

COVALENT STRUCTURES: SHAPE & BOND ANGLE.

VSEPR theory

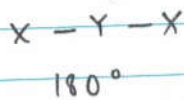
electrons will arrange themselves around a central atom as far apart as possible (because they repel)

- may be in BONDED pair, NON-BONDED pair or MULTIPLE BOND electron domains.

2x or 3x bond = ONE ELECTRON DOMAIN / CNC.

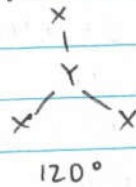
2 CNCs / Electron Domains

Linear



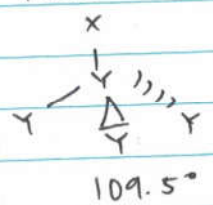
3 CNCs / EDs

Trigonal planar



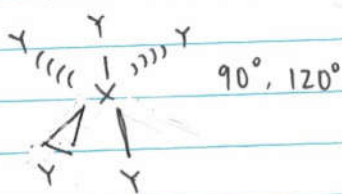
4 CNCs / EDs

Tetrahedral

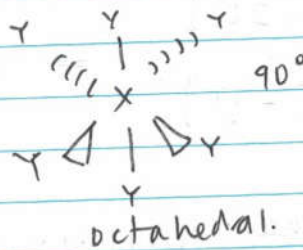


5 CNCs / EDs

Trigonal Bipyramid.



6 CNCs / EDs

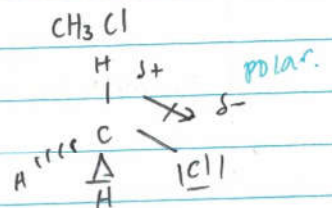
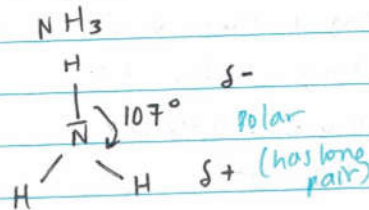
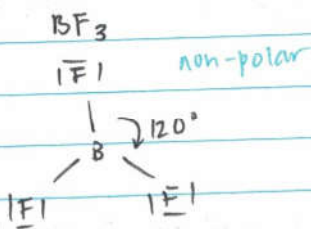


non bonded pairs also take up space.
angle ↓ 2.5 for each non-bond pair.

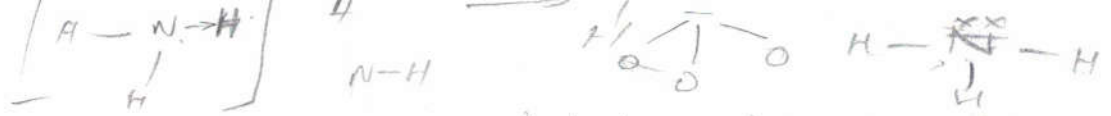
Bond polarity & Molecular polarity

molecular polarity is the result of individual bond polarity within a molecule AND the shape of the molecule itself

↳ some molecules may have polar bonds but are symmetrical = CANCEL.



- square planar (4 bpairs, 2 lps)
- trigonal pyramidal (3 pairs + 1 unpaired)
- non-linear / bent (2 pairs and 1+ 1 unpaired)
- v-shaped (117.5) (2, 2 BP + 1 lps)
- square pyramidal (5 pairs, 1 unpaired)



4 electron O. \rightarrow 1 is unbonded
 \rightarrow 0-bonded.

Resonance structures, Carbon + Silicon, **RESONANCE STRUCTURES**

one, ^{or} more possible positions for a **double bond** in a molecule/ion

giant covalent.

- ↳ individual structures are **resonance forms**
- ↳ actual structure is called **resonance hybrid**

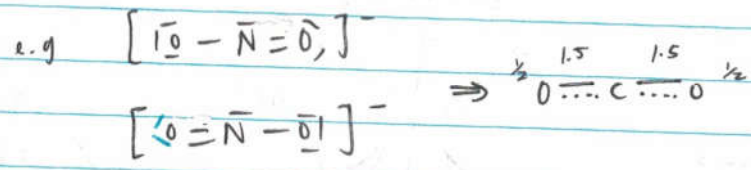
delocalisation of electrons

when there are a no. of p-orbitals (ie π bonds) on neighbouring atoms in a molecule/ion + a few resonance structures, can be drawn.

