

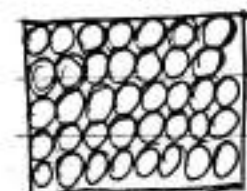
Atomic Theory

Appointment

In 1808 John Dalton put forward a theory and said → "Matter is made up of very small particles called atoms".

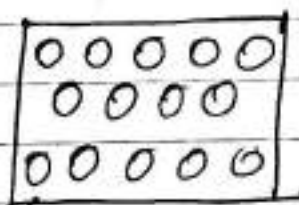
→ Matter is of 3 types → Solid
→ liquid
→ Gas

Arrangement of atoms in 3 types of matter :-

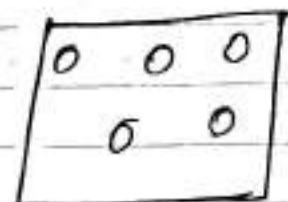


Solid

(close packing)



(Liquid)



(Gas)

→ Later as a result of large number of experiments conducted by Rutherford, Hasey, Bohr and others revealed that atom has a complex structure and atom can be further divided into small particles known as electron, proton and neutrons.

→ Electron :-

→ It is denoted by e^- .

→ Discovered by J.J. Thomson in cathode ray experiment

→ mass of $e^- \rightarrow 9.11 \times 10^{-31}$ kg

→ Charge of e^- is -ve and the value = -1.6×10^{-19} Coulomb

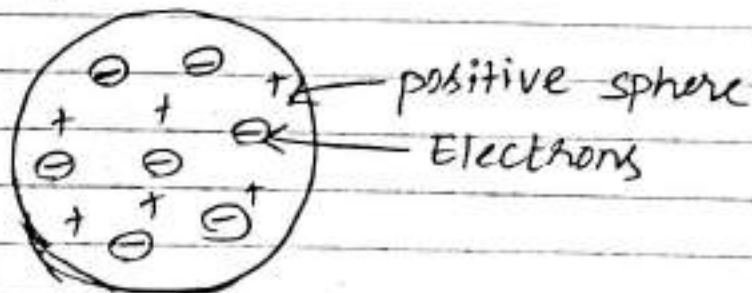
* Proton :-

- It is denoted by 'p'
- proton was discovered by E. Goldstein in cathode ray experiment.
- mass of proton = 1.672×10^{-27} Kg
- Charge of proton is +ve and the value is $+1.6 \times 10^{-19}$ - Coulomb.

* Thomson's atomic model :-

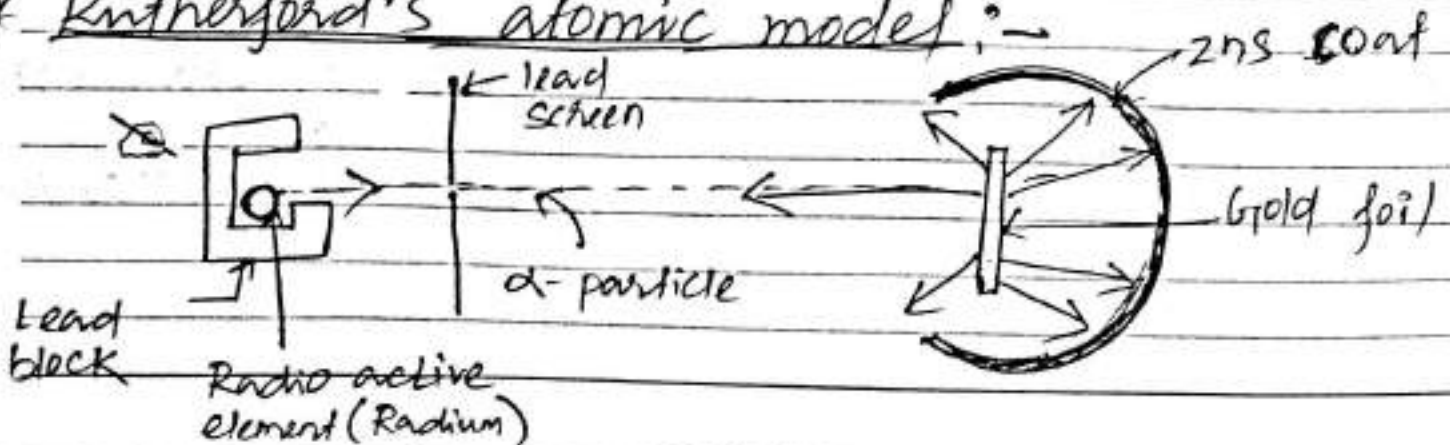
- The first atomic model was given by J.J Thomson in 1898 and is known as "plum pudding" or "raisin pudding" model

- According to him an atom consists of a sphere of positive electricity in which electrons are embedded like plum in puddings.



(Thomson's atomic model)

* Rutherford's atomic model :-



Q) In Rutherford's Gold foil Experiment, Rutherford bombarded a thin sheet of gold foil (thickness ≈ 0.0001 cm) with α -particles.

- α -particles are nothing but He^{2+} (positive charge) which are obtained from radioactive element Radium.
- A thin lead plate with a hole in it was there to form a beam of α -particles.
- A circular screen coated with ZnS (Zinc Sulphide) was placed on the other side of the foil.

Observation:-

Q) Write down the observation of Rutherford's Gold foil experiment.

- Ans - It was observed that most of the α -particles passed straight through the gold foil.
- very few α -particles were deflected at some angles after passing through the foil.
 - A very few i.e. one in 10,000 retraced their path.

* Conclusion:-

Q) What was the conclusion of Rutherford's Gold foil Experiment?

- Ans - From the experiment, it was concluded that
- Most of the space in an atom is empty.
 - There is a heavy +ve charge present at the centre of the atom which causes repulsion of α -particles.
 - Atom consists of two parts (i) Nucleus (ii) Extra nuclear part
 - Nucleus is a small positively charged centre carries almost whole mass of the atom.

Appointment

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→ The electron having negative charge is revolving around the nucleus in space.

→ The ^{circular} path in which the electron revolving ^{is} known ^{as} as orbits.

→ The centrifugal force and centrifugal force balances the revolution of electron around nucleus and prevents the electron to fall into the nucleus.

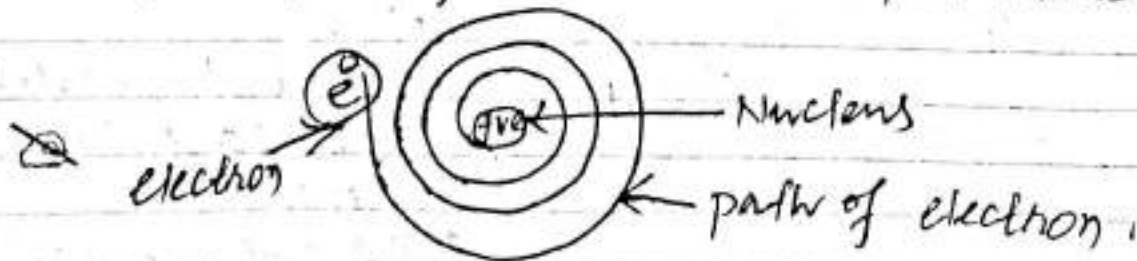
→ The atomic model given by Rutherford can be compared to Solar System.

