Electronic Configuration & Division of Elements into Blocks.

Elements of the periodic table has been divided into 4 blocks s, p, d and f depending on the nature of the atomic orbitals into which it differentiating electron the last electron enters.

The s-block element

In this element, the differentiating electron enters the "ns" orbital. Alkali and Alkaline earth metal of group 1a and 2a belong to this block, as you know the valence shell electronic configuration of this groups are ns and ns^2 respectively. We also know that each period of the periodic table belongs with Alkaline metals. All the elements in the s-blocks are metals.

The P-block element

The P-block element, in the element belonging to this block, the orbitals are successively filled. Thus, the element of the group 13 [III a], 14 [IV a], 15 [V a], 16 [VI a], 17 [VII a], 18 [0] are members of this block, since in the atoms of this element the differentiating electron enters the the np orbitals. The ns orbital in the atoms of this element are already completely filled so they have the valence shell electronic configuration as ns^2, np^{1-6} depending on the element.

Note that the element of the s and p blocks are also known as <u>Normal representative</u> or <u>Main group element</u>.

The d-block element.

The element in which the differentiating electrons enters (n-1)d orbitals are called the d-block. This element are placed in the middle of the periodic table between the s and p blocks element. The electronic configuration of the atoms of the element of this block can be represented by $(n-1)^{d1-10} ns^{0,1,2}$. This element which are also called <u>transitions elements</u> are divided into 4 series corresponding to filling of the 3d, 4d, 5d or 6d orbitals. While the 3d, 4d and 5d series consist of ten elements, the 6d series is incomplete and has only [illegible] elements ... actinium etc. Note that the d block elements are also known transition elements.

The f-block elements

~~The elements in which the extra electrons enters~~ the (n-2) f orbitals are called the f-blocks ~~element~~. The ~~atoms of this element have the ground~~ configuration; $(n-2)f^{1-14}(n-1)d^{0.1}ns^{2}$. This element belongs to two series depending upon the filling of 4f and 5f Orbitals. Element from Ce [z = 28]. to Lu [z = 71] are the members of the 4f series. While those from Th (z = 90) to Lr (z = 103) belongs to the 5F series. Element of 4f series which follows Lanthanium in the periodic table are called Lathanide whereas those of 5F series following Actinium are called Actinides. All these elements, are collectively refer to as inner transistion element because of filling an electron in an inner (n-2] f sub-shell. Note that f-blocks elements are also known as inner transistion elements.

LEWIS- STRUCTURE

Structural formular and Molecular Formular

Structural formular: It tells us about the covalent bond in a compound e.g CO_2. O=C=O

Molecular formular: It tells us about each atom in a molecules of the compound e.g CO_2 i.e It contains 1 atom of Carbon and 2 atoms of Oxygen

Lewis-structure: It shows not only the covalent bond but the valence electrons as well.

e.g CO_2 $|\bar{O} = C = \bar{O}|$

Rules for Writing lewis structures

Rule 1:

The total number of valence electron in any molecule or ion is equal to the sum of the valence electrons that each atom had before bonding, plus or minus charge in case of ions

Example:

(1) Calculate the total valence electron for the following (a) CO_2 (b) C_2H_2 (c) SO_4^{2-}

Soln/

(i) CO_2 C = 4

O = 6

Total valence electron = 4 + 2(6) = 16.

(b) C_2H_2 $C = 4$

$H = 1$

Total valence electron $= 4(2) + 1(2)$

$= 10$

(c) SO_4^{2-} $S = 6$

$O = 6$

Total valence electron $= 6 + 6(4) + 2$

$= 32$.

Rule 2:

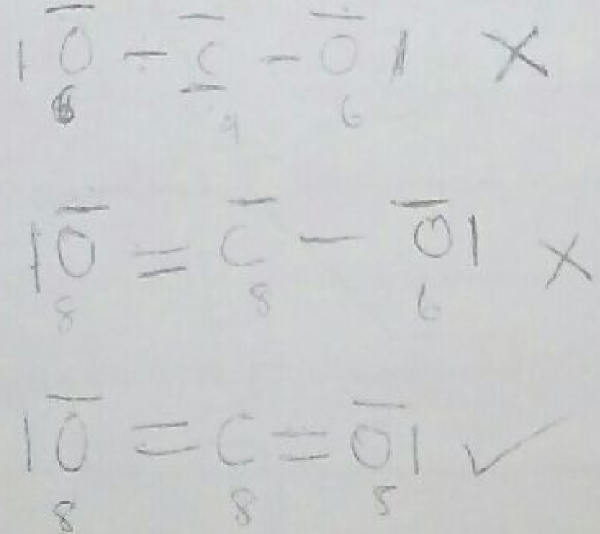
All the valence electrons calculated in group 1 and now inserted in pairs, in such a way that each atom has an octet configuration

e.g CO_2,

total valence $= 4 + 2(6) = 16$

16 valence electron ≡ 8 pair or 8 single bonds.

