CHM 101 SMART SHOP (Day 1) 1 (3 UNITS)

CHEMICAL BONDING (LEWIS CONCEPT OR THEORY)

A chemical bond is the force that acts between two or more atoms to hold them together as a stable molecule.

There are various types of chemical bonds. These include: Ionic or Electrovalent bonding, covalent bonding / Coordinate covalent bonding, Metallic bonding, Hydrogen bonding etc.

Valency is the combining power of an element eg Hydrogen has a valency of 1, Mg, and Ca have a valency of 2, Al has a valency of 3, NO_3^- has an electrovalency of -1, SO_4^{2-}, CO_3^{2-}, PO_4^{3-} have electrovalencies of -2, -2 and -3 respectively etc.

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Valence Electrons are the electrons on the outermost shell of an atom. eg Na: 2,8,1 or $1s^2 2s^2 2p^6 3s^1$, Mg: 2,8,2 or $1s^2 2s^2 2p^6 3s^2$, Al: 2,8,3 or $1s^2 2s^2 2p^6 3s^2 3p^1$

These examples show that the valence electrons of Na, Mg and Al are 1, 2 and 3 respectively. Similarly, the valence electrons for Cl, O, S etc are 7, 6 and 6 respectively. Thus the number of valence electrons can be derived from the group number.

The valence electrons involved in bonding are called bonding electrons while the remaining valence electrons that do not partake in bonding are referred to as non-bonding electrons. Eg

$:\ddot{Cl}(:)\ddot{Cl}:$ → Non-bonding electrons

1 - This type of bonding involves the transfer of electrons from one atom to ~~another~~ another.

- For example, consider an atom A with one (1) valence electron and another atom B with seven (7) valence electrons. A will transfer the one (1) valence electron to B to form an electrovalent or ionic bond.

$$A^{\times} \; + \; \ddot{B}: \longrightarrow A^{+} + :\ddot{B}:^{-} \quad \text{or} \quad \underset{\text{cation}}{A^{+}} + \underset{\text{Anion}}{B^{-}}$$

$$A^{+} + B^{-} \longrightarrow \underbrace{A^{+} \rightleftharpoons B^{-}}_{\text{Ionic bond}} \quad \text{or} \quad A^{+}B^{-}$$

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The electrostatic attraction between the cation (+) and anion (−) produced by electron-transfer constitutes an ionic or electrovalent bond.

Compounds containing such bonds are called Electrovalent or ionic compounds.

- For an ionic bond to be formed, the following conditions must be met:

i) atom A should possess 1, 2 or 3 valence electrons while atom B should have 5, 6 or 7 valence electrons. The elements in Group IA, IIA and IIIA satisfy this condition for atom A and those of groups VA, VIA and VIIA satisfy this condition for atom B.

ii) the ionization energy (IE) of atom A should be low,
- the electron affinity of atom B should be high.
- the electrostatic attraction between A^+ and B^- in the solid compound should also be high.

Factors Governing The Formation of An Ionic Bond

a) Ionisation Energy (IE): This is the energy needed to ~~remov~~ remove an electron from an atom of the element in the gaseous state, $M_{(g)} + IE \rightarrow M^{+}_{(g)} + e^{-}$. The lower the ionization energy, the greater will be the tendency of the metal atom of change ~~of~~ into cation and the greater will be the ease of formation of an ionic ...

This explains why alkali and alkaline earth metals form ionic bonds easily. The alkali metals form ionic bonds with more ease.

b) Electron Affinity (EA): This is the amount of energy released when an electron is added to a gaseous atom to form an anion. ($X_{(g)} + e^- \longrightarrow X^-_{(g)} + EA$.

The higher the electron affinity, the more the energy released and the more stable the anion will be. Group VIA and VIIA have higher electron affinity and higher tendency to form ionic bonds.

c) Lattice Energy: The amount of energy released when one mole of an ionic compound is formed from its cations and anions. The greater the lattice energy, the greater the strength of ionic bond. The value of lattice energy depends on two factors: size of the ions and charge on ions.

Examples of Ionic Compounds: These include NaCl, KCl, $MgCl_2$, MgO, CaO, Al_2O_3 etc.

Characteristics of Electrovalent/Ionic Compounds.
- solids at room temperature
- High Melting and Boiling points
- Hard and Brittle
- Soluble in water
- Conduct electricity in fuse/molten state or solution (Electrolytes)
- DO not exhibit isomerism
- Ionic reactions are fast.