* Drawback of Rutherford's atomic model :-

→ According to Clark Maxwell when a charged particle which moves under the influence of an attractive force, loses its energy continuously. As a result of which the e^- starts coming nearer to nucleus and the orbit of the e^- becomes coils like structure. However in real it doesn't occur. Rutherford could not explain the reason behind it.

→ He couldn't explain about the discontinuous hydrogen spectrum (Zeeman effect).

→ Further He couldn't explain the structure atom i.e. the distribution of electron around nucleus.



Bohr's model of the atom:—

Appointment

Neils Bohr adopted Rutherford's atomic model and Max Planck's quantum theory and proposed a theory known as Bohr's atomic model or "postulates of Bohr's atomic model".

Q) Write down the postulates of Bohr's atomic model.

Ans

→ According to Bohr an atom consists of a massive positively charged centre known as nucleus. The electron having -ve charge is moving around the nucleus in a fixed circular orbit without radiating energy. ~~is known as~~ The non-radiating orbit are known as stationary states.

→ Each of the fixed circular orbits are associated with some definite amount of energy and hence known as Energy level. The energy associated with each energy level goes on increasing with increase in distance from the nucleus.

→ The Energy level are named as K, L, M, N etc or numbered as 1, 2, 3, 4 - etc. The energy associated with each energy level are represented by the relation

$$E_n = \frac{-1312}{n^2} \text{ kJ mole}^{-1} \text{ (for H atom only)}$$

where n = no of shells (i.e. 1, 2, 3, ...)

Energy of $N > M > L > K$

→ Different Energy level are not equally spaced. It goes on decreasing with increase in distance of energy level from the nucleus.

→ Only those orbits are permitted whose angular momentum ($mvvr$) of an electron is a whole number multiple of $\frac{h}{2\pi}$

i.e. $mvvr = n \frac{h}{2\pi}$ Where m = mass of e^-

v = velocity of e^-

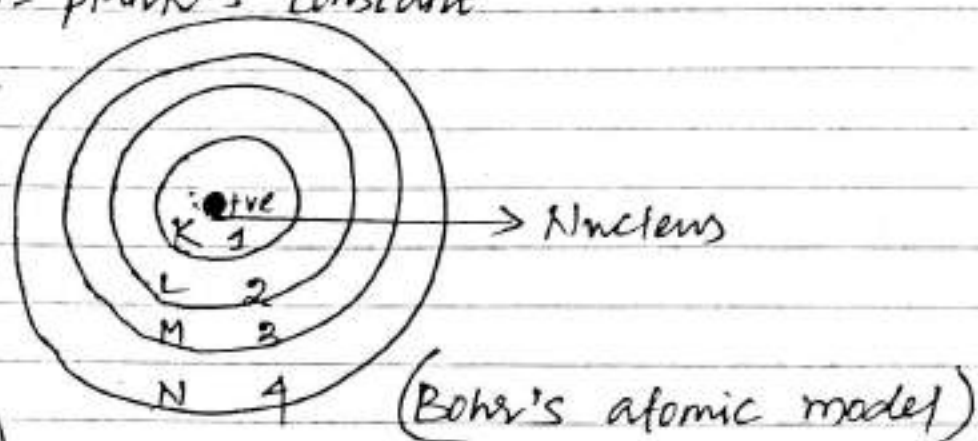
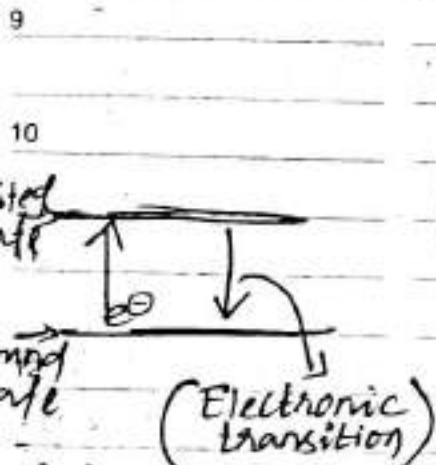
r = radius of energy level

$n =$

Saturday

Appointment

$n = \text{no of energy level}$
 $h = \text{plank's constant}$



→ An electron do not emit (lose) or absorb (except) Energy as long as it revolves in a particular orbit.

→ The electrons absorb or emit energy in the form of small packets known as quanta. When the electron absorbs energy it jumps from lower state to higher state. When it loses its energy it jumps from higher to lower state.

→ The lower energy level of electron is known as ground state. When the electron gets some energy from external source, ~~it~~ ^{it} get excited and jumps to higher energy level known as Excited state. However Excited state is not stable and hence the electron jumps down immediately in a fraction of second by losing energy in the form of light of suitable wavelength.

→ The movement of electron from one energy level to another is known as Electronic transition and the energy difference can be represented as $\Delta E = E_2 - E_1$

where as $\Delta E = \text{Change in Energy}$

$E_2 = \text{Energy of higher state}$

$E_1 = \text{Energy of lower state}$

→ According to Bohr's atomic model the position, velocity and momentum of any electron can be calculated accurately.

* Quantisation of Energy :-

Q// What do you mean by quantisation of Energy?
 → Quantisation means, the fixation of energy i.e. the quantity of energy can vary (diff.) only discontinuously to get a specific value.

Q// What is the basic difference between Rutherford's and Bohr's atomic model?

Ans - Rutherford's atomic model and Bohr's atomic model is ~~based on~~ basically differ by quantisation of energy, ~~transition~~ to Bohr's atomic model and angular momentum of electrons.

Drawbacks of Bohr's atomic model :-

- Bohr fails to explain the spectra of multi electron atoms (more than one electron)
- According to Bohr, the path of electron i.e. orbits are planar (2-Dimensional) however the electrons move around nucleus in space in 3-Dimensional space called orbitals.
- It fails to explain the cause of chemical combination in a molecule and the shape of molecules.
- According to Bohr's atomic model, the position and momentum of any electron can be calculated accurately. However, according to Heisenberg's uncertainty principle, it is impossible to measure the position and velocity of electron simultaneously.
- It can not explain the relative intensity of spectral lines.
- It fails to explain the splitting of spectral lines in a magnetic field.

Neutron:-

- Neutron was discovered by James Chadwick
 → mass of neutron - 1.675×10^{-27} Kg
 → specific charge \rightarrow zero (0)

Mass of an atom = No. of proton + No. of neutron

No. of neutron = Mass no - No. of proton

Q1) Calculate the no. of electron, proton and neutron of Na & Na⁺.

Ans - Na²³₁₁

no of e^- - 11

P - 11

$$\text{neutron} = \text{mass} - \text{no of P} \\ = 23 - 11 = 12$$

Na⁺

no of e^- = 11

P = 11

$$\text{neutron}(n) = 23 - 11 = 12$$

NOTE: any charge (+ve or -ve) doesn't effect the no. of 'p' & 'n'.

NOTE:-

→ +ve charge \rightarrow (Cation/Basic radical) - When an electron is removed from an atom it gives +ve charge on it.

