PROSPERITY ACADEMY

AS CHEMISTRY 9701

Crash Course

RUHAB IQBAL

CHEMICAL BONDING & INTERMOLECULAR FORCES

COMPLETE NOTES

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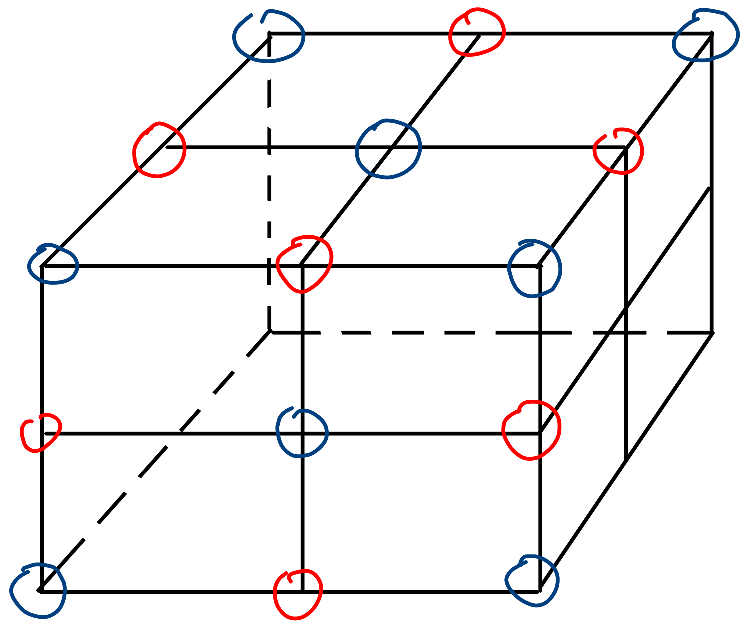
Chemical bonding and Intermolecular forces:-

Ionic bonding

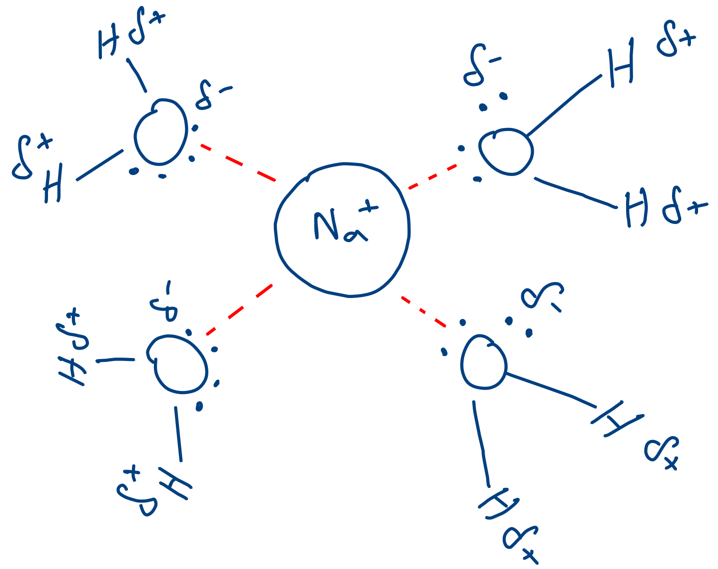
Ionic bond:- The strong electrostatic force of attraction between a positive metal cation and negative non metal anion.

- The metal atom gives its electron(s) to the non metal atom to achieve a stable configuration.
  ↳ cation ↳ anion
- Ionic compounds always form giant rigid lattice structures
- Ionic compounds have a high melting point as the ionic bonds require lots of energy to break
- Donot conduct in solid state but conduct in molten or aqueous state due to free moving ions.



○ : Cl^-
○ : Na^+

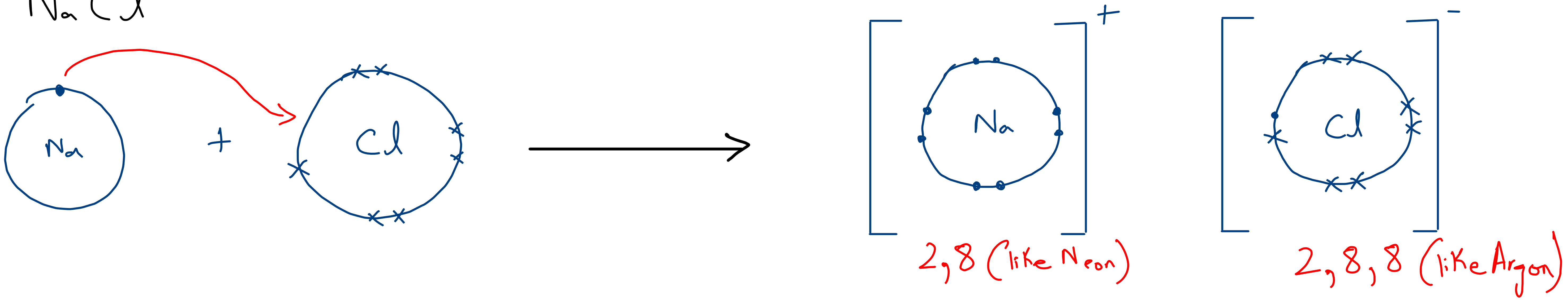
- Ionic compounds dissolve in water very well as water has hydrogen bonding and it attracts the ions favouring new bonds to form.



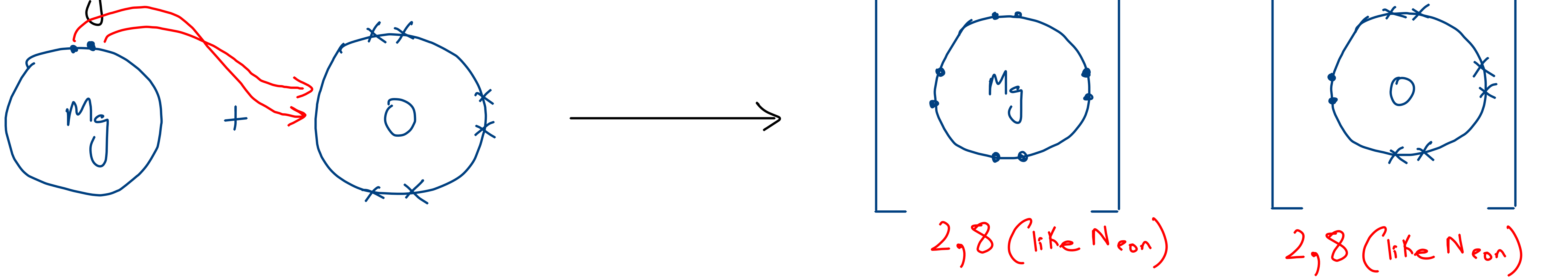
(New intermolecular forces are made favouring dissolving)

Dot and Cross diagrams:- (only showing outer shell electrons)

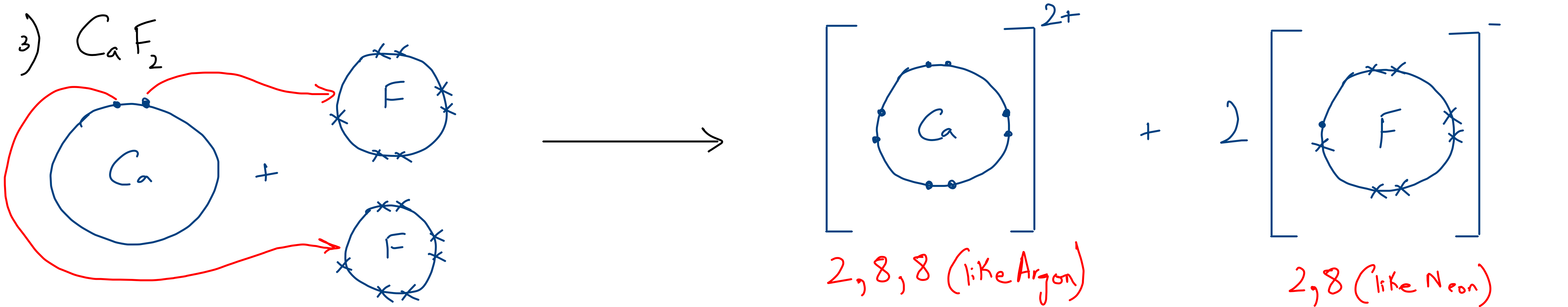
1) $NaCl$



2) MgO



3) CaF_2



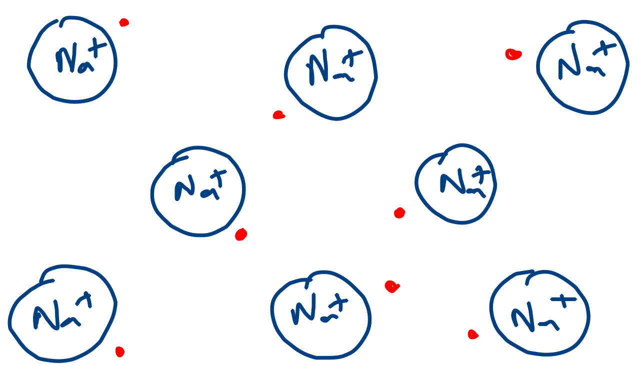
Metallic bonding:-

Metallic bond: Strong force of electrostatic attraction between positive cations and delocalised sea of electrons.

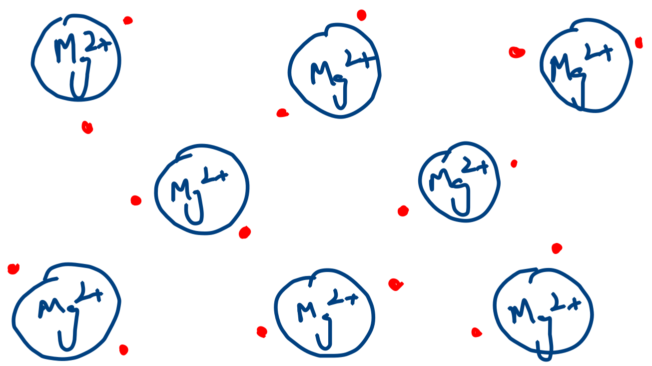
- The metal atoms lose their valence electrons to become positive cations and the delocalised sea of electrons form a "sea" around them.
- The cations are packed together in a regular lattice structure.
- Metals have high melting and boiling points as metallic bonds are very strong.
- Metals are good conductors due to free moving electrons in them in any state
- insoluble in water

Q. Show metallic bonding in

1) Na



2) Mg



The strength of metallic bonding increases with:

- greater cationic charge
- greater number of delocalised electrons
- smaller ionic radius

Rank $\overset{\text{radius decreases}}{\overrightarrow{\underset{+}{\underline{Na}},\ \underset{+2}{\underline{Mg}},\ \underset{+3}{\underline{Al}}}}$ in order of increasing melting points.

Ans. Na, Mg, Al

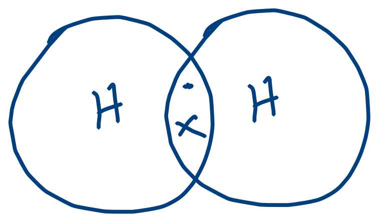
Covalent bonding:-

Covalent bond:- The force of attraction between nuclei and shared pair of electrons.

