Democritus The concept of an atom as a sphere and the basic unit of all matter.

Lavoisier All atoms of the same element have the same mass.

Joseph Proust The Law of Definite Proportions.

(For a given compound, the elements always combine in the same proportion).

John Dalton The Law of Multiple Proportions.

(Two elements A and B can form different compounds by combining in different proportions).

- Atoms are small, discrete, indivisible pieces of matter.
- All elements are made up of particles called atoms.
- An element's atoms are <u>identical</u> in size, mass, & chemical properties.

J. J. Thomson ('Plum Pudding' Model)

Experiment: Cathode ray tube

Discover the electron and determine the electron's charge-to-mass ratio.

Millikan

Experiment: Oil drop

Determine the charge on an electron and find the mass of the electron as approximately $1/2000^{th}$ of the mass of a hydrogen atom.

Ernest Rutherford

Experiment: Gold foil

Discover the atom's nucleus.

- The atom contains a tiny, dense center called the nucleus.
- The nucleus has essentially the entire mass of the atom.
  - The electrons weigh so little they give practically no mass to the atom.
- The nucleus is positively charged.

– The amount of positive charge balances the negative charge of the electrons.

– The electrons are dispersed in the empty space of the atom surrounding the nucleus.

James Chadwick

Discover the neutrons.

Hydrgen and oxygen can react to form two compounds A & B with different chemical and physical properties. Compound A has an O:H mass ratio = 8:1 and compound B has an O:H mass ratio = 16:1. This observation is consistent with the law of

A) Definite proportions.
B) Energy conservation.
C) Multiple proportions.
D) Mass conservation.

Chapter 1: Matter & Measurements
Lesson 1: Classification & States of Matter

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* Introduction :

Q 1. What is Chemistry ?

A 1. Chemistry is the study of matter and its Changes.

Q 2. What is the matter ?

A 2. The matter is anything that takes up space and has mass.

Q 3. What is the matter made from ?

A 3. The matter is made of atoms and molecules.

Where

Atoms are submicroscopic particles that constitute the basic building blocks of matter.

Molecules are group of atoms held together by Chemical bonds.

(held = bonded)

Now

Chemistry is the science that seeks to understand the behavior of matter by studying the behavoir of its atoms and molecules.

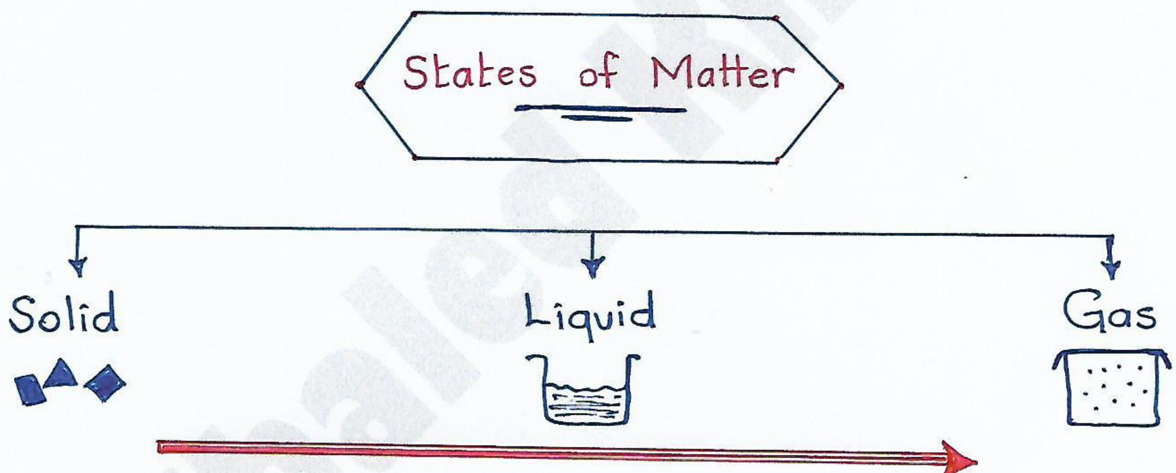
* Classification of Matter :

A) According to Composition :

Matter is classified according to the basic Components that make it up.

e.g. Carbon monoxide (CO) Contains Carbon atom and Oxygen atom held together by chemical bonding.

B) According to physical state :



State of matter is changed by varying The temperature.

In general;

Different structures of atoms and molecules lead to different properties of Matter.

Q.4 What are the differences between the three different states of matter ?

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A 4.

Comparison between the 3 states of matter :

Property	Solid	Liquid	Gas
1) Shape	Fixed shape (Rigid)	No fixed shape (shape of the Container)	No fixed shape (shape of whole Container)
2) Volume	Fixed volume	Fixed volume	No fixed volume (expands to fill the Container)
3) Motion of particles	Particles in fixed locations so, They Can only vibrate.	Particles Can freely move but with less degree as Compared to gas	Particles Can freely move Throughout the whole Container
4) Compressibility	Not easy to Compress	Not easy to Compress	Easy to Compress
5) Example	Ice	Liquid water	Water vapor

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Solid state

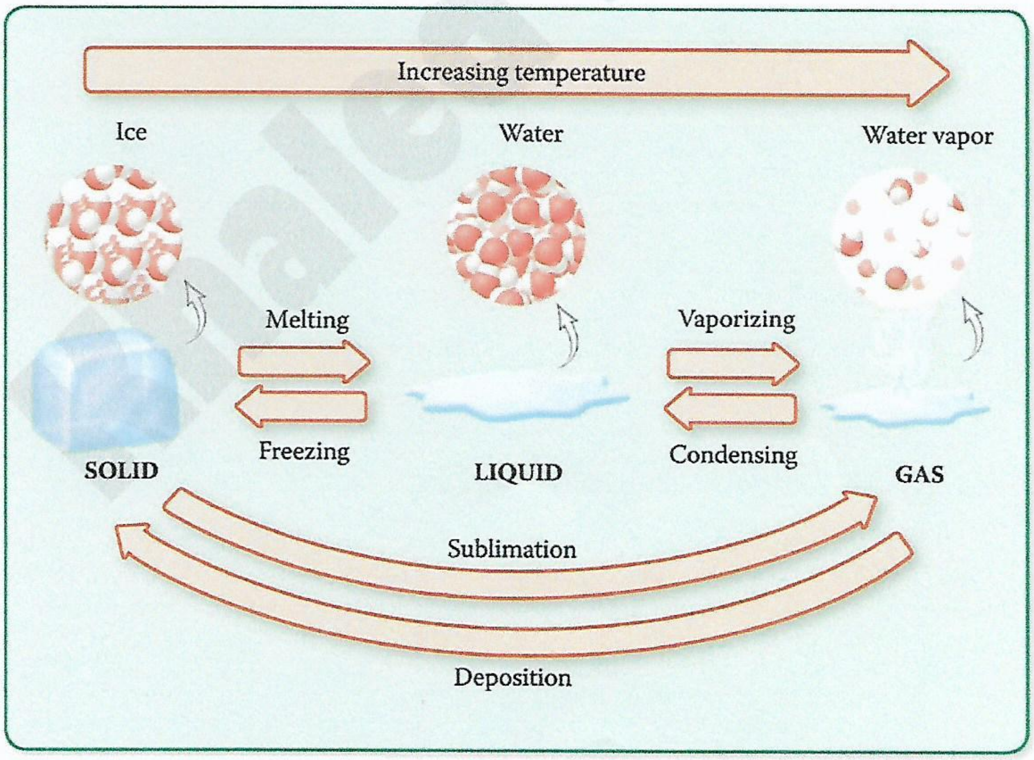
Crystalline

Atoms or molecules are arranged in patterns of long-range repeating order. e.g. Diamond

Amorphous

Atoms or molecules are not arranged in pattern of long-range repeating order. e.g. Glass

« Summary of state changes »



* Classification of Matter :

Matter

* Pure Substances *

They are composed of one substance and can't be physically separated.

Element

It can't be broken down by chemical reactions.

e.g. He , O_2 ,Fe

Compound

It can be chemically separated into its elements.

e.g. water can be chemically separated to H & O - atoms (water = H_2O)

* Mixtures *

They are composed of more than one substance and can be physically separated.

Homogeneous

It has uniform properties.

e.g. Tea - Sugar solution; air

Heterogeneous

It doesn't has uniform properties throughout.

e.g. Sandy water

End.

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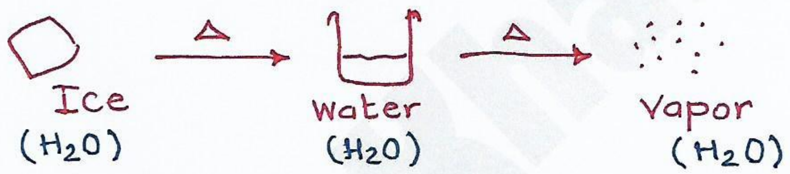
Chapter 1: Matter & Measurements
Lesson 2: physical & Chemical Changes and properties

Q1. What is the physical Change?

A1.

It is the Change that alters only the state (appearance) of matter but without change in its Composition.

e.g.



Q2. What is the Chemical Change?

A2.

It is the Change that alters the Composition of matter.

e.g.

Burning of fuel: $C_3H_{8(g)} \xrightarrow[\Delta]{(O)} CO_{2(g)} + H_2O_{(g)}$

Different molecules

As an evidence for the Chemical Change, the appearance and behavoir of matter are changed.

e.g.

Rusting of iron, a permanant color change can indicate that a chemical change is happened.

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Q 3. What is the difference between physical and chemical changes?

A3.

	physical Change	Chemical Change
Molecules	not changed	Changed
Composition	not changed	Changed

* Properties of matter :

Properties of Matter

physical properties

The property that a substance displays without change in its Composition. examples:

- Boiling point.
- Melting point.
- Density.
- Viscosity.
- Color.
- odour.

Chemical properties

The property that a substance displays with changing in its Composition through chemical Change. examples:

- Burning in air.
- Heat decomposition.
- Reaction with another substance
- flammability. ....

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* Matter and measurements :

* Work :

It is the action of Force through Distance.

* Kinetic energy : (KE)

It is the ability to do work (energy required to do work).

* Potential Vs. Kinetic energies :

Energy

Potential energy (PE)

The energy stored in the body by its position.

e.g.
Stored battery, raised ball

Kinetic energy (KE)

The energy that the matter possesses due to its motion.

→ $KE_{(Solid)} < KE_{(Liquid)} < KE_{(Gas)}$

→ As temperature increases, as KE increases.

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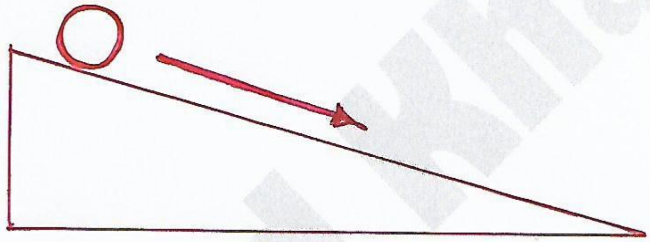
* Law of Conservation of Energy :

"Energy can be converted between potential and kinetic energy.
i.e. Energy is neither created nor destroyed.

In other words, Energy has many forms (Light, thermal, chemical, -------) and it can be transformed from one form to another.



Max PE
Min KE

Min PE
Max KE

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Chapter 1: Matter & Measurements

Lesson 1: Units of Measurements & Density Calculations

Q. What are the units of measurements ?

A.

They are the standard quantities that used to specify the measurements.

« Common Unit Systems »

- Imperial system
- Metric system

Scientists use the international system of units (SI) which is based on the metric system.

Measurement	Metric	SI
Length	meter (m)	meter (m)
* Volume	Liter (L)	Cubic meter (m^3)
* Mass	gram (g)	Kilogram (Kg)
* Temperature	Celsius (°C)	Kelvin (K)
Time	Second (s)	Second (s)

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N.B.

* Length :

Both metric and SI systems use meter (m).

but for small distances Centimeter can be used.

2.54 Cm = 1 in

1 m = 100 Cm

1 m = 39.4 in

1 m = 1.09 yd

* Mass & Weight :

Mass is the measure of the quantity of matter within the object.

while the weight is the measure of the gravitational pull on its matter.

1 Kg = 2.21 lb ; Kg = 1000 g or g = $\frac{1}{1000}$ Kg

* Volume :

It is the space that occupied by matter.

1 L = 1000 mL

1 L = 1.06 qt

946 mL = 1 qt

1000 L = 1 m^3

* Time :

1 day = 24 h

1 h = 60 min

1 min = 60 s

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* Temperature :

There are 3 Common units for temperature.

Celsius	Fahrenheit	Kelvin
(°C)	(°F)	(K)

• example :

Freezing (melting) point of water is 0°C , 32°F or 273 K.

Temperature - Units interconversions :

°C Vs. °F

$$°F = (1.8 \times °C) + 32$$

or

$$°C = \left(\frac{°F - 32}{1.8}\right)$$

°C Vs. K

$$K = °C + 273.15$$

or

$$°C = K - 273.15$$

Q. Convert 350 °F to °C

$$\because \; °C = \frac{°F - 32}{1.8} = \frac{350 - 32}{1.8} = 177\,°C$$

Q. Convert 298 K to °C

$$\because \; °C = K - 273.15 = 298 - 273.15 = 25\,°C$$

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* Prefixes & Prefix multipliers :

To increase or decrease the size of unit, some prefixes (power of 10) are used.

Prefixes that increase	Prefixes that decrease
Kilo (k) = 1000 or 10^3	deci (d) = 10^{-1}
mega (M) = 1000000 or 10^6	Centi (c) = 10^{-2}
giga (G) = 10^9	milli (m) = 10^{-3}
tera (T) = 10^{12}	micro (μ) = 10^{-6}
peta (P) = 10^{15}	nano (n) = 10^{-9}
	Pico (p) = 10^{-12}
	femto (f) = 10^{-15}

Now;

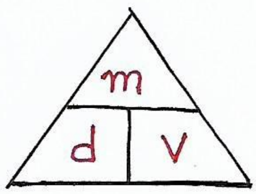
* Density:

It is the amount of matter in a substance per unit volume.

d = density

m = mass

V = volume



* Units of density are :

g/L for gases ; g/mL for Liquids ; g/cm^3 for Solids

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Q. A 0.258 g Sample of HDL has volume of 0.215 cm^3 what is the density, in g/cm^3, of the HDL sample?

A.

Given: 0.258 g HDL ; 0.215 cm^3 HDL.

Need: Density of HDL.

$$\because \text{Density} = \frac{\text{Mass}}{\text{Volume}} = \frac{0.258 \text{ g}}{0.215 \text{ cm}^3} = 1.20 \text{ g/cm}^3.$$

N.B. Relative density (to water)

Substance that has greater density than water $\Rightarrow$ Sinks but that with less density, $\Rightarrow$ floats over the water.

N.B. Density Calculations by volume displacement

In some Cases; the volume of object can be calculated by adding the object to water in graduated cylinder.

Q. What is the density of 48.0 g sample of metal if the level of water in a graduated cylinder rises from 25.0 mL to 33.0 mL after the metal is added?

A.

Volume of metal (by displacement) = 33.0 − 25.0 = 8 mL

$\because$

$$d = \frac{m}{V} = \frac{48.0 \text{ g}}{8.0 \text{ mL}} = 6.0 \text{ g/mL} \quad \text{or} \quad 6.0 \text{ g/cm}^3$$

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Chapter 2: Atoms, Molecules, ions & Periodicity

Lesson 2: Atomic theory & atomic structure

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* Law of Conservation of matter: "Lavoisier"

stated that:

"Matter is neither created nor destroyed in chemical reactions".

i.e.

Total mass of reactants = Total mass of products

or

Total number of reacting atoms = Total number of produced atoms.

* Atomic theory of matter " Dalton's theory ":

Dalton stated that:

- Atoms are small, discrete, indivisible pieces of matter.
- All elements are made up of particles called "atoms"
- An element's atoms are identical in size, mass, and chemical properties.
- Molecules (Compounds) are formed when two or more elements combine.
- Molecules are simple whole-number ratios of the combined elements

* Thomson's experiment:

Thomson conducted an experiment in which −ve charged electrode (cathode) is placed near to magnetic field and detector is placed at the other side to detect the "electron"
so,

Thomson's experiment confirms the presence of the electron.

* MilliKan oil experiment:

This experiment conducted to measure the charge of electron.

Result:

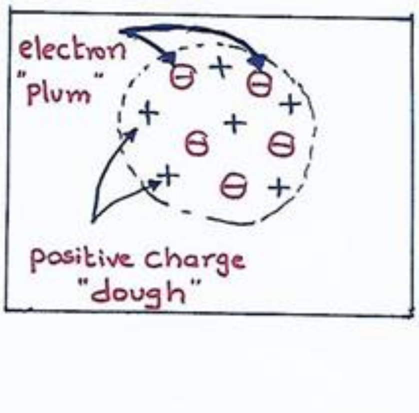
electron charge : -1.60×10^{19} C i.e. −ve charged.

electron mass : 9.1×10^{-28} g i.e. so small

* Plum - pudding Model by thomson :

Thomson proposed that :

The atom is composed of positive cloud of matter in which the electrons are embedded.



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* Rutherford's Gold Foil Experiment:

Conclusions of the experiment:

- Atom contains a tiny, dense center called "nucleus".
- the nucleus has essentially the mass of the atom.

* No mass for electrons "so little".

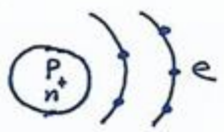
- The nucleus is positively charged and the amount of positive charge within the nucleus equals to the negative charges of electrons.
- Rutherford's student "James Chadwich" developed the nuclear theory and proposed that there are neutral particles within the nucleus called "neutrons".

Finally;

⇒ Atoms (particles) are composed of three subatomic particles:

Subatomic Particle	Charge	Mass (g)	Mass (amu)
* Proton	+1	1.67×10^{-24}	1
* Neutron	0	1.67×10^{-24}	1
* electron	−1	9.11×10^{-28}	Negligible

protons and neutrons are located within the nucleus (~ mass of atom) but the electrons are moving outside around the nucleus.



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« Atomic Number » Z

It is the number of protons inside the nucleus that determines the element's identity.

« Isotopes »

Some elements have atoms of different mass only because these atoms differ in the number of neutrons.

i.e. Same atomic number but different mass number

e.g.

$^{13}_{27}Al$

13 → Atomic number (Z) = No. of protons = No. of electrons.

27 → Mass number (A) = No. of protons + No. of neutrons.

example of isotopes :

	Carbon - 12 ;	Carbon - 13 ;	Carbon - 14
	$^{12}_{6}C$	$^{13}_{6}C$	$^{14}_{6}C$
Protons :	6 p^+	6 p^+	6 p^+
electrons :	6 e^-	6 e^-	6 e^-
Neutrons :	6 n	7 n	8 n

where ; $\text{No. of neutrons} = \overset{(A)}{\text{Mass No.}} - \overset{(Z)}{\text{Atomic No.}}$

* isotopes of hydrogen :

	$^{1}_{1}H$;	$^{2}_{1}H$;	$^{3}_{1}H$
protons = electrons	1	1	1
Neutrons	0	1	2
	Hydrogen	Deuterium	Tritium

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Q. Mention the numbers of protons, electrons & neutrons in:

a. $^{32}_{16}S$ $P = 16$; $e = 16$; $n = 32 - 16 = 16$

b. $^{65}_{29}Cu$ $P = 29$; $e = 29$; $n = 65 - 29 = 36$

c. $^{240}_{92}U$ $P = 92$; $e = 92$; $n = 240 - 92 = 148$

Q. What are the ions ?

Ions are atoms (or group of atoms) with positive (+) or negative (−) Charge and they are formed by losing or gaining electron (s).

e.g.

Losing: $_{11}Na^{0} \longrightarrow e^{-} + {}_{10}Na^{+}$ Cation

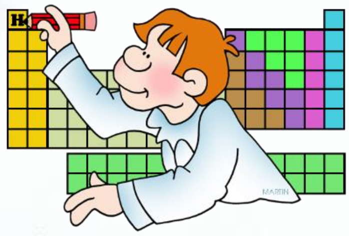
Gaining $_{17}Cl + e^{-} \longrightarrow {}_{18}Cl^{-}$ Anion

****

Ch. 2 : lesson 2 : "Periodic Table"

"Mendeleev" developed the first periodic table. He proposed that:

* When the elements are arranged in order of increasing atomic mass, Certain sets of properties recur periodically.
* Some elements have similar physical and chemical properties, he arranged them in Columns.



"Modern Periodic Table"

* The elements are arranged from left to right in increasing Atomic Number (no. of protons Z) rather than atomic mass as Mendeleev suggested.
* Rows in periodic table are Called "Periods"; while
* Columns in periodic table are Called "groups or families of elements with similar properties.
* The modern periodic table is organized in 8 main groups with letter A : 1A → 8A.
* There are series of elements lie between the main group 2 and 3, are called "Transition elements" with letter B.

« Classification of elements »

* Elements, in the periodic table, are classified into:

- Metals:

They lie on the lower left side and midde of the periodic table, they are:

- Good Conductors of heat and electricity.
- Malleable and ductile (shaped in sheets and wires).
- Shiny.
- losing electron(s) to form cations (+ve charged ions).
- 75% of the elements in the periodic table.
- All are solids except mercury Hg is liquid.

- Nonmetals:

They lie on the upper right side of the periodic table. They are:

- gases, Liquids, or solids.
- poor Conductors of heat and electricity.
- gaining electron(s) to form Anions (-ve Charged ions).
- not ductile and not malleable.

- Metalloids: (semimetals)

- They lie along zigzag line between metals and nonmetals, so
- they have mixed properties. They are solids
- They are poor of Conductor of heat
- They are known as semiconductors.

* Ions of the main group elements :

In general;

* AlKali metals (Group 1A) are very reactive and have a tendency to lose One electron and form +1 ions.
* AlKaline earth metals (Group 2A) are reactive but less than group 1A, they have tendency to lose two electrons and form +2 ions.
* Halogens (Group 7A) are very reactive and tend to gain One electron and form −1 ions.
* Noble gases or inert gases (Group 8A) are unreactive so, they have no tendency to lose or gain electrons.

In general;

* Main group elements that form cations (metals) carry charge equals to the group number.

e.g. Na lies in group 1A, so, forms cation with one +ve charge.
Mg " " " 2A, " " " " two " " .
Al " " " 3A; " " " " Three " " .

Thus; Na^+, Mg^{2+}, Al^{3+}

* Main group elements that form Anions (nonmetals) carry charge equals to (Group No. − 8)

e.g. Cl lies in group 7A, so forms anion with one −ve charge.
Charge of anion = 7 − 8 = −1 Cl^-

* Transition elements form various ions with different Charges so, No rule to predict.

Finally: metals and nonmetals tend to lose or gain electron(s) to reach the electronic configuration of the nearest noble gas.

* Atomic Mass and Isotopes:

The atomic mass of an element is not a simple whole number of atomic mass units (amu) but it is an average mass of all of its atoms. (isotopes).

Atomic Mass $= \sum$ (Natural abundance of isotope X Mass of isotope).

e.g.

Q. Calculate the atomic weight of Mg isotopes; Given that
79.0% Mg-24 ; 10.0% Mg-25 ; 11.0% Mg^{26}

A.

Firstly: Convert Percent $\xrightarrow[(\div 100)]{\text{to}}$ Decimal

Thus; Atomic weight $= (0.79 \times 24) + (0.10 \times 25) + (0.11 \times 26) = 24.3$

Q. Calculate atomic weight of Cl isotopes; Given that
75.77% Cl-35 and 24.23% Cl-37 ?

Ch.2: Lesson 6: Mole & Molar Mass

* Mole :

To Calculate the amounts of substances in Chemical reactions, Chemists use the standard quantity Called "Mole".

Where :

1 mole of any substance Contains Avogadro's number (6.022×10^{23}) of the particles of this substance (atoms or molecules).