i.e



(ii) ~~NH3~~ N_2 . N=5

~~total valence electron = 5(2) = 10~~

10 - valence electron ≡ 5 pairs

$|\bar{N} - \bar{N}|$ ×
6 6

$|\bar{N} = \bar{N}$ ×
8 6

$\bar{N} \equiv \bar{N}$ ✓ or $|N \equiv N|$ ✓
8 8 8 8

(iii) NH_3 N = 5 H = 1

valence electron = 5 + 1 × 3 = 8

8 valence electron = ~~2~~ 4 pairs

H — N — H
2 ‖ 8 2
H 4
×

H — N = H
2 ‖ 8
H 2
×

H — $\bar{N}$ — H
2 | 8 2
H

Rule 3:

If an octet is not possible, an outer shell of six or seven is permissible.

Note that: atoms in the second row of the Periodic table can hold more than 8 but atoms in the third row or below can hold ten or more. e.g.

Draw a lewis structure for antimony chloride ($SbCl_5$).

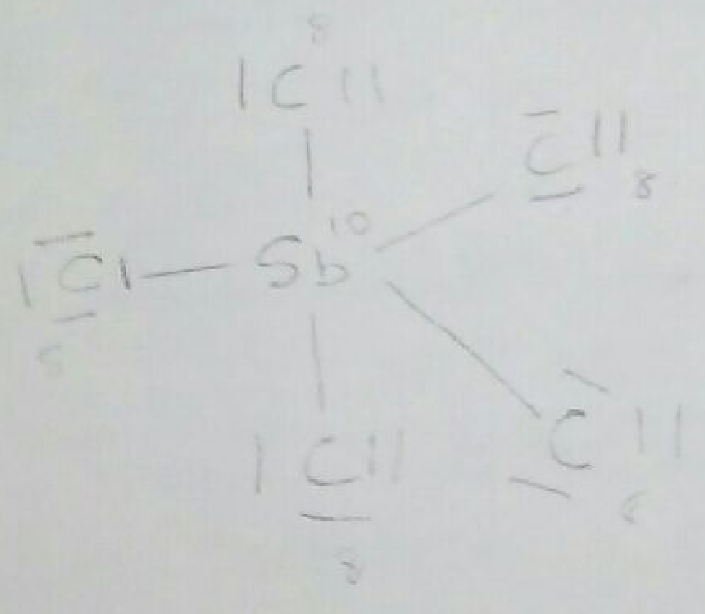
Solution

$SbCl_5$ Sb = 5 Cl = 7

total electrons = 5 + 5(7)

= 40

to valence electrons = 20 pairs



Rule 4: B

~~When all the valence electrons have been inse~~
it is necessary to inspect each atom to see if
~~must be given a formal charge.~~

The Formal charge for each atom is given b
the following formula ;-

Formal charge On a given atom = No of electrons in the metal atom before bonding − No of valency electrons − No o Covale bond on at

CHEMICAL BONDING

A molecule will only be formed if it is more stable and has a lower energy t the individual atoms.

Noble gases do not react with any other atom, [illegible] they are monoatomic and t are unreactive because the atoms already have low energy.