→ -ve charge \rightarrow (Anion/Acid radical) - When an atom gains an electron it gives -ve charge on it.

Q1) Calculate the no. of e, p & n of.

Ca, Ca^{2+} , F, O, O^{2-}

Ans Ca^{40}_{20} no of e $\rightarrow 20$
 p $\rightarrow 20$
 n $\rightarrow 40 - 20 = 20$

Ca^{2+} - no of e $= 20 - 2 = 18$
 p $= 20$
 n $= 40 - 20 = 20$

F^{19}_9 \rightarrow no of e $\rightarrow 9$
 p $= 9$
 n $= 19 - 9 = 10$

O^{16}_8 \rightarrow e $= 8$
 p $= 8$
 n $= 16 - 8 = 8$

Ans - O^{2-} no of e $= 8 + 2 = 10$
 p $= 8$
 n $= 16 - 8 = 8$

* Bohr-Bury Scheme :-

\rightarrow According to Bohr-Bury Scheme, the maximum number of electrons present in a shell can be calculated by the relation - $2n^2$
 where n = no of shell.

No. of shell (n)	Maximum no of e ⁻
For K-Shell (n=1)	$2 \cdot n^2 = 2 \cdot 1^2 = 2 \cdot 1 = 2 \cdot e^-s$
\rightarrow L-Shell (n=2)	$2 \cdot 2^2 = 2 \cdot 4 = 8 \text{ e}^-s$
M-Shell (n=3)	$2 \cdot 3^2 = 2 \cdot 9 = 18 \text{ e}^-s$
N-Shell (n=4)	$2 \cdot 4^2 = 2 \cdot 16 = 32 \text{ e}^-s$

→ The outermost orbit of any element cannot contain more than 8 e⁻s and the orbit immediately before it i.e. penultimate orbit cannot contain more than 18 e⁻s.

→ It is not necessary to complete any orbit before the ~~filling~~ the next orbit starts filling.

Q// What is Isotope? Give an example of it.

Ans ¹² Isotopes:-

It is defined as the atoms of the same element which have same atomic number but different mass numbers.

→ They have same physical and chemical properties.

e.g-1 H_1^1 (Hydrogen), H_1^2 (Deuterium), H_1^3 (Tritium)

e.g-2 = Cl_{17}^{35} , Cl_{17}^{37} .

Q// What is Isobar? Give an example of it.

Ans- It is defined as the atoms of the ~~same~~ different elements having same mass number but different atomic numbers.

e.g- Ar_{18}^{40} , Ca_{20}^{40}

→ They have different physical and chemical properties.

Q// What is Isotone? Give an example of it?

Ans- It is defined as the atoms of different elements having same numbers of neutrons.

e.g- Ge_{32}^{76} , As_{33}^{74}

(∵ Number of neutrons = 44)

* Pauli's Exclusion principle :-

According to Pauli's Exclusion principle, no two electrons in an atom can have all the quantum numbers alike.

e.g. for two electrons from K-shell

$n=1, l=0, m=0, S=+1/2$

$n=1, l=0, m=0, S=-1/2$

(Same value) (Same value) (Same value) (Different value)

for $(K, n=1)$
 $l = n - 1$
 $= 1 - 1 = 0$
 $l = 0$

Q1) What is Auf-Bau principle? Give a brief description of it.

Ans - According to this principle, the electrons are filled in various orbitals in order of their increasing energies.

→ Therefore the orbital having lowest energy will be filled first.

→ The energy associated with the orbitals can be calculated by the $(n+l)$ rule.

where n = no. of principal quantum no. (Shell)

i.e. $K=1, L=2, M=3, N=4$

l = Azimuthal quantum no. (Subshell)

$S=0, P=1, d=2, f=3$

e.g. 4s subshell is filled first than 3d subshell as 4s is associated with less energy than 3d.

for 4s, $(n+l)$ $4+0$ $= 4$	for 3d, $(n+l)$ $3+2$ $= 5$	Therefore energy of 3d > 4s
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Friday

Appointment → When the $(n+l)$ value for both the subshell will equal, the subshell having lesser 'n' value will have lower energy and will be filled first.

10 e.g- $3p = n+l$ $4s = n+l$
 $= 3+1$ $= 4+0$
11 $= 4$ $= 4$

⇒ $(n+l)$ value of $3p$ is equal to $4s$. In this case $3p$ will have lower energy than $4s$ and will be filled first.

Q1) What is Hund's rule?

According to this rule, no electron pairing takes place in $p, d,$ and f subshell until each degenerate orbital contains one electron.

16 NOTE:-

The energy of

17 $1s < 2s < 2p < 3s < 3p < 4s < 3d < 4p < 5s < 4d < 5p < 6s$

18 (By applying $(n+l)$ rule)

19 Electronic configuration:-

20 The arrangement of electron in various orbital is known as electronic configuration.

Rule

$n = 1$ (K-shell), 2 (L-shell), 3 (M-shell), 4 (N-shell)

$l = 0$ (s-subshell), $l = 1$ (p-subshell), $l = 2$ (d-subshell)

$l = 3$ (f-subshell)

S - $\boxed{\uparrow\downarrow}$ - maximum 2 electrons

P - $\boxed{\uparrow\downarrow}\boxed{\uparrow\downarrow}\boxed{\uparrow\downarrow}$ maximum 6 electrons
 $P_x P_y P_z$

d - $\boxed{\uparrow\downarrow}\boxed{\uparrow\downarrow}\boxed{\uparrow\downarrow}\boxed{\uparrow\downarrow}$ - maximum 10 electrons
 $d_{xy} d_{yz} d_{zx} d_{x^2-y^2} d_{z^2}$

f - $\boxed{\uparrow\downarrow}\boxed{\uparrow\downarrow}\boxed{\uparrow\downarrow}\boxed{\uparrow\downarrow}\boxed{\uparrow\downarrow}\boxed{\uparrow\downarrow}$ ← maximum 14 electrons

NOTE:-

According to Azimuthal quantum number
 S-subshell - spherical in shape.
 P - dumbbell
 d - double dumbbell
 f - complex.