π cloud forms rather than **π bond** = electrons are **delocalised!**

THE SPREADING OF p ELECTRONS ACROSS 2+ ATOMS.

- bonds are 1.5 bonds
- bond length/strength is between single/double bond
- very stable structure!



GIANT COVALENT STRUCTURES: ALLOTROPES OF CARBON

- diamond:** 4 Bonds between carbon, = strength, all sigma
- ↳ tetrahedral structure
 - ↳ sp^3 hybridized (4 electron domains)
 - ↳ $x e^-$ delocalized $\therefore x$ conduct \downarrow
 - ↳ strong covalent bonds = mp \uparrow \uparrow

covalently bonded **interlocking** structure is very strong.

- graphite:** 3 Bonds between carbon (sigma) = layers
- ↳ hexagons that interlock
 - ↳ sp^2 hybridized
 - ↳ one e^- is in p-orbital \rightarrow overlaps w/ other p-orbitals to form π cloud
 - ↳ within layer = strong **ABOVE + BELOW** layer.
 - between = weak London forces

✓ lubricant layers slip ✓ conducts \downarrow , 'p' electrons move

GIANT COVALENT STRUCTURES: GRAPHENE

graphene

- layer of planar carbon atoms in HEXAGONAL arrangement
- similar to graphite = sp^2 hybridized
- giant covalent structure (crystalline lattice, 2D)
 - ✓ conduct \uparrow , 'p' electrons delocalized
 - ✓ strong ✓ flexible ✓ transparent

C₆₀ / Bucky Balls

- spherical carbon molecules e.g. buckminsterfullerene (60 C atoms)
- 3 strong sigma σ bonds, arranged as pent/Hex-agons \rightarrow SMALL molecule.
 - \rightarrow join to make ball shape
 - \rightarrow sp^2 hybridized
- ✓ strong ✓ light but **weak** van der Waals forces hold molecules together
mpT much lower than graphite/diamond.
- unpaired 'p' electron - forms π cloud - conduct \uparrow - not as good as layers.

Silicon

- same as diamond BUT no allotropic behaviour - only covalent bonding.

SiO₂

- Si bonds to 4 O atoms in tetrahedral structure
- O bonds to 2 Si atoms.
- covalent GIANT structures ✓ HIGH mpT X conduct \uparrow

SIGMA & PI BONDS

single covalent bonds are sigma (σ) \rightarrow an axial, head-on/end-on overlap of \geq atomic/hybrid orbitals

\rightarrow electrons distributed symmetrically around a line $x-x$

a double covalent bond is a σ AND a sideways overlap of 2 non-hybridized p-orbitals
this happens ABOVE + below the sigma bond (π , pi bond)

in a triple covalent bond $\rightarrow \sigma + 2\pi$

HYBRIDISATION

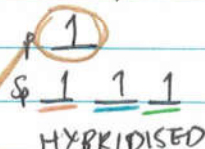
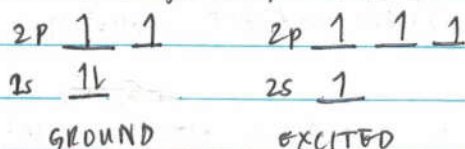
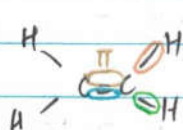
is the combining of atomic orbitals to produce new hybrid bonding orbitals
 \hookrightarrow when atoms covalently bond, they share electrons, but this happens by **OVERLAPPING** orbitals.

\hookrightarrow BUT: some are in...
① different orbitals (s, p, d, f)
② different shapes.

so, atoms create a NEW bonding orbital

a combination/HYBRID of types (s, p, d, f), energy and shape of orbital.

e.g. Ethene



tetrahedron \leftarrow

this leaves ONE unhybridised

trigonal planar \leftarrow

4 CNCS - $2p^3$ 'p' electron to form a pi bond!!

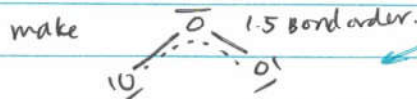
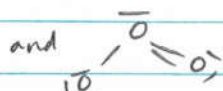
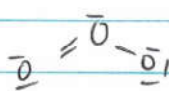
anything with 3 CNCS - sp^2

linear \leftarrow 2 CNCS - sp

lower energy level, more stable, harder to break

lone pairs don't count

OZONE (O_3)



bond strength =

bond has same energy

UV A - high wavelength, low frequency \therefore LOW ENERGY.

