Chapter one

Structure of the atom

Atomic structure

1- Greek philosophers:

Any piece of matter can be divided into smaller parts and each part can be subdivided into smaller parts which can't be divided this part is called Atom.

2- Aristotle

a-He refused the idea of Greek philosophers about the atom .

b- He supposed that all matters composed of 4 constituents which are (water, air , dusta and fire and postulated that the cheap metals such as iron or copper can be changed into precious ones like gold by changing the percentage of four constituents.

3- Boyl (lrish 1661)

i-He refused the Aristotle concept.
ii-He was the 1^{st} scientist to define the element as pure simple substance which can
not be analvsed into smple one by normal chemical methods.

4) Dulton s atom (Englsh 1803)

He supposed thet :

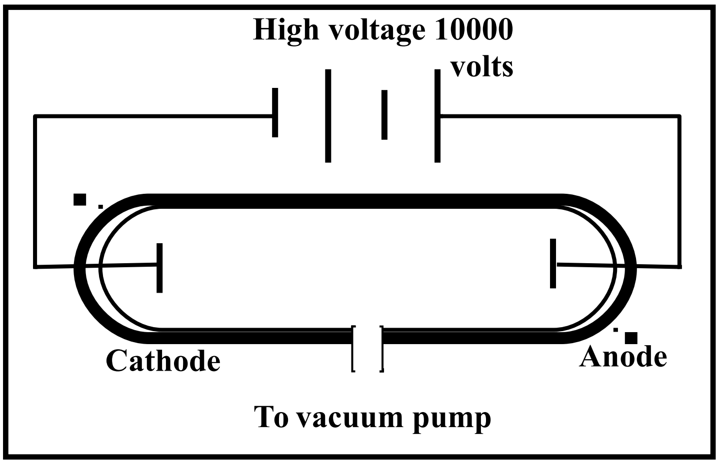
1- **The substance consists of very amall particles called atoms.**
2- **Every element consists of very small dense atoms which can t be divided.**
3- **Atoms f the same element are identical.**
4- **Atome of different are different.**

N.B **By the 19 th century**

Scientists had accepted the idea that elements consists of atoms but they knew very little about those atoms.

Cathode- ravs experiment

(discovery of the electron)



a- All gases under normal conditions of pressure and temp (76 cm. hg. 25c) don t conduct electricity .
b- If a glass tube evacuated from the gas to decrease its pressure to reach 0.1.1 ------------ 0.001 m.mhg

The gas will conduct electric current.

c- If the potential difference between the tow poles increases up 10.000 Volts a Flow of invisble rays are emitted from the cathode causing glowing to the wall of tube behind the anode and called cathode ray .

Properties of cuthode rays

1) Consists of tiny particles have mass and velocity.
2) Transfers in straigh lines glowing the glass facing the cathode.
3) Have negatve charge.
4) Have a thermal effect.
5) Affected by electric and Magnetic field.
6) Cathode rays don t change by changing either cathode material or type of the gas which proves that cathode rays take part in the structure of all substances.

5 – Thomson s atom 1897

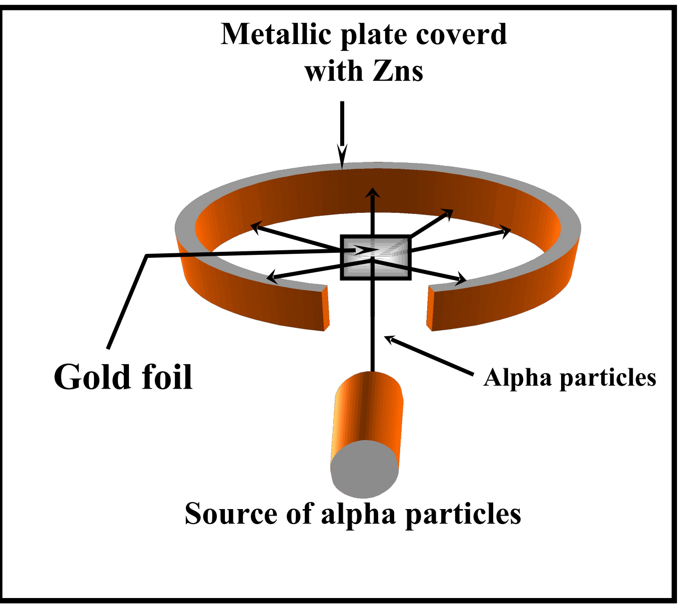
He conclude from the last experiment that

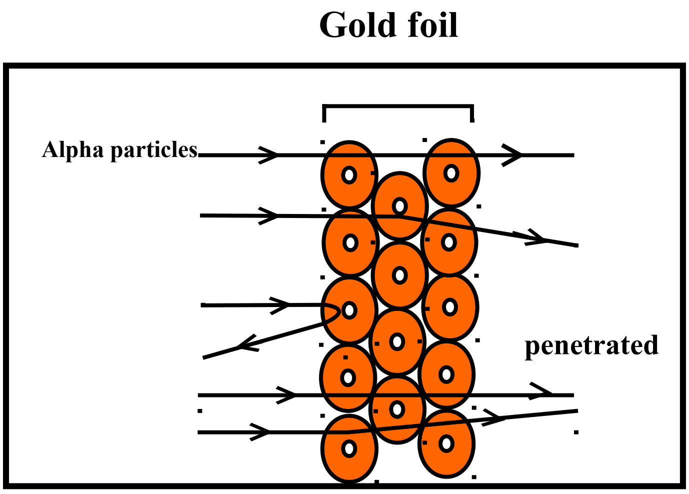
i- The atom is a homogenous sphere of positive electricity.

ii- Inside it there are negative electrons enough to make it electrically neutral.

Ruther ford s experiment

In 1911 Geiger and Marsden performed a famous exp According to suggestion of Ruther ford by the followng apparatus.





1 – He allowed alpha particles to hit a mtallic plate lined with Zinc suphide (glows when hits with alpha rays)

2 – On placing a gold foil in the front of alpha rays he concluded the following from the following observation.

Observation	Result
1- Most of alpha particles appeared in the same position before putting gold foil.	1- Most of the atom is a space not solid as explained by Dalton and thomson.
2- A very small percentage of alpha particles reflected back to appears as flashes on other side of sheet.	2- The atom has very small part with very small volume but high density.
3- Some flashes appeared on the sides of 1st site.	3- The dense part of the atom which concentrate in it most mass have same charge of alpha particle (+ve) which called nucleus of the atom.

Structure of the atom

Atomic No:

No of protons inside the nucleus or no of electrons inside the E. levels.

Mass No:

Sum of protons and neutrons inside the nucleus.

Isotopes:

Are different forms of atoms of the same element which have same atomic nos but different mass no.

Ex: hydrogen has 3 isotopes: ……….

Rutherford:

Rutherford was the 1st scientist who stated the concept of the atomic structure.

1- The Atom:

Although it has very small size but it has a complicated structure that resembles the solar system in which electrons revolve around the central nucleus in orbits as planets revolve around the sun.

2- The Nucleus:

Is much smaller than the atom. Located in the centre of the atom with (+ve) charge. There is a big space between the nucleus and orbits of electrons, so most of the atom is a space. Most mass of the atom is concentrated in the nucleus as mass of e is very small and can be neglected.

3- Electrons:

1- Have negligible mass compared to that of the nucleus.
2- No of electrons (-ve) equals no of ptotons (+ve) so the atom is electrically neutral.

3- Electrons revolve around the nucleus in a fixed orbit as electrons are affected by two forces equal in strength but in opposite direction, which are :
   a- Force of attraction of the nucleus to electrons.
   b- Centrifugal force due to velocity of electron around the nucleus.

Give reason: Electrons are not attracted to the nucleus.

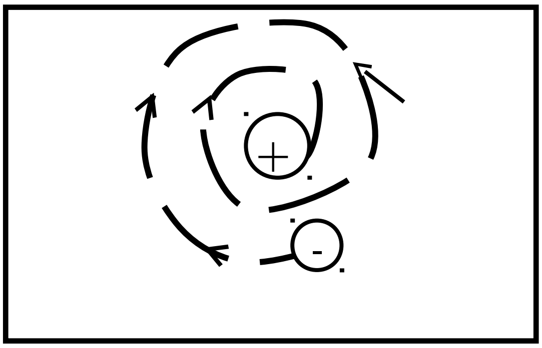
Explain: Structure of the atom in the view of rutherford.

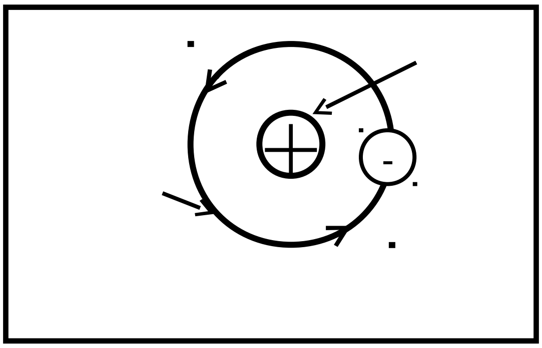
Objections on Rutherford's atomic model (Maxwell's theory):

Rutherford's concept was contradicted by maxwell's theory (Which was based on laws of newton mechanics and concerned with the movement of relatively large bodies).

- Which states that: " When an electrecally charged body moves in orbit, it will lose its energy gradually by emission of radiation causing gradual decrease in orbit radius".

By applying this theory on electron movement in Rutherford's atom, we would expect that electrons are in a state of continuous emission of radiation, so the atomic radius will decrease and electrons move in a spiral orbit until they hit the nucleus.





Give reason: Contradiction between the classical mechanical laws and rutherford.

Bohr's Atomic Model

Bohr's postulates:

1- A positively charged nucleus exists in the center of the atom.
2- Atom is electrically neutral as no of p^+ s equals to no of e's.
3- Electrons revolve around the nucleus in orbits due to centrifugal and attraction forces.

4- Electrons orbit the nucleus in a rapid movement without gaining or losing energy.
5- Electrons orbit the nucleus only in a definite allowed energy levels, so they can't be found at intermediate distance.
6- Each electron in the atom has a definite amount of energy depending on the distance between is E.L and nucleus. This energy increases as its radius increases.
7- It was found that the maximum no of energy levels in the heaviest known atoms in their ground state (unexcited) is only seven (K, L, M, N, O, P, Q). Each level has energy expressed by a whole no called principle Q. No.

Ex: The 1^{st} E. Level K its principle Q. no = 1
The 2^{nd} E. level L its principle Q. no = 2

8- If when atom is excited by heating (Quantum) or by electric discharge the electron will transfer to a higher E. level agrees with the absorbed quantum. The excited electron in the higher E. level is then unstable, so it returns to its original level losing the same quantum of energy, which it gained during excitation in the form of radiation have definite wavelength and frequency.

❖ Remarks:

1- The quantum: Is the amount of energy gained or lost when an electron jumps from one E. level to another.
2- The difference in energy between levels (Q) is not equal i.e. the difference in this energy decreases further from the nucleus. This means that the quantum of energy required to transfer an electron from one energy level to another is not equal.
3- The electron does not move from its level to another unless the energy absorbed or emitted is equal to the difference in energy between 2 levels i.e. one quantum.
(There is no half quantum for instance). Q can't be divided or doubled

Give reason: It is wrong to say that e' to be transferred from E.L (K) to E.L (M) needs amount of energy equals 2 quantum.

Excited Atom:

- It is an atom that acquired an amount of energy (Q) sufficient to transfer its e's from their original E.L to higher ones.

Advantages of Bohr:

Bohr's atomic theory succeeded in the following ways:

1- It explained hydrogen atom spectrum.
2- He introduced the idea of quantum no to detect energy of electrons in energy levels.
3- He proved that electrons during that electrons during rotation around the nucleus in ground state do not radiate energy, so they will not fall back to the nucleus .

(a reconcilion between Rutherford and Maxwell).

Disadvantages of Bohr's theory:

1 – If failed to explain the spectrum of any other element even that of He except hydrogen (Simplest Electronic System).

2 – He considered the electron as a (-ve) particle only and did not consider that it also has wave properties.

3 – He postulated that it is possible to determine precisely both speed and location of an electron at the same time. This is experimentally impossible because the apparatus used will change either speed or location of electron so the result won't be accurate.

4 – He described the electron when moving in a circular planer orbit, which means that hydrogen atom is planer. Later it was confirmed that hydrogen atom has 3 dimensional co-ordinates.

Bohr's Theory

The atomic spectrum

Studying and explaining the atomic spectrum was the key to he atomic structure in 1913 and deserved noble prize at 1922.

Atomic emission spectrum:

1 – By heating gases or vapours of substances to a high temperature (by heat or electricity) under low pressure it produces light.

2 – By using spectroscope we find that this light consists of a fixed number of coloured lines called line spectrum.

Line spectrum of any element:

By exp. Proved that spectrum line differs from one element to another like finger prints.

N.B:

Line spectrum of sun rays shows that composed of hydrogen and Helium.

The principles of Modern Atomic Theory:

1 – Dual nature of electron.

2 – The Heisenberg uncertainty principle.

3 – Finding the mathematical expression which describes the wave motion of electron, its shape and its energy.

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1 – The dual nature of the electron

The experimental data showed that the electron has a dual nature i.e

a) It is a material particle. b) It has wave properties.

* De Broglie principle:

Every moving body (such as electron or the nucleus of an atom or whole molecule)is associated with (accompanied by) a wave motion (or matter waves) which has some properties of light waves.

The matter wave motion differs from electromagnetic waves n …………

Matter Waves	Electromagnetic Wave
1- They are not separating from the moving body	1- They are separated from the moving body
2- Their speed is not equal to the speed of light	Their speed is equal to the speed of light

2- The Heisenberg uncertainty principle: (quantum mechanics)

It is practically impossible to determine both position and the velocity of the electron exactly (precisely) at the same time. We can only say that it is probably to a greater r lesser extent to locate the electron in this or in that place. This is to speak in terms of probability seems to be more precise.

3. The wave equation for motion of electron inside the atom:

(Schrodinger wave mechanics theory): He applied the ideas of Blanck, Einestein, De Broglie and Heisenberg so he shows that:

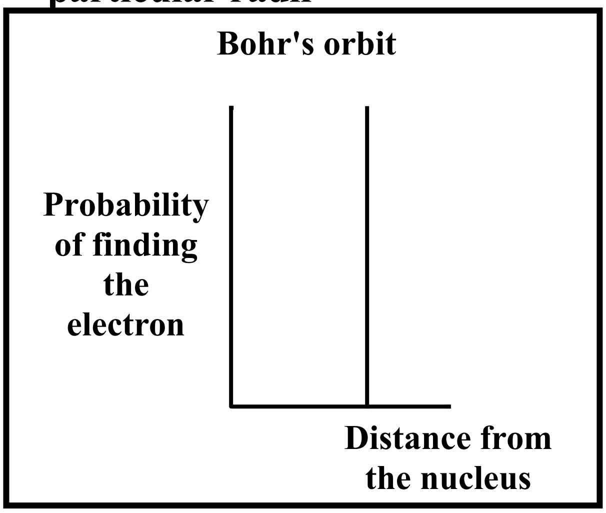
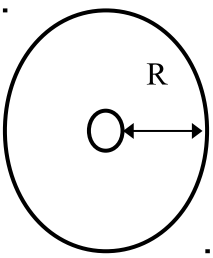
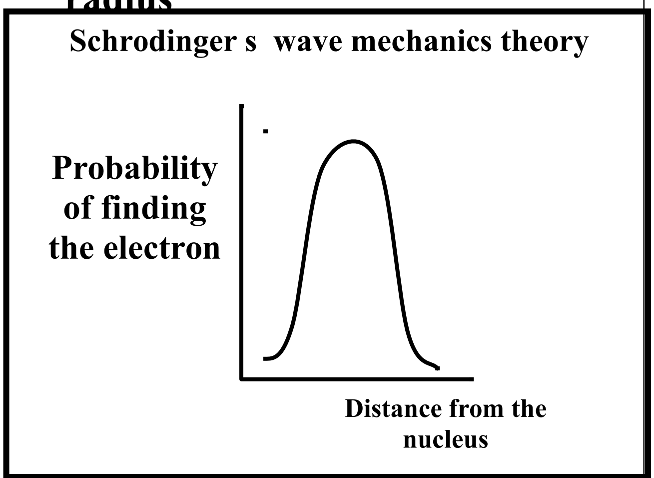
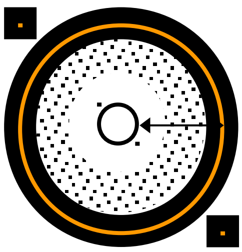
❖ It is possible to determine the allowed energy levels of the electron and define the region of space around the nucleus where it is most probable to find the electron in each energy level.

❖ The electronic motion around nucleus has a wave properties therefore the position to use the term electron cloud to describe any orbital.

Electron Cloud: (used to describe any orbital)

"Area of space around the nucleus where there is a great probability for finding electrons in all direction and all positions."

❖ The difference between the orbit and orbital concepts according to both Bohr and the wave mechanics theories:

Bohr s theory	Schrodinger s wave mechanics theory
• It is a circular planer orbit with particular radii   	• It is an electron cloud used to describe any orbital • Electron cloud the regions of high density of dots represent the region of high probability of finding the (e) from which it is possible to define the atomic radius  

❖ **The mathematical solution of the Schrodinger equation introduced four numbers which are called quantum numbers.**

- Quantum Numbers:

These nos define the volume of space (orbital) where there is maximum probability of finding electrons. Besides, they define the energy, shape and direction of orbitals.

