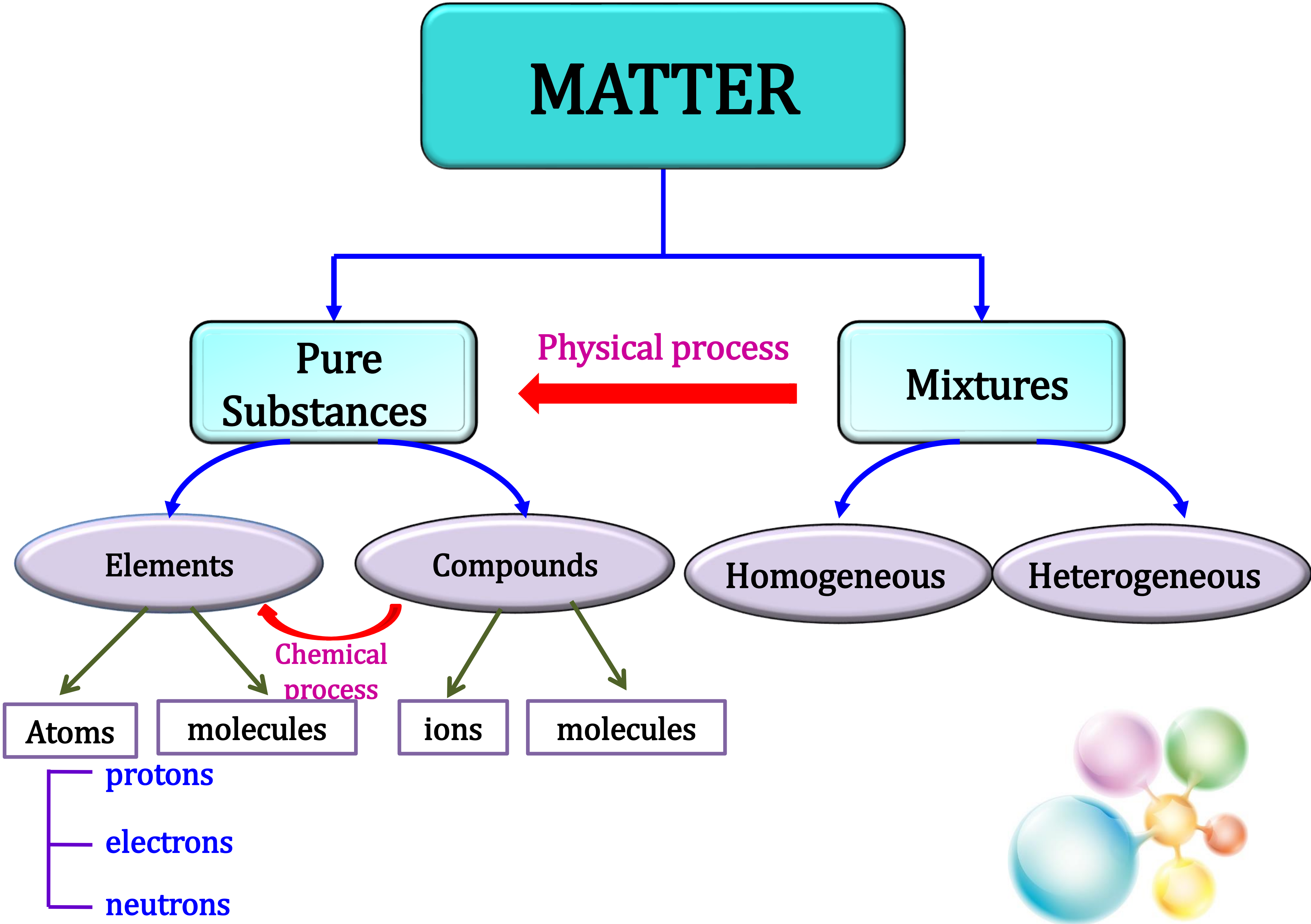
1.1 Atoms and molecules



MATTER

- **Anything that has mass and volume**




MATTER
Pure Substances
Physical process
Mixtures
Elements
Compounds
Homogeneous
Heterogeneous
Chemical process
Atoms
molecules
ions
molecules
protons
electrons
neutrons

ISOTOPE

- Two or more atoms of the same element having same proton number but different nucleon number.

@

- Two or more atoms of the same element having same number of protons but different number of neutrons.

EXAMPLE 1 :

$^{1}_{1}H$ — Protium

$^{2}_{1}H$ (D) — Deuterium

$^{3}_{1}H$ (T) — Tritium

$^{235}_{92}U$ $^{238}_{92}U$

Isotopes Notation @ Atomic Symbol

The symbol for an atom:

Nucleon Number → A

Proton Number → Z

$${}^{A}_{Z}\mathrm{X}$$

X ← Element Symbol

Example: ${}^{235}_{92}\mathrm{U}$ ${}^{1}_{1}\mathrm{H}$

PROTON NUMBER (Z)

- Number of protons in the nucleus of an atom of an element.
- Also called Atomic Number.

EXAMPLE:

All carbon atoms ($Z = 6$) have 6 protons

All oxygen atoms ($Z = 8$) have 8 protons

All uranium atoms ($Z = 92$) have 92 protons

NUCLEON NUMBER (A)

- Total number of protons and neutrons in the nucleus of an atom of an element

Nucleon Number = Number of protons + number of neutrons

$$\text{Number of Neutrons} = \underset{(A)}{\text{Nucleon Number}} - \underset{(Z)}{\text{Proton Number}}$$

- Nucleon number also called mass number

EXAMPLE 2 :

Nucleon number of chlorine, $A = 35$

$$^{35}_{17}Cl$$

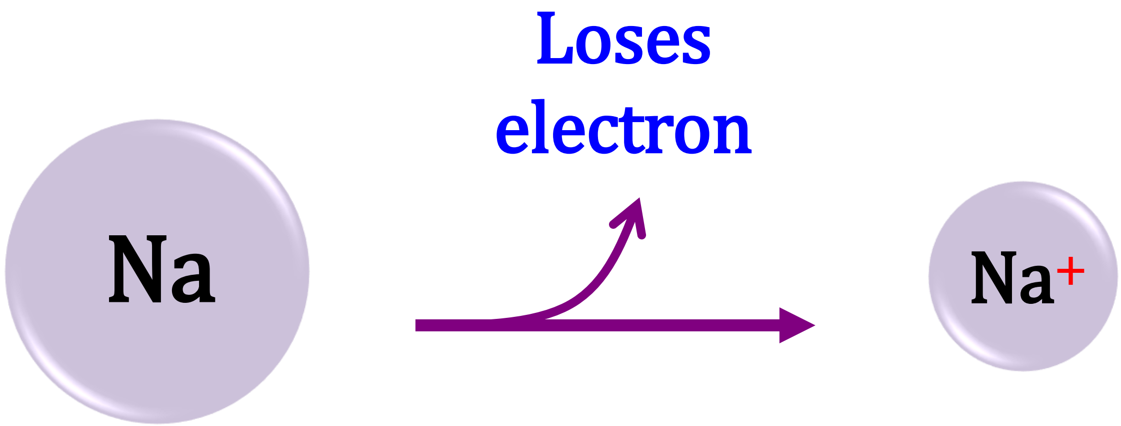
Proton number of chlorine, $Z = 17$

Number of neutrons
$= A - Z$
$= 35 - 17$
$= 18$

CATION

Cation

- ion with a positive charge
- If a neutral atom loses one or more electrons it becomes a cation.



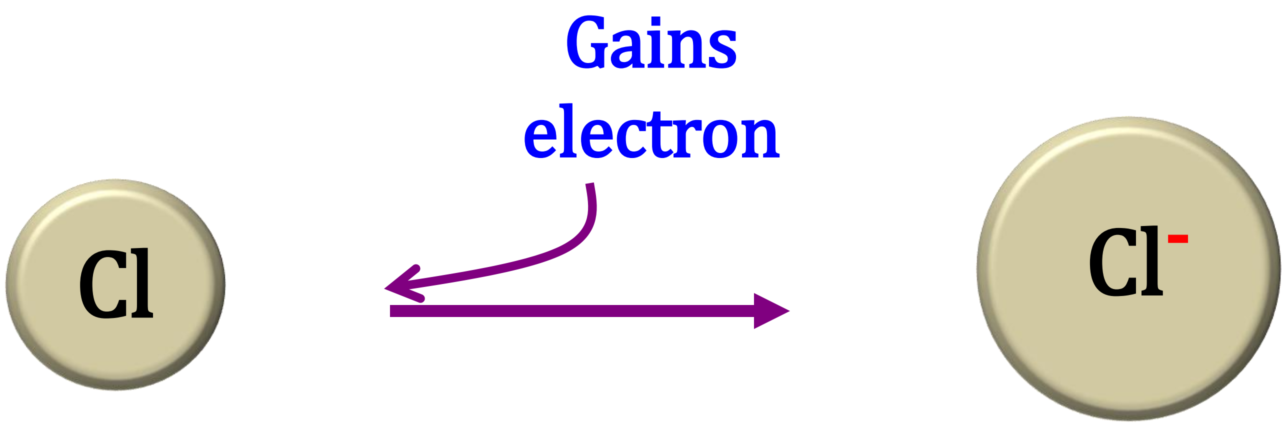
11 protons
11 electrons

11 protons
10 electrons

ANION

Anion

- ion with a negative charge
- If a neutral atom gains one or more electrons it becomes an anion.



17 protons
17 electrons

17 protons
18 electrons

NOTES !!

- The proton number, Z, is the **nuclear charge** and also **the number of electrons** in a neutral atom of the element.

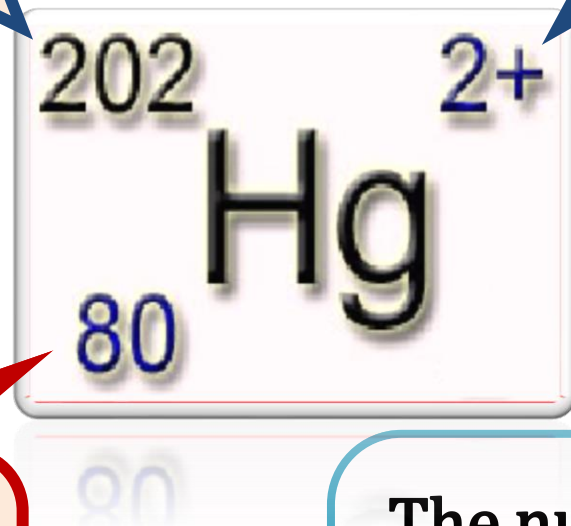
- No. of proton = no. of electron → *0 charge* ⇒ neutral

- No. of proton > no. of electron → *+ve charge* ⇒ cation
  (*atom lost electrons*)

- No. of proton > no. of electron → *-ve charge* ⇒ anion
  (*atom gained electrons*)

Nucleon number of mercury,
A = 202

Total charge on the ion



Proton number of mercury,
Z = 80

The number of neutrons
= A – Z
= 202 – 80
= 122

- Calculate the number of electrons of ion mercury, Hg^{2+} .

$$^{202}_{80}Hg^{2+}$$

Total charge on the ion

Tips:
Charge positive: (+ve)
Cation- atom lost electron

No. of proton > no. of electron

The number of electrons
= 80 - 2
= 78

EXAMPLE 3 :

Write the appropriate notation for each of the following species :

Species	Number of :			Isotope Notation
	Proton	Neutron	Electron	
A	2	2	2	
B	1	2	0	
C	1	1	1	
D	7	7	10	

Answer

Write the appropriate notation for each of the following species :