So,

$$\boxed{1 \text{ Mole} = 6.022 \times 10^{23} \text{ particles}} \quad \ldots. (1)$$

In periodic table; each element has molar mass (e.g. Carbon has atomic mass 12); it means the mass of element per one mole.

So,

$$\boxed{\text{Moles} = \frac{\text{Mass}}{\text{Molar Mass}}} \quad \ldots\ldots (2)$$

where;

Mass : the mass given for element or Compound

Molar Mass : is the atomic mass of element (from Periodic table) or Molecular mass of Compound (Sum of atomic masses of all molecule atoms).

Examples :

Q1. How many Mg atoms are in 0.200 g ?

A.

Step 1 : Convert mass (g) to moles (use Eq. 2)

$\because \quad \text{Moles} = \dfrac{\text{Mass}}{\text{Molar Mass}} = \dfrac{0.200\ g}{24.31} = 8.23 \times 10^{-3}\ mol$

From Periodic Table. $^{24}_{12}Mg$

Step 2: Convert moles to atoms:

(Eq. 1) $\Rightarrow$ 1 mole $= 6.022 \times 10^{23}$ atoms

Thus:

$$1 \text{ mole} \longrightarrow 6.022 \times 10^{23} \text{ atoms}$$
$$8.23 \times 10^{-3} \text{ mol} \longrightarrow \ ? \text{ atoms}$$

$\therefore$ No. of Mg atoms $= (8.23 \times 10^{-3}\ mol)(6.022 \times 10^{23}\ atoms/mol)$
$= 4.95 \times 10^{21}$ atoms

Q2. How many H_2O molecules are in 5 g?

A.

Step 1: Convert mass to moles: $\quad \text{moles} = \dfrac{\text{Mass}}{\text{Molar Mass}}$

but for Compounds, you need to Calculate Molar Mass (Molecular Mass);

Molecular Mass of $H_2O = (2 \times 1) + (1 \times 16) = 18\ g$

Remember Well:

$$\text{Molecular Mass} = \sum (\text{No. of element} \times \text{Atomic Mass of element})$$

(From Periodic Table)

Now; $\quad \text{Moles} = \dfrac{\text{Mass}}{\text{Molar Mass}} = \dfrac{5}{18} = 0.278\ mol$

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Step 2 : Convert moles to molecules

$\because$ 1 mole $\longrightarrow$ 6.022×10^{23} molecules
0.278 mol $\longrightarrow$? molecules

$\therefore$ Molecules of water $= (0.278 \text{ mol})(6.022 \times 10^{23} \text{ molecules})$
$= 1.67 \times 10^{23}$ molecules.

Q.3 How many grams of CO_2 are in 6.75×10^{22} molecules of CO_2 ?

A.

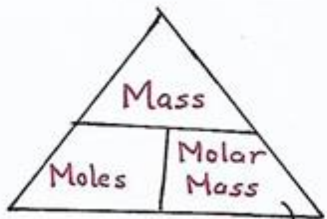
Here, we will do the opposite ;

step 1 : Convert molecules to moles :

1 mole $\longrightarrow$ 6.022×10^{23} molecules
From given ? mole $\longrightarrow$ 6.75×10^{22} molecules

$$\therefore \text{Moles } (CO_2) = \frac{6.75 \times 10^{22}}{6.022 \times 10^{23}} = 1.12 \times 10^{-1} \text{ moles } CO_2$$

Step 2 : Convert moles to mass



from Periodic Table

$\because$ Mass (CO_2) = Moles X Molar Mass

but: Molar Mass $(CO_2) = (1 \times 12) + (2 \times 16) = 44$ g/mol

$\therefore$ Mass $(CO_2) = (1.12 \times 10^{-1})(44) = 4.93$ g #

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* Problems :

Q.1 How many moles of Cl_2, are present in 71.0 g Cl_2 ?

A.

$\because$ Relation between moles & Mass is: $\text{Moles} = \dfrac{\text{Mass}}{\text{Molar Mass}}$

so,

firstly, Calculate Molor mass of $Cl_2 = 2 \times 35.5 = 71$ g/mol

so,

$$\text{Moles}(Cl_2) = \frac{71.0 \text{ g}}{71 \text{ g/mol}} = 1 \text{ moles}$$

Q.2 How many atoms of Al, are present in 54 g Al ?

A.

Hint: to Calculate No. of atoms, you have to Calculate moles.

$$\because \text{Moles}(Al) = \frac{\text{Mass}}{\text{Molor Mass}} = \frac{54}{27} = 2 \text{ moles.}$$

(Molar Mass of Al from periodic table is 27 g/mol)

but from Avogadro : $\because$ 1 mole $\longrightarrow 6.022 \times 10^{23}$ atoms Al

$\therefore$ 2 mole $\longrightarrow$?

$$\text{No. of atoms (Al)} = 2 \times 6.022 \times 10^{23} = 12.044 \times 10^{23} \text{ atoms}$$

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Ch.2 : Lesson 7 : Electron Configuration

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* Introduction:

It is very important to Know the number and the arrangement of the electrons in the atom to predict how the element will react.

* Quantum Mechanics:

The electron's location in the atom is described by four quantum numbers n, l, m_l, and m_s.

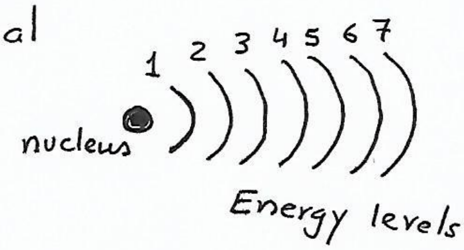
where; according to: Pauli exclusion principle

« No two electrons in an atom can have the same four quantum numbers »

1. Principal quantum number (n):

It indicates the electron's principal energy level. (distance from nucleus).

$n = 1 \longrightarrow 7$



2. Angular quantum number (l):

It indicates the electron's orbital type s, P, d, F. $l = 0 \longrightarrow n-1$

l-value	0	1	2	3
Orbital	s	p	d	f

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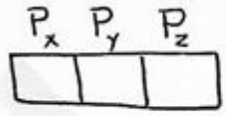
3. Magnetic quantum number: m_l

It indicates the orientation of orbital in space.

$m_l : -l \rightarrow 0 \rightarrow +l$

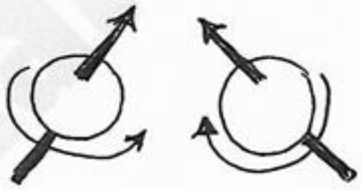
e.g.

if $l = 1$; $m_l = -1, 0, +1$



4. Spin quantum number: m_s

It indicates the direction of electron spin in its orbital.



$m_s = +\frac{1}{2}$ or $-\frac{1}{2}$

$\boxed{\uparrow\downarrow}$

Again,

Pauli exclusion principal

"No two electrons can have the same four quantum numbers".

« Electronic Configuration »

From the above quantum mechanics, there are four different types of orbitals in the atom : S , P , d , F .

Remember Well; the max number of electrons in these orbitals is :

$$S^2 \; ; \; P^6 \; ; \; d^{10} \; ; \; F^{14}$$

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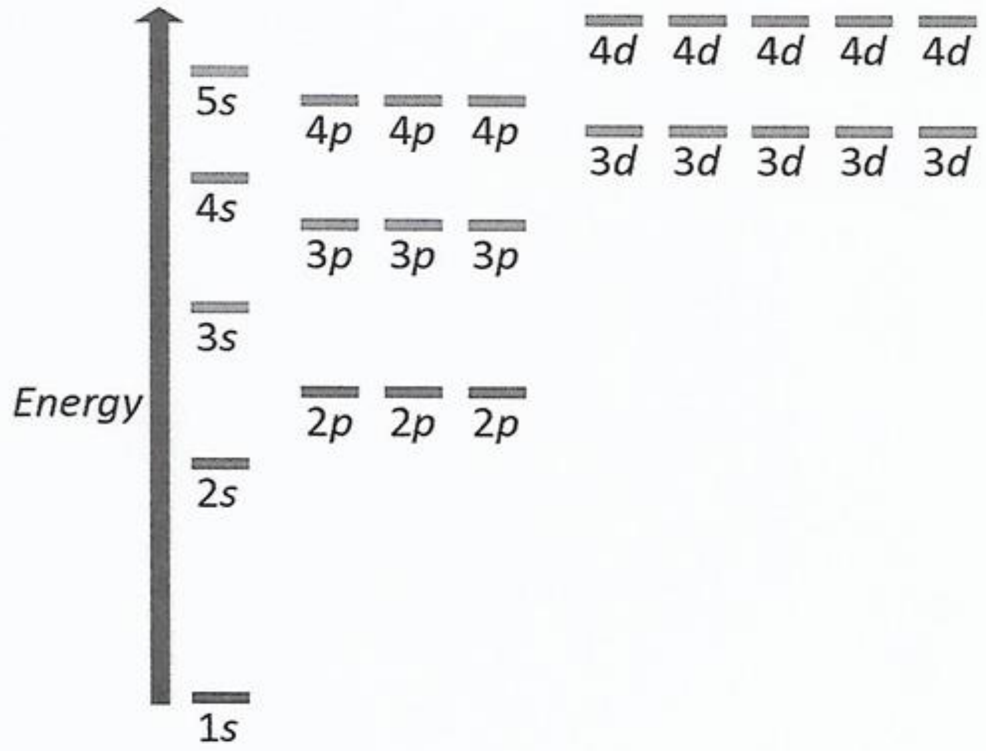
* Energy ordering of electrons :

"Aufbau principle" or "Building up principle"

The electrons fill the orbitals in order of increasing energy, starting with the lowest energy level orbital available.



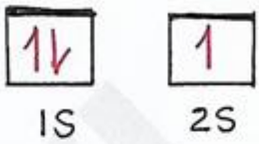
e.g.

- $^{2}_{4}He$ has 2 electrons ; so it electronic Configuration : $1S^2$
- ^{3}Li has 3 electrons ; " " " " : $1S^2\,2S^1$
- ^{7}N has 7 electrons ; " " " " : $1S^2\,2S^2\,2P^3$
- ^{11}Na has 11 electrons ; " " " " : $1S^2\,2S^2\,2P^6\,3S^1$

* Orbital diagrams

For example for ^{3}Li ; its electronic Configuration $1S^2\ 2S^1$

we Can use the orbital diagram :



where ;

each box represents one orbital and

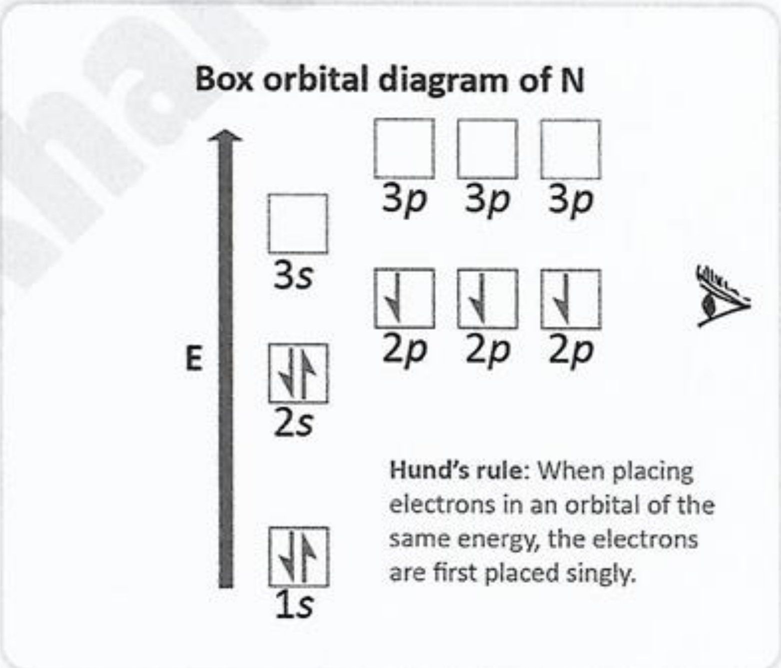
half_arrow represents one electron (↿)

"Hund's rule"

"When placing electrons in an orbitals of same energy, the electrons are first placed singly"

e.g.

For nitrogen ^{7}N has electronic Configuration : $1S^2\ 2S^2\ 2p^3$



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* Valence electrons :

The sum of electrons that are present in the outer energy level of the atom.

example ;
Sodium atom ^{11}Na has electronic configuration :

$$^{11}Na : 2 , 8 , 1$$

$$\underbrace{1s^2 \ 2s^2 2p^6}_{\text{Core electrons (10)}} \ \underbrace{3s^1}_{\text{Valence electrons (1)}}$$

The valence electrons participate in bonding, forming cations (if lost) and anions (if gained).

example ; Silicon atom ^{14}Si has electronic configuration

$$^{14}Si : 2 , 8 , 4$$

$$\underbrace{1s^2 \ 2s^2 2p^6}_{\text{Core electrons (10)}} \ \underbrace{3s^2 \ 3p^2}_{\text{Valence electrons (4)}}$$

We can simply substitute the core electrons with the nearest noble gas (of same number of electrons)

So,

^{11}Na : $1s^2 2s^2 2p^6 3s^1$ or $[Ne] 3s^1$

^{14}Si : $1s^2 2s^2 2p^6 3s^2 3p^2$ or $[Ne] 3s^2 3p^2$

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* Electron Configuration and Ions :

During the chemical reactions, the electrons are removed from or added to the Valence shell (outermost energy level) forming cations or anions respectively.

Ions

Cations

- are +ve charged atoms
  (No. of electrons < No. of protons)
- Metals form cations

e.g.
Sodium : ^{11}Na : $1s^2\ 2s^2 2p^6\ 3s^1$
$[Ne]\ 3s^1$

Na atom loses one electron to be stable (like $[Ne]^{10}$).

Na^+ : [Ne]

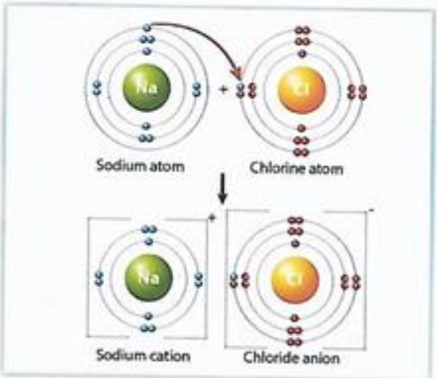
Anions

- are −ve charged atoms
  (No. of electrons > No. of protons)
- Nonmetals form anions.

e.g.
Chlorine: ^{17}Cl : $1s^2\ 2s^2 2p^6\ 3s^2\ 3p^5$
$[Ne]\ 3s^2\ 3p^5$

Cl atom gains one electron to be stable (like $[Ar]^{18}$).

Cl^- : $[Ne]\ 3s^2 3p^6$ or [Ar]



Ch.2 : Lesson 8 : Periodic Trends

* Introduction :

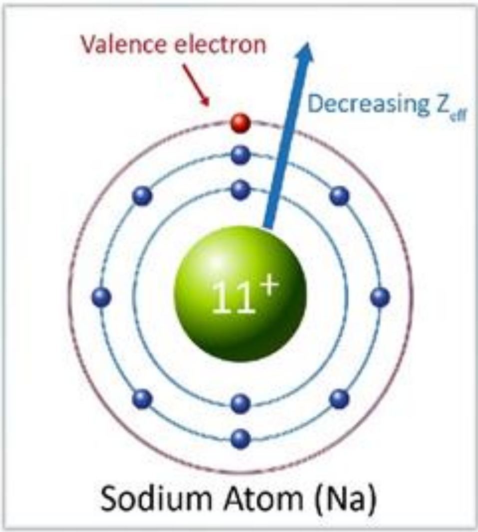
The number of valence electrons (Group No.) in an element determines its reactivity.

« Effective Nuclear Charge » Z_{eff}

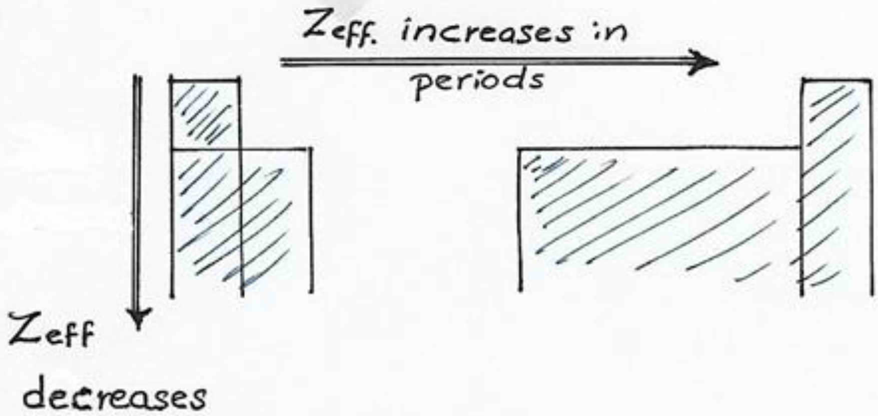
It is the pulling force an electron "feels" from the protons in the nucleus.

The Closer electrons to nucleus feel a greater pulling force, so, more energy is needed to remove.

Also, the electrons in the valence shell have less Z_{eff} so, they are easily removed.



In general; Z_{eff} increases across the periods and decreases going down the groups.



* Nuclear Charge Shielding effect :

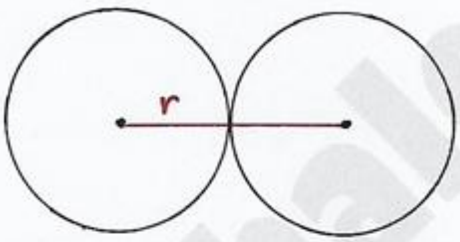
The inner (Core) electrons in an atom shield the outer electrons from the positive nuclear charge.

but Z_{eff} increases across the period due to the incomplete shielding by core electrons in subshells.

shielding ability of subshell : $s > p > d > f$

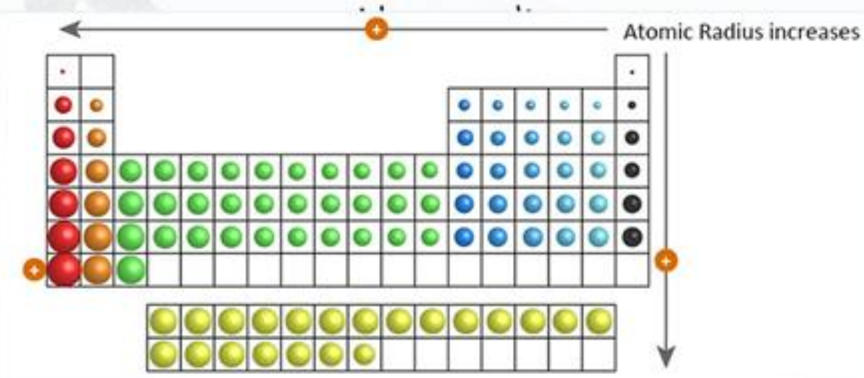
* Atomic radius :

Because the atoms don't have a definite edge, so the atomic radius is calculated when the atom is bounded to another atom.



N.B.

Atomic radius is ½ the distance between two bonded atoms.



Atomic radius is highly affected by Z_{eff}.

* Ionization energy : (IE)

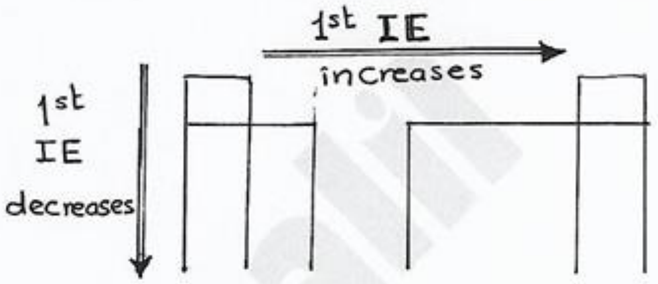
IE is the energy required to remove an electron from an atom in the gas phase. It is endothermic (absorbed).

The 1^{st} IE increases across the period and decreases going down the group.

$$M_{(g)} + E \longrightarrow M^{+}_{(g)} + e^{-} \quad ; \Delta E = +ve$$

Q. Explain why noble gases have max ionization energy?

A. Noble gases have completed energy levels with electrons so, they are very stable, so not easy to remove electron from their orbitals.



$$1^{st}\ IE \propto \frac{1}{\text{Atomic Radius}}$$

* Electron Affinity: (EA)

It is the change in energy (ΔE) when an atom gains an electron to form

$$A_{(g)} + e^- \longrightarrow A^-_{(g)} \quad ; \ \Delta E = +ve \text{ or } -ve$$

Electron affinity (ΔE) depends on the e^- repulsions and the volume of the atom.

Across the period, As Z_{eff} increases as EA increases and in groups Z_{eff} decreases down the group so, EA decreases.

$$EA \propto Z_{eff}$$

EA decreases

Electron Affinities (kJ/mol)

1A	2A	3A	4A	5A	6A	7A	8A
H −73							He >0
Li −60	Be >0	B −27	C −122	N >0	O −141	F −328	Ne >0
Na −53	Mg >0	Al −43	Si −134	P −72	S −200	Cl −349	Ar >0
K −48	Ca −2	Ga −30	Ge −119	As −78	Se −195	Br −325	Kr >0
Rb −47	Sr −5	In −30	Sn −107	Sb −103	Te −190	I −295	Xe >0

Increasing EA

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* Exceptions:

* EA (nonmetals) > EA (metals)
  form anions — form cations

* EA is greatest for halogen group.
  So, halogens form highly stable anions.

* EA of Fluorine < EA of Chlorine
  because F atom has smaller volume so, greater electron repulsion.

* EA (Group 1) > EA (Group 2)
  because group 2 elements have a full valence shell (s).

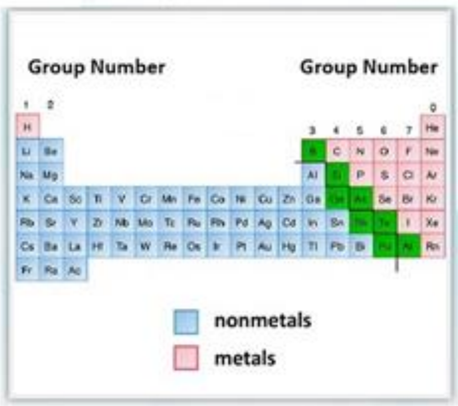
* EA (Noble gases) $\simeq$ 0
  because noble gases (Group 8) have full valence shells so, they are quite stable energetically.

* Properties of Metals & Nonmetals:

Metallic character of an element relates to how closely its properties match the typical properties of metal.

