CHEMICAL BONDING

- Electrophiles are ~~elements~~ atoms having positive charge due to loss of electrons.
- Nucleophiles are atoms having negative charge due to gain of electrons.
- Electronegativity is the ability ~~to~~ of an atom to attract to itself an electron pair shared with another atom in a chemical bond.
- The greater it's value, greater the power of an atom to attract the electrons in a covalent bond towards itself.
- It's value increases up each group.
- It's value increases from Grp I to Grp 17.
- Fluorine is the most electronegative element.
- Value decreases down a group.
- Halogens have highest values of electronegativity.
- Non-metals have higher value while metals have lesser value of electronegativity.

~~Electron~~

Factors influencing it.

- Nuclear charge: atoms in same period with greater +ve nuclear charge are more likely to attract bonding pair of electrons.
- Atomic radius: Inversly proportional. As atomic radius increases, nuclear force of attraction to other electrons is less, so electronegativity is less.

/.

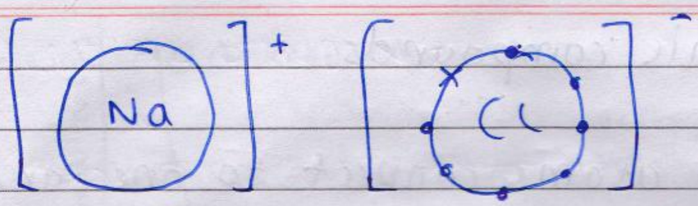
- Shielding ÷ Higher shielding effect means lower nuclear force of attraction and thus less electronegativity.
- Oxidation number ÷ If ON is +ve, the ~~at~~ element lacks electrons so electronegativity will be more. If its +2 it's even more than when its +1.
- The most commonly used scale of electronegativity values is called the Pauling electronegativity scale, N.

$$\underset{2.5}{C} \overset{\times}{\underset{\cdot}{\text{———}}} \underset{2.1}{H}$$

Since electronegativity of carbon > hydrogen, the electrons will shift towards C if they were to combine.
Thus AlF_3 is more ionic than $AlCl_3$ even though they are both covalent bonds.
$AlCl_3$ – $E(Cl) - E(Al) = 3 - 1.5 = 1.5$
AlF_3 – $E(F) - E(Al) = 4 - 1.5 = 2.5$ ✓

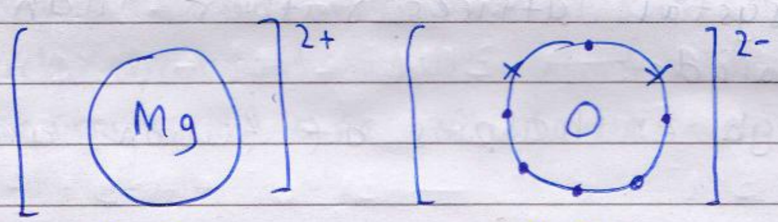
Ionic Bonding (Electrovalent bond)

- +ve ions are formed when atoms lose e^- (Cations)
- -ve ions are formed when atoms gain e^- (Anions)
- Individual ions are never stable. An opposite charge ion with same charge must be present to stabalize the ion.
- The force of attraction between +ve and -ve ions is called electrostatic force of attraction.

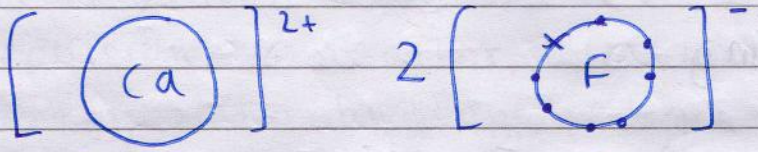


2,8 (Ne) 2,8,8 (Ar)

When ions are formed the atoms resemble configuration of their nearest noble gas.



2,8 (Ne) 2,8 (Ne)



2,8,8 (Ar) 2,8 (Ne)

~~Isoelectronic elements are~~

The ions in a compound having same electronic structure are termed as isoelectronic ions.
Eg = MgO. These form very stable bonds (lattices ~~ar~~ and thus their MP and BP is very high.

Properties of ionic compounds

- They form when atoms connect to one another by ionic bonds.
- Ionic bonds are the strongest chemical bonds.
- One atom have partial +ve charge while other has partial -ve charge.
- Ionic compounds have high MP and BP.
- They form crystal lattices rather than amorphous solids.
- They have high enthalpies of fusion and vaporization.
- They are hard and brittle.
- They conduct electricity when in water.

Covalent Bonding

(i) Element with high IE → difficult to lose e^-
(ii) Element with low EA (electron affinity)
↳ difficult to gain e^-

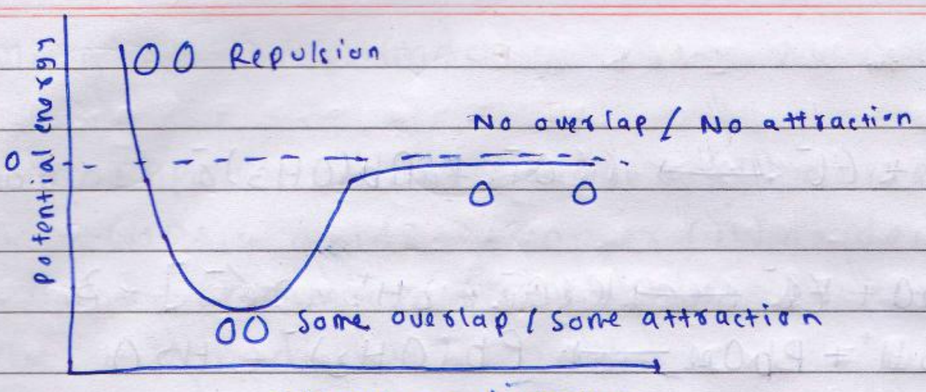
Since, it difficult to lose and gain electrons, the electrons are shared between the elements and thus covalent bond is formed.

Electron affinity is the amount of energy released/changed/involved when electrons are added to one mole of +ve gaseous ions to form negative ions / ions with lesser charge.

Electrons that gain e^- easily have high EA.

$F_{at\ radius} < F^-_{at\ radius}$

Bonds like $C \equiv C$ have nucleus very close which repel each other making $C \equiv C$ unstable. thus C≣C can't exist.



Effective covalent bonds will only form when they are at a certain threshold distance and such that the two atoms have lost their energy and are at lower energy states.

In electron affinity once an electron is added to O^- to form O^{2-}, the extra e^- will reduce at. radius and thus ~~there'll be~~ less there'll be repulsion to another electron and thus more energy is to be supplied.
Addition of first e^- is exothermic while second e^- is endothermic.

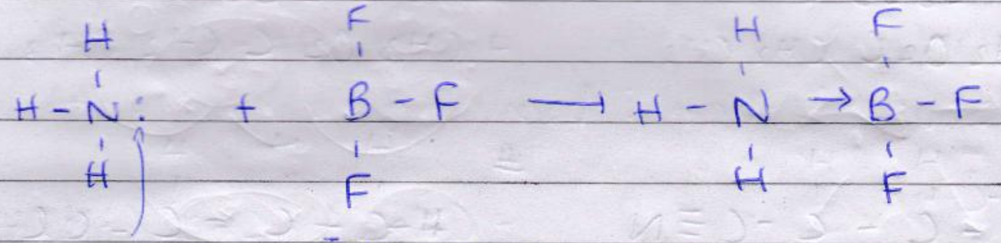
Properties of covalent compounds

- They have relatively low MP. and BP.
- Most covalent compounds are gaseous.
- Have low enthalpy of fusion.
- Tend to be soft & flexible.
- They also tend to be more flammable.
- They don't separate in water and thus aren't electrical conductors.
- They also don't dissolve in water very well.

CHEMICAL BONDING...

Coordinate covalent bond.

- When there is sharing of e^- but only one of the elements contributes the e^-.



lone pair

e^- rich

This has $2e^-$

↓

Lewis base

$\bar{e}$ deficient as all $3e^-$ are used and only $6e^-$ are left still needs $2e^-$

↓

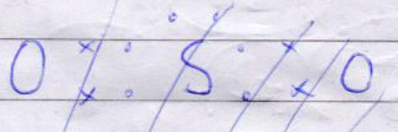
Lewis acid

coordinate covalent bond / dative bond

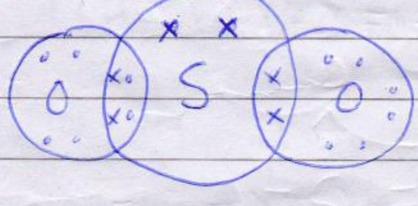
→ indicates contribution.

Ammonium boron trifluoride (Salt)

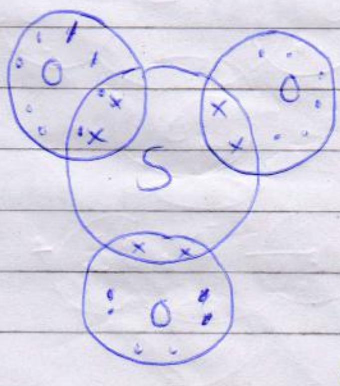
- Only pair of electrons can be contributed. single e^- donation isn't possible



SO_2 ..



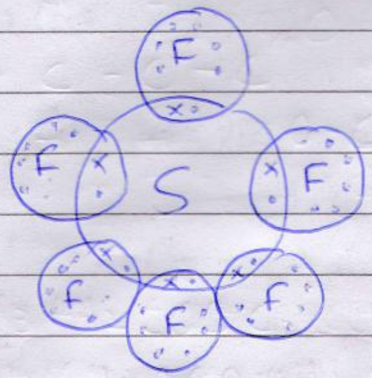
O=S→O (SO_2)



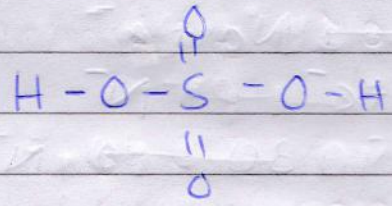
O

↑

O=S→O (SO_3)