1- **Principle Q.no (n).**
2- **Subsidiary or orbital (azimuthal) Q. no (L).**
3- **Magnetic Q. no (m).**
4- **Spin Q. no (ms).**

1- Principle Q No (n):

Bohr had used this no to define the following:

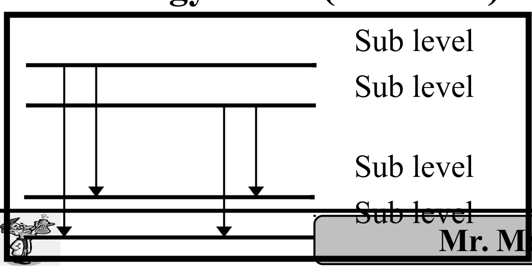
1- **Order of principle E. levels their number in the heaviest known atom in the ground state is seven.**
2- **No of electrons required to fill a given E. level = two times the square of the level no (2n2).**

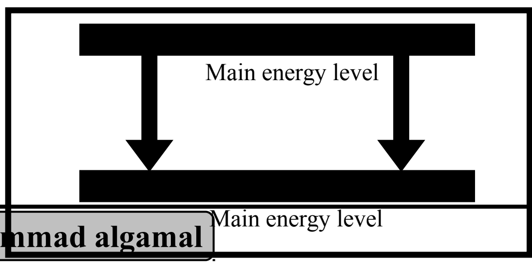
-1st E.L	K	Is filled with	2 electrons
-2nd E.L	L	Is filled with	8 electrons
-3rd E.L	M	Is filled with	18 electrons
-4th E.L	N	Is filled with	32 electrons

- **But this rule does not apply to the last three levels (O, P, Q). However, the atom becomes unstable if no of electrons exceeds 32 electrons on any level.**

2. Subsidiary Q No (L):

1- **Used to detect the no of sub levels in each E. level.**
2- **The energy sub levels take the symbols s, p, d, f. this is shown by the scientist Somerfield. When he used a spectroscope which has a high resolving power, he found that the single line (which represents electron transition between two different energy levels) is indeed a number of fine spectral lines which represents electron transition between very near energy levels (sublevels).**





3- **No of sublevels in each energy level = order of principle energy level (n).**

-1st E.L	K	has 1 sub level	1s
-2nd E.L	L	has 2 sub level	2s, 2p
-3rd E.L	M	has 3 sub level	3s, 3p, 3d
-4th E.L	N	has 4 sub level	4s, 4p, 4d, 4f

N.B:

- **Energy of sub levels of same E. level is not equal.**

f>d>p>S

- **Energy of same sublevels but of different E. levels also differ in energy.**

Ex: **4d>3d 4p>3p>2p**

- **There is a small difference in the energy between sub-levels.**

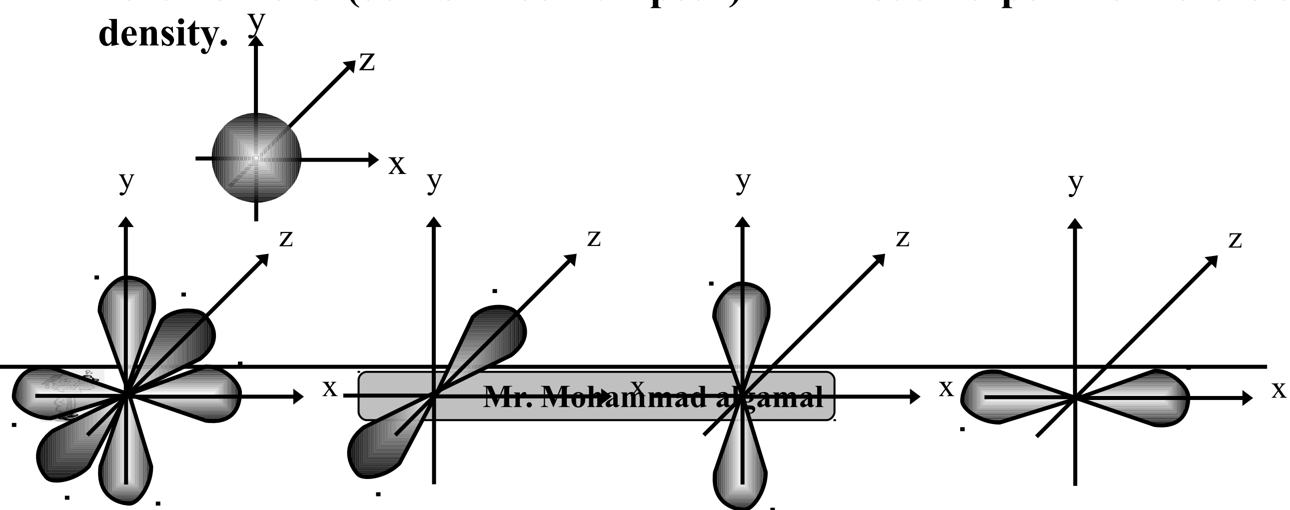
3. Magnetic Q No (m):

Detected by Zieman when he exposed spectral line to strong magnetic field, he found that each line divides into many lines, so he concluded that each E. sublevel has no orbitals.

- ***Magnetic Q No is characterized by:***

1- **Used to detect no of orbitals in each E. sub level and their direction in space.**
2- **Sublevel (S) has one orbital of spherical symmetrical shape.**
3- **Sublevel(P) has 3 orbitals.**

- **Each orbital (P_x,P_y,P_z) is perpendicular to the other two.**
- **Also the electron cloud of each orbital takes the form of 2 pears meeting head to head (dumb – bell shaped) at a node i.e point of zero electron density.**



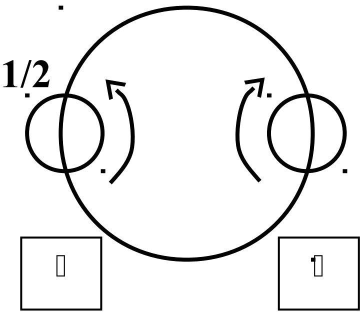
4- Sublevel (d) has 5 orbitals.
5- Sublevel (f) has 7 orbitals.
6- Orbitals of the same sub level are equal in energy and shape.

Ex: $p_x = p_y = p_z$

7- No of orbitals in each E. level = square order of E. level = n^2.

4-<u>Spin Q. No (ms):</u>

- **Detects the direction in which the electron spins around its axis during its rotation around the nucleus.**
- **Each orbital can be saturated by 2 electrons, one electron spins around its axis clockwise while the other electron spins anti – clockwise in order to from 2 opposite magnetic fields to decrease the force of repulsion between them which keep the atom stable.**
- **It has only two possible values + 1/2 – 1/2**



<u>Give reason</u>: Each orbital carries 2 electrons although they are negatively charged.

<u>Summary of the relationship between the principle E.L, sub levels orbitals and no of electrons:</u>

1- No of energy sublevels = order of principle level (n).
2- No of orbitals within a principle level = square the no of the level (n^2).
3- No of electrons occupying a given E. level = two times the square order of this level ($2n^2$).

Ex:

Order of E.L.	E.L	No of sub-levels (n)	No of orbitals (n2)	No of electrons (2n2)
1st	K	1 (1s)	1	2
2nd	L	2 (2s, 2p)	4	8
3rd	M	3 (3s, 3p, 3d)	9	18
4th	N	4 (4s, 4p, 4d, 4f)	16	32

❖ **Principle of distributing electrons:**

- **Aufbau (building – up) principle:**

"Electrons occupy energy sublevels in an ascending order according to increasing energy where the lowest energy sublevel is filled 1st".

Ex: 4s is filled before 3d as energy of 4s < 3d.

K	1S			
L	2S			2p
M	3S			3p
N	4S		3d	4p
O	5S		4d	5p
P	6S	4F	5d	6p
Q	7S	5F	6d	

Examples:

Na 11	$1s^2$	$2s^2$	$2p^6$	$3s^1$				
Ca 20	$1s^2$	$2s^2$	$2p^6$	$3s^2$	$3p^6$	$4s^2$		
Zn 30	$1s^2$	$2s^2$	$2p^6$	$3s^2$	$3p^6$	$4s^2$	$3d^{10}$	

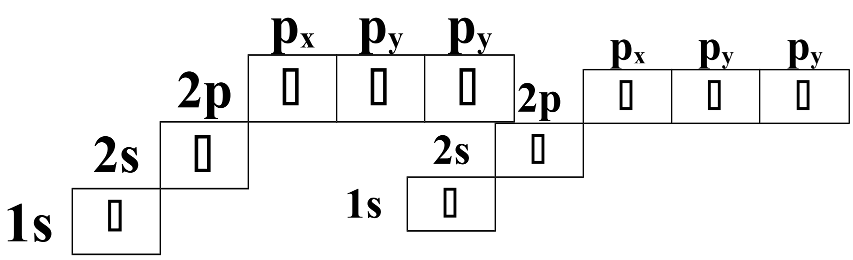
- Hund's rule:

State that: "No electron pairing takes place in a given sublevel until each orbital contains one electron."

Ex:

$_7N$	$1s^2 – 2s^2 – 2p^3$

$_8O$	$1s^2 – 2s^2 – 2p^4$



- **Atom is stable when the outer sub-level is half completely filled with e's.**

❖ Remarks:-

1- **Electrons are preferred to be unpaired before pairing because according to Hund's rule on pairing electrons in the same orbitals, they will repel decreasing stability of the atom.**
2- **Electrons prefer to be paired with another electron than to transfer to a higher sub-level, as the energy needed to transfer it to a higher sub-level.**
3- **Also the spin of single electrons must be in the same direction because this gives the atom more stability.**

Another E. configurations (to the nearest noble gas)

Ex:

Li 3 : (He_2) ,$2S_2$

Na 11 : (Ne_{10}) , $3S_1$

K19 : (Ar_{18}), $4S_1$

Questions

1. **Write short notes on:**
   1- **Hund's rule.**
   2- **Aufbau principle.**
   3- **Heisenberg uncertainty principle.**
   4- **Electron cloud.**
   5- **Defects of Bohr's theory.**
   6- **Excited atom.**

2. **Write scientific term:**
   1- **Area of space around nucleus which has probability for finding electron.**

2- Sublevel can be saturated with 2 electrons.
3- Sublevel can be saturated with 6 electrons.
4- Sublevel can be saturated with 10 electrons.
5- Sublevel can be saturated with 14 electrons.
6- Electrons occupy sublevels according to increasing energy.
7- Q. number which detects no of sub levels.
8- Q. number which detects no of orbitals and their direction in space.
9- Q. number used to detect direction of electron around its axis.
10-Amount of energy gained or lost to jump electron from an energy level to another.

3. Complete:
1- Max. No of energy levels in the heaviest known atom is ……………
2- An element is atomic no 16. The electrons can be distributed in ………….. E. levels …………. Sub levels and ………….orbitals.
3- Orbitals of the same sub level are …………. In energy.
4- Energy of electron depends on …………….and increases as ………………

4. Compare between:
1- Orbit and orbital.
2- Principle, Q. no and subsidiary Q. no.
3- Mag. Q. no and spin Q. no.

5. Give reason:
1- Its wrong to consider the electron as (-ve) particle only.
2- Any orbital can be saturated by 2 electrons although they are negatively charged.
3- The relation 2n2 can't be applied to O, P and Q levels.
4- Sub level 4s is filled with 10 electrons.
5- Sub level d is filled with 10 electrons.
6- Its impossible to detect both velocity and location of electron at the same time .
7- Energy level N is filled with 32 electrons .
8- Electrons of the same orbital spin in opposite direction .
9- Electrons of the same sublevel prefer to be unpaired before pairing.
10-Electrons prefer to be paired in a given sublevel than to transfer to a higher sub l- level .
11- Spectral line for any element characterizes it .
12-Explain rutherfor's atomic theory also write result of its experiment .

Vi – write electronic configuration for :-

Na_{11} , p_{15} , Cl_{17} , Fe_{26} , Cr_{24}

VII- Choose:

1- 1st scientist defines the element is ……………
a) Dalton b) Rutherford c) Boyl d) Thomson

2- Substance composed of 4 components which are (water, air, dust and fire) was the idea of ……….
a) Bohr b) Rutherford c) Dalton d) Aristole

3- To prove that cathode rays taking part in all substances ……………
  a) Have thermal effect.
  b) Transfer in straight line.
  c) Have tiny porticles.
  d) Don't changing substance of cathode or kind of gas.

4- On heating gases or vapours under low pressure at high temperature
  a) Absorbs light. b) Gives light.
  b) Gives gamma rays. D) Gives alpha rays.

- Show by experiment how can obtain cathode rays.
- Explain Thomson's atomic structure.

From Rutherford's experiment:
Explain the following results:
1- Most of the atom is a space not solid sphere.
2- The atom has part characterized by more dense and have very small volume (nucleus).
3- Charge of the dense part which have most mass of atom must have +ve charge.

Chapter two

Classification of elements and the long periodic table

The long form periodic table:
It depends on the building up principle (Auf - bau) The elements are arranged in an ascending order according to their atomic number (each element have one electron more than the element before it) .

The long form periodic table : the elements were arranged in an ascending order according to their atomic numbers

The periodic table is divided into four main blocks

1- S – block elements
2- P – block elements
3- d – block elements
4- f – block elements

1s	**2s < 2p**	**3s < 3p**	**4s < 3d < 4p**
5s < 4d < 5p	**6s < 4f < 5d < 6p**	**7s < 5f < 6d**	

$s^1s^2$1234567

sP P P P P P 1

d d d d d d d d d d 1 2 3 4 5 6 7 8 9 10

kcolB P

kcolb S

kcolb d

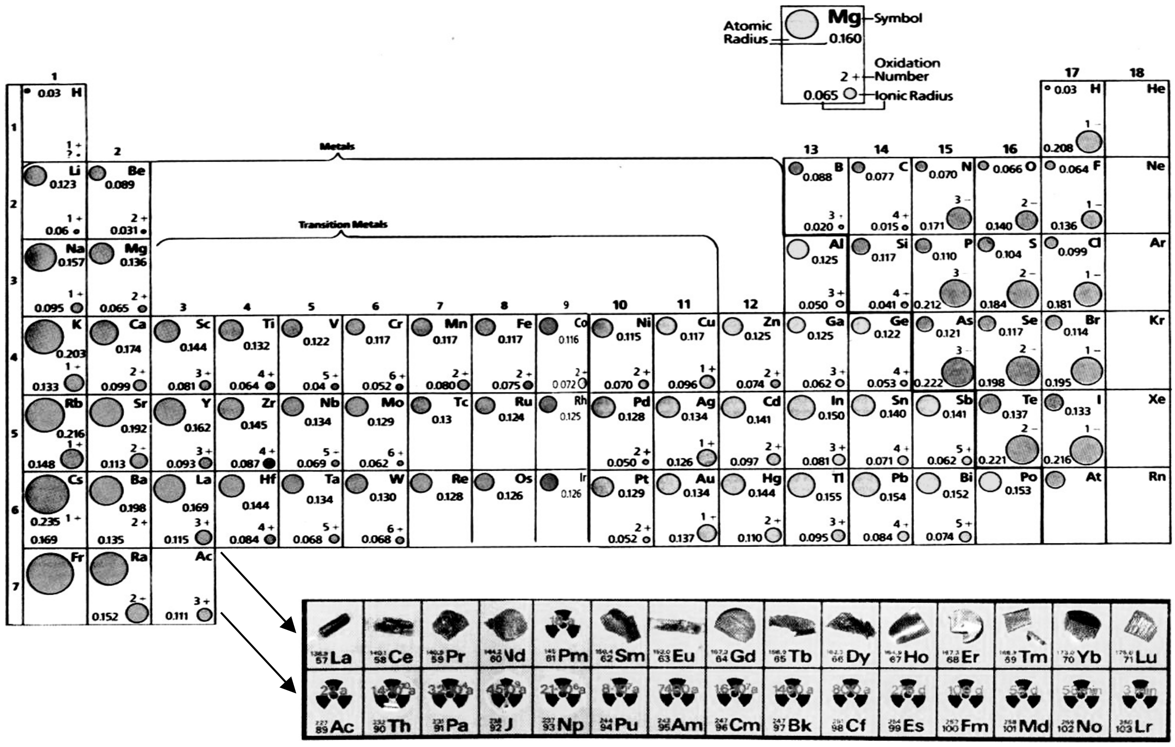
$f^1f^2f^3f^4f^5f^6f^7f^8f^9f^{10}f^{11}f^{12}f^{13}f^{14}$67

kcolb f

Sequence of energy sublevels

1S < 2S < 2P < 3S < 3P < 4S < 3d < 4P < 5S < 4d < 5P < 6S < 4f < 5d < 6P < 7S < 5f < 6d

Description of long periodic table :



1- S – block consists of two groups because the S – sublevel consists of one orbital which is filled with two electrons only .

2- P – block consists of (6) groups because the P – sublevel consists of three orbitals which filled with six electrons .

3- d – block consists of (10) groups because the d – sublevel consists of five orbitals which are filled with ten electrons .

4- f block are separated from the table so that the table is not too wide (long)

5- The first period contains two elements because it consists of elements of the sublevel 1S = 2 electrons .

6- The second period contains eight elements because it consists of the sublevel (2S + 2P) = 8 electrons .

7- The third period contains eight elements because it consists of element of sublevel (3S + 3P) = (2 + 6) = 8 electrons .

8- The fourth period contains eighteen elements because it consists of elements of the sublevels (4S + 3d + 4P) = 2 + 10 + 6 = 18 electrons .

9- The fifth period contains (32) elements because it consists of elements of the sublevels (5S + 4d + 5P) = 2 + 10 + 6 = 18 electrons .

10- The six period contains (32) elements because it consists of elements of the sublevels (6S + 4P + 5d + 6P) = 2 + 14 + 10 + 6 = 32 electrons .

How can you find the location and the type of element in the periodic table ?

1- Write the electronic configuration of element in quantum levels .
2- Number of period = The maximum value of principle energy level (quantum number) .

Example :

Find the number of period and group for each of the following element :
Na : The atomic number = 11
Cl : The aomic number = 17

Noble gases : They are the elements of the last column of the P – block all their energy levels are completely filled with electrons .

The representative elements : They are the elements of main group (S – and P – blocks) all their energy levels are completely filled with electrons except for the external energy level .

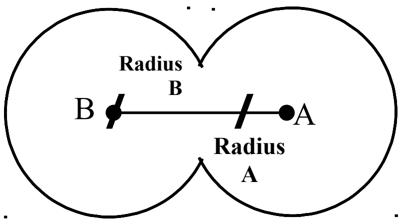
The transition elements : They are the elements of the d – block all their energy levels are completely filled with electrons except for the two external energy levels .

The inner transition elements : They are the elements of the f – block all their energy levels are completely filled with electrons except for the three external levels .

Trends and periodicity of properties in the periodic table

The atomic radius :
We can not determine the atomic radius because electron has a wave motion so it is impossible to determine exactly the location of an electron around the nucleus .

The atomic radius :
It is half the distance between centers of two similar atoms in a diatomic molecule .



The bond length : It is the distance between the nuclei of two bonded atoms .

There are many methods to measure the bond length such as :

1- X – ray . 2- Electron diffraction .

Examples

1- The bond length in the chloride molecule Cl – Cl is 1.98 Å and the length between carbon and chloride atoms C – Cl is 1.76 Å . Calculate the atomic radius of carbon .

solution

The atomic radius of chlorine = $\frac{1.98}{2}$ = 0.99 Å

The atomic radius of carbon = 1.76 – 0.99 = 0.77 Å

3- The bond length in the molecule of NH_3 is 1.0 Å and the bond length in the molecule of H_2 is 0.6 Å . Calculate the bond length in nitrogen molecule (N_2) ?

solution

The atomic radius of hydrogen = $\frac{0.6}{2}$ = 0.3 Å

The atomic radius of nitrogen = 1 – 0.3 = 0.7 Å

The bond length of the N_2 = 0.7 × 2 = 1.4 Å

Atomic radius decrease in period by increasing atomic number ?

Because of increase in the atomic number gradually makes to increase the positive nuclear charge therefore the attractive force of the nucleus for electrons will be increased and atomic radius decreased .

Atomic radius increases in group by increasing of atomic number why ?

Because the increase in atomic number makes to increase the number of energy level screen the attractive force of the nucleus for valence electrons and increase the repulsion force between electrons therefore the atomic radius increases .

N.B.