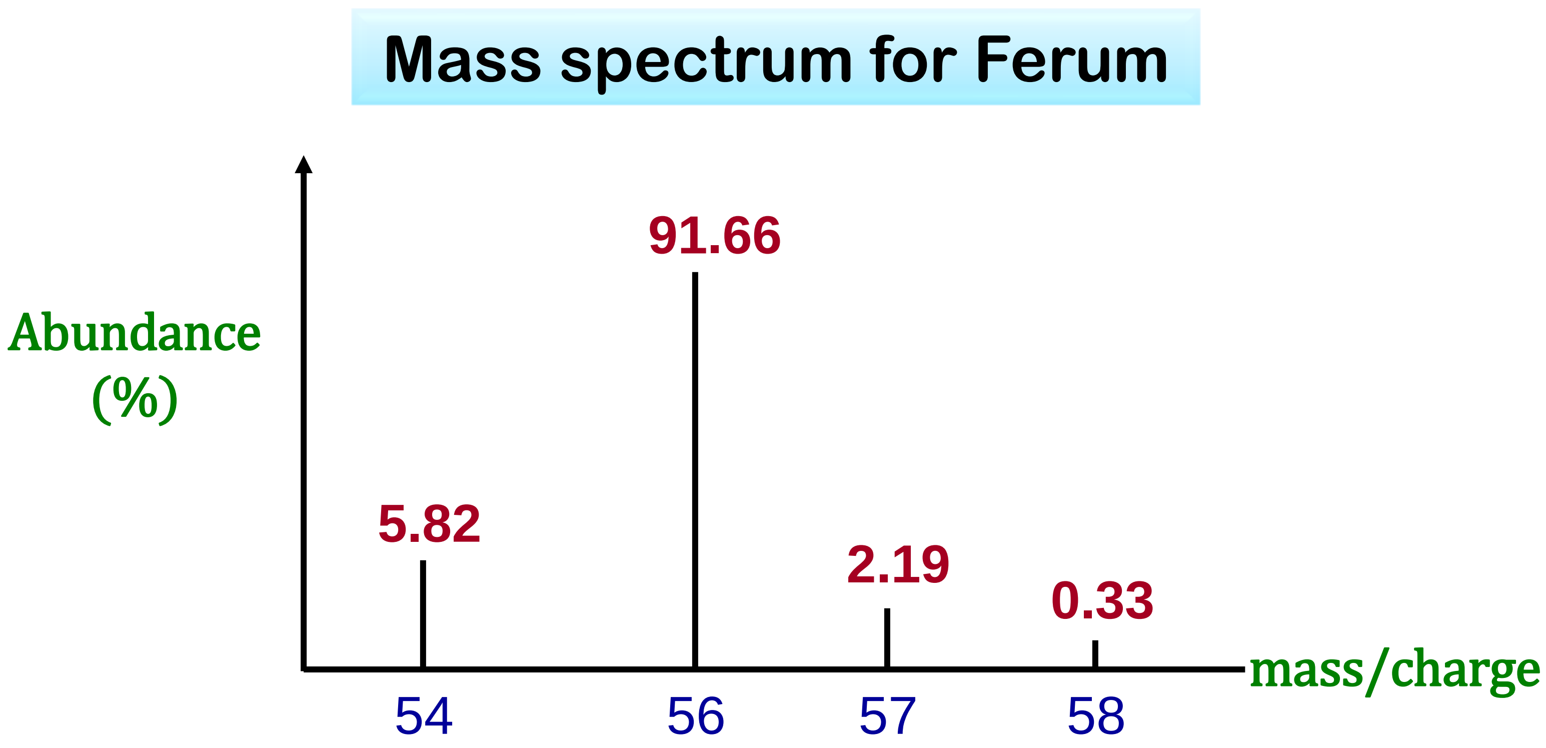
Species	Number of :			Isotope Notation
	Proton	Neutron	Electron	
A	2	2	2	$^{4}_{2}A$
B	1	2	0	$^{3}_{1}B^{+}$
C	1	1	1	$^{2}_{1}C$
D	7	7	10	$^{14}_{7}D^{3-}$

EXERCISE 1

Which of the following species has the correct number of electrons and neutrons?

	SPECIES	ELECTRON	NEUTRON
A)	${}^{24}_{11}Na^{+}$	10	13
B)	${}^{32}_{16}S^{2-}$	16	16
C)	${}^{37}_{17}Cl^{-}$	16	20
D)	${}^{58}_{26}Fe^{3+}$	29	32

INTERPRET MASS SPECTROMETER



- The height is proportional to the amount of each isotope present.

Information of A Mass Spectrum

- **height** of each line = **abundance** of each isotope.
- numbers of peaks = types of isotopes.
- ratio of m/e for each species is found from the value of the accelerating voltage associated with a particular peak.
- height of the peak is directly proportional to its abundance

AVERAGE ATOMIC MASS

- The average of mass of its naturally occurring isotopes weighted according to their abundances

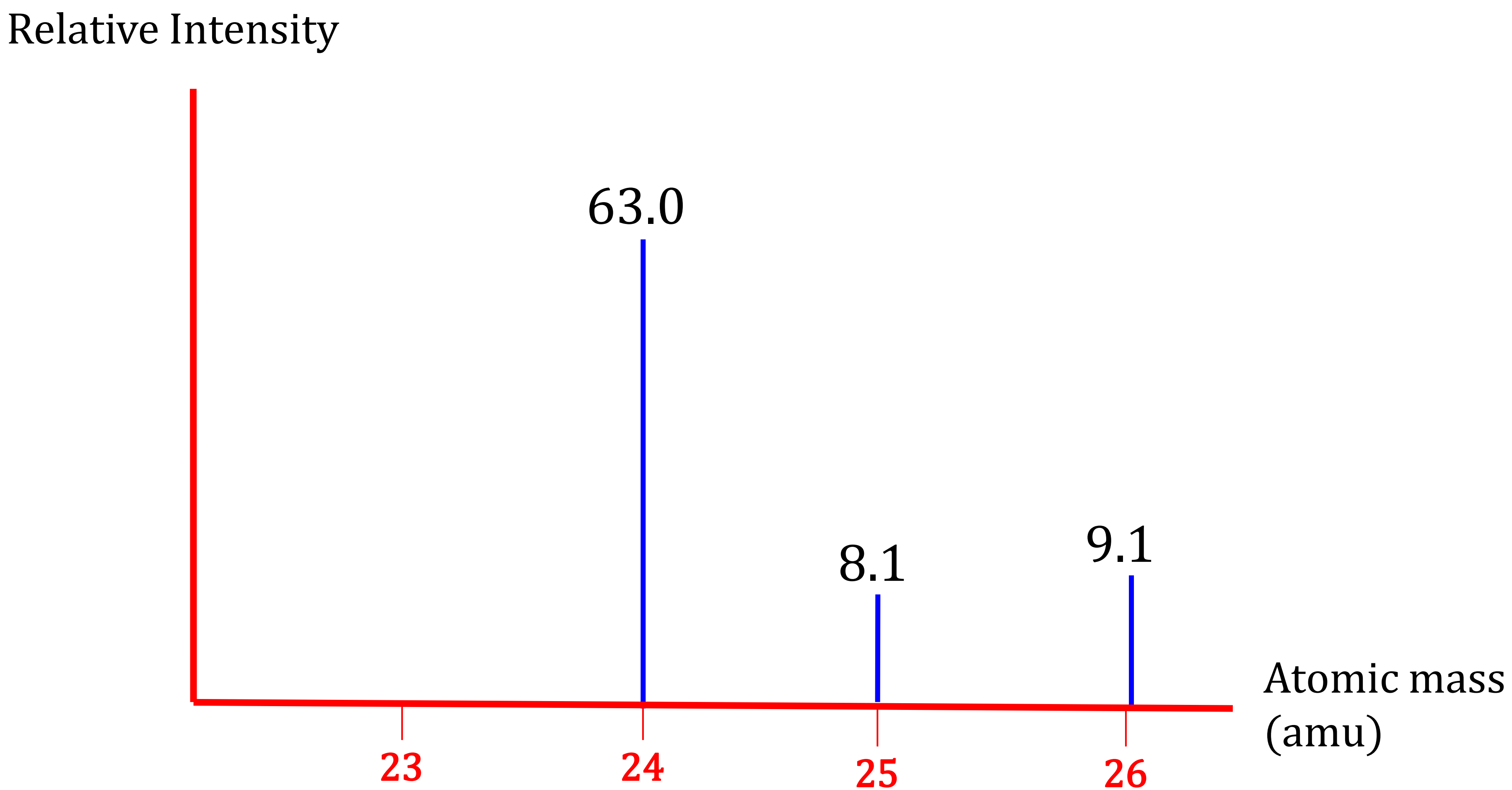
$$Average\ atomic\ mass = \frac{\sum(abundance \times isotopic\ mass)}{\sum abundance}$$

@

$$Average\ atomic\ mass = \frac{\sum(\%\ abundance \times isotopic\ mass)}{\sum \%\ abundance}$$

EXAMPLE 4:

Calculate the **average atomic mass** of Mg.



Answer

$$\boldsymbol{Average\ atomic\ mass} = \frac{\sum(\boldsymbol{abundance} \times \boldsymbol{isotopic\ mass})}{\sum \boldsymbol{abundance}}$$

$$Average\ atomic\ mass = \frac{(63.0\ \times 24) + (8.1\ \times 25) + +(9.1\ \times 26)}{63.0 + 8.1 + 9.1}$$

= 24.33 amu

EXAMPLE 5:

Nitrogen, N ($Z = 7$) has two naturally occurring isotopes. Calculate the percent abundances of $\boldsymbol{^{14}N}$ and $\boldsymbol{^{15}N}$ from the following:

atomic mass (average) of $N = 14.0067$ amu;
isotopic mass of $^{14}N = 14.0031$ amu;
isotopic mass of $^{15}N = 15.0001$ amu.

Answer

Let : **abundance of $^{14}N = x\ \%$,**
abundance of $^{15}N = 100\% - x\ \%$

Average atomic mass of nitrogen

$= (\ \%\text{ of } ^{14}N \text{ x isotopic mass of } ^{14}N) + (\%\text{ of } ^{15}N \text{ x isotopic mass of } ^{15}N)$

$$14.0067\ amu = \left[\frac{x}{100} \times 14.0031\ amu\right] + \left[\frac{100-x}{100} \times 15.0001\ amu\right]$$

$$1400.67 = 14.0031x + 1500.01 - 15.0001x$$

$$x = 99.64$$

Thus,
$x = \%$ abundance of ^{14}N
$= \mathbf{99.64\ \%}$

&

abundance of $^{15}N = 100 - x$
$= 100 - 99.64$
$= \mathbf{0.36\ \%}$

Relative Mass

- The mass of an atom relative to another atom can be determined experimentally using mass spectrometer.

- Currently, carbon-12 isotope is used as standard to measure relative atomic mass.

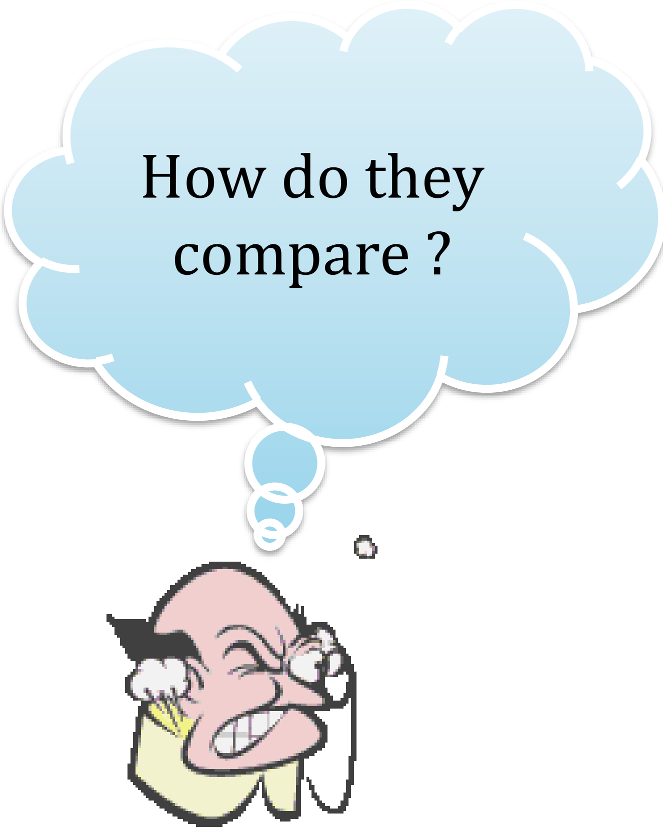
- Relative atomic mass is dimensionless.