Metallic Character decreases

Metallic Character increases



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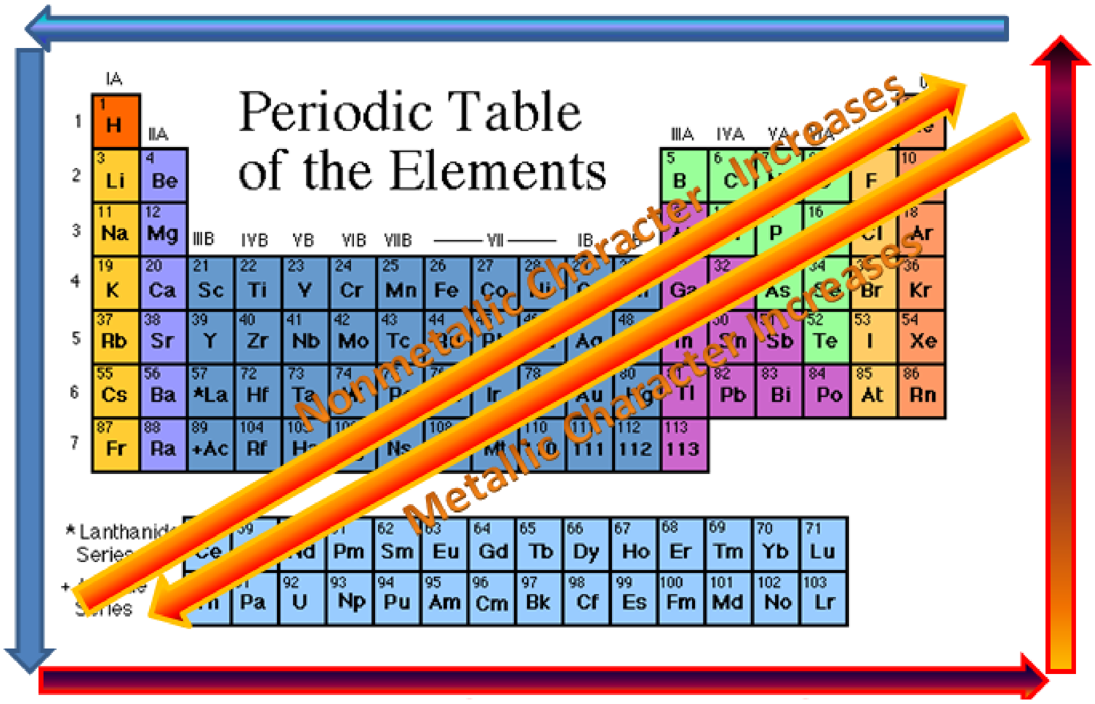
قسم الكيمياء - كلية العلوم

General Trends in the Periodic Table

Atomic Radius / **Bond length increase**



Electron Affinity /Electronegativity/ Ionization Energy

Effective Nuclear Charge (Zeff) **increase**

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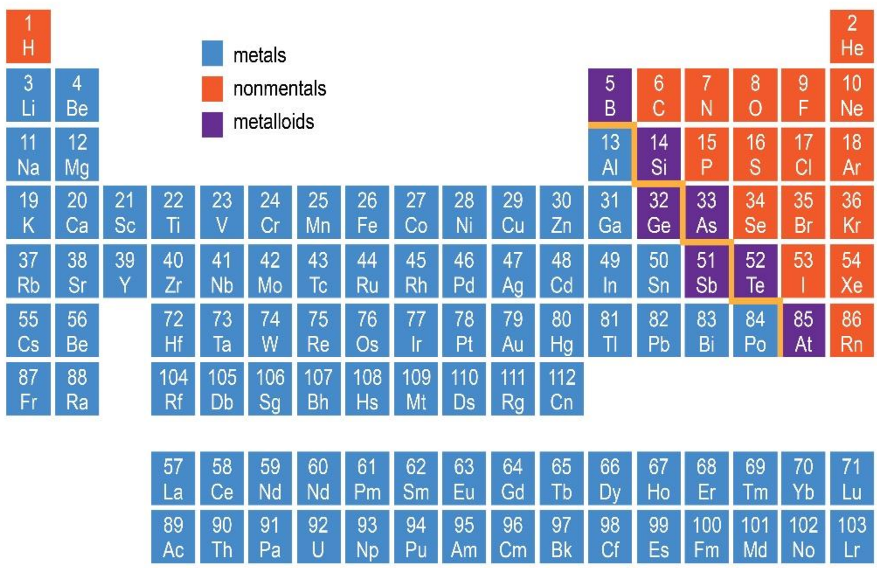

* Chapter 3 : Stoichiometry

* Lesson 9 : Emperical, Molecular, & Structural formulas

* Compounds are made of elemental atoms held together by chemical bond ; Compounds have two or more elements.

Q. what is the chemical bond ?

A. Chemical bonds are the forces of attraction between the atoms. which Come from the attraction between protons & electrons.

Types of Chemical bonds

- Covalent
- Ionic

Covalent

It is formed by equal sharing of electrons between the two atoms

e.g. H_2O (water) is Covalent Compound

H_2O
(2 : 1)

* Covalent Compounds may have Single, double or triple bonds.

e.g.

H−H   O=O   N≡N

Ionic

It is formed by tranfer of one or more electrons from one atom to the other forming Cation and anion

e.g. Sodium Na^{11} Chlorine Cl^{17}

2,8,1 → 2,8,7

Na^+ : 2,8 Cl^- : 2,8,8

So, Na^+Cl^-
ionic Compond.

* The properties of the formed Compound are completely different from its elements

Property	H_2	O_2	H_2O Water molecule
* B.p.	$-253\,^\circ C$	$-183\,^\circ C$	$100\,^\circ C$
* State	Gas	Gas	Liquid
* flammability	Explosive	Necessary for Combustion	used to extinguish flame

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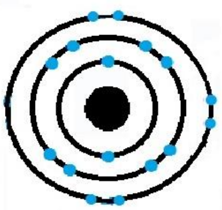
* Octet rule :

Atoms, of main-group elements, tend to Combine in such way that each atom has 8 electrons in its valence shell (as the nearest noble gas).

Thus;

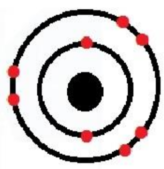
In Covalent Compounds, two non metals of same electron affinity Combine by equal sharing of electrons and fill their valence shell with 8 electrons to be stable.



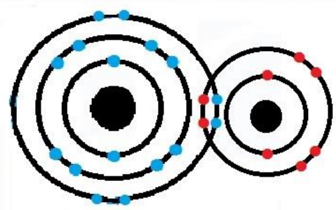
^{16}S : 2, 8, 6

By sharing of two electrons

^{8}O : 2, 6



S^{2-} : 2, 8, 8



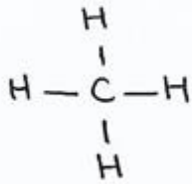
O^{2-} : 2, 8

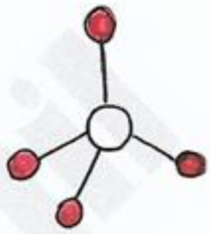
(*) Representing Compounds:

Compounds are generally represented by chemical formula, where the chemical formulas show the elements that are in the Compound and use the Letter symbol of the element.

e.g.

Methane CH_4





Molecular formula   Structural formula   Ball-stick model

* Molecular Formula:

It shows the actual number of atoms of each element present in the Compound.

e.g.

$\underbrace{H_2O}_{2H\quad 1O}$; $\underbrace{K_2SO_4}_{2K\quad 1S\quad 4O}$; $\underbrace{Mg(NO_3)_2}_{Mg\quad 2N\quad 6O}$

Some Complex ions: SO_4^{--} ; NH_4^{+} ; NO_3^{-} (polyatomic ions)

(Sulfate) (Ammonium) (Nitrate)

* Emperical Formula:

It shows only the simplest ratio of atoms of each element in the Compound.

e.g. Molecular Formula: H_2O_2 ; C_6H_6 ; $\underbrace{C_2H_4 , C_3H_6 ; C_4H_8}_{CH_2}$

Emperical Formula: HO ; CH ;

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* Structural Formula:

It uses lines to represent the Covalent bonds and shows how the atoms in the molecule are Connected.

Where;

Shape of Line	No. of Shared e^-	Name of Covalent bond
Single line —	1 pair = 2 e^-	Single H—H
Double line =	2 pairs = 4 e^-	double O=O
Triple line ≡	3 pairs = 6 e^-	Triple N≡N

Q. What is the difference between atomic element and molecular element?

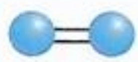
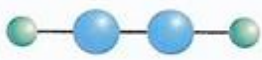
A.

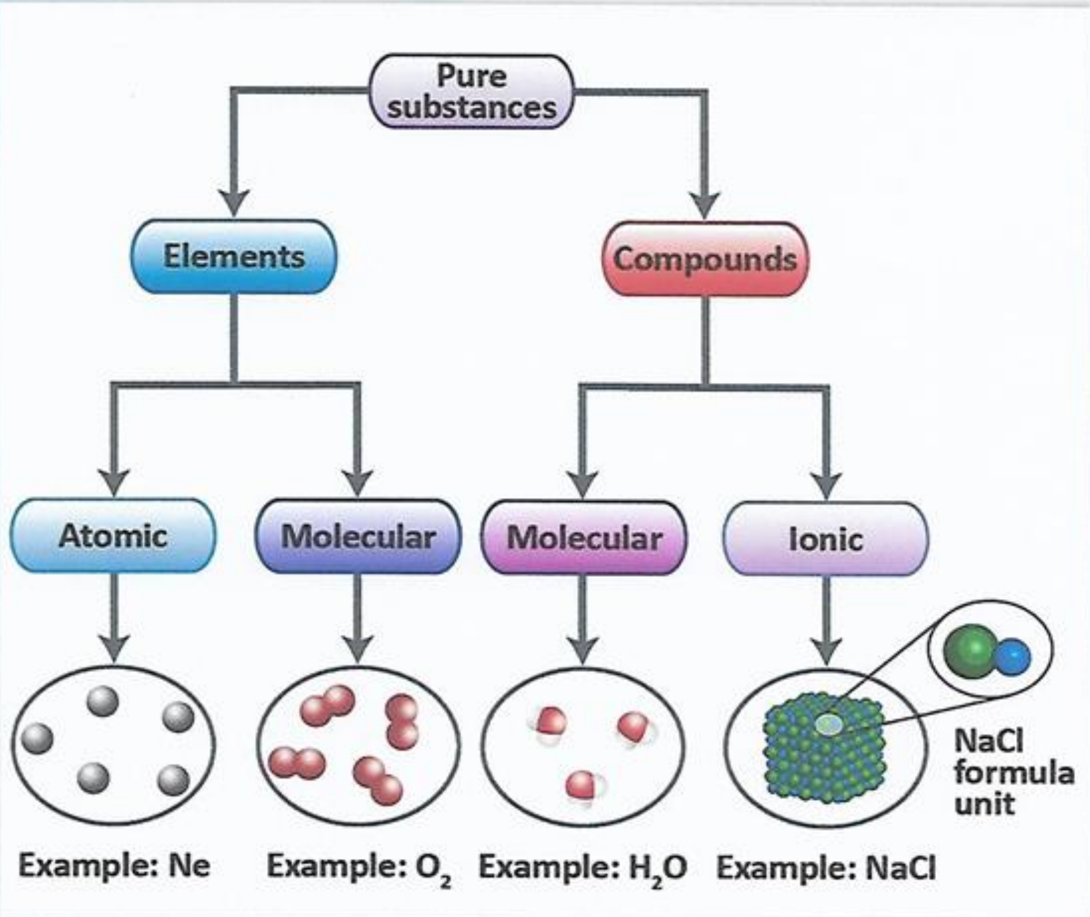
Atomic element Composed of single type of atoms. e.g. C ; He.

Molecular element is Composed of multi-atom molecules (with more than one atom). e.g.

diatomic molecules like O_2 and Cl_2.

polyatomic molecules like S_8 and P_4.

Chemical Formula	Structural Formulas	
H_2	H—H	
O_2	O=O	
H_2O	H—O—H	
H_2O_2	H—O—O—H	



« Ch 3: Stoichiometry »

Lesson 10 : Compound Formulas and Naming

* Introduction :

Sodium Chloride ($NaCl$) is made of sodium and Chlorine atoms. it has a common name : Table salt and scientific name : Sodium Chloride.

Most of the Compounds don't have Common names, so they are named according to scientific rules (scientific names).

* Ionic compounds (Formulas & Names)

Ionic Compound = (+ve metal Cation) + (−ve nonmetal Anion)

the formula unit of ionic Compound must have an equal number of +ve and −ve charges, so the formula unit has net charge = Zero

$$\underbrace{Na^{+} \quad Cl^{-}}_{0} \qquad\qquad \underbrace{Br^{-} \quad Mg^{2+} \quad Br^{-}}_{0}$$

$$NaCl \qquad\qquad\qquad MgBr_2$$

⇨ How to write the formula of ionic Compound Aluminum oxide ?

e.g. 1

Aluminum Oxide
- Al (Metal) → Group 3A → Al^{3+}
- O (Nonmetal) → Group 6A → O^{2-}

Al^{3+} O^{2-}

Al_2O_3

e.g. 2

Calcium Chloride
- Ca → Group 2A → Ca^{2+}
- Cl → Group 7A → Cl^{-}

Ca^{2+} Cl^{-}

$CaCl_2$

* Formula-to-Name Rules for Ionic Compounds :

* For Cations:

- **A** Metal (with invariant charge) → Name of metal
- **B** Metal (with variable charge) → Name of metal + Charge
- poly atomic ion → Name of polyatomic ion

* For Anions :

- For Nonmetal → Stem of nonmetal name + ide
- poly atomic ion → Name of polyatomic ion

N.B. polyatomic ions are several atoms attached together by Covalent bonds as one ion.

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* Metal ions with variable charge (can have more than one possible charge) are named as :

Name of metal + Roman numeral charge

e.g.

$Fe^{2+} \longrightarrow$ Iron (II) ; $Cu^{+} \longrightarrow$ Copper (I)

N.B. the charge can be determined from the charge of anion.

* Metal ions with invariant charge (can have only one possible charge) so, they are named without the oxidation number (charge)

e.g.

$Li^{+} \longrightarrow$ Lithium ; $Ca^{2+} \longrightarrow$ Calcium

Variable Change Examples			
Metal	**Ion**	**Name**	**Older name***
Chromium	Cr^{2+}	Chromium(II)	Chromous
	Cr^{3+}	Chromium(III)	Chromic
Iron	Fe^{2+}	Iron(II)	Ferrous
	Fe^{3+}	Iron(III)	Ferric
Cobalt	Co^{2+}	Cobalt(II)	Cobaltous
	Co^{3+}	Cobalt(III)	Cobaltic
Copper	Cu^{+}	Copper(I)	Cuprous
	Cu^{2+}	Copper(II)	Cupric
Invariant Change Examples			
Metal	**Ion**	**Name**	**Group number**
Li	Li^{+}	Lithium	1A
Na	Na^{+}	Sodium	1A
K	K^{+}	Potassium	1A
Mg	Mg^{2+}	Magnesium	2A
Ca	Ca^{2+}	Calcium	2A
Al	Al^{3+}	Aluminum	3A
Zn	Zn^{2+}	Zinc	*
Ag**	Ag^{+}	Silver	*

* Naming of Binary Ionic Compounds:

① Binary Ionic Compound = Metal Cation + Nonmetal Anion

② Name the metal cation firstly, then the name of nonmetal anion.

③ For metal cation:

Cation

Invariant Charge → Name of metal

Variant → Name of metal + Charge

④ For nonmetal anion:

Name of nonmetal + end (ide)

(For metals with invariant Charge)

Name of cation (metal)	Base name of anion (nonmetal)+ *-ide*
Cesium $Cs = Cs^{+}$ Because it is group 1A	Fluorine $F = F^{-}$ Because it is group 7A
Cs^{+} = Cesium	F^{-} = fluoride

Cesium fluoride

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⇒ Naming binary ionic compound with variable charge metal: (Transition element)

e.g. CuF_2

1) Metal Cation + Nonmetal Anion
   Cu F

2) Group No. of F is 7A ; so it will be F^-
   Thus; Cu must be Cu^{2+} (to balance 2 F ions)

3) Name Cation : Cu^{2+} = Copper (II)

4) Name Anion : F^- = Fluoride

5) Final Name : Copper (II) Fluoride

⇒ Naming the compounds containing polyatomic ions:

- Polyatomic ions are single ions that contain more than-one atom.
- Name the ionic compound : name of cation + name of anion

* Remmber well:

- Carbonate : CO_3^{2-}
- Bicarbonate : HCO_3^-
- Hydroxide : OH^-
- Nitrate : NO_3^-
- Nitrite : NO_2^-
- Sulfate : SO_4^{2-}
- Sulfite : SO_3^{2-}
- Ammonium : NH_4^+

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* Naming Molecular Compounds :

* Molecular Compounds are composed of two or more nonmetals.

Rules :

1) Write the name of element with the smallest group number first.
2) If the two elements are in the same group, write the name of element with the greatest row number first.
3) Use the prefixes that indicate the number of atoms present.

Name : Prefix – Name of the 1st element + Prefix – Name of the 2nd element + ide

➢ If there is only one atom for the first element, don't use the prefix "mono".

Prefix	No.
mono	1
di	2
tri	3
tetra	4
penta	5
hexa	6
hepta	7
octa	8
nona	9
deca	10

examples;

NI_3 : nitrogen triiodide

PCl_5 : phosphorus pentachloride

P_4S_{10} : tetraphosphorus decasulfide

* Naming Acids :

- Acids are molecular compounds that produce H^+ when dissolved in water e.g. $HCl_{(aq.)}$; $H_2SO_{4\,(aq.)}$
- Acids have sour taste and can dissolve many metals. such as Zn, Fe, or Mg but not Au, Ag & Pt.

Types of Acids

Binary acids	Oxyacids
$(H^+) + $ (nonmetal anion)	(H^+) + Poly atomic anion
e.g. HCl	e.g. HNO_3 , H_2SO_4

* Acids are produced by adding H^+ ion to anion;

$$H^+ + \underset{\text{Nitrate}}{NO_3^-} \longrightarrow \underset{\text{Nitric acid}}{HNO_3}$$

$$H^+ + \underset{\text{Chloride}}{Cl^-} \longrightarrow \underset{\text{Hydrochloric acid}}{HCl}$$

****

Ch. 3. : Stoichiometry

Lesson 11 : Composition of Compounds & Chemical equations

* Introduction :

Chemical formula, e.g. H_2O, is useful to Know the no. of atoms and their ratio. Also, the chemical formulas are used to show the reactions in the form of "Chemical equations".

Q. What is the molecular mass or molecular weight ?

A.

It is the mass of an individual molecule (formula unit).

i.e. the mass of a molecule is the sum of the masses of the atoms that make it up (expressed in amu).

example; water has formula unit : H_2O

$$\therefore \text{Molecular mass} = \sum \text{No. of element atoms} \times \text{its atomic mass}$$

Given: $H = 1 ; O = 16$

$$\therefore \underset{(H_2O)}{\text{Molecular mass}} = \underbrace{(2 \times 1)}_{H} + \underbrace{(1 \times 16)}_{O} = 18 \text{ amu (atomic mass unit)}.$$

So Molar mass of water is 18 g/mol.

* Mass Percent Composition :

It is the mass percent of one element to the total mass of all elements in the molecule.

$$\% \text{ Mass of element } (Z) = \frac{\text{Mass of element } (Z)}{\text{Molar mass of Compound}} \times 100$$

Example;

Calculate the mass % of hydrogen in water ?

A.

From Periodic Table

Water has formula unit (H_2O) $\therefore$ $H = 1$; $O = 16$

$$\% \text{ Mass } (H) = \frac{\text{Mass of } H}{\text{Molar mass of } H_2O} \times 100$$

$$= \frac{2 \times 1}{18} \times 100 = 11.1\%$$

Examples; Calculate the molar mass of each of the following:

1) K_2S

Molar mass = 2 X Atomic mass (K) + 1 X Atomic mass (S)

$= (2 \times 39.1) + (1 \times 32.1) = 110.3$ g/mol

2) $Ca_3(PO_4)_2$

Molar mass $= (3 \times 40.1) + 2 [(1 \times 31) + (4 \times 16)]$

$= 120.3 + 2 [31 + 64]$

$= 120.3 + 2 (95) = 120.3 + 190$

$= 310.3$ g/mol .

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Q. Calculate the mass % of oxygen in $HOCH_2CH_2OH$?

A.

firstly; $HOCH_2CH_2OH \Rightarrow C_2H_6O_2$

From Periodic Table
$C = 12$
$H = 1$
$O = 16$

$\therefore$ Molar mass $= (2 \times 12) + (6 \times 1) + (2 \times 16) = 62$ g/mol

$$\% \text{Mass}(O) = \frac{\text{Mass of O}}{\text{Molar Mass}} \times 100 = \frac{2 \times 16}{62} \times 100 = 51.6\,\%$$

* Emperical formula :

It is the simplest whole-number ratio of the atoms in a compound.

- Emperical formula Can be Calculating by Knowing the mass percent of each element in the compound.

(*) A Compound Contains 28 % nitrogen and 72 % Oxygen. Calculate its emperical formula ?

A.

1. Assuming the given percent as part of 100 g of compound. (Mass).
   so, 28 g nitrogen and 72 g oxygen.

2. Convert the mass to moles for each element.

$$\text{moles}(N) = \frac{\text{Mass}}{\text{Atomic mass}} = \frac{28}{14} = 2 \text{ mol}$$

$$\text{moles}(O) = \frac{72}{16} = 4.5 \text{ mol}$$

Now ; $N_2O_{4.5}$ is not accepted. (fraction ??)

3. Multiply all mole ratios by the suitable number to get whole numbers for all elements.

so, multiply with 2 : $2(N_2O_{4.5}) \Rightarrow N_4O_9$

∴ Emperical formula of the Compound is N_4O_9.

Q. A compound contains 60% Carbon, 4.48 g hydrogen and 35.52% oxygen. Calculate its emperical formula?

H = 1
C = 12
O = 16

A.

Step 1: Assuming % as Mass (g)

	C	H	O
% Percent:	60	4.48	35.52
Mass :	60	4.48	35.52

Step 2: Convert Mass to moles

C	H	O
$\frac{60}{12}$	$\frac{4.48}{1}$	$\frac{35.52}{16}$
5 mol	4.48 mol	2.22 mol

To

Step 3: Divide by the smallest number (2.22)

C	H	O
$\frac{5}{2.22}$	$\frac{4.48}{2.22}$	$\frac{2.22}{2.22}$
2.25	2	1

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Now ; $C_{2.25}\ H_2\ O_1$

It is not accepted (fraction).

Step 4 : Multiply with the suitable number to get all whole-number ratio.

so, Multiply by 4 : $C_9\ H_8\ O_4$

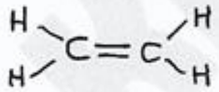
Decimal	Multiply by
0.20	5
.25	4
.33	3
.50	2

Now ;

Q. What is the difference between emperical and molecular formulas ?

"Emperical Formula"

It is the simplest whole-number ratio of the atoms of the elements in the compound.

e.g.  C_2H_4

∴ Emperical formula : CH_2

"Molecular Formula"

It is the actual number of atoms of elements in the compound. It is a multiple of the emperical formula.

Molecular formula = Emperical formula X No.

so, if we have an emperical formula CH_2

so, Molecular formula may be :

CH_2 , C_2H_4 , C_3H_6 , C_4H_8 or .....

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* Chemical Reactions and Equations :

- Chemical reactions involve rearrangement of atoms to produce new molecules.
- Chemical equations are a shorthand way to describe the chemical reactions.

* Chemical equations give basic information about the reaction :

- Formulas of reactants and products.
- States of reactants and products (s), (l), (g) or (aq.)
  Solid Liquid gas aqueous
- Relative numbers of reactants and products in the reaction.
- Weights of reactants and products.
- Required conditions : Δ (heat) ; $h\nu$ (light) ; ......etc.

* Balanced chemical equations :

Chemical equation must be balanced, in such way that an equal number of atoms of each element on both sides of the arrow.

e.g.

$$CH_{4(g)} + 2\,O_{2(g)} \longrightarrow CO_{2(g)} + 2\,H_2O_{(g)}$$

1C, 4H, 4O   1C, 4H, 4O

Q. Balance the following chemical equation :

$$Al_{(s)} + O_{2(g)} \longrightarrow Al_2O_{3(s)} \quad ?$$

A.

$$4\,Al_{(s)} + 3\,O_{2(g)} \longrightarrow 2\,Al_2O_{3(s)}$$

4Al, 6O   4Al, 6O

Chapter 4 : Chemical Bonding ; Chemical Reaction
Lesson 12 : Reaction Stoichiometry, Limiting Reactants and Percent Yield

* Introduction :

- Stoichiometry is the calculation of the relative quantities of the reactants and products in a chemical reaction.