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COVALENT BO[illegible] – [illegible] SHARING BOND[illegible]

This involves sharing of electron pairs between two atoms (A and B). In ~~this case~~ simple covalent bond both atoms donate electrons to be shared.

E.g A+ x. B →  Shared electron pair or A — B covalent

The shared pair is indicated by a dash (—) between the two bonded atoms. The compounds containing covalent bonds are called covalent compounds.

* A covalent bond may be defined as the attractive force between atoms created by sharing of an electron pair.

– The conditions favourable for the formation of covalent bonds are:

(i) Number of valence electrons. Each ~~of~~ of the atoms A and B should have 5, 6 or 7 valence electrons ~~so that both~~ will achieve stable octet by sharing 3, 2 or 1 electron-pair. H has one electron in the valence shell and attains duplet. The non-metals of groups VA, VIA and VIIA satisfy this condition.

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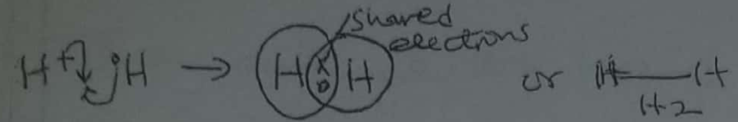
(ii) Equal Electronegativity: Atom A will not transfer electrons to atom B if ~~both~~ have the same electronegativity values and hence electron sharing will take place. This is very possible if ~~both~~ atoms are of the same element.

(iii) Equal sharing of Electrons: Atoms A and B should have equal (or almost equal) electron affinities so that they can attract bonding electron pairs equally. Thus, equal sharing of electron pairs will form non-polar covalent bond.

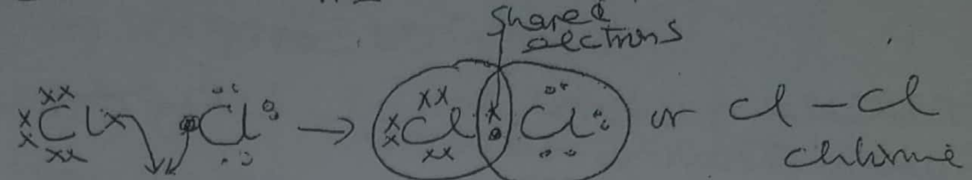
Examples of covalent compounds

H_2, Cl_2, H_2O, NH_3, CH_4, $\underbrace{O_2, N_2, CO_2}$

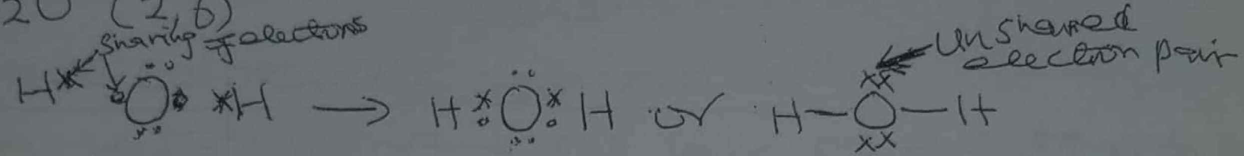
Multiple covalent cpds



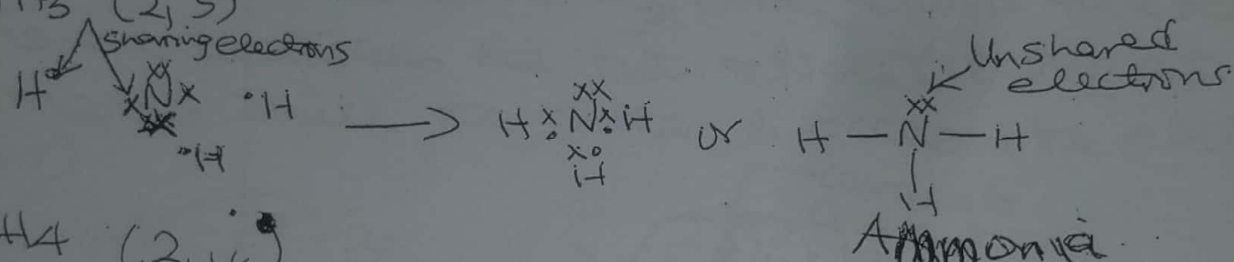
Cl_2 (2,8,7)