SF_6



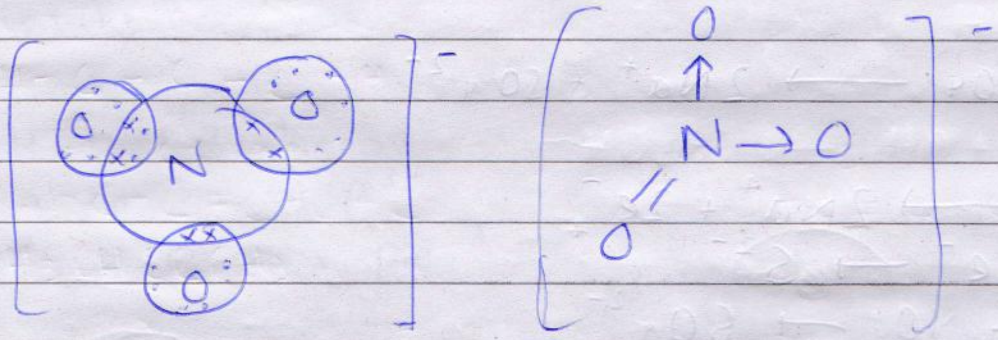
H_2SO_4

$HNO_3 \rightleftharpoons H^+ + NO_3^-$

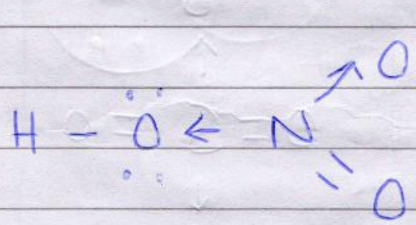
(i) $H \longrightarrow H^+ + e^-$

(ii) $N + e^- \longrightarrow N^-$

(iii) $N^- + 3O \longrightarrow NO_3^-$



$H^+ + NO_3^- \longrightarrow HNO_3$



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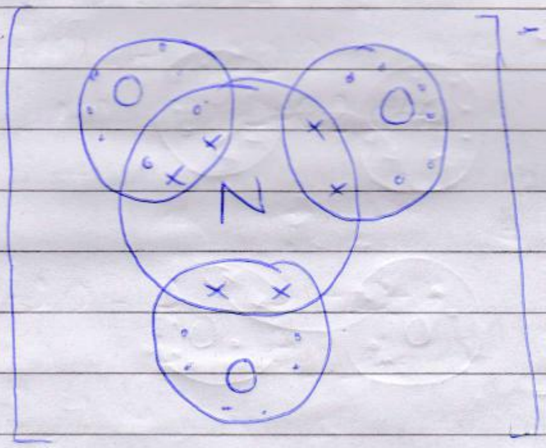
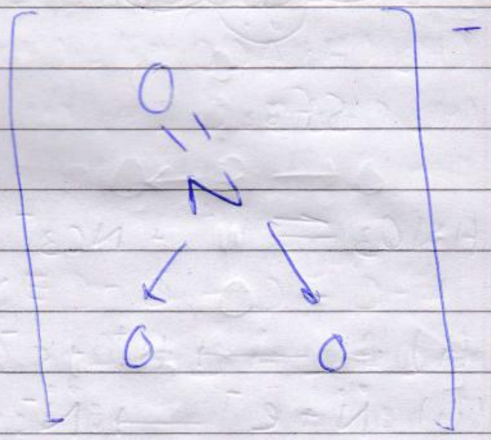
$NaNO_3 \rightleftharpoons Na^+ + NO_3^-$

(i) $Na \longrightarrow Na + e^-$
(ii) $N + e^- \longrightarrow N^-$
(iii) $N^- + 3O \longrightarrow NO_3^-$

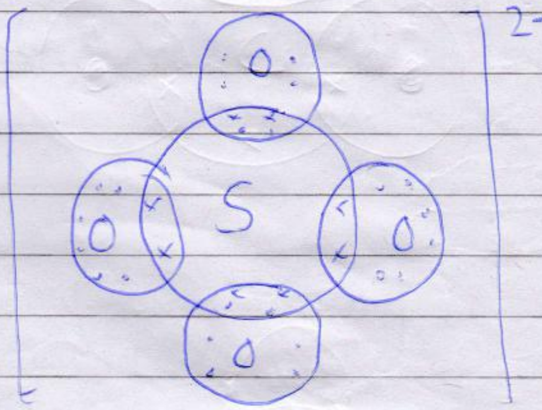


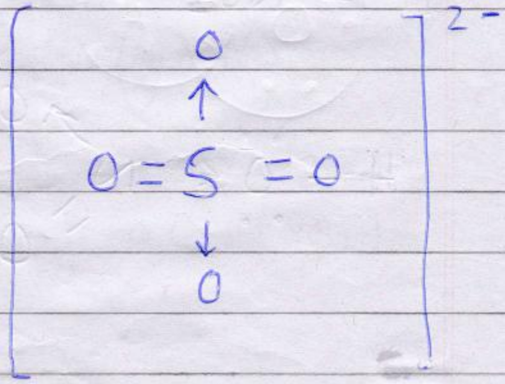


Na_2SO_4, K_2CO_3, H_3PO_3, H_3PO_4

(i) $Na_2SO_4 \longrightarrow 2Na^+ + SO_4^{2-}$

$2Na \longrightarrow 2Na^+ + 2e^-$
$S + 2e^- \longrightarrow S^{2-}$
$S^{2-} + 4O \longrightarrow SO_4^{2-}$





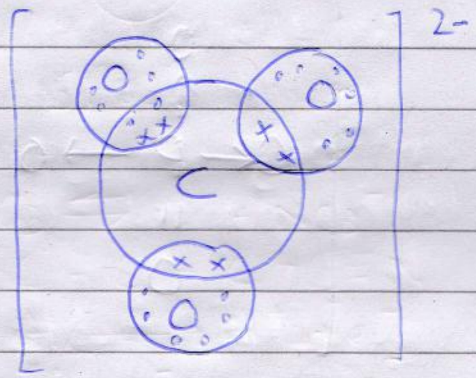
$2Na^+ + SO_4^{2-} \longrightarrow Na_2SO_4$

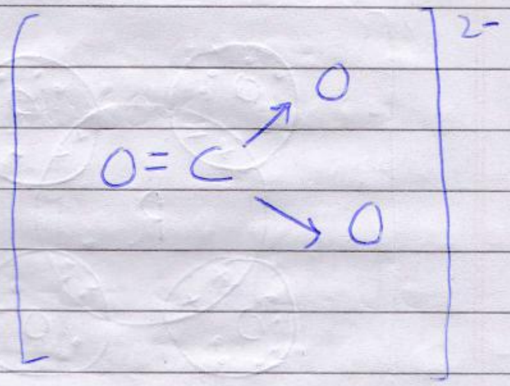
(ii) K_2CO_3

$\cancel{2K^+ + CO_3}\ 2K \longrightarrow 2K^+ + 2e^-$

$C + 2e^- \longrightarrow C^{2-}$

$C^{2-} + 3O \longrightarrow CO_3^{\,2-}$





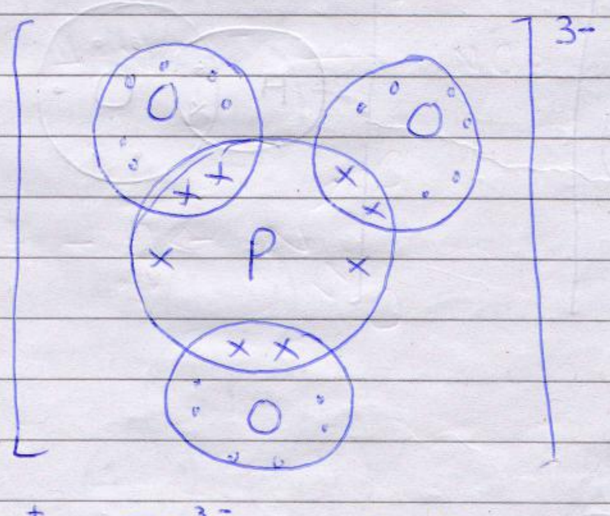
$2K^+ + CO_3^{\,2-} \longrightarrow K_2CO_3$

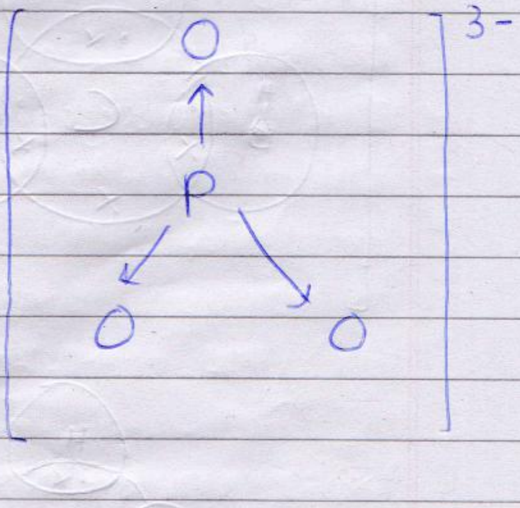
(iii) H_3PO_3

$3H \longrightarrow 3H^+ + 3e^-$

$P + 3e^- \longrightarrow P^{3-}$

$P^{3-} + 3O \longrightarrow PO_3^{\,3-}$

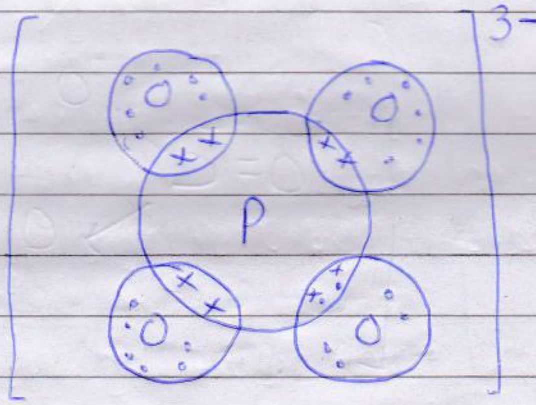


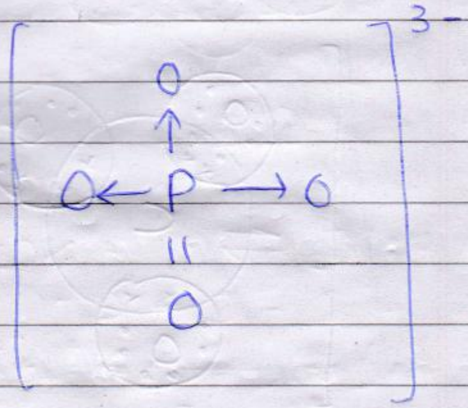


$3H^+ + PO_3^{\,3-} \longrightarrow H_3PO_3$

(iv) H_3PO_4

$3H \longrightarrow 3H^+ + 3e^-$
$P + 3e^- \longrightarrow P^{3-}$
$P^{3-} + 4O \longrightarrow PO_4^{3-}$