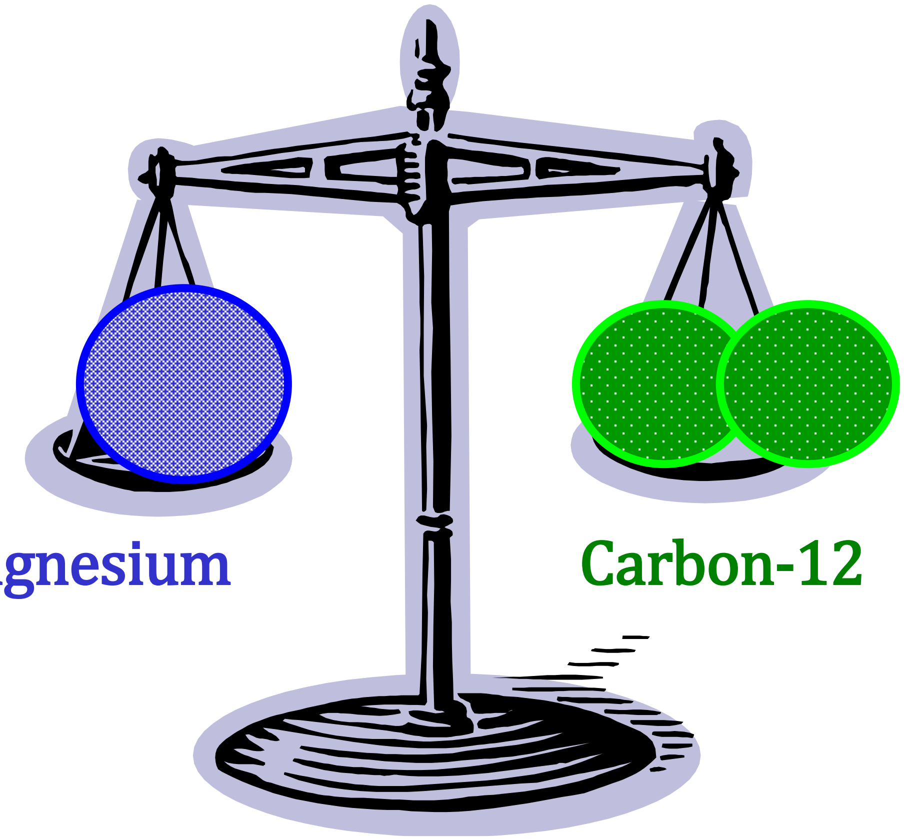
RELATIVE ATOMIC MASS, A_r

- A mass of one atom of an element compared to one twelfth mass of one atom of carbon-12 atom

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

- Relatives atomic mass is a ratio, hence, it has no 





1 atom Mg have the same weight as 2 atoms of ^{12}C

$$A_r \text{ of Mg} = 2 \text{ x } 12$$
$$= 24 \text{ amu}$$

Relative Atomic Mass of Mg

$$Relative\ atomic\ mass\ of\ Mg = \frac{mass\ of\ one\ atom\ of\ Mg}{\frac{1}{12} \times mass\ of\ one\ atom\ {}^{12}C}$$

$$Relative\ atomic\ mass, A_r of\ Mg = \frac{24\ \cancel{amu}}{\frac{1}{\cancel{12}} \times \cancel{12}\ \cancel{amu}}$$

$$= \mathbf{24}$$

EXAMPLE 6:

Determine the *relative atomic mass* of an element Y if the ratio of the atomic mass of Y to carbon-12 atom is 0.75.

Answer

Since the atomic mass ratio is 0.75

$$\frac{\text{mass of one atom } Y}{\text{mass of one atom 12 C}} = 0.75$$

$$\text{mass of one atom } Y = 0.75 \text{ x } 12.00 \text{ amu}$$
$$= 9.0 \text{ amu}$$

$$\therefore \text{ relative atomic mass, } A_r\ Y = \frac{9.0\ amu}{\frac{1}{12} \times 9.0\ amu}$$
$$= \mathbf{9.0}$$

RELATIVE MOLECULAR MASS, M_r

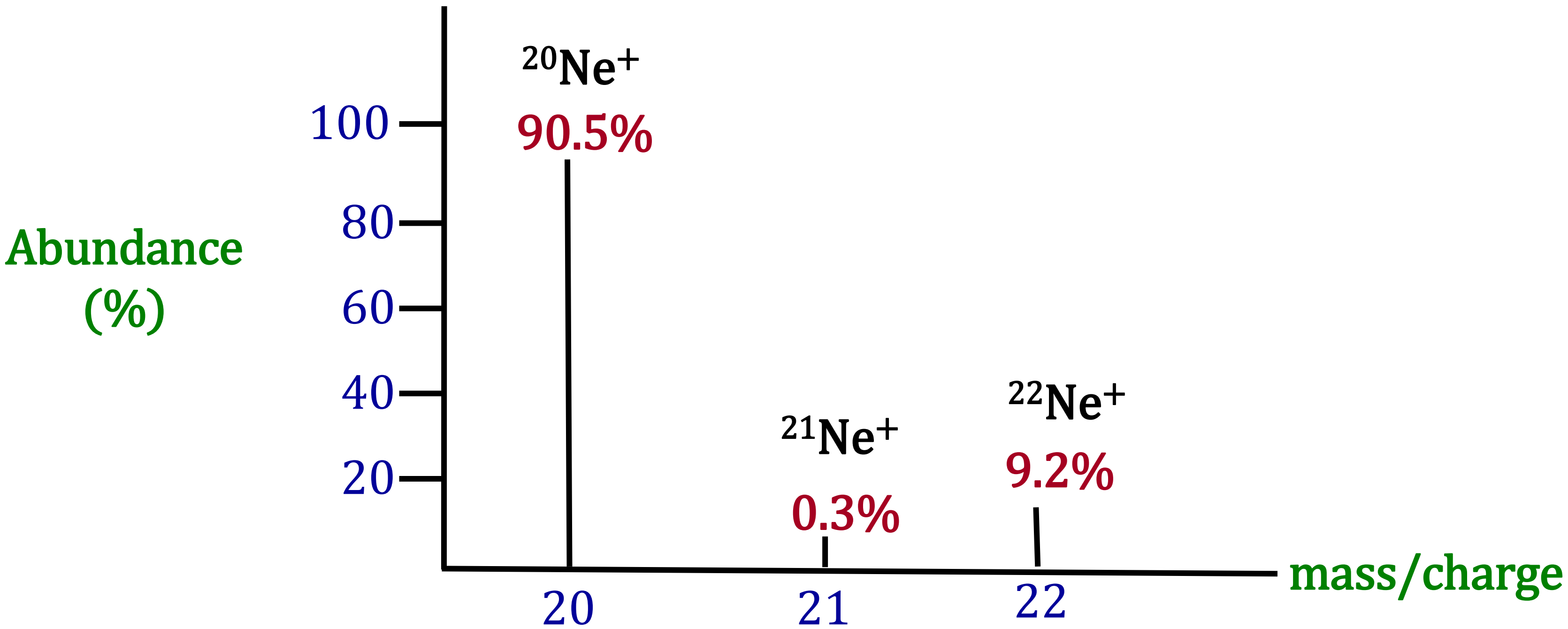
- A mass of one molecule of a compound compared to one twelfth mass of one atom of carbon-12 atom

Relative molecular mass, $M_r = \dfrac{\textbf{mass of one molecule of a compound (amu)}}{\frac{1}{12} \times \textbf{mass of one } {}^{12}\textbf{C atom (amu)}}$

- **Relatives molecular mass is a ratio, hence, it has <u>no</u> ~~UNIT~~**

EXAMPLE 7:

Calculate the relative atomic mass of neon from the mass spectrum.



Answer

$$\textit{Average atomic massof Ne} = \frac{\sum(\%\ abundance\ \times isotopic\ mass)}{\sum \%\ abundance}$$

$$= \frac{(90.5\ \times 20\ u) + (0.3\ \times 21\ u) + (9.2\ \times 22\ u)}{90.5 + 0.3 + 9.2}$$

$$= 20.2\ \mathrm{u}$$

$$\textit{Relative atomic mass of Ne} = \frac{\textit{mass of one atom of Ne}\ (u)}{\frac{1}{12} \times \textit{mass of one}\ {}^{12}C\ atom(u)}$$

$$= \frac{20.2\ \cancel{u}}{\frac{1}{\cancel{12}} \times \cancel{12.0}\ \cancel{u}}$$

$$= \mathbf{20.2}$$

NOTE: *If the atomic mass is not given, the mass number can be used as atomic mass.*

EXERCISE 2:

Copper occurs naturally as mixture of 69.09% of ^{63}Cu and 30.91% of ^{65}Cu. The isotopic masses of ^{63}Cu and ^{65}Cu are 62.93 u and 64.93 u respectively. Calculate the relative atomic mass of copper.

$A_r\ Cu = 63.55$

EXERCISE 3:

Naturally occurring iridium, Ir is composed of two isotopes, ^{191}Ir and ^{193}Ir in the ratio of 5:8. The relative isotopic mass of ^{191}Ir and ^{193}Ir are 191.021 u and 193.025 u respectively. Calculate the relative atomic mass of Iridium.

$A_r\ Ir = 192.254$

EXERCISE 4:

The atomic masses of ${}_{3}^{6}Li$ and ${}_{3}^{7}Li$ are 6.0151 a.m.u. and 7.0160 a.m.u. respectively. What is the relative abundance of each isotope if the relative atomic mass of lithium is 6.941?

$\%\ of\ {}_{3}^{6}Li\ =\ 7.49\ \%$
$\%\ of\ {}_{3}^{7}Li\ =\ 92.51\ \%$

1.2 MOLE CONCEPT

LEARNING OUTCOMES

At the end of this topic, students should be able to:

a) Define the terms empirical and molecular formulae. (C1)

b) Determine empirical and molecular formulae from mass composition or combustion data. (C3)

c) Determine the empirical formula (formula unit) from experiment (C3)

d) Define each of the following concentration measurements:(C1)

i. Molarity (M)

ii. Molality (*m*)

iii. Mole fraction (*X*)

iv. Percentage by mass (%w/w)

v. Percentage by volume (%v/v)

LEARNING OUTCOMES

At the end of this topic, students should be able to:

e) Calculate each of the following concentration measurements: (C3, C4)

   i. Molarity (M)
   ii. Molality (*m*)
   iii. Mole fraction (*X*)
   iv. Percentage by mass (%w/w)
   v. Percentage by volume (%v/v)

CHEMICAL FORMULA

Empirical Formula

- formula that shows the simplest ratio **of all elements** in a molecule.

Molecular Formula

- formula that shows the actual number of atoms of each element in a molecule.

EXAMPLE 1:

Empirical Formula	Molecular Formula
H_2O	H_2O
CH_2O	$C_6H_{12}O_6$
NH_2	N_2H_4

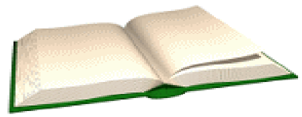
- The relationship between empirical formula and molecular formula is :

$$\boxed{\text{Molecular formula} = n\ (\text{Empirical formula})}$$

EXAMPLE 2:

A sample of hydrocarbon contains 85.7% carbon and 14.3% hydrogen by mass. Its molar mass is 56 g mol^{-1}. Determine the empirical formula and molecular formula of the compound.