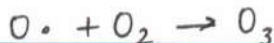
B

C - low wavelength

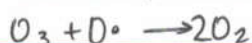
ozone stops B/A

oxygen stops C

formation



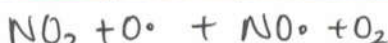
depletion



reaction with CCl_2F_2



NO_x



OVERALL



depletion catalysed

METALLIC BOND

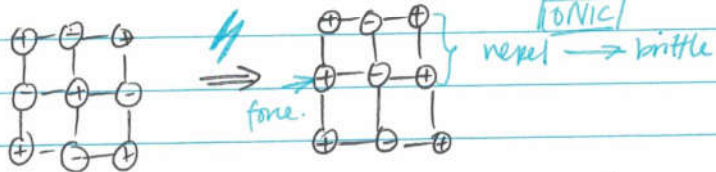
metals: generally have low electronegativity (and like to be oxidized)



metallic bond = bond between positive metal ion and a sea of delocalised electrons.

BONDING IS MOBILE

↳ when you HIT them.



but metals are MALLEABLE (make into shapes)

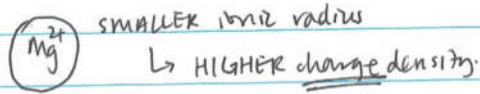
↳ e^{-} constantly moving, keeping the bonds

what makes metal strong?



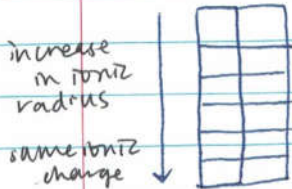
metalliz

DUCTILE (pull into wires)



↳ HIGHER charge density

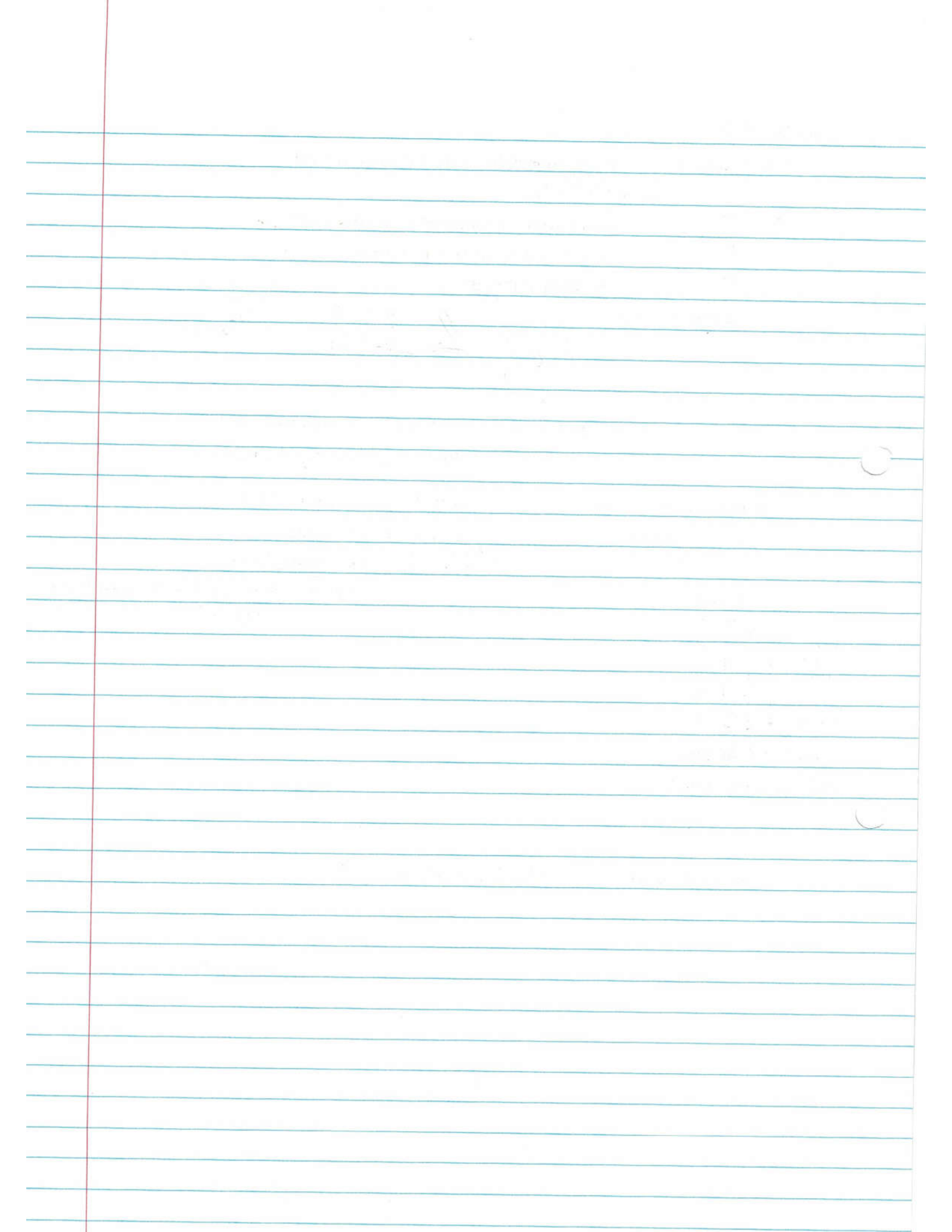
↳ same charge, (same charge, but larger atom)
(2+) →