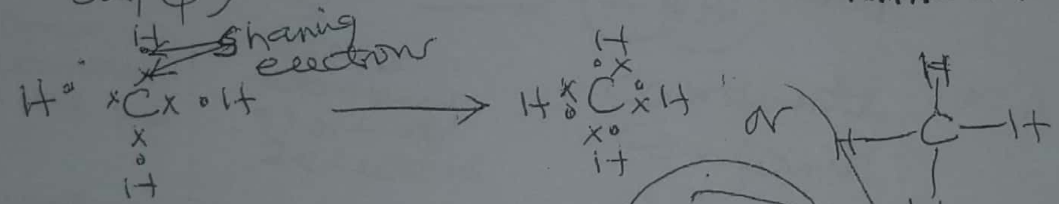
H_2O (2,6)



NH_3 (2,5)



CH_4 (2,4)



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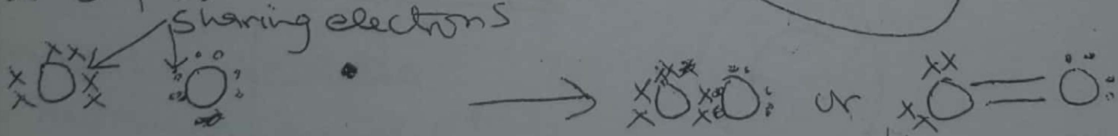
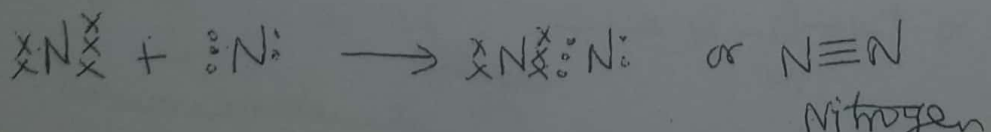
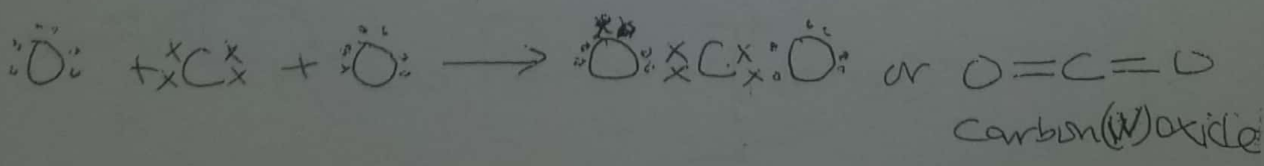
O_2 (2,6)



N_2 (2,5)



CO_2 (2,4)

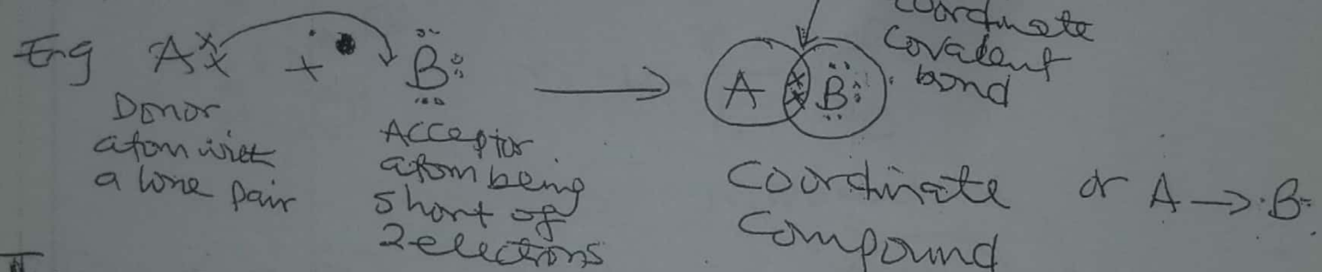


Properties of covalent compounds:

- Gases, liquids or solids at room temperature.
- Low melting and boiling points.
- They are neither hard nor brittle
- Soluble in organic solvents
- Non electrolytes (Do not conduct electricity in solution or molten/fused states)
- Exhibit isomerism
- Molecular reactions.