The low energy of the noble gases is associ with their having complete outer shell elec

Fot

TYPES OF BOND

1. There are 3 different ways by which atom attain a stable electronic configuration; by-:

(i) Loosing electron

(ii) Sharing electron

(iii) gaining electron

TYPES OF ELEMENT

Element may be divided into;

(i) Electropositive element [they loose electrons]

(ii) Electronegative element [they gain electrons]

(iii) Element with little tendency to loose or gain electrons

Electrovalent bond; combination of electropositive & electronegative

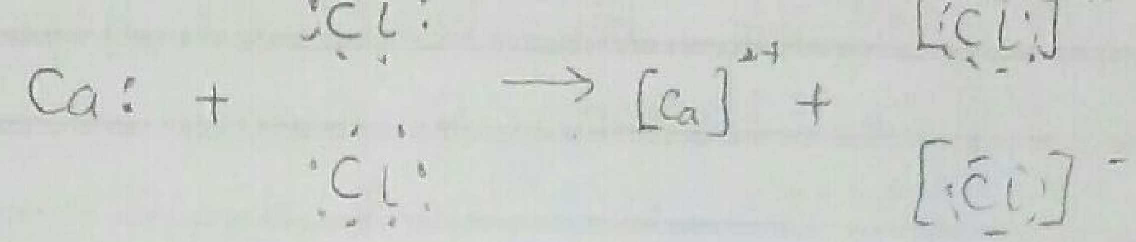
Covalent bond; electro negative + electro negative

Metallic bond; electron positive + electron Positive

Ionic bond involves a complete transfer of one or more electron(s) from one atom to another e.g NaCl

Note: When Na atom looses electron it becomes stable element which is Neon.

Lewis structure can be used to represent the formation of $CaCl_2$. Ca has 2 electrons in its outer shell, it is an electropositive so each Ca looses 2 electrons to two chloride atoms forming the chloride of 2 positive ions



Xteristics of Ionic bond.

1 They have high melting & boiling points bcoz (i) their is strong electrovalent bond between the ions e.g the boiling points of Na is 1467°C while the melting point is 801°C. (ii) The high energy is needed to break down the crystal lattice.

- majority of them are solid at room temperature.

2. They are easily soluble in water in which two energies are involved.

(i) lattice energy (ii) hydration energy

E.g Sodium chloride is soluble in H_2O because lattice energy is less than hydration energy

Lattice energy is the energy needed to break ~~up bond while hydration energy is the~~ energy liberated on hydration of ions. ~~When lattice energy greater than hydration~~ energy, the ionic compound form is insoluble.

i.e lattice energy $<$ hydration energy = Soluble

lattice energy $>$ hydration energy = Insoluble

3 They are good electrolyte i.e they conduct electricity

Covalent bond

When two atoms react together, both have a tendency to gain electron but neither atom has any tendency to lose electron. In such case, atom share electron in order to obtain a noble gas configuration. The bond holding such atoms together when pairs of electrons are shared between atoms is called Covalent bond.

Eg H_2 & Cl_2

$$:\ddot{Cl}\cdot + \cdot\ddot{Cl}: \longrightarrow :\ddot{Cl}:\ddot{Cl}:$$

Each Cl atom gives a share of one of each electron to another.

H· + ·Ö· + ·H → H:Ö:H

or

H—Ö—H

~~Multiple Covalent bond~~

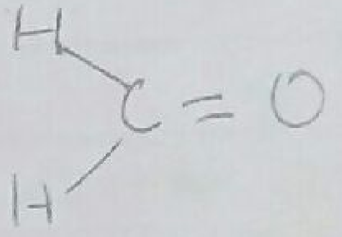
A single bond have one pair of electr
i.e one pair of electron is shared to form
one bond.

Double bond; four electrons are share
to form 2 bonds

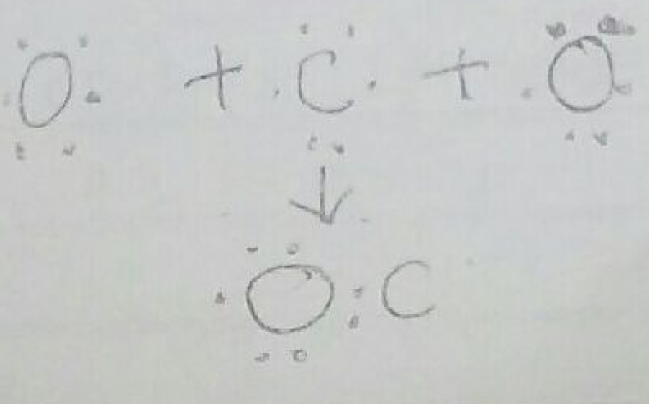
Tripple bond; Six electrons are share
to form 3 bonds;