DECREASE CHARGE DENSITY

LOWER MELTING POINT.

alloys are mixed metals or a metal with non-metallic element

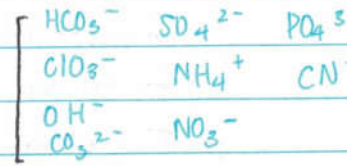


BONDING

IONIC BONDING & STRUCTURE.

metals donate valence electrons. \rightarrow +ve ions (cations)
 non-metals gain valence electrons = FULL OUTER SHELL.
 \rightarrow -ve ions (anions)
 \rightarrow held together by electrostatic forces.

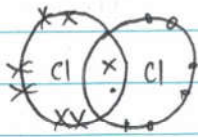
ions w/ more than 1 element = polyatomic.



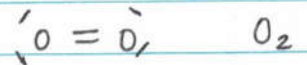
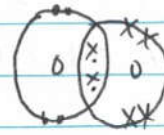
- an ionic bond is due to the electrostatic attraction between oppositely charged ions.
- they are usually solids with lattice structures, and have solid crystalline properties
 \rightarrow \therefore have HIGH melting points due to strong electrostatic attraction.
- conduct electricity in aqueous/molten state - ions free to move
- dissolve in water due to polar nature
- electrolytes solvents (power source + industrial process)

COVALENT BONDING

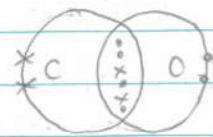
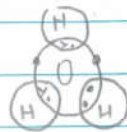
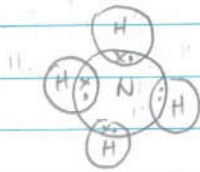
happens between a PAIR of electrons and positively charged nuclei held together by mutual attraction of the pair of electrons to the nuclei.



and may have multiple bonds



coordinate bonds are formed when ONE ATOM supplies BOTH shared electrons to form covalent bond. (NH_4^+ , H_3O^+ , CO)



bond length decreases
 bond strength increases

as the no. of shared electrons $\uparrow\uparrow$

$$\left[\begin{array}{l} \text{bond length} \\ \propto \\ \frac{1}{\text{bond strength}} \end{array} \right]$$

IONIC OR COVALENT?

electronegativity is the ability of an atom to attract an electron (covalent bond)

large electronegativity = ionic. ($\Delta > 1.8$)

\rightarrow closest to Fluorine = most electronegative.