Electronic configuration:-

$H_1 - 1s^1$
 $He_2 - 1s^2$
 $Li_3 - 1s^2 2s^1$

$Be_4 - 1s^2 2s^2$

$B_5 - 1s^2 2s^2 2p^1$

$C_6 \rightarrow 1s^2 2s^2 2p^2$

$N_7 \rightarrow 1s^2 2s^2 2p^3$

$O_8 \rightarrow 1s^2 2s^2 2p^4$

$F_9 - 1s^2 2s^2 2p^5$

$Ne_{10} - 1s^2 2s^2 2p^6$

$Na_{11} \rightarrow 1s^2 2s^2 2p^6 3s^1$

$Mg_{12} \rightarrow 1s^2 2s^2 2p^6 3s^2$

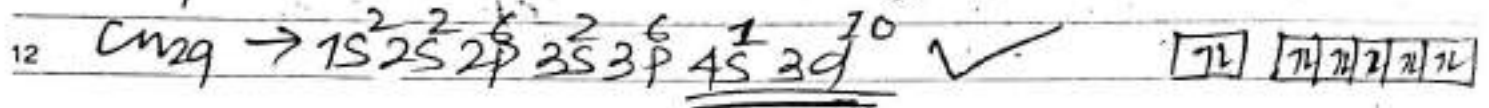
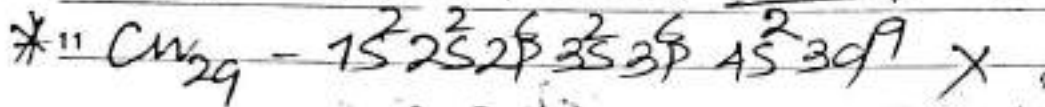
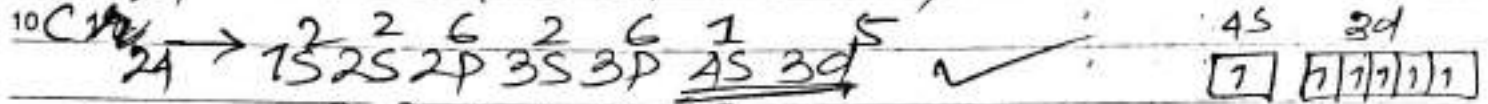
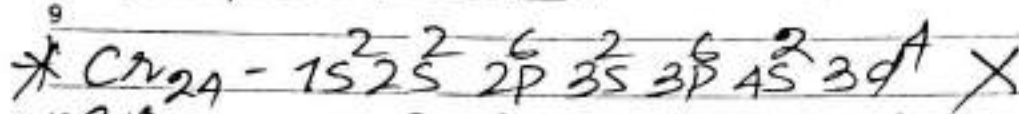
$Al_{13} \rightarrow 1s^2 2s^2 2p^6 3s^2 3p^1$

$Si_{14} \rightarrow 1s^2 2s^2 2p^6 3s^2 3p^2$

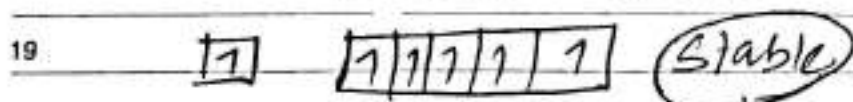
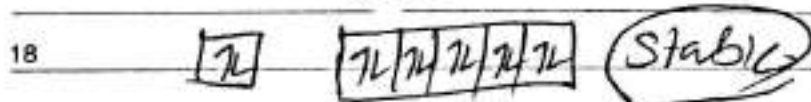
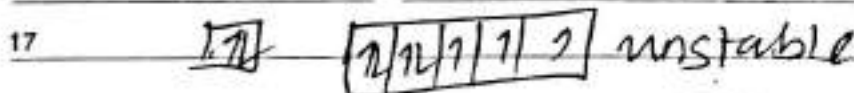
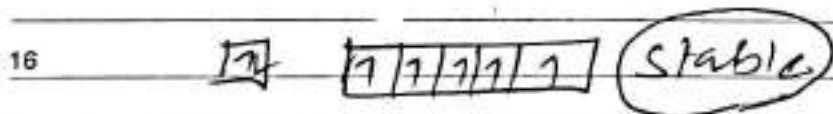
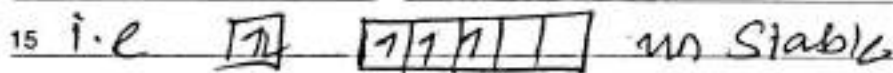
$P_{15} \rightarrow 1s^2 2s^2 2p^6 3s^2 3p^3$

SB on

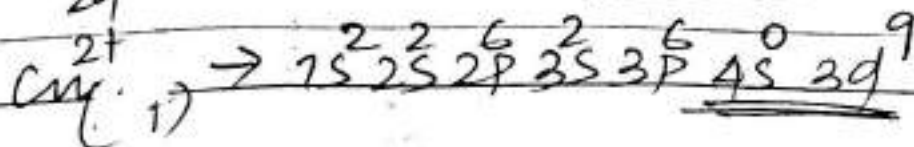
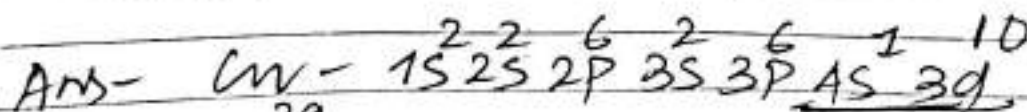
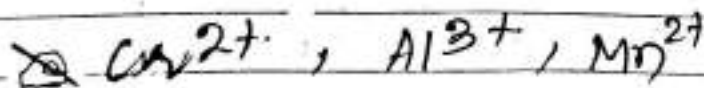
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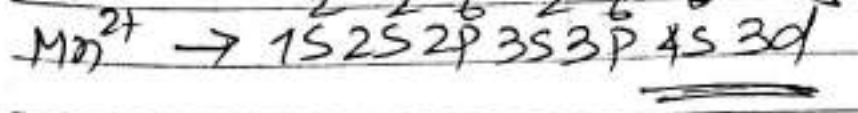
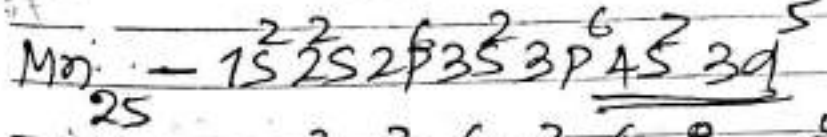
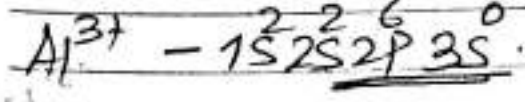
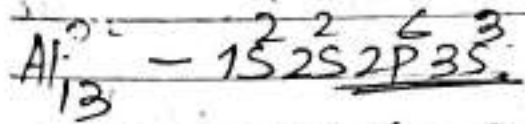
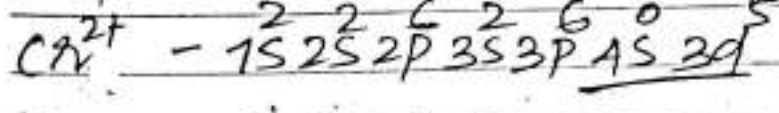
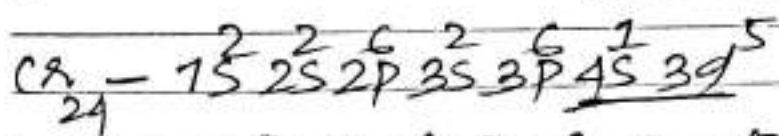
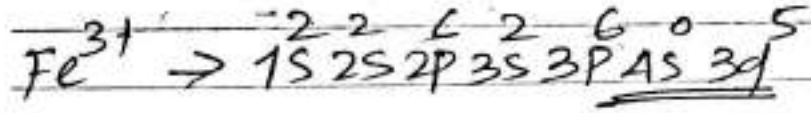
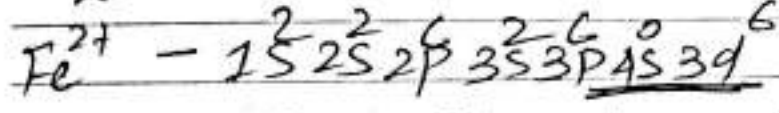
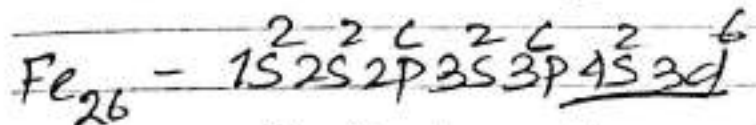
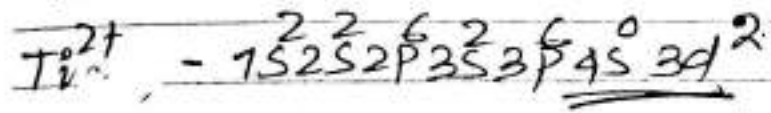
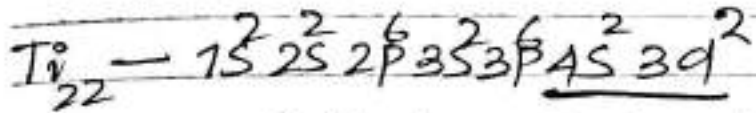
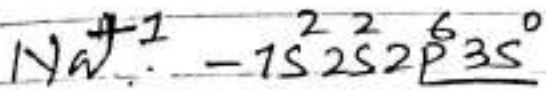
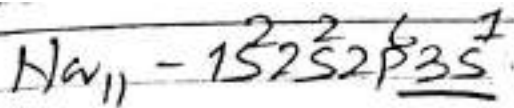