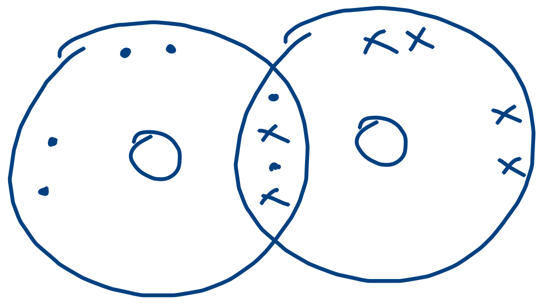
- Atoms share their electrons with each other to achieve a stable configuration.
- Can have giant covalent structures or simple molecular structures
- Usually insoluble in water

Dot and Cross diagrams:-

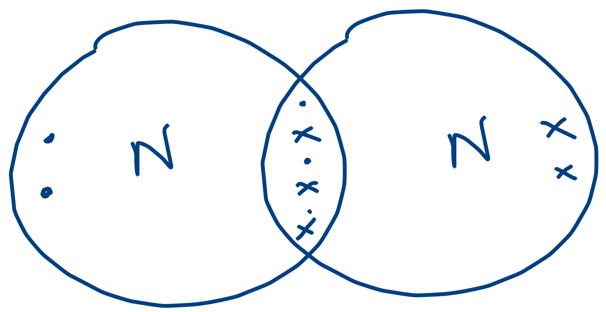
1) H_2



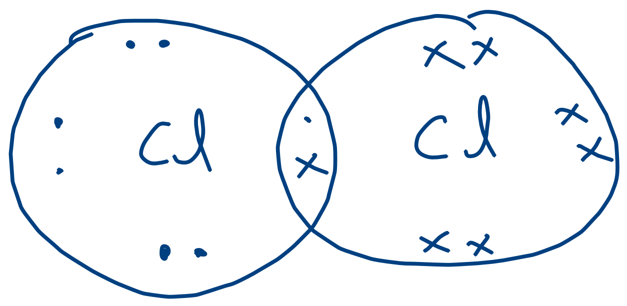
2) O_2



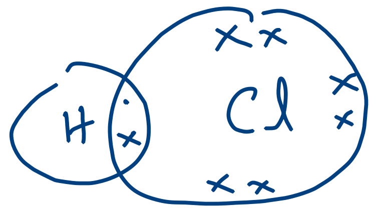
3) N_2



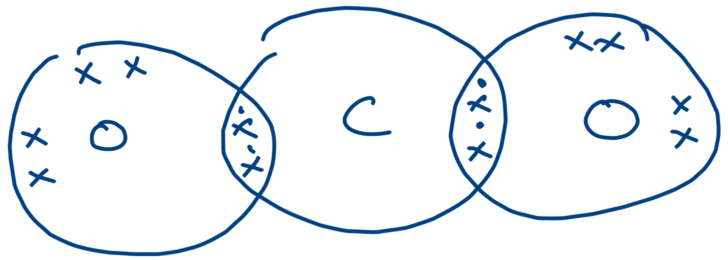
4) Cl_2



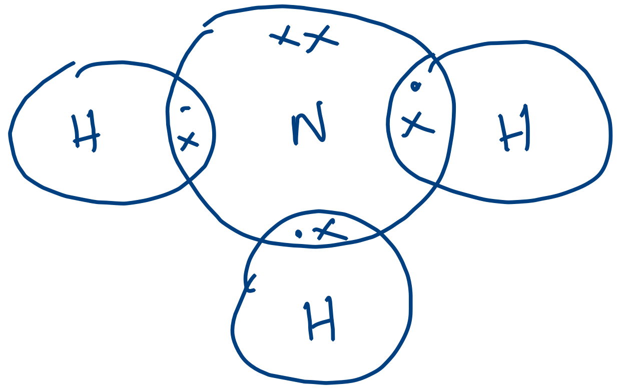
5) HCl



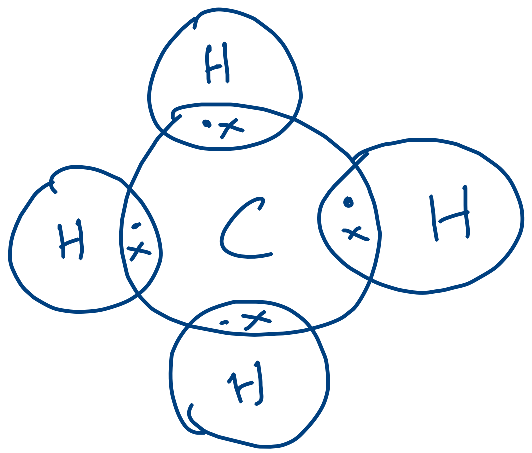
6) CO_2



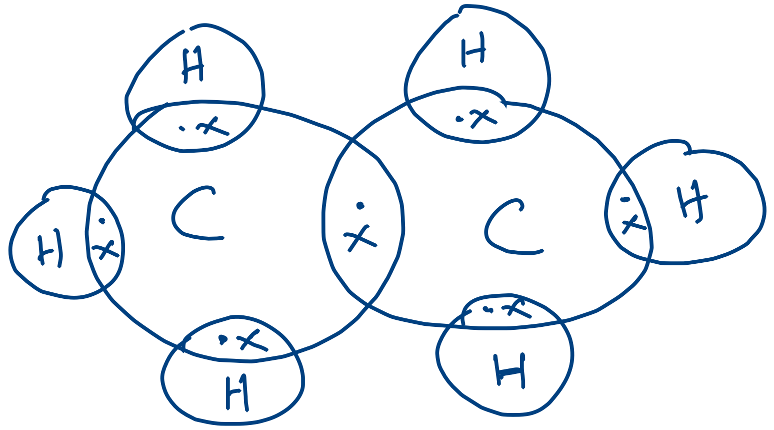
7) NH_3



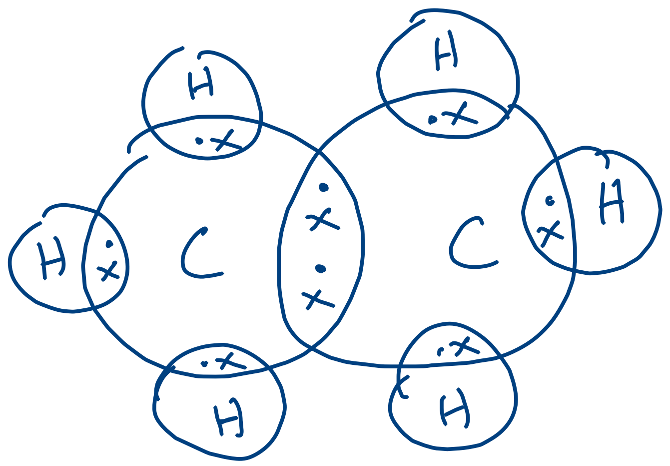
8) CH_4



9) C_2H_6 (ethane)



10) C_2H_4 (ethene)

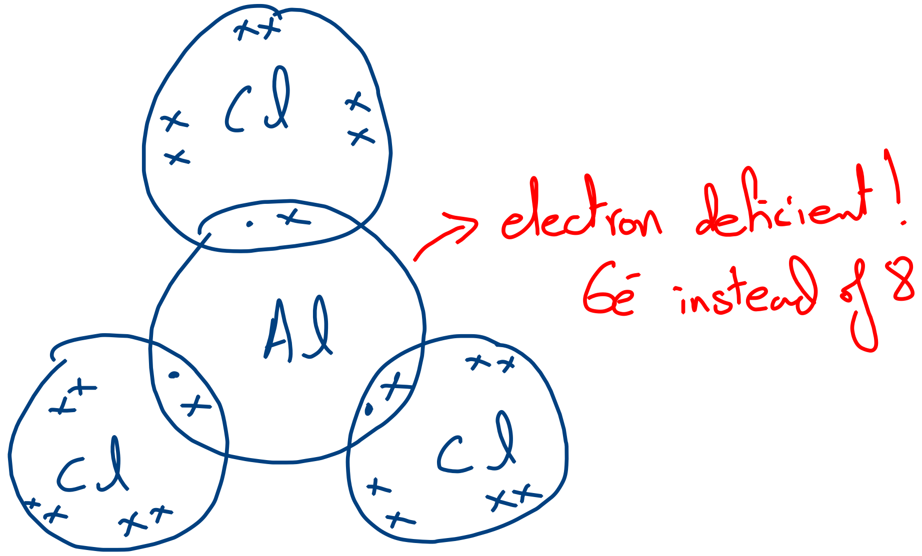


Dative bonding / Coordinate bonding :-

Dative bond :- A dative bond is formed when an atom uses its lone pair and contributes both electrons in the shared pair of electrons.

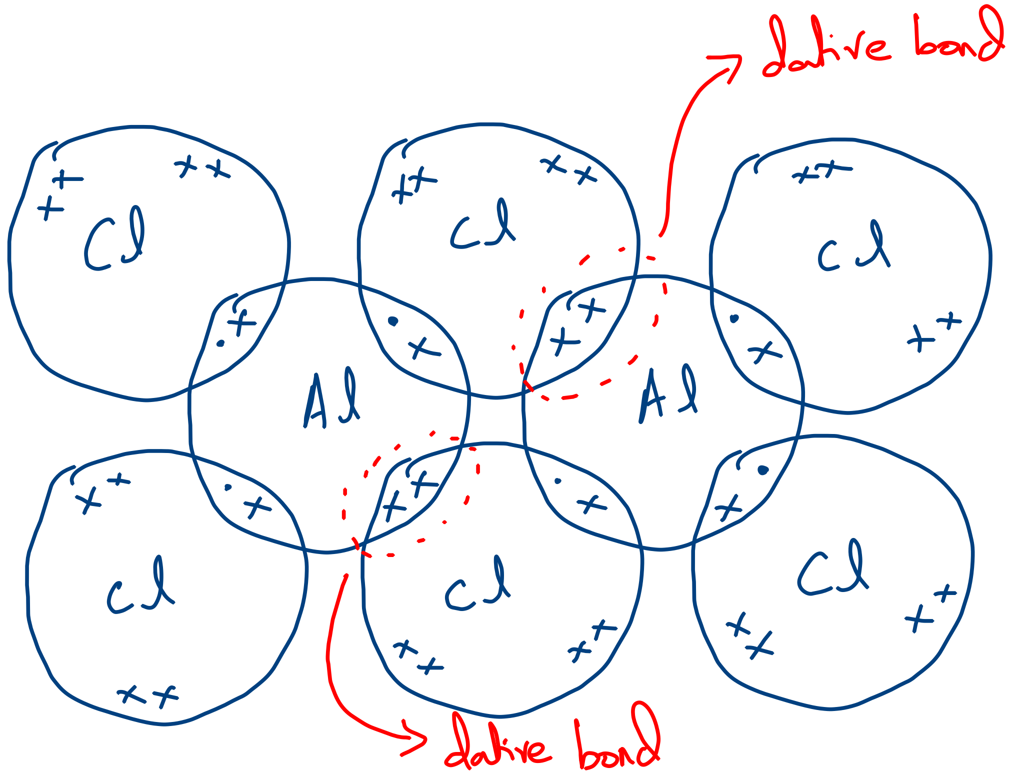
Lone pair : A pair of valence electrons not bonded to any other species. (Unshared pair of electrons)

Very important course example :- Al_2Cl_6



$AlCl_3$

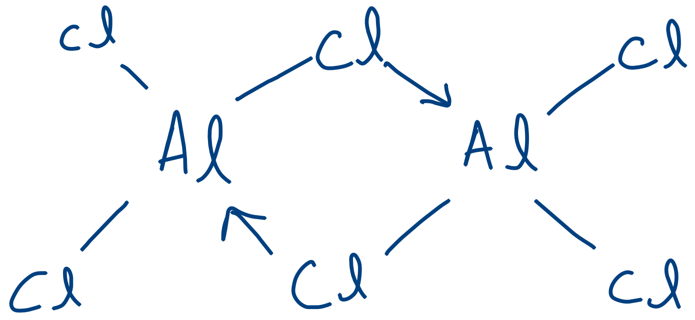
dimerises



- This molecule is very unstable
  1) Only exists in gaseous state at very high temperatures
  2) Aluminium is electron deficient.