\rightarrow similar electronegativity = less polar.

greater Δ in electronegativity = more polar (e.g. water)

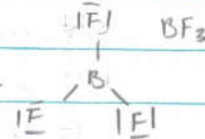
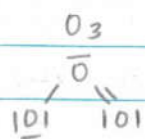
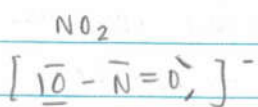
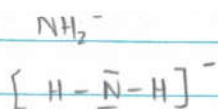
COVALENT STRUCTURES

Electrons exist in **pairs** (xx, x^o, -)

- most elements follow the **octet rule** (CONF. conference)
- others can have **less than 8e** (Be, B, Al, beryll))
- others expand **make more than 8e** (P, S, Cl)

working out lewis structures:

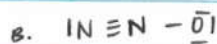
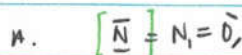
- total no. of valence electrons $\div 2 =$ pairs
- link the central atom + other atoms.
- add more pairs round each atom (unless H, Be (4e), B & Al (6e))
- add double bonds. Ⓢ if P, S, Cl: add electrons to central atom.



But what if there's more than one lewis structure? Ans: **FORMAL CHARGE**.

e.g. N_2O

↳ FC = charge of atom if all molecule had = electronegativity.



FC = no. of valence e⁻ - no of bonds - no. of bonding e⁻ (NON-)

a. FC([N]) = 5 - 2 - 4 = -1

FC(N) = 0

FC(N) = **(-2)** too far from 0.

(N₁) = 5 - 4 - 0 = +1

(N₂) = +1

(N₁) = +1

(O) = 6 - 2 - 4 = 0

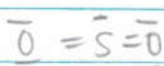
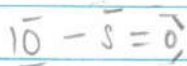
(O) = -1

(O) = +1

lower FC = most stable structure (less charge transfer)

A + B have same lowest FC, so you must choose the one where the negative values of FC is on the MOST electronegative atom

O is the most electronegative. ∴ B is preferred.



FC(O) = 6 - 1 - 6 = (-1)

FC(O) = 6 - 2 - 4 = 0

(S) = 6 - 3 - 2 = (2+)

(S) = 6 - 4 - 2 = 0

(O₂) = 6 - 2 - 4 = (0)

(O) = 6 - 2 - 4 = 0

INTERMOLECULAR FORCES

forces **between** covalent molecules.

ALL = VAN DER WAALS FORCE

- London forces
- dipole dipole
- H+ bonds.

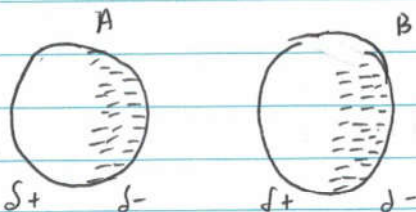
simple molecular compounds

STRONG INTRA-molecular forces (within molecule)

WEAK INTER-molecular forces

STRONG COVALENT bonds BETWEEN ATOMS.

LONDON FORCES



movement of electrons creates a temporary dipole which INDUCES a dipole in a neighbouring molecule.
temporary induced dipole.

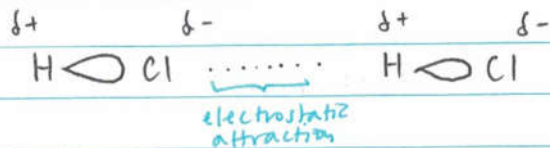
STRENGTH = more electrons (Mr)

= surface area of contact

(ability to get closer)

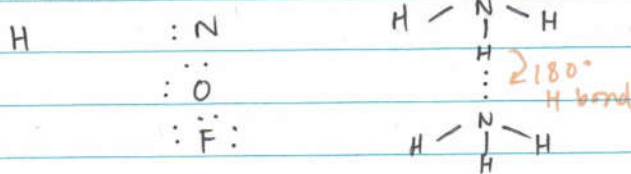
↑ CONTACT ↑ STRENGTH ↑ MP.

DIPOLE-DIPOLE



formed due to a permanently induced dipole (the molecules are polar)
STRENGTH: depends on difference in electronegativity (how polar)

HYDROGEN BONDS

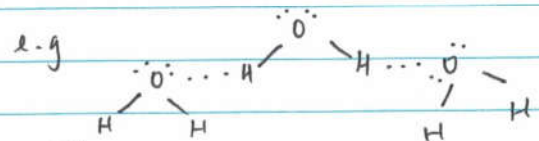


it is an interaction between a hydrogen atom and a lone pair of N, O or F.

STRENGTH: no. of H bonds
no. of lone pairs.

properties of hydrogen bonds

- HIGH B.P
- HIGH surface tension (adhesive)
- SOLID less dense than liquid.



open lattice = every H can bond w/ another atom.