⇒ This is because the orbitals filled partially i.e. half filled and fully filled are more stable.



Q11 write down the electronic configuration of $Cu, Cu^{2+}, Na^+, Ti^{2+}, Fe^{2+}, Fe^{3+}$,





— 0 —

2

Probable Questions

9. long type 'a' (6 marks)

10. What are the observations and conclusions of Rutherford's Gold foil experiment.

2. Write down the postulates of Bohr's atomic model

2.12 Write down the drawbacks of Bohr's atomic model

A. Write down the drawbacks of Rutherford's Gold foil Experiment

(4 marks)

14. 3d subshell has greater energy than 4s. Explain briefly.

2. Write down the Bohr-Bury Scheme.

(2 marks)

17. What do you mean by (i) Isotope (ii) Isobar (iii) Isotone. Give an example of each.

2.18 Calculate the no of e^- , p & n of Ca^{2+} , Fe^{2+} , Na^+

3. When an electron gains energy, it jumps from _____ to _____ & when it loses energy, it jumps from _____ to _____

4.20 Write down the electronic configuration of Cr^{2+} , Al^{3+} , Cu^{2+} .

5. What is Pauli's exclusion principle?

6. What is (M) rule?

7. What is Hund's rule

Chemical Bonding

Appointment

9

Most of the elements in periodic table do not exist in free state. They always found in combined state. However few elements are there which do not combine with each other or with other elements. They always found in independent state.

e.g. He_2

Ne_{10}

Ar_{18}

Kr_{36}

Xe_{54}

Rn_{86}

→ These elements ^{are} found in 18th group.¹³
They are also known as zero group elements or inert gases or noble gases.

→ They are the most stable elements of periodic table.

16

Definition:-

Chemical bonding is defined as the force of attraction which holds together two or more atoms in a fixed ratio to form a molecule.

18

Q11 Why do atoms combine?

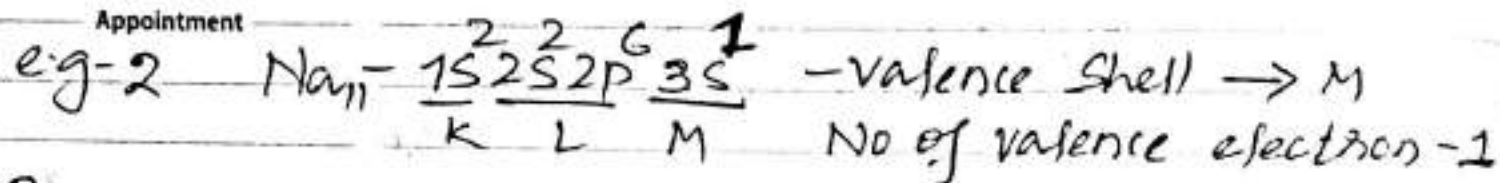
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→ In order to get stable by acquiring nearest inert gas configuration, the atoms combine to form molecule.

Valence Shell:-

The outermost shell of an atom is known as valence shell and the electrons present in it is known as valence electrons.

e.g. $\text{O}_8 - \underset{\text{K}}{1s^2} \underset{\text{L}}{2s^2} \underset{\text{M}}{2p^4} \therefore \rightarrow \text{L} (2s, 2p)$ valence shell
 \rightarrow no. of valence electrons - 6



Imp Valency:-

It is defined as the number of electron.

lose, gain or shared by an atom in order to acquire nearest inert gas configuration and get stable.

NOTE:-

When two atoms combine with each other,

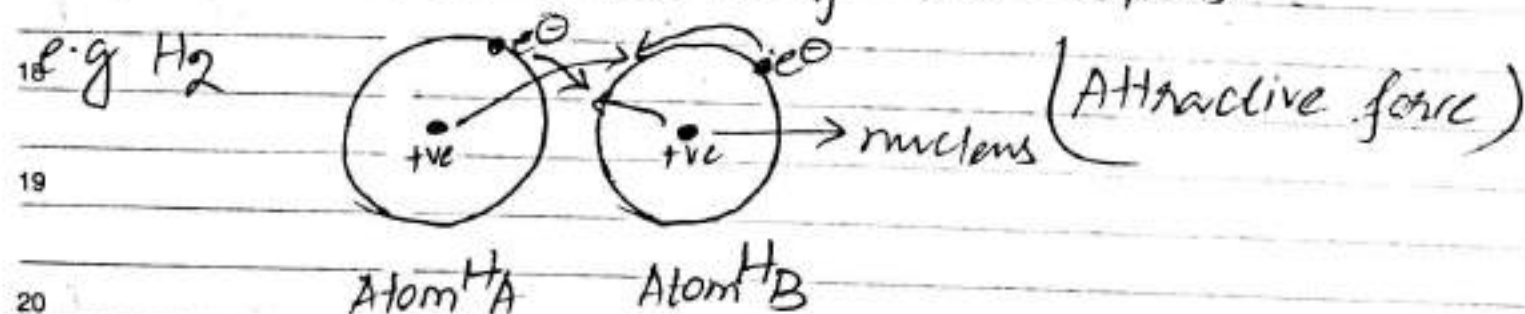
two types of force come into play.

(i) Attractive force

(ii) Repulsive force

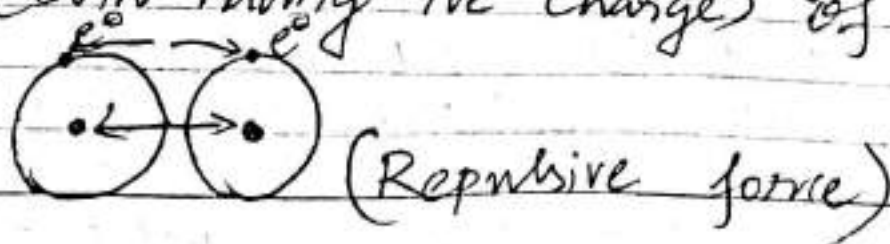
* Attractive force:-

Attractive force takes place between the electron and nucleus of two atoms.