Answer



Assume: mass of hydrocarbon = 100 g

Element	C	H
Mass (g)	85.7	14.3
Number of mole (mol)	$= \frac{mass}{molar\ mass}$ $= \frac{85.7}{12.0}$ $= 7.1417$	$= \frac{mass}{molar\ mass}$ $= \frac{14.3}{1.0}$ $= 14.3$
Simplest ratio	$= \frac{7.1417}{7.1417}$ $= 1.0$	$= \frac{14.3}{7.1417}$ $= 2.0$
Empirical Formula	CH_2	

- How to find **molecular formula** when given **molar mass** and **empirical formula**?

$$(\mathbf{CH_2})_{\boldsymbol{n}} = \textbf{molar mass}$$

$$[(\mathbf{12.0} + \mathbf{2}(\mathbf{1.0})]\mathbf{n} = \mathbf{56}$$

$$= \frac{56}{14.0}$$

$$\boxed{\boldsymbol{n} = \mathbf{4}}$$

Molecular formula = n (Empirical formula)

Molecular formula = 4 (CH_2)

Molecular Formula = C_4H_8

EXAMPLE 3:

A white solid was analysed and found to contain 40.0% Carbon, 6.7% Hydrogen and 53.3% oxygen by mass. What is the empirical formula of the substance?

Answer

Total % by mass = 40.0% + 6.7% + 53.3 % = **100 %**

Assume: mass of hydrocarbon = 100 g

Element	C	H	O
Mass (g)	**40.0**	**6.7**	**53.3**
Number of mole (mol)	$= \frac{mass}{molar\ mass}$ $= \frac{40.0}{12.0}$ $= \mathbf{3.3333}$	$= \frac{mass}{molar\ mass}$ $= \frac{6.7}{1.0}$ $= \mathbf{6.700}$	$= \frac{mass}{molar\ mass}$ $= \frac{53.3}{16.0}$ $= \mathbf{3.3312}$
Simplest mole ratio	$= \frac{3.3333}{3.3312}$ $= \mathbf{1}$	$= \frac{6.700}{3.3312}$ $= \mathbf{2}$	$= \frac{3.3312}{3.3312}$ $= \mathbf{1}$
Empirical Formula	CH_2O		

Empirical Formula By Combustion Data

EXAMPLE 4:

Combustion of 2.30 g of an organic sample, X, yields 3.30 g CO_2 and 1.80 g of H_2O. Determine the empirical formula of X.

Strategy : Relate the information given

Organic X contains elements: C, H & O

Mass of organic X = 2.30 g	Mass of CO_2 = 3.30 g
Mass of C + H + O = 2.30 g	Mass of H_2O = 1.80 g

Mass of C = ?

Mass of H = ?

Mass of O = ?

Answer

Mass of C = ?

mass of CO_2 → mol of CO_2 → mol of C → mass of C

$$\text{Mole of } CO_2 = \frac{3.30\ \cancel{g}\ CO_2}{44.0\ \cancel{g}\ mol^{-1} CO_2} = 0.075 \text{ mol } CO_2$$

1 mol of CO_2 contains 1 mol of C

0.075 mol of CO_2 contain **0.075 mol of C**

Mass of carbon = no. of mole C x molar mass C

= 0.075 x 12.0

$\boxed{\text{Mass of carbon} = 0.90 \text{ g C}}$

Mass of H = ?

mass of H_2O → mol of H_2O → mol of H → mass of H

$$\text{Mole of } H_2O = \frac{1.80\ \cancel{g}\ H_2O}{18.0\ \cancel{g}\ mol^{-1} H_2O} = 0.10 \text{ mol } H_2O$$

1 mol of H_2O contains 2 mol of H

0.10 mol of H_2O contain 0.20 mol of H

$\boxed{\text{Mass of hydrogen} = 0.20 \times 1.0 = 0.20 \text{ g H}}$

Mass of O = ?

Mass of C + H + O = 2.30 g

Mass of O = 2.30 – 0.90 – 0.20

$\boxed{\text{Mass of O} = 1.20 \text{ g O}}$

Element	C	H	O
Mass (g)	0.90	0.20	1.20
Number of mole	0.90 / 12.0 = 0.075	0.20 / 1.0 = 0.200	1.20 / 16.0 = 0.075
Simplest mole ratio	0.075/ 0.075 = 1 1 x 3 = 3	0.200/ 0.075 $\boxed{= 2.667}$ 2.667 x 3 = 8	0.075/ 0.075 = 1 1 x 3 = 3
The empirical formula is	$C_3H_8O_3$		

Never round off values close to whole number in order to get a simple ratio but multiply the value by a factor until you get a whole number.

Another method of calculation

FROM **EXAMPLE 4**

Combustion of 2.30 g of an organic sample, X, yields 3.30 g CO_2 and 1.80 g of H_2O. Determine the empirical formula of X.

Answer

Mass of C = ?

$$\text{Mass of C} = \frac{12.0\ \cancel{\text{g mol}^{-1}}}{44.0\ \cancel{\text{g mol}^{-1}}} \times 3.30\ \text{g}$$

$$\text{Mass of C} = \mathbf{0.90\ g}$$

Combustion of 2.30 g of an organic sample, X, yields 3.30 g CO_2 and 1.80 g of H_2O. Determine the empirical formula of X.

Mass of H = ?

$$\text{Mass of H} = \frac{2.0\ \text{g}\ \cancel{\text{mol}^{-1}}}{18.0\ \text{g}\ \cancel{\text{mol}^{-1}}} \times 1.80\ \text{g}$$

$$\textbf{Mass of H} = \mathbf{0.20\ g}$$

Mass of O = ?

Same as previous method

$$\text{Mass of C + H + O} = 2.30\ \text{g}$$

$$\text{Mass of O} = 2.30 - 0.90 - 0.20$$

$$\textbf{Mass of O} = \mathbf{1.20\ g}$$

CONCENTRATION UNITS

Solute

- is the substance being dissolved and present in the smaller amount

Solvent

- is the substance doing the solving and present in the larger amount

Solution

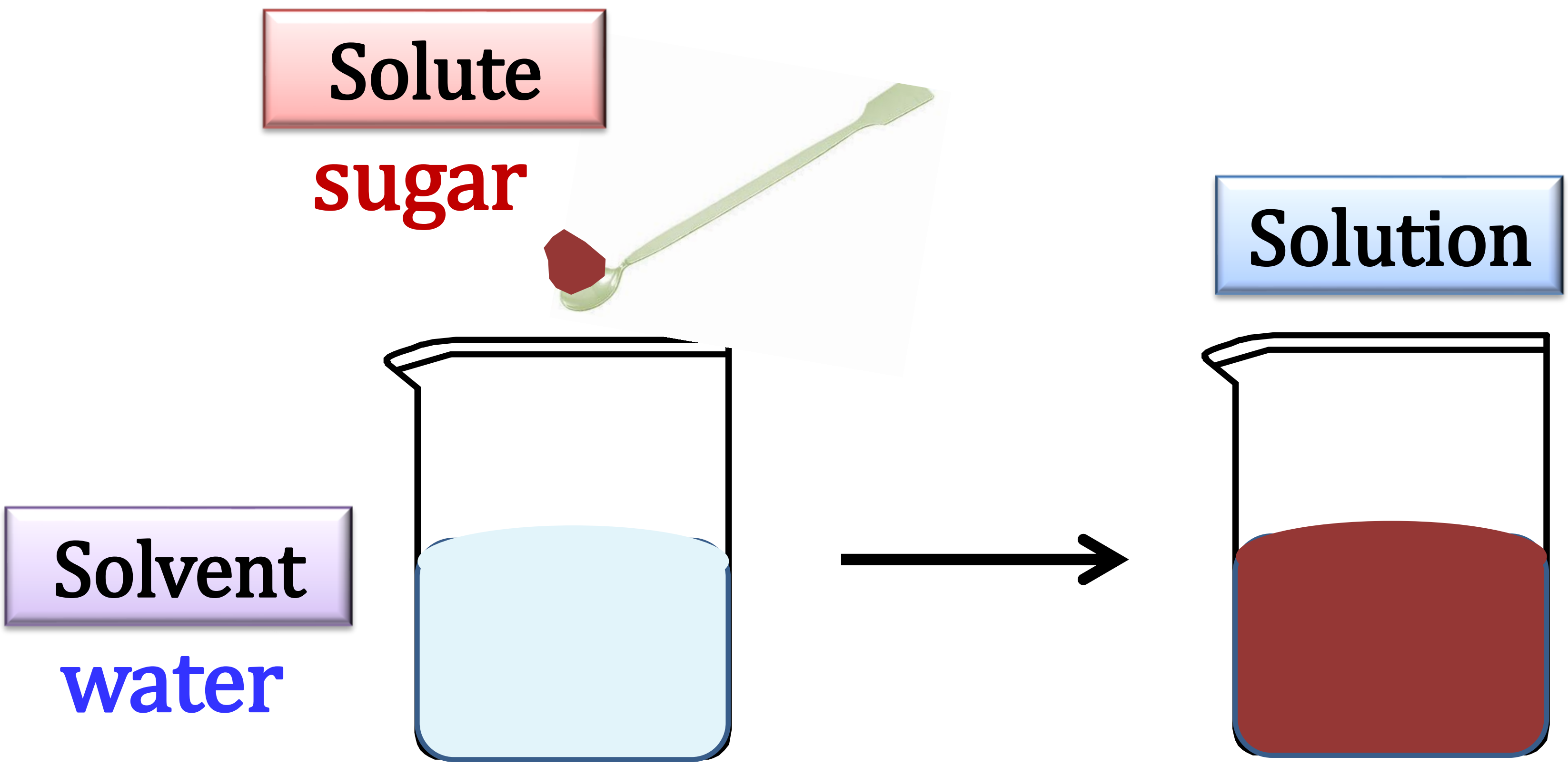
- is a homogenous mixture of 2 or more substances that formed when amount of solute dissolves completely in a solvent

Concentration

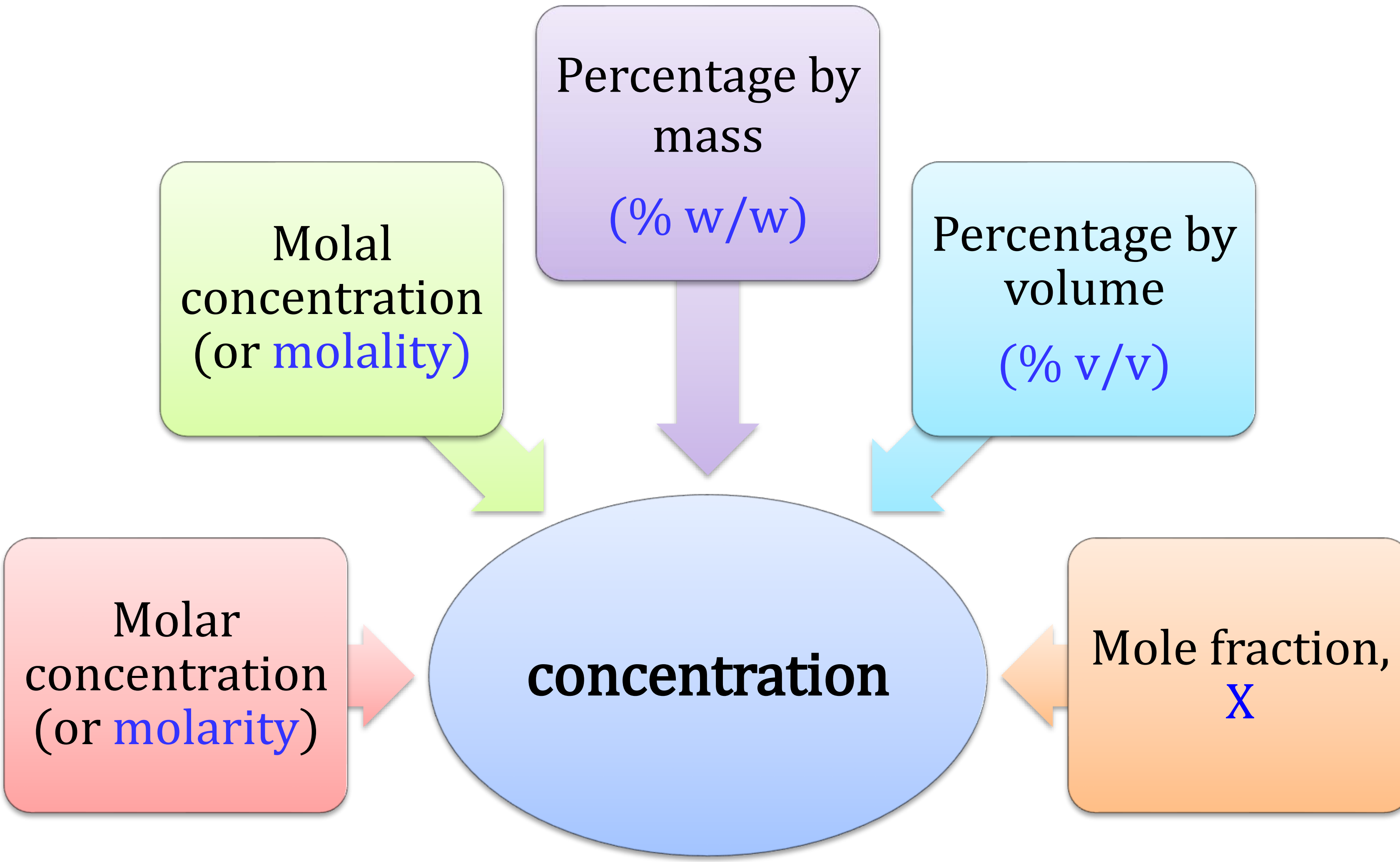
- is the amount of solute present in a given quantity of solvent in a solution

EXAMPLE 5:

WATER + SUGAR $\longrightarrow$ SUGAR SOLUTION



Expression of Concentration



a MOLARITY, M

- The number of mole of solute dissolved in 1 L of solution.