1- The cation (+ve ion) radius is smaller than that of its atom ?
This is due to the increasing of positive charges of the protons which attract the valency electrons leading to a decrease in the cation radius .

2- The anion's radius is bigger than that of its atom ?
This is due to the increase in the number of the negative charges in the shells therefore the repulsive force between the electrons increases so the shells move a part and this leads to increasing of the anionic radius than that of its atom .

Ionization potential

It is the amount of energy required to remove the smallest bounded electron completely from an isolated gaseous atom

$M \longrightarrow M^+ + e^-$ $\Delta H = +$ ve KJ/Mole

The atom losses electrons and converted into positive ions . It has a positive value (ΔH = + ve KJ/Mole) .

1- **The first ionization energy** : **It is the energy required to remove one electron from neutral atom to form a cation (+ ve) with one positive charge .**

$M \longrightarrow M^+ + e^-$ $\Delta H = +$ ve KJ/Mole

2- The first ionization energy of noble gas is very high ?

Due to the stability of their electronic configuration because it is difficult to remove an electron from completely filled shell .

3- The ionization energy of element of group (5A)

(N_7 – P_{15} – As_{33} – Sb_{51} – Bi_{83}) is much greater than any element have the same period because the outer most energy sublevel (P) has three electrons and it is half filled with electrons (nP) and this gives the atom of the element some extra stability so the ionization energy is greater .

4- The ionization energy of sodium is much smaller than that of chlorine ?

Because the atomic size of chlorine is smaller than that of sodium so the attractive force of the nucleus on the valence electrons in the case of chlorine is more strongly and the electrons valence need a higher energy to be separated from the atom .

5- Ionization energy increases period ?

Because the positive nuclear charge gradually increases with the increase of the atomic number led to decrease the atomic radius and increase the attractive force of the nucleus on the valence electrons therefore the electrons needed m large high energy to remove (separated) from the atom .

6- Ionization energy decreases in group ?

Due to the increase in the atomic size and screen of the attraction force of the nucleus on the valence electrons therefore the electrons needed a smaller value of energy to separated from the atom .

<u>7- The ionization energy of elements of group (2A) is much greater than any element have the same period ?</u>
Because the outer most energy level (S) of the element of group (2A) is completely filled with electron (nS^2) and this gives the atom of the element some extra stability so the ionization energy is much greater .

<u>Electron affinity</u>

It is the amount of energy released when an extra electron is added to a neutral gaseous atom to form an anion (- ve ion)
$X + e^- \longrightarrow X^- + \text{energy} \qquad \Delta H = -\text{ ve KJ/Mole}$

<u>G.R.F.</u> In the horizontal periods electron affinity increases with the increase in atomic number ?
Due to the atomic radius (size) gradual decrease so it becomes easier for the nucleus to attract the new electron .

<u>G.R.F.</u> The electron affinity decreases in group ?
Due to the increase of the atomic volume with increase atomic number and this leads to the screeing of the attraction force of the nucleus on the valence electrons .

<u>Exception cases :</u>
Beryllium has a relatively high of electrons affinity due to the stability of its atom that has completely filled orbitals ($1S^2$, $2S^2$) ?
Because the outer most energy sublevel (nS) is completely filled with electron and it gives the atom some extra stability .

<u>Elements of the fifth group (N_7 , P_{15}) have a lower value of electron affinity</u>
Because the outer most energy sublevel (nP) has three electrons and it is half filled with electrons it gives the atom some extra stability (N_7 : $1S^2$, $2S^2$, $2P^2$).

<u>Noble gases have not (small) electron affinity</u>

Because all energy sublevels are completely filled with electrons which gives the atoms great stability .

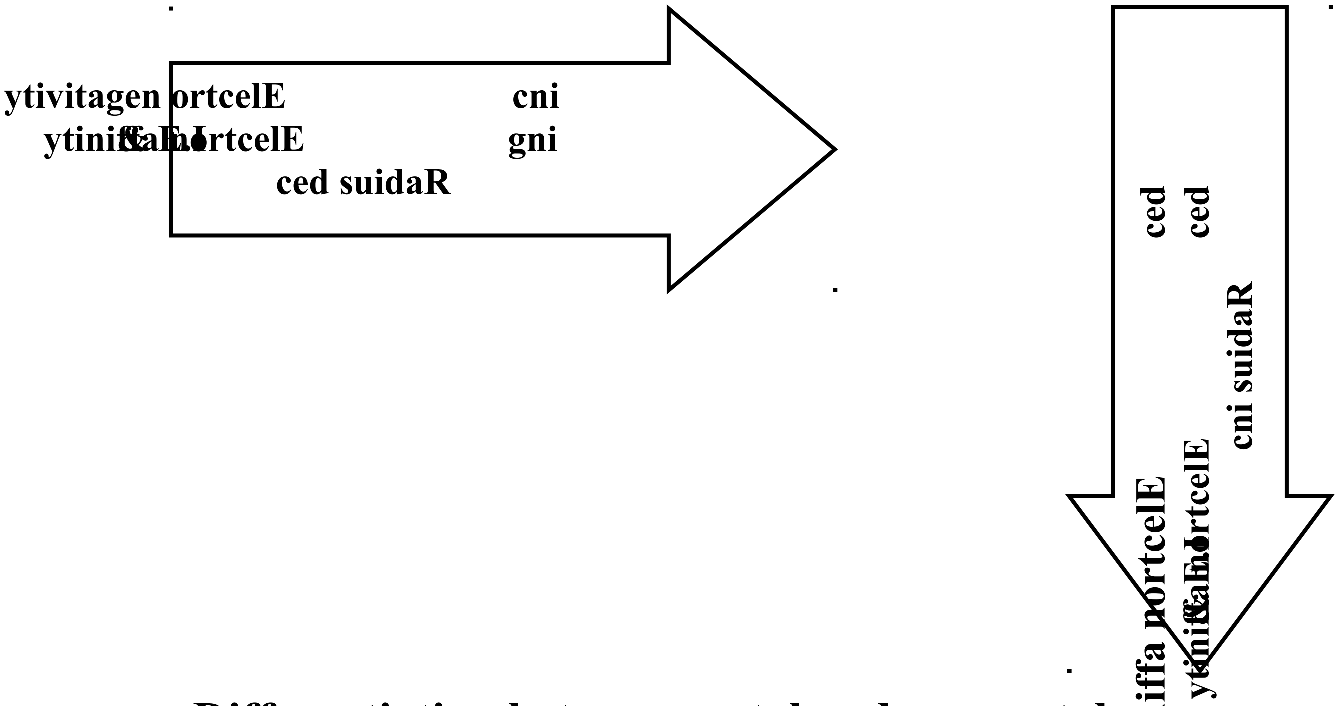
Electron affinity of Fluorine (F_9) is less than that of chlorine (Cl_{17}) ?
Because the atomic radius (size) of fluorine atom is smaller than that of chlorine atom and when fluorine atom gains electron it is affected by a great repulsion force bigger than that in chlorine atom and fluorine atom is very small size .

Electro negativity

The tendency of an atom to attract the electrons of chemical bond to itself .
It is the querage of the ionization potential and electron affinity .

Compare between ionization energy , electron affinity and electro negativity

Ionization energy	Electron affinity	Electro negativity
It is the amount of energy needed to remove the least connected electron bond in a single atom	It is the amount of energy released when an extra electron is added to a neutral single atom to form an ion	It is tendency of an atom to attract the electrons of chemical bond to itself
It refers to the atom in its single state	It refers to the atom in its single state	It refers to the atoms which linked together in the molecule
It is inversely proportional to the atomic radius	It is inversely proportional to the atomic radius	It is inversely proportional to the atomic radius
The atom losses electrons and converted into positive ion	The atom gains electrons and convers into negative ion	
It has a positive value	It has a negative value ΔH = - value type of reaction is exothermic	



Differentiation between metal and non metal

Metals	Non metals
Valence shell has less than help its capacity of electrons (1 or 2 or 3)	Valence shell has more than half its capacity of electrons (5 or 6 or 7)
They one called electropositive elements .	They are called electronegative elements
They have Relatively large atomic radius therefore the ionization energy and electron affinity and electro negativity have smell values	They have small atomic radius ionization energy and electro negativity and electron affinity have high value
They have a good conductivity of electricity .	they have a bad conductivity of electricity .

N.B.

1- Fluorine is the strongest non metal while caesium is the strongest metals Because fluorine has smallest radius while caesium has the biggest radius .

2- Metals are considered as electropositive elements because metals lose electrons to form positive ions .

3- Non metals are considered as electronegative elements because non metals gains electrons to from negative ions .
4- Metal are good conductors of electricity because they have few valence electrons which can transfer easily from one position to another in the metal structure .

5- Non metals are bad conductors of electricity because their valence electrons are strongly bounded to the nucleus due to the small atomic size therefore it is difficult for the valency electrons to be transferred .

6- Metals have small values for ionization energy and electron affinity because they have large atomic radius .

7- The strongest metal lie at the bottom on the left and side of the periodic table because of the increasing in the atomic number the atomic radius increases gradually so the attractive force of the nucleus to the valency electrons decreases therefore it is very easy for the atom to lose the valency electrons so the metalic property increases .

Metalloids

1- They are elements whose valency shell contains 4 electrons .
2- Sometimes they act as metal (when they gains electrons) .
3- They act as semiconductors (Boron – silicon) which are used in or transistors and knows as semiconductors electronic instruments .

metals $\xrightarrow[\text{in period}]{\text{decreases}}$ metalloids $\xrightarrow[\text{in period}]{\text{increase}}$ non metal

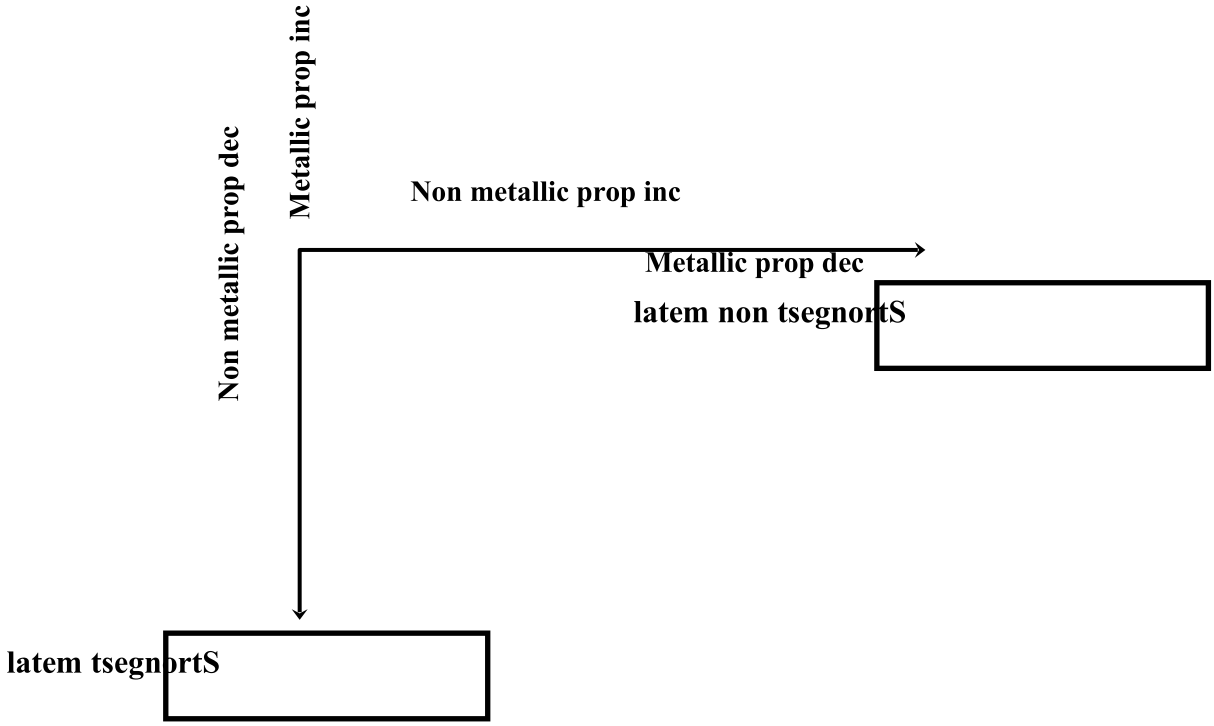
* Because the atomic size decrease and the attractive force of nucleus on the valence electron will be increased therefore it is difficult for the atom to lose the valence electron .

Metalloids increase in period because the atomic size decrease and the attractive force of nucleus to electron will be increased therefore it becomes easier for the nucleus to gain a new electron .

In groups : Metals increase in group because the increase in the atomic number makes to increase of energy levels and screen (decrease) the attractive force of the nucleus on the valence electrons therefore it is easy for the atom to lose the valence electrons .

In groups : non metals decrease because the increase in the atomic number makes to increase the energy level and screen the attractive force of nucleus on the valence electrons therefore it is difficult for the atom to gain a new electron .

Fluorine is considered as the strongest non – metal because the atomic size of fluorine is very small therefore the attractive force of the nucleus to the electrons will be increased therefore it is very easy for the atom to gain a new



electrons .

Acidic – basic properties

Acidic oxides	Basic oxides
They are non – metalic oxides such as CO_2 , SO_2 , SO_3 , P_2O_5	**They are metalic oxides such as Na_2O , K_2O , MnO , CaO , BaO**
They dissolve in water to form acids $Co_2 + H_2O \rightarrow H_2CO_3$ $SO_2 + H_2O \rightarrow H_2SO_3$ $SO_3 + H_2O \rightarrow H_2SO_4$ $P_2O_5 + H_2O \rightarrow 2H_3PO_4$	**Some basic oxides dissolve in water to form alkalies and others are not** $Na_2O + H_2O \rightarrow 2NaOH$ $K_2O + H_2O \rightarrow 2KOH$ $CaO + H_2O \rightarrow Ca(OH)_2$ $MgO + H_2O \rightarrow Mg(OH)_2$
They react with alkalis to form salt and water $CO_2 + 2NaOH \rightarrow Na_2CO_3 + H_2O$ $SO_2 + 2NaOH \rightarrow Na_2SO_4 + H_2O$	**They react with acids to form salt and water** $Na_2O + 2Hcl \rightarrow H_2O + 2Nacl$ $MgO + H_2SO_4 \rightarrow MgSO_4 + H_2O$
They do not reaction with acids	**They do not react with alkalis**

<u>The oxyacids</u> : are acids that contain hydrogen, oxygen and a third element usually a non-metal .

- It can take the following symbols: - M On (OH)m

where , (M) is the atom of the element .
(n) is the number of oxygen atoms .
(m) is the number of hydroxyl groups .

The strength of oxyacids are depends on the number of oxygen atoms which does not linked with hydrogen atoms when this number increase , the strength of the acid .

Acid Name	The number of free atoms of Oxygen	Acidic Property
H_4SiO_4	-	Weak acid
H_3PO_4	1	Moderate acid
H_2SO_4	2	Strong acid
$HClO_4$	3	Very strong acid

<u>Amphoteric oxides :</u>

1- They react with bases and acids to form (give) salt and water in both cases .

2- They have both acidic and basic properties such as Al_2O_3 , ZnO , Sb_2O_3 , SnO .

How can you obtain CuO from mixture of ($CuO + Al_2O_3$) ?

3- KOH is strong base than of NaOH because the atomic radius of K is greater than that of Na so the bond strength between K and hydroxide group (OH^-) decreases therefore it is easy to separate hydroxide ions to form strong base .

4- HF is weak acid because the atomic size of fluorine atom is very small so it has ability to attract the electrons bond to itself therefore it is difficult to separate hydrogen ions to form strong acid .

5- HI is stronger acid because the atomic (size) radius of iodine atom is largest to it has a weak ability to attract the electron bond to itself therefore it is easy to separate hydrogen (H) ions to form strong acid .

<u>Oxidation number</u> : at last we defined it as it is the number of hydrogen atoms that combine with or can be replaced by an atom of the element .

<u>The modern definition</u> : the number of single (unpaired) electrons in the valence shell of the atom .

N_7	**$1S^2$**	**$2S^2$**	**$2p^3$**	**Trivalet**
O_8	**$1S^2$**	**$2S^2$**	**$2p^4$**	**Divalent**
F_9	**$1S^2$**	**$2S^2$**	**$2p^5$**	**Mono valet**

<u>Oxidation number</u> : it is a number that refers to the electric charge (+ve or -ve) that atom would have in the compound .

<u>Rules for assigning oxidation numbers</u>

Oxidation number of Oxygen = 2- in most of its compounds except in peroxides e.g. (hydrogen peroxide H_2O_2) , sodium peroxide Na_2O_2 and potassium peroxide K_2O_2 is (1-) supper oxide e.g. KO_2 the oxidation number of oxygen is (½-)

The oxidation number of oxygen = 2+ in OF_2 because the electro negativity of fluorine is higher than the electron negativity of oxygen .

The oxidation number of hydrogen in its hydrides = 1- such as NaH , KH , CaH_2 .

The oxidation number of any chemical compound = zero because the algebraic sum of the oxidation number of its atom = zero .
Such as NaCl , HCl , CuSO4

The oxidation number of any element in its pure state = zero .
e.g. Fe , Cl , Na , O_2S , P4 .

The oxidation number for any atomic group (poly atomic ion) = the number of charge on the group (ion) such as OH^- , $SO4^{--}$, $CO3^{--}$, $NH4^+$, $PO4^{---}$.

Examples

Calculate the oxidation number of chlorine in these compound , and for sulphure

Nacl Nacl = zero +1 +Cl = 0 Cl = -1	**$NaClO_2$ $NaClO_2$ = zero 1+Cl 2-×2 = zero Cl = +3**	**$NaClO_4$ $NaClO_4$= zero +1 Cl -8 = 0 Cl = +7**	**H_2S = zero 2×1+S= zero S = 2-**	**$SO3^{--}$ SO3 = -2 S -6 = -2 S = 4**
K_2S K_2S = zero C + S = 0 S = -2	**$Na_2S_2O_3$ $Na_2S_2O_3$ = zero 2×1+S_2 -6= zero S = 2**	**S = zero**	**H_2SO_4= zero 2 + S-8= 0 S = 6**	**$SO4^{--}$ SO4 = -2 S -8 = -2 S = +6**

<u>The oxidation process</u> : it is the process of losing electrons due to increase the oxidation number for the element .

<u>The Reduction process</u> : it is the process of gaining electrons due to decrease the oxidation number for the element .

Oxidation	Reduction
Zero $\longrightarrow$ (+) value	(+) value $\longrightarrow$ zero
(+) small value $\longrightarrow$ (+) big value	(+) big value $\longrightarrow$ (+) small value
(-) value $\longrightarrow$ zero	zero $\longrightarrow$ (-) value
(-) big value $\longrightarrow$ (-) small value	(-) small value $\longrightarrow$ (-) big value

Explain the type of change (oxidation or Reduction for carbon in this Reaction .