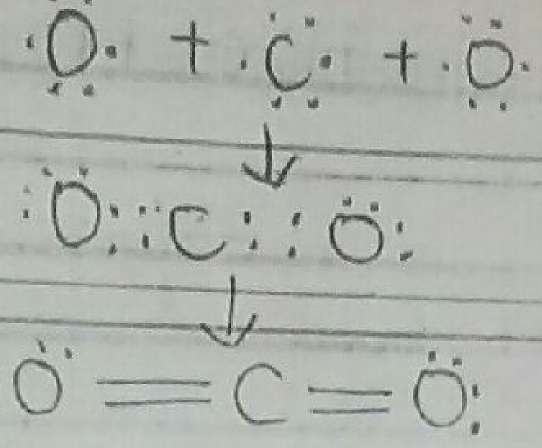
NB! N_2 has 3 bonds / triple bond
Carbon Monoxide [CO]
C≡O has triple bond

Bromine hydride


has double bond

FORMAL DEHYDE





Characteristics of Covalent bond.

1. Covalent compounds have lower boiling & melting point than ionic bond. because There is weak intermolecular bond except diamond & graphite, bcoz their atoms are held together in a continuous giant structure.

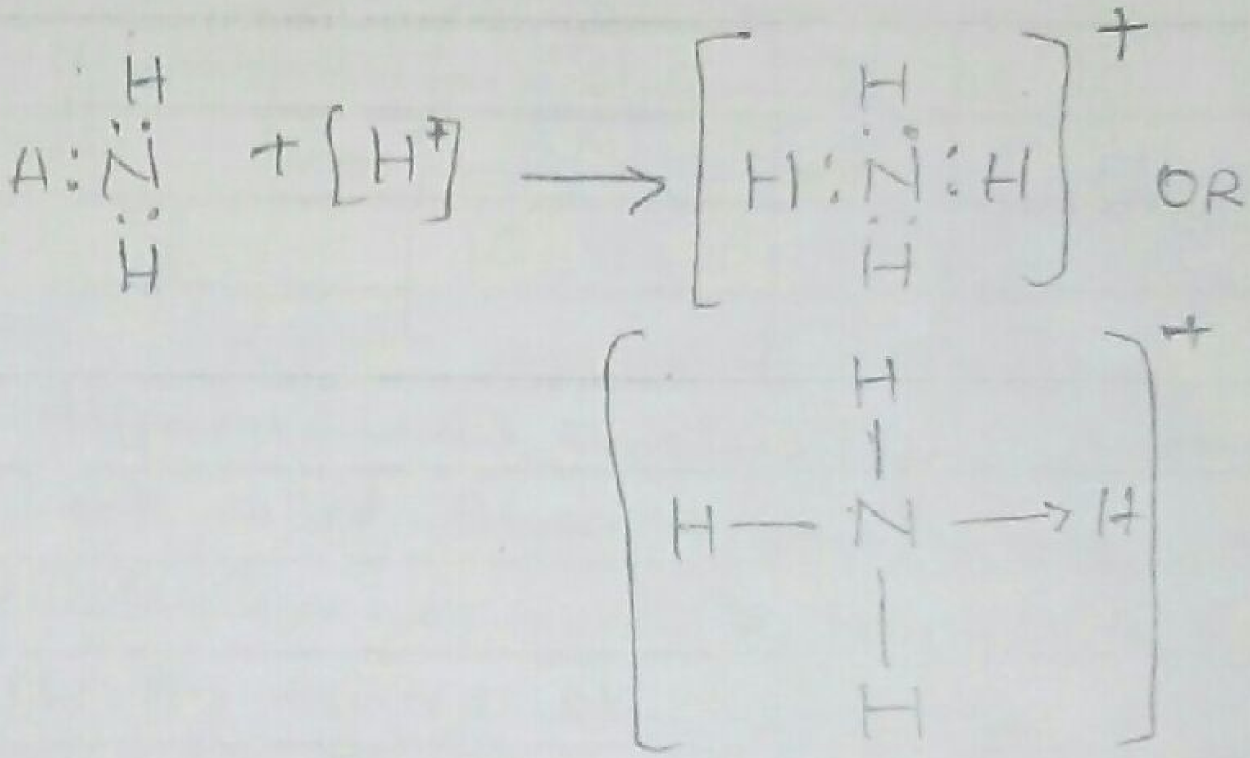
2. Covalent compound dissolve in non-polar solvent while ionic compund dissolve in polar solvent. Example of non-polar solvent molecule bromine,