- Aluminium is no more electron deficient and so the molecule is now more stable

Al_2Cl_6 can also be shown as



- Arrow represents dative bond
- Arrow points from donor to acceptor

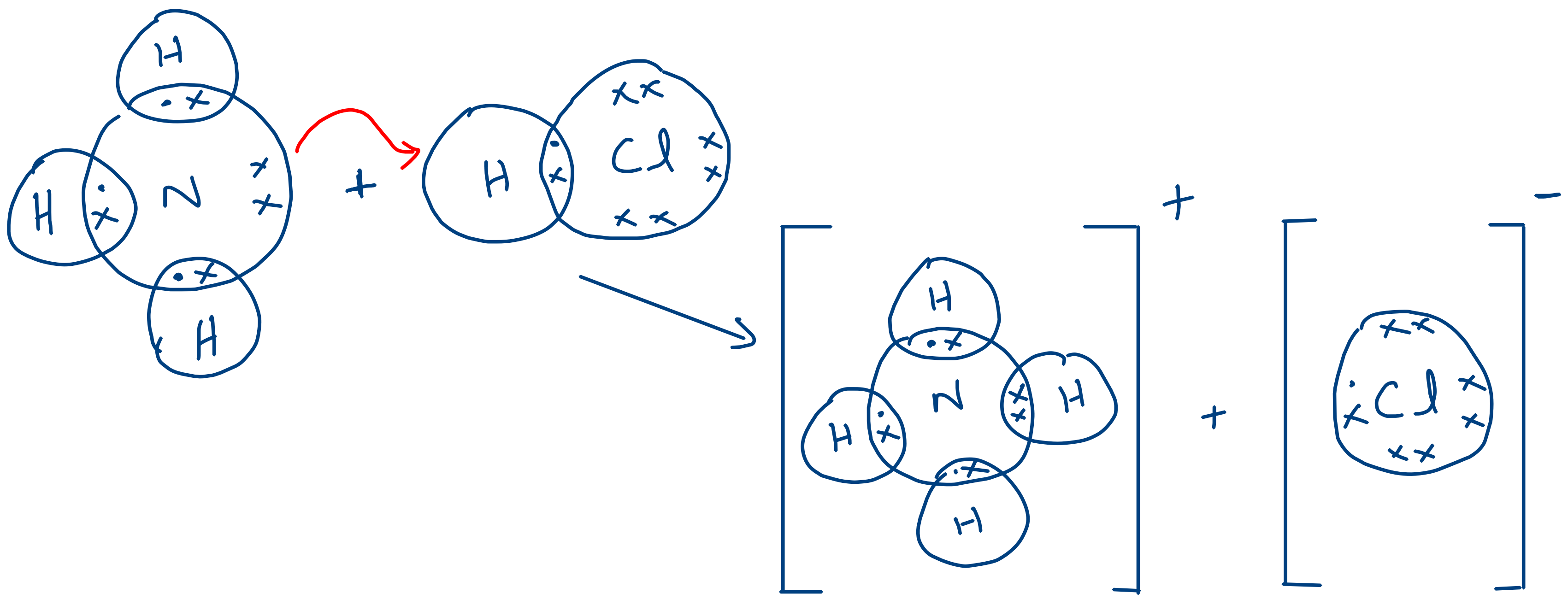
What favours the formation of dative bonds?

- Electron deficiency
- presence of lone pair

- Incase of a base accepting a hydrogen, it is feasible due to:
  1) lone pair
  2) an ionic bond can form which is more stable than a covalent bond.

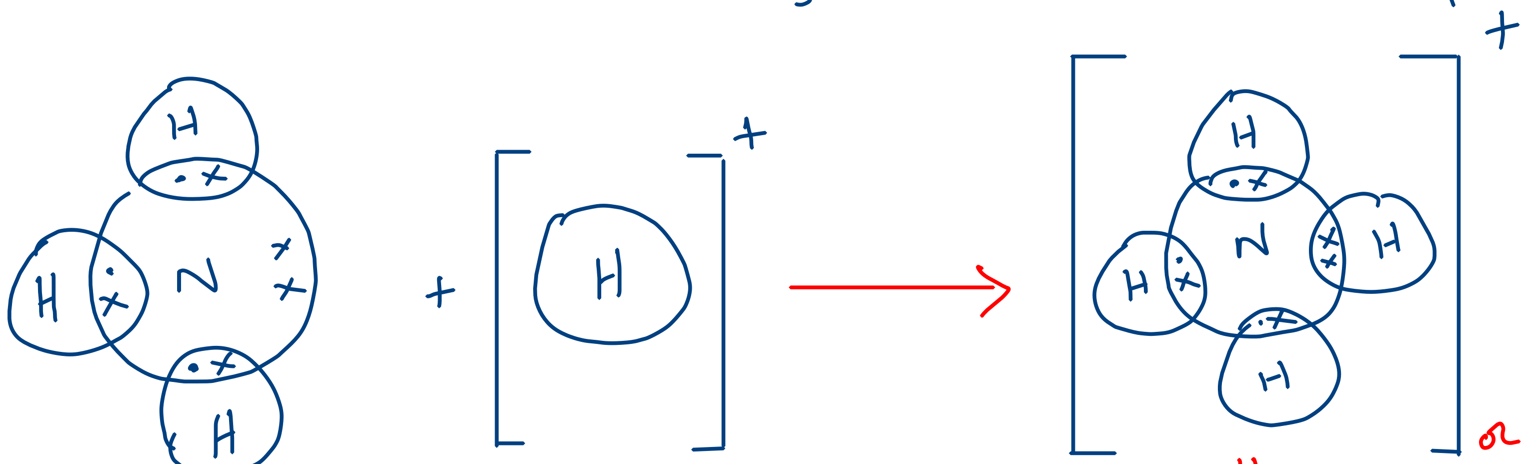
More examples :-

1) $NH_3 + HCl \longrightarrow NH_4^+Cl^-$

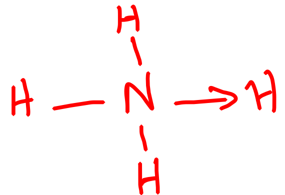


ionic equation: $NH_3 + H^+ + \cancel{Cl^-} \longrightarrow NH_4^+ + \cancel{Cl^-}$

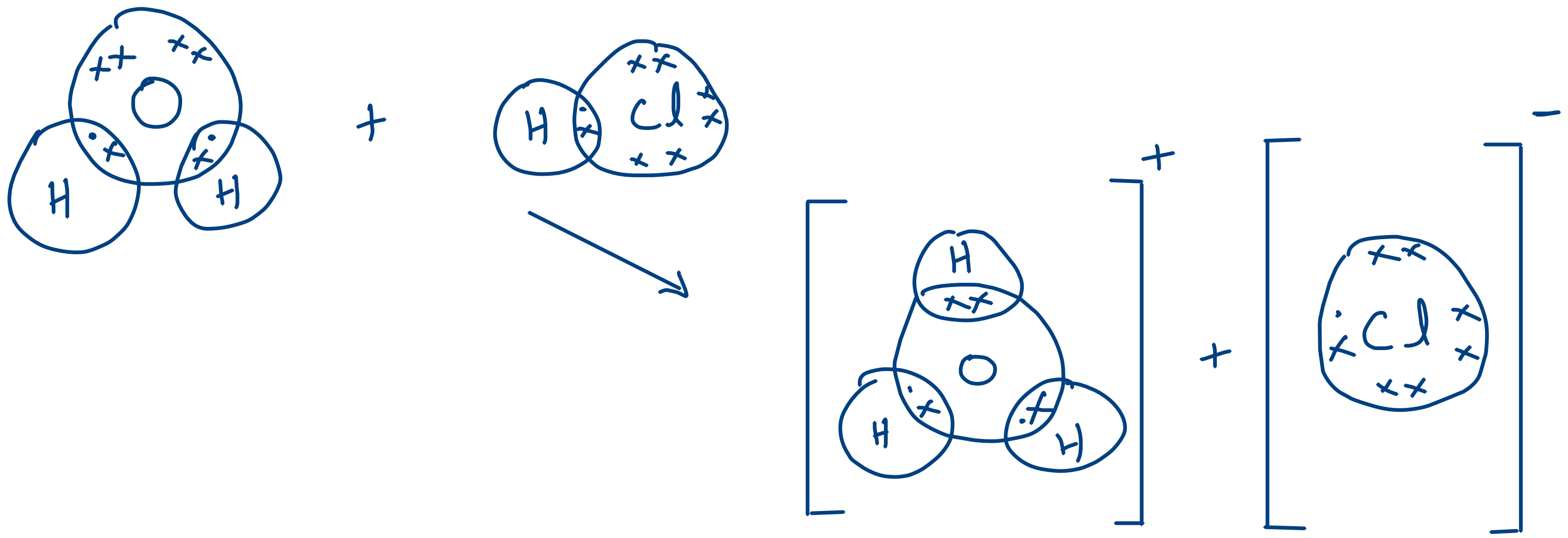
$NH_3 + H^+ \longrightarrow NH_4^+$



or

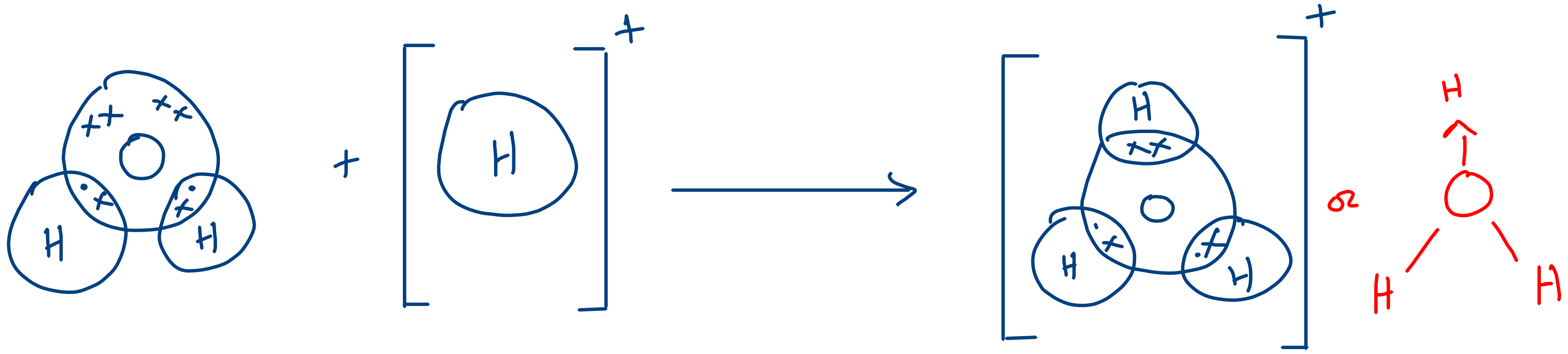


2) $H_2O + HCl \longrightarrow H_3O^+ + Cl^-$

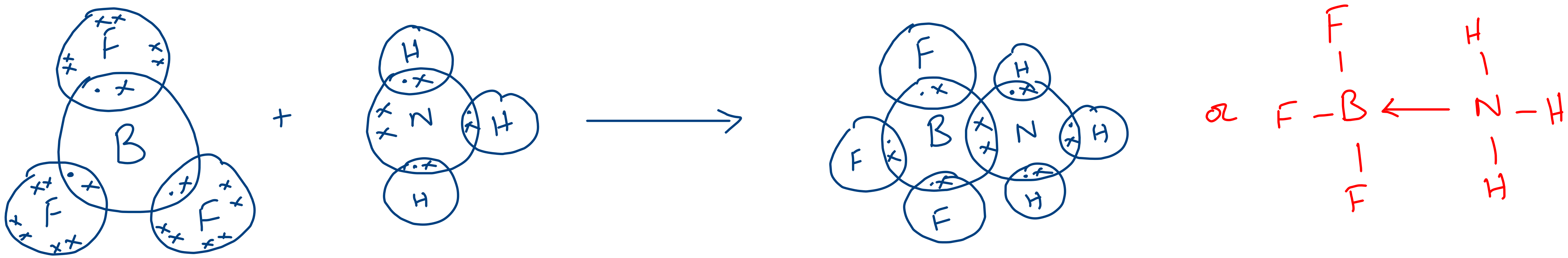


ionic equation: $H_2O + H^+ + \cancel{Cl^-} \longrightarrow H_3O^+ + \cancel{Cl^-}$

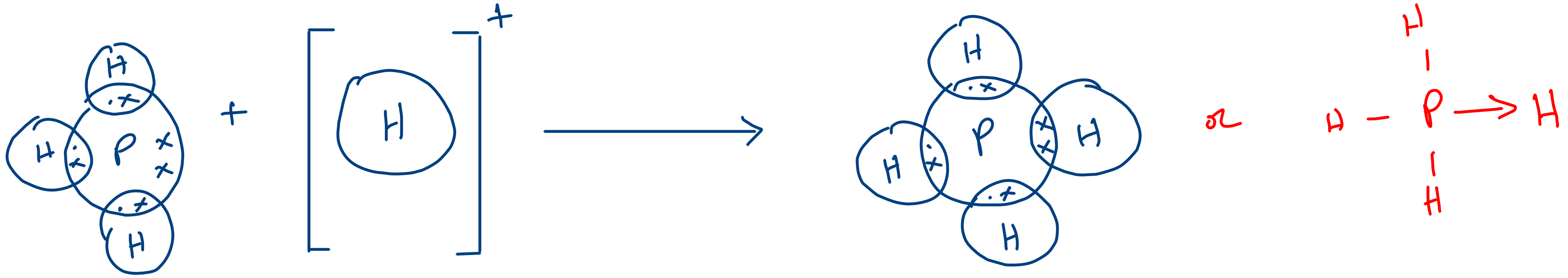
$H_2O + H^+ \longrightarrow H_3O^+$



3) $BF_3 + NH_3 \longrightarrow F_3BNH_3$



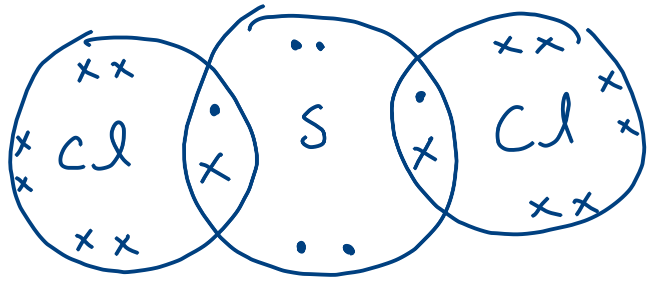
4) $PH_3 + H^+ \longrightarrow PH_4^+$



(b) SCl_2 is formed in the following reaction.