$C_2H_5OH \longrightarrow$ $CH_3CHO \longrightarrow$ $CH_3COOH \longrightarrow$ CO_2

$C = -2 \longrightarrow$ $C = +1 \longrightarrow$ $C = zero \longrightarrow$ $C = +4$

Oxidation Oxidation Oxidation

Explain the type of change (oxidation or reduction) that of chromium and iron in this reaction .

$K_2CR_2O_7 + 6FeCl_2 + 14HCl \longrightarrow 2KCl + CrCl_3 + 6FeCl_3 + 7\ H_2O$

Solution

$K_2Cr_2O_7 \xrightarrow{\text{Reduction}} CrCl_3$ **$FeCl_2 \xrightarrow{\text{Reduction}} FeCl_3$**

$Cr = +6 \longrightarrow Cr = +3$ **$Fe = +2 \longrightarrow Fe = +3$**

R

Gradation of oxidation number in periodic table :

Group	I	II	III	I V	V	V I	V II
Element	Na	Mg	Al	Si	P	S	Cl
Oxide Oxidation Number	Na_2O 1 +	MgO + 2	Al_2O_3 + 3	SiO_2 + 4	P_2O_5 + 5	SO_3 + 6	Cl_2O_7 + 7

In the case of hydride

- The oxidation number of element (metal) of 1st and 3rd groups in their compounds agrees with group number (it takes positive oxidation number to which they belong)
- Most elements in the middle of the table have variable oxidation number in different compounds .

The oxidation number of the elements in groups from IV to VII = the group number -8 = - value because the elements gain electrons to complete a full shell .

Group	**I**	**II**	**III**	**I V**	**V**	**V I**	**V II**
Element	Li	Be	B	C	N	O	F

Hydride Oxidation Number	LiH	BeH_2	BH_3	CH_4	NH_3	H_2O	HF
	1 +	+ 2	+ 3	+ 4	- 3	- 2	- 1

The oxidation number of noble gases of (group zero) is zero because they do not combine to form compounds .

Questions on Chapter 2

A- Complete the following :

1- The formula of silicon dioxide is .... and the oxidation number of silicon in it is ....

2- The formula of beryllium hydride is ...., and the oxid. no. of Beryllium is .... , while the formula of ammonia is .... and the oxid. no. of nitrogen in it is …………..

3- The oxidation number of oxygen in most of its compounds is .... but in peroxides is ……...... while in the elementary state is ....

4- The oxidation number of hydrogen in most of its compounds is .... but in metal hydroxide is ……………..

5- The decrease of atomic radius of nonmetals leads to the .... of ionization potential and electron affinity.

B- Write the scientific term (or rule which explains each of the following statements:

1- Numbers which identify the orbitals, their shape and their orientation.

2- The electron is a particle which has a mass and carries a negative charge.

3- Orbitals are filled with one electron first before gaining of electron tack place.

4- Each moving body is associated with a wave motion which has some of wave properties.

5- Every mobile body has a wave motion with light wave properties.

6- The elements where the (f) and (d) orbitals are occupied but not completely filled with electrons in their atomic or ionic state.

7- Electrons occupy the orbitals in the order of increasing orbital energy, the lowest energy orbitals are filled first.

8- Half the distance between the centers of two similar atoms in diatomic molecule.

9- Elements of valence shell filled by more than its half.

10- The energy required to separate electrons less in connection by single atom in gaseous state.

11- The amount of energy released when an extra electron is added to a neutral gaseous atom .

12- The region of space around the nucleus where it is most probable to find the electron .

C- Choose the best answer:

1- The sublevel (d) contains .... orbitals, it can accommodate up to .... electron . (2 - 3 - 5 - 10 - 14)

2- An element of atomic number 10, the number of sublevels filled with electrons are .... (2 - 3 - 5)

3- The valency shell of the atoms of fifth group elements contain ....

a) np4 b) np5 c) np d) np3

4- Group 5-A elements are classified as .... block . (S-P-d-f)

5- Sublevel (f) saturated with number of electrons equals ..... .

a) 6 b) 16 c) 14 d) 32

6- The fourth energy level is saturated by a number of electrons equals ....

a) 8 b) 16 c) 32 d) 64

7- Electron affinity in group decrease ....

a) up on increasing atomic number only

b) up on decreasing atomic number only

c) up on increasing atomic volume only

d) a and c

8- The oxidation number of hydrogen in hydrogen molecule is (1, 2, 3, 0) and its oxidation number of hydrogen chloride is (1, 2, 3, 0).

9- The oxidation number of the first group elements equal ... .

a)+2 b)+1 c)-2 d)-1

10- The oxidation number of oxygen in H2O2 is ....

a)+2 b)+1 c)-2 d)-1

. ... The atomic number of element is 29, so, its electronic structure is -11

a) (Ar) $3d^{10}$, $4S^{1}$

b) (Kr) $4P^{4}$, $5S^{2}$

c) (Xe) $6P^{6}$, $6S^{2}$

12- The electron affinity of Fluorine as compared to chlorine is ....

a) Less b) more c) the same d) non of the above

13- The oxidation number of Sulphur in (SO_4) is ....

a) 2+ b)4+ c)2- d) 6+

14- The metallic property increase in the group elements by ....

a) increasing of atomic number only .
b) increasing of ionic potential .
c) increasing of boiling degree .

d) when the radius decrease .

15- The element whose atomic number is 24, its correct electronic configuration is ....

a) (Arlg) 4S2, 3d" b) (Arlg) 4S\ 3d'

c) (Arlg) 4S°, 3d4 d) (Arlg) 4S2, 3d'

16- The element 86X, its electronic configuration in the outer most energy level is ....

a) $7S^{1}$ b) $3d^{5}$ c) $4S^{2}$, $4P^{5}$ d) $6S^{2}$, $6p^{6}$

17- The element X, its atomic number is 73 is considered from .... series.

a) Lanthanides b) Main transition elements

c) Representative elements d) actinide

18- $CO_2$is .... oxide.

a) Amphoteric b) Acidic c) Neutral d) Basic

19- Na_2O is .... oxide .

a) Amphoteric b) Acidic c) Neutral d) Basic

20- $A1_2\ O_3$ is .... oxide.

a) Amphoteric b) Acidic c) Neutral d) Basic

21- The element has the outer most electronic configuration 75 is found in ....

a) the first period, group 7 b) the six period, group 1

c) the seventh period, group 1 d) the six period, group 7

22- The spin quantum number (Ms) decide ....

a) Energy sublevel and their number

b) Number of electrons in each principle energy level

c) Direction of electron movement

23- The weakest acid is ....

a) HC1 b) HBr c) HF

24- Orbitals of the same energy sublevel are ....

a) different in energy but similar in shape

b) similar in energy but different in shape

c) similar in energy and shape

25- The oxidation number of oxygen in OF_2 is ....

a) -2 b) -1 c) +2

26- The element having the electronic configuration (Ar) $3d^5$ $4S^1$ is...........

a) Representative b) Inner transition c) Main transition

27- The maximum number of electrons that saturate, the energy level (n) is ....

a) 2n b) 2n2 c) 3n2

28- Elements of group III are of the block ....

a)S b)P c)d d)F

29- In Mosely periodic table, the elements are rearranged according to:

a) mass numbers b) atomic numbers

c)atomic weight d)a & b

30- The radius of the atom is :

a) the distance between the nucleus and the farthest electron in the atom,

b) half the distance between two combining atoms in a diatomic molecule,

c) the distance between the centers of two similar atoms in a diatomic molecule

d) half the distance between centers of two similar atoms in a diatomic molecule

31- In the same group as we go downwards the ionization energy:

a) increases b)decreases c) not change

32- The electron affinity of fluorine as compared to chlorine is:

a) less b)more c) the same d) a & b

33- Metal ions are usually :

a) neutral b) positive c) negative d) acidic

- While nonmetal ions are usually:

a) positive b) negative c) neutral d) basic

34- The element where the outermost shell contain more than 4 electrons are :

a) metals b) nonmetals c)metalloids d) inert gas

35- The small radius of nonmetal make the value of ionization energy and electron affinity:

a) small b) great c) moderate d) no change

36- Na_2O is :

a) acidic oxide b) basic oxide c) amphoteric d) neutral

37- $A1_2O_3$ is:

a) acidic oxide b) basic oxide c) amphoteric d) neutral

38- P_2O_5 is:

a) acidic oxide b) basic oxide c) amphoteric d) neutral

39- The perchloric acid $HClO_4$ as compared to the suphuric acid H_2SO_4 is:

a) stronger b) weaker c) the same strength

40- The oxidation number of oxygen in potassium oxids K_2O is :

a) (1+) b) (2-) c) (1-) d) (2+)

41- The element which is found in the second period and has many oxidation numbers (5+ ——>3-):

a) nitrogen b) carbon c) aluminium d) sulphur

42- The element which has(1+)as oxidation number and sometimes (1-) is:

a) lithium b) hydrogen c) chlorine d) cesium

43- The element which has the electronic configuration $IS^2\ 2S^2\ 2P^6$ is:

a) sodium b) neon c) iron d) magnesium

and it is in the (2^{nd} , 3^{rd} , 4^{th}) period .

44- The oxidation no. of chlorine in $HClO_4$ is :

a) (1-) b) (1+) c) (5+) d) (7+)

45- The oxidation no. of manganese in $KMnO_4$ is:

a) (1-) b) (4+) c) (6+) d) (7+)

46- The oxidation on. of oxygen in H_2O_2 is:

a) (2-) b) (1-) c) (1+)

47- The oxidation on. of nitrogen in NH_2OH is:

a) (3-) b) (2-) c) (1-) d) (1+)

48- The oxidation on. of hydrogen in MgH_2 is:

a) (1+) b) (1-) c) (2+) d) (2-)

49- The value of ionization energy in the period increases as the: series

a) atomic radius decreases b) atomic ardius increases

c) atomic number increases d) a & c

50- The element (X) its atomic number is 13 is considered

a) lanthanide's b) main transition element

c) representative element d) actinide

51- The 4th period in the periodic table contains elements

a) 8 b) 18 c) 16 d) 32

52- The 6th period in the periodic table contains elements,

a) 16 b) 32 c) 48 d) 30

53- Electronegativity increases across the periods

a) atomic radius increases b) atomic radius decreases

c) atomic number decreases d) a & c

54- Electron affinity increases across the period as :

a) atomic number increaes b) atomic number decreases

c) atomic radius decreases d) a & c

55- The ionization energy of magnesium I .. than that of chlorine

a) more b) less c) equal

<u>D- Write the scientific explanation of the following :</u>

1. The radius of a cation is smaller than that of the corresponding atom .

2. The oxidation number of elements of the first group (1A) in the periodic table is (1+).
3. The oxidation of iron II ion oxidation is easier than Manganese II ion.
4. It is impossible to determined the place and speed of the electron in the same time .
5. P - sublevel saturated with 6 electrons .
6. Physically, it is impossible to measure the atomic radius of the atom.
7. Line spectra is characteristic for any element.
8. Quantum numbers can not used to find the number of electrons in energy level more than 4.
9. Elements of S, P blocks called representative elements .
10. Element of f block are called inner transition elements.
11. Electron prefers to paired with another electron in the same level instead of intering a new sublevel .
12. The third ionization potential of Magnesium is greater than first ionization potential .
13. The ionization potential of noble gas of group zero is very high .
14. In periods of periodic table the ionization energy increases as we move to right i.e. as the atomic radius decreases .
15. Non-metals do not conduct electricity.
16. In groups of periodic table the ionization energy decreases with the increasing the atomic number .
17. No electron pairing takes place in a given sublevel until one electron each orbital contain one electron .
18. Both fluoride -ve ion and sodium +ve ion have the same number of electron.
19. Spin quantum number has only two values .
20. All nobles gases has oxidation number equal zero .
21. Strong acids, strong base lies at the bottom of the periodic table.
22. Potassium hydroxide is strong base than that of sodium hydroxide.
23. Fluorine is considered as the strongest non-metal .
24. Iodine is considered as the weakness non-metal .

25. Fluorine is considered the most electronegative element .
26. Chlorine has a higher electronegative than that of sodium in sodium chloride .
27. Electron affinity of fluorine is less than that of chlorine .
28. Carbon has higher electron affinity.
29. Fluorine has higher electron affinity.
30. Beryllium has lower electron affinity.
31. Noble gas (neon) has not electron affinity.
32. Nitrogen-atom is trivalent while fluorine atom is monovalent.
33. Sulphur has divalent and hexavalent.
34. Cesium is strong metal while fluorine is a strong non-metal.
35. The P-block elements contain six groups.
36. The metals are electropositive elements.
37. The bond in hydrogen molecule is shorter than the bond in chlorine molecule.
38. Barium oxide is basic oxide.
39. Sulphur dioxide is acidic oxide.
40. Cesium hydroxide is stronger alkali than potassium hydroxide.
41. It is impossible to find the following energy sublevels in any atom : IP - 2S - 3S - 3P - 5d.
42. Sodium $_{11}Na$ is softer than Aluiminium $_{13}A1$.
43. Elements of group 1-A have the largest atomic radius of atoms in the period table.
44. The two electrons in an orbital of one energy sublevel do not repel each other.
45. Carbon dioxide is an acidic oxide while potassium .oxide is a basic oxide.
46. Aluimnium oxide is considered an amphoteric oxide. Show by equation .

<u>1-A elements in the periodic table regarding each of the following :</u>

Atomic radius – electron affinity – ionization energy

B -**Compare** between Atomic radius - electron affinity - ionization energy

(G) What is meant by : the dual nature of electron - Hund's rule - Electro negativity - lonization potential - oxidation number - Amphoteric oxide - line spectra - the excited atom - Quantum - De Broglie principle - Auf bau principle - inner transistion elements - electron affinity -oxidation process - Reduction process - Heisenberg uncertainty principle - transition element - noble gas -electron cloud .

Chapter three

Chemical reaction

Chemical reaction:

- Is the reaction in which the bonds of reactants are broken forming new bonds in products.
- (Inert gases are chemically inactive). Because their outer energy levels are completely filled with electrons (stable elements) where they have high I.P. & low E.A.
- On mixing iron fillings with sulphur the result will be a mixture not a compound.

Because: There is no chemical bond formed between iron and sulphur.

But if this mixture is heated enough to form new bond the result will be a compound called iron sulphide FeS.

Type of chemical bonds:

1) Ionic bond.
2) Covalent Bond.
3) Co-ordinate bond.
4) Hydrogen bond.
5) Metallic Bond.

1)Ionic bond:

- This bond is usually formed between metals and nonmetals. It is known that atoms of metals are characterized by large volumes (atomic radius). Accordingly, their ionization energies are low. This facilitates the loss of their few electrons of the outermost shell. Metal atoms are then changed to cations with an identical electron structure to the nearest noble gas in the periodic table.
- On the other hand, nonmetal atoms are characterized by their small volumes.

  Accordingly, their electron affinities are high this facilitates the gaining of electrons (those lost by metal atoms). Non-metal atoms are changed to anions with an identical electron structure to the nearest noble gas in the periodic table.

- **Consequently, an electrostatic attraction occurs between (+ve) cations & (-ve) anions. It is called the ionic bond.**
- **This means that ionic bond has no materialistic existence.**

Examples: Formation of NaCl:

$Na \longrightarrow Na^+ + e^-$ $\qquad$ $Cl + e^- \longrightarrow Cl^-$

2, 8, 1 $\qquad$ 2, 8 $\qquad$ 2, 8, 7 $\qquad$ 2, 8, 8

$Na^+ + Cl^- \longrightarrow NaCl$

- **Ionic bond is formed between atoms when the difference in E.N between them is higher than 1.7**
- **As the difference in E.N. between atoms increases, the strength of the ionic bond increases which increases the melting point, boiling point and degree of conductivity.**

	Na	Mg	Al	Cl
Electro negativity	**0.9**	**1.2**	**1.5**	**3**
Different in electronegativity	**Nacl 3-.09=2.1**	**Mgcl2 3-1.2=1.8**	**Alcl 3-1.5=1.5**	
Melting point (C)	**810**	**714**	**190**	
Boling point(C)	**1465**	**1412**	**Changing directly from solid to gas (sublimes)**	
Conductivity of electricity	**Very good conductor**	**Good conductor**	**Does not conduct (covalent bond)**	

General properties of ionic compounds:

(1) Structure:

-These are crystals that are condtructed of collections of cations and anions bound by electrostatic forces in crystal lattice containing the ion in a regular pattern.

(2) Melting and boiling points:

-Ionic compounds generally have high melting and boiling points because a great amount of energy is needed to break down the crystal lattic and

overcome the strong electrostatic attraction force between cations and anions.

II- covalent Bond:

-Formed between atoms of non-metals of the same element (have the same electrogativity) or between atoms of different elements have difference in E.N. less than 1.7 it occurs by sharing of valence electrons and is divided into two types:

1- Pure covalent:

Formed between 2 similar atoms have the same E.N(difference in E.N = Zero)

Ex: F_2, Cl_2, O_2, H_2

-In this case, the two atoms have the same E.N. (same ability to attract the pair of electrons to itself). Thus, the electron pair spends the same time in the vicinity of each atom and the net charge on each atom is zero.

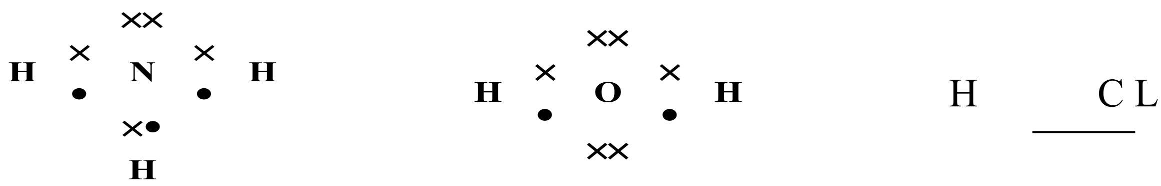
2- Polar covalent:

-Formed between 2 atoms have difference in E.N less than 1.7

Ex: HCl molecule

-In this case, because chlorine atom has more E.N., so it has greater ability to attract the pair of electrons of the covalent bond (i.e. the electrons spend more time at a chlorine atom). As a result, chlorine atom acquires a partial negative charge (-δ) and not complete one (as in the case of chloride ion Cl-), while hydrogen atom acquires a partial positive charge (+δ).

Polar molecules:



Ammonia (NH_3) Water (H_2O) Hydrogen chloride

General properties of covalent compounds:

1)Electrical conductivity:

-Ions are responsible for electrical conductivity in solutions. Since covalent compounds are not normally ionizable, so they do not conduct electric current as liquids or in aqueous solution.

2)Melting and boiling points:

-Covalent compounds are characterized by their relatively low melting & boiling points due to the very weak attraction force between their molecules which needs low amount of thermal energy to be separated.

Give reason: Solution of HCL in benzene does not conduct electricity, but its solution in

H_2O is a good conductor of electricity.