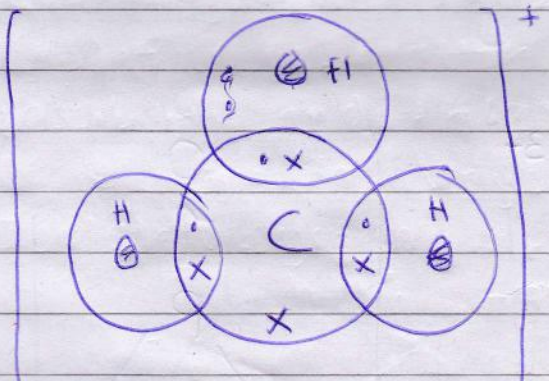


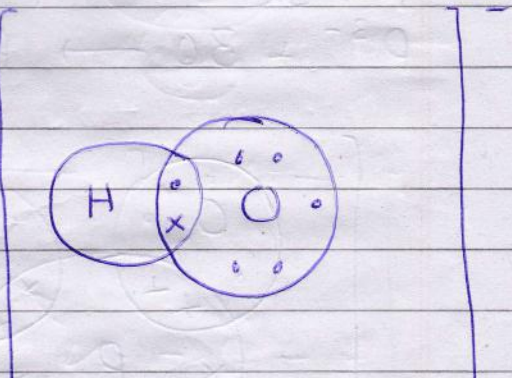


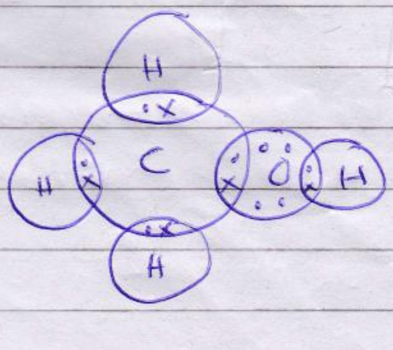
$3H^+ + PO_4^{3-} \longrightarrow H_3PO_4$

CH_3OH

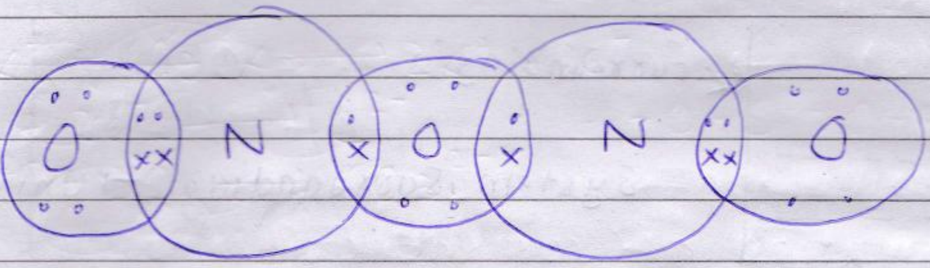
$CH_3^+ + OH^- \longrightarrow CH_3OH$



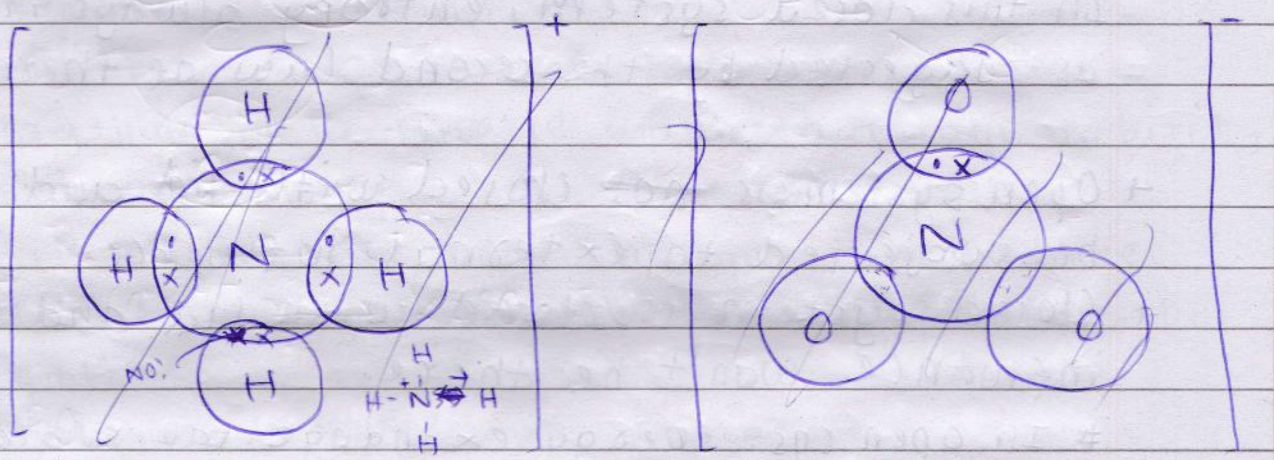


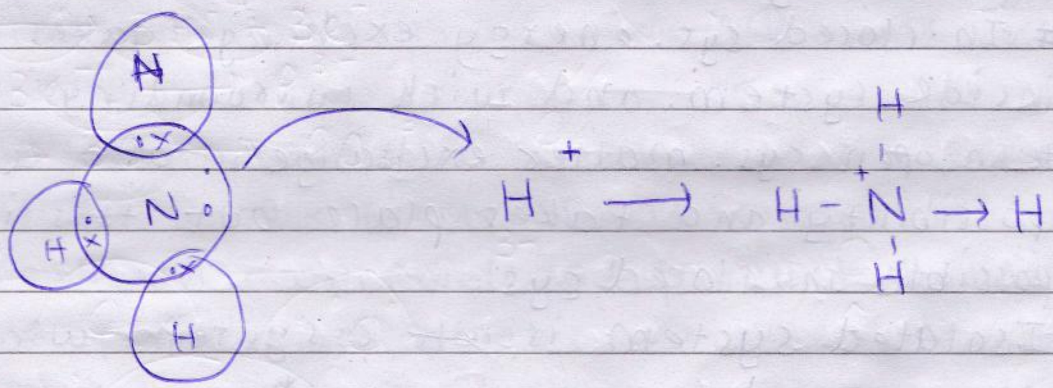


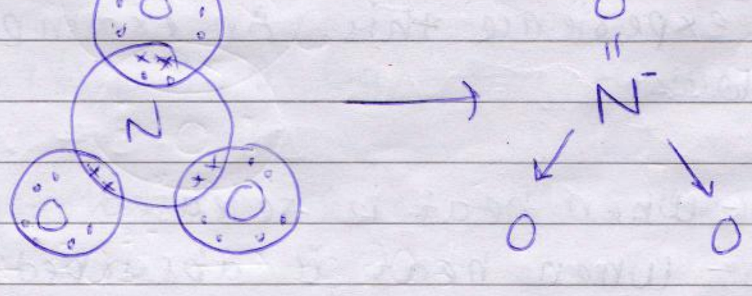
N_2O_3



NH_4NO_3

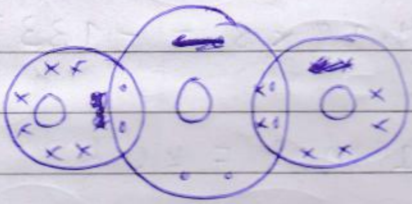






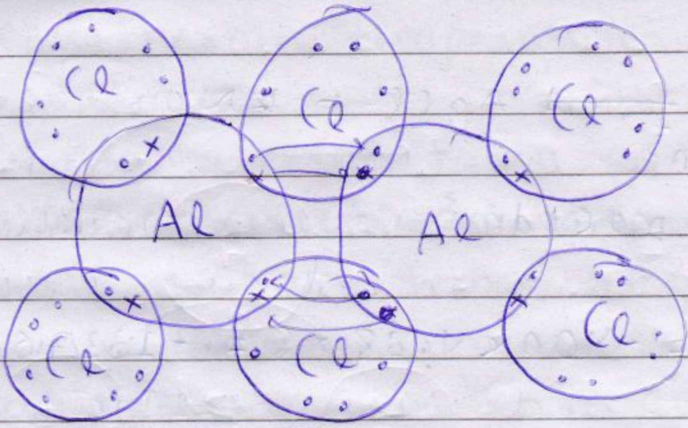
CHEMICAL BONDING...

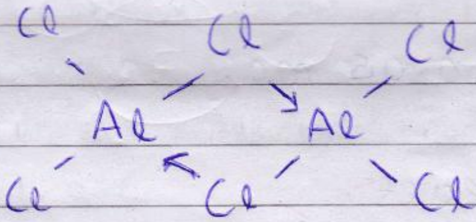
O_3



$:\ddot{O} \leftarrow \ddot{O} = \ddot{O}:$

Al_2Cl_6





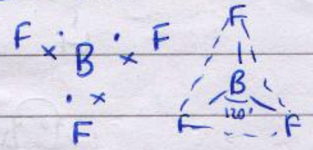
Properties of coordinate covalent ~~bo~~ compounds.

- It is semi polar and non ionizable bond.
- It is directional in nature so coordinate compounds have definite molecular geometry and many show geometrical isomerism.
- Their properties fall between true ionic and true covalent compounds.

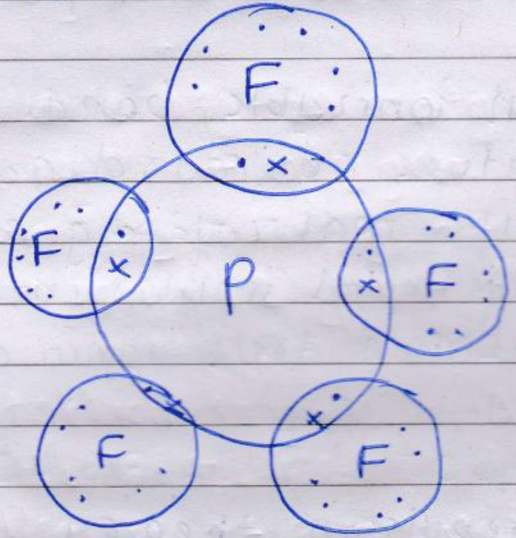
Valence shell electron pair repulsion theory

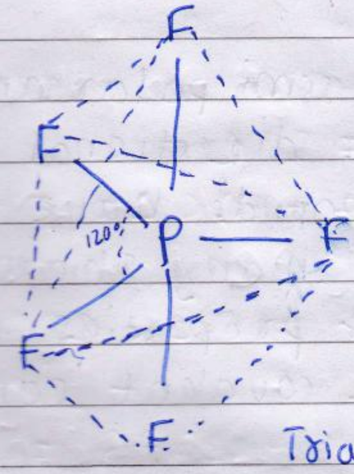
- The shapes of molecules
- In 1874, Dutch chemist Jacobus van't Hoff suggested for the first time that molecules possess a definite, unique, 3d shape whose shape is determined by angles between bonds ~~wh~~ and the bond angles are determined by arrangement of electrons.
- Angles between pairs will depend on number of pairs around the atom.
- Electrons try to be farthest due to repulsion.