* Repulsive force:-

Repulsive force takes place between two electrons (both having negative charge) and two nuclei (Both having +ve charge) of two atoms.



Types of Chemical Bonds :- :-

- There are three main types of chemical Bonds.
- (i) Ionic bond or Electrovalent bond 9
 - (ii) Covalent Bond 10
 - (iii) Co-ordinate bond or Dative bond. 11

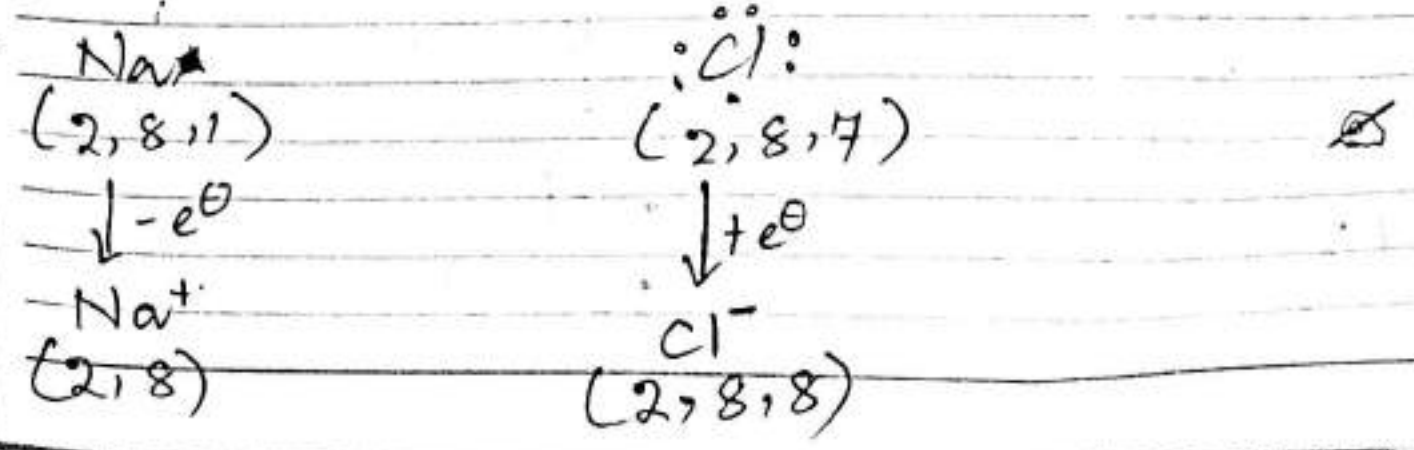
(i) Ionic Bond :-

The bond which is formed by the transfer of one or more electron from one atom to another or the bond which is formed by the loss and gain of electron between two atoms is known as Ionic bond and compound formed is known as Ionic compound. 14

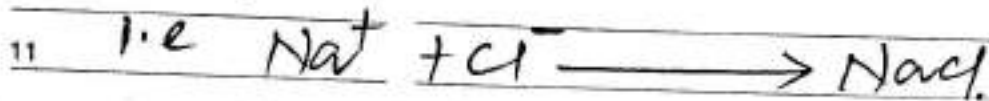
ex- Formation of Sodium chloride. (NaCl)



In NaCl, Na lose one electron to get nearest inert gas configuration (i.e. Ne) and form a cation Na⁺. Where as Cl gain one electron from Na to get nearest inert gas configuration i.e. (Ar) and form an anion. 20



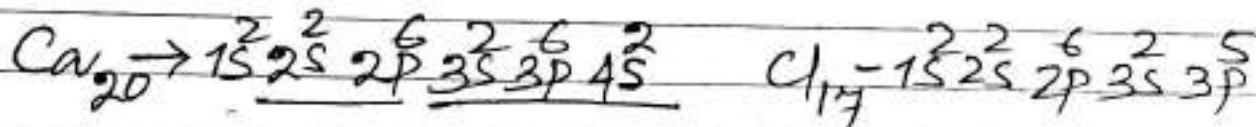
→ As the opposite charge attract each other, the Na^+ and Cl^- get attracted by electrostatic force of attraction by decrease of energy and form NaCl molecule.



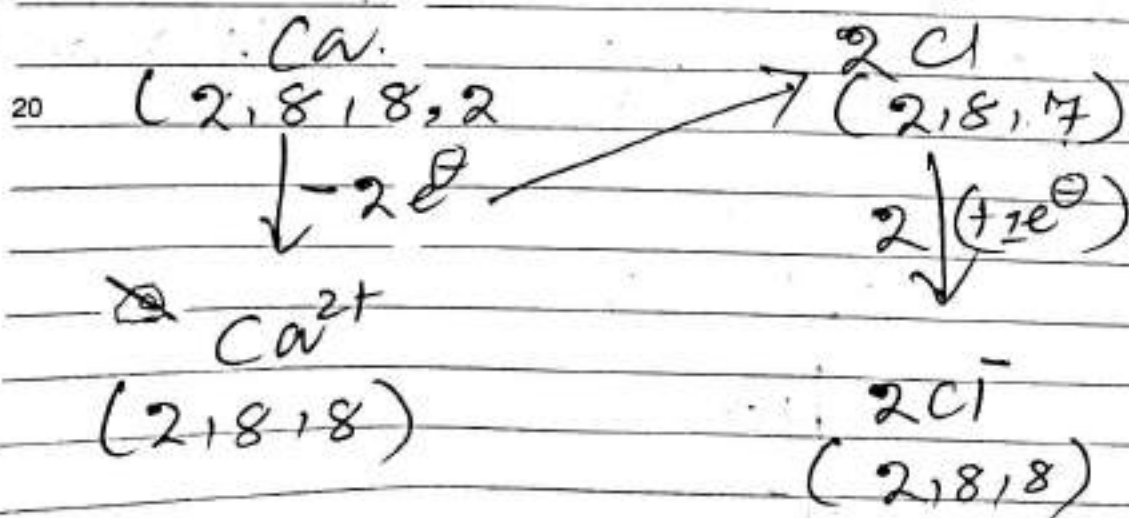
Ex. 2

Q// Explain the ionic bonding in CaCl_2 .