Answer: Because HCL is a polar covalent compound which can diffuse in benzene (non polar solvent) but can't be ionized into (+ve) ion (H+) & (-ve) ion (CL-), so it can't conduct electricity. While HCL in H2O (polar solvent) can be ionized into ions which can migrate to the opposite electrodes and so conducts electricity.

$$HCL + H_2O \longrightarrow H_3O^+ + CL^-$$

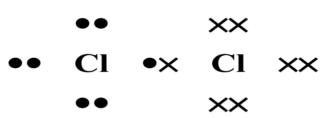
Explanation of covalent bonds:

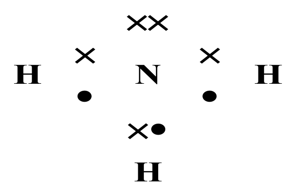
(1)Electronic theory of valency: (Octet rule Theory)

-Lewis and kosel scientist supposed that:

" all atoms of elements have tendency to reach the octet structure of the outer energy level for the nearest inert gas expect ……..''

Example:





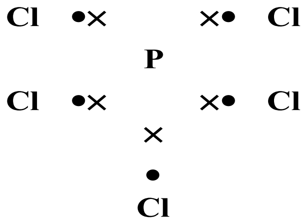
•Defects of electronic theory of valency:

(1) It failed to explain the binding in many molecules.

Which No of e is around central atom is more or less in which than 8

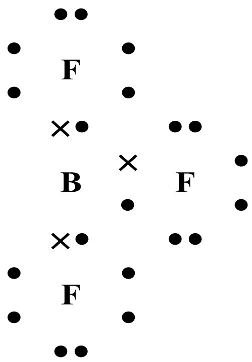
(i) In PCL5:

Phosphorus is surrounded by 10 electrons.



(ii) In BF3:

Boron is surrounded by 6 electrons.



(2) It couldn't explain some properties of molecules as stereo structure and angles between bonds.

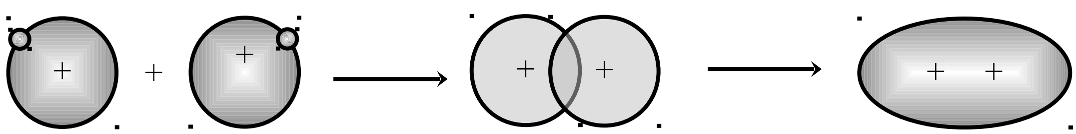
Give reason: **Octet rule can't be applied for PCI and BF3**.

•The valency Bond Theory: (V.B.T.)

-Electrons has wave property so the formation of covalent bond as a result of overlapping of an atomic orbital of an atom with an unpaired electron, with another orbital in another orbital in another atom has an unpaired electron to form a molecular orbital contains a pair of electrons.

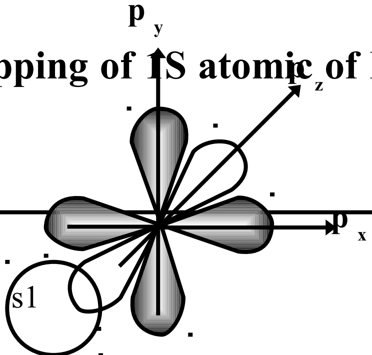
1-H_2 Molecule:

H2 molecule is formed as a result of overlapping of the e' of 1s orbital of each atom

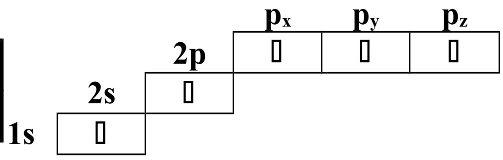


2-HF molecule: 1H 1S

HF molecule is formed as a result of overlapping of 1S atomic of H atom with 2P atomic orbital of F atom.



$_9F$	$1s^2 - 2s^2 - 2p^5$
$_1H$	$1s^1$



3-NH3 molecule:

H_1 : $1S^1$

N_7 $1S^2$, $2S^2$, $2P^1_x$, $2P^1_y$, $2P^1_z$

•NH_3 is formed as a result of overlapping of

P_x , p_y ,p_z Orbitals of (N) atom with 3(1S) orbitals of hydrogen atoms.

Q: How does the valence bond theory explain the structure of methane?

-There are 2 single electrons in carbon atom, but in methane molecule, the carbon atom forms 4 covalent bonds. So, the carbon atom must have 4 single electrons. How?? By exciting one electron from 2S to the vacant orbital 2P.

-Now, the carbon atoms has 4 single electrons, but they aren't equivalent in energy as one electron is located in 2S orbital which is lower in energy than 2P orbital. Then they must be = in energy. How?? By hybridization between one orbital of 2S and 3 orbitals of 2P forming 4 orbitals equivalent in shape and energy.

-Each of the hybridized orbitals in a carbon atom contains a (-ve) electron. These orbitals must go as far a part as possible from the other orbitals to decrease the repulsion forces between orbitals. When the angles between orbitals are 109 28, they will be more stable (less repulsive) compared to angles of 90 (an alternate structure). To complete the methane molecule, the four equivalent electrons of the four hybridized orbitals of the carbon atom can overlap with the 4(1S) electrons of the 4 hydrogen atom.

•Explain methane Molecular Structure: (CH4):

(1) Type of hybridization: SP^3.

(2) Angle between bonds: 109- 28

(3)Stereo structure: tetrahedron pyramid.

(4)Bonds: Single covalent bond of the type sigma.

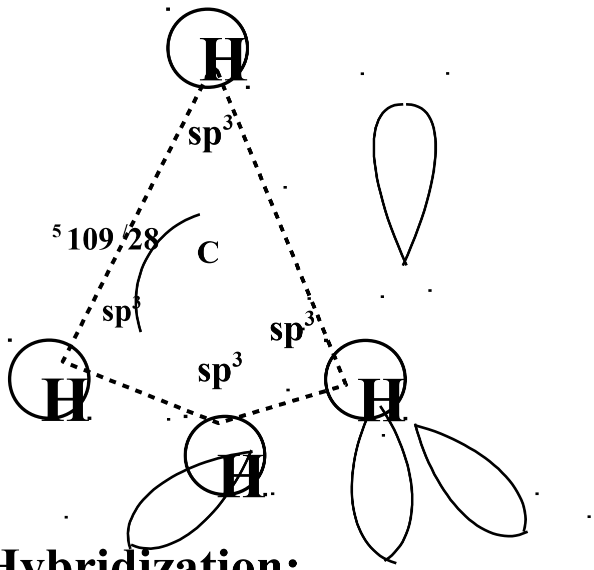
4 (c – H) bonds formed du to overlapping of 4SP^3 orbitals of one carbon atom with 4(1S) orbitals of 4 hydrogen atoms

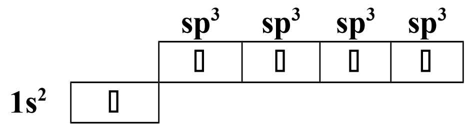
- **Methane is chemically in active, due to the presence of 4 sigma bonds in its structure which are very strong (can't be broken easily), so great amount of energy is needed to break them down.**

How does the valence bond theory explain the structure of methane:

$2p^2$ $2s^2$ $1s^2$ Carbon atom in the ground state

$2p^3$ $2s^1$ $1s^2$ Excited carbon atom





Hybridized carbon atom

•Hybridization:

"Is the combination of orbitals of close energy in the same atom to form a number of equivalent orbitals that can take part in chemical combination."

•Properties of hybridization:

(1)Hybridization occurs between orbitals of close energy for the same atom.

(2)Hybridization occurs after excitation.

(3) Number of hybridized orbitals equal number of pure orbitals taking part in hybridization.

Example:

Hybridization of 1S with 1P gives 2SP orbitals.
Hybridization of 1S with 2P gives $3SP^2$ orbitals.
Hybridization of 1S with 3P gives $4SP^3$ orbitals.

(4) Hybridized orbitals are equal in shape and energy; also angles between them are equal.
(5) The shape of the hybridized molecular orbitals differ from these of the pure atomic orbitals forming them. The hybridized molecular orbitals must protrude to the outside to be more capable of overlapping than the pure atomic orbitals.

Molecular Orbital Theory: (M.O.T)

- Considers the molecule as one unit (or a big atom with multi – nuclei) in which some of atomic orbitals of the combined atoms overlap forming molecular orbitals.
- The molecular orbitals have symbols sigma (δ) & (Л)

Compare between sigma & (pi) bonds:

Sigma Bond (δ)	Pi-Bond (Л)
1- It is formed by overlapping of atomic orbitals head to head.	1- It is formed by overlapping of atomic orbitals side by side .
2- Overlapped orbitals are one the same axis (same line)	2- Overlapped orbitals are parallel .
3- Collinear overlap .	3-Collinear overlap .
4- Strong due to great orbital overlapping (high electronic density) .	4-weak due to lees orbital overlapping (LOWER
5- Between(a)pure – hybridized orbitals (b)Hybridized-hybridized orbitlas	
6- Makes organic compounds less active	6-Makes organic compounds more active

Explain ethylene molecular structure : { $C_2 H_4$ } (ethane)

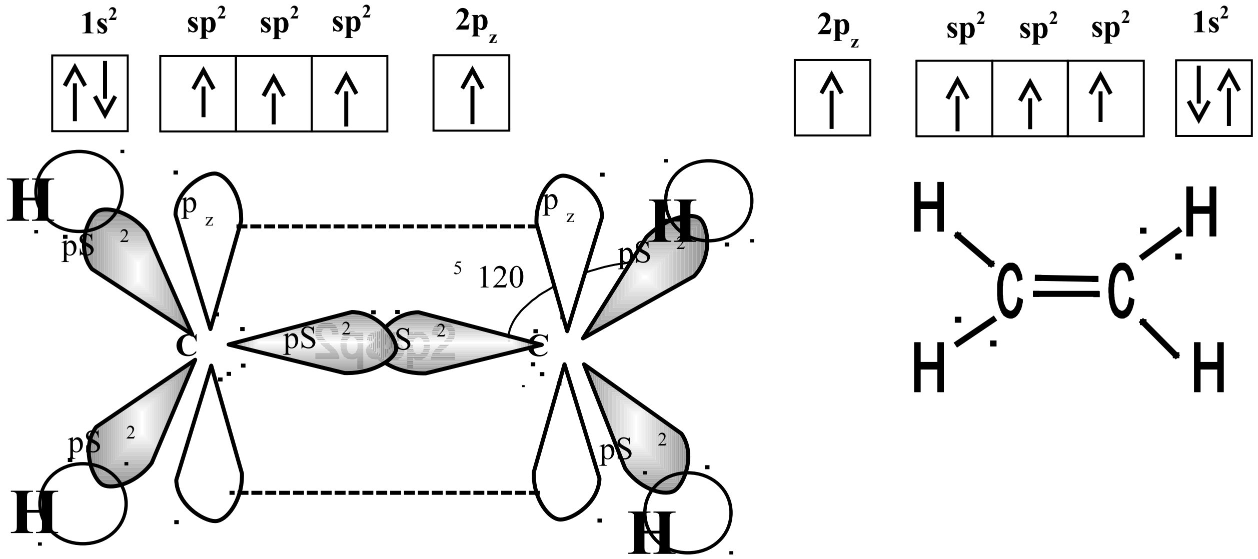
(1) type (kind) of hybridization = sp^2

(2) angle between bonds = 120 º

(3) stereo structure = planer triangular structure .

(4) bonds : di – cover bond of the type <u>sigma</u> and <u>pi</u> .

- {C——C } formed due to overlapping of one sp^2 orbital of a c atom with another sp^2 orbital of a another c atom . { sigma }
- {C——C } formed due to overlapping of pure $2p_z$ orbital of a c atom with another $2p_z$ orbital of a another c atom . { pi }
- {C ——H } formed due to overlapping of one sp^2 hybridized orbital of a c atom with one pure (1S)orbital of H atom . { sigma }



<u>Explain acetylene molecular structure :</u>

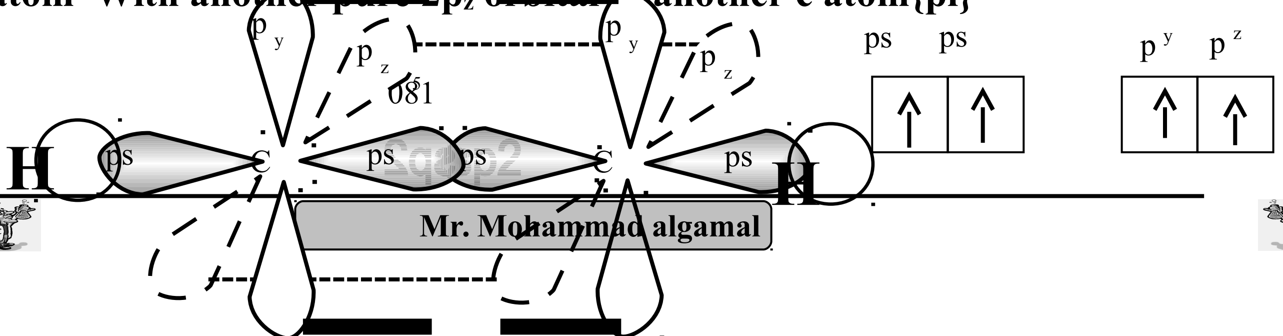
(1) type of hybridization : sp

(2) angle between bonds : 180o

(3) stereo structure : linear structure .

(4) bonds : tri – covalent bond .

- { C— C } formed due to overlapping of one sp orbital of c atom With one sp orbital of another c atom . { sigma }
- { C—H} formed due to overlapping of one sp orbital of c atom With pure 1S orbital of one hydrogen atom . { sigma }
- { C— C} formed due to overlapping of one $2p_y$ orbital of c atom With another pure $2p_y$ orbital of another C atom{ pi }
- { C— C} formed due to overlapping of one pure $2p_z$ orbital of c atom With ano[redacted]$2p_z$ o[redacted]f another c atom{pi}



H C≡C−H

Point of comparison	Methane CH_4	ethylene C_2H_4 (ethane)	Acetylene C_2H_2 (ethane)
No and type of hybridization	$1s+3p=4sp^3$	$1s+2p = 3sp^2$	$1s+1p = 2sp$
Angle between bonds	109o 28	120o	180o
Stereo structure	Tetrahedron pyramid	Planer triangle	Liner

III – Co – ordinate bond :-

" is a type of covalent bond formed between 2 atoms on of theme has one Orbital containing alone pair of electrons which is called donar atom , while The other atom has a vacant orbital called acceptor atom '

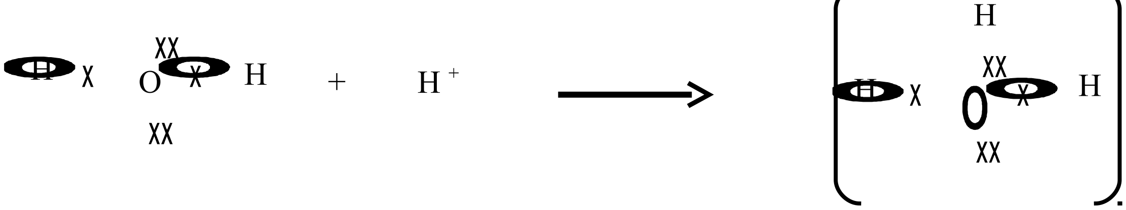
The lone pair of electrons are original from one atom .

Example :

<u>(1) hydronium l on (hydroxonium) H3O+</u>

Is formed when a strong acid dissolved in water :

(donor) (proton acceptor) (Hydronium ion)



<u>Give reason:</u> proton of stong acid does not exist freely in water

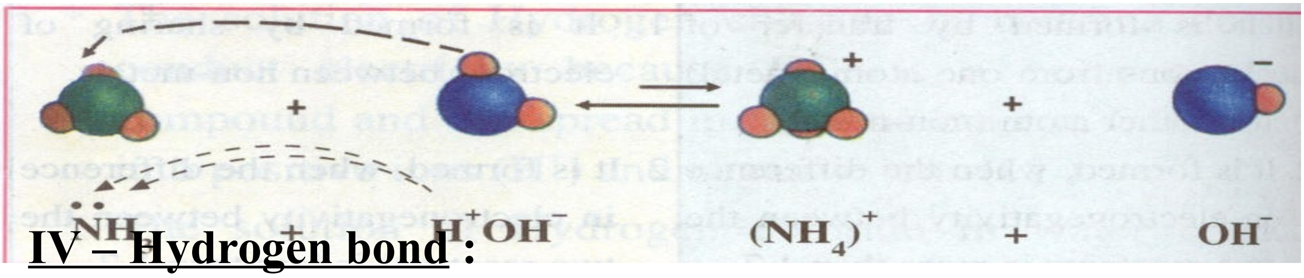
(3) <u>Ammonium lon (NH4) + :</u>

-in the last example , proton is acceptor while central atom is donor like oxygen in H_3O^+ , phosphorous in PH_4^+ & nitrogen in NH_4^+.

-also types of bonds in the last examples are polar covalent and co – ordinate bonds .

Q : compare between covalent and co-ordinate bonds .

Definition with examples .



IV – Hydrogen bond :

***is a bond formed between polar molecules in which hydrogen atoms lies between to atoms of high electron gativity as (oxygen) or (fluorine) , so the hydrogen atom binds with one atom by polar covalent bond and binds with the second atom by hydrogen bond .**

****So hydrogen atom acts as a bridge to bind molecules together .**

Explanation of hydrogen bond in water :

(1) oxygen atoms has small volume , so it has high electronegativity (3.5) , while electronegativity of hydrogen is 2.1 . so oxygen atom will carry a -8 charge , while hydrogen atom will carry a (+ s) charge .

(2) hydrogen bond is formed due to the attraction force between one hydrogen atom of one molecule and one molecule and one oxygen atom of another molecule , so molecule of water are collected by hydrogen bonds , so water exists in a liquid state and has high boiling point .

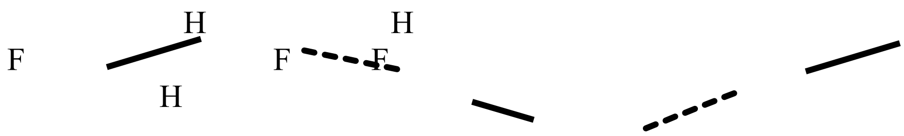
O — H --- O H

H H

<u>Give reason</u>: Although molecular weight of water (H_2O) is very small (18) but it exists in a liquid state and boils at 100 C, while molecular weight of hydrogen sulphide (H_2S) is (34) but it exists in a gaseous state and boils at (-61 C).

Answer: Due to the presence of big difference in E.N. between hydrogen and oxygen and so formation of hydrogen bond between molecules of water.

- ## Hydrogen Bond in HF:



Give reason: Although sugar is covalent compound but it dissolves in water.

Answer: Due to formation of H2 bond between hydroxyl group of sugar & oxygen of H_2O, but its solution is a bad conductor of electricity because it can't be ionized.

- ## Properties of hydrogen bond:

1- Strength of H-Bond depends on the difference in electronegativity increases, the strength also increases and the boiling point will be high as in water.

2- H-Bond is longer than covalent bond.