~~O = O~~

Molecule	No. of e^- pair around central atom	shape
Cl-Be-Cl	2	Linear
	3	Trigonal planar
$H-\overset{H}{\underset{H}{C}}-H$	4	Tetrahedral

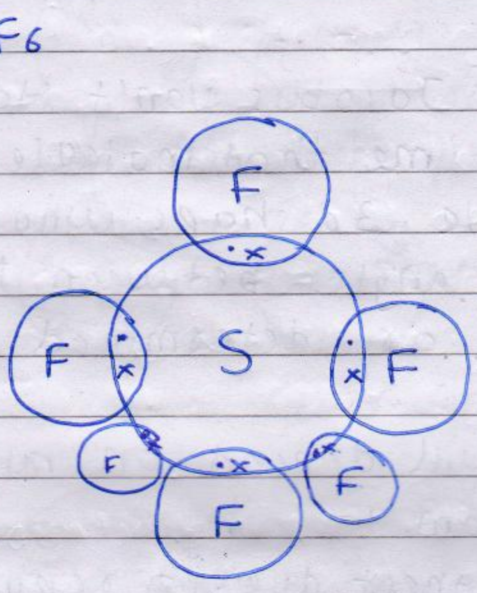
PF_5

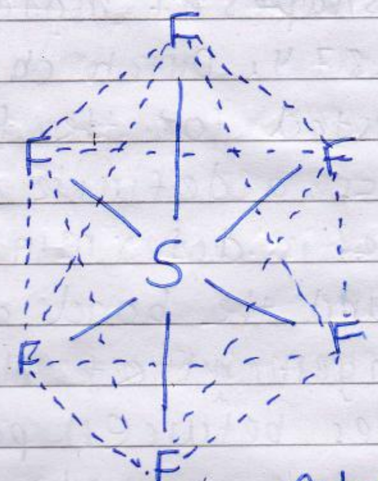




Triangular Bipyramid

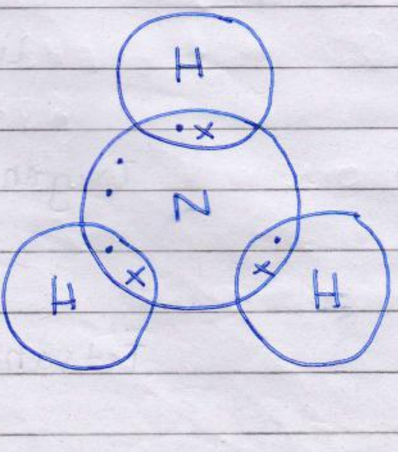
SF_6

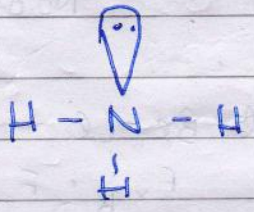


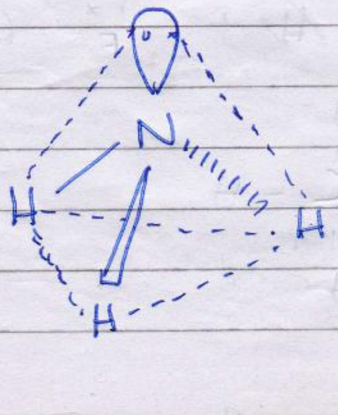


Octahedral / Square bipyramidal

NH_3







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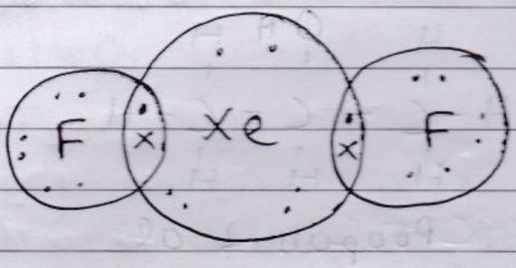
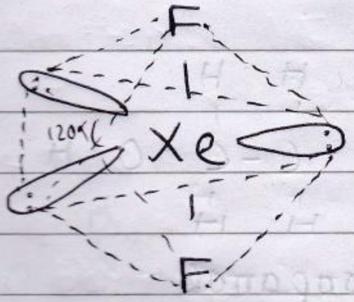
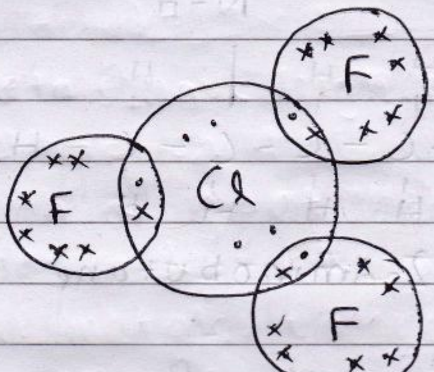
CHEMICAL BONDING...

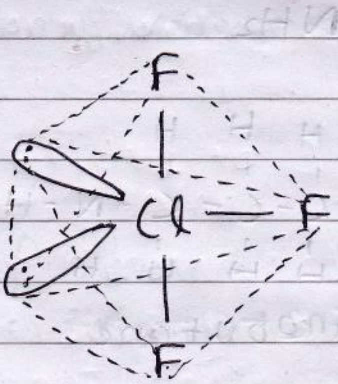
XeF_2





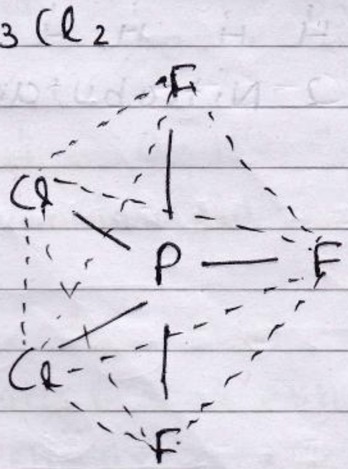
ClF_3





lone pair - lone pair > lone pair - bond pair > bond pair - bond p.

PF_3Cl_2

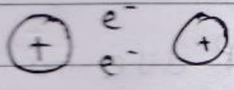


Here chlorine is kept as such because Cl–P–F bond angle is always 120°, but if Cl were on top Cl–P–F bond angle is 90°, which is less. This is case for different atoms.

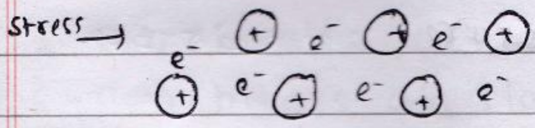
Metallic bonding

Study of metallic ~~properties~~ bonding helps explain metallic properties like conduction of heat and electricity, malleability · ductility.

In a pure metallic element, all atoms are identic
If we look closely at most metallic lattices we find that atoms are arranged similar to argon atoms in solid argon.



A basic metal molecule. Here electrostatic force is neutralized and thus the electrons remain stationary.
This attraction between electron cloud and metal ions is metallic bonding.
They are very strong as they resemble ionic bonding



- When stress is applied the atoms slide, thus creating malleability.
- When electricity (electrons) are supplied to metal one electron, displaces other and thus conduction is possible.
- Delocalized electrons don't belong to any particular atom.
- These delocalized electrons displace each other when potential difference is supplied.

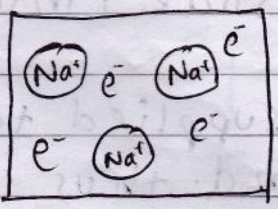
Some attractions

~~Ar~~

Substance	Formula	Attraction	MP
Argon	Ar	Weak dipole	Increases ↓
Hydrogen Chloride	HCl	Dipole - Dipole	
Water	H_2O	H_2 bonding	
Iodine	I_2	Strong dipole	
Sucrose	$C_{12}H_{22}O_{11}$	Strong H^+ bonding	

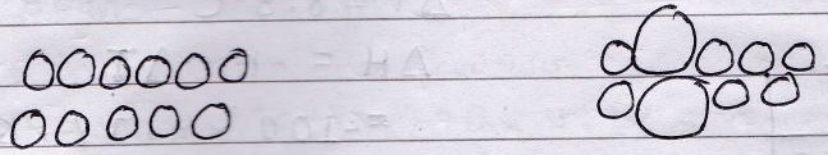
+ Metals have high MP and BP due to great attraction however some covalent compounds with a giant lattice (macromolecules) like diamond tend to have even higher MP & BP.
+ In metals, valence electrons don't belong to any of particular nuclei so they are free and mobile.
+ When photons strike metals, e^- get exicted and jump to higher energy level which is followed by almost immediate reemission of visible light causing lusture

Na-metal



+ Mercury is liquid because it is big atom but loses less electrons thus attraction between Hg^+ & e^- is less making Hg^+ farther apart.

+ They also have high tensile strength due to strong attraction.



pure metal — alloy

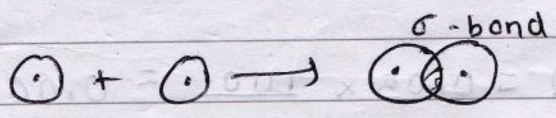
Alloys

+ Metallic substance made by mixing and fusing two or more metals or a metal & non metal to observe qualities like strength, hardness.
+ They generally have superior properties than its parent metal.
+ Resist corrosion.
+ Example Brass and Bronze made of Copper and Zinc in different ratios.

The overlap of orbitals

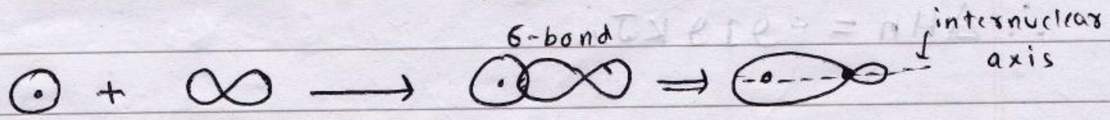
σ - bond

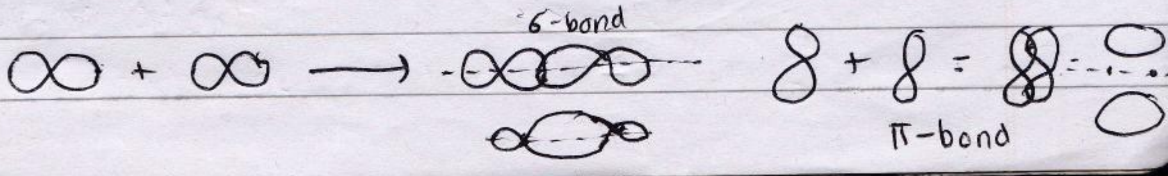
i) s-s orbitals overlap
ii) s-p orbitals overlap
iii) p-p orbitals overlap

i) H_2 : $1s^1$, $1s^1$ 

$H + H \rightarrow H_2$

internuclear distance (bond length)

ii) HF : 

iii) Cl_2 : 

BONDING EXERCISE

1. Solid NaCl and MgO are held by ionic bonds. Write symbols and electronic configuration for each of the ions.

a) i) Na^+ (2,8) ii) Cl^- (2,8,8)
iii) Mg^{2+} (2,8) iv) O^{2-} (2,8)

b) Explain what holds Na^+ and Cl^- ions together in the solid crystal.

→ Na^+ and Cl^- ions are of opposite charges. So, they attract each other. This coloumbic force of attraction is what holds the ions together.

c) NaCl melts at 1074 K ; MgO melts at 3125 K. Both have identical structures. So, why is there such a difference.