- Unit: mol L^{-1} @ mol dm^{-3} @ molar, M

$$Molarity, M = \frac{moles\ of\ solute\ (mol)}{volume\ of\ solution\ (L)}$$

$$1\ L = 1\ dm^3$$
$$1\ ml = 1\ cm^3$$
$$1\ dm^3 = 1000\ cm^3$$

EXAMPLE 6:

1.46 molar glucose ($C_6H_{12}O_6$)

or

1.46 mol/L glucose ($C_6H_{12}O_6$)

1.46 mole of glucose ($C_6H_{12}O_6$) solute

1 L of the solution

EXAMPLE 7:

A matriculation student prepared a solution by dissolving 5.528×10^{-3} mol of sodium carbonate, Na_2CO_3 in 250.0 cm^3 of water. Calculate its molarity.

Answer

Solute = sodium carbonate, Na_2CO_3 **Solvent** =water, H_2O

Solution = sodium carbonate solution,Na_2CO_3 + H_2O

- V of solution: $250.0\ cm^3 = 250\ ml = 0.25\ L$
- $$Molarity = \frac{moles\ of\ solute\ (mol)}{volume\ of\ solution\ (L)}$$

$$= \frac{5.528 \times 10^{-3}\ mol}{0.25\ L} = 0.022\ mol\ L^{-1}$$

EXAMPLE 8:

A student prepared a solution of NaCl by dissolving 1.461 g of NaCl in a 250 mL volumetric flask. What is the molarity of this solution.

Answer

Solute = ?	Solvent = ?	Solution = ?
$NaCl$	H_2O	$NaCl + H_2O$

- $$Mole\ of\ NaCl = \frac{Mass\ of\ NaCl\ (g)}{Molar\ mass\ of\ NaCl\ (g\ mol^{-1})}$$

$$= \frac{1.461\ \cancel{g}}{58.5\ \cancel{g}\ mol^{-1}}$$

$$= \mathbf{0.0250\ mol\ NaCl}$$

- Moles of NaCl = 0.0250 mol
- V of NaCl solution : 250.0 cm^3 = 250 ml = 0.25 L

$$\boxed{Molarity, M = \frac{Moles\ of\ NaCl\ (mol)}{Volume\ of\ NaCl\ solution\ (L)}}$$

$$= \frac{0.0250\ mol}{0.250\ L}$$

= 0.10 mol/L NaCl solution @

= 0.10 M NaCl solution

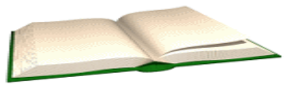
EXAMPLE 9:

How many grams of calcium chloride, $CaCl_2$ should be used to prepare 250.00 mL solution with a concentration of 0.500 M.

Answer

Solute = ?	Solvent = ?	Solution = ?
$\downarrow$	$\downarrow$	$\downarrow$
$CaCl_2$	H_2O	$CaCl_2 + H_2O$

Answer



- $n_{CaCl_2} = M_{CaCl_2} \times V_{solution}$

$$= 0.500 \text{ mol}/\cancel{\text{L}} \times 250 \times 10^{-3} \cancel{\text{L}}$$

$$\mathbf{n_{CaCl_2} = 0.125 \text{ mol}}$$

- Mass of $CaCl_2 = n_{CaCl_2} \times$ molar mass $CaCl_2$

$$= 0.125 \cancel{\text{mol}} \times 111.1 \text{ g}/\cancel{\text{mol}}$$

Mass of $CaCl_2$ = 13.89 g

b MOLALITY

- the number of moles of solute per 1 kg of solvent in a solution.
  - (unit: mol kg^{-1} @ molal @ m)

$$Molality, m = \frac{moles\ of\ solute\ (mol)}{mass\ of\ solvent (kg)}$$

<u>Mass</u> of solution = <u>mass</u> of solute + <u>mass</u> of solvent

<u>Volume</u> of solution ≠ <u>volume</u> of solvent

EXAMPLE 10:

Calculate the molality of a solution prepared by dissolving 0.288 mol of $CaCl_2$ in 271 g of water?

Answer

Solute = ?	→	$CaCl_2$	Solvent = ?	→	H_2O

- Moles of $CaCl_2$ = ? Mass of water = ?
- Moles of $CaCl_2$ = 0.288 mol Mass of water = 271/1000 kg
- $$Molality, m = \frac{moles\ of\ solute\ (mol)}{mass\ of\ solvent(kg)}$$

$$= \frac{0.288\ \text{mol}}{271\ \times\ 10^{-3}\ kg}$$

$$= \mathbf{1.06\ mol\ kg^{-1}}$$

EXAMPLE 11:

Calculate the molal concentration of ethylene glycol ($C_2H_6O_2$) solution containing 8.40 g of ethylene glycol in 200 g of water. The molar mass of ethylene glycol is 62 g/mol.

Answer

Solute = ? ⇒ $C_2H_6O_2$ **Solvent = ?** ⇒ H_2O

- ➢ Moles of $C_2H_6O_2$ = ? mol
- Mass of water = (200/1000) kg

➢ No. of moles of $C_2H_6O_2$

$$= \frac{8.40\ \cancel{g}}{62\ \cancel{g}/mol}$$

$$= 0.1355\ mol$$

➢ $Molality, m = \dfrac{moles\ of\ solute\ (mol)}{mass\ of\ solution (kg)}$

➢ $Molality, m = \dfrac{moles\ of\ C_2H_6O_2\ (mol)}{mass\ of\ water (kg)}$

$$= \frac{0.1355\ mol}{0.2\ kg}$$

$$= 0.68\ mol\ kg^{-1}$$

EXERCISE 1:

A solution containing 8.89 g glycerol, $C_3H_8O_3$ in 75.0 g of ethanol, C_2H_6O. What is the molality of the solution?

Answer: 1.29 molal

c MOLE FRACTION, X

- Mole fraction is the ratio of the number of moles of one component to the total of number of moles present.

$$Mole\ fraction\ of\ A, X_A = \frac{n_A}{n_{Total}}$$

Where;

- n_A = the no. of moles of component A in the mixture
- n_{Total} = the total no. of moles in all component in the mixture
  $= n_A + n_B + \ldots\ldots$

- If a solution containing A, B and C:

  - $$Mole\ fraction\ of\ A, X_A = \frac{n_A}{n_A + n_B + n_C}$$

    $$= \frac{n_A}{n_{Total}}$$

- ✓ No unit for mole fraction
- ✓ Mole fraction always smaller than 1
- ✓ Total mole fraction in mixture = 1

EXAMPLE 12:

What is the mole fraction of $CuCl_2$ in a solution prepared by dissolving 0.30 mole of $CuCl_2$ in 40.0 mole of H_2O ?

Answer

- $Mole\ fraction\ of\ CuCl_2 = \dfrac{n_{CuCl_2}}{(n_{CuCl_2} + n_{H_2O})}$

$$= \frac{0.30\ \cancel{mol}}{(0.30 + 40.0)\cancel{mol}}$$

$$= \mathbf{0.0074}$$

EXAMPLE 13:

A sample of ethanol, C_2H_5OH contains 200.0 g of ethanol and 150.0 g of water. Calculate the mole fraction of;

a) Ethanol

b) water

in the solution.

Answer

$$n_{ethanol} = \frac{mass\ (g)}{molar\ mass\ (g\ mol^{-1})}$$

$$= \frac{200.0\ g}{46.0\ g\ mol^{-1}}$$

$$\boldsymbol{n_{ethanol} = 4.3478\ mol}$$

$$\boldsymbol{X_{ethanol} = \frac{n_{ethanol}}{n_{ethanol} + n_{water}}}$$

$$= \frac{4.3478\ mol}{(4.3478 + 8.3333)\ mol}$$

$$\boldsymbol{X_{ethanol} = 0.3429}$$

$$n_{water} = \frac{mass\ (g)}{molar\ mass\ (g\ mol^{-1})}$$

$$= \frac{150.0\ g}{18.0\ g\ mol^{-1}}$$

$$\boldsymbol{n_{water} = 8.3333\ mol}$$

$$\boldsymbol{X_{water} = 1 - X_{ethanol}}$$

$$= 1 - 0.3429$$

$$\boldsymbol{X_{water} = 0.6571}$$

d PERCENTAGE BY MASS (%w/w)

- The ratio of the mass of a solute to the mass of the solution, multiplied by 100

- $Percentage\ by\ mass, \%\,{}^{w}/_{w} = \dfrac{mass\ of\ solute\ (g)}{mass\ of\ solution\ (g)} \times 100$

@

- $Percentage\ by\ mass, \%\,{}^{w}/_{w} = \dfrac{mass\ of\ solute\ (g)}{mass\ of\ solute + mass\ of\ solvent\ (g)} \times 100$

Since, Mass of solution = mass of solute + mass of solvent

Example 14:

Given that 10% percent by mass of NaOH in the solution.

From equation :

- $Percentage\ by\ mass, \%\ ^{w}/_{w} = \frac{\boldsymbol{mass\ of\ solute}\ (g)}{mass\ of\ solution\ (g)} \times 100$

$\Rightarrow$ 10 g of NaOH dissolved in 100 g of solution

$\Rightarrow$ 10 g of NaOH dissolved in 90 g of solvent (water)

Assume mass of solution = 100 g
Mass of solute, NaOH = 10 g

mass of solvent = Mass of solution - mass of solute
= 100 g - 10 g
= 90 g

Example 15:

A sample of 0.892 g of potassium chloride, KCl is dissolved in 54.3 g of water. What is the percent by mass of KCl in this solution?

Solute = ? ; Solvent = ? ; Solution = ?
Solute = KCl ; Solvent = H_2O ; = KCl + H_2O

$$Percentage\ by\ mass, \%\,{}^{w}/_{w} = \frac{mass\ of\ solute\ (g)}{mass\ of\ solution\ (g)} \times 100$$

$$\%\,{}^{w}/_{w} = \frac{0.892\ g}{(0.892 + 54.3)\ g} \times 100$$

$$= \mathbf{1.616\ \%}$$

EXERCISE 2:

A solution is made by dissolving 4.2 g of sodium chloride, NaCl in 100.00 mL of water. Calculate the mass percent of sodium chloride in the solution.