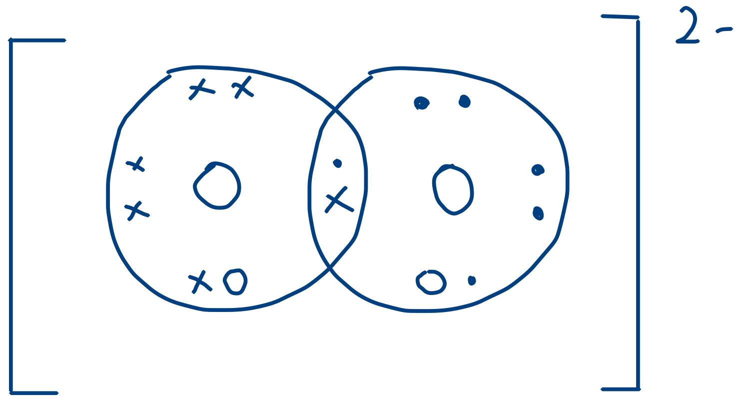
$$S_2Cl_2(l) + Cl_2(g) \rightleftharpoons 2SCl_2(l) \qquad \Delta H = -40.6\,kJ\,mol^{-1}$$

(i) Complete the 'dot-and-cross' diagram to show the bonding in a molecule of SCl_2. Show outer electrons only.



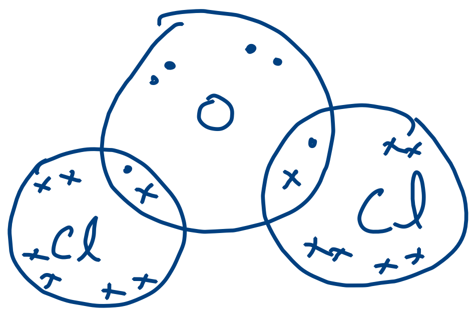
(iii) The peroxide ion, O_2^{2-}, has a single covalent bond between the two oxygen atoms. Each oxygen atom carries a negative charge.

Draw a 'dot-and-cross' diagram for the peroxide ion. Show outer electrons only.



O : electrons from another species

(i) Draw a 'dot-and-cross' diagram of Cl_2O. Show outer-shell electrons only.

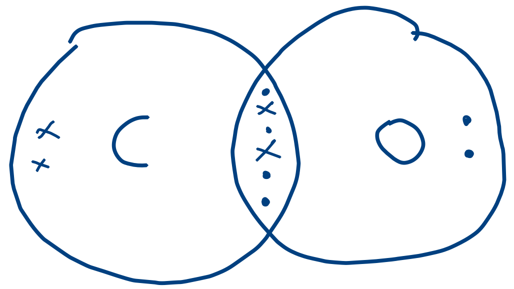


(b) Carbon monoxide, CO, is a gas at room temperature and pressure. It contains a coordinate bond.

(ii) Draw a 'dot-and-cross' diagram to show the arrangement of outer electrons in CO.

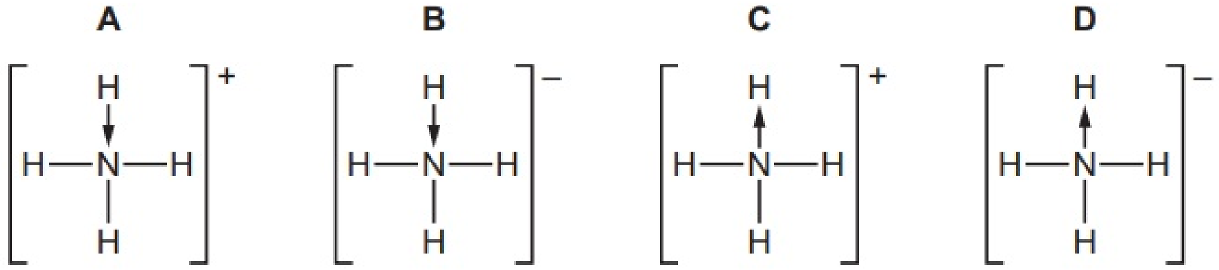
Show the electrons belonging to the C atom as **×**.

Show the electrons belonging to the O atom as **●**.



18 The dative covalent bond can be represented by an arrow, →. The arrow points towards the atom receiving the lone pair.

Which diagram of an ammonium ion is correct?



6 Which type of bonding is **never** found in elements?

A covalent

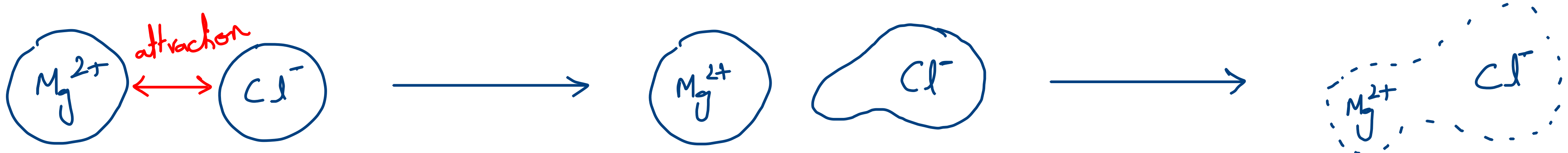
B ionic

C metallic

D van der Waals' forces

Polarisation of ions:-

In ionic bonding, the cation can pull some of the electron density from the anion, distorting its shape.



Some degree of covalent character developed

Polarising power of cation: The tendency of a cation to polarise an anion

A cation will have a greater polarising power if:

1) It has a greater positive charge
2) It has a smaller ionic radius.

→ Across a period, for example period 3

Na^+, Mg^{2+}, Al^{3+}, Si^{4+} → Polarising power increases

charge increases

ionic radius decreases

Polarisability of an anion: The ease with which an anion can be polarised

An anion will have greater polarisability if:

1) Greater negative charge
2) Greater ionic radius

→ Across a period, for example period 3

N^{3-}, O^{2-}, F^- → Polarisability decreases

charge decreases

ionic radius remains roughly same

Q. Rank the anions in order of increasing polarisability: F^-, Cl^-, S^{2-}

smaller radius — bigger radius — greatest charge so greatest polarisability

Ans. F^-, Cl^-, S^{2-}

Q. What factors favour ionic bonding? (Opposite of polarisation power and polarisability)

For cation:
- smaller positive charge
- greater cationic radius

For anion:
- smaller negative charge
- smaller anionic radius.

That is why $NaCl$ & $MgCl_2$ are ionic and Al_2Cl_6 and $SiCl_4$ are covalent

Polarity in Covalent molecules:-

Electronegativity:- The tendency of an atom to attract a shared pair of electrons towards itself.

Learn this: F O N Cl
$\longrightarrow$ order of electronegativity

F: most electronegative   O: 2nd   N & Cl: about the same (3rd and 4th)

[C and H and I have similar electronegativity]

use this to work out trend

Trends of electronegativity across the periodic table:-

The Periodic Table of Elements

Key: atomic number / atomic symbol / name / relative atomic mass

Group 1	2	3	4	5	6	7	8	9	10	11	12	13	14	15	16	17	18
							1 H hydrogen 1.0										2 He helium 4.0
3 Li lithium 6.9	4 Be beryllium 9.0											5 B boron 10.8	6 C carbon 12.0	7 N nitrogen 14.0	8 O oxygen 16.0	9 F fluorine 19.0	10 Ne neon 20.2
11 Na sodium 23.0	12 Mg magnesium 24.3											13 Al aluminium 27.0	14 Si silicon 28.1	15 P phosphorus 31.0	16 S sulfur 32.1	17 Cl chlorine 35.5	18 Ar argon 39.9
19 K potassium 39.1	20 Ca calcium 40.1	21 Sc scandium 45.0	22 Ti titanium 47.9	23 V vanadium 50.9	24 Cr chromium 52.0	25 Mn manganese 54.9	26 Fe iron 55.8	27 Co cobalt 58.9	28 Ni nickel 58.7	29 Cu copper 63.5	30 Zn zinc 65.4	31 Ga gallium 69.7	32 Ge germanium 72.6	33 As arsenic 74.9	34 Se selenium 79.0	35 Br bromine 79.9	36 Kr krypton 83.8
37 Rb rubidium 85.5	38 Sr strontium 87.6	39 Y yttrium 88.9	40 Zr zirconium 91.2	41 Nb niobium 92.9	42 Mo molybdenum 95.9	43 Tc technetium –	44 Ru ruthenium 101.1	45 Rh rhodium 102.9	46 Pd palladium 106.4	47 Ag silver 107.9	48 Cd cadmium 112.4	49 In indium 114.8	50 Sn tin 118.7	51 Sb antimony 121.8	52 Te tellurium 127.6	53 I iodine 126.9	54 Xe xenon 131.3
55 Cs caesium 132.9	56 Ba barium 137.3	57–71 lanthanoids	72 Hf hafnium 178.5	73 Ta tantalum 180.9	74 W tungsten 183.8	75 Re rhenium 186.2	76 Os osmium 190.2	77 Ir iridium 192.2	78 Pt platinum 195.1	79 Au gold 197.0	80 Hg mercury 200.6	81 Tl thallium 204.4	82 Pb lead 207.2	83 Bi bismuth 209.0	84 Po polonium –	85 At astatine –	86 Rn radon –
87 Fr francium –	88 Ra radium –	89–103 actinoids	104 Rf rutherfordium –	105 Db dubnium –	106 Sg seaborgium –	107 Bh bohrium –	108 Hs hassium –	109 Mt meitnerium –	110 Ds darmstadtium –	111 Rg roentgenium –	112 Cn copernicium –	113 Nh nihonium –	114 Fl flerovium –	115 Mc moscovium –	116 Lv livermorium –	117 Ts tennessine –	118 Og oganesson –

$\xleftarrow{\text{decreases}}$

decreases $\downarrow$

Across a period, electronegativity increases as:

1) nuclear charge increases
2) shielding remains constant
3) atomic radius decreases
   ↳ so more attraction between shared pair and nucleus

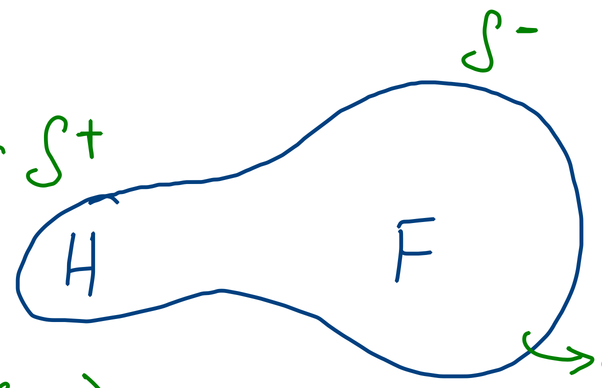
Down a group, electronegativity decreases as:

1) nuclear charge increases
2) shielding also increases

] — cancel each other out

3) atomic radius increases due to increase in number of shells $\longrightarrow$ so less attraction between shared pair and nucleus.