→ The magnitude of charges attracting each other is twice that of NaCl in MgO. So, the ionic bond in MgO is stronger than that of NaCl, and thus it requires more heat energy to break its bonds.

2. Write down the electronic structures of the following ions and also in terms of s, p & d orbitals.

a) S^{2-}
2,8,8
$1s^2 2s^2 2p^6 3s^2 3p^6$

b) F^-
2,8
$1s^2 2s^2 2p^6$

c) Ca^{2+}
2, 8, 8
$1s^2 2s^2 2p^6 3s^2 3p^6$

d) Fe^{2+}
~~2, 8, 8, 6~~ 2, 8, 14
$[Ar]\ 3d^6$

e) Cr^{3+}
2, 8, 11
$[Ar]\ 3d^3$

f) Ni^{2+}
2, 8, 16
$[Ar]\ 3d^8$

g) I^-
2, 8, 18, 18, 8
~~[Kr]~~ $[Kr]\ 5s^2 4d^{10} 5p^6$

h) Sn^{2+}
2, 8, 18, 18, 2
$[Kr]$ ~~$5s^2 3d^{10}$~~ $5s^2 4d^{10}$

i) Sn^{4+}
2, 8, 18, 18
$[Kr]\ 4d^{10}$

j) Te^{2-}
2, 8, 18, 18, 8
$[Kr]\ 5s^2 4d^{10} 5p^6$

k) Ag^+
2, 8, 18, 18
$[Kr]\ 4d^{10}$

3. Write formulae for following ionic compounds.

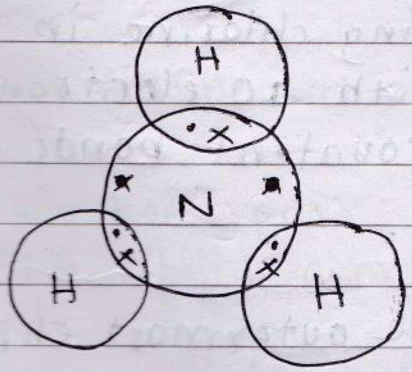
a) sodium sulphide → Na_2S.
b) caesium bromide → $CsBr$
c) Magnesium nitride → Mg_3N_2
d) chromium (III) oxide → Cr_2O_3
e) tin (II) chloride → $SnCl_2$
f) nickel (II) selenide → $NiSe$
g) aluminium oxide → Al_2O_3
h) copper (I) iodide → CuI

4. A covalent bond consists of a pair of shared electrons between two atoms. How does this shared pair hold the two atoms together?

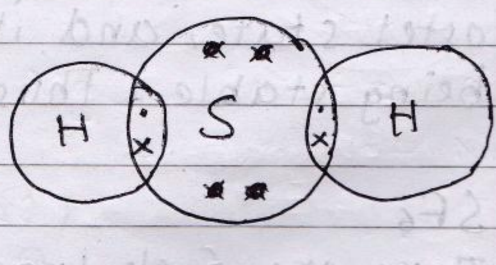
→ The pair of shared electrons is required for both atoms to attain stable electronic configuration. However, except for in diatomic molecules, one atom attracts the electrons with more force, causing them to shift slighty. However, the other atom also doesn't want to let go of the electrons due to it being in stability. This creates ~~an~~ a stable state, where both the atoms are sharing the electrons, thus holding them together.

5. Draw dots and crosses diagrams for:

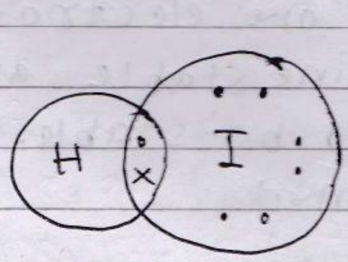
(i) NH_3



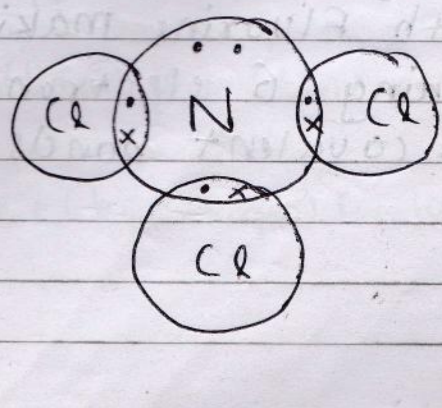
(ii) H_2S



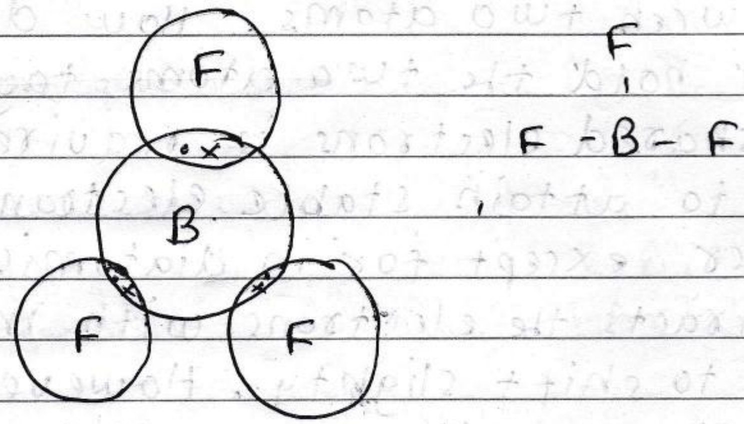
(iii) HI



(iv) NCl_3



(v) BF_3



6. Explain how the covalent bonding in the given compounds arises.

a) PCl_5

→ There are 5 electrons in the outermost shell of Phosphorus. However it can hold 5 more electrons and extend to $4s^2$ orbital to be stable with configuration $4s^2$. Thus, it shares one electron each with chlorine, leaving chlorine in stable octet state, and itself with 10 electrons and being stable. Thus five covalent bonds are formed.

b) SF_6

→ There are 6 electrons in the outermost shell of Sulfur. But it can hold 6 more electrons due to extended d subshell, to attain stable configuration. Thus, it shares one electron each with Fluorine making Fluorine stable and gaining 6 electrons itself to be stable. Thus six covalent bonds are formed.

c) XeF_4

→ Xenon initially has 8 electrons in its outermost shell. However, it also can use other unused subshells in 4^{th} shell, and thus hold more than 8 electrons. So, it will form four covalent bonds to combine with four fluorine atoms, thus having a total of 12 electrons itself.

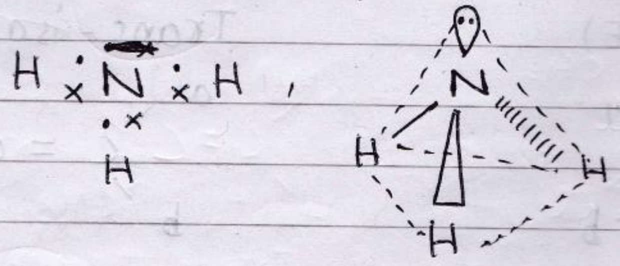
7. Xenon forms several compounds, but neon dosen't. Suggest why it is so?

→ It is so because neon has a configuration of $1s^2 2s^2 2p^6$ which means in both of its shells all the subshells are used up. So it ~~dosen't~~ cannot utilize any other subshell to accomodate electrons to form bonds. However Xenon can utilize unused subshells in 4^{th} shell to accomodate electrons.

8. Please fill the table shown.

→ AX_4 - bond pair

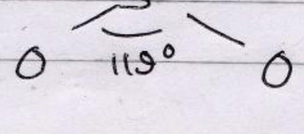
→ AX_3E - (bond pair, lone pair)



→ AX_2E_2 - trigonal planar, (bond pair, lone pair)

→ AX_3 - bond pair

→ AX_2E - trigonal planar, (bond pair, lone pair)



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METALLIC BONDING EXERCISE

1. K, Ca and Sc follow each other in periodic table. Their MP and BP is given.

	Potassium	Calcium	Scandium
MP	336 K	1112 K	1814 K
BP	1033 K	1757 K	3104 K

a) Write their electronic configuration.
K - $1s^2 2s^2 2p^6 3s^2 3p^6 4s^1$
Ca - $1s^2 2s^2 2p^6 3s^2 3p^6 4s^2$
Sc - $1s^2 2s^2 2p^6 3s^2 3p^6 4s^2\ 3d^1$

b) Each of them are held together by metallic bonding. Explain the nature of it using K as example.
→ Metallic bonding can be explained using a model. In this model, the atoms of metals lose electrons to form an electron soup. The positively charged ions attract the electrons from this soup.
In potassium, the K^+ ions attract an electron due to their charges. However, the same e^- is attracted by other K^+ ions with equal magnitude leaving the electron still and the atoms to bond with this attractive force.

c) The metallic bonding isn't completely broken until a metal boils. State what happens to the bond strength as you go from K to Sc.
→ Since, the amount of energy required to break the bond (BP) increases, the strength of bond also increases from K to Sc.

d) Explain the effect from part c.

→ The energy required to break the attractioen between the electrons and ions in a metallic lattice increases from K to Sc. This is because as more energy is supplied, the atoms vibrate more rapidly thus making them separate.

1a) Mg has higher MP and BP than Na. This can be explained in terms of electronic structure, packing and atomic radii. Explain why they cause the MP and BP of Mg to be higher.

Na - $1s^2 2s^2 2p^6 3s^1$

Mg - $1s^2 2s^2 2p^6 3s^2$

In Mg, the 3s orbital is compeletly filled. Thus, this causes the electrons to be attracted by more force which in turn decreases the atomic radius of Mg than of Na. So, the electrostatic force of attraction between atom and e^- in Mg is more than that of Na. Also, in the lattice two Mg atoms share $2e^-$ while in Na, two atoms only share $1e^-$. Thus the bond force for Na is weaker than that of Mg. This in turn causes MP & BP of Mg to be higher.

2. Explain why metals are good conductors of electri-city and heat.

→ In a metallic lattice, there are some delocalized/ free electrons which don't belong to any particular atom. When a potential difference is supplied to a metal, the free e^- can accelerate and help transfere electricty from one point to another fast, making them good conductors of electricity. Similarly when heat is provided, the same free electrons move due to kinetic energy and can tranfer this heat between atoms easily making them good conductors of heat as well.

3a) Pure metals are malleable and ductile. Explain the two words.

→ Malleable means very easy to turn into thin sheets ~~by~~ and ductile means very easy to draw into wires.

b) If a metal is subjected to small stress it will return to its original shape when stress is removed however larger stress causes perman-ent deformation. Explain this using diagrams.