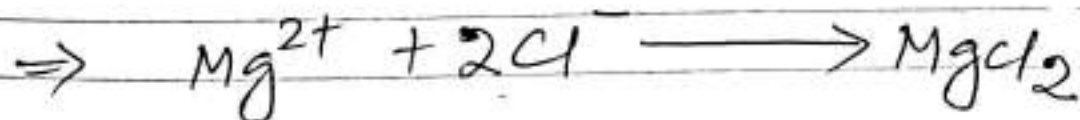
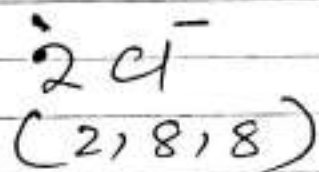
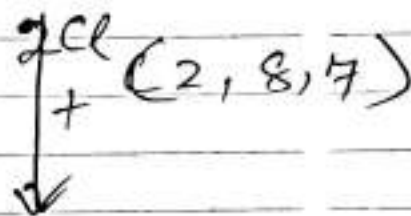
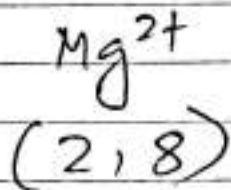
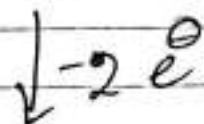
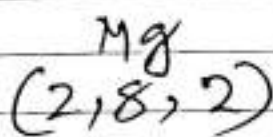
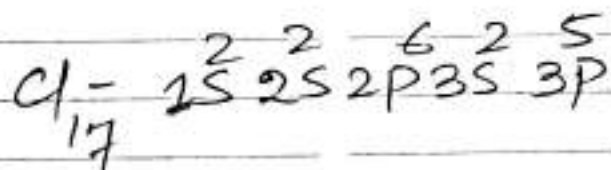
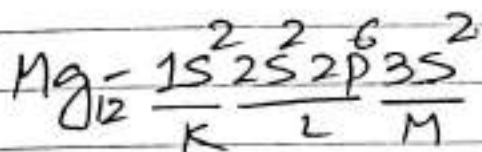
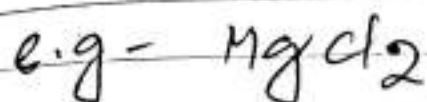
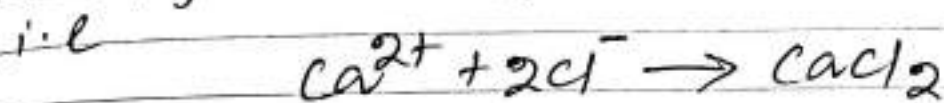
Ans:- Formation of CaCl_2



In CaCl_2 , Ca loses two electrons to get nearest inert gas configuration (i.e. Ar_{18}) and form a cation. Where as 2 Cl atoms except 1 electron each from Ca in order to get stable by attaining nearest inert gas configuration i.e. (Ar_{18}) and form anion (2Cl^-)



→ As the opposite charge attract each other the Ca^{2+} & 2Cl^- get attracted by electrostatic force of attraction by decreasing energy and form CaCl_2 molecule.



Characteristics of Ionic Bond

- This type of Bond is formed between two dissimilar atoms.
- The Ionic compound is found in solid state.
- These are hard and rigid.
- These are crystalline in nature.
- They have high melting and boiling point.
- They are polar in nature and are soluble in polar solvent. However insoluble in non-polar solvent like organic solvents.
- The Ionic compounds are bad conductor of electricity in solid state. However it conduct electricity in fused or molten state.
- Ionic compound do not show Isomerism.

NOTE :-

- (i) Isomerism :- compounds having same molecular formula different structure.
- (ii) lone pair of electron :- The valence electrons which do not take part in bonding.

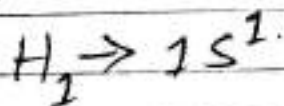
Covalent Bond:-

→ It is defined as the chemical bond which is formed by the mutual sharing of valence electrons between two atoms.

→ The compound formed is known as covalent compound.

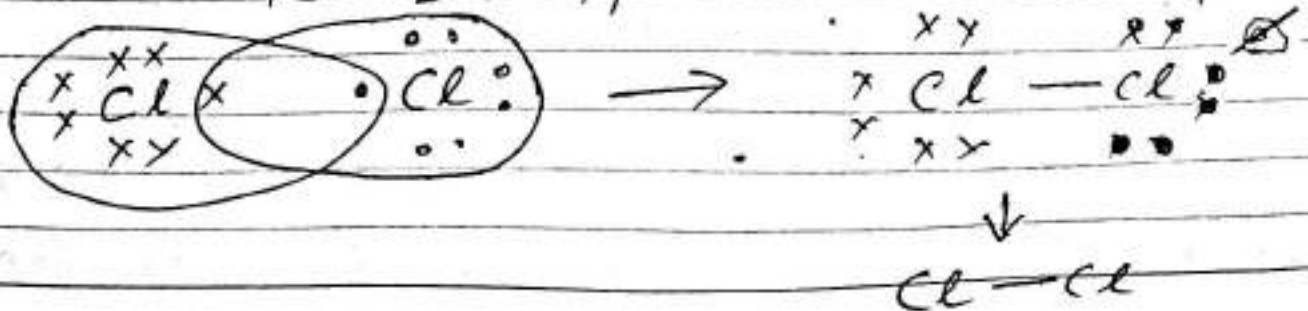
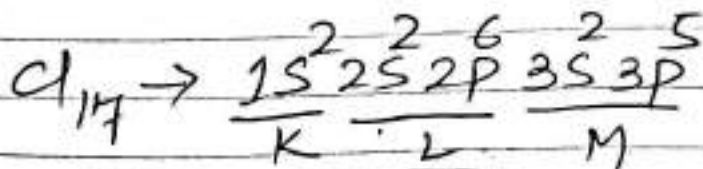
→ This type of bonding is formed between two similar or dissimilar atoms.

e.g-1. H_2 molecule



In Hydrogen molecule each Hydrogen required one more electron to become stable. Hence each hydrogen atom share their one electron to form a covalent bond betⁿ them.

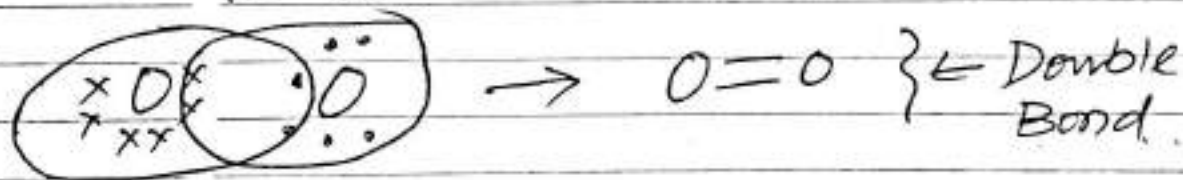
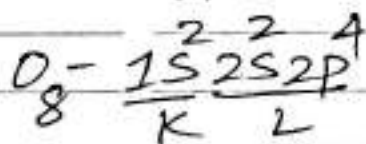
e.g-2 → Cl_2 molecule



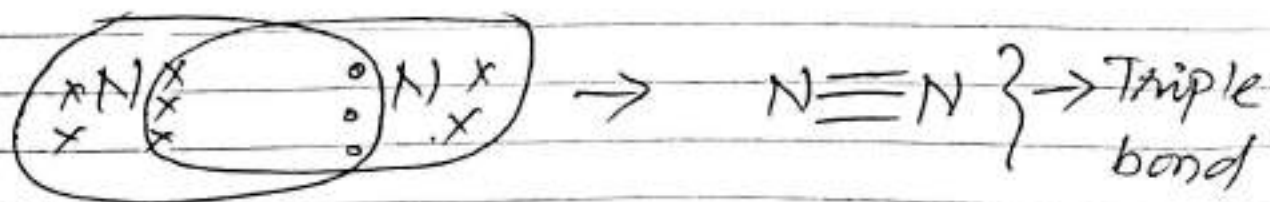
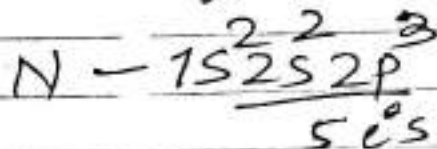
Appointment