Answer : 4.0%

e PERCENTAGE BY VOLUME (% v/v)

- the ratio of the volume of a solute to the volume of the solution, multiplied by 100

> $$Percentage\ by\ volume, \%\ {}^{v}/_{v} = \frac{volume\ of\ solute}{volume\ of\ solution} \times 100$$

- Most often used for liquids and gas

Example 16:

A 200 mL of perfume contains 28 mL of alcohol. What is the % concentration of alcohol by volume in this solution?

Answer

- $Percentage\ by\ volume, \%\,{}^{v}/_{v} = \dfrac{volume\ of\ solute}{volume\ of\ solution} \times 100$

$$= \frac{volume\ of\ alcohol}{volume\ of\ perfume} \times 100$$

$$= \frac{28\ mL}{200\ mL} \times 100$$

$$= \mathbf{14\ \%}$$

Example 17:

A 350 cm^3 sample of vinegar contains 2.10 cm^3 of ethanoic acid. What is the concentration of ethanoic acid by volume in this sample?

Answer

- $Percentage\ by\ volume, \%\, {}^{v}/_{v} = \dfrac{volume\ of\ solute}{volume\ of\ solution} \times 100$

$$\%\, {}^{v}/_{v} = \frac{2.10\ cm^3}{350\ cm^3} \times 100$$

$$= \mathbf{0.6\ \%}$$

EXERCISE 3:

Suppose you have 265.5 mL of an aqueous ethanol solution that is 30.0 % ethanol by volume. How much ethanol (in mL) is in the bottle?

Answer: 79.65 mL

- For question that give concentration and density, can make assumptions based on concentration.

Unit Concentration	Formula	Assumption (Suggestion)
$Molarity$	$\frac{\boldsymbol{moles\ of\ solute\ (mol)}}{\boldsymbol{volume\ of\ solution\ (L)}}$	$\boldsymbol{Volume\ of\ solution = 1\ L}$
$Molality$	$\frac{\boldsymbol{moles\ of\ solute\ (mol)}}{\boldsymbol{mass\ of\ solvent (kg)}}$	$\boldsymbol{Mass\ of\ solvent = 1\ kg}$
$\%\ ^{w}/_{w}$	$\frac{\boldsymbol{mass\ of\ solute\ (g)}}{\boldsymbol{mass\ of\ solution\ (g)}} \times 100$	$\boldsymbol{Mass\ of\ solution = 100\ g}$
$\%\ ^{v}/_{v}$	$\frac{\boldsymbol{volume\ of\ solute (ml)}}{\boldsymbol{volume\ of\ solution\ (mL)}} \times 100$	$\boldsymbol{Volume\ of\ solution = 100\ mL}$

EXERCISE 4:

An aqueous solution of ethylene glycol used as an automobile engine coolant is 40.0% $HOCH_2CH_2OH$ by weight and has a density of 1.05 g/ml. What are the;

a) Molarity
b) Molality

of $HOCH_2CH_2OH$ in the solution

a) Molarity, $M = 6.774$ mol/L

b) Molality, $m = 10.753$ mol/kg

EXERCISE 5:

1. An 8.00%(w/w) aqueous solution of ammonia has a density of 0.9651 g mL^{-1}. Calculate the

   a) Molarity

   b) Molality

   c) mole fraction of the NH_3 solution

Answer

1 : a) 5.12 mol kg^{-1} , b) 4.54 mol L^{-1} ,c) 0.0843

1.3 STOICHIOMETRY

LEARNING OUTCOMES

At the end of this topic, students should be able to:

a) Write a balanced chemical equation :
   - a) by inspection method
   - b) by ion-electron method (redox equation) (C3)

b) Define: (C1)
   - i. limiting reactant
   - ii. percentage yield.

c) Perform stoichiometric calculations using mole concept including limiting reactant and percentage yield. (C3, C4)

Balancing Chemical Equation

Chemical equation denotes a chemical reaction :

$$x\mathrm{A} + y\mathrm{B} \longrightarrow z\mathrm{C} + w\mathrm{D}$$

(reactant) (product)

$x, y, z,$ and w = stoichiometric coefficients

Two methods of balancing chemical equation:

a) Inspection method

b) Ion-electron method

Inspection Method

Write down the unbalanced equation. Write the correct formulae for the reactants and products.

Balance element/atom

i. Balance the metallic element, followed by non-metallic atoms.
ii. Balance the hydrogen and oxygen atoms.

Check to ensure that the total number of atoms of each element is the same on both sides of equation.

EXAMPLE 1:

- start with C

$$\underset{\text{2 C atom}\quad\text{6 H atoms}}{C_2H_6} + O_2 \longrightarrow \underset{\text{1 C atom}}{CO_2} + \underset{\text{2 H atoms}}{H_2O}$$

$$C_2H_6 + \frac{7}{2}O_2 \longrightarrow \underbrace{2CO_2 + 3H_2O}_{\text{7 O atoms}}$$

$$2C_2H_6 + 7O_2 \longrightarrow 4CO_2 + 6H_2O$$

Reactants	Products
4 C	4 C
12 H	12 H
14 O	14 O

EXAMPLE 2:

- **Balance** pure elements such as **Al**.

$$Al + H_2SO_4 \longrightarrow Al_2(SO_4)_3 + H_2$$

1 Al atom

2 Al atoms

- **Balance polyatomic ions** as a unit

$$2Al + 3H_2SO_4 \longrightarrow Al_2(SO_4)_3 + 3H_2$$

6 H atoms

1 SO_4^{2-} ions

3 SO_4^{2-} ions

3 H atoms

- **Balance** H.

Reactants	Products
2 Al	2 Al
6 H	6H
3 S	3 S
12 O	12 O

EXERCISE 1:

Balance these equations.

a) $Al(s) + O_2(g) \rightarrow Al_2O_3(s)$

b) $N_2(g) + H_2(g) \rightarrow NH_3(g)$

c) $C_6H_6(l) + O_2(g) \rightarrow CO_2(g) + H_2O(l)$

Answer

EXERCISE 2:

Balance the following chemical equation :

1. $Fe_2O_3 + HCl \longrightarrow FeCl_3 + H_2O$
2. $NH_3 + CuO \longrightarrow Cu + N_2 + H_2O$
3. $C_2H_5OH + O_2 \longrightarrow CO_2 + H_2O$
4. $AgNO_3 + Na_2CrO_4 \longrightarrow Ag_2CrO_4 + NaNO_3$
5. $Al_2(SO_4)_3 + KOH \longrightarrow Al(OH)_3 + K_2SO_4$

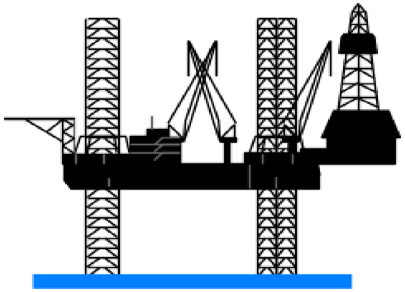
Answer

Balancing Redox Equation

Redox Reaction

- Redox reaction is a reaction that involves both reduction and oxidation reactions.

Oxidation	Reduction
An increase in oxidation number	A decrease in oxidation number
Loss of electrons	Gain of electrons



OIL - Oxidation Is the Loss of electrons

RIG - Reduction Is the Gain of electrons

Oxidation Number

- **Oxidation numbers** of any atoms can be determined by applying the following **rules**:

1. For **monoatomic ions**,

   oxidation number = the **charge** on the ion

Example:

Monoatomic ions	oxidation number
Na^{+}	+1
Cl^{-}	-1
Al^{3+}	+3
S^{2-}	-2

2. For free elements, oxidation number on each atom = 0

Example:

Elements	oxidation number
Na	0
O_2	0
Br_2, P_4, S_8	0

3. For most cases, oxidation number for;

Example:

Atom	oxidation number
O	-2
H	+1
F, Cl, Br, I	-1

Exception:

1. H bonded to metal

Example: NaH, MgH_2

oxidation number for **H** = **-1**

2. Halogen bonded to oxygen

- oxidation number for halogen = **+ve**

Example: Cl_2O_7

oxidation number for **Cl** = **+7**

3. In a neutral compound

- The total oxidation number of each atoms that made up the molecule is zero.

Example :

Determine the oxidation number of iodine, I in $NaIO_3$

$$\text{Total oxidation no.} = +1 + I + 3(-2) = 0$$

$$I = +5$$

4. The oxidation number of polyatomic ions

- The total oxidation number of each atoms that made up the ion is equal to the net charge **of the ion**.

Example :

Determine the oxidation number of chromium, **Cr** in $Cr_2O_7^{2-}$

Total oxidation no. $= 2Cr + 7(-2) = -2$

$Cr = +6$

EXAMPLE 3:

Balance the chemical equation in **acidic solution** using the Ion–electron Method/Half – Reaction Method

$$Fe^{2+} \quad + \quad Cr_2\,O_7{}^{2-} \longrightarrow Fe^{3+} + Cr^{3+}$$

Oxidation

$\boxed{+2}$ $\longrightarrow$ $\boxed{+3}$

$$Fe^{2+} \quad + \quad Cr_2O_7^{\;2-} \longrightarrow Fe^{3+} + Cr^{3+}$$

$\boxed{+6}$ Reduction $\boxed{+3}$

STEPS:

1. Separate the equation into **two half reactions**.

Oxidation : $Fe^{2+} \longrightarrow Fe^{3+}$

Reduction : $Cr_2O_7^{\;2-} \longrightarrow Cr^{3+}$

2. Balance the atoms other than O and H in each half-reaction.

$$Fe^{2+} \rightarrow Fe^{3+} \quad \text{(no change)}$$

$$Cr_2O_7^{2-} \rightarrow 2Cr^{3+}$$

3. Add H_2O to balance O atom & add H^+ to balance H atoms.

$$Cr_2O_7^{2-} \rightarrow 2Cr^{3+} + 7H_2O$$

$$14H^+ + Cr_2O_7^{2-} \rightarrow 2Cr^{3+} + 7H_2O$$

4. Add e^- to one side of each half-reaction to balance the charges on the half-reaction.

$$Fe^{2+} \rightarrow Fe^{3+} + \underbrace{e^-}_{\boxed{-1}}$$

$+2$ $\qquad$ $\underbrace{+3 \quad -1}_{+2}$

$$\underbrace{6e^- + 14H^+ + Cr_2O_7^{2-}}_{+6} \rightarrow 2Cr^{3+} + 7H_2O$$

$\boxed{-6}$ $\quad$ $+14$ $\quad$ -2 $\qquad$ $+6$ $\quad$ 0

$\underbrace{-6 + 14 - 2}_{+6}$

Tips: $1e^- \equiv -1$

5. **Multiply** each half-reaction by **an integer**, so that **number of electron lost** in one half-reaction **equals** the number **gained** in the other.

$$\left[\; Fe^{2+} \rightarrow Fe^{3+} + e^{-} \;\right] \times 6$$

$$6Fe^{2+} \rightarrow 6Fe^{3+} + 6e^{-}$$

$$6e^{-} + 14H^{+} + Cr_2O_7^{2-} \rightarrow 2Cr^{3+} + 7H_2O$$

6. **Combine** the two half-reactions and **cancel out** the species that appear **on both sides** of the equation.

Oxidation: $6Fe^{2+} \rightarrow 6Fe^{3+} + \cancel{6e^-}$

Reduction: $\cancel{6e^-} + 14H^+ + Cr_2O_7^{2-} \rightarrow 2Cr^{3+} + 7H_2O$

$$14H^+ + Cr_2O_7^{2-} + 6Fe^{2+} \rightarrow 6Fe^{3+} + 2Cr^{3+} + 7H_2O$$

acidic solution / acidic medium

Check that the number of atoms and the **charges** are balanced.