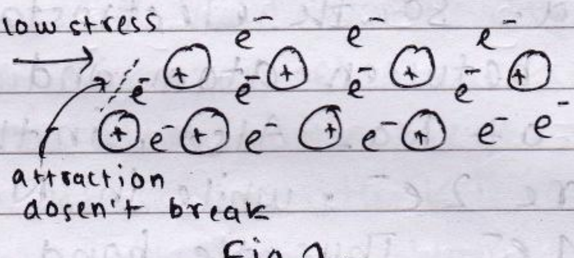

Fig 1.

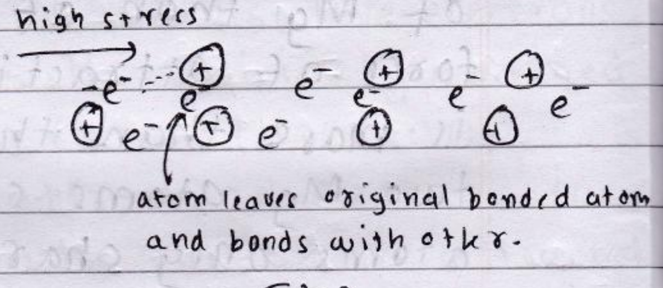

Fig2.

In Fig 1, the low strese isin't enough to break the metallic bonding between atoms so it dosen't undergo permanent deformation, however in Fig 2, the high stress is enough to break the bonding between two atoms and bond the atoms with new atoms. Thus even when stress is removed, deformation remains as the atoms have bonded with new atoms.

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c) When a blacksmith works with metals, they are heated, hammered and re heated. Explain what's happening in terms of ~~electro~~ structure of metal.

→ The metals need to be reshaped into desirable shape. When they are heated, their bonding in lattice becomes weaker, so it is easier to deform them with ~~sm~~ lesser stress. But the metals lose heat to their surrounding fast so they need to be reheated to work with as the lattice bonding ~~becaus~~ becomes stronger again and it needs to ~~me~~ be made weaker.

d) Why is brass harder than copper and zinc, its constituent metals.

→ Brass is harder than its constituent metals as the lattice of brass consists of atoms of both copper and zinc, which are of different sizes. This uneveness is lattice makes it more difficult to slide the lattic layers, thus increasing hardness.

CHEMICAL BONDING

Hybridization

- It is the mixing of orbitals.
- Orbitals must be in same shell to undergo hybridization

```
      H
      |
  H - C - H
      |
      H
```

All four covalent bonds are similar as there is same bond length, bond energy, same distance from each other

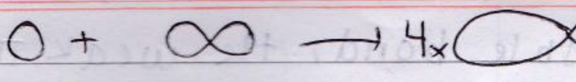
$C - 1s^2 \underbrace{2s^2 2p^2}$ 2nd shell electrons can hybridize but not first.
$H - 1s^2$
So, there is overlapping of p-orbital of C and s-orbital of H.

C - [↑↓] [↑↓] [↑ | ↑ |] Ground state configuration can't form 4 covalent bonds as there are only 2-unpaired electrons.

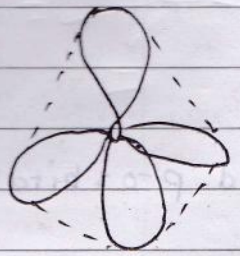
C - [↑↓] [↑] [↑ | ↑ | ↑] Exicted state configuration when one e^- from 2s orbital goes to 2p. But still the bonds won't be similar as 2s orbitals ~~have~~ has less energy than 2p.
(2s) (2p)

- So, to make the bonds similar the 2s and 2p orbitals combine to form sp^3 hybridization

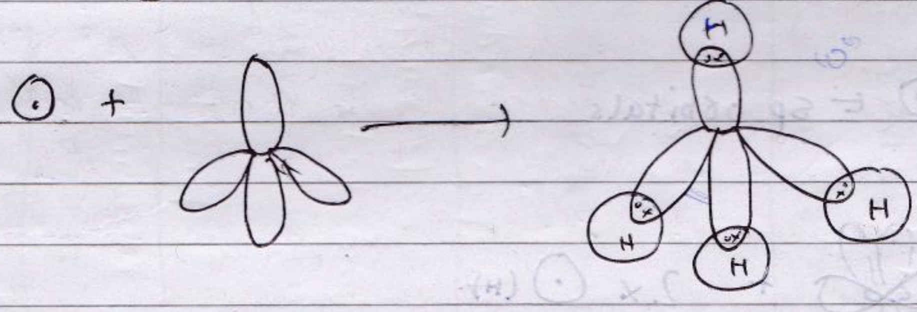
[↑↓] [↑ | ↑ | ↑ | ↑] This orbital is called sp^3 orbital
(2s) $\underbrace{\quad}_{2p}$

 sp^3

Orientation of orbitals will be. (in methane)



Tetrahedral shape and 109°28' angle.
Degenerate orbitals (with same energy)

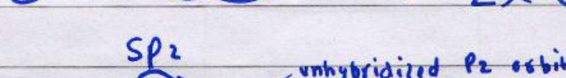
4 

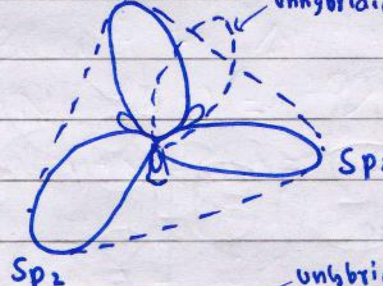
$C - 1s^2 2s^2 2p^2$

i) C – [↑↓] [↑↓] [↑ | ↑ |]

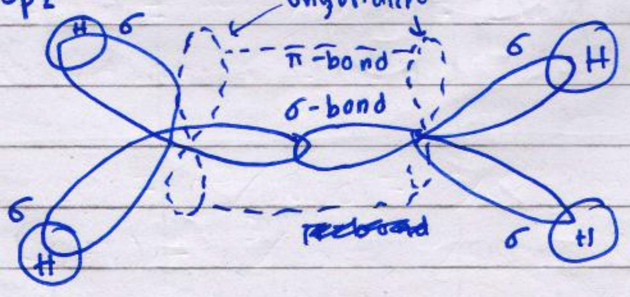
ii) C – [↑↓] [↑] [↑ | ↑ | ↑]

iii) C – [↑↓] [↑ | ↑ | ↑] [↑] ← Unhybridized ~~subshe~~ orbital
sp^2 sp^2 sp^2

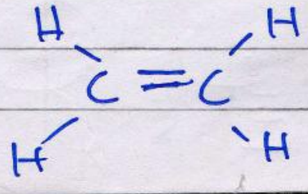
iv)  sp^2



Triangular planar shape with 120° angle between them.



5-sigma bonds.

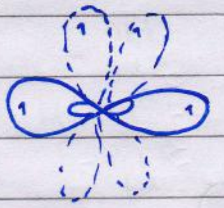


While rotating this double bond, the weak π-bond will break, disintegrating the molecule. Thus, double bond has restricted rotation.

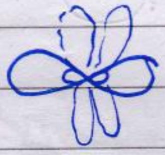
i) C - $\boxed{\uparrow\downarrow}$ $\boxed{\uparrow\downarrow}$ $\boxed{\uparrow \mid \uparrow \mid }$

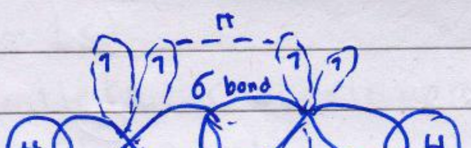
ii) C - $\boxed{\uparrow\downarrow}$ $\boxed{\uparrow}$ $\boxed{\uparrow \mid \uparrow \mid \uparrow}$

iii) C - $\boxed{\uparrow\downarrow}$ $\underset{sp\quad sp}{\boxed{\uparrow \mid \uparrow}}$ $\boxed{\uparrow \mid \uparrow}$ ← Unhybridized p-orbitals



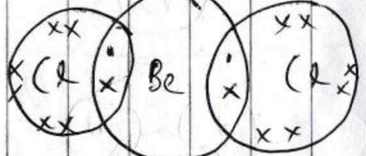
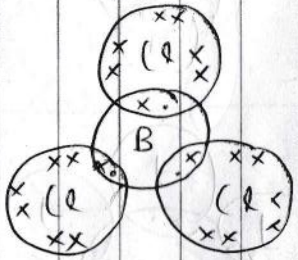
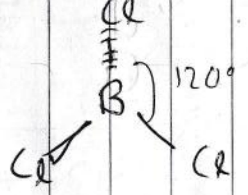
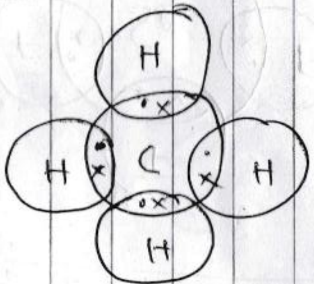
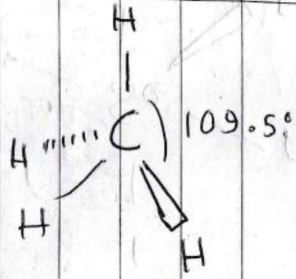
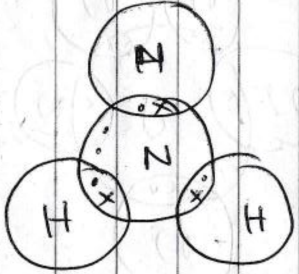
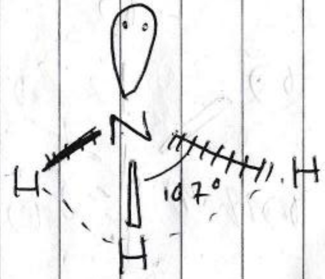
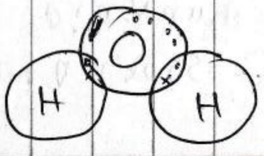
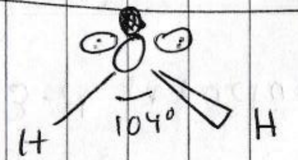
← sp orbitals