3. They are non-electrolyte i.e they do not conduct electricity

DATIVE / COORDINATE BOND

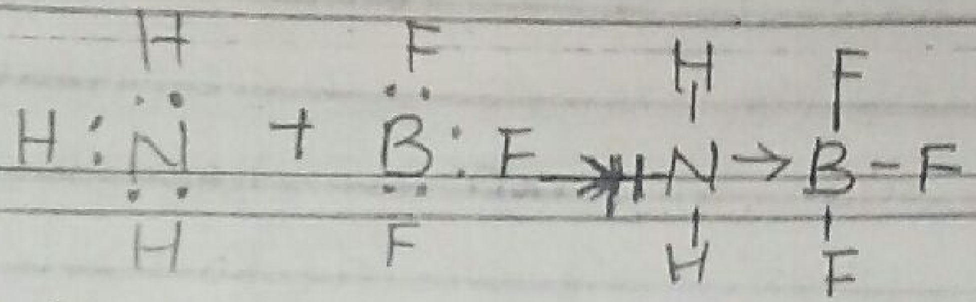
It is a covalent bond where the electron pair shared originate from one atom and not from the other. They are identical to normal covalent bond. Good example of compound that has dative bond is NH_4^+

Representing using Lewis structure, we have,



i.e the electron used to form dative bond comes from only nitrogen and it is called Lone pair electron. It is different from covalent because it is indicated using an "arrow".

2 BF_3 is another good example of dative bond



Reaction of Ammonia and BF_3, the electron used to form the bond is from NH_4^+

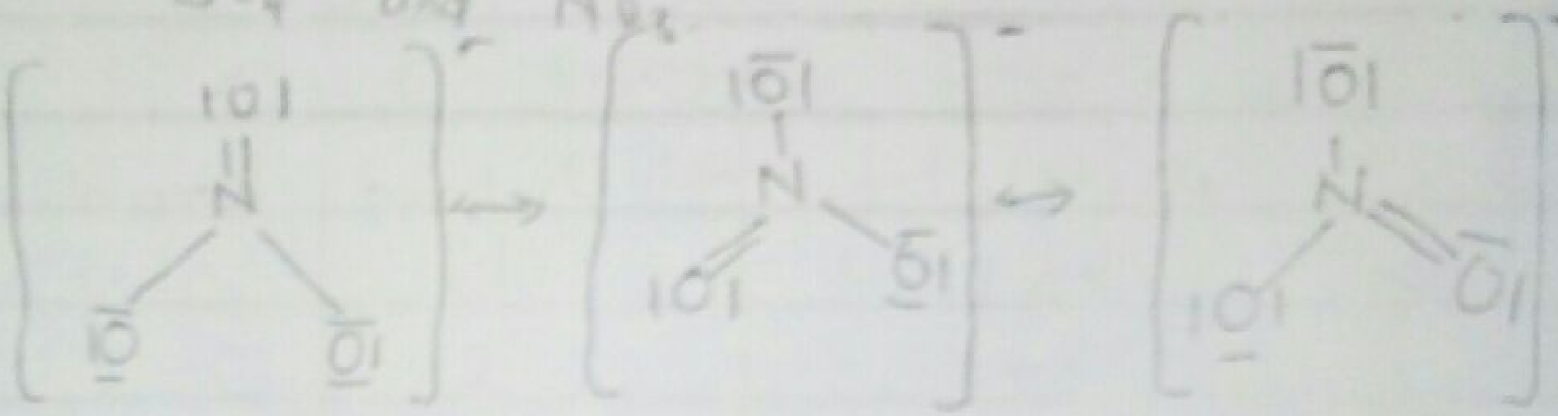
HYDROGEN BOND.

It is a Dipole - Dipole attraction which occurs between hydrogen atom attached to a strongly electronegative atom. They are usually chlorine, oxygen or nitrogen. Hydrogen bond is a very weak bond.

Resonance Structures

We have found that it is possible for a molecule or ion to have more than one structures because there may be more than one way of distributing or positioning the electron pairs such than an octet of electron surrounds each atom e.g.