Discussing polarity in H – F: Fluorine is much more electronegative than H and so pulls most of the electron density towards itself.



partially positive ($\delta +$)

electrons spend more time here so this becomes partially negative ($\delta -$)

This can be drawn as:

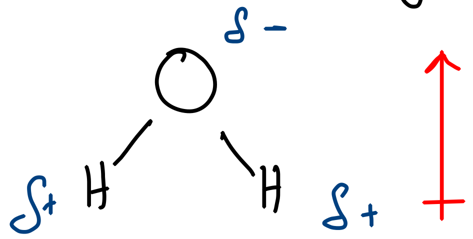
$$\overset{\longmapsto}{\underset{\delta+}{H} - \underset{\delta-}{F}}$$

arrow points from $\delta +$ to $\delta -$

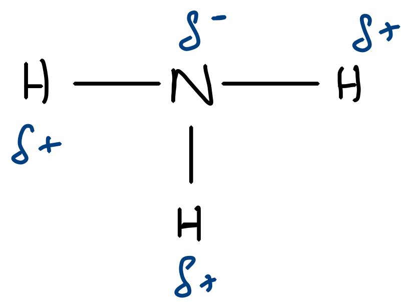
Q. Predict the partial positive and negative charges on:

1) $\underset{\delta+}{H} - \underset{\delta-}{Cl}$ $\longmapsto$

2)



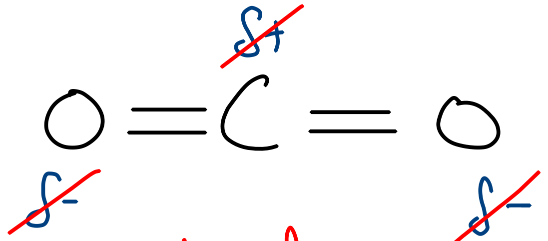
3)



4) H — H
non polar

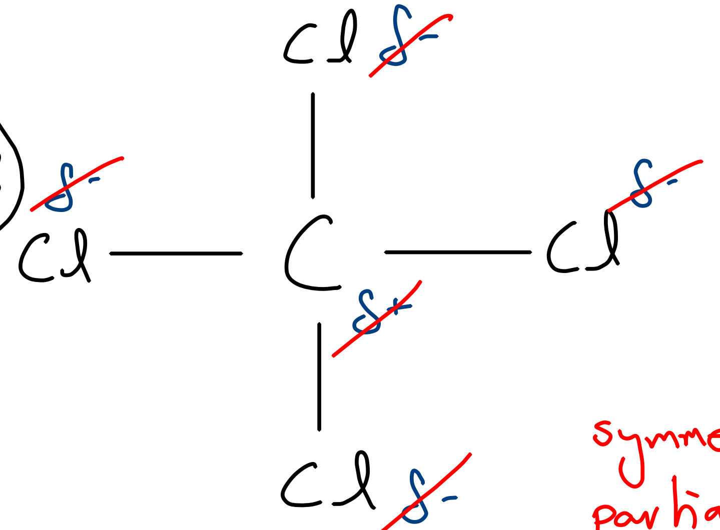
Same species bonded together are non polar for e.g. O_2, N_2, Cl_2, F_2

5)



symmetrical so no overall partial charges so therefore non polar

6)



symmetrical so no overall partial charges so therefore non polar

Behaviour of Polar molecules in an electric field:-

Pauling Electronegativity scale:- (No need to learn)

H 2.1																	He ---
Li 1.0	Be 1.5											B 2.0	C 2.5	N 3.0	O 3.5	F 4.0	Ne ---
Na 0.9	Mg 1.2											Al 1.5	Si 1.8	P 2.2	S 2.5	Cl 3.0	Ar ---
K 0.8	Ca 1.0	Sc 1.3	Ti 1.5	V 1.6	Cr 1.6	Mn 1.5	Fe 1.8	Co 1.8	Ni 1.8	Cu 1.9	Zn 1.6	Ga 1.6	Ge 1.8	As 2.0	Se 2.4	Br 2.8	Kr 3.0
Rb 0.8	Sr 1.0	Y 1.2	Zr 1.4	Nb 1.6	Mo 1.8	Tc 1.9	Ru 2.2	Rh 2.2	Pd 2.2	Ag 1.9	Cd 1.7	In 1.7	Sn 1.8	Sb 1.9	Te 2.1	I 2.5	Xe 2.6
Cs 0.7	Ba 0.9	La-Lu 1.1-1.2	Hf 1.3	Ta 1.5	W 1.7	Re 1.9	Os 2.2	Ir 2.2	Pt 2.2	Au 2.4	Hg 1.9	Tl 1.8	Pb 1.8	Bi 1.9	Po 2.0	At 2.2	Rn ---
Fr 0.7	Ra 0.9	Ac-No 1.1-1.7															

- If the difference in electronegativity values is greater than 1.8, then compound will be ionic
- If the difference in electronegativity values is lesser than 1.8, then compound will be covalent

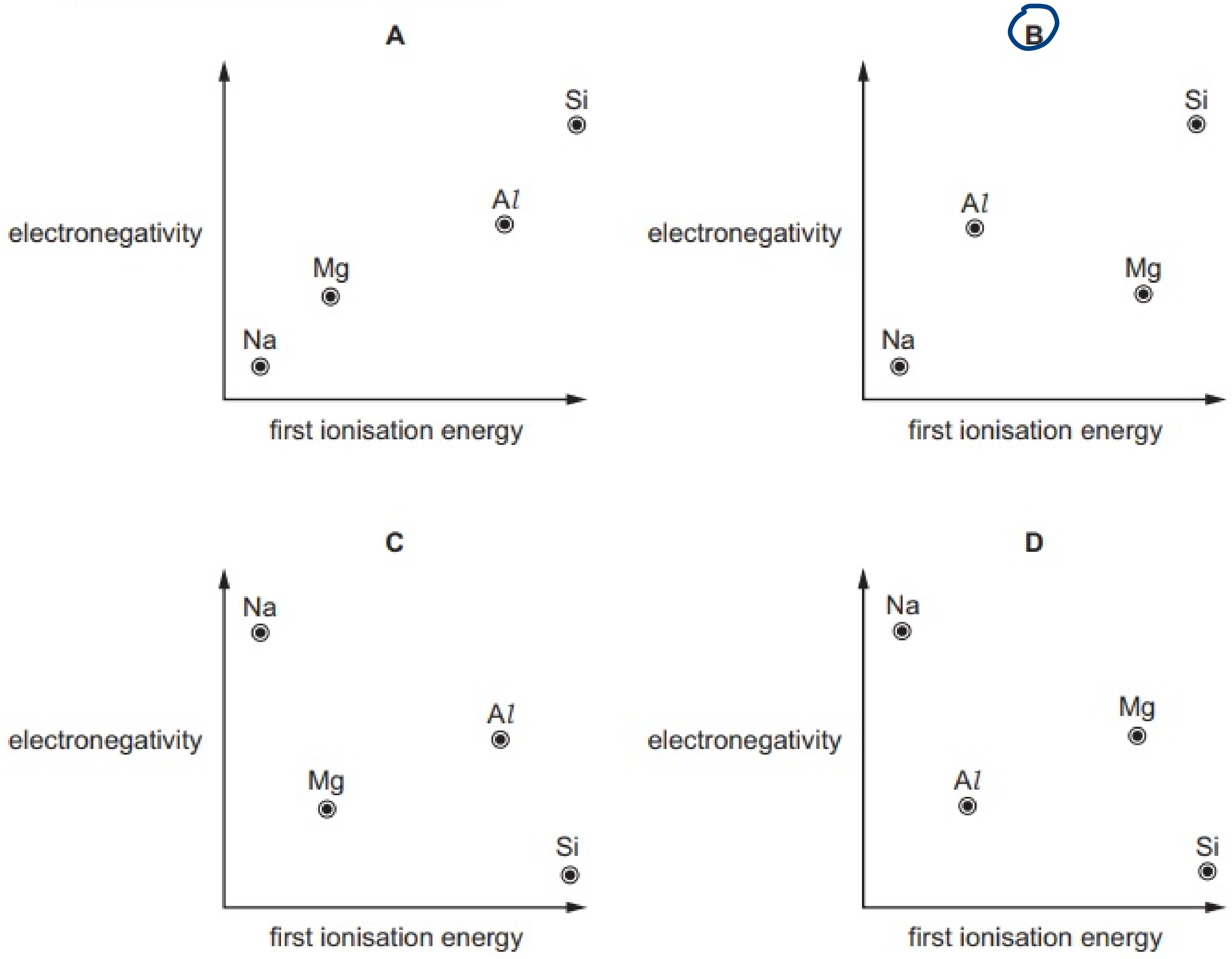
$AlCl_3$: $3 - 1.5 = 1.5$
covalent

AlF_3 : $4 - 1.5 = 2.5$
ionic

32 Which molecules have no overall dipole moment?

1. boron trifluoride ✓
2. methane ✓
3. phosphorus pentafluoride ✓

12 Which diagram correctly shows the electronegativity of the elements Na, Mg, Al and Si plotted against their first ionisation energies?



9 At 200 °C aluminium chloride is a gas with M_r = 267.

What is the number of covalent bonds, dative covalent bonds and lone pairs of electrons in one molecule of aluminium chloride at 200 °C?

	covalent bonds	dative covalent bonds	lone pairs
A	6	2	0
B	6	2	16
C	6	2	18
D	3	0	9

31 Which gaseous molecules are polar?

1 ammonia

2 hydrogen sulfide

3 boron trifluoride

32 A carbon monoxide molecule, CO, has three bonds between the carbon atom and the oxygen atom.

Which features are present in one CO molecule?

1 two lone pairs of electrons

2 a co-ordinate (dative covalent) bond from oxygen to carbon

3 two π bonds

2 Which compound has the least covalent character?

A $\overset{+}{Li}Cl$ B $\overset{2+}{Mg}Cl_2$ C $\overset{+}{K}Cl$ D $\overset{4+}{Si}Cl_4$

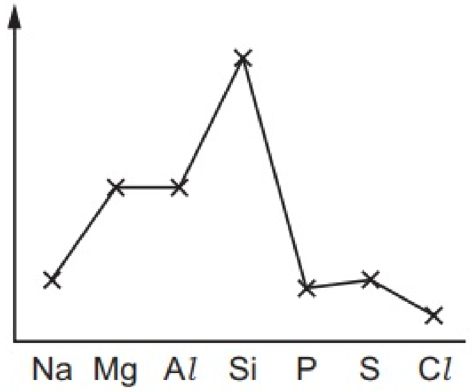
polarisability: charge ↑↓
: radius ↑↓

polarising power: charge ↑↓
radius ↓↑

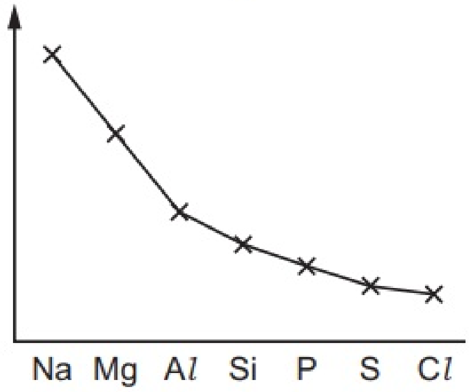
13 The graphs show trends in four physical properties of elements in Period 3, excluding argon.

Which graph has electronegativity on the y-axis?

A



B



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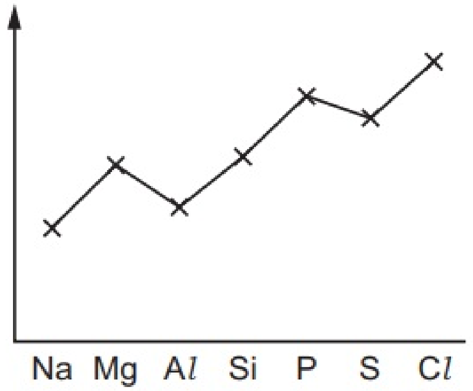
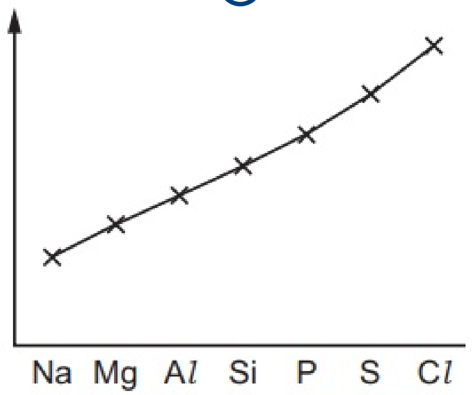


D



Intermolecular forces: Forces of attraction between 1 molecule and a neighbouring molecule. (These are broken in simple molecules when melting and boiling)

Permanent dipoles:- 2 polar molecules will attract one another and the force of attraction is Known as a permanent dipole

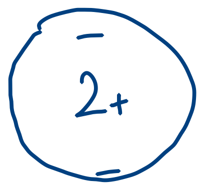
$$\text{i)}\quad \overset{\delta+}{H} — \overset{\delta-}{Cl} \text{------} \overset{\delta+}{H} — \overset{\delta-}{Cl}$$

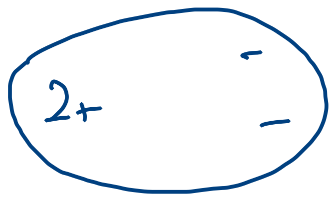
- Molecule must be polar

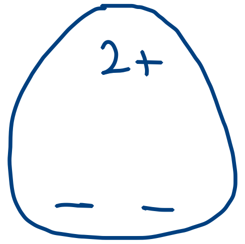
Temporary dipoles:- The temporary force of attraction formed between electrons and nuclei due to electron density not being symmetrical at any instant.