3- H-Bond is much weaker than covalent bond.

4- H-Bond has several forms:

A- Straight line. B- Closed ring. C- Open net.

	Covalent bond	H2 Bond
B.L.	1 A	3 A
Strength in (k.j)	418	21

V- Metallic Bond (between atoms of metal in the metallic structure):

"Is formed from electron cloud of the free valence electrons around (+ve) metal ions."

- **The free valency electrons of the outer shell are associated together forming an electron cloud which decreases the repulsion force between (+ve) ions in the metallic structure. The strength of the metallic bond depends on no of free valnce electrons. As the no of free valence electrons increases, the atoms of metal will be strongly bonded, so the metal will be harder, of higher melting & boiling points and higher thermal and electrical conductivity.**

- **Give reason: elements of group IA as Na are soft and have low melting point while elements of group IIIA as AI are hard and have high melting point.**

- **Answer: In case of Na: due to weak metallic bond which depends only on one electron from ns, while in case of Al: due to strong metallic bond which depends on three valency electrons of ns, np.**

Give reason: elements of 1st transition series are hard except Cu is relatively soft and has low melting point.

Answer: in case of T.E: due to strong metallic bond as it depends on electrons of 4s & 3d but Cu_{29} ( ) due to weak metallic bond which bond which depends only on one electron of 4S.

- **Explain types of bond in the following:**

Nacl	**water**	**hydronium ion**	**chlorine molecule**
Iron	**Aluminum**	**Ammonium chloride**	**Hydrogen Fluoride**

- **Note:**
  - **Ionic compounds dissolve in polar solvent (H2O).**

- **Polar compounds as HCL dissolve in polar and non polar solvents.**

Questions

(1) Define each of the following:

1- Ionic bond.
2- Co-ordinate bond.
3- Chemical reaction.
4- Hybridization.
5-Covalent bond.
6-H2-bond.
7-Octet rule.

(2) Give reason:

1. Octet rule theory can't be applied for PCL5 or BF3.
2. The pi-bond is weaker than sigma bond.
3. HCL solution in water conducts electricity but its solution in benzene does not conduct electricity.
4. Ionic compounds have higher melting point than covalent compounds.
5. NaCl molten is a good conductor of electricity while AlCl3 solution is a poor conductor of electricity.
6. H2O has smaller molecular weight thas H2S gas, but it exists in a liquid state and has high boiling point.
7. Elements of 1st transition series are hard metals except 29Cu is relatively soft.
8. In ethylene, hybridization of carbon orbitals is SP2, while in acetylene is of SP.
9. Metals are good conductors of electricity.
10. Both fluoride (negative ion) and sodium (positive ion) have the same number of electrons.

(3) Write the scientific name:

1- A bond is formed by the overlapping of orbitals side by side.
2- A bond is formed by the overlapping of orbitals head to head.
3- An ion formed by combination between water and proton.
4- The bond in which the pair of electrons arises from one atom.
5- The bond which is formed between two atoms having the same electronegativity.
6- The bond which is formed between two atoms where the difference in electronegativity is higher than 1.7.

(4) Compare between:

1. Sigma and Pi Bonds.

2. Covalent and ionic compounds.
3. Valency bond theory and molecular orbital theory.
4. Methane – ethylene and acetylene according to number of hybridized orbitals, kind of hybridizayion, stereo structure, angles between bonds.

- Write defects of electronic theory of valency.
- Explain the formation of H2, HF and NH3 molecules by using valency bond theory.

(5) Explain types of bonds in:
- Hydrogen chloride HCL. - Methane. - Acetylene.
- Ethylene. - Amonia NH3. - Ammonium hydroxide.
- Chlorine gas Cl2. - Sodium chloride. - Water.

(6) The following elements: A1 , B11 , C17

1- Show how can you get:

a) Ionic compound. B) polar covalent compound.

c) pure covalent compound.

2- What will happen by combination between A&C then dissolving the result in water and put a litmus paper?

(7) Choose the correct answer from the following:

1. For the elements 9A, 10 B, 11C, which of the following can happen?

a) B combines with C.	b) B can react with itself.
c) A combines with B.	d) A reacts with C.

2. An elements has atomic number 9. If two atoms combine with each other, what is the type of the bond that will be found in the molecule?

a) Metallic.	b) Co-ordinate.
c) Ionic.	d) Covalent.

3. Covalent compounds are characterized by:

a) Weak intermolecular attractive forces.	b) Good electrical conductivity.
c) Polar bonds.	d) Solubility in polar solvents.

4. The bond in a hydrogen fluride molecule is a polar covalent bons because the two atoms differ in:

a) Their position in the periodic table.	b) Electronegativity.
c) Electron affinity.	d) Ionization potential.

5. Hybridized (sp) molecular orbitals will have:

a) Three orbital lobes.	b) Two orbital lobes.
c) A linear shape.	d) Two orbitals and a linear shape.

6. In the acetylene molecule we notice that:

a) Between the two carbon atoms there is a double bond, one is sigma the other is pi.
b) Between the two carbon atoms there is a triple bond, one is sigma and two are pi.
c) Each carbon atom is SP hybridized.
d) B and c are correct.

7. When sodium reacts with hydrogen, hydrogen becomes:

a) A positive ion.	b) An atom carrying a partial positive charge.
c) A negative ion.	d) An atom carrying a partial negative charge.

8. Two oxygen atoms combine to form a molecule, then:

a) Each atom shares an electron to give one covalent bond.
b) One atom gives a pair of electrons to the other atom.
c) Each atom shares two electrons.
d) Polar double covalent bond is formed between the two atoms.

Chapter 4

The main group elements of the periodic table

1- S – Block elements : elements of group I Alkali metals .
2- P – Block elements : elements of group V.A

First : elements of S – Block

Elements of (1A) group :
Elements of (1A) group are considered as alkali metals because their oxides dissolve in water easily forming strong Alkalis .
1- Lithium ⟶ Li ⟶ no using
2- Sodium ⟶ Na ⟶ Rock salt (NaCl)
3- Potassium ⟶ K ⟶ in sea water KCl and carnallite ($KClMgCl_2.6\ H_2O$) .
4- Rubidium ⟶ Rb ⟶ no using
5- Caesium ⟶ Cs ⟶ no using
6- Franciuim ⟶ Fr ⟶ Radioactive element it is produced from diseintigration of actinium

$${}_{89}Ac^{227} \longrightarrow {}_{87}Fr^{223} + {}_{2}He^{4}$$

General properties of elements of group 1A

1- Every element consists of one electron in the outer most energy level they are characterized by :
A- Every element lies in the beginning of new period .
B- Oxidation number in their compounds is equal (1+).
C- They are chemically very active due to the presence of one electron in the outer mast energy level which can by easily lost and they have very low ionization potential .
D- The first ionization energy low while second ionization energy is high because in the first ionization energy it is easy to lose the valencey electron but the second ionization energy result from the breaking up of a completely filled shell .

2- Most of their compounds are ionic: -
They can lose the electrons from their outer most energy level easily to form positive ions which have the same electronic structure of noble gas which preceds it .

3- They are very strong reducing agent because they have a large atomic radius (or volume) and small ionization energy so they lose the electrons from their outer most energy level easily .

4- They are most (soft) metals with low melting and boiling points due to the decreasing in the strength of the metalic bond between atoms since they have only one electron in the outer most energy level .

5- They have a large atomic radius because each element occupied the begining of its period .

6- **Elements of group (1A) are considered of the highest electropositive metals because they can easily lose the valency electron .**

7- Potassium and Caesium are used in photoelectric calls because the atoms of these elements have a large atomic radius and small ionization energy so when they are exposed to light they lose the electrons from their outer most energy level easily .

6- They have characteristic colours when the atom gains an amount of energy which is sufficient to transfer electrons to higher energy levels they give a characteristic colours : dry test

Element	Colour
Lithium	Crimson
Sodium	Golden yellow
Potassium	Pale violet
Calcium	Bluish violet

7- They are kept under liquid hydrocarbons .

Sodium is kept under kerosine because it is a very active metal which can react with air and water so it is stored under kerosine .

8- Action of atmospheric air :

All elements lose their metalic luster because they reacts easily with air to form metal oxide .

* Reaction with nitrogen of air to form (give) lithium nitride .

$$6\,Li + N_2 \longrightarrow 2\,Li_3N$$

G.R.F :

Lithium nitride is used a fertilizer ?

This is Because lithium nitride decomposes when the soil is irrigated giving ammonia (fertilizer) .

$$Li_3N + 3H_2O \longrightarrow NH_3 + 3LiOH$$

9- Reaction with water

$$2\,Na + 2H_2O \longrightarrow 2\,NaOH + H_2 + E$$

Sodium reacts with water forming sodium hydroxide and large amount of energy which is enough to cause the burning of hydrogen evolves with an explosion so sodium fires are not extinguished by water .

10- Reaction with oxygen :

$$4\,Li + O_2 \longrightarrow 2\,Li_2O$$

$$2\,Na + O_2 \longrightarrow Na_2O_2$$

$$2\,K + O_2 \longrightarrow 2\,KO_2$$

Potassium super oxide is used in submarines and aeroplanes in closed atmospheres because it reacts with exhaled carbon dioxide giving oxygen required for breathing :

$$4KO_2 + 2CO_2 \xrightarrow[\text{Cat.}]{CuCl_2} 2K_2CO_3 + O_2$$

11- Reaction with acides

$$2Na + 2HCl \longrightarrow 2NaCl + H_2$$

12- Reaction with hydrogen (to form hydrides)

$2Li + H_2 \longrightarrow 2LiH$

$2Na + H_2 \longrightarrow 2NaH$

$NaH \xrightarrow{\text{Electrolysis}} Na^+ + H^-$

Go towards cathode Go towards anode

N.B : Hydrides are ionic compounds because they produced from the reaction of element with hydrogen such as NaH , LiH .

13- Reaction with halogens :

All elements of group 1A are reacts with halogens forming very stable ionic halides .

$2Na + Cl_2 \longrightarrow 2NaCl$

$2K + Br_2 \longrightarrow 2KBr$

14- Reaction with other non – metal :

$2Na + S \longrightarrow Na_2S$ (Sodium sulphide)

$3K + P \longrightarrow K_3P$ (Potassium phosphate)

15- Action of heat on metal carbonates :

All alkali metals carbonates do not decompose when heated except Lithium carbonate .

$Li_2CO_3 \xrightarrow{\text{Heat}} Li_2O + CO_2$

16- Action of heat on metals nitrates :

They decompose partially giving metal nitrite and oxygen

$2NaNO_3 \xrightarrow{\Delta} 2NaNO_2 + O_2$

Sodium nitrate is not used in the manufacture of bombs because a great explosion happens when potassium nitrate decomposes by heat

$2KNO_3 \xrightarrow{\Delta} 2KNO_2 + O_2$

Extraction of metals

Alkali metals are not found in elemental state in nature because these metals are easily to losing their valence electron and oxidized in atmospheric air forming the oxide .

Elements of group (1-A) are extracted from their ores by electrolysis because they strongest reducing agent and can not be reduced from their ores by any reducing agents other than electrolysis .

Anhydride : Compounds which dissolve in water giving acid or alkali .

Commonly used sodium compounds

1- sodium hydroxide NaOH

a) **preparation in industry :**

by the electrolysis of sodium chloride solution

b)properties:

1- a white hygroscopic solid compound
2- it has a corrosive effect on skin
3- it dissolves easily in water forming an alkaline solution through an exothermic dissolution

3- it react with acids forming the sodium salt of the acid and water

$$NaOH + HCL \longrightarrow NaCL + H_2O$$

$$2NaOH + H_2SO_4 \longrightarrow Na_2SO_4 + 2H_2O$$

Uses :

1- NaOH used in many industries as : Soap , synthetic silk and paper
2- it used in purify petrol
3- detection of basic radicals (cations):-

detection of of copper II (Cu^{++})

salt solution + NaOH it gives a blue p.p.t turns black by heating

$$CuSO_4 + 2NaOH \longrightarrow Cu(OH)_2 + Na_2SO_4$$

$Na_2SO_4 \longrightarrow CuO + H_2O$

Detection of aluminium AL3+

Salt solution + NaOH gives a white p.p.t dissolves in excess of NaOH

$AlCl_3 + 3NaOH \longrightarrow 3NaCl + Al(OH)_3$

$Al(OH)_3 + NaOH \longrightarrow NaALO_2 + H_2O$

2- sodium carbonate Na_2CO_3

the hydrated salt $Na_2CO_3.10H_2O$ is known as washing soda

a) preparation :

1- **in laboratory : by passing CO2 gas through a hot solution of NaOh , the solution is left to cool , white crystal of Na2CO3 are separated**
2- **in industry : (solvay methode)**

$NH_3 + CO_2 + Nacl + H_2O \longrightarrow NaHCO_3 + NH_4Cl$

$2NaHCO_3 \longrightarrow Na_2CO_3 + CO_2 + H_2O$

Properties:

1- **white powder , easily dissolves in water . its solution has an alkaline effect**
2- **it is not affected by heat i.e it melts without decomposition**
3- **it react with acid , and CO2 evolves**

$Na_2CO_3 + 2HCl \longrightarrow 2Na_2CO_3 + CO_2 + H_2O$

Uses :

1- **paper industry**
2- **water softening**
3- **textile industry**

4- manufacture of glass

Elements of group (5A)

Nitrogen N_7 : non – metal – diatomic – gas in atmospheric air 80% .

Phosphor P_{15} : non – metal – Calcium phosphate ($Ca_3(PO_4)_2$) Apatite $CaF_2Ca_3(PO_4)_2$ (4 atom) .

Arsenic As_{33} : metalloid – Arsenic sulphide As_2S_3 – vapour (4atom As_4)

Antimony Sb_{51} : metalloid–Antimony sulphude Sb_2S_3 – vapour atoms Sb_4

Bismuth Bi_{83} : metal forming a crystal lattice – weak to conduct electricity – vapour (2atom)

General properties

1- **Oxidation number** : Elements of group [5 – A] have several oxidation numbers because they gain electrons from 1 to 3 through covalent sharing or electrons from 1 to 5 electron and reach to the stability state .

2- **With oxygen** : All elements of this group form oxides are acidic (decreases with increasing the atomic number) such as N_2O_3, N_2O_5, P_2O_3, P_2O_5 while other are amphoteric Sb_2O_3 or Bi_2O_3 or basic (increases with increasing the atomic number) Bi_2O_3 .

3- **With hydrogen** : Most of elements of this group reacts with hydrogen to form hydrides such as NH_3 , PH_3 , phosphene , Arsine AsH_3

These compounds (NH_3- PH_3) can form coordinate bonds due to presence of pair of electrons in valence shell so it can give this electrons to the outer atoms or ions to form coordinate bond

$NH_3 + H^+ \longrightarrow NH_4$, $PH_3 + H^+ \longrightarrow PH_4$

These compounds are basic because atom of element has one pair of electrons donated to positive proton of hydrogen which is found in the molecule of water therefore the negative hydroxyl group separated from molecule of water .

$NH_3 + H^+OH^- \longrightarrow NH_4^+ \quad OH^-$

- The polarity of hydrogen compounds in this group decreases with increasing atomic number .

- The thermally stability and the solubility in water are decreases with increasing the atomic in this group (NH_4^+) is more polarity than (PH_4^+) is more polarity than (AsH_4^+)

Allotropy

It is the presence of the element in more than one form having the same chemical properties but different physical properties .
Both nitrogen (gas) and bismuth (metal) have not allotropic .

Forms :

Solid non - metal	Allotropic forms
Phosphorus	white – red – violet.
Arsenic	black – grey – yellow.
Antimony	yellow – black.

Nitrogen N_2

Properties of nitrogen

1- Reaction of nitrogen with hydrogen

$$N_2 + 3H_2 \xrightarrow{\text{Electric spark}} 2NH_3$$

2- Reaction of nitrogen with oxygen

$$N_2 + O_2 \xrightarrow{\text{Electric arc.}} 2NO$$

The colour of nitric oxide (colourless) turns brown when it is exposed to atmospheric because nitric oxide is oxidized to form nitrogen dioxide when it exposed to air .

$$2NO + N_2 \longrightarrow 2NO_2$$

3- Reaction of nitrogen with metals

$$3M + N_2 \xrightarrow{\Delta} Mg_3N_2$$

magnesium nitride

$$Mg_3N_2 + 6H_2O \longrightarrow 2NH_3 + 3Mg(OH)_2$$

4- Reaction of nitrogen with calcium carbide ($CaCN_2$) to form calcium cyanamide ($CaCN_2$) is used as agricultural fertilizer because it reacts with water irrigating to from ammonia gas fertilizer .

$$CaC_2 + N_2 \longrightarrow CaCN_2 + C$$

$$CaCN_2 + 3H_2O \longrightarrow CaCO_3 + 2NH_3$$

Important nitrogen compounds

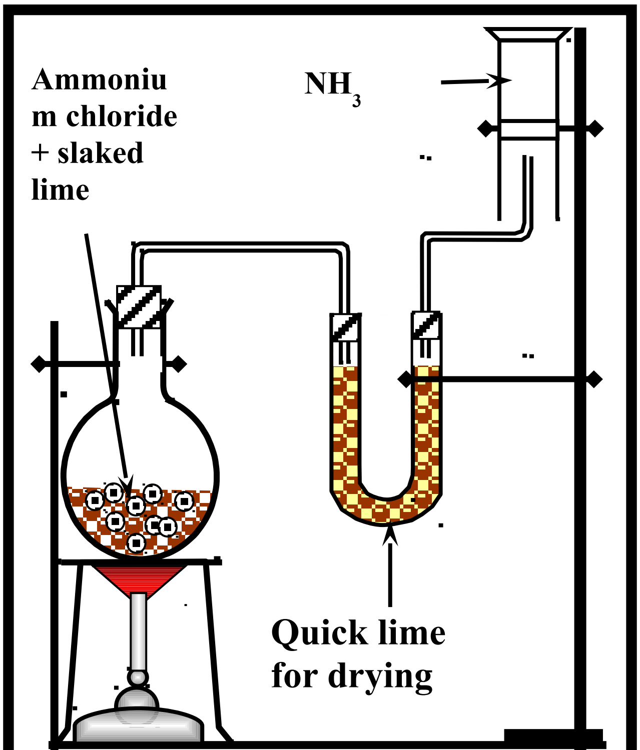
1- Ammonia gas (NH_3)

Preparation ammonia gas in lab

By heating a mixture of ammonium chloride and salked lime ($Ca(OH)_2$)

$$2NH_4Cl + Ca(OH)_2 \longrightarrow CaCl_2 + 2NH_4OH$$

$$2NH_4OH \xrightarrow{Heat} 2NH_3 + 2H_2O$$



Preparation og ammonia in lab

Ammonia gas is dried by passing it in quick lime (CaO) because quick lime dose not react with ammonia gas conc. H_2SO_4 in not used for dring ammonia gas because it reacts with acid forming $(NH_4)_2SO_4$ due to the basic property of ammonia .