SO_4^{2-} and NO_3^-



The different Lewis structure for Nitrate ion (NO_3^-) are called Resonance structure and the actual structures are said to be resonance hybrid that is, the blended structure of two or more Lewis structure. Resonance occur when more than valid Lewis structure, that differ only in the allocation of electrons can be written for a particular molecule. Thus in resonance structure all atoms must be in the same positions, only the position of the electrons are different

Lewis Structure and Bond Properties. 17

In lewis structures, a pair of electron that is shared between two atoms are called bonding Pairs and are generally referred to as Covalent bond. Electrons that are not shared are called Lone Pairs or non-bonding Pairs. Where there are two shared pairs of electrons we have double covalent bonds while three pairs of electrons shared between bonded two atoms, they result in the formation of triple bonds. The number of bonds between two atoms in a molecule or ion is called Bond Order

eg C − C bond order = 1
C = C bond order = 2
C ≡ C bond order = 3

Bond Length

Bond length is defined as the distance btw the nuclei of two bonded atoms to a good approximation, the bond length can be taken as the sum of covalent radii of the two bonded atoms.

e.g If covalent radii of H,F and Cl

are 0.37Å, 0.64Å and 0.99Å.

The distances between the centre of two hydrogen (H_2), F_2 and Cl_2 are

$H_2 = 0.37 + 0.37 = 0.74Å$

$F_2 = 0.64 + 0.64 = 1.28Å$

$Cl_2 = 0.99 + 0.99 = 1.98Å$

Also consider the following

C−C, 1.54Å

C=C, 1.34Å

C≡C 1.20Å

bond	bond Order	bond length
C−C	1	1.54Å
C=C	2	1.34Å
C≡C	3	1.20Å

From the tables, we can say that the bond length varies with the bond order i.e As the bond order increases, the bond length decreases.

Bond Strength

The strength of a particular bond can be measured by it bond energy. Bond energy is the energy required to break one mole of

a bond of a given type.
Consider the table below:

Bond	Bond length	Bond energy
C-C	1.54 Å	347 kJ/mol
C=C	1.34 Å	614 kJ/mol
C≡C	1.20 Å°	839 kJ/mol.

From the table, it can be seen that as the bond energy increases, the bond length decreases. Thus the stronger the bond, the greater the energy needed to break such that bond strength of chemical bonds:

C≡C > C=C > C-C
839 kJ/mol 614 kJ/mol 347 kJ/mol

Bond	Bond length	Bond order
C-C	1.54 Å°	1
C=C	1.34 Å°	2
C≡C	1.20 Å°	3

As bond order increases, bond length decreases

Bond	bond length	bond energy
C-C	1.54 Å°	347 kJ/mol
C=C	1.34 Å°	614 kJ/mol
C≡C	1.20 Å°	839 kJ/mol.

→ As the bond length decreases, the bond energy increases and vice versa.

$C \equiv C > C = C > C - C$

HYBRIDISATION

The Process of mixing two or more atomic orbitals to produce a new orbital called Hybrid Orbitals. Example of Atomic orbitals S, P, D, F

$1s + 1p \longrightarrow sp$ hybrid orbitals

$1s + 2p \longrightarrow sp^2$ hybrid orbitals

$1s + 3p \longrightarrow sp^3$ hybrid orbitals

$1s + 2p + 1d \longrightarrow sp^2d$ hybrid orbitals

$1s + 3p + 2d \longrightarrow sp^3d^2$

mation of Sp^3 hybrid orbitals

Carbon

Carbon (ground state) $= 2s^2 2p^2$

$P \longrightarrow P_x, P_y, P_z$

Carbon (ground state) $= 2p_x^1 \quad 2p_y^1 \quad 2s^2 \xrightarrow{\text{excited}} 2p_x^1 \quad 2p_y^1 \quad 2p_z^1 \quad 2s^1$ } mixing Sp^3 hybrid orbitals.

g CH_4, H_2O, NH_3.

Formation of Sp^2 hybrid Orbitals

Consider boron with ground state electronic configuration of $2s^2 2p^1$

Boron (ground state)

$2p^1$ — $2s^2$ —excited→ $2p_y^1$, $2p_x^1$, $2s^1$ } mixing Sp^2 hybrid orbitals.

e.g BF_3.

Formation of Sp hybrid Orbitals.

Consider Berryllium.

$Be = 2s^2$

Be (ground state)

$2p_x^0$ — $2s^2$ —excited→ $2p_x^1$, $2s^1$ } mixing of these two orbitals give rise to Sp hybrid orbitals.