Suppose helium:-

At some instant, helium might be like

1) 

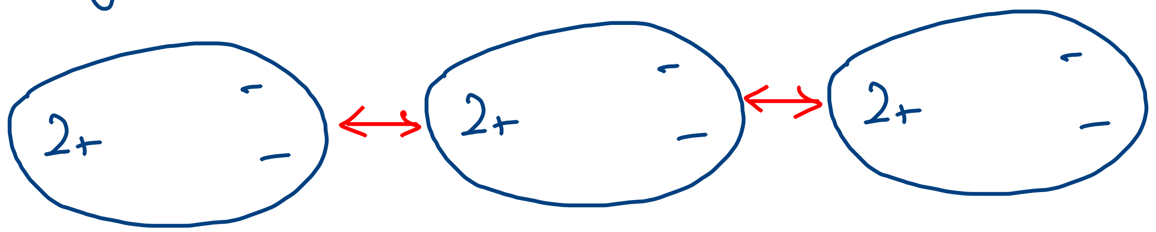
2) 

3) 

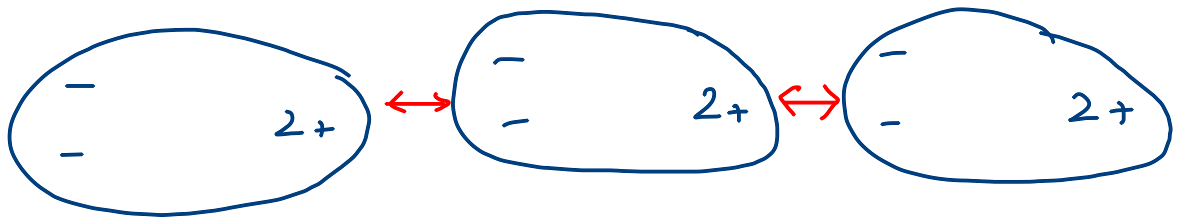
etc

- Temporary dipoles exist in every substance

temporary dipoles can setup like



or



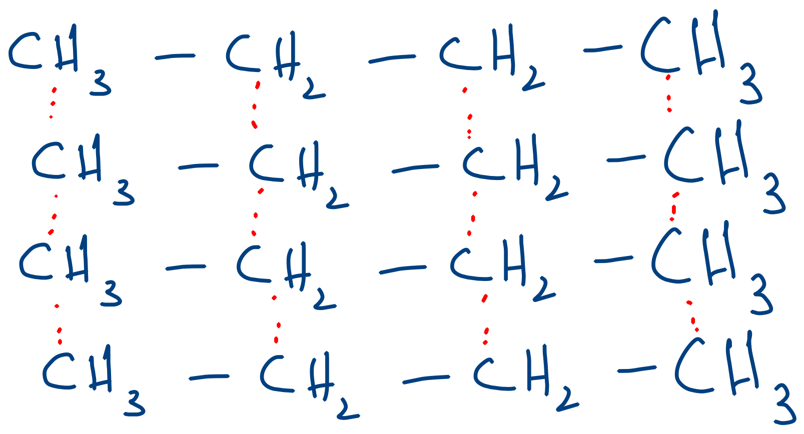
* As the number of electrons increase, the temporary dipole forces get stronger

for e.g.

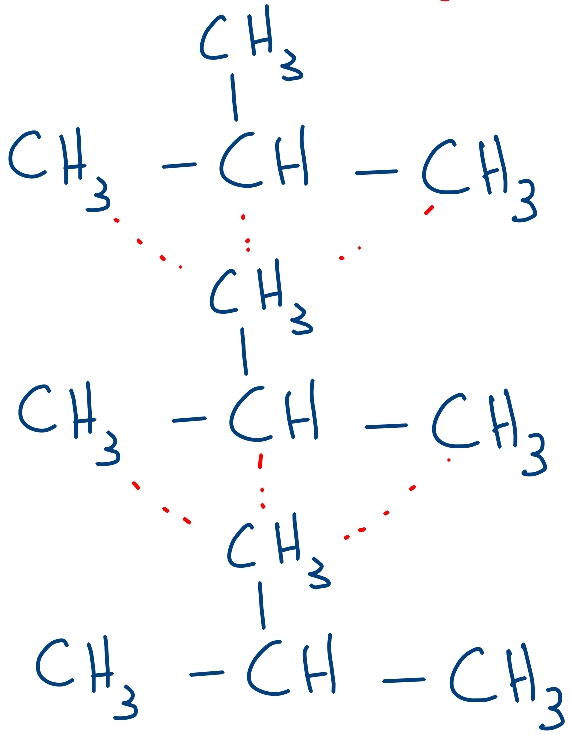
F_2	He	melting point / boiling point increases as number of electrons increase
Cl_2	Ne	
Br_2	Ar	
I_2	Kr	

CH_4	melting point / boiling point increases as number of electrons increase
C_2H_6	
C_3H_8	
C_4H_{10}	
C_5H_{12}	

* Straight chained molecules experience stronger temporary dipoles.



More Surface Area
temporary dipoles stronger
m.p / b.p higher

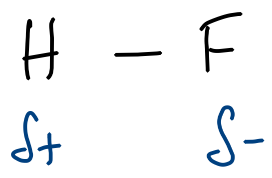


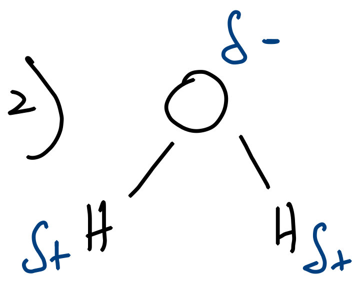
Less surface area
temporary dipoles weaker
m.p / b.p lesser

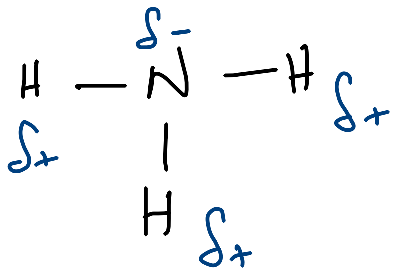
- Temporary dipoles exist in all molecules
- Temporary dipoles are weaker than permanent dipoles
- Collectively, temporary and permanent dipoles are known as Van der Waal's forces

Hydrogen bonding:- Force of attraction between a hydrogen bonded to a very electronegative atom and a neighbouring lone pair

- Hydrogen bonding only occurs in molecules where hydrogen is directly bonded to F or O or N

1) 

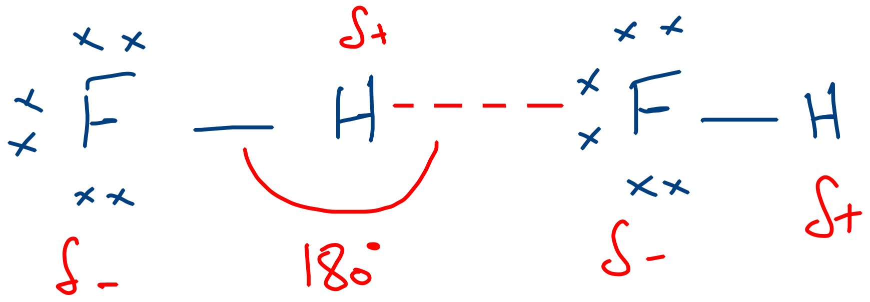
2) 

3) 

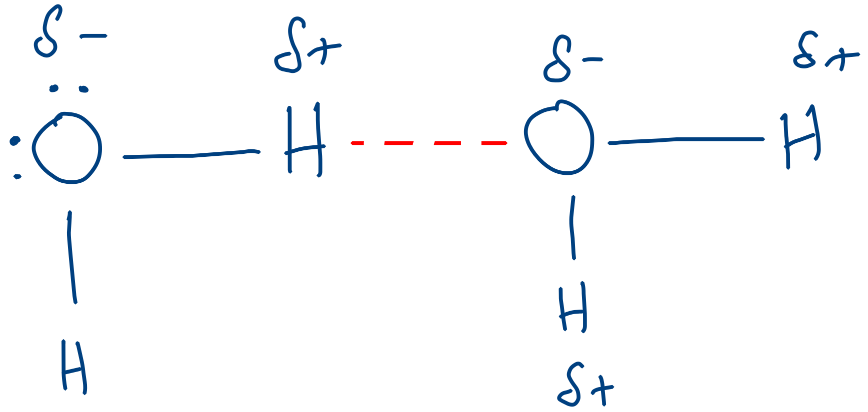
- F, O and N are very electronegative and give H a very significant δ+ charge.

Q. Draw hydrogen bonding between

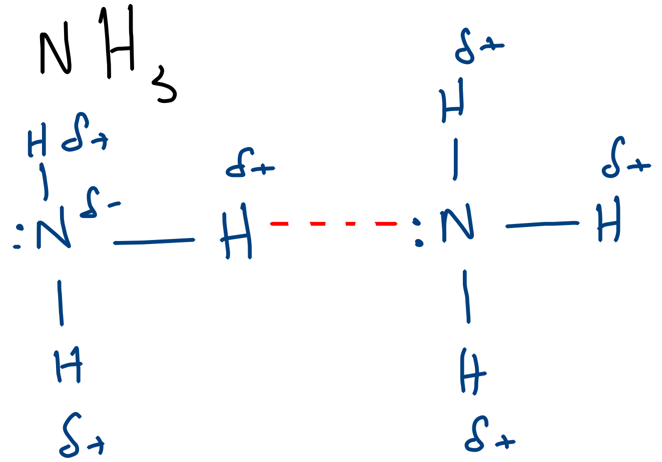
1) H - F



2) H_2O

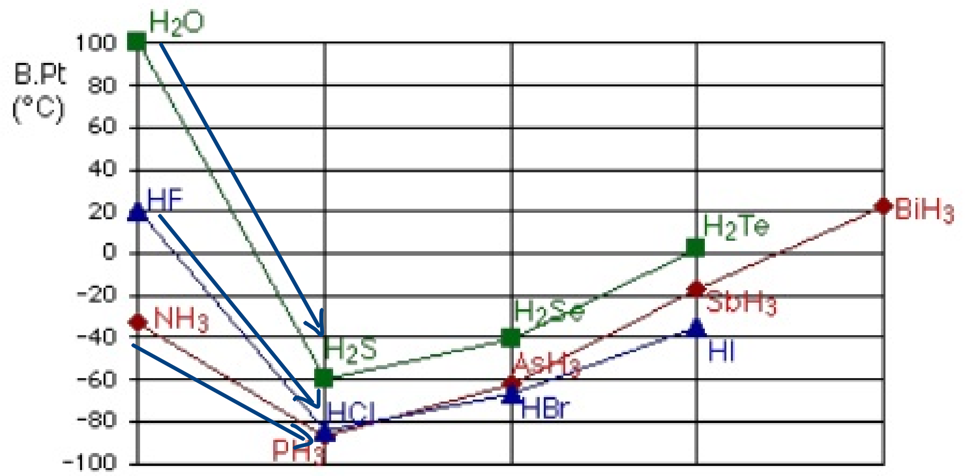


3) NH_3



1) Show hydrogen in middle
2) Show atom bonded to hydrogen in straight line
3) Show atom with lone pair attracting hydrogen also in straight line

∗ Hydrogen bonding is the strongest intermolecular force!