- Balanced equations

The coefficients in a balanced chemical equation show the relative amounts (in moles) of reactants and products.

e.g.

$$2\ C_8H_{18\,(l)} + 25\ O_{2\,(g)} \longrightarrow 16\ CO_{2\,(g)} + 18\ H_2O_{(g)}$$

2 mol 25 mol 16 mol 18 mol

Thus;

Balanced equations outline how reactants combine to form products.

i.e. 2 mol C_8H_{18} react with 25 mol O_2 (2 : 25)
also, 2 mol C_8H_{18} give 16 mol CO_2 (2 : 16)
2 mol C_8H_{18} give 18 mol H_2O (2 : 18)

} Stoichiometric Ratio

Q. Suppose that we burn 22 mol C_8H_{18}, Calculate the formed amount of CO_2 ?

From the equation : $2\ C_8H_{18} \longrightarrow 16\ CO_2$

2 mol — 16 mol

22 mol — ?

$$\text{moles } (CO_2) = \frac{22 \times 16}{2} = 176 \text{ mol } CO_2$$

Q. Calculate the mass of CO_2 produced if 114.22 g of C_8H_{18} was burnt? (Given that: Molar mass C_8H_{18} = 114.22 g/mol) & Molar mass CO_2 = 44 g/mol

From the equation :

$$2\,C_8H_{18} \longrightarrow 16\,CO_2$$

2 mol → 16 mol

so, Firsty Convert mass 114.22 g of C_8H_{18} to moles :

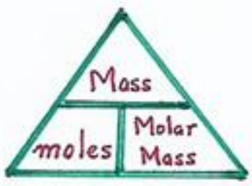
$$\text{moles}\ (C_8H_{18}) = \frac{\text{Mass}}{\text{Molar Mass}} = \frac{114.22}{114.22} = 1\ \text{mol}$$

So;

$$2\,C_8H_{18} \longrightarrow 16\,CO_2$$

2 mol — 16 mol

1 mol — ?



$$\therefore\ \text{moles}\ (CO_2) = \frac{1 \times 16}{2} = 8\ \text{mol}$$

Now;

Convert moles of CO_2 to the required mass:

$$\text{Mass}\ (CO_2) = 8 \times 44 = 352\ \text{g}$$

N.B. If you asked about the no. of CO_2 molecules produced use Avogadro's number (6.022×10^{23} molecules/mol).

$$\therefore\ \text{Molecules of } CO_2 = \text{moles} \times \text{Avogadro's No.}$$

$$= 8 \times 6.022 \times 10^{23} = 4.82 \times 10^{24}\ \text{molecules.}$$

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Q. In the following photosynthesis process; how many grams of glucose ($C_6H_{12}O_6$) could be synthesized from 37.8 g CO_2 ?

Eqvation: $6\,CO_2 + 6\,H_2O \longrightarrow C_6H_{12}O_6 + 6\,O_2$ Molor mass ($C_6H_{12}O_6$) = 180.2

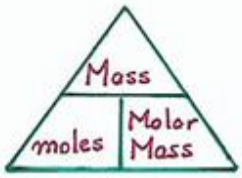
A. 〃 〃 (CO_2) = 44

From the equation: $6\,CO_2 \longrightarrow C_6H_{12}O_6$

6 mol 1 mol

But we have the mass of CO_2 (37.8 g), so you have to convert it to moles:

$$\text{moles } (CO_2) = \frac{\text{Mass}}{\text{Molor Mass}} = \frac{37.8}{44} = 0.859 \text{ moles}$$



Thus;

from the stoichiometric ratio:

$6\,CO_2 \longrightarrow C_6H_{12}O_6$

6 mol 1 mol

0.859 mol ?

$$\therefore \text{moles}_{C_6H_{12}O_6} = \frac{0.859}{6} = 0.143 \text{ moles}$$

but the question asks about the mass of glucose.

so, we have to convert moles to mass using Molar mass:

$$\text{Mass } (C_6H_{12}O_6) = 0.143 \times 180.2 = 25.8 \text{ g}$$

Q. Estimate the amount of CO_2 (g) produced in 2007 by the combustion of 3.5×10^{15} g of gasoline ?

A. Ratio: 2 mol C_8H_{18} : 16 mol CO_2

1 mol C_8H_{18} = 114.22 g

1 mol CO_2 = 44 g

Step 1: Convert 3.5×10^{15} g to moles of gasoline

$$\text{moles (gasoline)} = \frac{3.5 \times 10^{15}}{114.22} = 3.06 \times 10^{13} \text{ mol}$$

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Step 2 :

From stoichiometric ratio: $2\ mol\ C_8H_{18} : 16\ mol\ CO_2$

$3.06 \times 10^{13}\ mol$ ⟶ ? mol

$$\therefore \text{moles}(CO_2) = \frac{(3.06 \times 10^{13}) \times 16}{2} = 2.45 \times 10^{14}\ mol$$

Step 3: Convert moles of (CO_2) to mass :

$$\text{Mass}(CO_2) = \text{moles} \times \text{Molar mass}$$
$$= (2.45 \times 10^{14})(44\ g/mol) = 1.08 \times 10^{16}\ g$$
$$\simeq 1.1 \times 10^{16}\ g\ CO_2$$

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* Limiting Reactant :

the limiting reactant determines the amount of product that will be produced in a reaction.

Example ; $3\,Ca + 2\,P \longrightarrow Ca_3P_2$

Stoichiometric ratio 3 : 2

If we start with 2 moles Ca and 2 moles of P ;

So, the amount of Ca isn't enough to react with the 2 moles of P.

Thus; Ca controls the production of Ca_3P_2 (will completely used up).

So, Ca is the limiting reactant (P is excess reactant).

* Excess Reactant :

This is any reactant that occurs in a larger quantity than is required for the reaction to take place.

* Theoretical Yield :

This is the amount of product that can be produced in a chemical reaction, based on the amount of limiting reactant.

* Actual Yield :

This is the amount of product actually produced in a reaction (usually < theoretical yield).

* Percentage Yield :

$$\%\,Yield = \frac{Actual\ Yield}{Theoretical\ Yield} \times 100$$

Q. If we have 5 molecules of CH_4 and 8 molecules of O_2; Which is the limiting reactant in the following reaction:

$$CH_{4(g)} + 2\,O_{2(g)} \longrightarrow CO_{2(g)} + 2\,H_2O$$

A.

To decide which one (CH_4 or O_2) is the limiting reactant Compare the amount of product that produced from each.

" Product amount "

Depending on CH_4 amount

Eq.	CH_4 $\longrightarrow$	CO_2
Ratio	1	1
Given	5	?

Expected amount (CO_2) = 5 molecules

Depending on O_2 amount

Eq.	$2\,O_2$ $\longrightarrow$	CO_2
Ratio	2	1
Given	8	?

Expected amount (CO_2) = 4 molecules

Thus; O_2 amount is enough only to produce 4 molecules of CO_2 (less amount); so. O_2 is the limiting reactant.

#

Q. Ammonia, NH_3, was synthesized according to the following reaction using 86.3 g NO and 25.6 g H_2 :

$$2\,NO_{(g)} + 5\,H_{2(g)} \longrightarrow 2\,NH_{3(g)} + 2\,H_2O_{(g)}$$

What is the theoretical yield of NH_3 in grams?

Given that: molar mass NO = 30.01 g/mol
" " H_2 = 2.02 g/mol
" " NH_3 = 17.03 g/mol

A. $2\,NO_{(g)} + 5\,H_{2\,(g)} \longrightarrow 2\,NH_{3\,(g)} + 2\,H_2O_{(g)}$

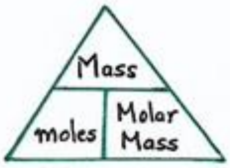
* Step 1:

Hint:

For NO : $g\ NO \xrightarrow[(\div 30.01)]{} mol\ NO \xrightarrow[\substack{2:2 \\ (\times 2/2)}]{\text{From Eq.}} mol\ NH_3$

Smallest $\xrightarrow{(\times 17.03)}$ $g\ NH_3$

For H_2 : $g\ H_2 \xrightarrow[(\div 2.02)]{} mol\ H_2 \xrightarrow[\substack{5:2 \\ (\times 2/5)}]{\text{From Eq.}} mol\ NH_3$



« Which is the limiting reactant ? »

For NO:

	$2NO$	$\longrightarrow$	$2NH_3$
Ratio:	2 moles		2 moles
Given	$\frac{86.3}{30.01}$ moles		?

moles (NH_3) = 2.8757 mol

less product is produced from the limiting reactant NO

For H_2 :

	$5H_2$	$\longrightarrow$	$2NH_3$
Ratio:	5 moles		2 moles
Given	$\frac{25.6}{2.02}$ moles		?

moles (H_2) = 5.0693 mol

* Step 2:

To Calculate theoritical yield (mass of product) ; Convert the produced moles of product (2.8757 mol) to mass :

Theoretical yield (NH_3) = 2.8757 X Molar mass

= 2.8757 X 17.03 = 49.0 g

Q. 53.2 g Na and 65.8 g Cl_2 react according to the equation:

$$2\,Na_{(s)} + Cl_{2\,(g)} \longrightarrow 2\,NaCl_{(s)}$$

- Which is the limiting reactant?
- The actual yield was 86.4 g of NaCl. What is the theoretical yield and percentage yield?

Atomic Mass Na = 23
Molar mass Cl_2 = 70.9

A.

Which is the Limiting reactant

For Na:

Eq.	$2Na \longrightarrow$	$2\,NaCl$
Ratio	2 mol	2 mol
Given	$\frac{53.2}{23}$	?

$\therefore$ moles (NaCl) = 2.31

For Cl_2:

Eq.	$Cl_2 \longrightarrow$	$2\,NaCl$
Ratio	1 mol	2 mol
Given	$\frac{65.8}{70.9}$	?

$\therefore$ moles (NaCl) = 1.85 Small.

$\therefore$ Cl_2 is the limiting reactant.

Thus;

Theoretical yield (NaCl) = moles X Molar mass = 1.85 X 58.5 = 108 g

From the given actual yield (NaCl) = 86.04 g

$$\therefore \quad \%\ \text{yield} = \frac{\text{Actual yield}}{\text{Theoretical Yield}} \times 100$$

$$= \frac{86.4}{108} \times 100 = 80.0\,\%$$

Ch. 4: Chemical bonding & Chemical Reaction
Lesson 13: Solution Concentration & Types of aqueous solutions

* Solution Concentration & Solution Stoichiometry

Solution is a homogenous mixture of two or more substances.

Solution = Solvent + Solute(s)
(largest amount) (small amounts)

e.g.

- Sugar solution : water is solvent & sugar is solute.
- Air : N_2 gas is Solvent & other gases (O_2, CO_2, H_2O...) are solutes.

* Solution Concentration:

Solute to solution relationship.

« Molarity »

It is the number of moles of solute per liter of solution.

$$\boxed{M = \frac{\text{No. of moles of Solute}}{\text{Volume of solution (L)}}}$$

Unit: mol/L

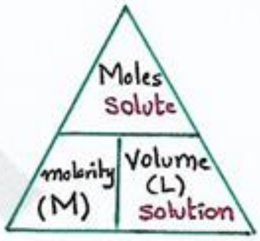
* Prepare 1 L of a 1.00 M NaCl solution:

* Weigh out and 1.00 mol of NaCl (58.44 g) in graduated flask then add water to dissolve the salt, then add to the mark 1 L.

Q. Find the molarity (M) of a solution that has 25.5 g KBr dissolved in 1.75 L solution? Given: molar mass (KBr) (1 mol KBr = 119.00 g)

A.

$$\because M = \frac{\text{moles of solute (KBr)}}{\text{Volume of Solution (L)}} = \frac{n_{solute}}{V_{Soln.(L)}}$$



$$\because n_{solute} = \frac{\text{Mass}}{\text{Molar Mass}} = \frac{25.5}{119.00} = 0.214 \text{ mol}$$

$$\therefore M = \frac{0.214}{1.75} = 0.122 \text{ M}$$ (Unit: mol/L or M)

N.B. Most solutions are between 0 and 18 M, so the answer makes sense.

Q. How many liters of 0.125 M NaOH Contain 0.255 mol NaOH?

A.

0.125 M NaOH means that: 0.125 mol NaOH $\longrightarrow$ 1 L solution
0.255 mol NaOH $\longrightarrow$?

$$V_{Solution} = \frac{0.255 \times 1}{0.125} = 2.04 \text{ L solution (Acceptable)}$$

---

* **Solution Dilution:**

Often, solutions are stored as concentrated stock solutions.

To make solutions of lower concentrations, only more solvent is added.

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Thus; the amount of solute doesn't charge, Just the volume of Solution

Moles of solute before dilution = Moles of solute After dilution

$$\boxed{M_1 \cdot V_1 = M_2 \cdot V_2}$$

Q. Prepare 3.00L solution of a 0.500 M $CaCl_2$ from 10.0 M stock solution ?

A.

diluted Solution (1)	stock solution (2)
V = 3 L ; M = 0.5	V = ? ; M = 10.0

$$\because \quad M_1 \cdot V_1 = M_2 \cdot V_2$$

$$(0.5 \times 3) = (10.0 \times V_2) \Rightarrow V_2 = \frac{0.5 \times 3}{10} = 0.150 \text{ L}$$

Thus;

we take 0.150L from the stock solution and add water to reach volume of 3.00 L.

Q. How would you prepare 200 mL of 0.25 M NaCl solution from a 2.0 M solution ?

A.

Diluted solution	stock solution
V_1 = 200 ml ; M_1 = 0.25 M	V_2 = ? ; M_2 = 2 M

$$\because \quad M_1 \cdot V_1 = M_2 \cdot V_2$$

$$(200 \times 0.25) = (2 \times V_2) \Rightarrow V_2 = \frac{200 \times 0.25}{2} = 25 \text{ mL}.$$

Thus;

take 25 ml of 2.0M solution, and add water up to 200 mL.

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* Types of aqueous solutions and solubility :

Case A : Dissolve NaCl salt in water ⟶ Homogeneous solution .

Case B : " Sugar $C_{12}H_{22}O_{11}$ in water ⟶ " " .

Now;

How do solids such as salt and sugar dissolve in water ?

OR :

What happens when a solute dissolves ?

A.

In general, dissolution of solute in solvent depends on :

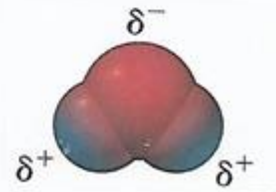
A) Attractive forces between solute molecules .

B) " " " Solvent " .

C) " " " solute and solvent molecules .

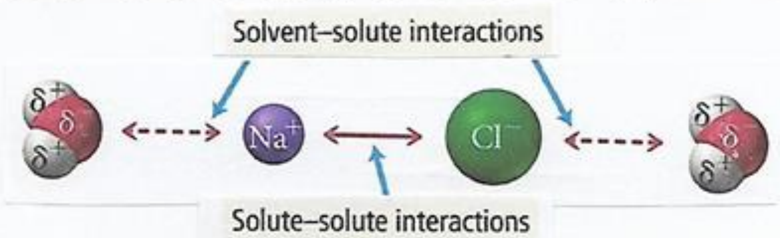
If the attraction forces between solute and solvent are strong enough , the solute will dissolve .

* Charge distribution in water molecule :



In water molecule ; there is an uneven distribution of electrons in such way that oxygen has partial –ve charge and hydrogen has partial +ve charge , Thus water is a polar molecule .

* Solute and Solvent interactions : NaCl in water



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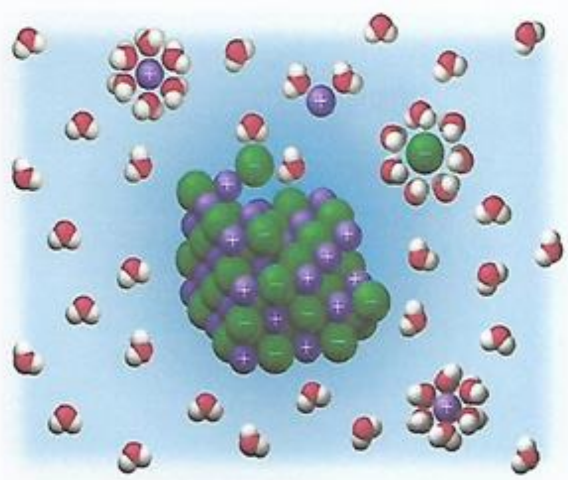
Dissolution of an Ionic Compound



* Dissolving NaCl in water:

When NaCl is added to water:

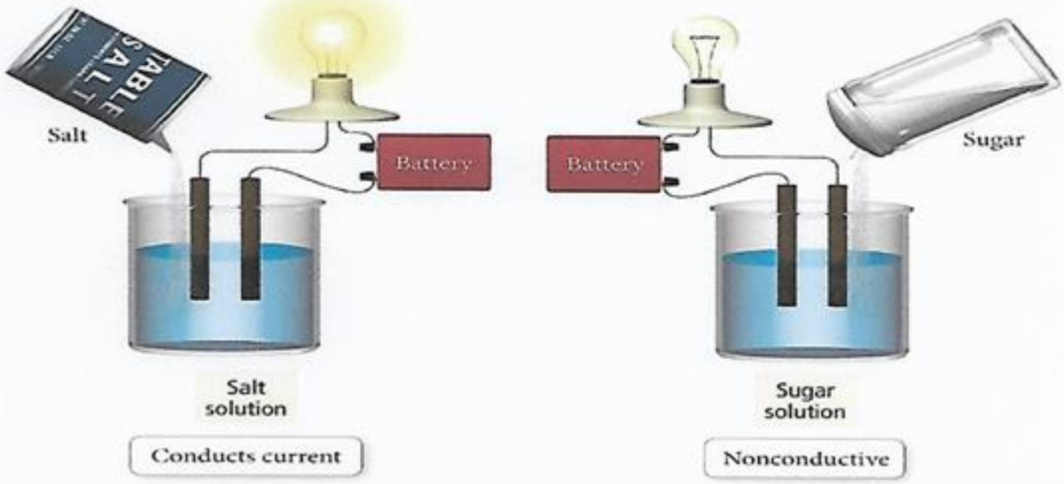
1. Each ion (Na^+ or Cl^-) is attracted toward H_2O molecules and away from the solid Crystal.
2. When the ion enters the solution, it is surrounded by water molecules Through the opposite Charged pole.
3. As result, solution with free moving ions is formed and can Conduct the electricity.

* Compare electrolyte and nonelectrolyte solutions?

	Stronge electrolytes	Weak electrolytes	Nonelectrolytes
Ionization	ionize Completely	ionize partially	don't ionize at all
Examples	Soluble salts and Strong acids or bases $CuCl_2 \rightarrow Cu^{2+} + 2Cl^-$	Weak acids or bases $CH_3COOH \rightleftharpoons CH_3COO^- + H^+$	polar substances like sugar and alcohol $C_6H_{12}O_{6(s)} \rightarrow C_6H_{12}O_{6(aq.)}$
Electrical Conductivity	good Conductors for electricity	weak Conductors	don't Conduct electricity

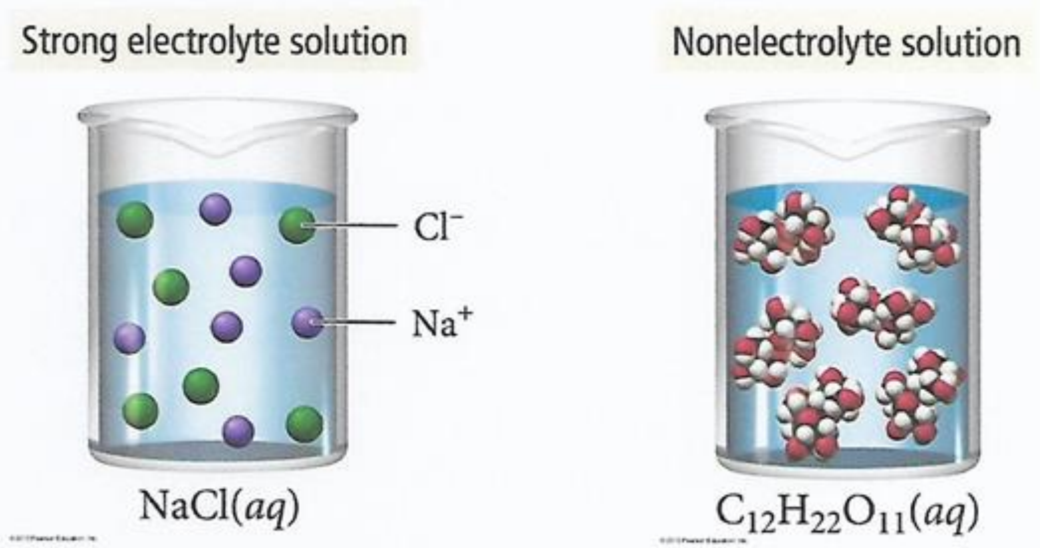
Electrolytes and Nonelectrolytes

Electrolyte and Nonelectrolyte Solutions



- **A solution of salt (an electrolyte) conducts electrical current**
- **A solution of sugar (a nonelectrolyte) does not.**

Salt (Ionic compound) vs. Sugar Dissolved in Water



Salts (ionic compound) ionize when

Molecular compounds do not

Ch. 4: Chemical Bonding & Chemical Reaction

Lesson 14: Acid-Base Reactions & Redox Reactions

* Acid-Base Reactions:

They are very common in chemistry.

e.g. HCl "hydrochloric acids" exists in stomach and helps in food digestion. When it is leaked into the esophagus, it causes the "heartburn". So, Antacids are used to neutralize the stomach acid by Acid-Base chemical reaction.

* Arrhenius definitions for acid and base:

Acid: A substance that produces H^+ ions in aqueous solution.

Base: A substance that produces OH^- ions in aqueous solution.

Examples;

$$\underset{\text{Acid}}{HCl_{(aq.)}} \longrightarrow H^+_{(aq.)} + Cl^-_{(aq.)} \quad ; \quad \underset{\text{Base}}{NaOH_{(aq.)}} \longrightarrow Na^+_{(aq.)} + OH^-_{(aq.)}$$

$$\Rightarrow \underset{\text{Acid formula}}{HA} \longrightarrow \underset{\text{proton}}{H^+} + A^- \qquad \underset{\text{Base formula}}{BOH} \longrightarrow B^+ + \underset{\text{hydroxide ion}}{OH^-}$$

Acid-Base "Neutralization" Reaction:

General equation: Acid + Base $\longrightarrow$ Salt + Water

e.g. $HCl_{(aq.)} + NaOH_{(aq.)} \longrightarrow NaCl_{(aq.)} + H_2O$

Net ionic equation: $H^+_{(aq.)} + OH^-_{(aq.)} \longrightarrow H_2O_{(L)}$

N.B.

* Protons released by acids associate with water (H_2O) in solution to form Hydronium ions " H_3O^+ "

$$\text{Acid} \longrightarrow \underset{\text{proton/water}}{H^+ + H_2O} \longrightarrow \underset{\text{hydronium ion}}{H_3O^+_{(aq.)}}$$

So,

$$HCl_{(aq.)} + H_2O_{(l)} \longrightarrow H_3O^+_{(aq.)} + Cl^-_{(aq.)}$$

* Hydroxide ions released by bases in water

$$\text{Base} \longrightarrow \underset{\text{hydroxide/water}}{{}^-OH + H_2O} \longrightarrow OH^-_{(aq.)}$$

But !!

Ammonia does not contain hydroxyl group (OH) but produces hydroxide ions (OH^-) in water, thus it is a base.

$$\underset{\text{Ammonia}}{NH_3} + H_2O \longrightarrow NH^+_{4\,(aq.)} + {}^-OH_{(aq.)}$$

* Polyprotic acids :

Some acids contain more than ionizable proton

e.g. Sulfuric acid (H_2SO_4) is diprotic acid

- $H_2SO_{4\,(aq.)} \longrightarrow H^+_{(aq.)} + HSO^-_{4\,(aq.)}$ (1^{st} step is strong)
- $HSO^-_{4\,(aq.)} \rightleftharpoons H^+_{(aq.)} + SO^{--}_{4\,(aq.)}$ (2^{nd} step is weak)

Similarly; Bases can release more than one ^-OH per mole.

e.g.

$$Sr(OH)_{2(aq.)} \longrightarrow Sr^{2+}_{(aq.)} + 2\,OH^-_{(aq.)}$$

Q. How to write the acid - Base equations?

A.