→ Each chlorine atom has 7 electrons in their valence shell and require one more electron to acquire the nearest stable inert gas configuration i.e of Ar. Therefore, each hydrogen share one of their valence electrons between them to form covalent bond.

e.g-3 - O_2



e.g-4 - N_2



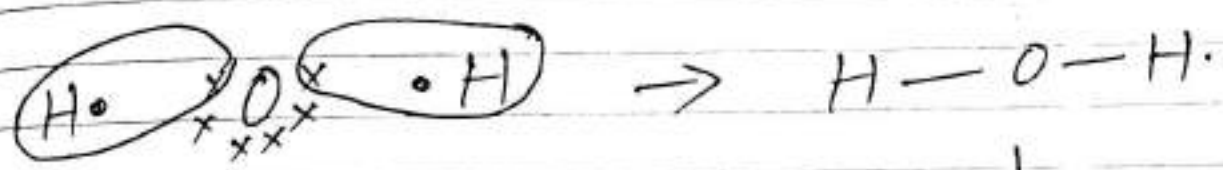
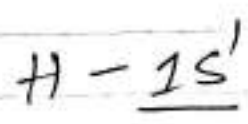
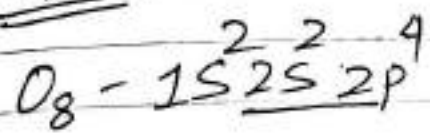
- NOTE In (i) '—' ← σ bond
 (ii) '=' ← 1 σ and 1 π bond
 (iii) '≡' ← 1 σ and 2 π bond

→ σ bond is strong due to head on overlapping

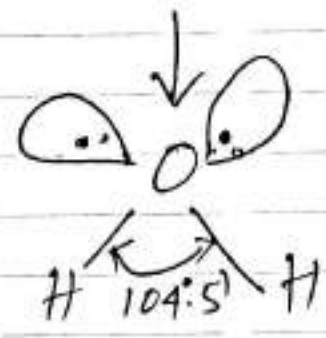
While π bond is weak due to sidewise overlapping.

Q1/ Explain the formation of H_2O molecule.

e.g-5
Ans H_2O



Shape - 'V' shape
Bond angle - 104.5°

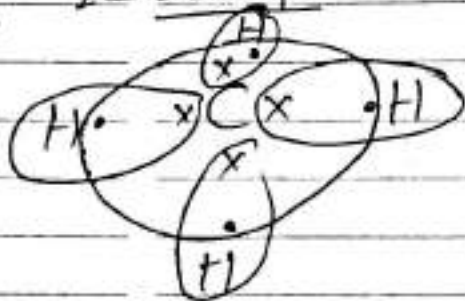


In water molecule, there are one oxygen and two hydrogen atoms are present. The oxygen atom has 6 electrons in their valence shell and require two more electrons to get stable. Similarly, the hydrogen atoms have one electron in their valence shell and require one more electron to get stable. So each hydrogen atom share its valence electron with the valence electron of oxygen to form a covalent bond.

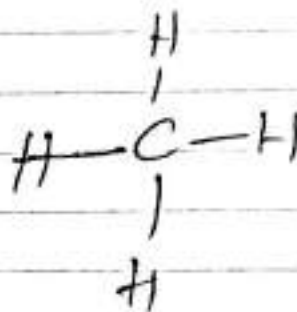
e.g-6. CH_4 - Methane

H - $1s^1$

$\text{C} - 1s^2 2s^2 2p^2$

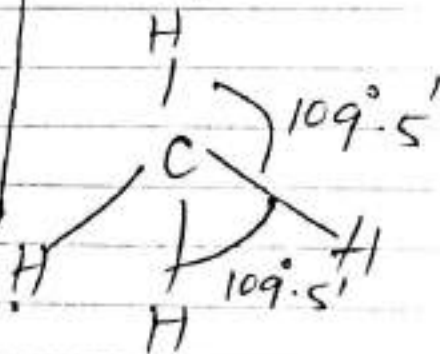


→



↓

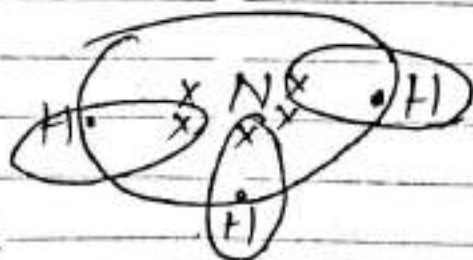
Shape - Tetrahedral
Bond angle - 109.5°



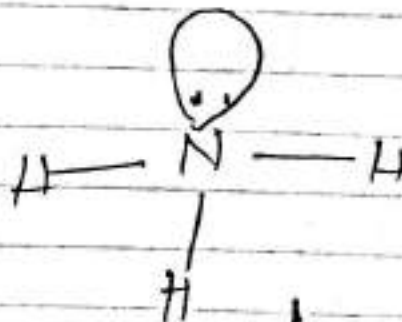
e.g-7. NH_3

$\text{N} - 1s^2 2s^2 2p^3$

H - $1s^1$

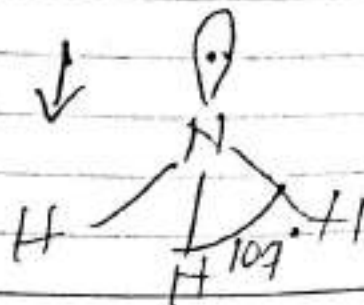


→



↓

Shape - Pyramidal
Bond angle - 107°



Chemical Bonding

Mar 17	Sun	Mon	Tue	Wed	Thu	Fri	Sat	Sun	Mon	Tue	Wed	Thu	Fri	Sat
			1	2	3	4	5	6	7	8	9	10	11	
12	13	14	15	16	17	18	19	20	21	22	23	24	25	
26	27	28	29	30	31									

February 2017

4

Saturday

Appointment

Long type

1. What is ionic bonding? Explain the formation of NaCl and CaCl_2 .

Short type

1. What do you mean by chemical bonding?

2. What is valency.

3. Write the Lewis structure of H_2O . What is the shape and bond angle of H_2O .

Long Type :-

1. What is covalent bonding? Explain the bonding with an example. (H_2 , Cl_2 , O_2 , N_2 , H_2O , CH_4)

2. What is co-ordinate bonding? Explain the bonding with an example. (SO_2 , NH_4^+)

3. Give a comparison between Ionic bonding and co-valent bonding.

SHORT TYPE Q :- (4)

1. Write the formation of NH_3 , CH_4 , SO_2 , NH_4^+ .

2. What is the structure of and bond angle of NH_3 & CH_4 .

3. Write down the characteristics of co-ordinate Bonding.