Q. Why does H_2S have a lower boiling point than H_2O?

Ans. H_2O has hydrogen bonding while H_2S can at best only have permanent dipoles

Q. Why do boiling points increase from HCl to HBr to HI?

Ans. The number of electrons increase so stronger temporary dipole forces are achieved.

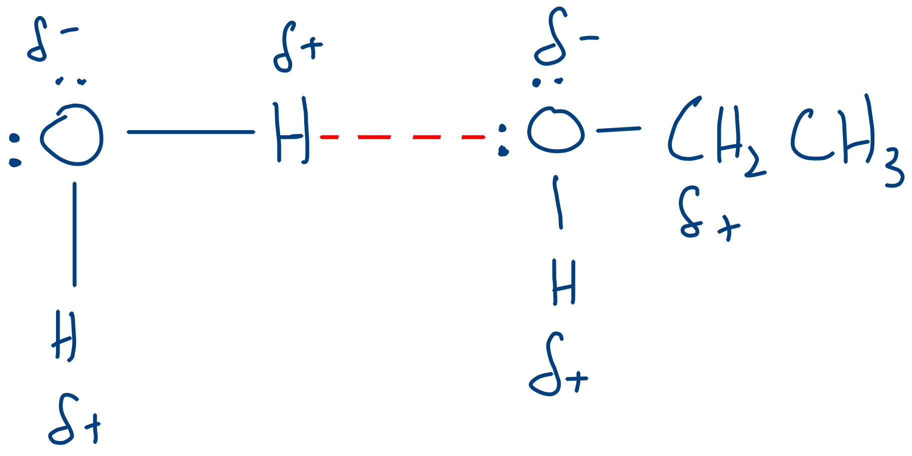
Many of water's properties are due to hydrogen bonding :-

1) Water has high m.p and b.p than other similar isoelectronic molecules → Same electron number

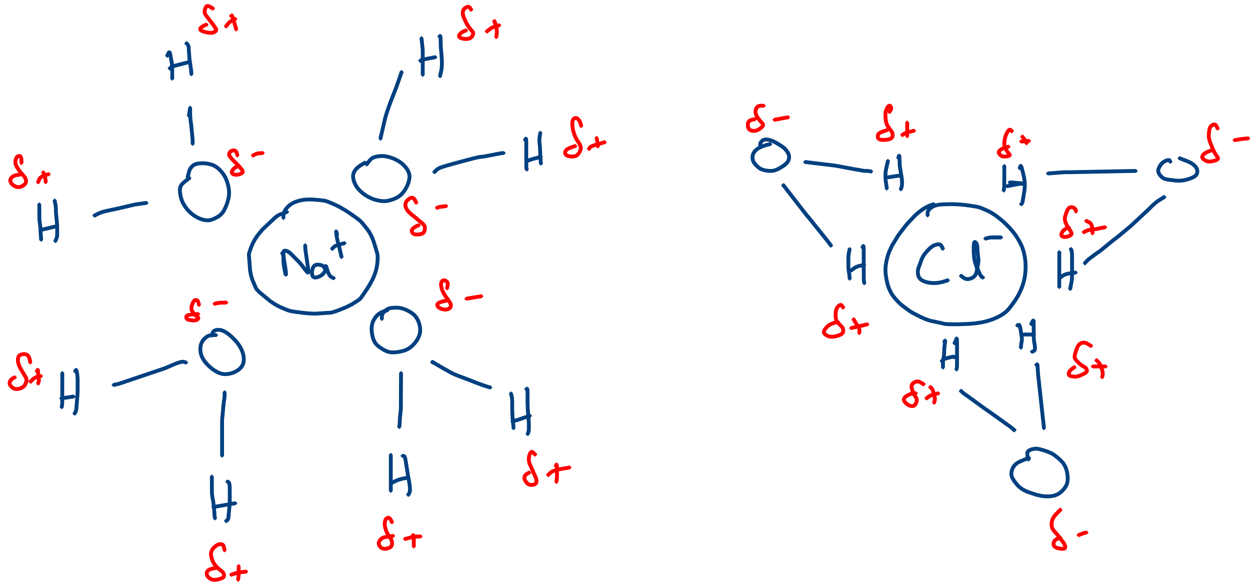
2) Water has high viscosity as hydrogen bonding does not allow molecules to slide over each other easily

3) Water has high surface tension

4) Water is a good solute as it can form hydrogen bonds to other polar molecules and can form other stronger intermolecular forces to ions.



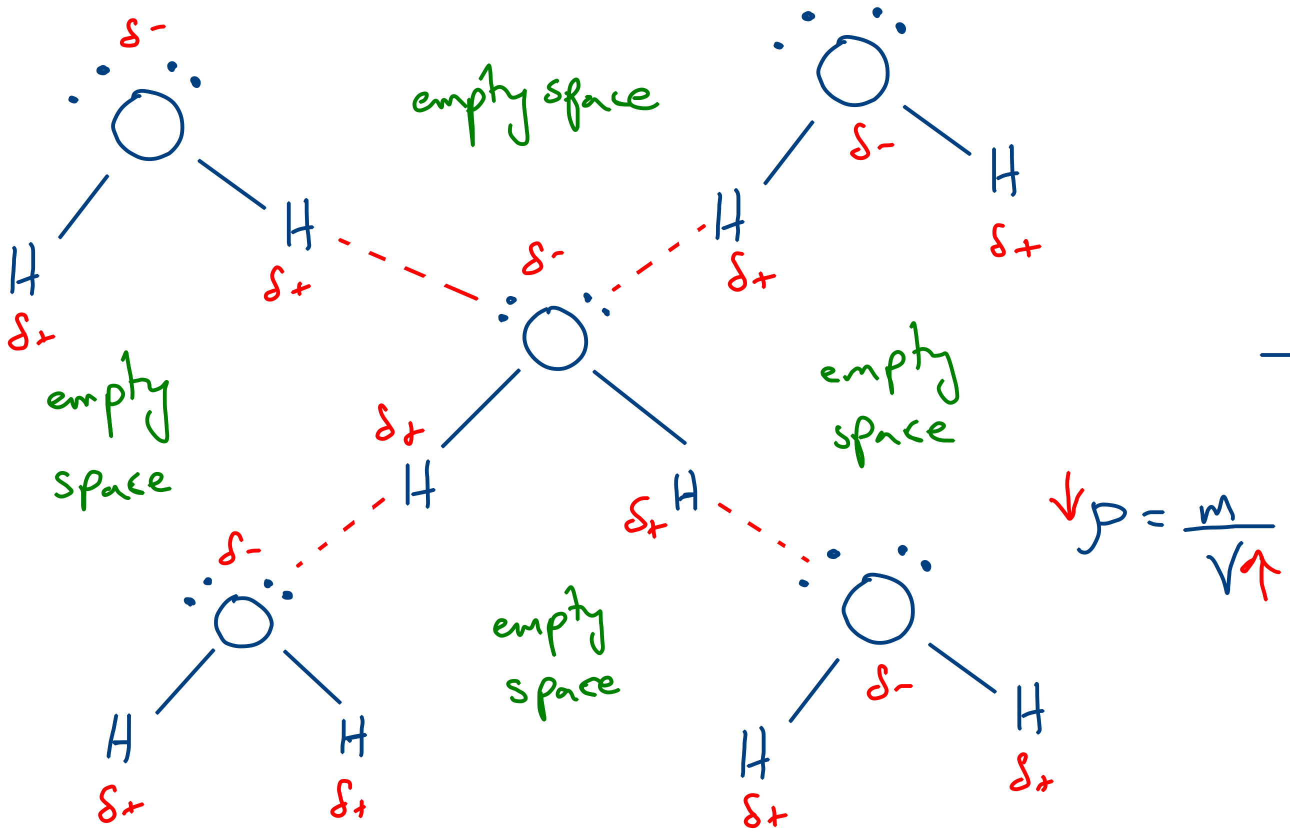
Water dissolving ethanol



Water dissolving NaCl.

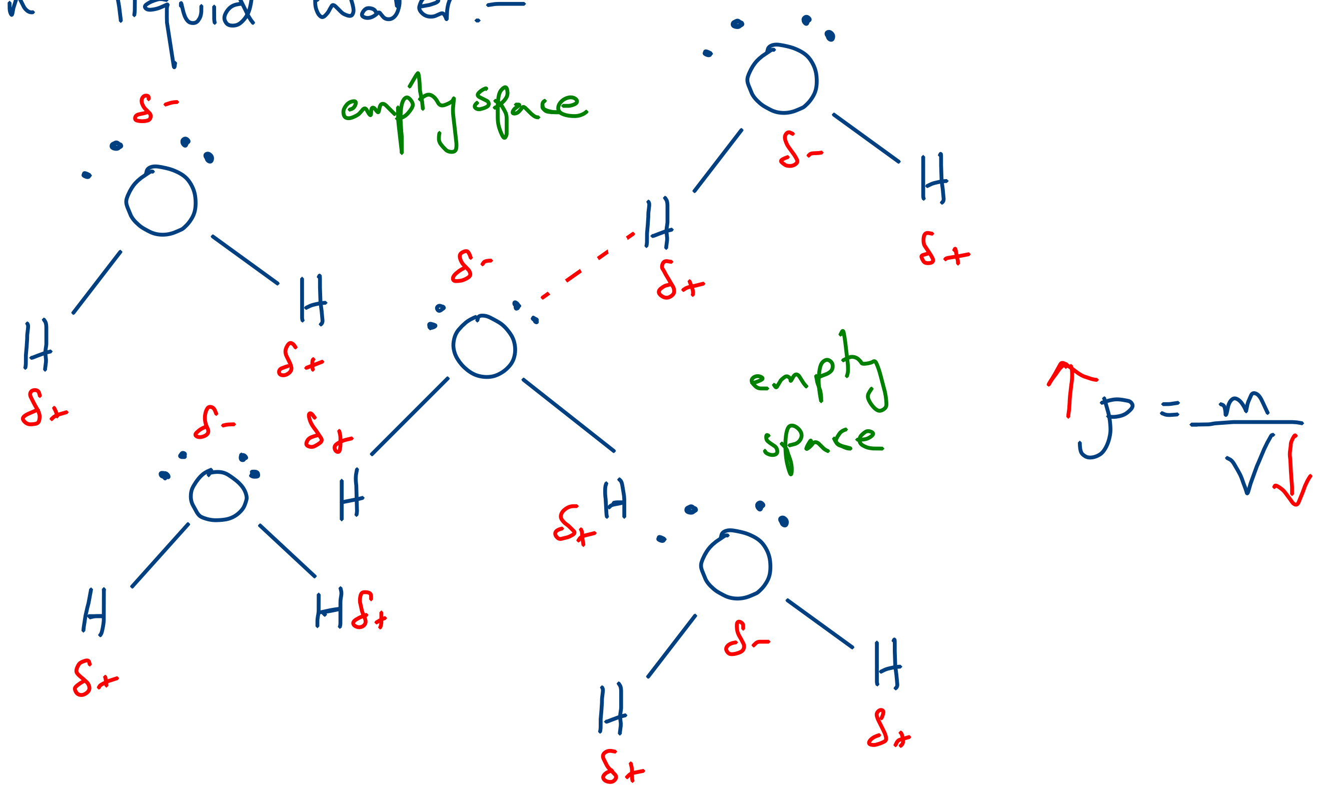
5) The anomalous expansion and ice being less dense than water are also effects of hydrogen bonding

In solid ice:



$\xrightarrow{\text{melts}}$

In liquid water:-



Some of the hydrogen bonds break so molecules are closer together now increasing strength →

* Remember: temporary dipoles < permanent dipoles < hydrogen bonding < giant covalent < ionic & metallic bonding

3 In which pair does the second substance have a **lower** boiling point than the first substance?

temp↓ Permanent↑

A C_2H_6 and C_2H_5Cl

hydrogen↑

B CH_3OCH_3 and C_2H_5OH

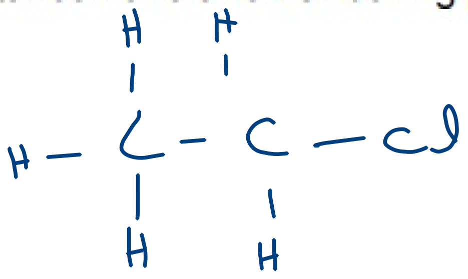
temp↓

C Ne and Ar

less-↓ more-↑

D CH_3NH_2 and C_2H_6

permanent and hydrogen   temp



8 In which change are **only** temporary dipole-induced dipole forces overcome?

A $C_2H_5OH(l) \rightarrow C_2H_5OH(g)$ ✗

B $H_2O(s) \rightarrow H_2O(l)$ ✗ } hydrogen bonding

C $O_2(s) \rightarrow O_2(l)$ ✓

D $C_4H_{10}(l) \rightarrow C_4H_{10}(s)$ not overcome

5 Which compound has the highest melting point?

A H_2O **B** NH_3 **C** HCl **D** I_2

has 2 lone pairs so can form more hydrogen bonds

only 1 lone pair

2 Which physical properties of water are due to hydrogen bonding?

1 Water has a higher boiling point than H_2S. ✓

2 Ice floats on water. ✓

3 The H-O-H bond angle in water is approximately 104°. ✗

6 The gecko, a small lizard, can climb up a smooth glass window. The gecko has millions of microscopic hairs on its toes and each hair has thousands of pads at its tip. The result is that the molecules in the pads are extremely close to the glass surface on which the gecko is climbing.