DATIVE/CO-ORDINATE COVALENT BOND

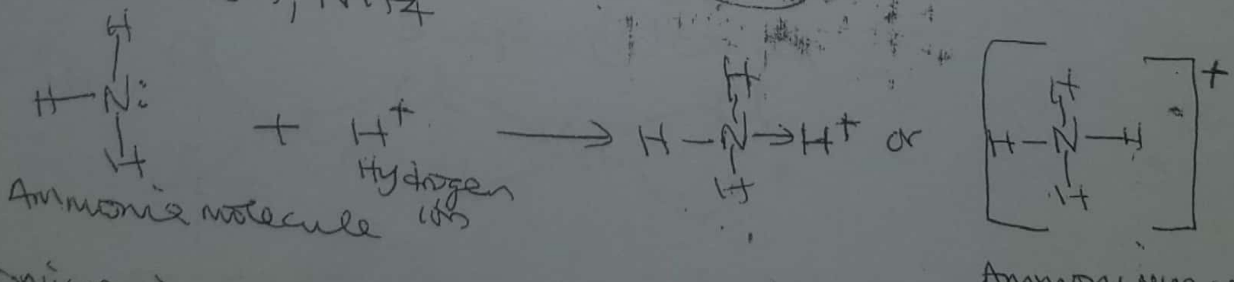
In dative bonding the pair of electrons to be shared are donated by one of the atoms and then shared between the two atoms (i.e one atom, A donates while the other, B accepts the electron pair).

Eg 

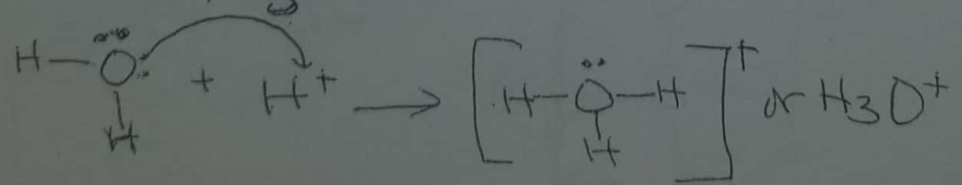
* The molecule or ion that contains the donor atom is called the ligand.

Examples include:

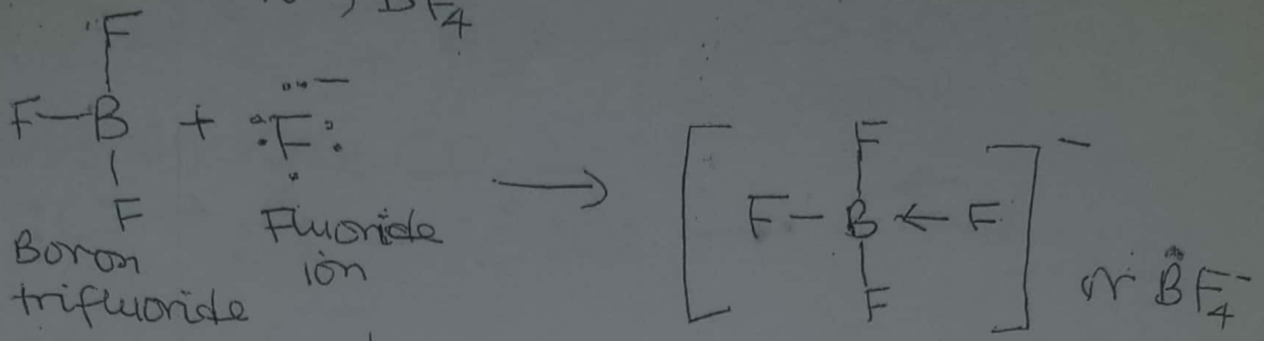
Ammonium ion, NH_4^+



Hydronium ion, H_3O^+



Fluoroborate ion, BF_4^-

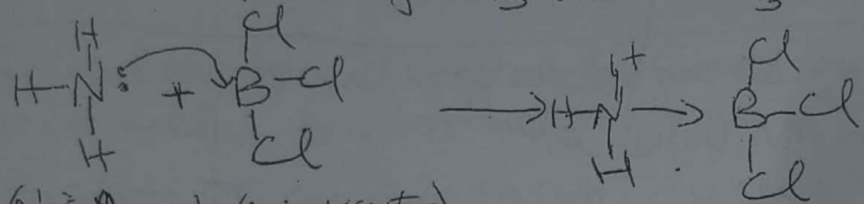


Boron trifluoride (B acts as Acceptor)