Example; Reaction of hydrogen iodide (HI) with barium hydroxide $Ba(OH)_2$

Step 1: Main unbalanced equation:

$$\underset{\text{(Acid)}}{HI_{(aq.)}} + \underset{\text{(Base)}}{Ba(OH)_{2(aq.)}} \longrightarrow \underset{\text{(Salt)}}{BaI_{2(aq.)}} + \underset{\text{(water)}}{H_2O_{(L)}}$$

Step 2: Balance the equation: "Molecular equation"

$$2HI_{(aq.)} + Ba(OH)_{2(aq.)} \longrightarrow BaI_{2(aq.)} + 2H_2O_{(L)}$$

Step 3: Net ionic equation (By removing spectator ions of salt)

$$H^+_{(aq.)} + OH^-_{(aq.)} \longrightarrow H_2O_{(L)}$$

Some common acids

Name of Acid	Formula
Hydrochloric acid	HCl
Hydrobromic acid	HBr
Hydroiodic acid	HI
Nitric acid	HNO_3
Sulfuric acid	H_2SO_4
Perchloric acid	$HClO_4$
Acetic acid	$HC_2H_3O_2$ (weak acid)

Some common bases

Name of Base	Formula
Sodium hydroxide	$NaOH$
Lithium hydroxide	$LiOH$
Potassium hydroxide	KOH
Calcium hydroxide	$Ca(OH)_2$
Barium hydroxide	$Ba(OH)_2$
Ammonia	NH_3 (weak base)

« Oxidation _ Reduction Reactions » "Redox" Reactions

They are the reactions in which electron(s) are transferred from one reactant to the other.

examples; Rusting of iron & combustion of gasoline (octane).

where;

Oxidation: gain of oxygen or loss of hydrogen.

Reduction: gain of hydrogen or loss of oxygen.

$$Fe_{(s)} + 3\,O_{2\,(g)} \longrightarrow 2\,Fe_2O_3$$ Rusting of iron

gain of Oxygen "Oxidation"

* Redox Reactions without Oxygen:

During the formation of ionic compounds, metal loses electron(s) "Oxidation" and nonmetal gains the electron(s) "Reduction".

e.g. NaCl

(metal) $\overset{11}{Na}_{(s)}$ (2, 8, 1) + $\overset{17}{Cl}_{(g)}$ (2, 8, 7) (nonmetal)

$\downarrow$ Redox

Na^{+} (2,8) Cl^{-} (2,8,8)

$\downarrow$

$NaCl_{(s)}$

Na (metal) loses electron "Oxidized" and Cl (nonmetal) gains this electron "reduced".

Equation: $$2\,Na_{(s)} + Cl_{2\,(g)} \longrightarrow 2\,NaCl_{(s)}$$

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* Very Important Note :

* The element which is oxidized is Known as "Reducing agent or Reductant e.g. Na is oxidized but reduces Cl.
* The element which is reduced is Known as "Oxidizing agent or Oxidant e.g. Cl is reduced but oxidizes Na.

* Rules for Assigning Oxidation state :

It is very important to know the oxidation number (state) to identify the redox reactions that occur between nonmetals.

Rule (1) : Oxidation state of an atom and free element is Zero.

Rule (2) : Oxidation state of monoatomic ion is equal to its charge.

Rule (3) : Sum of oxidation states of all atoms in:

- Neutral molecule or formula unit is Zero.
- An ion is equal to its charge.

Rule (4) : For metals, oxidation number equals to the group number.

e.g. Group 1A have +1 oxidation state.

Group 2A have +2 oxidation state.

Rule (5) : For nonmetals

Oxidation States of Nonmetals		
Nonmetal	**Oxidation State**	**Example**
Fluorine	-1	MgF_2 -1 ox state
Hydrogen	+1	H_2O +1 ox state
Oxygen	-2	CO_2 -2 ox state
Group 7A	-1	CCl_4 -1 ox state
Group 6A	-2	H_2S -2 ox state
Group 5A	-3	NH_3 -3 ox state

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- Sulfur as single ion has (−2) oxidation state but in polyatomic ion is different. e.g. H_2SO_4

→ H : is +1 (group 1A)
→ O : is −2 (group 6A)

Sum of Charges = Zero (Rule 3)

$$2(+1) + 1(X) + 4(-2) = 0$$

$$\therefore X = +6 \qquad \therefore S = +6$$

Q. Is the following reaction a redox reaction?

$$C + 2S \longrightarrow CS_2$$

A.

From the rules of oxidation states;

$$\underset{(0)}{C} + 2\underset{(0)}{S} \longrightarrow \underset{(+4)}{C}\underset{(-2)}{S_2}$$

Oxidation
Reduction

Where;

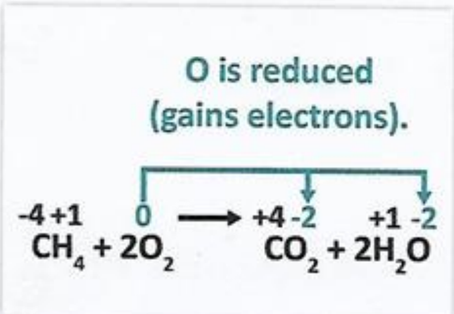
Oxidation : loss of electron(s) or increase in oxidation state. (C)

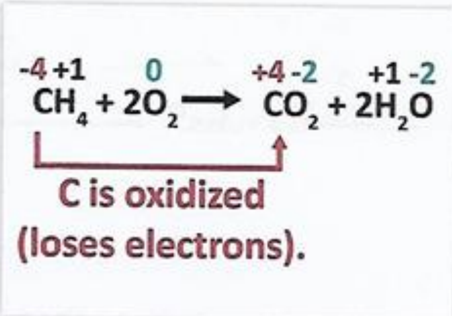
Reduction : gain of electron(s) or decrease in oxidation state. (S)

Q. Identify which element oxidized and that reduced in the following reaction?

$$CH_4 + 2O_2 \longrightarrow CO_2 + 2H_2O$$

A. From the rules of oxidation states;





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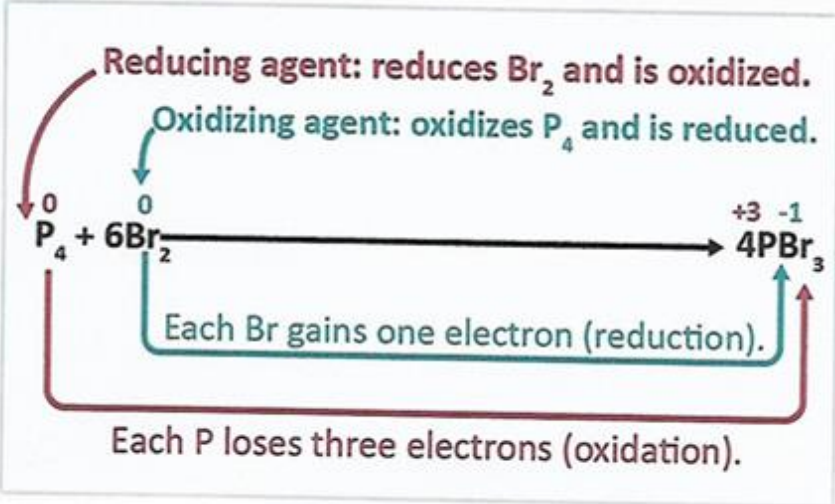
Q. In the following reaction, which substance is the oxidizing agent and which is the reducing agent?

$$P_4 + 6\,Br_2 \longrightarrow 4\,PBr_3$$

A.



Thus;

* The element which is oxidized ⇒ Reducing agent or Reductant.
* " " " " reduced ⇒ Oxidizing agent or Oxidant.

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Chapter 4 : Chemical bonding
Lesson 15 : Types of bonds

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* Introduction :

As we Know ; Chemical Compounds are formed by chemical Combination "bonding" between their atoms , but HOW ??

* Types of Chemical bonds :

Bond	Ionic	Covalent	Metallic
1. Element	Metal + Nonmetal	Two nonmetals	Metals
2. Bonding	e^- - transfer from metal (low IE) to nonmetal (high E.A) Thus, electrostatic attraction between Cation and anion	sharing of e^-'s between the two nonmetals (high EA), Thus attraction through shared electrons.	Valence electrons are shared as Pool between the metal atoms. "Nuclei in sea of electrons" "Electron Sea Model"
3. Example	$Na^+ Cl^-$ Cation Anion	H_2O water	$Na_{(s)}$ Sodium metal

* Classification of Covalent bonds :

They are three different types of Covalent bonding based on the no. of shared electrons .

Covalent bond.	Single	Double	Triple
1. No. of Shared e⁻'s	Two	four	Six
2. No. of Shared e⁻ pairs	One	Two	Three
3. No. of Lines	Single Line	Double Line	Triple Line
4. Examples	H—H ; C—C C—H ; H—Cl	O=O ; C=O C=C ; C=N	N≡N ; C≡N C≡C

In general;

The type of bonding used is the suitable way for the atom to reach "Octet rule" (or the noble gas e^- configuration).

e.g.

$^{8}O: 1s^2\ 2s^2\ 2p^4$ So, No. of valence electrons = 6 (Group No. 6A)

So, Oxygen atom makes bond in such way that it gains two electrons to reach the "Octet rule"

$$^{8}O: 1s^2\ 2s^2\ 2p^4 \xrightarrow[\text{sharing of } 2e^- \text{ / Covalent bond}]{\text{gains } 2e^- \text{ / ionic bond}} O^{2-}: 1s^2\ \underbrace{2s^2\ 2p^6}_{\text{octet}}$$

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* Lewis Representations:

Lewis structures allow us to easily see the number of valence electrons in an atom. Lewis notation uses two dots (duet) to represent a pair of electrons.

e.g.

$^{8}O : 1s^2 2s^2 2p^4$ Valence electrons = 6

Thus;

Lewis structure for O-atom is : 

N.B. Dots represent the electrons that the element could lose and empty spaces indicate the electrons that are needed to satisfy the octet rule.

Examples:

(Put the electrons firstly single then make pairing)

N.B.

Two elements not subject to Octet rule : H & He because they have only one E-level of max e^--capacity = 2

H : hydrogen can only lose or gain one electron.

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* Lewis structures for ionic Compounds:

Lewis structure represents ionic bonding by moving electron dots from metal to nonmetal, where the lewis structure of anion is usually written in brackets with charge in its upper right corner:

$$\underset{2,8,8,1}{K\cdot} \quad \underset{2,8,7}{\cdot\ddot{\underset{\cdot\cdot}{Cl}}:} \xrightarrow[\text{bonding}]{\text{Ionic}} \underset{2,8,8}{K^+} \quad \underset{2,8,8}{\left[:\ddot{\underset{\cdot\cdot}{Cl}}:\right]^-}$$

Both are "octet"

Another example;

Ionic Compound of sodium and sulfur needs two Na-atoms and one S-atom, Explain Why ??

Na-atom must lose one valence electron to obtain "octet", and.
S-atom " gain two electrons to obtain "octet"
Thus; Two Na-atoms are needed to combine with one S-atom.

$$2\,\dot{Na} + \cdot\ddot{\underset{\cdot}{S}}: \xrightarrow[\text{Transfer}]{\text{Ionic bonding}\ 2e^-} Na^+ \left[:\ddot{\underset{\cdot\cdot}{S}}:\right]^{2-} Na^+$$

* Lattice Energy:

Ions are arranged in a pattern "Crystal lattice" where the crystal lattice maximizes attractions between cations and anions leading to more stable arrangement.
extra stability of crystal lattice is measured as "Lattice Energy".

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N.B.

⇒ Lattice Energy increases with:

* Increasing the charge on the ion.
* Decreasing the size of ion.

* **Lattice energy:** the energy needed to separate a mole of a solid ionic compound into its gaseous ions.

Q. What are the properties of ionic compounds?

A.

- Physical properties:
  a- high melting points > 300°C.
  b. high boiling points.
  c. Hard and brittle solids.

- Solid ionic compounds:
  a. All are crystalline at room temperature.
  b. don't conduct electricity, but in liquid state they are strong electrolyte "Conductors"
  c. Liquid states are thermal insulators.

- Solubility of ionic compounds:
  Many solid ionic compounds are soluble in water.

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Chapter 4: Chemical Bonding
Lesson 16: Factors Affecting Bond Type

* Introduction:

Oxygen and sulfur (Group 6A) form covalent compounds with hydrogen in atomic ratio 1:2 (H_2O & H_2S). However, their reactivity is quite different (H_2S is more reactive).

Reactivity depends on:

1. Type of bonds within the compound.
2. polarity.
3. Bond energy (energy required to break the bonds).
4. Energy required to separate the molecules.

* Covalent Bonding & Lewis Structures:

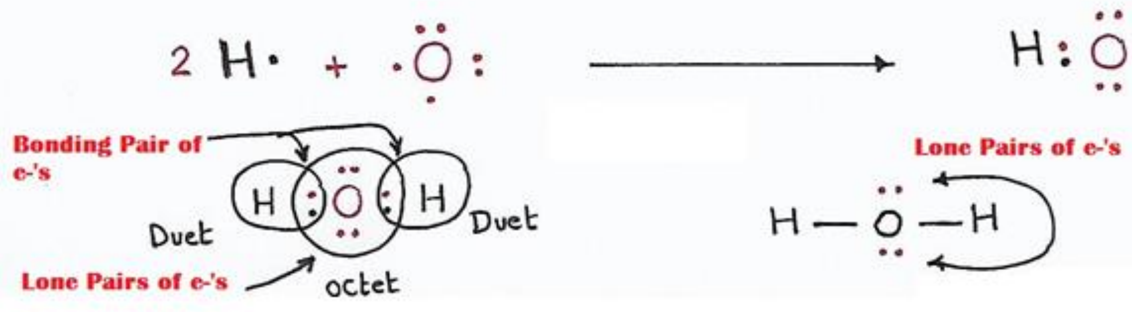
Molecular compounds contain covalent bonds (of shared electrons).

Lewis theory represents the covalent bonds by drawing the neighboring atoms to share e⁻-pairs to obtain octet rule (or duet for hydrogen).

Example; water H_2O

Lewis structures for atoms

Lewis structures for molecule



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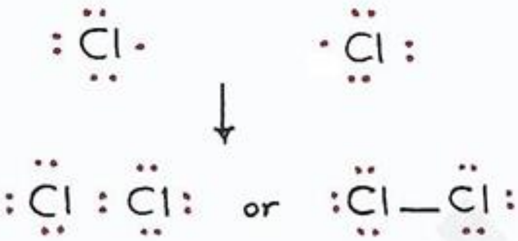
Example; Diatomic Halogen Molecules

Lewis structure :



Both are Octet

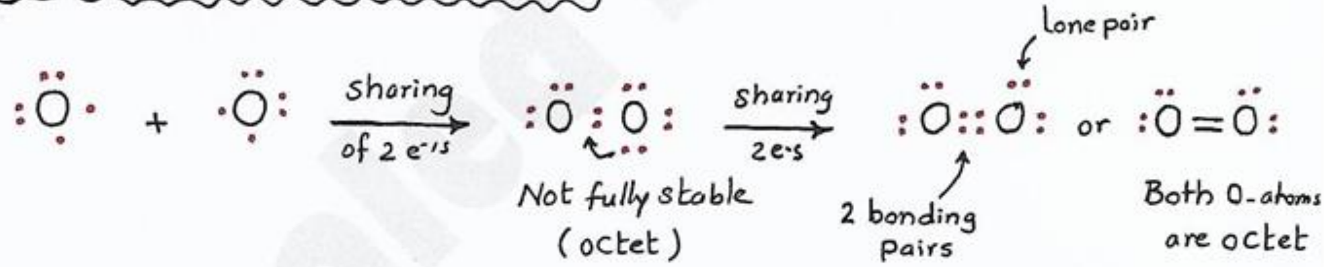
Q. Draw Lewis structure for H_2 molecule?

$$\dot{H} + \dot{H} \longrightarrow H:H \quad \text{or} \quad H-H$$

H obeys duet rule

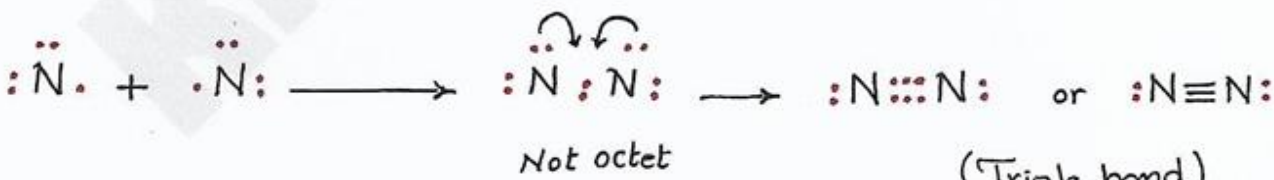
* Lewis structures for Multiple bonds :

Oxygen :



(Double bond)

Nitrogen :



(Triple bond)

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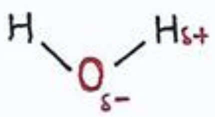
* Electronegativity : (E.N.)

It is the relative ability of atoms to attract shared electrons.

N.B.

1. EN (nonmetal) > EN (metal)
2. F-atom has the highest EN-value (4.0) while Cs and Fr atoms have the lowest EN-value (0.7).
3. As EN difference increases as the bond polarity increases.



; $\overset{\delta+}{H}-\overset{\delta-}{F}$ or $H \longrightarrow F$

. Electronegativity values and Trends in periodic table :

Electronegativity increases across the period

Electronegativity decreases down the group

H 2.20																	He
Li 0.98	Be 1.57											B 2.04	C 2.55	N 3.04	O 3.44	F 3.98	Ne
Na 0.93	Mg 1.31											Al 1.61	Si 1.9	P 2.19	S 2.58	Cl 3.16	Ar
K 0.82	Ca 1.0	Sc 1.36	Ti 1.54	V 1.63	Cr 1.66	Mn 1.55	Fe 1.83	Co 1.88	Ni 1.91	Cu 1.9	Zn 1.65	Ga 1.81	Ge 2.01	As 2.18	Se 2.55	Br 2.96	Kr 3.0
Rb 0.82	Sr 0.95	Y 1.22	Zr 1.33	Nb 1.6	Mo 2.16	Tc 1.9	Ru 2.2	Rh 2.28	Pd 2.2	Ag 1.93	Cd 1.69	In 1.78	Sn 1.96	Sb 2.05	Te 2.1	I 2.66	Xe 2.6
Cs 0.79	Ba 0.89	La 1.1	Hf 1.3	Ta 1.5	W 2.36	Re 1.9	Os 2.2	Ir 2.2	Pt 2.2	Au 2.4	Hg 1.9	Tl 1.8	Pb 1.8	Bi 1.9	Po 2.0	At 2.2	Rn
Fr 0.7	Ra 0.9	Ac 1.1															

* Electronegativity & Bond Polarity :

* The bond polarity depends on the elements' electronegativities.

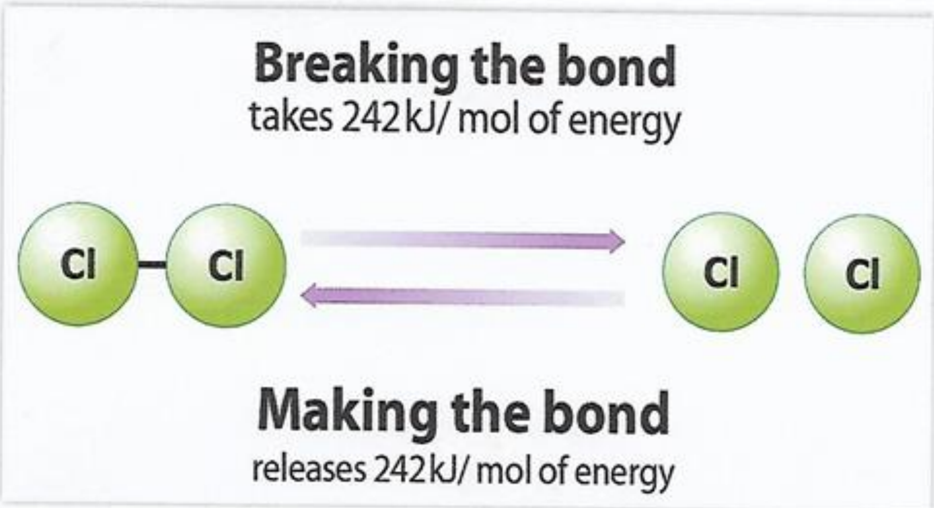
Type of bond

$\Delta EN = 0$	$\Delta EN = 0.1-0.4$	$\Delta EN = 0.5-1.9$	$\Delta EN \geq 2.0$
∴ Pure Covalent (Equal sharing of e⁻'s)	Nonpolar Covalent	Polar Covalent	Ionic (Transfer of e⁻'s)
e.g. H — H	Cl — Br	$\overset{\delta+}{H}-\overset{\delta-}{Cl}$	$Na^+ \; Cl^-$
$\Delta EN = 0$	$\Delta EN = 0.2$	$\Delta EN = 0.96$	$\Delta EN = 2.23$

* Bond Energies :

The amount of energy, in gaseous state, that it takes to break 1 mol of a bond in a compound.

In general, Chemical reactions involve breaking of bonds in reactants and making new bonds in products.



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In general, the more electrons two atoms share, the stronger the covalent bond.

$$C{\equiv}C\ (837\ kJ) > C{=}C\ (611\ kJ) > C{-}C\ (347\ kJ)$$
$$C{\equiv}N\ (891\ kJ) > C{=}N\ (615\ kJ) > C{-}N\ (305\ kJ)$$

* Trends in bond strength:

* Bond Lengths :

Bond Length is the distance between the nuclei of bonded atoms.

In general,

lewis theory predicts that the more electrons two atom share, the shorter the bond should be when comparing bonds of like atoms.

Bond length : (pm)

$$\underset{(120)}{C{\equiv}C} < \underset{(134)}{C{=}C} < \underset{(154)}{C{-}C}$$

Bond Length for halogens :

$$\underset{(143)}{F{-}F} < \underset{(199)}{Cl{-}Cl} < \underset{(228)}{Br{-}Br} < \underset{(266)}{I{-}I}$$

* Trends in bond length :

By Comparing the average bond length of a bond between two particular atoms in a variety of Compounds.

* Bond length decreases across the period:

$$\underset{154}{C-C} > \underset{147}{C-N} > \underset{143\ pm}{C-O}$$

* Bond length increases down the group :

$$\underset{144}{F-F} < \underset{198}{Cl-Cl} < \underset{228\ pm}{Br-Br}$$

* In general, as bonds get longer, they also get weaker.

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Chapter 5 : Aqueous Solutions & Acid-Base Equilibria

Lesson 17 : Chemical Equilibrium

* Introduction :

Ammonia (NH_3) is manufactured from the reversible reaction between hydrogen and nitrogen gases

$$N_{2(g)} + 3\,H_{2(g)} \rightleftharpoons 2\,NH_{3\,(g)}$$

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* Reversible reactions :

They are the chemical reactions where the reactants form products that, in turn, react together to give the reactants back. These reversible reactions are indicated by a double-headed arrow (forward & backward directions)

* Chemical equilibrium :

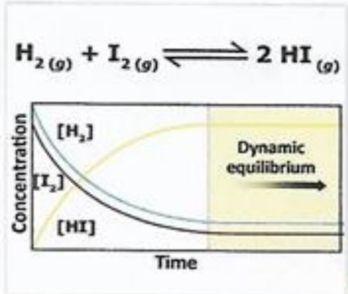
It is the point at which the rate of forward and that of backward are equal.