Left side | Right side

$(14 \times 1) + (-2) + (6 \times 2) = 24 = (6 \times 3) + (2 \times 3)$

Example: For Basic Solutions

- Balance the chemical equation in basic solution using the Ion–electron Method / Half–Reaction Method

$$Fe^{2+} \quad + \quad Cr_2O_7^{\,2-} \longrightarrow Fe^{3+} + Cr^{3+}$$

- Follows the steps in acidic medium(step 1 –step 6) followed by these additional steps;

1. **Add to both sides** of the equation the **<u>same number</u>** of **OH^-** as there are **H^+** $\Rightarrow$ to eliminate ALL H^+
2. **Combine** H^+ and OH^- to form H_2O.
3. **Cancel** any H_2O that you can.

7. **Add to both sides** of the equation the **same number** of **OH^-** as there are **H^+** $\Rightarrow$ to eliminate ALL H^+

Equation from acidic solution :

$$14H^+ + Cr_2O_7^{2-} + 6Fe^{2+} \rightarrow 6Fe^{3+} + 2Cr^{3+} + 7H_2O$$

In this case: **add 14 OH^-**

$$14OH^- + 14H^+ + Cr_2O_7^{2-} + 6Fe^{2+} \rightarrow 6Fe^{3+} + 2Cr^{3+} + 7H_2O + 14OH^-$$

8. **Combine** H^+ and OH^- to form H_2O.

$$\underbrace{14OH^- + 14H^+}_{} + Cr_2O_7^{2-} + 6Fe^{2+} \rightarrow 6Fe^{3+} + 2Cr^{3+} + 7H_2O + 14OH^-$$

$$14H_2O + Cr_2O_7^{2-} + 6Fe^{2+} \rightarrow 6Fe^{3+} + 2Cr^{3+} + 7H_2O + 14OH^-$$

9. **Cancel** any H_2O that you can.

$$\overset{7}{\cancel{14}}H_2O + Cr_2O_7^{2-} + 6Fe^{2+} \rightarrow 6Fe^{3+} + 2Cr^{3+} + \cancel{7}H_2O + 14OH^-$$

$$7H_2O + Cr_2O_7^{2-} + 6Fe^{2+} \rightarrow 6Fe^{3+} + 2Cr^{3+} + 14\,OH^-$$

basic solution / basic medium

Check that the number of atoms and the **charges** are balanced.

Left side | Right side

$$(0) + (-2) + (6 \times 2) = 10 = (6 \times 3) + (2 \times 3) + (-1 \times 14)$$

Acidic solution

EXAMPLE 4:

Reduction

$$MnO_4^- + Fe^{2+} \longrightarrow Mn^{2+} + Fe^{3+}$$

Oxidation

Answer

- **Reduction :** $MnO_4^- \longrightarrow Mn^{2+}$

$$MnO_4^- \longrightarrow Mn^{2+} + 4H_2O$$

$$\underbrace{MnO_4^- + 8H^+}_{-1 + 8 = +7} \longrightarrow \underbrace{Mn^{2+} + 4H_2O}_{2 + 0 = +2}$$

$$\boxed{MnO_4^- + 8H^+ + 5e^- \longrightarrow Mn^{2+} + 4H_2O}$$

- **Oxidation :** $Fe^{2+} \longrightarrow Fe^{3+}$

  $+2 \qquad\qquad +3$

  $Fe^{2+} \longrightarrow Fe^{3+} + e^{-}$

Balance the electron

$$MnO_4^{-} + 8H^{+} + 5e^{-} \longrightarrow Mn^{2+} + 4H_2O$$

$$Fe^{2+} \longrightarrow Fe^{3+} + e^{-}$$

$$5(Fe^{2+} \longrightarrow Fe^{3+} + e^{-})$$

$$\boxed{5Fe^{2+} \longrightarrow 5Fe^{3+} + 5e^{-}}$$

Combine two half-equations

$$MnO_4^- + 8H^+ + \cancel{5e^-} \longrightarrow Mn^{2+} + 4H_2O$$

$$5Fe^{2+} \longrightarrow 5Fe^{3+} + \cancel{5e^-}$$

$$\underbrace{MnO_4^- + 5Fe^{2+} + 8H^+}_{\boxed{-1 \ + \ 10 \ + 8 \ = +17}} \longrightarrow \underbrace{Mn^{2+} + 5Fe^{3+} + 4H_2O}_{\boxed{+2 \ + \ 15 \ + \ 0 \ = \ +17}}$$

☞ *Check for net charge and atoms*

EXAMPLE 5:

Basic solution

Reduction

$$MnO_4^- + C_2O_4^{2-} \longrightarrow MnO_2 + CO_3^{2-}$$

Oxidation

Answer

- Reduction : $MnO_4^- \longrightarrow MnO_2$

$$\underbrace{MnO_4^- + 4H^+}_{-1 + 4 = +3} \longrightarrow \underbrace{MnO_2 + 2H_2O}_{0}$$

$$\boxed{MnO_4^- + 4H^+ + 3e^- \longrightarrow MnO_2 + 2H_2O}$$

- **Oxidation :** $C_2O_4^{2-} \longrightarrow CO_3^{2-}$

$$C_2O_4^{2-} \longrightarrow 2CO_3^{2-}$$

$$\underbrace{C_2O_4^{2-} + 2H_2O}_{-2} \longrightarrow \underbrace{2CO_3^{2-} + 4H^+}_{-4\,+\,4\,=\,0}$$

$$\boxed{C_2O_4^{2-} + 2H_2O \longrightarrow 2CO_3^{2-} + 4H^+ + 2e^-}$$

$$2(MnO_4^- + 4H^+ + 3e^- \longrightarrow MnO_2 + 2H_2O)$$

$$3(C_2O_4^{2-} + 2H_2O \longrightarrow 2CO_3^{2-} + 4H^+ + 2e^-)$$

- $2MnO_4^- + \cancel{8H^+} + \cancel{6e^-} \longrightarrow 2\,MnO_2 + \cancel{4H_2O}$
- $3C_2O_4^{2-} + \overset{2}{\cancel{6}}H_2O \longrightarrow 6CO_3^{2-} + \overset{4}{\cancel{12}}H^+ + \cancel{6e^-}$

$$\boxed{2MnO_4^- + 3C_2O_4^{2-} + 2H_2O \longrightarrow 2\,MnO_2 + 6CO_3^{2-} + 4H^+}$$

☞ *Check for net charge and atoms*

$$2MnO_4^- + 3C_2O_4^{2-} + 2H_2O \longrightarrow 2MnO_2 + 6CO_3^{2-} + 4H^+$$

- Add the same number of OH^- as H^+ on both sides of equation

$$2MnO_4^- + 3C_2O_4^{2-} + 2H_2O + 4OH^- \rightarrow 2MnO_2 + 6CO_3^{2-} + \underbrace{4H^+ + 4OH^-}_{4H_2O}$$

$$2MnO_4^- + 3C_2O_4^{2-} + \cancel{2H_2O} + 4OH^- \longrightarrow 2MnO_2 + 6CO_3^{2-} + \underset{2}{\cancel{4}}H_2O$$

$$2MnO_4^- + 3C_2O_4^{2-} + 4OH^- \longrightarrow 2MnO_2 + 6CO_3^{2-} + 2H_2O$$

Keep in mind

Balancing Redox equation

$$MnO_4^- + C_2O_4^{2-} \longrightarrow MnO_2 + CO_3^{2-}$$

Separate the equation into two half reaction (Oxidation and Reduction)

↓

Balance atoms other than O and H

↓

Add H_2O to balance O atoms and add H^+ to balance H atoms

↓

Add electron to balance charge

↓

Combine two half-reactions and cancel out the species that appear on both sides of the equation.

❑ Difference between balancing redox reactions in acidic and basic medium

$$2MnO_4^{-} + 3C_2O_4^{2-} + 2H_2O \longrightarrow 2MnO_2 + 6CO_3^{2-} + 4H^{+}$$

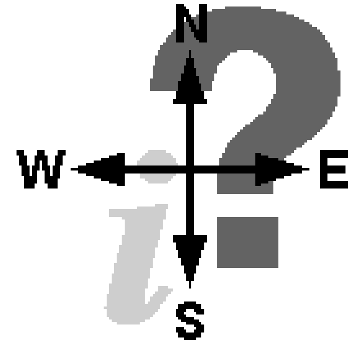
Acidic medium

Basic medium

Add $4OH^{-}$ on both side

$$2MnO_4^{-} + 3C_2O_4^{2-} + 4OH^{-} \longrightarrow 2MnO_2 + 6CO_3^{2-} + 2H_2O$$

EXERCISE 3:



Balance the following redox equation :

1. $Zn(s) + H^+(aq) \longrightarrow Zn^{2+}(aq) + H_2(g)$

2. $S_2O_3^{2-} + I_2 \longrightarrow S_4O_6^{2-} + I^-$

3. $C_2O_4^{2-} + MnO_4^- \longrightarrow CO_2 + Mn^{2+}$ (acidic solution)

4. $Cr_2O_7^{2-} + Fe^{2+} + H^+ \longrightarrow Cr^{3+} + Fe^{3+} + H_2O$

5. $CrO_2^- + ClO^- \longrightarrow CrO_4^{2-} + Cl^-$ (basic solution)

Answer

Balance the following redox equation :

1. $Zn(s) + 2H^+(aq) \longrightarrow Zn^{2+}(aq) + H_2(g)$

2. $2S_2O_3^{2-} + I_2 \longrightarrow S_4O_6^{2-} + 2I^-$

3. $5C_2O_4^{2-} + 2MnO_4^- + 16H^+(aq) \longrightarrow 10CO_2 + 2Mn^{2+} + 8H_2O$

(acidic solution)

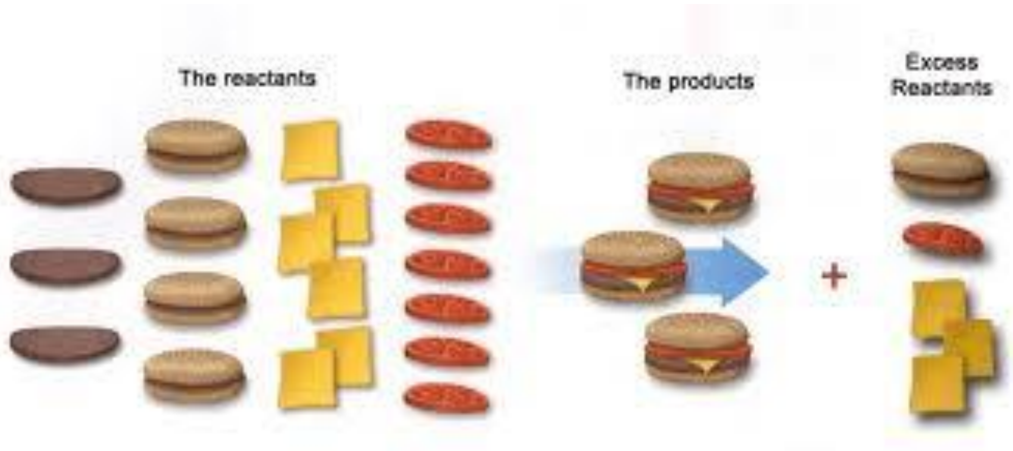
4. $Cr_2O_7^{2-} + 6Fe^{2+} + 14H^+ \longrightarrow 2Cr^{3+} + 6Fe^{3+} + 7H_2O$

5. $2CrO_2^- + 3ClO^- + 2OH^- \longrightarrow 2CrO_4^{2-} + 3Cl^- + H_2O$

(basic solution)

LIMITING REACTANT

- Reactant that is completely consumed in a reaction and limit the amount of products formed.
- Reactants used up first in a reaction
- Determine the amount of products formed

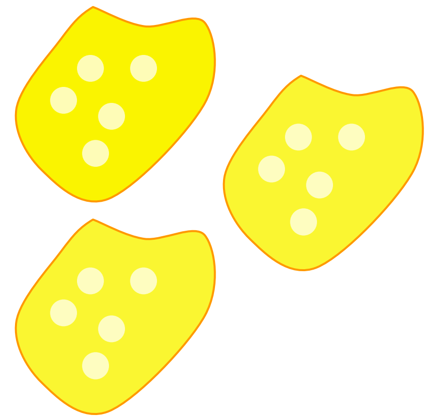


The Cheese Sandwich Analogy

9 slices of Bread

3 slices of cheese

reactants



Excess Reactant:
bread

Limiting reactant:
cheese

product

PERCENTAGE YIELD

- Percentage yield is the **percent of the actual yield of a product to its theoretical yield**

$$\boldsymbol{Percentage\ yield = \frac{actual\ yield}{theoretical\ yield} \times 100}$$

- The **amount of product** predicted by a **balanced equation** is the **<u>theoretical yield.</u>**

- The theoretical yield is never obtained because:

1. The reaction may undergo side reaction.
2. Many reaction are reversible.
3. There may be impurities in the reactants.
4. The product formed may react further to form other product.
5. It may be difficult to recover all of the product from the reaction medium.