Fluoride ion

Fluoroborate ion

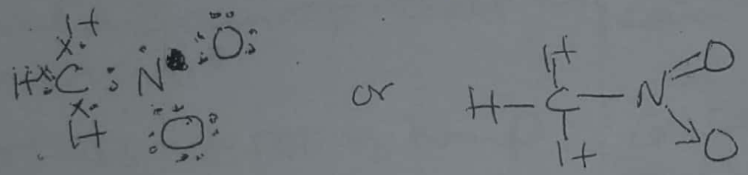
Addition compound of NH_3 with BCl_3



(N is Donor) (B is Acceptor)

Addition compound.

Nitromethane, CH_3NO_2



Nitromethane.

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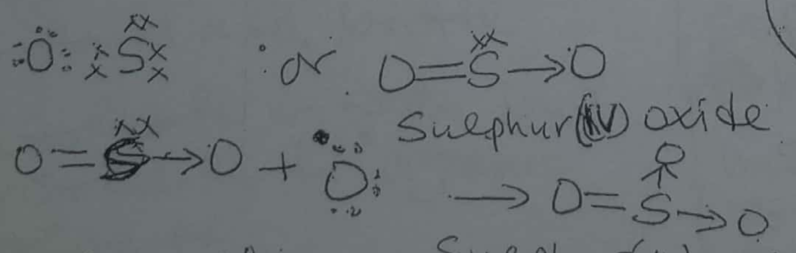
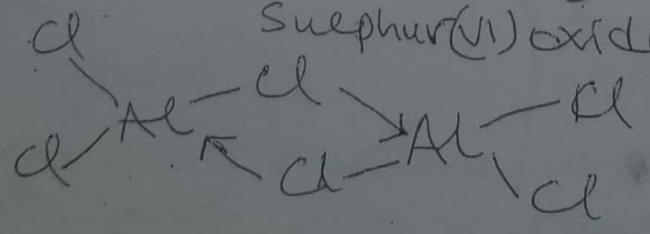
SO_2 and SO_3



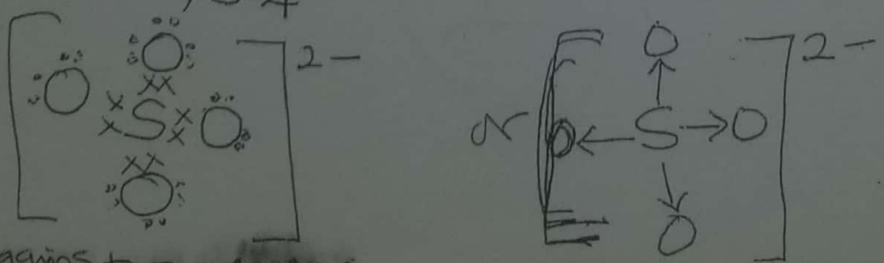
Sulphur(IV) oxide

Sulphur(VI) oxide.

Al_2Cl_6



Sulphate ion, SO_4^{2-}



S gains two electrons from metal atom and

Sulphate ion

Ozone, O_3

$$O=O + O \rightarrow O=O\rightarrow O$$

ozone

Carbon(II) oxide, CO

$$C + O \rightarrow C:O \rightarrow C\overset{\leftarrow}{=}O$$

Comparison between Ionic and Covalent Bonds

Ionic or Electrovalent Bond	Covalent Bond
- Formed by transfer of electrons from a metal to a non-metal	Formed by sharing of electrons b/w non metal atoms
- Consists of an electrostatic force b/w (+) and (−) ions	Consists of shared pairs of electrons b/w atoms.
- Non-rigid, ~~and~~ non-directional and cannot result in isomerism.	Rigid ~~and~~ directional and can cause stereoisomers
- Solids at room temp	Gases, liquids or ~~soft~~ solids
- High melting and boiling points	Low melting and boiling points
- Hard and brittle	Soft and much easily broken
- Soluble in water but insoluble in organic solvents	Insoluble in water but soluble in organic solvents
- Electrolytes	Non-electrolytes
- Undergo ionic reactions that are fast	Undergo molecular reactions that are slow

Hydrogen bonding is the electrostatic attrac[tion] between an H-atom covalently bonded to a hi[ghly] electronegative atom X and a lone pair of ele[ctrons] of X in ~~another~~ molecule. It is often repre[sen]ted by dotted lines.

$$X^{\delta-} - H^{\delta+} + :X^{\delta-} - H^{\delta+} \longrightarrow :X^{\delta-} - H^{\delta+} \cdots :X^{\delta-} - H^{\delta+}$$

Electrostatic attrac[tion]

or ~~X—H—X~~ Hydrogen bond

$X - H \cdots X - H$

N/B (i) Only O, N and F with very high electro negativity and small ~~size~~ atomic sizes are capable of forming H-bonds.