At equilibrium : $\boxed{\text{Rate of forward} = \text{Rate of backward}}$

* Dynamic equilibrium :

For a reversible reaction under equilibrium At first, the concentration of reactants decreases till reach the equilibrium, then the reverse reaction will start to increase till reach the state of dynamic equilibrium.

* Under Chemical or dynamic equilibrium, the equilibrium (forward and backward) rates are equal but oftenly the concentrations of reactants and products are not.



* Representing equilibrium :

In equilibrium; the molar concentrations are shown in brackets.

e.g. $[H_2]$ for hydrogen gas ; $[CaCl_2]$ in aqueous solution.

Consider: $$H_{2(g)} + I_{2(g)} \rightleftharpoons 2HI_{(g)}$$

$[H_2]$ and $[I_2]$ decrease with time and $[HI]$ increases, till an equilibrium is reached at:

$$\underset{(K_f)}{\text{Rate of forward}} = \underset{(K_r)}{\text{Rate of backward}}$$

Dynamic equilibrium can be expressed numerically by "equilibrium Constant"

For a general reaction:

$$aA + bB \rightleftharpoons cC + dD$$

Where; a, b, c & d are the relative stoichiometric coefficients in the chemical equation.

* Equilibrium Constant : K

It is the ratio, at equilibrium, of the concentrations of the products raised to their coefficients divided by the concentrations of the reactants raised to their coefficients.

$$\boxed{K_{eq} = \frac{[C]^c\,[D]^d}{[A]^a\,[B]^b}}$$

Products

Reactants

* example ;

1) $2\,N_2O_{5\,(g)} \rightleftharpoons 4\,NO_{2\,(g)} + O_{2\,(g)}$

$$K_{eq} = \frac{[NO_2]^4\,[O_2]}{[N_2O_5]^2}$$

2) $CH_3OH_{(g)} \rightleftharpoons CO_{(g)} + 2H_{2\,(g)}$

$$K = \frac{[CO]\,[H_2]^2}{[CH_3OH]}$$

* Significance of the equilibrium constant (K) :

$K_{eq} \ll 1$

- forward reaction doesn't proceed (More Reactants)
- Equilibrium position favors reactants (to left).

$K_{eq} \simeq 1$

- Neither direction is favored.
- Equilibrium position isn't favor any direction

$K_{eq} \gg 1$

- forward reaction proceeds to Completion (more Products)
- Equilibrium position favors products (to right).

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* Equilibrium Constant and Chemical Equations :

* Consider the chemical equilibrium : $A + 2B \rightleftharpoons 3C$;

$$K_{forward} = \frac{[C]^3}{[A][B]^2}$$; For : $3C \rightleftharpoons A + 2B$

- Equilibrium constant for the reverse equation : $\boxed{K_{reverse} = \frac{1}{K_{forward}}}$

- Multiplying the equation by factor (n) :

$nA + 2nB \rightleftharpoons 3nC$; $\boxed{K' = \frac{[C]^{3n}}{[A]^n [B]^{2n}} = K^n}$

- Adding two equations of K_1 & K_2 ; the overall equilibrium constant is their multiplication.

$A \rightleftharpoons 2B$; $K_1 = [B]^2/[A]$

$2B \rightleftharpoons 3C$; $K_2 = [C]^3/[B]^2$

$A \rightleftharpoons 3C$; $K_{overall} = K_1 \times K_2 = \frac{[C]^3}{[A]}$

---

* Equilibrium Constant and Pressure :

For gaseous reactions ; partial pressure of gas is proportional to its concentration ; so, the equilibrium constant (K_p) is given as :

e.g. $H_{2\,(g)} + I_{2\,(g)} \rightleftharpoons 2HI_{(g)}$

$$K_c = \frac{[HI]^2}{[H_2][I_2]} \quad \& \quad K_p = \frac{P_{HI}^2}{P_{H_2} \cdot P_{I_2}}$$

* Calculate K_c for the following reaction :

$$H_{2(g)} + I_{2(g)} \rightleftharpoons 2HI_{(g)}$$

$[H_2] = 0.22$ M ; $[I_2] = 0.22$ M ; $[HI] = 1.56$ M

A.

$$K_c = \frac{[HI]^2}{[H_2][I_2]} = \frac{(1.56)^2}{(0.22)(0.22)} = 50.28 \; ; \; K_c \gg 1$$

forward is favored.

* Reactions involving Solids :

For : (1) $2CO_{(g)} \rightleftharpoons CO_{2(g)} + C_{(s)}$

In equilibrium; the concentration of solid isn't change (also, pure liquid) so, equilibrium constant doesn't imply the concentrations of pure solids or liquids.

$$K_c = \frac{[CO_2][C]}{[CO]^2} \text{ (incorrect)} \; ; \; K_c = \frac{[CO_2]}{[CO]^2} \text{ (correct)}$$

(2) $CO_{2(g)} + H_2O_{(L)} \rightleftharpoons H^+_{(aq.)} + HCO_3^-{}_{(aq.)}$

$$K_c = \frac{[H^+][HCO_3^-]}{[CO_2]}$$

(3) $CaCO_{3(s)} \rightleftharpoons CaO_{(s)} + CO_{2(g)}$; $K_c = [CO_2]$

(4) $2CO_{(g)} \rightleftharpoons CO_{2(g)} + C_{(s)}$; $K_c = \frac{[CO_2]}{[CO]^2}$

Q. What would the K_c value be given the following concentrations at equilibrium for the chemical reaction at room temperature?

$$2\,NOCl_{(g)} \rightleftharpoons 2\,NO_{(g)} + Cl_{2\,(g)}$$

$[NOCl] = 1.34\ M$; $[NO] = 0.66\ M$; $[Cl_2] = 0.33$

A.

$$K_c = \frac{[NO]^2[Cl_2]}{[NOCl]^2} = \frac{(0.66)^2(0.33)}{(1.34)^2} = 0.08 \quad \ll 1$$

Backword direction is favored. (left)

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Chapter 5 : Aqueous Solutions & Acid-Base Equilibria
Lesson 18 : Le Châtelier's Principle

* Le chatelier's Principle:

When a chemical system at equilibrium is disturbed (Changed), the system shifts in a direction that minimizes that disturbance. so, a system tends to maintain the equilibrium state.

* Factors that can disturb the system at equilibrium :

* Changing the concentration of reactant or product.
* Changing the volume (or Pressure).
* Changing the temperature.

1) Equilibrium and Concentration Change :

Changing the concentration of a reactant or product will disturb the equilibrium state.

« Reaction Quotient » Q

For a reaction : $aA + bB \rightleftharpoons cC + dD$

at equilibrium

$$K_{eq.} = \frac{[C]^c\,[D]^d}{[A]^a\,[B]^b}$$

at any given time

$$Q = \frac{[C]^c\,[D]^d}{[A]^a\,[B]^b}$$

(✱) **K** describes the reaction that is at equilibrium, while
Q " " " is at any state.

(✱) So, the reaction quotient (Q) is a comparable value to the equilibrium constant (K) and it is used to determine if the system at equilibrium or not.

➤ For a reaction:

$$N_2O_{4(g)} \rightleftharpoons 2\,NO_{2(g)}$$

Adding $[N_2O_4]$		Adding $[NO_2]$
Not at equilibrium	At equilibrium	Not at equilibrium
$Q < K$	$Q = K$	$Q > K$
- Reaction shifts right to decrease $[N_2O_4]$ and reach the equilibrium.	$Q = \frac{[NO_2]^2}{[N_2O_4]} = K$	- Reaction shifts left to decrease $[NO_2]$ and reach the equilibrium.

2) Equilibrium and Pressure (volume) change:

- Increasing the pressure (or decreasing the volume) will cause the system (reaction) to shift to the side of the fewest gas moles and vice versa

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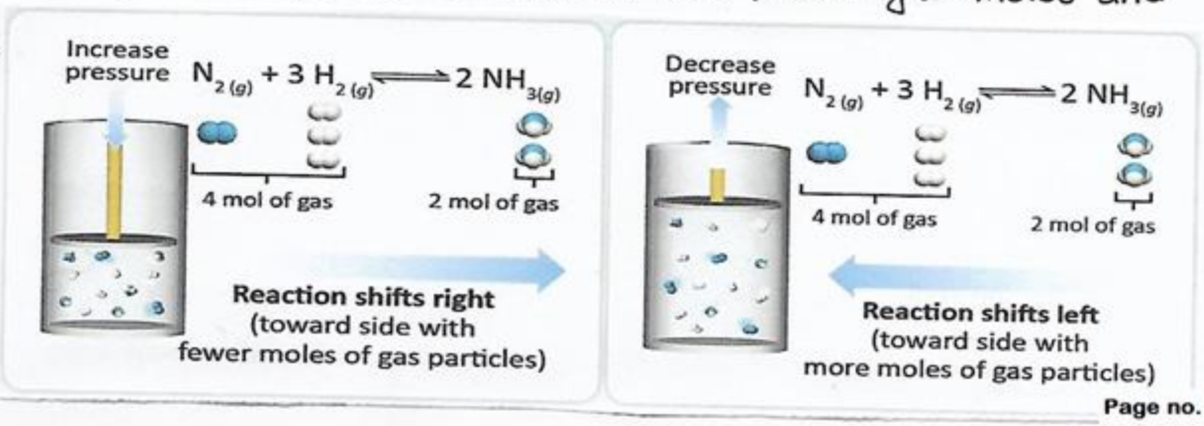

N.B.

For a reaction with equal number of gaseoues moles on both sides, changing the pressure will not cause shifting of the reaction in any direction.

$$2\,HBr_{(g)} \rightleftharpoons H_{2\,(g)} + Br_{2\,(g)}$$

3) Equilibrium and Temperature change:

Types of thermal Reactions

"Exothermic" $\Delta H = -ve$

$$A + B \longrightarrow C + D + \underset{(product)}{heat}$$

Thus; adding heat (product) will shift the reaction to left.

e.g.

$$N_{2(g)} + 3H_{2(g)} \rightleftharpoons 2NH_{3(g)} + heat \quad \oplus$$

Adding heat shifting the reaction to left (smaller K).

OR

$$N_{2(g)} + 3H_{2(g)} \rightleftharpoons 2NH_{3(g)} + heat \quad \ominus$$

Removing heat shifting the reaction to right (larger K).

i.e. More Ammonia (NH_3).

"Endothermic" $\Delta H = +ve$

$$A + B + \underset{(Reactant)}{heat} \longrightarrow C + D$$

Thus; adding heat (as reactant) will shift the reaction to right (larger K).

e.g.

$$N_{2(g)} + O_{2(g)} + heat \rightleftharpoons 2NO_{(g)}$$

$\oplus \longrightarrow$ right ($K\uparrow$)

left ($K\downarrow$) $\longleftarrow \ominus$

Removing heat shifting the reaction to left (smaller K).

N.B. Equilibrium Constant (K) is temperature dependent.

يعتمد على درجة الحرارة فقط

* Nature of acids and bases :

Acids

General properties:

* Taste sour.
* React with active metals Al, Zn, Fe, but not Cu, Ag, Au.
* Corrosive
* React with carbonates, and producing CO_2 gas.

$$CaCO_3 + 2HCl \longrightarrow CaCl_2 + H_2O + CO_2$$

Bakingsoda or Chalk

* turn Litmus paper to Red.

Bases

General properties:

* Taste bitter.

AlKaloids are the basic plant product

* Solutions are slippery or gelatinous to touch.
* Red litmus paper turns Blue.

In general: Acid + Base ⟶ Ionic Salt + Water

* Arrhenius definition of acid and base :

- Acid : A substance that produces H^+ ions in aqueous solution.
- Base : " " " OH^- " " " .

Examples;

$$\underset{\text{Acid}}{HCl_{(aq.)}} \longrightarrow H^+_{(aq.)} + Cl^-_{(aq.)} \quad \& \quad NaOH_{(aq.)} \longrightarrow Na^+_{(aq.)} + OH^-_{(aq.)}$$

But Arrhenius definition doesn't explain ammonia (NH_3) as base.

* Bronsted-Lowry definition of acids and bases :

Acids

* Acids are proton (H^+) donors.

e.g.

$$HCl_{(aq.)} + H_2O_{(L)} \longrightarrow H_3O^+_{(aq.)} + Cl^-_{(aq.)}$$

Acid Base

Bases

* Bases are proton (H^+) acceptors.

e.g.

$$NH_{3\,(aq.)} + H_2O_{(L)} \longrightarrow NH_4^+{}_{(aq.)} + {}^-OH_{(aq.)}$$

Base Acid

N.B.

Some substances (like water, H_2O) can act as acid or base in some chemical reactions.

* Conjugate Acid-Base pairs :

The two substances that related to each other by transfer of a proton (H^+) are called a conjugate acid-base pair.

* Conjugate base : anything remains after removal of H^+ from acid.
* Conjugate acid : " " " removal of ^-OH from base (or accepting H^+ from acid).

$$H_2SO_{4\,(aq)} + H_2O_{(l)} \longrightarrow HSO^-_{4\,(aq)} + H_3O^+_{(aq)}$$

$H_2SO_{4\,(aq)}$	$H_2O_{(l)}$	$HSO^-_{4\,(aq)}$	$H_3O^+_{(aq)}$
Acid	Base	Conjugate base	Conjugate acid

$$HCO^-_{3\,(aq)} + H_2O_{(l)} \rightleftharpoons H_2CO_{3\,(aq)} + OH^-_{(aq)}$$

$HCO^-_{3\,(aq)}$	$H_2O_{(l)}$	$H_2CO_{3\,(aq)}$	$OH^-_{(aq)}$
Base	Acid	Conjugate acid	Conjugate base

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(*) Some chemical groups and their Conjugate acid and conjugate base that they can form :

More H^+ is acid.

Conjugate Acid		Conjugate Base
NH_4^+	NH_3	NH_2^-
H_3O^+	H_2O	OH^-
H_3PO_4	$H_2PO_4^-$	HPO_4^{2-}
H_2O	HO^-	O^{2-}

Less H^+ is base.

****

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Ch. 5 : Aqueous Solutions & Acid-Base Equilibria

Lesson 19 : Acid strength and pH scale

* Introduction:

Lemon juice (citric acid) is known as weak acid (weaker than HCl), so

Q. How can you measure and compare the acidity of acids?

A. Acidity can be measured using the pH scale.
we can identify the acid by using Litmus paper but to measure and compare its degree of acidity, pH-Scale is used.

Acids — Bronsted-Lowry "proton (H^+) donors"

"Strong Acid"

* It ionizes completely in water.

$$HCl_{(aq.)} + H_2O_{(l)} \longrightarrow H_3O^+_{(aq.)} + Cl^-_{(aq.)}$$

1.0 M → 1.0 M

* Single arrow → Complete ionization

All of HCl molecules ionize to produce H_3O^+.

* General equation (strong Acid)

$$\underset{\text{(strong Acid)}}{HA_{(aq.)}} + H_2O_{(l)} \longrightarrow H_3O^+_{(aq.)} + \underset{\text{(weak Conjugate base)}}{A^-_{(aq.)}}$$

"Weak Acid"

* It is partially ionized in water

$$HF_{(aq.)} + H_2O_{(l)} \rightleftharpoons H_3O^+_{(aq.)} + F^-_{(aq.)}$$

1.0 M → << 1.0 M

* Equilibrium arrow ($\rightleftharpoons$) → partial ionization

Only some of HF molecules ionize to produce H_3O^+.

* General equation (weak Acid)

$$\underset{\text{(weak acid)}}{HA_{(aq.)}} + H_2O_{(l)} \rightleftharpoons H_3O^+_{(aq.)} + \underset{\text{(strong Conjugate base)}}{A^-_{(aq.)}}$$

N.B. Every strong acid has weak conjugate base &
" weak " " strong " " .

* Diprotic acid :
The acid which has two ionizable protons.

⇒ e.g. H_2SO_3 (sulforous acid); H_2CO_3 (Carbonic acid)

Both are diprotic acids (produce two protons H^+ in water).

Thus;

* HCl, HNO_3 (are monoprotic acids)
* H_2SO_3, H_2SO_4, H_2CO_3 are diprotic acids.
* H_3PO_4 (phosphoric acid) is triprotic acid.

Very Imp. Note: H_2SO_4 (diprotic), its 1st ionization is strong but the 2nd is weak:

- 1st-ionization : $H_2SO_{4(aq.)} + H_2O_{(l)} \longrightarrow H_3O^+_{(aq.)} + HSO^-_{4(aq)}$
- 2nd-ionization : $HSO^-_{4(aq.)} + H_2O_{(l)} \rightleftharpoons H_3O^+_{(aq.)} + SO^{--}_{4(aq.)}$

pH-Scale

It is a compact way (scale) to specify and compare the acidity of solution.

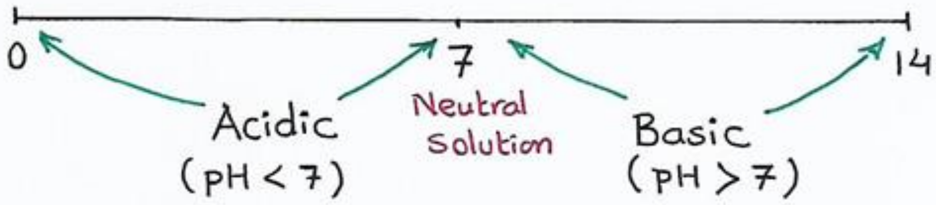
$$\boxed{pH = -\log[H_3O^+]} \text{ For Acids}$$

Example;

For a solution of $[H_3O^+] = 10^{-3}$ M

$\therefore pH = -\log[H_3O^+] = -\log(10^{-3}) = 3$

* pH Scale:



(N.B.) - pH unit corresponds to a 10-fold change in $[H_3O^+]$

- Highly Concentrated acids can have -ve pH-value.

e.g.

if $[H_3O^+] = 2\ M \implies pH = -\log[H_3O^+] = -\log(2) = -0.3$

* POH Scale:

It is analogous to pH Scale, but based on $[OH^-]$ instead of hydronium ion $[H_3O^+]$.

$$\boxed{pOH = -\log[\bar{O}H]}$$ For Bases.

* pH & pOH relationship:

① $\boxed{pH + pOH = 14}$

② $\boxed{[H_3O^+][OH^-] = 10^{-14}}$

Q. Calculate pH of a solution at 25°C with $[OH^-] = 1.3\times10^{-2}\ M$

A. For Base $[OH^-]$

$pOH = -\log[O\bar{H}] = -\log(1.3\times10^{-2}) = 1.89$

But $pH + pOH = 14$

$pH = 14 - pOH = 14 - 1.89 = 12.11$ > 7 Basic

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Q. Calculate the H_3O^+ concentration for a solution with $pH = 4.80$.

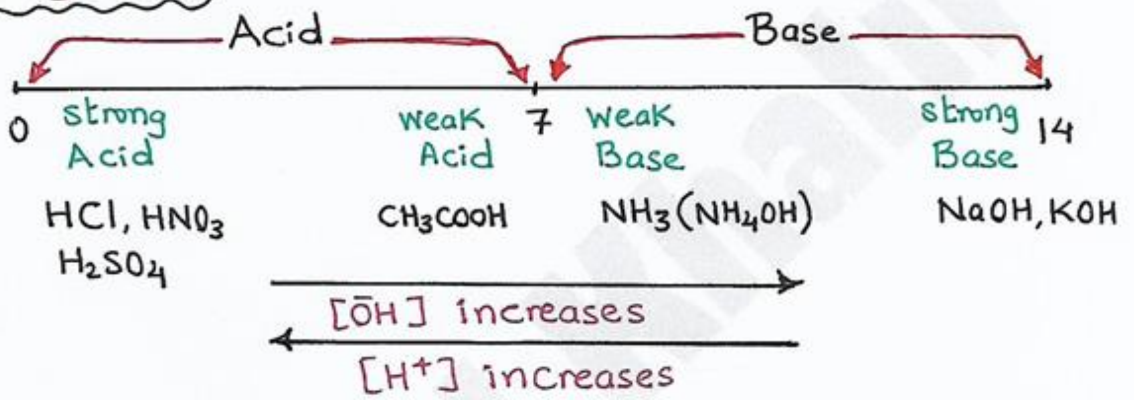
A.

$\because \; pH = -\log [H_3O^+]$

$-pH = \log [H_3O^+] \Rightarrow [H_3O^+] = 10^{-pH} = 10^{-(4.80)}$

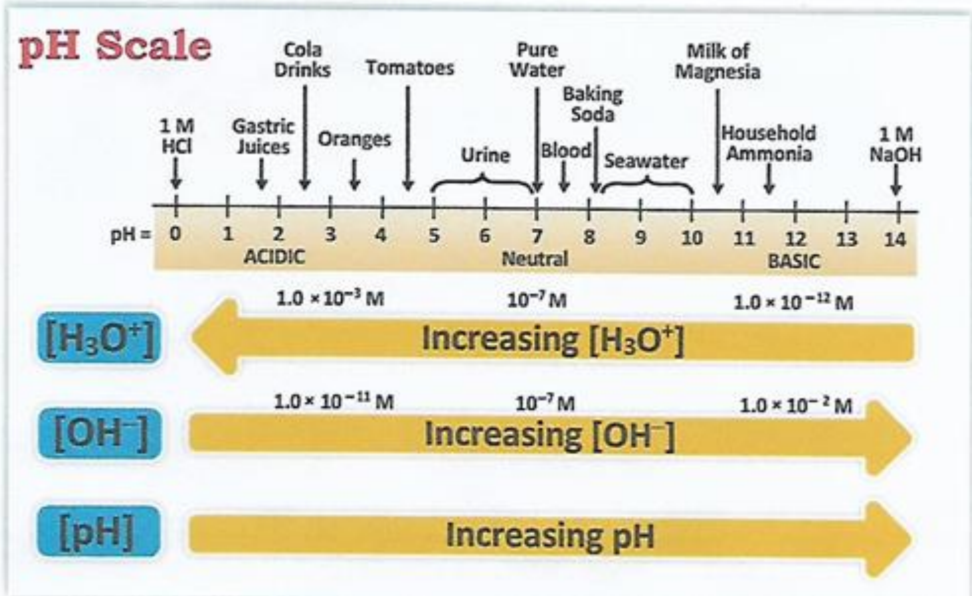
$= 1.6 \times 10^{-5}$ M

* Measuring pH :



N.B.

* As $[H^+]\uparrow$ $pH\downarrow$ $\rightarrow$ strong acid, e.g. HCl
* As $[H^+]\downarrow$ $pH\uparrow$ $\rightarrow$ weak acid. e.g. Acetic acid CH_3COOH.



د. خالد خليل
قسم الكيمياء - كلية العلوم

Bases

Bronsted. Lowry
" proton (H^+) acceptor "

"Strong Base"

* It ionizes Completely in water.

$$NaOH_{(aq.)} \longrightarrow Na^+_{(aq.)} + OH^-_{(aq.)}$$

* All of NaOH molecules ionize to produce $\bar{O}H$ - ions.

"Weak Base"

* It is partially ionzed in water.

$$NH_{3(aq.)} + H_2O_{(l)} \rightleftharpoons NH^+_{4(aq.)} + OH^-_{(aq.)}$$

* Only Some $\bar{O}H$ are produced.

Lewis Acids & Bases definitions

Acid = electron pair acceptor & Base = electron pair donor.