Ammonia gas is collected by down – word displacement of air because it is lighter than air or density of NH_3 is less than air .

Properties of NH_3 gas

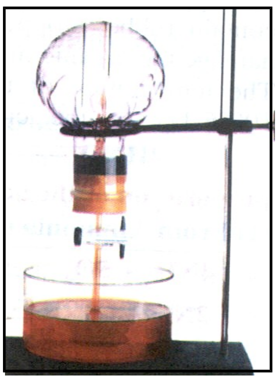
1- It is colourless and pungent smell .

2- It is easily soluble in water to from NH_4OH which turns the red litmus solution into blue .

$$NH_3 + H_2O \longrightarrow NH_4OH$$

***Experiment to prove that NH_3 gas is soluble in water and its solution has alkaline effect .**

1- Setup the apparatus as show in figure the lower bottle contains litmus


(The fountain experiment)

<u>G.R.F :</u> Ammonia is considered anhydride base ?

<u>Preparation of ammonia gas in industry (Haber's method)</u>

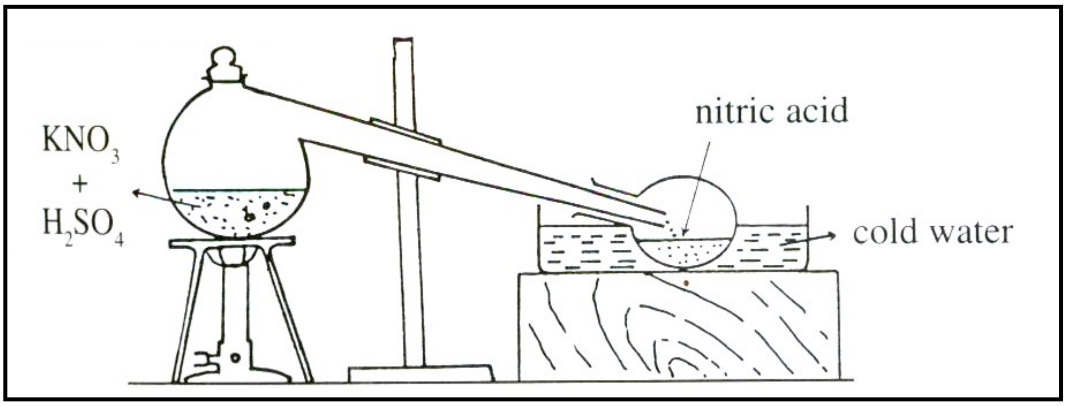
From nitrogen and hydrogen in presence of catalyst (iron) at 500ºC under 200 atmospheric pressure .

$N_2 + 3H_2 \longrightarrow 2NH_3$

<u>2- Nitric acid HNO₃</u>

1- preparation of nitric acid in lab

$2KNO_3 + H_2SO_4 \xrightarrow[\text{Conc}]{\text{Heat}} K_2SO_4 + 2HNO_3$



The apparatus for preparation of nitric acid does not contain rubber stopper because the vapours of nitric acid damage the organic materials as rubber .
The temperature of exp. dose not exceed more than 100ºC because the acid is decomposed thermally .

2- Preparation of the acid in industry

<u>Properties of acid :</u>

1- <u>Action of heat</u> :
It decomposed by heat giving nitrogen dioxide (NO_2), oxygen and water

$4HNO_3 \longrightarrow 4NO_2 + O_2 + 2H_2O$

2- Nitric acid is an oxidizing agent because it is reduced in to different product depends on :

a- The activity of reducing agent (the metal) .
b- The presence of some impurities in the metal .
c- Concentration of the acid .
d- Temperature of reaction :

- <u>Metals above hydrogen .</u>

$$2Fe + 8\cancel{HNO_3} \xrightarrow{Dil} 2Fe(NO_3)_3 + 4H_2O + 2NO$$

<u>- Metals below hydrogen in chemical series .</u>

$$3Cu + 8HNO_3 \xrightarrow{Dil} 3Cu(NO_3)_2 + 4H_2O + 2NO$$

$$Cu + 4HNO_3 \xrightarrow{Conc} Cu(NO_3)_2 + 2H_2O + 2NO_2$$

Copper reacts with nitric acid although it is below hydrogen in the electro chemical series because the acid reacts with copper as oxidizing agent i.e. it oxidize the copper to copper oxide which reacts with acid .
The above reaction is used to differentiate between dil. and canc. HNO_3 :

Experiment	Dilute HNO_3	Conc. HNO_3
1- put a piece of copper to each of them .	**Nitric oxide gas colourless is formed that turns in to nitrogen dioxide gas . $Cu + HNO_3 \xrightarrow{Dil} Cu(NO_3) + 2NO + 2H_2O$**	**Nitrogen dioxide gas (reddish brown fumes) are formed . $Cu + 4HNO_3 \xrightarrow{Conc} Cu(NO_3)_2 + 2NO_2 + 2H_2O$**

<u>The passivating effect</u> : Some metals (such as Fe – er – Al) are not affected by the concentrated nitric acid (HNO_3) due to the formation of layer of the metal oxide and this layer is non porous so it protects the metal from further reaction .

Economic importance of 5th group elements

Nitrogen : The manufacture of Ammonia – nitric acid – nitrogenous fertilizers .

Phosphorus : The manufacture of matches , rat- poison , several military industries , phosphorus fertilizers , many alloys such as phosphorus bronze (Cu + Sn + P) and incendiary bombs .

Antimony : It is used with lead in accumulators antimony sulphide (Sb_2O_3) is used for dying .
Bismuth : Alloys of bismuth , lead cadmium and tin are characteriesed by their low melting point .

Questions

Show by symbolic equations only:

1. The reaction between nitrogen and lithium, then adding water.
2. Passing of CO2 gas on a hot solution of NaOH.
3. Adding dilute cold sulphuric acid on sodium peroxide.
4. Exposing a piece of sodium in air for a long time.
5. Reaction between lithium and hydrogen, then adding water.
6. Adding water to sodium hydride, sodium oxide, sodium peroxide and potassium superoxide.
7. Heating lithium carbonate and sodium nitrate.
8. Heating a mixture of ammonium chloride and sodium nitrate.
9. Heating of magnesium with nitrogen, then adding water.
10. Reaction of calcium carbide and nitrogen in presence of electric arc, then adding water.
11. Heating a mixture of ammonium chloride and calcium hydroxide, then dissolve the resulting gas in water. .
12. Reaction of nitric acid with Cu, Fe.
13. Dissolving calcium Cyanamid in water.
14. Effect of heat on nitric acid.

How can you get each of the following:

1. Oxygen from sodium nitrate.
2. Ammonia gas from lithium. .
3. Oxygen from potassium.
4. Nitrogen dioxide from conc. Nitric acid.
5. Ammonia gas from nitrogen.
6. Nitrogen from sodium nitrite.
7. Ammonia from calcium carbide.
8. Ammonia from magnesium.
9. Caustic soda in industry.
10. Washing soda:
    i- In industry.
    ii- In lab.

- From nitric acid how can you get Nitric oxide, Nitrogen dioxide, Hydrogen.
- From conc. Sulphuric acid, quick lime, water, sodium nitrate and ammonium chloride how can you get nitrogen, ammonia, oxygen, nitric acid?
- In lab. With drawing, how to prepare Nitrogen, Ammonia gas, Nitric acid?

Give reason for:

1. The Chemical reactivity of the alkali metals.
2. Potassium superoxide is used submarines and high altitude aeroplanes.
3. We must not extinguish sodium fires with water.
4. Caesium metal is used in photoelectric cells.
5. The weak strength of the metallic bond between atoms of 1^{st} group metals.
6. Ammonium ion is more basic than phosphonium ion.
7. There are various oxidation numbers for nitrogen.
8. A co-ordinate covalent bond in ammonium ion is formed.
9. Calcium Cyanamid is used as a fertilizer.
10. Sodium is kept under kerosene.
11. Sodium is not kept under water.
12. A white layer is formed on the surface of a piece of sodium when is left exposed to air for a long time.
13. Potassium nitrate is used in production of bombs while sodium nitrate is not used.
14. 1^{st} group elements have oxidation number (+1) and are found at the beginning af a new period.
15. 1^{st} group elements are strong reducing agents.
16. 1^{st} group elements are soft metals.

17. 1^{st} group elements form ionic compounds.
18. Sodium metal is soft magnesium is mild while aluminum is hard.
19. Although copper is less reactive than hydrogen but it reacts with conc. Nitric acid.
20. Calcium Cyanamid is used in production of ammonia.
21. Ammonia gas is not collected by down displacement of water but it is collected by down displacement of air.
22. Ammonia gas is not dried by conc. Sulphuric acid.
23. 1^{st} group elements hydrides are strong reducing agents.
24. Iron does not react with conc. HNO3
25. Aluminum vessels are used to keep conc. HNO3.
26. Oxidation number of nitrogen is (+ ve) in oxygenated compounds , but its oxidation number is (- ve) in hydrogenated compounds .
27. Gp I a separated by electrolysis of their molten halides not their molten halides not their solution .
28. Gp I a are extracted by electrolysis not by reduction of their oxides .
29. Ammonia gas is a reducing agent .
30. Nitrogen reacts only under high temp .
31. Its difficult to extract alkali metals by ch . methods .
32. Urea is preferred in hot regions .
33. Con H C I in used to detect presence of ammonia .
34. Alloys of bi , lead , cadmium , and tin preferred making fuses .

<u>Define :</u>

1- Allotropy .
2- Anhydride .

How to differentiate between :

1- Copper sulphate and aluminum sulphate
2- Sodium nitrate and Nitrate
3- Conc . and dilute H N O 3 .

<u>How to identify :</u>

1- Copper in its salt solution .
2- Aluminum in its salt solution .
3- Nitrate ion
4- Nitrate ion

5- Ammonia gas

Give one use for :

1- Cs
2- Potassium Nitrate
3- Phosphorus
4- Caustic soda
5- Washing soda
6- Platinium wire

Detect which ox / red agent by equation :

1- Nitrate acid
2- Sodium peroxide
3- Potassium super oxide
4- Sodium hydride
5- Ammonia
6- Sodium nitrate

Chapter 5
Transition elements

- The electronic configuration and oxidation states.
- General properties of the transition elements.
- The iron metal.
- Extraction of iron from its ores.
- Production of iron.
- The alloys.
- The properties of iron.

Group	3B	4B	5B	6B	7B	8			1B	2B
Element	$_{21}Sc$	$_{22}Ti$	$_{23}V$	$_{24}Cr$	$_{25}Mn$	$_{26}Fe$	$_{27}Co$	$_{28}Ni$	$_{29}Cu$	$_{30}Zn$

The main transition elements are those of (d) block , they consists of three horizontal series which are located in the periods 4 , 5 and 6 .

1st transition series $Sc_{21} \longrightarrow Zn_{30}$
2nd transition series $Y_{39} \longrightarrow Cd_{48}$
3rd transition series $La_{57} \longrightarrow Hg_{80}$

Economic importance of elements of the first transition series

The element	Symbol	Economic importance (uses)
Titanium	Ti_{22}	Manufacture of rockets and super sonic aircrafts
Vanadium	V_{23}	Manufacture of steel Manufacture of car springs V_2O_5 is used as a catalyst in production of sulphuric acid (H_2SO_4)
Chromium	Cr_{24}	It is used in plating metals Manufacture of stainless steel Manufacture of heating coils Cr_2O_3 is used in making pigments $K_2Cr_2O_7$ is used as an oxidizing agent Potassium chromium is used as mordant dyes

		in clothes
Manganese	Mn_{25}	It is used in the production of steel $KMnO_4$ used as oxidizing agent as an antiseptic
Iron	Fe_{26}	Manufacture of cars , ships and tools
Cobalt	Co_{27}	It is used in forming several alloys Cobalt compounds are used in colouring glass $CoCl_2$ is used in secret ink and atmospheric predictions It is used in malignancy diagnosis It is used in treatment detect creak and photographing It is used in studying the humidity of the soil
Nickel	Ni_{28}	It is used in plating metals It is used in forming several alloys It is used as a catalyst in hydrogenation of oils It is used in manufacture of vessel which are used to store HF and NaOH
Copper	Cu_{29}	It is used in the manufacture of electric wires Manufacture of many alloys The anhydrous $CuSO_4$ is used in the detection of water

G.R.F :

1- Scandium has limited uses ?

Because it is not an abundant element in the earth's crust.

2- Titanium is preferred in the manufacture of rockets and super sonic aircrafts ?

Because it can keep strength at high temperature .

3- Vanadium is used in the manufacture of steel ?

Because it gives the steel hardness and corrosion resistance also it is used in manufacture of car springs because it has a corrosion resistance .

4- Chromium is highly reactive but it resists the effect of atmosphere ?

Due to the formation of non – porous and oxide layer on its surface which prevents further reaction with oxygen in the air .

5- Chromium is used to plate other metals to resist corrosion and make them more attractive .

6- Manganese is used in production of steel ?
Because it reacts with oxygen during production process and prevents the formation of bubbles in steel during cooling .

7- Cobalt II Chloride is used in manufacture of secret ink ?
Because aqueous solution of cobalt II chloride has a pale pink colour which does not appear in writing then changing to deep blue on dehydration by heating therefore the writing appears .

$$CoCl_2.6H_2O \xrightarrow{\Delta} CoCl_2 + 6\,H_2O$$

Pale pink colour deep blue colour

8- Anhydrous cobalt II chloride is used in atmospheric predictions ?
Because certain papers are converted with blue chloride when turned to pale pink this indicate that the increase in reactive humidity in air and possibility of raining .

$$CoCl_2 + 6H_2O \longrightarrow CoCl_2\,.6H_2O$$

9- Nickel is used in plating metals to protect them from oxidation and corrosion and to give them coating .

10- Caustic soda and liquid hydrogen fluoride are kept (stored) in containers made of nickel ?
Because of its resistance to corrosion and it is not affected by them .

11- Cr_{24} and Cu_{29} are abnormal from except electronic configuration ?
In case of chromium (Cr_{24}) Cr_{24} : (Ar_{18}) $4S^1$ $3d^5$ so these sublevels 4S and 3d becomes half filled this makes the atom has less energy and more stability .
In case of Copper (Cu_{29}) Cu_{29} : (Ar_{18}) $4S^1$ $3d^{10}$ so the sublevels 4S becomes half filled and sublevels 3d become completely filled this makes The atom has less energy and more stability

Question : Explain why iron II is easily oxidized to iron III (where as) while Mn II in not easily oxidized to Mn III ?

Answer :

The electronic configurtionof iron atom is : $Fe_{26} = (Ar_{18})\ 4S^2\ 4d^6$

$$Fe^{++}\ 3d^6 \xrightarrow{\text{Oxidation}} Fe^{+++}$$

Iron III ions is more stable as the 3d sublevel is half filled it is more stable and the oxidation of iron (II) is facilitated but in the case of manganese atom the electronic configuration is :

$Mn_{25} = (Ar_{18})\ 4s^2\ 3d^5$

$$Mn^{++}\ 3d^5 \xrightarrow{\text{Oxidation}} Mn^{+++}\ 3d^4$$

Half- filled more stable **less stable**

12- The element of the first transition series loses its 4S electrons before losing the 3d electrons ?

Because the 4S electrons has energy less than the energy of 3d .

13- The transition elements of the first series gives a (+2) oxidation state ?

Because when they gain energy they will lose two electrons from the sublevel 4S at first .

14- Scandium (Sc_{21}) can not give oxidation state (+2) but it gives oxidation state (+3) only ?

Because when the atom gains energy it loses two electrons from the sublevel 4S then one electron from 3d sublevel to be more stable .

15- Transition elements are characterized by having variable oxidation states ?

Because the two sublevels 4S and 3d of really equal energy and their electrons are lost in sequence when the atom is oxidized .

16- Scandium can not give oxidation state (+4) ?

Because the amount of energy needed to obtain this ion is greater due to the braking of a complete energy level .

General definition for transition element :

It is the element that in completely filled (d) or (f) sublevel in either the free or in one of its oxidation states .
The coinage metals whish are (Cu_{29} - Ag_{47} - Au_{79}) are considered as transition elements ?

Although the orbitals of (d) sublevel are completely filled with electron because in their higher oxidation state Cu^{++} , Ag^{++} , Au^{+++} contain (9) or (8) electrons in d sublevel .
Cu^{++} (Ar_{18}) $3d^9$ $4s^0$
Ag^{++} (kr_{18}) $4d^9$ $5s^0$
Au^{+++} (Xe_{54}) $5d^8$ 6s

Zn_{30} , Cd_{48} , Hg_{80} , are non transition elements because the orbitals of (d) sublevel are completely filled with electrons in their free states or in their higher oxidation states .
Zn^{++} (Ar_{18}) $4S^0$ $3d^{10}$
Cd^{++} (Kr_{36}) $5S^0$ $4d^{10}$
Hg^{++} (Xe_{54}) $5d^{10}$ $6S^0$

Transition elements (Fe , Co , Ni) are used in the manufacture of alloys ?
Because these elements have nearly equal atomic radius.

The general properties of the transition

1-The densities of transition elements increase with the increase of atomic number due to the increase of atomic mass while the atomic radius is nearly constant .

2- melting and boiling point

The element of the first transition series have high melting and boiling point because they have strong metallic bonds resulting from sharing of both the 4S and 3d electrons .

3- magnetic properties:-

3-Transition elements are paramagnetic due to the presence of unpaired electrons in the (d) or bitals .

Paramagnetic substance	Diamagnetic substance
It is the substance which is attracted to the magnetic field due to the presence of unpaired electrons in (d) orbitals .	It is the substance which is repelled (not attracted) to the magnetic field due to the presence of electron which are paired in all (d) orbitals .
The magnetic moment is equal to the number of unpaired electrons in (d) sublevel such as Cu^{++} , Fe Ni , Mn , Co , Sc	The magnetic moment is equal zero because electrons are paired in all orbitals such as zinc , sc^{+++}

Which of the following elements paramagnetic and which is diamagnetic
Zn , Cu^{++} , Fe^{++}
Write the order of momentum of these element .

Zn d^{10}
Number of unpaired electrons 0
Magnetic momentum 0
Diamagnetic

Cu^{++} d^{9}
Number of unpaired electrons 0
Magnetic momentum 1
Paramagnetic

Fe^{++} d^{6}
Number of unpaired electrons 4
Magnetic momentum 4
Paramagnetic

The colours : atoms or ions of transition elements are colour due to the presence of unpaired electrons in the orbitals of (d) sublevel .

<u>The complementary colour</u> : The colour that is not absorbed by the substance .

Absorbed colour	Complementary colour
Orange	Blue
Yellow	Purple
Red	Greenish blue
Green	Red puprple

When the white light falls on the substance it absorbs the amount of energy is sufficient to excited unpaird electrons in the sublevels therefore the substance appears whit complementary colour to absorbed colour .