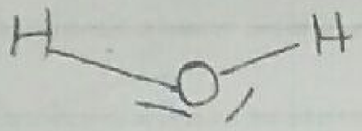
Electron Pair Repulsion theory and the shapes of Molecules.

EPRT

Used to Predict the shape or geometry of molecules, Also called Valence shell electron Pair repulsion Model (VSEPR).

When we used this theory, we focus on the central atom of the molecule. This theory states that all valence shell electron pairs surrounding the central atom locate theselves in such a manner as to be as far away from each other as possible.

Two distinct types of electron pairs may surround the central atom e.g H_2O



bonds (a) ≡ Covalent bonds due to bonding electron pairs

bond (b, c) ≡ Non-bonding electron pairs (unshared pairs).

Molecules and their Geometry.

Cordination number; is the number of bonds attached to the central metallic atom in a compounds

AB_2 Molecules.

B – A – B

Coord No = 2

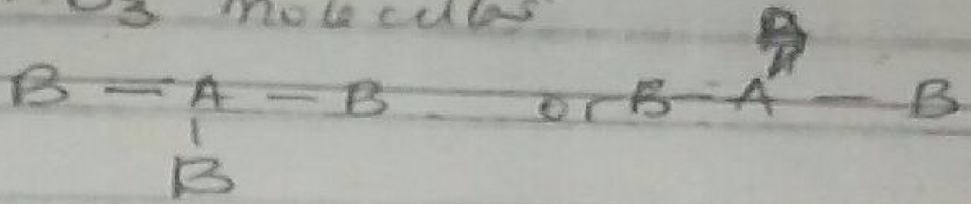
Shape = linear

bond angle = 180°

e.g BeF_2 ; $BeCl_2$

F – B – F , Cl – Be – Cl

AB3 molecules



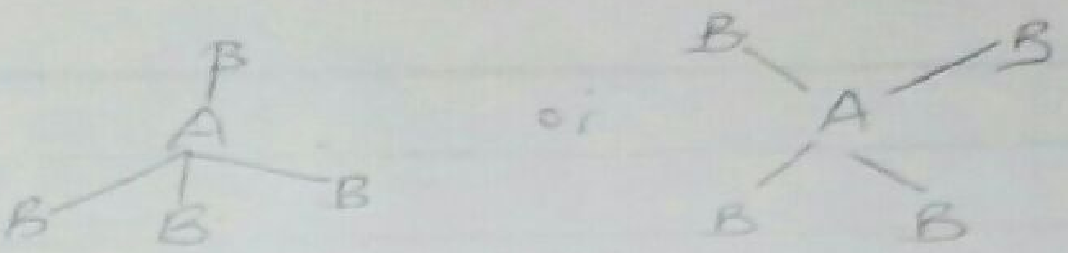
Coord No = 3

Shape = triangular

bond angle = 120°

e.g BF_3

AB4 molecules



Coord. No = 4

Shape — tetrahedral / square planner

tetrahedral	Square planner
bond angle 109°	90°
No isomer	isomer — cis / trans
e.g CH_4	e.g $Pb(NH_3)_2Cl_2$

AB5 molecules

Coordinate No = 5

Shape — trigonal bipyramid / square pyramid

Bond angle = 90°, 120°.

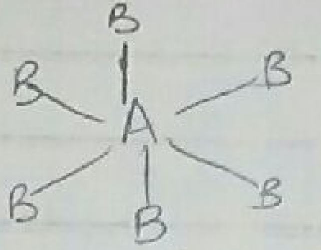
~~AB_6 Molecules~~

Coordinate No = 6

~~Shape = Octahedral~~

bond angle = 90°

eg SF_6, $Co(NH_3)_4Cl_2$.



SF_6, $Mn(H_2O)_4Cl_2$, $Co(NH_3)_4Cl_2$

Isomer → Cis / trans

Solid Structures.

Primitive → 1, P

Body centered - 2, I

face centered - 4, F

End centered - 2, C

SYSTEM

There are 14 ~~geaptno~~ Bravais Lattices for solid compound, there are 7 systems

1. Cubic, Unit cell, axes and angle

are = the length ; a = the breadth, b = the height c

$\alpha = \beta = \gamma = 90°$ [All the angles are equal]