- The **amount product actually** obtained in a reaction is the **<u>actual yield.</u>**

STOICHIOMETRIC CALCULATIONS

- Quantitative study of reactants and products in a chemical reaction.

$$xA + yB \longrightarrow zC + wD$$

reactants products

$$xA + yB \longrightarrow zC + wD$$

Determining Limiting Reactant

Stoichiometric Method

1. **Write** complete equation
2. **Calculate moles** of reactants, A and B.
3. **Calculate** amount of reactant B **required to react completely** with reactant A
4. **Compare** the amount of B required (needed) with that available (given) in the system.

If B (given) < B (needed)

B = limiting reactant

If B (given) > B (needed)

B = excess reactant ⇒ A = limiting reactant

EXAMPLE 6:

	$H_2\,(g)$ +	$F_2\,(g)$ →	$2HF\,(g)$
Initial amount: (given)	3 mol	2 mol	4 mol

➢ **Determine the limiting reactant:**

From the equation,

1 mol of H_2 needed 1 mol of F_2

3 mol of H_2 needed

$$\frac{3\ \cancel{\text{mol}\ H_2} \times 1\ \text{mol}\ F_2}{1\ \cancel{\text{mol}\ H_2}}$$

$$= 3\ \text{mol}\ F_2\ \text{(needed)}$$

$$H_2\,(g) \quad + \quad F_2\,(g) \quad \rightarrow \quad 2HF\,(g)$$

Initial amount: (given)	3 mol	2 mol	4 mol

mole of F_2 needed (3 mol) > mole of F_2 given (2 mol)

$\therefore$ Limiting reactant = F_2

Excess reactant = H_2

EXAMPLE 7:

For the following reaction :

$$2NH_3(g) + 3CuO(s) \longrightarrow N_2(g) + 3Cu(s) + 3H_2O(l)$$

18.0 g NH_3 and 90.0 g CuO are allowed to react.

[A_r Cu = 63.6, H = 1.0, N = 14.0, O = 16.0]

a) Determine the limiting reactant.

b) Calculate the mass of N_2 gas formed.

c) Determine the mass of the excess reactant remain after the completion of the reaction.

Answer

(a) Determine the limiting reactant.

- $number\ of\ moles = \dfrac{mass\ (g)}{molar\ mass\ (g\ mol^{-1})}$

- $\boldsymbol{n_{NH_3}} = \dfrac{18.0\ g}{17.0\ g\ mol^{-1}}$

  $= \mathbf{1.059}\ \boldsymbol{mol}\ (\boldsymbol{given})$

- $\boldsymbol{n_{CuO}} = \dfrac{90.0\ g}{79.6\ g\ mol^{-1}}$

  $= \mathbf{1.131}\ \boldsymbol{mol}\ (\boldsymbol{given})$

$$2NH_3(g) + 3CuO(s) \longrightarrow N_2(g) + 3Cu(s) + 3H_2O(l)$$

- From the equation:

$$2\ mol\ of\ NH_3\ react\ with\ 3\ mol\ of\ CuO$$

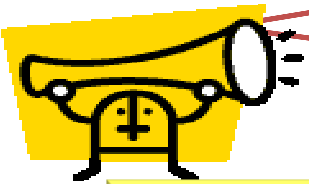
$$1.059\ mol\ of\ NH_3\ react\ with\ \frac{1.059\ mol\ of\ NH_3}{2\ mol\ of\ NH_3} \times 3\ mol\ of\ CuO$$

$$= 1.5885\ \text{mol CuO}\ \textbf{(needed)}$$

mole of CuO needed (**1.5885 mol**) > mole of CuO given(**1.131 mol**)

$\therefore$ **CuO is the limiting reactant**

(b) Calculate the mass of N_2 gas formed.



Limiting reactant limits the amount of product formed

$$2NH_3(g) + 3CuO(s) \longrightarrow N_2(g) + 3Cu(s) + 3H_2O(l)$$

[Tips: compare limiting reactant with product]

- **From the equation:**

3 mol of CuO produce 1 mol N_2

1.131 mol of CuO produce $\dfrac{1.131 \text{ mol CuO}}{3 \text{ mol CuO}} \times 1 \text{ mol } N_2$

$= 0.377 \text{ mol } N_2$

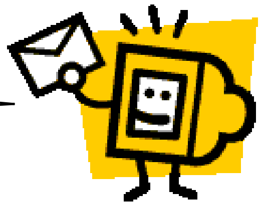
- Mass of N_2 formed $= n_{N2} \text{ x Molar mass } N_2$

$= 0.377 \text{ mol} \times 28 \text{ g mol}^{-1}$

$= \mathbf{10.56\ g}$

(c) Determine the mass of the excess reactant remain after the completion of the reaction.

Only the limiting reactant is completely consumed



$$2NH_3(g) + 3CuO(s) \longrightarrow N_2(g) + 3Cu(s) + 3H_2O(l)$$

- **From the equation:**

3 mol of CuO react with 2 mol of NH_3

1.131 mol of CuO react with $\dfrac{1.131 \text{ mol CuO}}{3 \text{ mol CuO}} \times 2 \text{ mol } NH_3$

$= \mathbf{0.754 \text{ mol } NH_{3\,(reacted)}}$

- Mass of NH_3 reacted $= n_{NH3}$ x Molar mass NH_3

$= 0.754 \text{ mol} \times 17 \text{ g mol}^{-1}$

$= 12.818 \text{ g}$

$\therefore$ Mass of NH_3 excess $= (18 - 12.818) \text{ g}$

$=$ **5.182 g**

EXAMPLE 8:

If 3.7 g sodium metal (Na) and 4.3 g chlorine gas (Cl_2) react to form NaCl, what is the theoretical yield? If 5.5 g NaCl was formed, what is the percentage yield?

[A_r Na = 23.0, Cl = 35.5]

Answer

- $$mole\ of\ Na = \frac{mass\ of\ Na\ (g)}{molar\ mass\ of\ Na(g\ mol^{-1})}$$

$$= \frac{3.7\ g}{23.0\ g\ mol^{-1}}$$

= **0.1609 mol Na (given)**

- $$mole\ of\ Cl_2 = \frac{mass\ of\ Cl_2(g)}{molar\ mass\ of\ Cl_2(g\ mol^{-1})}$$

$$= \frac{4.3\ g}{71.0\ g\ mol^{-1}}$$

= **0.0606 mol Cl_2 (given)**

From the equation: $2Na(s) + Cl_2(g) \rightarrow 2NaCl(s)$

$$2\ mol\ of\ Na\ react\ with\ 1\ mol\ of\ Cl_2$$

$$0.1609\ mol\ of\ Na\ react\ with\ \frac{0.1609\ mol\ of\ Na\ \times 1\ mol\ of\ Cl_2}{2\ mol\ of\ Na}$$

$= 0.0804$ mol Cl_2 (needed)

mole of Cl_2 needed (0.0804 mol) > mole of Cl_2 given (0.0606 mol)

$\therefore$ Cl_2 is limiting reactant

- From the equation: $2Na(s) + Cl_2(g) \rightarrow 2NaCl(s)$

$$1\ mol\ of\ Cl_2\ produce 2\ mol\ of\ NaCl$$

$$0.0606\ mol\ of\ Cl_2\ produce\ \frac{0.0606\ \cancel{mol\ of\ Cl_2} \times 2\ mol\ of\ NaCl}{1\ \cancel{mol\ of\ Cl_2}}$$

$$= 0.1212\ mol\ NaCl$$

- Mass of NaCl $= n_{NaCl}$ x Molar mass of NaCl

$= 0.1212\ \cancel{mol} \times 58.5\ g/\cancel{mol}$

$= 7.0902$ g NaCl (theoretical yield)

- $$\textbf{Percentage yield} = \frac{\textit{actual yield}}{\textit{theoretical yield}} \times 100$$

$$= \frac{5.5\ g}{7.0902\ g} \times 100$$

$$= \ 77.57\ \%$$

EXERCISE 4:

Hydrogen reacts with oxygen to produce water as in the reaction equation below:

$$2H_2(g) + O_2(g) \rightarrow 2H_2O(g)$$

Assume that before the reaction takes place, there is 20 g H_2 gas and 224 g O_2 gas.

i. Determine the limiting reagent in this reaction. Show your calculation.

ii. Determine the quantity in moles of H_2 and O_2 consumed in the reaction.

iii. Calculate the mass of H_2O produced at the end of the reaction.

iv. Calculate the mass of the excess reagent left after the reaction has completed.

Answer ii) 10, 5 iii) 180 g iv) 64 g

GLOSSARY

BIL	TERM	SYMBOL/FORMULA	DEFINE
1.	ISOTOPE	-	Two or more atoms of the same element having same proton number but different nucleon number. Or Two or more atoms of the same element having same number of protons but different number of neutrons.
2.	PROTON NUMBER	Z	Number of protons in the nucleus of an atom of an element.
3.	NUCLEON NUMBER	A	Total number of protons and neutrons in the nucleus of an atom of an element.
4.	RELATIVE ATOMIC MASS	A_r	A mass of one atom of an element compared to one twelfth mass of one atom of carbon-12 atom

GLOSSARY

BIL	TERM	SYMBOL/FORMULA	DEFINE
5.	**RELATIVE MOLECULAR MASS**	M_r	A **mass** of **one molecule of a compound compared to one twelfth mass of one atom of carbon-12 atom**
6.	**EMPIRICAL FORMULA**	-	Formula that shows the **simplest ratio of all elements** in a molecule.
7.	**MOLECULAR FORMULA**	-	**Total number of protons and neutrons in the nucleus of an atom of an element.**
8.	**SOLUTE**	-	Is the **substance** being dissolved and present in the **smaller amount**
9.	**SOLVENT**	-	Is the **substance** doing the solving and present in the **larger amount**

GLOSSARY

BIL	TERM	SYMBOL/FORMULA	DEFINE
10.	**SOLUTION**	-	Is a **homogenous mixture of 2 or more** substances that formed when amount of **solute** dissolves completely in a **solvent**
11.	**CONCENTRATION**	**-**	Is the **amount of solute** present in a given **quantity of solvent** in a solution
12.	**MOLARITY**	M @ mol L^{-1} @ mol dm^{-3}	The **number of mole** of **solute** dissolved in 1 L of solution.
13.	**MOLALITY**	m / mol kg^{-1}	The **number of moles of solute** per **1 kg of solvent** in a solution.
14.	**MOLE FRACTION**	X	The **ratio** of the **number of moles of one component** to the **total of number of moles present**.

GLOSSARY

BIL	TERM	SYMBOL/FORMULA	DEFINE
15.	**PERCENTAGE BY MASS**	$\%{}^{w}/_{w}$	The **ratio** of the **mass of a solute** to the **mass of the solution**, multiplied by 100
16.	**PERCENTAGE BY VOLUME**	$\%{}^{v}/_{v}$	The **ratio** of the **volume of a solute** to the **volume of the solution**, multiplied by **100**
17.	**LIMITING REACTANT**	-	Reactant that is **completely consumed** in a reaction and **limit the amount of products** formed.
18.	**PERCENTAGE YIELD**	-	The **percent of the actual yield of a product to its theoretical yield**