(ii) H-bond is longer and much weaker than normal covalent bonds

(iii) H-bond like a covalent bond has a preferred bonding direction.

Examples:

HF ; $\overset{\delta+}{H} - \overset{\delta-}{F} \cdots \overset{\delta+}{H} - \overset{\delta-}{F} \cdots \overset{\delta+}{H} - \overset{\delta-}{F} \cdots$

H_2O ; $\overset{\delta+}{H} - \overset{\delta-}{\ddot{O}}: - H^{\delta+}$ $\quad H - \ddot{O}: \cdots H - \ddot{O}: \cdots H - \ddot{O}: \cdots$ (each O also bonded to H)

Liquid Water

NH_3 ; $H - N: \cdots H - N: \cdots H - N: \cdots$ (each N also bonded to two H)

2-Nitrophenol etc.

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Types of H-bonding

- Intermolecular Hydrogen Bonding Eg HF, H_2O, NH_3
- Intramolecular Hydrogen Bonding Eg 2-Nitrophenol, 2-Nitrobenzoic acid

Properties of H-Bonding

- Abnormally high boiling and melting points
- High solubilities
- Three dimensional crystal structure

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METALLIC BONDING:

This is the type of bonding which holds ~~the~~ atom together in metal crystals.
Many theories have been ~~proposed~~ proposed to explain the metallic bonding. Only the electron sea model will be considered here.

The Electron Sea Model

Metal atoms have the following characteristics:

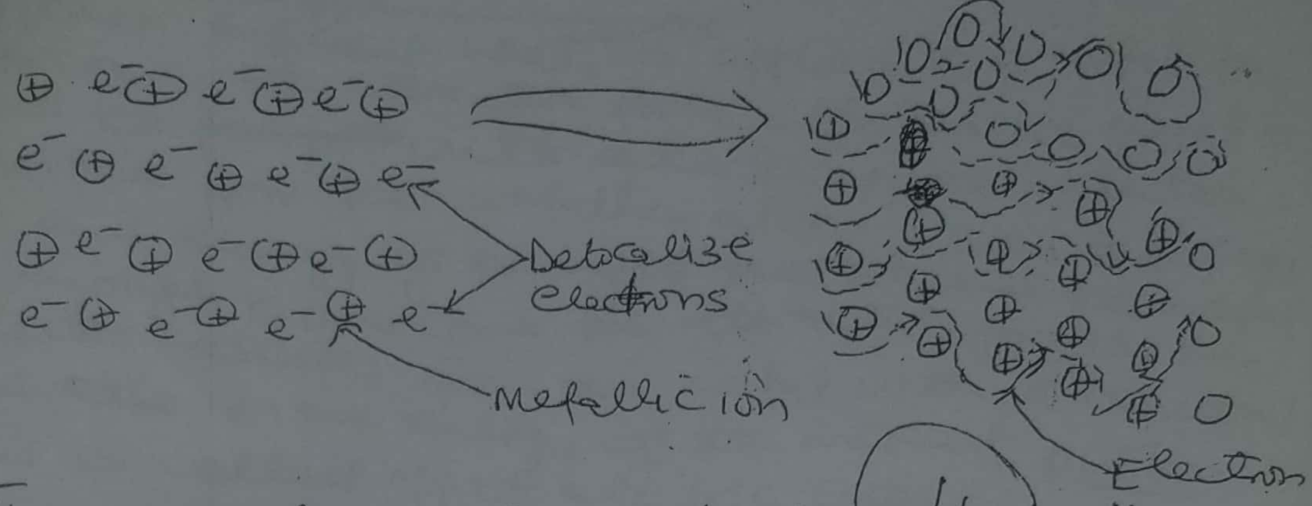
i) Low ionization Energies: This signifies that the valence electrons ~~can~~ in metal atoms can easily be separated.

ii) A number of vacant electron orbitals in their outermost shells; Eg Mg with electron configuration $1s^2 2s^2 2p^6 3s^2 3p^0$ has three vacant 3p orbitals in its outer electronic shell.