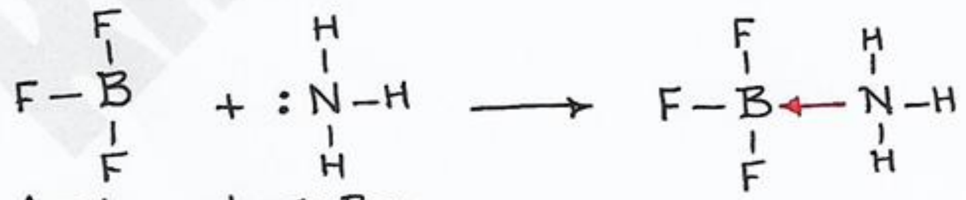
e.g.

$$H^+ + :NH_3 \longrightarrow [H-NH_3]^+ \quad \text{or} \quad NH_4^+$$

Ammonium

According to Lewis model;

A substance doesn't need to Contain hydrogen to be an Acid.



Lewis Acid (e⁻ pair acceptor) Lewis Base (e⁻ pair donor) Adduct

* In general, Lewis acid has an empty orbital that can accept the e⁻ pair from Lewis base.
* Lewis acids : BF_3, $FeCl_3$, $AlCl_3$ (e⁻ poor).
* Lewis bases : NH_3, OH^- (e⁻ rich).

Ch. 6: Reaction Kinetics & Thermodynamics

Lesson 20: Energy, work, and Heat

- « Law of Conservation of energy » or "first law of thermodynamics"

  The energy of Universe is Constant; Energy can be neither created nor destroyed.

- « Work »

  It is the energy transfer by a force acting through a distance.

- Energy:

  It is the capacity to do work.

Energy

Kinetic energy	Potential energy
It is the energy of motion of an object.	It is the stored energy in an object due to its position

- Heat:

  It is the thermal energy transferred across a boundary of one region of matter to another.

- SI - unit for work, energy, and heat is the Joule (J)

1 Calorie = 4.184 Joule

$$\boxed{1 \text{ Cal.} = 4.184 \text{ J}}$$

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* System and surroundings :

* System is the region in space in which a definite quantity of matter is studied energetically.
* Surroundings : are everything else or the rest of the universe with which the system can exchange energy.

* Thermodynamics :

It is the study of energy that is exchanged between the system and the surroundings.

$\text{System} \xrightarrow{E} \text{Surroundings}$ « Exothermic »

$\text{System} \xleftarrow{E} \text{Surroundings}$ « Endothermic »

From 1^{st} law of thermodynamics :

" Energy of the universe is Constant " $\therefore \Delta \text{Energy}_{(universe)} = 0$

$$\Delta E_{system} + \Delta E_{surroundings} = 0$$

* Internal Energy :

It is the sum of Kinetic and potential energies of all the particles that compose a system.

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* Change in internal energy (ΔE) depends only on the amount of energy in the system at the begining and at the end.

$$\boxed{\Delta E = E_{final} - E_{initial}}$$

* State of the system (chemical reaction) is specified (described) by factors like temperature, pressure, concentration and phase.

* State function :

It is a mathematical function that depends only on the initial and final conditions (and not the process or path).

$$\Delta E = E_{final} - E_{Initial}$$

$$\boxed{\Delta E_{Reaction} = E_{products} - E_{reactants}}$$

* "Energy exchange"

Energy is exchanged between a system and its surroundings through heat and work.

q = heat (transfer of thermal energy).

w = work.

where q and w are not state functions and depend on the process.

$$\boxed{\Delta E = q + w}$$

$$\text{System} \underset{\text{Work }(w)}{\overset{\text{Heat }(q)}{\longleftrightarrow}} \text{Surroundings}$$

* Sign Convention for q, w, and ΔE : V. Imp.

q (heat)	+ system *gains* thermal energy	- system *loses* thermal energy
w (work)	+ work done *on* the system	- work done *by* the system
Δ*E* (change in internal energy)	+ energy flows *into* the system	- energy flows *out* of the system

* Heat exchange :

Heat is the exchange of thermal energy between a system and its surroundings caused by temperature change.

- Temperature : It is the measure of the thermal energy within a sample of matter.

- Heat exchange : It occurs when a system and its surroundings have a difference in temperature.

- Heat Flow : Heat flows from the matter of high temperature to that of low temperature till reach same temperature (or thermal equilibrium).

Ch.6: Reaction Kinetic & Thermodynamics

Lesson 21: Heat Capacity

* Reaction Kinetics :

It means how fast the reaction goes from reactants to products. Reaction Kinetics describe :

* Reaction rate (Fast or Slow).
* Reaction Mechanism (Steps of the reaction).

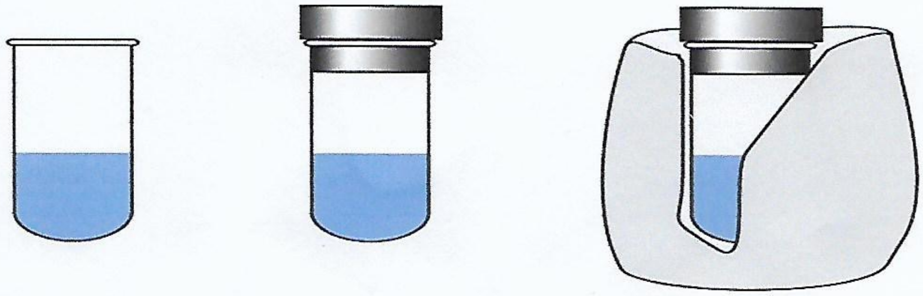
* Thermodynamics :

It deals with the reaction's Energy requirements. It tells us if the reaction is favored or not.

* Types of systems and the transfer of energy and matter :

Transfer between system and surrounding

System	Energy	Matter
Open	Yes	Yes
Closed	Yes	No
Isolated	No	No



Open Closed Isolated

* Internal energy : E

$$\boxed{E = PE + KE}$$

PE : potential energy ; KE : Kinetic energy

But

Internal energy change is state function :

$$\Delta E = E_{final} - E_{Initial}$$

For chemical reactions;

$$\boxed{\Delta E = E_{products} - E_{reactants}}$$

ΔE

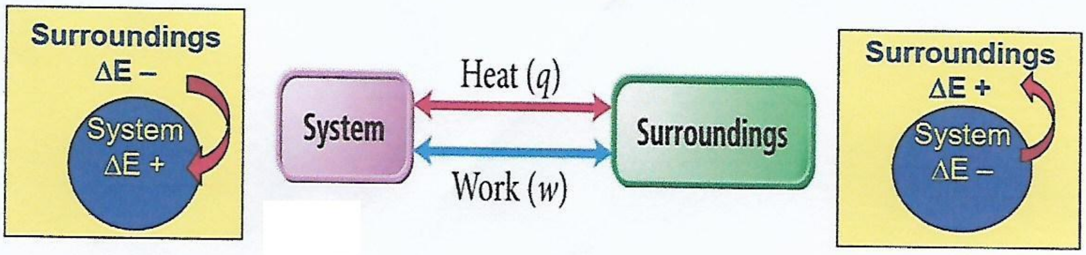
$E_{products} < E_{Reactants}$

$\Delta E = -ve$

$E_{products} > E_{Reactants}$

$\Delta E = +ve$

* Energy exchange between the system and surroundings :



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Thus, we have to understand how the Heat (q) and work (W) are exchanged between the system and surroundings.

* Heat (q):

It is the exchange of thermal energy between a system and surrounding. by the equation:

$$\boxed{q = C \times \Delta T}$$

heat capacity — Temperature change

* Heat Capacity (C):

It is the quantity of heat required to change the temperature of a system by 1°C.

$$C = \frac{q}{\Delta T} = J/°C$$

* Specific heat Capacity (C_s):

It is the quantity of heat required to change the temperature of 1 g of substance by 1°C (J/g.°C).

$$\boxed{q = m \times C_s \times \Delta T} \quad (1)$$

* Work (Pressure-volume work): (W)

In chemical reactions, Pressure-volume work occurs by force which is accompained with change in volume against the external pressure:

$$\boxed{\Delta W = -P\Delta V} \quad (2)$$

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N.B. (P-V) work is calculated in units of atm.L

So, to convert it to Joule, we multiply by 101.3 because

$$101.3\ J = 1\ atm.L$$

Q. If a ballon is inflated from 0.100 L to 1.85 L against an external pressure of 1 atm, how much work is done?

A. Given: $P = 1\ atm$; $V_1 = 0.10\ L$; $V_2 = 1.85 L$

$\because \quad W = -P\,\Delta V$

$\Delta V = V_2 - V_1 = 1.85 - 0.1 = 1.75 L$

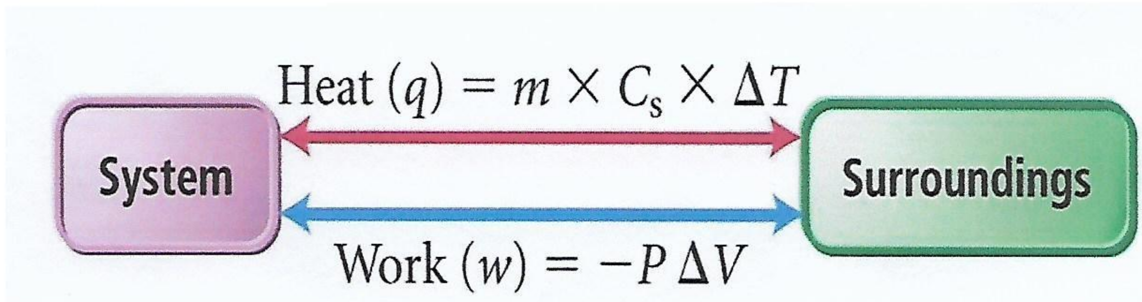
$\therefore W = -(1\ atm)(1.75\ L) = -1.75\ atm.L$

$\because \quad 1\ atm.L \longrightarrow 101.3\ J$

$\quad -1.75 \longrightarrow ?$

$\therefore W = -1.75 \times 101.3 = -177\ J$

Energy exchange $(q + W)$



* Measuring ΔE for chemical reaction under constant volume :

At constant volume ; $\Delta V = 0$; so $W = 0$ $(W = -P\Delta V)$

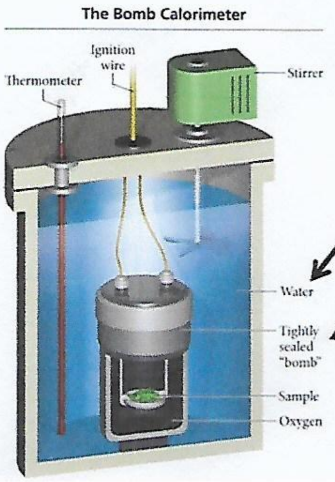
$$\because \Delta E_{rxn} = q_v + W$$

$$\boxed{\Delta E_{rxn} = q_v}$$

In Calorimetry; the exchanged thermal energy between the system (reaction) and the surrounding is measured by observing ΔT using Bomb Calorimetry.

Bomb Calorimetry

The basic principals involved in measuring the heat of Combustion (calorific value) of solid and liquid fuels, foodstuffs and other combustible materials in a bomb calorimeter.



Some heat from the reaction warms water; therefore:

q_{water} = (specific heat)(water mass)(ΔT)

Some heat from the reaction warms the "calorimeter bomb"; therefore:

q_{bomb} = (heat capacity, J/°C)(ΔT)

So, total heat evolved

$q_{total} = q_{water} + q_{bomb}$

Ch.6 : Reaction Kinetics & Thermodynamics

Lesson 22: Enthalpy & Catalysis

* Enthalpy (H) :

Enthalpy (heat content) is the sum of internal energy (E) and the product of its pressure and volume :

$$\boxed{H = E + PV}$$

Where;

Internal energy (E), pressure, volume and H are all state functions.

At constant pressure ; ΔH (heat evolved) of a chemical reaction is given by :

$$\Delta H = \Delta E + P\Delta V$$

but $\Rightarrow W = -P\Delta V$

$$= (q_p + W) + P\Delta V$$

$$= q_p + W - W$$

$$\therefore \boxed{\Delta H = q_p}$$

Thus, at constant pressure

ΔH of the reaction can be calculated by observing the change in ΔT (q)

Q. A piece of zinc weighing 35.8 g was heated from 20.00 °C to 28.00 °C. How much heat was required? The specific heat of zinc is 0.388 J/g.°C .

A. Given: $m = 35.8\ g$; $C_s = 0.388\ J/g.^\circ C$; $\Delta T = 28 - 20 = 8^\circ C$.

$$\because\ q = m \cdot C_s \cdot \Delta T = (35.8\ g)\left(0.388 \frac{J}{g.^\circ C}\right)(8^\circ C)$$

$$= 111\ J$$

* Enthalpy change of chemical reaction :

ΔH is state function and thermo property of matter (Extensive physical property).

For chemical reaction: $\Delta H = H_{final} - H_{initial}$

$$= H_{products} - H_{reactants}$$

ΔH

$H_{final} < H_{initial}$	$H_{final} > H_{initial}$
$\Delta H = -ve$	$\Delta H = +ve$
Exothermic reaction	Endothermic reaction

* Hess's law : For chemical reaction:

$$\boxed{\Delta H_{rxn} = \sum \Delta H_{(products)} - \sum \Delta H_{(reactants)}}$$

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* A Comparison between endothermic and exothermic reactions :

"Endothermic"	"Exothermic"
- The reaction vessel cools.	- the reaction vessel warms.
- Heat is absorbed.	- Heat is evolved.
- Energy is added to system.	- Energy is subtracted from the system.
- q is +ve sign.	- q is -ve sign.

* Q. Classify each process as exothermic or endothermic

- Your hand gets cold when you touch ice.
  (heat leaves your hand and moves to ice) ∴ Exothermic
- The ice gets warmer when you touch it.
  (heat flows into the ice) ∴ Endothermic
- Water boils in a Kettle being heated on a stove.
  (heat flows into water to boil it) ∴ Endothermic
- Water vapor condenses on a cold pipe.
  (heat leaves water vapor to be condensed to liquid). Exothermic
- Ice cream melts.
  (heat flows into the ice cream to melt it) Endothermic

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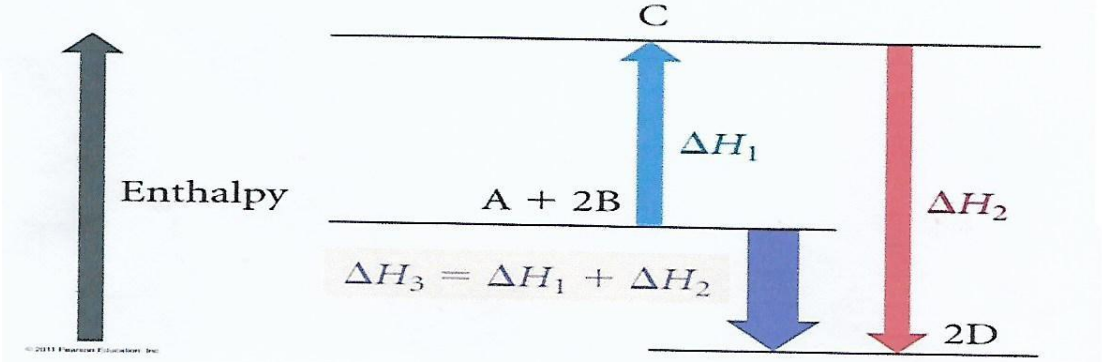
Hess's Law

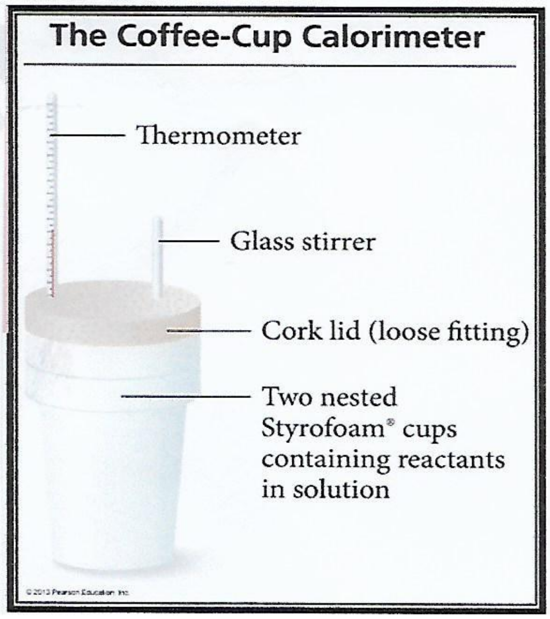
The change in enthalpy for a stepwise process is the sum of the enthalpy changes of the steps.



* Constant pressure Calorimetry: Measuring ΔH_{rxn}

Coffee - Cup Calorimeter is used to measure the enthalpy change (ΔH) for chemical reactions in solution.

$$q_{reaction} = -\,q_{solution}$$
$$= -(mass_{soln.} \times Cs_{soln.} \times \Delta T)$$



- Heat evolved (or absorbed) causes a temperature change in the solution.

Thus, the heat gained by a solution equals to that lost by the reaction and vice versa.

$$\therefore\; q_{rxn} = q_{soln}$$

At constant pressure

$$\Delta H_{rxn} = q_p = q_{rxn}$$

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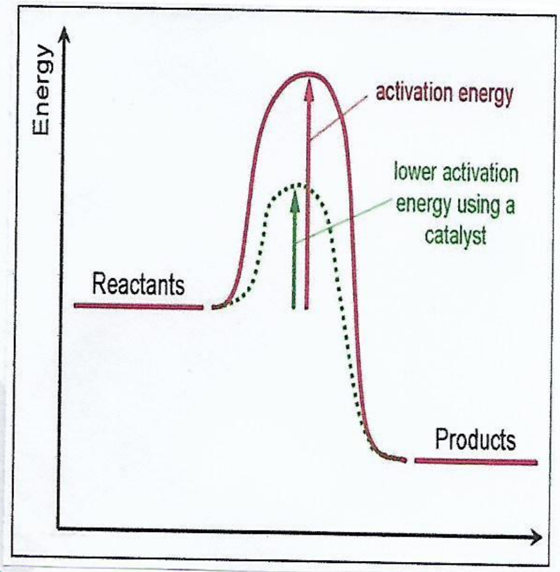
* Catalysis :

A Catalyst is a substance that accelerates the rate of a chemical reaction without itself being transformed or consumed.

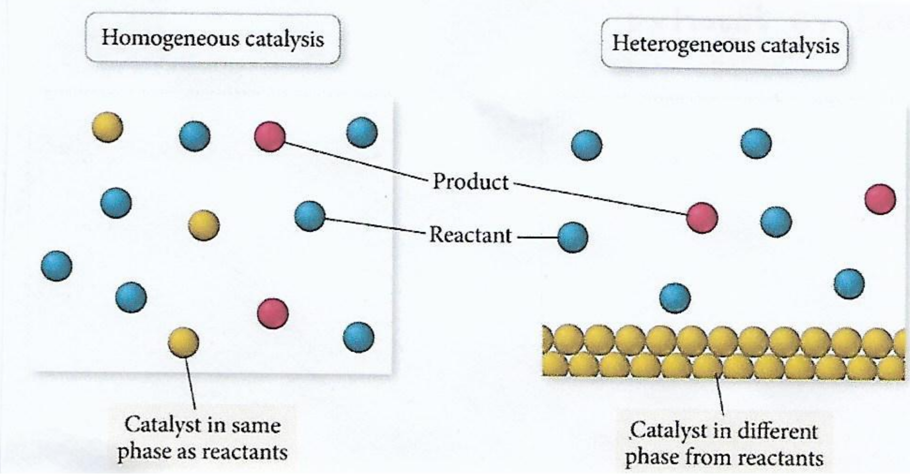
* Role of Catalyst is to provide an alternative mechanism for the reaction with lower activation energy.



* Catalyst is Consumed in the early mechanism step, then made in a later step.

Catalysts

Homogeneous Catalysts	Heterogeneous Catalysts
- They are in the same phase as the reactants.	- They are in a different phase as the reactants.
e.g. $Cl_{(g)}$ in destruction of $O_{3(g)}$	e.g. Solid Catalytic Converter in a car's exhaust system.

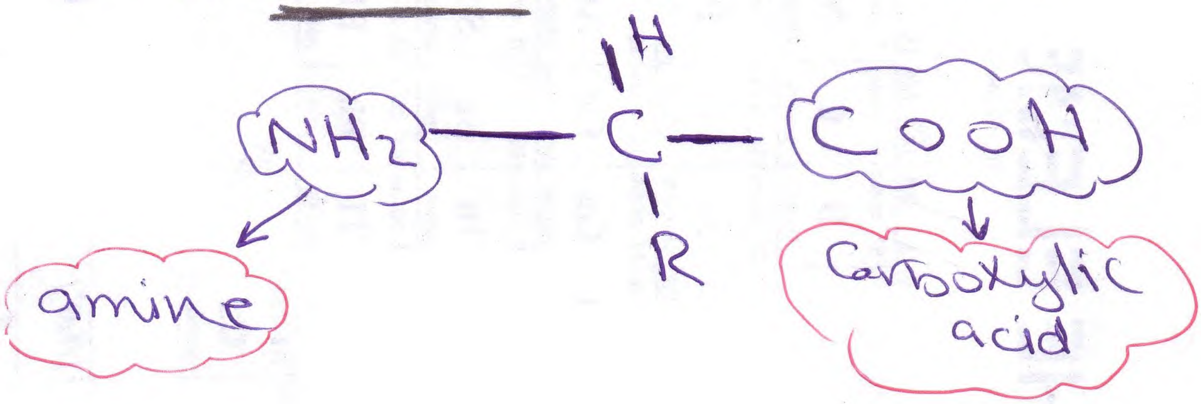


Ch. 7 Part 2 — BioChemistry الكيمياء الحيوية

Summary

(1) Proteins :

* Amino acids :- The Building Blocks of Proteins :-



* Peptide bond : or → (Amide Linkage)

it is The amide group :

→ Formed by Condensation Reaction تفاعل التكاثف

Takes Place between The –COOH group of one amino acid with The –NH_2 group of another amino acid. →

examples : Alanine and glycine

- Proteins are Polymers of amino acids Linked by The Peptide Bond.

- PolyPeptides :
  Formed when a large Number of amino acids are linked by Peptide bonds.

- Proteins : are linear (non-branched) PolyPeptides.

② Carbohydrates : الكربوهيدرات
(hydrats of Carbon: CHO)

Building Blocks of Carbohydrates are monosaccharides :

Classes (Types)

① monosaccharides :
(The most simple)
Can't be broken-down to smaller Carbohydrates :

e.g : ① hexoses:
glucose, Fructose, galactose

② Pentoses:
Ribose, deoxiribose

② Disaccharides
* Sucrose : Table sugar
(glucose + Fructose)
* Lactose : milk sugar
(glucose + galactose)
* Maltose : malt sugar
(glucose + glucose)

③ Poly-saccarides
① From animal :-
ex: glycogen
② From Plants :-
Starch
Cellulose

③ Lipids :- الدهون

non Polar organic Compounds :-

made From:

The Building Blocks of Lipids are:-

Fatty Acids + glycerol

Linked by: Ester Linkage.

* Triglycerides:-

The most Common glycerides and used For energy storage in plants and animals:- (two types):-

Butter (مثل الزبدة) الدهون ① Fats: Solid → From animals

(مثل الزيوت النباتية) الزيوت ② oils: Liquid → From plants.

④ Nucleic Acids :- الأحماض النووية

molecules That store information in The nucleous of The Cell.