Cu^{++} (Ar_{18}) $4S^0$ $3d^9$ due to the presence of unpaired electron in the orbitals of (d) sublevel when the light falls on the cu^{++} ions unpaired electron absorb amount of energy equal to the energy of the orange colour therefore unpaired electron can excite and jump to a higher energy level and it appear whit complementary colour which is blue .

Co^{++} ions has a pale pink colour $\cancel{Co}_{27} \longrightarrow$ co^{++} due to the presence of unpaired electrons in the orbitals of (d) sublevel when the light falls on the co^{++} ions some of these electrons absorb the amount of energy which is equal to the energy of the green colour therefore unpaired electrons can excite and jump to a higher energy level and it appears with complementary colour which is red .

Ions of non – transition elements (representative elements) are colourless because they are needed a large amount of energy higher than energy of the visible light to excite the electrons to higher energy level and orbitals of (d) sublevel are empty or completely filled with electrons .

The transition elements have catalytic activity due to the presence of unpaired electrons in the orbitals of (d) sublevel which can be used in formation of bonds between the reacting molecules and the atoms of the metals surface therefore the concentration of reactants increase on the catalyst surface so the speed of reaction increase .

Iron $_{26}Fe^{56}$

The ore	Chemical symbol	Scientific name	Colour
Magnetite	Fe_3O_4	Magnetic iron oxide	Black
Haematite	Fe_2O_3	Iron III oxide	Blood red
Limonite	$Fe_2O_5 . 3H_2O$	Hydrate iron II oxide	Yellow
Siderite	$FeCO_3$	Iron II carbonate	Yellowish grey

The extraction of iron from the ores :

The process of extraction include three stages :

1- Dressing of iron ore .

2- Reduction of iron ore .

3- Production on of iron .

4- Roasting process .

1- Crushing process : It's the process of converting the large size of the iron ore to small size to be easily reduced .

2- **Purification and concentration of the ore by using two methods** :

This process removes most of the impurities such as rocks by using one or more of several mechanical and physical processes

3**- Sintering process : A process of converted the fine particles of the iron ore to large particles to be easily reduced .**

4**- Roasting process : Heating the ores in air strongly to :**

A- Increase the percentage of iron in the ore .

$$FeCO_3 \xrightarrow{\Delta} FeO + CO_2$$

$$\xrightarrow{(O)} Fe2O3$$

42% Fe **69% Fe**

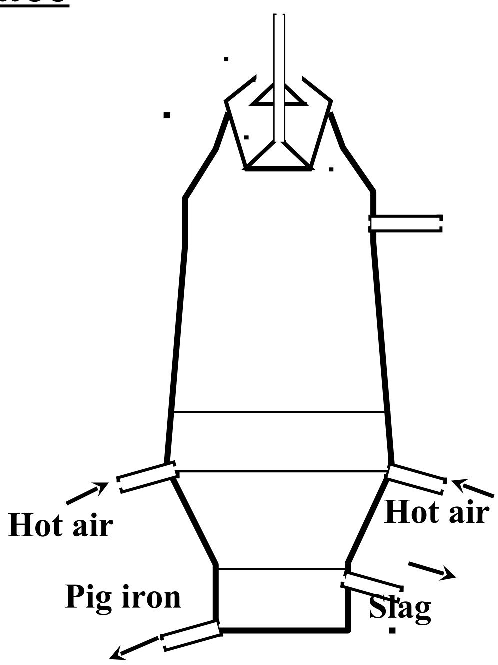
$$Fe_2O_3 . 3H_2O \xrightarrow{\Delta} 3H_2O + Fe_2O_3$$

60% Fe **69% Fe**

B- Oxidation of some impurities : such as (S , P)

$$S + O_2 \xrightarrow{\Delta} SO_2$$
$$4P + O_2 \xrightarrow{\Delta} 2P_2O_5$$

The blast furnace



1- The charge (load)
Iron III oxide Fe_2O_3 , Coke (C) , limestone ($CaCO_3$) .

2- Source of energy :
An air blast is supplied through nozzles at the bottom of the furnace .

3- Reducing agent :
(Co) carbon monoxide from coke the role of coke in blast furnace it is oxidized by hot air to CO_2 which is reduced by coke to produce co which reduces iron III to iron

$$C + O_2 \xrightarrow{Heat} CO_2$$
$$2CO_2 + C \xrightarrow{Heat} 2CO$$
$$3Fe_2O_3 + CO \xrightarrow{230\text{-}300^oC} 2\,Fe_3O_4 + CO_2$$
$$Fe_3O_4 + CO \xrightarrow{400\text{-}700^oC} 3FeO + CO_2$$
$$Feo + CO \xrightarrow{Above\ 700^oC} Fe + CO_2$$

The Role of lime stone ($CaCO_3$)

It decomposed by heat to calcium oxide and carbon dioxide

$$CaCO_3 \xrightarrow{Heat} CaO + CO_2 .$$

Calcium oxide reacts with acidic oxides

$$CaO + SO_2 \xrightarrow{\Delta} CaSiO_3$$

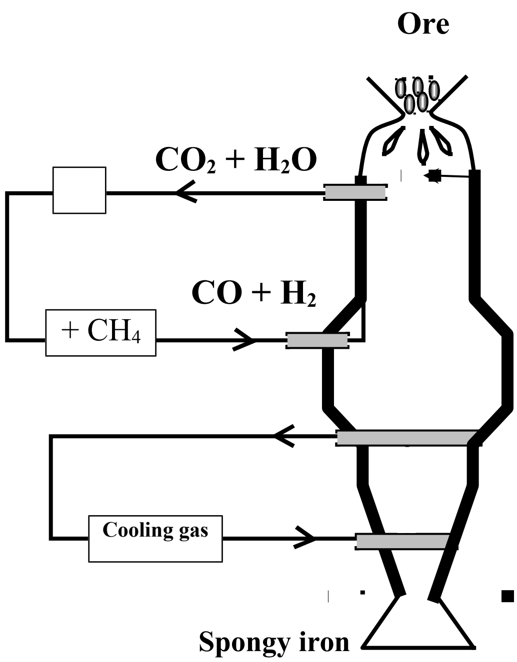
$$CaO + Al_2O_3 \xrightarrow{\Delta} Ca(AlO_2)_2$$

- The slage floats on the top of molten iron and protects it from oxidation by the current of air .

- The slage is used in brick industry , cement and asphalt .

- The type of iron produced : pig iron .
- It contains 95% iron , 4% carbon , 1% impurities .
- The company : Egyptian company in Helwan .

Midrex furnace



1- The charge (Heamtite) .
2- Reducing agent :
(mixture of CO + H2) From natural gas

The Role of natural gas in midrex

Furnace : It combine with CO_2 gas and H_2O vapour to form reducing agent

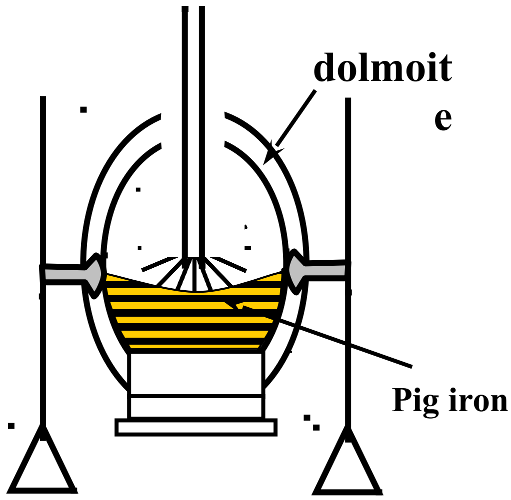
$$CH_4 + CO_2 + H_2O \xrightarrow{Cat.} 3CO + 5H_2$$

$$2Fe_2O_3 + 3CO + 3H_2 \longrightarrow 4Fe + CO_2 + 3H_2O$$

- The iron hammered to separate the impurities
- The type of iron produced : spongy iron
- The company : National company of steel in Dekhela Alexanderia .

Iron production

The oxygen converters
Production of steel by using oxygen converters



The oxygen converters

1- The charge : Molten of big iron to save energy required for melting iron if it is solidified .

2- The furnace is lined with dolmoite .
The Role of dolomite ($CaCO_3 + MgCO_3$) it decomposes by heat forming CO_2 and CaO , MgO (basic oxides which) react with acidic oxides of impurities to produce slage .

3- The oxidizing agent : Current of pure oxygen the rote of ferromanganese alloy (Fe , Mn , C) is added to prevent the formation of

gas bubbles in steel and increases the hardness of steel because manganese combines with the oxygen in steel .

4- Pure oxygen is preferred to air because air contains different gases specially nitrogen which does not help in oxidation of impurities .

Galvanization : Converting steel with zinc to protect it from rusting .

Alloys

It is a substance formed from two metals and sometimes formed from metals and other element in a certain ratio .

Methods of preparation of alloys :

1- Melt method :
The metals are melt together and then leave the metal to cool gradually .
2- Electro position method :
Two metals or more are deposited at the same time from solution on the surface other elements such as electro plating iron handles with brass (Cu + Zn) .

Types of alloys

1- Interstitial alloy .
2- Substitional alloy .
3- inter metalic alloy .

Interstitial alloy	Substitional alloy	Intermetalic alloy
It is formed when a metal introduces to the structure of another pure metal	It is formed when a some atom of the crystalline lattice of pure metal is replaced by the atoms of other metal added .	It is formed when two metals or more combine with each other chemical compounds with new properties
Such as separated carbon in iron carbon alloy	such as stainless steel – copper – gold alloy – iron – nickel alloy which they are distinguished by : a- The same crystalline structure .	Such as alloy of cementite Fe_3C which distinguished by : a- It is hardness (solid). b- It is valency which does not obey the low

	b- The same chemical properties . c- The same diameter .	of valency . c- It can not be formed by the metals of the same group in the periodic table .

<u>N.B.</u> : To differentiate between two alloys one of them contain separated carbon (interstitial alloy) and other contains combinated carbon and a bad odour is ppt. in the case of combinated carbon .

<u>Properties of iron</u>

<u>A) Physical properties :</u>

1- It is not hard , malleable and ductile .
2- Good electric conductor and heat .
3- Melt at 1500ºC .
4- It has a density 7.87 gm/Cm^3 .
5- It is a transition element .

1- Effect of air
a- Dry air .   b- Moist. air .   c- Hot air .

- Dry air has no effect .

- Moist. air it formed in iron dust ($Fe(OH)_3$)
Iron dust is consider as a corrosion layer because the moist. air covered the surface of iron with porous layer , water vapour can pass through it and cause the corrosion of iron from the inside .

- Hot air : $3Fe + 2O_2 \xrightarrow{\Delta} Fe_3O_4$

<u>2- Effect of water</u>
Iron does not react with cold water or hot water but when water vapour is passed on red hot iron , magnetic iron oxides is formed and hydrogen gas evolves .
$3Fe + 4H_2O \xrightarrow{\Delta} 2Fe_3O_4 + 4H_2$

<u>3- Reaction with acid</u>
- With dilute sulphuric acid
$Fe + H_2SO_4 \xrightarrow{dil} FeSO_4 + H_2$

- With dilute hydrochloric acid

$$Fe + 2HCl \xrightarrow{dil} FeCl_2 + H_2$$

- With conc. Sulphuric acid

$$Fe + H_2SO_4 \xrightarrow{conc} FeSO_4 + Fe_2(SO)_3 + SO_2 + H_2$$

- With conc. nitric acid

$$Fe + HNO_3 \xrightarrow{conc}$$ **No reaction . Why ?**

Due to formation of a thin non – porous layer of oxide which protects iron from further oxidation .

4- Reaction with non – metals

1- With chlorine (Cl_2)

$$2Fe + 3Cl_2 \xrightarrow{conc} 2FeCl_3$$ **Iron III chloride**

because chlorine is a strong oxidizing agent so it prevents the formation of iron II salt .

2- With sulphure

$$Fe + S \xrightarrow{\Delta} FeS$$

Because sulphure is reducing agent so it prevents the formation of iron III salt .

Oxides of iron

Iron II oxide FeO	Iron III oxides Fe_2O_3	Magnetic oxide Fe_3O_4
By Reducing of iron III oxide or magnetic iron oxide $Fe_2O_3 + H_2 \xrightarrow{\Delta}$ FeO + H_2O $Fe_3O_4 + H_2 \rightarrow$ 2FeO + H_2O $Fe_3O_4 + CO \xrightarrow{\Delta}$ FeO+ CO_2 by Heating iron II oxalate COO Fe $\xrightarrow{\Delta}$ CO_2+CO+FeO COO Iron II oxide is formed because CO is reducing agent it prevents the lormatgon of iron III oxide	By Heating iron II sulphate $FeSO_4$ green $\rightarrow$ Fe_2O_3 + SO_2 + SO_3 Red 2- by Heating iron III hydroxide $Fe(OH)_3 \rightarrow Fe_2O_3 + H_2O$ By oxide or magnetic iron oxide $FeO + O_2 \rightarrow Fe_2O_3$ $Fe_3O_4 + O_2 \rightarrow Fe_2O_3$	By passing hot air in red hot iron $Fe + O_2 \rightarrow Fe_3O_4$ $Fe_3O_4 + O_2 \rightarrow Fe_2O_3$ $Fe_3O_4 + H_2 \rightarrow FeO + H_2O$ Starting with Fe how can you obtain the oxides of iron ? By passing water vapour in red hot iron $Fe + 4H_2O \rightarrow Fe_3O_4 + CO_2$ By reducing of iron III oxide $Fe_2O_3+CO_2 \rightarrow Fe_3O_4+CO_2$

FeO	Fe2O3	Fe3O4
It is a black solid oxide	It is a blood red colour oxide	It is a black colour
Effect of heat $FeO+ O_2 \rightarrow Fe_2O_3$ $FeO + H_2SO_4 \rightarrow FeSO_4 + H_2O$	Effect of heat $Fe_2O_3 \rightarrow Fe_3O_4+ O_2$ $Fe_2O_3 + H_2SO_4$ $Fe_2O_3(SO_4)_3 + H_2O$	Effect of heat Fe_3O_4+ 1/2 $O_2 \rightarrow Fe_2O_3$ $Fe_3O_4 + H_2SO_4$ $Fe_2(SO_4)$ $FeSO_4 + H_2O$ Two types of iron salts because FeO, Fe_2O_3

Diagram shows the important reaction of iron

Blast furnace	Midrex furnace	Oxygen converter
1- Charge (load) $Fe_2O_3 + C + CaCO_3$	Charge [load] Fe_2O_3	Charge [load] molten pig iron
Reducing agent Co from coke	Reducing agent a mixture of $CO + H_2$ from natural gas	reducing agent . —
Oxidizing agent —	Oxidizing agent —	Oxidizing agent current pure oxygen
Type of iron produced pig iron	Type of iron produced pig iron	Type of iron produced steel iron

Detection of iron II and iron III cations:

1- iron II cation Fe^{2+}

1- adding sodium or ammonium hydroxide to iron II salt solution , a white green P.P.T of iron II hydroxide is formed

$$FeSO_4 + 2\,NaOH \longrightarrow Na_2SO_4 + Fe(OH)_2$$

2- iron III cation Fe^{3+}

adding sodium or ammonium hydroxide to iron III salt solution , areddish brown p.p.t of iron III hydroxide is formed

$$FeCl_3 + 3\,NH_4OH \longrightarrow 3NH_4Cl + Fe(OH)_3$$

Questions

1- Define the following:

a) Transition element.
b) Oxidation number.
c) Para and diamagnetism.
d) Pig iron.
e) Spongy iron.

2- Write on each of the following:

a) Iron ore-dessing.
b) Reduction of iron ores.
c) Steel manufacture by oxygen converter.
d) Alloys.

Choose:

Nickel	for weathering forecast.
Cobalt	for Hydrogenation of oil.
Chromium	for air crafts industry.
Titanium	for electroplating and dyes.
	For diagnosis and treatment of Cancer.

Nickel	rockets industry.
Manganese	production of vessels to keep HF liquid.
Titanium	production of steel on oxygen converter.
Vanadium	for electric wires

Copper	making secret ink.
Nickel	for making car springs.
Cobalt	making vessels to keep HCL liquid
	Making dry cells.

Show what will happen in each of the following with written symbolic equation:

1- Passing steam on red – hot iron then adding conc. HCL.
2- Passing Cl2 gas on red – hot iron then adding ammonia solution.
3- Heating iron (II) oxalate in absence of air then adding dilute HCl.

Adding excess dilute H2SO4 to iron then divide the solution into 2 parts.
I- heating the 1st part strongly.
ii- adding to the 2nd part blue litmus solution.

5 –Heating iron with conc H_2SO_4.

Give reason:

1. coinage metals (Au, Ag, Cu) are transision element.
2. Iron is a transition element.
3. Transition elements have nearly constant atomic sizes.
4. Scandium has only oxidation state (3+)
5. Transition elements have several oxidation state & catalytic activity.
6. Chromium is a reactive metal but it resits effect of air.
7. Titanium is preferred than AL in making rockets and aircrafts.
8. Cu and Cr have iregular E. configuration.
9. Metals are preferred to be used in the form of alloys.
10. On adding excess of HCL to iron, a black ppt is formed.
11. $Fe_2(SO_4)_3$ is a paramagnetic substance but $ZnSO_4$ is a diamagnetic substance.
12. Cobalt (II) chloride is used in producation of secret link.
13. Ferromanganese alloyis added in production of steel in oxygen converter.
14. Copper with gold forms a substitutional alloy.
15. on the reaction between dilute H_2SO_4 and Fe, $FeSO_4$ is formed.
16. on heating iron with Cl_2, iron (III) chloride is formed not $FeCl_2$.
17. On heating $FeSO_4$, a red ppt. Is formed.
18. On heating iron (II) oxalate in absence of air a black ppt. Is formed.
19. On heating iron (II) oxalate in absence of air iron (II) oxide is formed not iron (III) oxide.
20. Hydrated titanium (III) ion in ($Ti(H_2O_6)$ is purple while hydrated titanium (IV) ion in ($Ti(H_2O_6)$ is colourless.
21. The effective main T.E are only 27not 30.

Show how can you get each of the following:

1. The three iron oxides from iron.
2. Iron from (II) oxalate.
3. Iron (II) hydroxide from iron.
4. Iron (II) oxide from iron (II) oxalate.
5. Iron (II) chloride from iron (II) sulphate how to getiron (II) oxide, magnetic iron oxide.

Write the role of the following:

1. Lime stone in production of iron in blast furnace.
2. Coke in production of iron in blast furnace.
3. Ferromanganese alloy in oxygen converter.
4. The lining (dolomite) in oxygen converter.

5. Slag.
6. Natural gas in midrex furnace.
7. Oxygen in oxygenated converter.

Compare between:

(1) Blast furnace and midrex furnace according to:
i- charge
ii- reducing agent.
iii- kind of iron produced.

(2) pig and spongy iron.
(3) intermetallic, interstitial and substitutional alloy.

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How can you differ between:

1- Dilute and conc. H_2SO_4 by using iron.
2- Two alloys of iron and carbon.
3- iron I I and I I I sulphate .

With my best wishes

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