There is considerable overlapping of vacant orbitals on one atom with similar orbitals of adjacent atoms, through out the metal crystal. Hence, it is possible for an electron to be delocalized and move freely in the vacant molecular orbital passing through the entire metal crystal. The delocalized electrons belong to the whole crystal and no longer to the metal atoms.

Due to the delocalization of valence electrons, the positive metal ions produced remain fixed in the crystal lattice while the delocalized electrons are free to move about in the vacant space in between. The metal is thus pictured as a network of lattice of ...

in a "sea of electrons" or "gas of electr...
This simple model of metallic bonding is called the Electron sea Model.



The Electron sea Model

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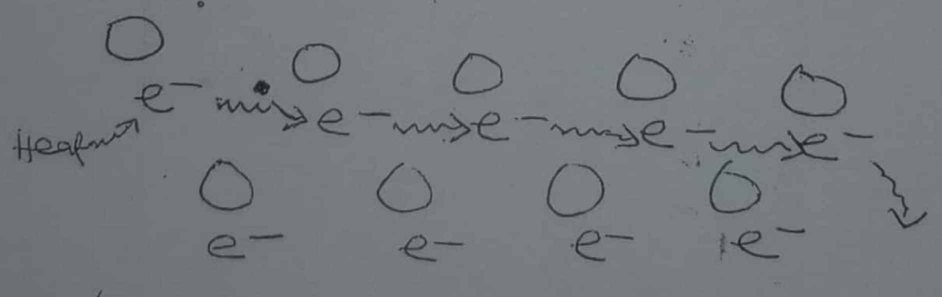
(iii) Heat conductivity: When a metal is heated at one end, the heat is carried to the ~~other~~ end. The mobile electrons in the electron sea at one end of the metal absorb heat energy and increase their vibrational motion. They collide with adjacent electrons and transfer the added energy to them. Thus electron mobility allows heat transfer to the ~~other~~ end.



Conduction of Heat through a metal.

iv) Ductility and Malleability: This can also be explained using the ~~E~~lectron sea Model. In metals, the positive ~~electrons~~ ions are surrounded by the sea of electrons that flows around them. If one layer of the metal ions is forced across the ~~other~~, by hammering, the internal ~~position~~ structure remains unchanged. The ...

crystal lattice is restored. This allows for it to be ductile and malleable.

(i) Emission of Electrons;

When sufficient heat energy is applied to a metal to overcome the attractive force between the metal ic and the ~~positive~~ outer electrons, the electron i emitted from the metallic atom. When the frequency and energy of the light that ~~strikes~~ strikes the metal is enough to overcome the attractive forces, the electron escapes from the metal ~~with~~ a resultant decrease in the energy of the incident photon & this is called "Photoelectric Effect".

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...ch reaction, ΔG is negative at all temperature

when a reaction is exothermic ($\Delta H = -ve$) and ΔS is negative, but ΔH is far more negative than $T\Delta S$ (at our temperature T) e.g

$$2Mg_{(s)} + O_{2(g)} \rightarrow 2MgO_{(s)} \quad \Delta H = -ve, \Delta S = -ve.$$

$$H_{2(g)} + \tfrac{1}{2}O_{2(g)} \rightarrow H_2O_{(l)} \quad \Delta H = -ve, \Delta S = -ve.$$

For endothermic reaction ($\Delta H = +ve$) will be spontaneous if entropy change ΔS is large and positive usually at a high temperature T e.g.

$$CaCO_{3(s)} \longrightarrow CaO + CO_{2(g)} \quad \Delta H\ +ve, \Delta S = +ve$$

$$T\Delta S > \Delta H \quad \therefore \quad \Delta H - T\Delta S < 0$$

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i) When two or more gases are mixed at constant volume, heat is neither liberated nor absorbed i.e $\Delta H = 0$, the driving force is due to an increase in entropy at constant volume.

$$\Delta G = 0 - T\Delta S = -ve$$

Non - Spontaneous reactions

A reaction will not be spontaneous at all temperatures, if it is endothermic (positive ΔH) and the entropy change ΔS is negative. Under these conditions, the free energy change, ΔG is positive.

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