→ Consist of:

Nitrogen Base + sugar + Phosphate (only in Nucleotides)

القاعدة النيتروجينيه

I) Nitrogen Bases:

a) Purines:

- **Adenine** (A)
  - in: DNA & RNA
- **Guanine** (G)
  - in: DNA & RNA

b) Pyrimidines

- **Thymine** (T)
  - in: DNA only
- **Cytosine** (C)
  - in: DNA & RNA
- **Uracil** (U)
  - in: RNA only

*note

- **Nucleosides:**
  Nitrogen Base + Sugar

- **Nucleotides:** → Bulding Blocks of Nucleic Acids
  Nitrogen Base + Sugar + Phosphate

II) Pentose Sugar :

- Ribose → in RNA
- Deoxy Ribose → in DNA

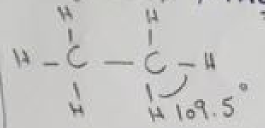
Organic and Biological Chemistry:-

* Hydrocarbons « H + C »

→ Alkane (C_nH_{2n+2})
→ Alkene (C_nH_{2n})
→ Alkyne (C_nH_{2n-2})
→ Aromatic (C_nH_n)

« nonPolar covelent »

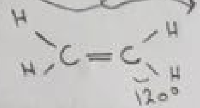
→ **Alkane**



- Single bond
- saturated with H → مشبع
- tetra Hydral shape → هرم رباعي
- C_nH_{2n+2}
- sp^3 → hyPrazation

« يحصل له دوران » rotation

→ **Alkene**



- Double bond
- unsaturated → غير مشبع
- C_nH_{2n}
- sp^2

→ **Alkyne**

180°
$H-C\equiv C-H$

- Trouble bond
- unsaturated
- C_nH_{2n-2}
- sp

→ Ar.H
« Aromatic »

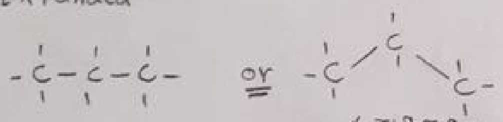
* (IUPAC)

Prefix + base + suffix.

* isomers:- have the same molecular furmola but different stractural formal (atoms are bonded in different order.

* ways to show bonding connection:-

1. Expanded

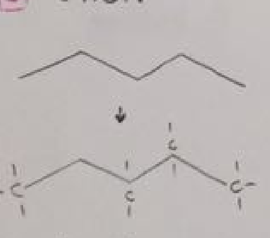


« zigzag »

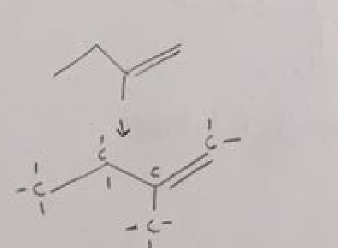
2. Condensed

$CH_3CH_2CH_2CH_3$ or $CH_3(CH_2)_2CH_3$

3. stick

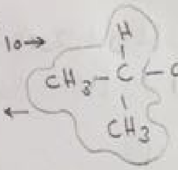


$CH_3CH_2CH_2CH_3$



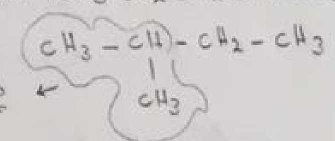
$CH_3-CH_2-C(CH_3)=CH_2$

- butane
  → normal butane → C_4H_{10} ($CH_3CH_2CH_2CH_3$)
  → iso butane → C_4H_{10} →



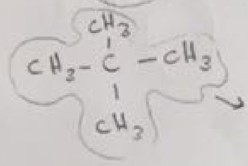
iso ← اسمه

- Pentane
  → normal $CH_3CH_2CH_2CH_2CH_3$
  → iso



iso ←

  → neo



→ neo

* البيوتان : 2 isomers
* البنتان : 3 isomers

⟸ بدلاً من اختراع common names الـ isomers

فى التسميه نستخدم الـ (IUPAC)

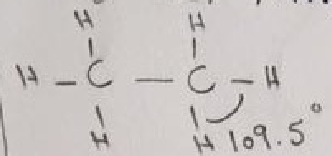
Organic and Biological Chemistry :-

* Hydrocarbons ((H + C))

→ Alkane (C_nH_{2n+2})
→ Alkene (C_nH_{2n})
→ Alkyne (C_nH_{2n-2})
→ Aromatic (C_nH_n)

((nonPolar Covelent))

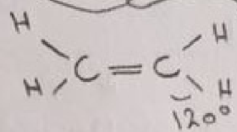
→ **Alkane**



- Single bond
- Saturated with H → مشبع
- tetra Hydral shape → هرم رباعي
- C_nH_{2n+2}
- SP^3 → hyPrazation

« يحصل له دوران »
rotation.

→ **Alkene**



- Double bond
- unsaturated → غير مشبع
- C_nH_{2n}
- SP^2

→ **Alkyne**

- Trouble bond
- unsaturated
- C_nH_{2n-2}
- SP

→ Ar. H
((Aromatic))

* (IUPAC)

Prefix + base + suffix.

* isomers :- have the same molecular furmola but different stractural formal ((atoms are bonded in different order.

* ways to show bonding connection :-

→ normal butane → C_4H_{10} ($CH_3CH_2CH_2CH_3$)

Ar. H — sp
«Aromatic»

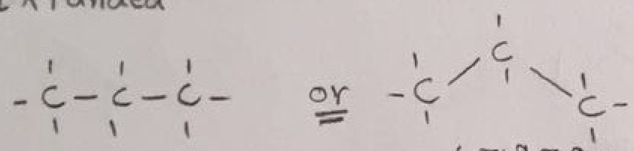
* (IUPAC)

Prefix + base + suffix.

* isomers:- have the same molecular furmola but different stractural formal «atoms are bonded in different order.

* ways to show bonding connection:-

[1] Expanded

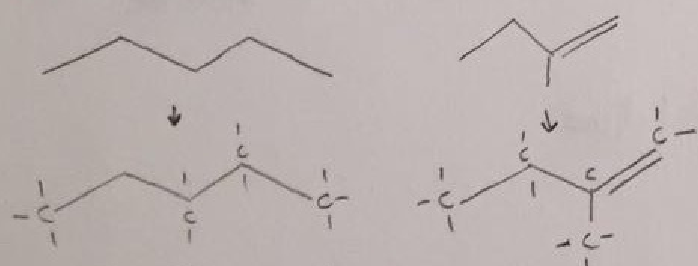


«zigzag»

[2] Condensed

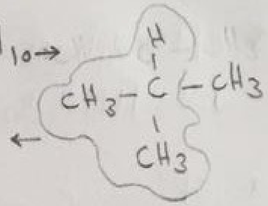
$CH_3CH_2CH_2CH_3$ or $CH_3(CH_2)_2CH_3$

[3] stick



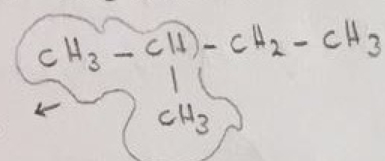
$CH_3CH_2CH_2CH_3$ $\quad$ $CH_3-CH_2-C(CH_3)=CH_2$

- butane
  - normal butane → C_4H_{10} $(CH_3CH_2CH_2CH_3)$
  - iso butane → C_4H_{10} →



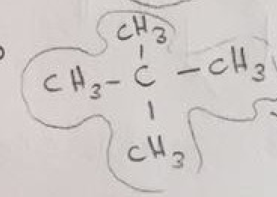
iso اسمه

- Pentane
  - normal $CH_3CH_2CH_2CH_2CH_3$
  - iso



iso

  - neo



neo

* البيوتان : 2 isomers
* البنتان : 3 isomers

⟸ بدلاً من اختراع common names للـ isomers
فرو التسمية بتبعنا الـ (IUPAC)

* IUPAC Name :-

→ Prefix + base + Suffix

1. Choose the largest continue chain
base وأعد كم كربون واكتب الـ

2. اكتب الـ suffix
ane ene yne

3. Prefix :- Alkyle group or Hallogens

$$\overset{1}{CH_3}-\underset{\underset{CH_3}{|}}{\overset{2}{CH}}-\overset{3}{CH_2}-\underset{\underset{CH_2CH_3}{|}}{\overset{4}{CH}}-\overset{5}{CH_2}\overset{6}{CH_2}\overset{7}{CH_3}$$

methyle ← CH_3 CH_2CH_3 → Ethyle

طريقة التسمية :-

1. 7 carbons → heptane
2. alkyle group اسماء الـ
3. **4 - ethyl - 2 - methyl heptane**

$$\underset{1}{CH_3}-\overset{\overset{CH_3}{|}}{\underset{\underset{\underset{2}{CH_3}}{|}}{C}}-\underset{3}{CH_2}-\underset{\underset{\underset{4}{CH_2CH_3}}{|}}{CH}-\underset{5}{CH_2}-\underset{6}{CH_2}-\underset{7}{CH_3}$$

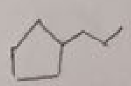
4 - ethyl - 2,2 - dimethyl heptane

* بين الارقام → « , »
بين الارقام والحروف → « - »

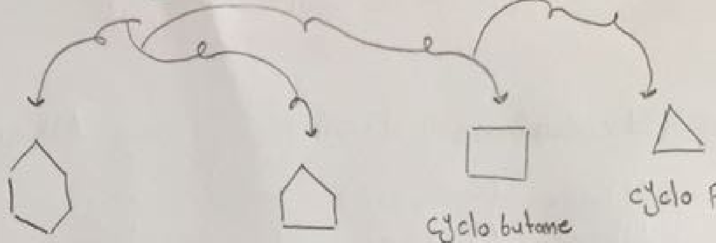
* sp^3 → σ → سجما
* sp^2 → σ π → سجما + باي
* sp → σ 2π → سجما + 2 باي

- 
1- ethyl - 3 - methyle cyclo hexane

- 
1,3 - methyl cyclo hexan

- 
Propyle cyclo Pentane

* Cycloalkanes ((C_nH_{2n})) nonPolar solvent



cyclo hexane   cyclo Pentane → The most stable member rings ..

cyclo butane   cyclo Probane → ring strain
متلاوية او متوترة اكثر

* كل ما انزل « unstable »

- 
cyclo Probyle butane

- 
2 - cyclo Propyle hexane

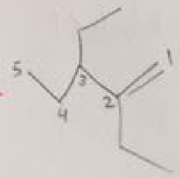
* Alkyne

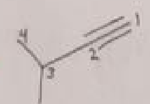
- 
6,6 - dimethyle - 3 - heptane

* Aromatic

benzen → 
سريعة دمين « unreative »

- 
3 - methyl - 1 - hexane.

- 
2,3 - diethyl - 1 - Pentene.

- 
$CH_3-\underset{\underset{CH_3}{|}}{CH}-C\equiv CH$
3 - methyl - 1 - butyne

- 
2 - hexene

* IUPAC Name :-

→ Prefix + base + suffix

[1] Choose the largest continue chain

وأعـد كم كربون واكتب الـ base

[2] اكتب الـ suffix

ane ← suffix → ene → yne

[3] Prefix :- Alkyle group or Hallogens

$$\overset{1}{CH_3}-\overset{2}{CH}(CH_3)-\overset{3}{CH_2}-\overset{4}{CH}(CH_2CH_3)-\overset{5}{CH_2}\overset{6}{CH_2}\overset{7}{CH_3}$$

methyle ← $\underline{\underline{CH_3}}$ $\underline{\underline{CH_2CH_3}}$ → Ethyle

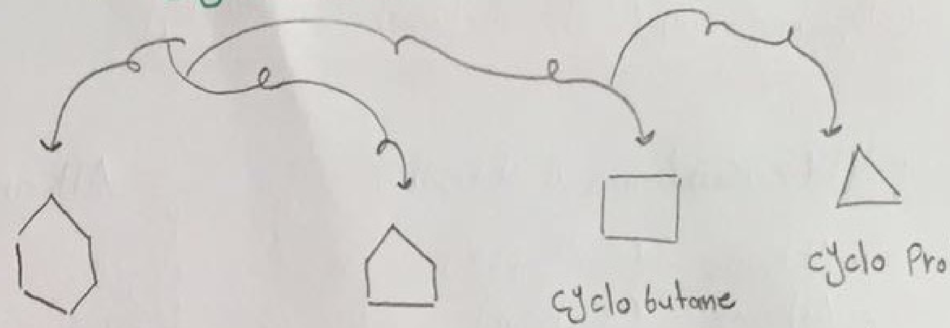
مؤيمة المتسبه :-

1. 7 carbons → heptane
2. اسماء الـ alkyle group
3. $\boxed{\text{4 - ethyl - 2 - methyl heptane}}$

$$\underset{1}{CH_3}-C(CH_3)(\underset{2}{CH_3})-\underset{3}{CH_2}-CH(\underset{4}{CH_2CH_3})-\underset{5}{CH_2}-\underset{6}{CH_2}-\underset{7}{CH_3}$$

$\boxed{\text{4 - ethyl - 2,2 - di methyl heptane}}$

* Cyclo alkanes «C_nH_{2n}» nonPolar solvent



cyclo hexane cyclo Pentane → The most stable member rings ..

cyclo butane cyclo Probane → ring strain

مقاومه او منفذ اكثر

* كل ما اقل «unstable»

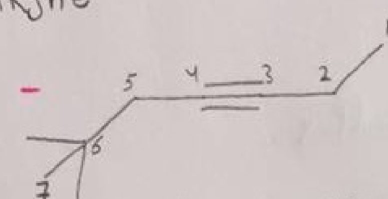
- 

cyclo Probyle butane

- 

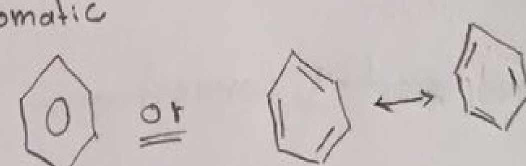
2 - cyclo Propyle hexane

* Alkyne

- 

6,6 - dimethyle - 3 - heptane

* Aromatic

benzen → 

مستقره «unreative»

= 7 carbons → heptane

2 alkyle group اسماء ال

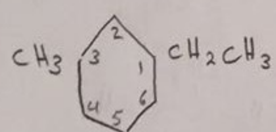
3 4 - ethyl - 2 - methyl heptane

$$\begin{array}{ccccccccccccc} & & CH_3 & & & & & & & & & & \\ & & | & & & & & & & & & & \\ * \; CH_3 & - & C & - & CH_2 & - & CH & - & CH_2 & - & CH_2 & - & CH_3 \\ 1 & & | & & 3 & & | & & 5 & & 6 & & 7 \\ & & CH_3 & & & & CH_2CH_3 & & & & & & \\ & & 2 & & & & 4 & & & & & & \end{array}$$

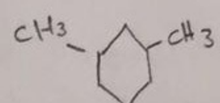
4 - ethyl - 2,2 - dimethyl heptane

* بين الارقام → ‹‹ , ››
بين الارقام والحروف → ‹‹ - ››

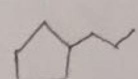
* $sp^3 \rightarrow \sigma \rightarrow$ سجما
* $sp^2 \rightarrow \sigma \, \pi \rightarrow$ سجما + باي
* $sp \rightarrow \sigma \, 2\pi \rightarrow$ سجما + ٢ باي

- 

1- ethyl - 3 - methyle cyclo hexane

- 

1,3-methyl cyclo hexan

- 

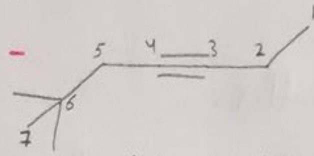
Propyle cyclo Pentane

3 4

cyclo Probyle butane

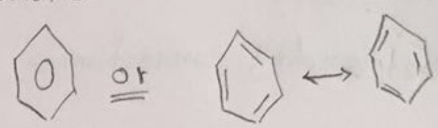
2 - cyclo Propyle hexan

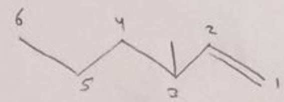
* Alkyne

- 

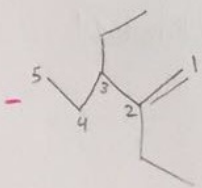
6,6 - dimethyle - 3 - heptane

* Aromatic

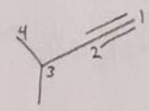
benzen →  بسوله دبين ‹‹ unreative ››

- 

3 - methyl - 1 - hexane.

- 

2,3 - diethyl-1- Pentene.

- 

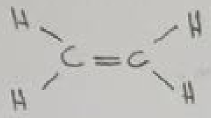
$$\begin{array}{l} CH_3 - CH - C \equiv CH \\ \qquad\;\; | \\ \qquad CH_3 \end{array}$$

3 - methyl - 1 - butyne

- 

2- hexene

* Alkene

⇒ simplest one  C_2H_4

common name :- ethylene

IUPAC :- ethene

⇒ ترقيم الرابطة في المركبات الا فيما أكثر من 4 كربونات

- -C-C=C → Propene
- -C-C-C=C → 1- Butene
- C-C=C-C → 2- Butene

C_4H_8 isomers

 1- Butene

 2- Butene

* cis and trans → موجودة فقط في ال alkene

* Alkyne

H-C≡C-H C_2H_2

common name :- acetylene

IUPAC :- ethyne

- -C-C≡C-H → Propyne
- C-C-C≡C- → 1- Butyne
- C-C≡C-C- → 2- Butyne

* Reaction of Alkane :-

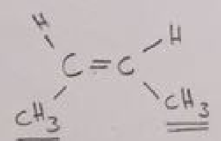
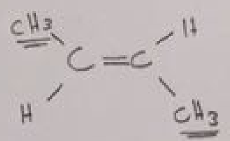
- unreactive
- Combustion
- They make great nonpolar solvent

C-C , C-H sgma

زوايا قوية

Ex :-

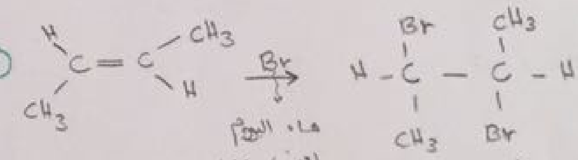
C_4H_8

* $CH_3-CH_2-CH=CH_2$ 1-butene
* $CH_3-CH=CH-CH_3$ 2-butene
*  cis-2-butene
*  trans-2-butene

أكثر stable

* Reaction of Alkene

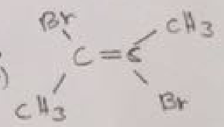
- Addition Reaction :-

① 

ماء البروم يعني (H_2)

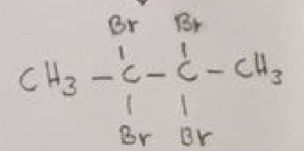
2,3 - dibromo butane

- unsaturated معروفة دايماً للتفاعلات الا ضافة

② $CH_3-C\equiv C-CH_3$ $\xrightarrow[(H_2O)]{Br}$ 

2-butyne

ممكن تختص 2 mol of Br_2

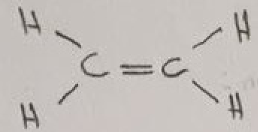


2,2,3,3 - tetrabromo butane

- الهالوجينات تكسر ال π bond ..

* Alkene

⇒ simplest one



C_2H_4

common name :- ethylene

IUPAC :- ethene

⇒ تقيم الرابطة في المركبات الا فيما أكثر من 4 كربونات

- -C-C=C → Propene
- -C-C-C=C → 1- Butene
- C-C=C-C → 2- Butene

1- Butene, 2- Butene → C_4H_8 isomers

1- Butene

2- Butene

* cis and trans → موجودة فقط في ال alkene

Ex :- C_4H_8

* Alkyne

H - C≡C - H C_2H_2

common name :- acetylane

IUPAC :- ethyne

- -C-C≡C-H → Propyne
- -C-C-C≡C- → 1- Butyne
- -C-C≡C-C- → 2- Butyne

* Reaction of Alkane :-

- unreactive
- Combustion
- They make great nonpolar solvent

C-C , C-H sgma
روابط قوية

* Reaction of Alkene

- Addition Reaction :-

Electronegativity

Nonmetal Oxidation

Bond Energy

ملاحظة: اتجاه الأسهم تعني الزيادة increasing.

Electronegativity

Nonmetal Oxidation

Bond Energy

Bond Length

1	2	3	4	5	6	7	8	9	10	11	12	13	14	15	16	17	18
1 **H** 1.008																	2 **He** 4.0026
3 **Li** 6.94	4 **Be** 9.0122											5 **B** 10.81	6 **C** 12.011	7 **N** 14.007	8 **O** 15.999	9 **F** 18.998	10 **Ne** 20.180
11 **Na** 22.990	12 **Mg** 24.305											13 **Al** 26.982	14 **Si** 28.085	15 **P** 30.974	16 **S** 32.06	17 **Cl** 35.45	18 **Ar** 39.948
19 **K** 39.098	20 **Ca** 40.078	21 **Sc** 44.956	22 **Ti** 47.867	23 **V** 50.942	24 **Cr** 51.996	25 **Mn** 54.938	26 **Fe** 55.845	27 **Co** 58.933	28 **Ni** 58.693	29 **Cu** 63.546	30 **Zn** 65.38	31 **Ga** 69.723	32 **Ge** 72.630	33 **As** 74.922	34 **Se** 78.97	35 **Br** 79.904	36 **Kr** 83.798
37 **Rb** 85.468	38 **Sr** 87.62	39 **Y** 88.906	40 **Zr** 91.224	41 **Nb** 92.906	42 **Mo** 95.95	43 **Tc** (98)	44 **Ru** 101.07	45 **Rh** 102.91	46 **Pd** 106.42	47 **Ag** 107.87	48 **Cd** 112.41	49 **In** 114.82	50 **Sn** 118.71	51 **Sb** 121.76	52 **Te** 127.60	53 **I** 126.90	54 **Xe** 131.29
55 **Cs** 132.91	56 **Ba** 137.33	57-71 *	72 **Hf** 178.49	73 **Ta** 180.95	74 **W** 183.84	75 **Re** 186.21	76 **Os** 190.23	77 **Ir** 192.22	78 **Pt** 195.08	79 **Au** 196.97	80 **Hg** 200.59	81 **Tl** 204.38	82 **Pb** 207.2	83 **Bi** 208.98	84 **Po** (209)	85 **At** (210)	86 **Rn** (222)
87 **Fr** (223)	88 **Ra** (226)	89-103 #	104 **Rf** (265)	105 **Db** (268)	106 **Sg** (271)	107 **Bh** (270)	108 **Hs** (277)	109 **Mt** (276)	110 **Ds** (281)	111 **Rg** (280)	112 **Cn** (285)	113 **Nh** (286)	114 **Fl** (289)	115 **Mc** (289)	116 **Lv** (293)	117 **Ts** (294)	118 **Og** (294)

* Lanthanide series	57 **La** 138.91	58 **Ce** 140.12	59 **Pr** 140.91	60 **Nd** 144.24	61 **Pm** (145)	62 **Sm** 150.36	63 **Eu** 151.96	64 **Gd** 157.25	65 **Tb** 158.93	66 **Dy** 162.50	67 **Ho** 164.93	68 **Er** 167.26	69 **Tm** 168.93	70 **Yb** 173.05	71 **Lu** 174.97
# Actinide series	89 **Ac** (227)	90 **Th** 232.04	91 **Pa** 231.04	92 **U** 238.03	93 **Np** (237)	94 **Pu** (244)	95 **Am** (243)	96 **Cm** (247)	97 **Bk** (247)	98 **Cf** (251)	99 **Es** (252)	100 **Fm** (257)	101 **Md** (258)	102 **No** (259)	103 **Lr** (262)

Bond Length

Alkane	Alkene	Alkyne
Single bond with only one sigma (σ).	**Double** bond with only one sigma (σ) and one pi (π).	**Triple** bond with only one sigma (σ) and two pi (π).
CnH2n+2 Ex: C_6H_{14}	CnH2n Ex: C_6H_{12}	CnH2n-2 Ex: C_6H_{10}
Sp^3 - hybrid They have a Tetrahedral geometry.	Sp^2	Sp
Saturated.	Unsaturated.	Unsaturated.
The bond **angle** for a single bond is 109.5°.	The bond **angles** between the double bond are 120°.	The bond **angles** between the triple bond are 180°.
Free Rotation.	• No free rotation. • Nomenclature of Alkenes :(تسميات الألكينات) - Cis-Alkenes: have the carbons in the chain on the **same** side of the molecule. - Trans-Alkenes: have the carbons in the chain on the **opposite** side of the molecule.	It doesn't matter.

سارا الزليباني