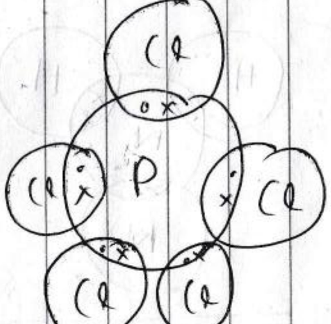
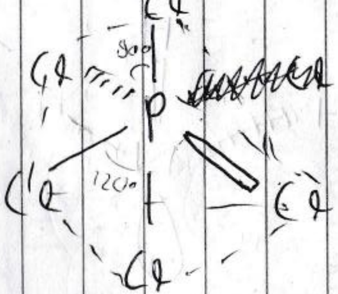
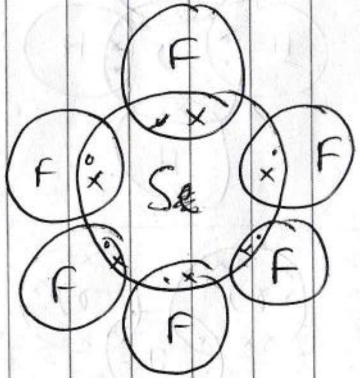
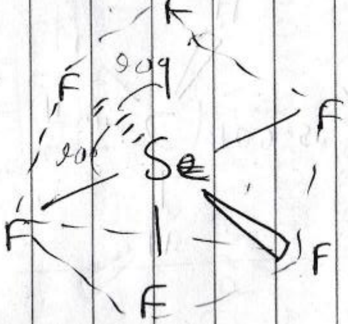
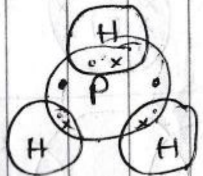
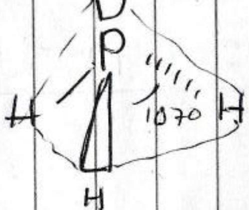
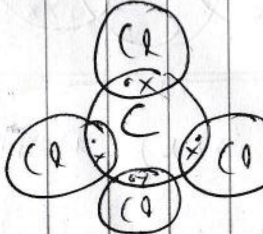
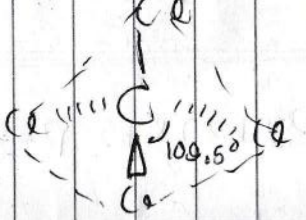
iv) 2x  + 2x ⊙ (H)



H−C≡C−H

180°

Molecule	Dot and Cross Diagram	No. of bond e^- pair around central atom	No. of lone pair	Shape	3-d structure	Hybrid-isation
$BeCl_2$		2	0	$Cl \overset{180°}{—Be—} Cl$ Linear	$Cl - \overset{180°}{Be} - Cl$	Sp
BCl_3		3	0	Trigonal Planar		Sp^2
CH_4		4	0	Tetrahedral		Sp^3
NH_3		3	1	Trigonal Pyramid		Sp^3
H_2O		2	~~6~~ ~~8~~ 2	bent, angular V-shaped		Sp^3

Molecule	Dot & Cross Diagram	No. of e^- pairs around central atom	No. of lone pair	Shape	3-d structure	Hybridization
PCl_5		5	0	Trigonal Bypyramid		sp^3d
$S\cancel{e}F_6$		6	0	Se Bypyramidal ($\cancel{D}$ Bitetrahedral) Octahedral		sp^3d^2
PH_3		4	1	Trigonal pyramid		sp^3d
CCl_4		4	0	Tetrahedral		sp^3

Date ______
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Hybridization in water

$O - 2,6$, $1s^2, 2s^2 2p^4$

O - [↑↓] [↑↓] [↑↓ | ↑ | ↑] - Ground state

O - [↑↓] [↑↓ | ↑↓ | ↑ | ↑]

S P sp^3

sp^3 hybridization

$Be - 2,2$, $1s^2 2s^2$

Be - [↑↓] [↑↓] ~~- Ground state~~ [| |] - Ground state

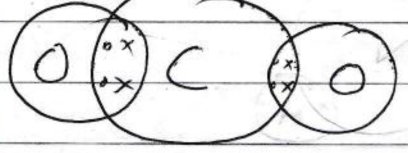
Be - [↑↓] [↑] [↑ | |]

Be - [↑↓] [↑ | ↑] [|] - Hybridized

sp

Sp hybridization

CO_2

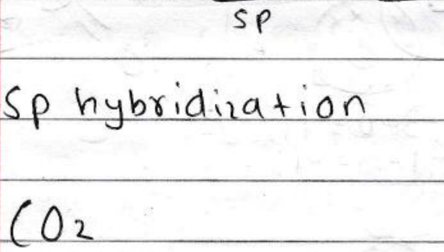


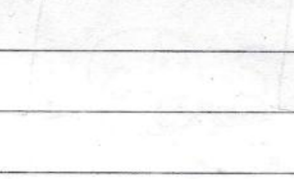
180°

O = C = O Sp^2

Linear

H_2S



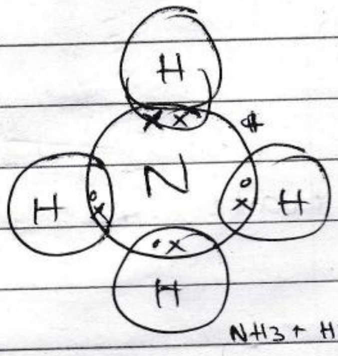


V-shaped

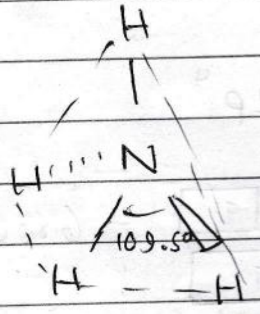
4 8th Dec 2021

Date ____ Page ____

NH_4^+



$NH_3 + H^+$

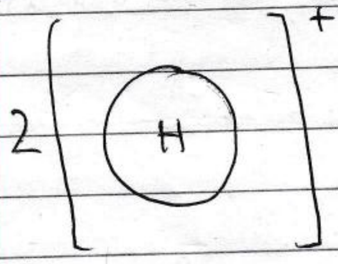


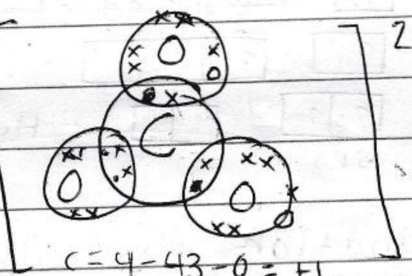
Sp^3

Tetrahedral

H_2CO_3

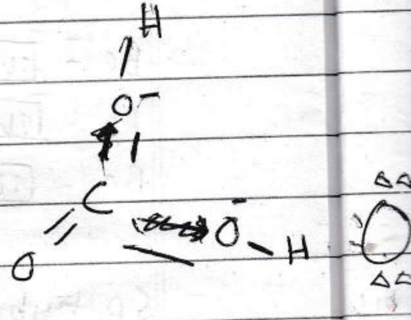
~~2H~~ $H + H + CO_3$



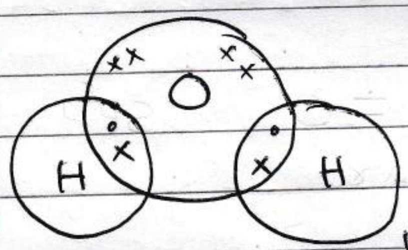


$C = 4 - 43 - 0 = +1$

$O = 6 - 1 - 6 = -1$

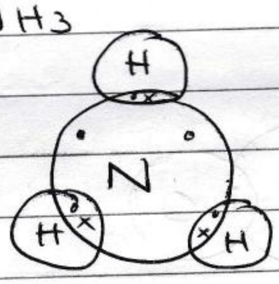


H_2O

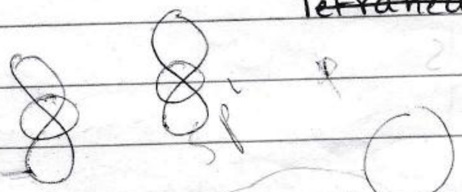


V-shaped

NH_3

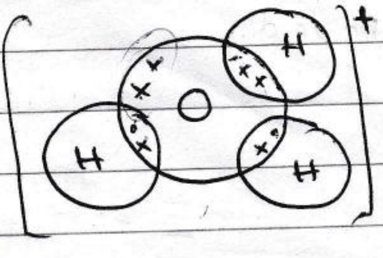


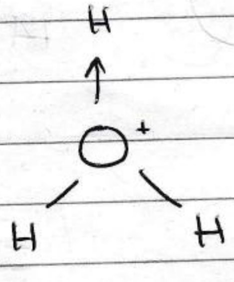
Trigonal planar
~~Tetrahedral~~



H_3O^+

$H^+ + H_2O \longrightarrow H_3O^+$





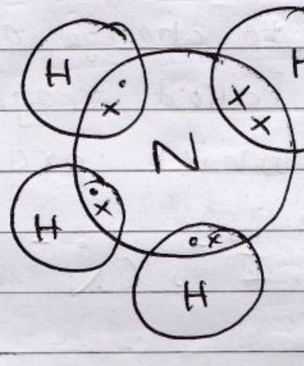
O has +ve charge as it has donated one e^- to datively bonded H

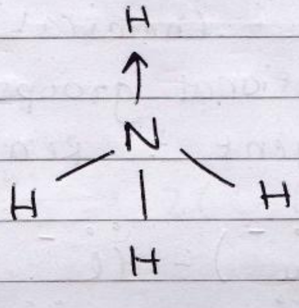
Triangular pyramid.

NH_4^+

Formalcharge = Valence e^- − bonded e^- − lone pair e^-

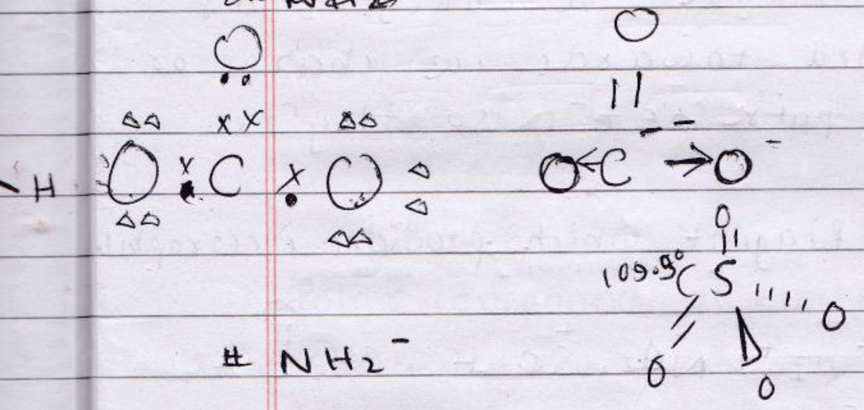
$H^+ + NH_3 \longrightarrow NH_4^+$



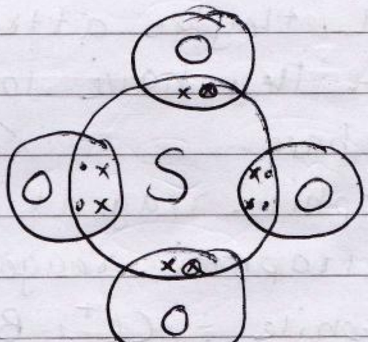


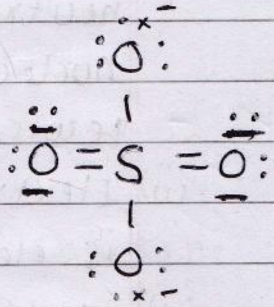
$POCl_3$
NO_2
NO_2^-
NO_3^-

~~NH_2^-~~



SO_4^{--}

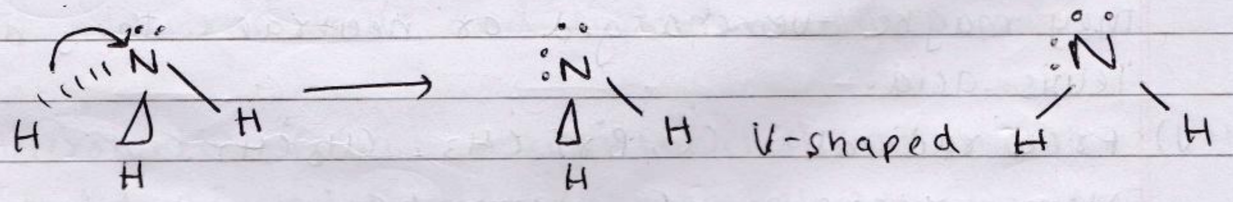




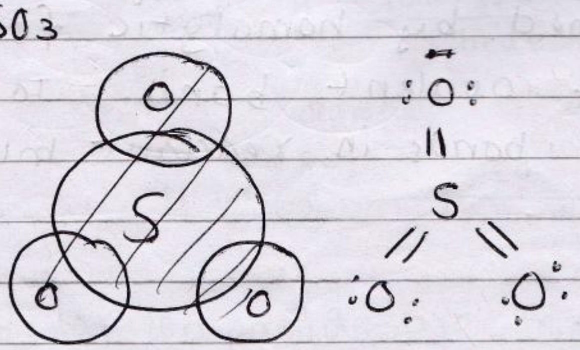
S formal charge = 6 − 6 − 0 = 0

NH_2^-

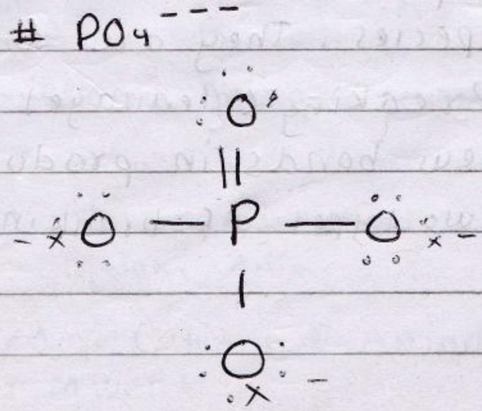
NH_3 ~~$-H^+$~~ $\longrightarrow NH_2^- + H^+$



SO_3



PO_4^{---}



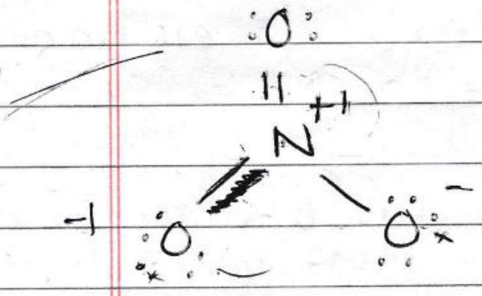
H_2O Date ______ Page ______

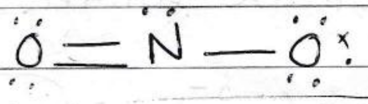
HYBRIDIZATION . . . NO_2

NO_2^-

\# NO_3^- \# NO_2^-

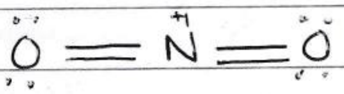
NO_3^-

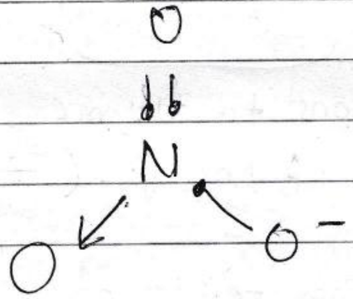




$1s^2 \boxed{2s^2\ 2p_x^2\ 2p_y^2\ 2p_z^2}$

\# NO_2^+ (nitronium)





FC = 5 - 3 - 2 = 0

Electronegativity and bonding

Electronegativity of an atom is the measure of ability of ~~sh~~ an atom to attract a shared pair of e^-s towards itself. It increases across a period and decreases down the group.

2.1 H						
1.0 Li	1.5 Be	2.0 B	2.5 C	3.0 N	3.5 O	4.0 F
Na	Mg	Al	Si	P	S	Cl

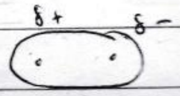
Linus pauling studied ~~the pauling e~~ electronegativity & introduced pauling electronegativity scale.

Polar bond

Covalent bond in which the bonding electrons are distributed unequally due to difference in electroneg-ativity of the bonded atoms.

$^{\delta+}H$ —— $Cl^{\delta-}$

2.1   3.5



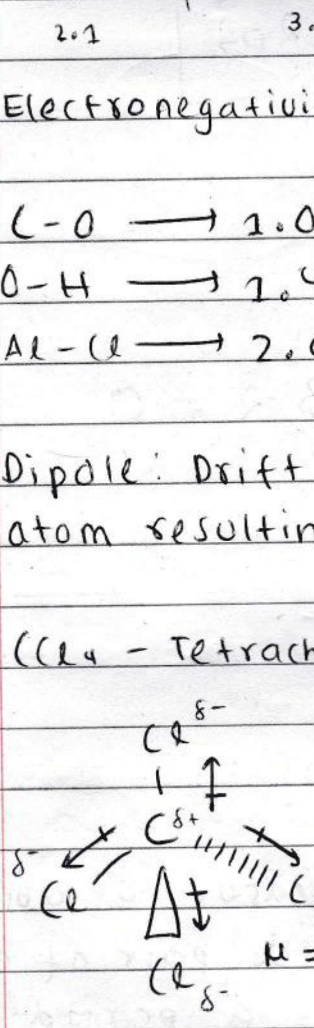
Centre of +ve & -ve charge don't coincide

Electronegativity difference = 3.5 − 2.1 = 1.4 ..

C−O ⟶ 1.0
O−H ⟶ 1.4
Al−Cl ⟶ 2.0

Dipole: Drift of bonded electrons to the more electronegative atom resulting in separation of charge. ($+\!\!\longrightarrow$) (μ)

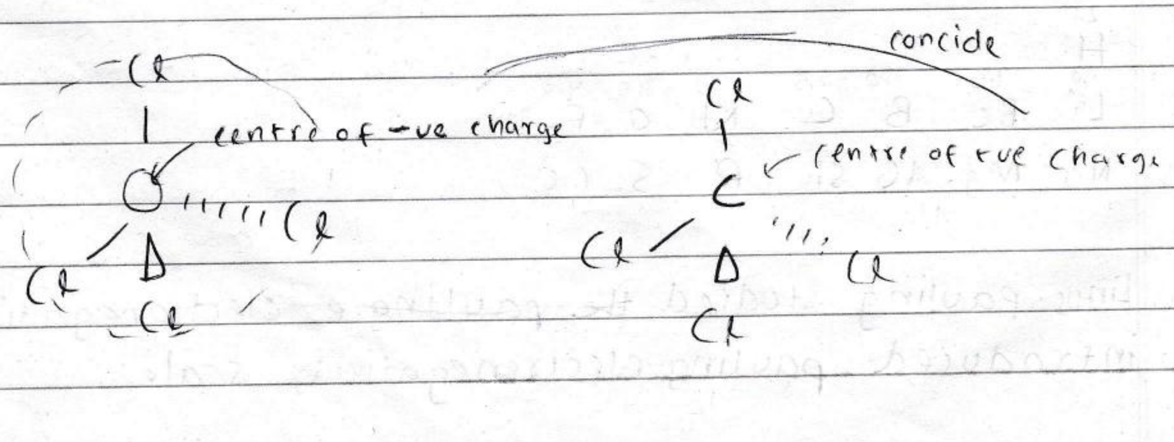
CCl_4 − Tetrachloromethane (Carbon tetrachloride)



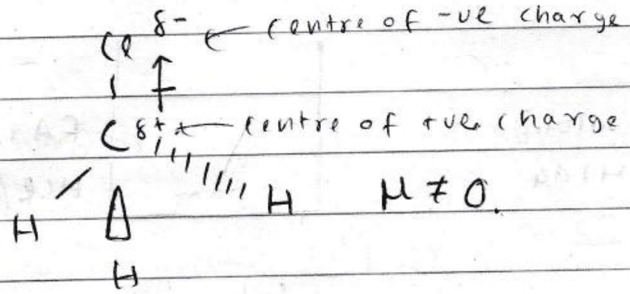
Net dipole ~~mou~~movement : Vector sum of the dipole moment of the individual bonds in the molecule.

$\mu = 0$ Centre of +ve and −ve charge concide.

Dipole movement: Measure of the polarity of the molecule.



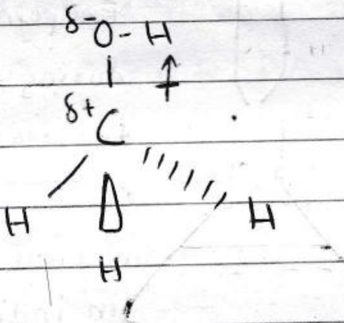
CH_3Cl - Chloromethane



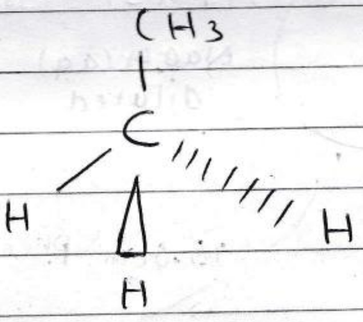
Polar molecule.

$\mu \neq 0$.

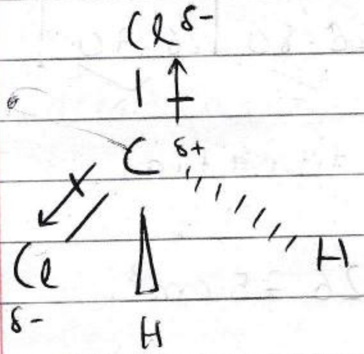
CH_3OH, C_2H_6, CH_2Cl_2, SF_4



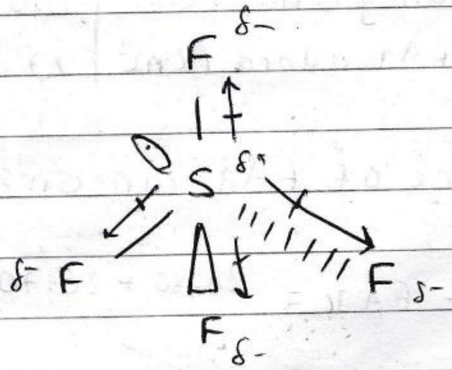
Tetrahedral
Polar



Tetrahedral
Non-polar



Tetrahedral
Polar



~~Tetrahedral~~ Trigonal pyramid
~~Non~~ polar