What is the attraction between the gecko's toe pads and the glass surface?

A co-ordinate bonds **C** ionic bonds

B covalent bonds **D** van der Waals

14 Which statement explains why iodine is less volatile than chlorine?

A Chlorine is more electronegative than iodine and so has more repulsion between its molecules.

B The greater number of electrons in iodine leads to larger temporary dipole-induced dipole forces.

C The I–I bond energy is smaller than the Cl–Cl bond energy.

D The iodine molecules have stronger permanent dipole-permanent dipole forces.

5 Descriptions of the bonding in three substances are given.

substance 1 strong covalent bonds between atoms, permanent dipole-dipole attractions between molecules

substance 2 strong covalent bonds between atoms, weak forces between molecules

substance 3 strong covalent bonds between atoms, hydrogen bonding between molecules

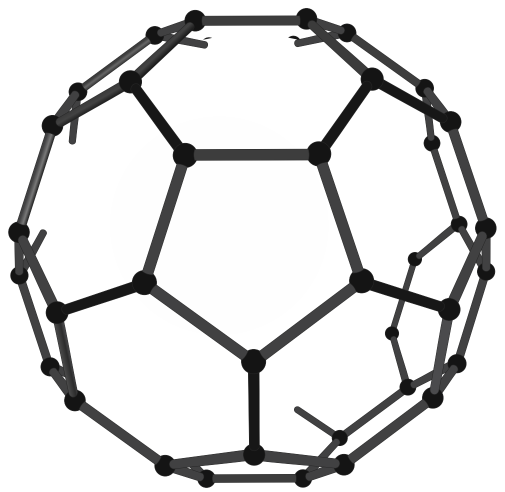
Which compounds could be substances 1, 2 and 3?

	substance 1	substance 2	substance 3
A	CH_3OH ✗	Al_2Cl_6 ✓	CH_2Cl_2 ✗
B	Al_2Cl_6 ✗	CH_2Cl_2 ✓	CH_4 ✗
C	CH_2Cl_2 ✓	CH_4 ✓	CH_3OH ✓
D	CH_4 ✗	CH_3OH ✓	H_2O ✓

Simple molecular structures :- (Small covalent molecules)

- Molecules such as H_2O, CO_2, Cl_2, I_2 exist as simple molecular structures.
- Held together by weak intermolecular forces
- Upon melting or boiling, these weak intermolecular forces are overcome (low m.p and b.p)
- Donot conduct electricity
- Usually insoluble in water
- In the solid state, they arrange in a regular lattice, held together by weak intermolecular forces.

Buckminster fullerene (C_{60}) :-

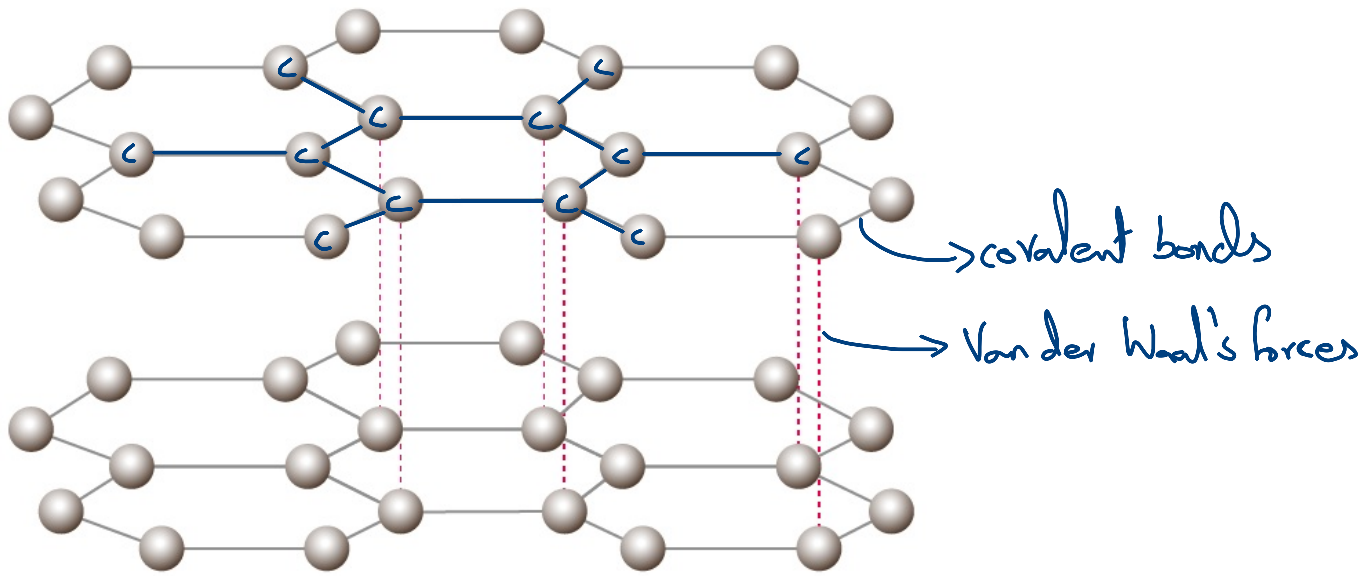


- Simple molecular
- Carbon atoms have pentagonal and hexagonal arrangements (1 C atom is bonded to 3 C atoms)
- 1 delocalized electron per carbon but it still cannot conduct as the electron cannot move from atom to atom → does not conduct electricity
- temporary dipoles holding together → low m.p. and b.p.
- insoluble in water
- allotrope of carbon

Giant Covalent Structures :- (Big covalent molecules)

- Substances such as diamond, graphite, SiO_2, graphene and nano tubes exist as giant structures.
- Held together by many covalent bonds
- Upon melting or boiling, covalent bonds are broken.
- May or may not conduct electricity
- Insoluble in water
- In solid state, they arrange in a giant lattice, held together by strong covalent bonds.

Graphite :-



- Giant covalent
- Carbon atoms have hexagonal arrangement (1 C is bonded to 3C)
- 1 delocalized electron per carbon and it can move from atom to atom so conducts electricity
- covalent bonds holding together (high m.p and b.p)
- insoluble in water
- Sheets can slide over each other due to van der Waals forces
- allotrope of carbon

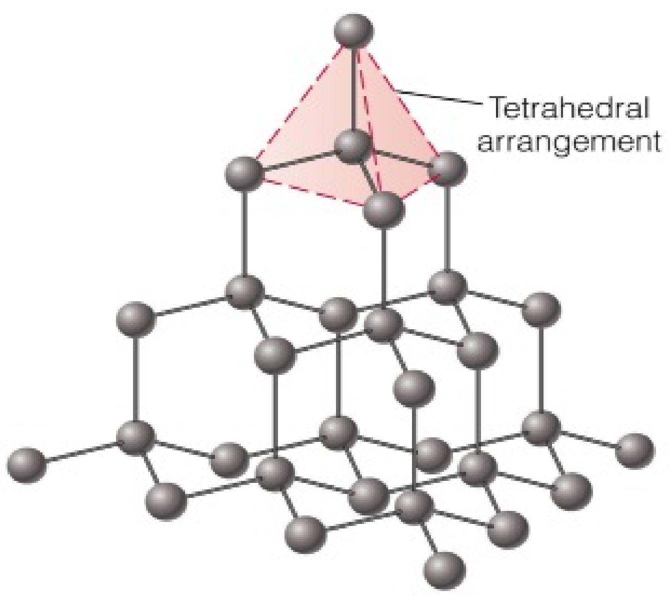
Graphene: It is just a single sheet of graphite.

- Same properties as graphite

Nanotubes:- It is a graphene sheet rolled up

- Same properties as graphite and graphene
- Highest strength to weight ratio of any substance.

Diamond:-



- Giant covalent
- Carbon atoms in tetrahedral arrangement (1C bonded to 4C)
- No delocalised electrons → so does not conduct electricity
- Covalent bonds holding together (high m.p and b.p)
- insoluble in water
- Hardest material known
- Allotrope of carbon.

Forces overcome during melting and boiling in different materials:-

Simple molecular: intermolecular forces
Giant covalent: Covalent bonds
Ionic bonding: ionic bonds
Metals: metallic bonds

9 Materials can be classified by their chemical structures. Four common types of structure are metallic, ionic, simple molecular and giant molecular.

Some physical properties of four substances are shown in the table.

Which substance has a simple molecular structure?

	melting point / °C	effect of adding water	electrical conductivity
A	64	reacts	good when solid
B	113 ✓	insoluble ✓	always poor ✓
C	767	soluble	good when solid
D	1600	insoluble	always poor

5 The table shows some properties of four substances.

Which substance could be potassium iodide?

	melting point of solid / °C	electrical conductivity when molten
A	–66	poor
B	–39	good
C	680	good
D	1600	poor

7 Which pair of substances are both simple molecular?

A C_{60} and graphene

B C_{60} and iodine

C graphene and graphite

D graphite and iodine

33 Carbon exists in several different forms. Two of these forms are buckminsterfullerene and graphene. Buckminsterfullerene is a fullerene allotrope of carbon.

Which statements about buckminsterfullerene and graphene are correct?

1 Both have delocalised electrons. ✓

2 Buckminsterfullerene has a giant molecular structure. ✗

3 The carbon atoms in graphene form a tetrahedral lattice. ✗

D

3 Solid carbon dioxide, CO_2, is similar to solid iodine, I_2, in its structure.

Which statement about solid CO_2 and solid SiO_2 is correct?

A Both solid CO_2 and solid SiO_2 exist in a lattice structure.

B Both solid CO_2 and solid SiO_2 have a simple molecular structure.

C Both solid CO_2 and solid SiO_2 have atoms joined by single covalent bonds.

D Both solid CO_2 and solid SiO_2 change spontaneously to gas at s.t.p..

33 Carbon can exist as allotropes which include graphite, diamond and a fullerene.

Which statements are correct?

1 All three allotropes contain covalent bonds. ✓

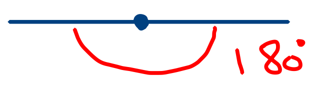
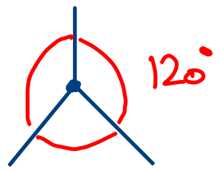
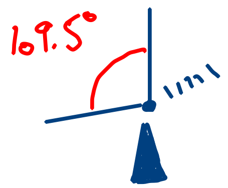
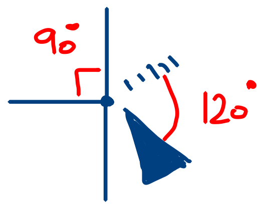
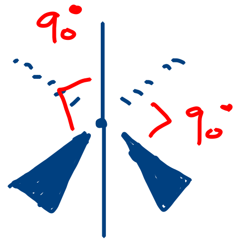
2 All three allotropes are giant molecular. ✗

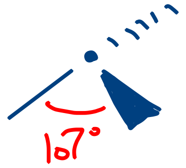
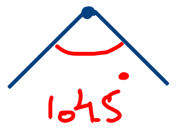
3 All three allotropes have delocalised electrons. ✗

D

Shapes of molecules:-

Shapes are deduced using the VSEPR theory
→ electron pairs around an atom repel each other
→ bonding and lone pairs arrange themselves to be as far apart as possible due to repulsion.

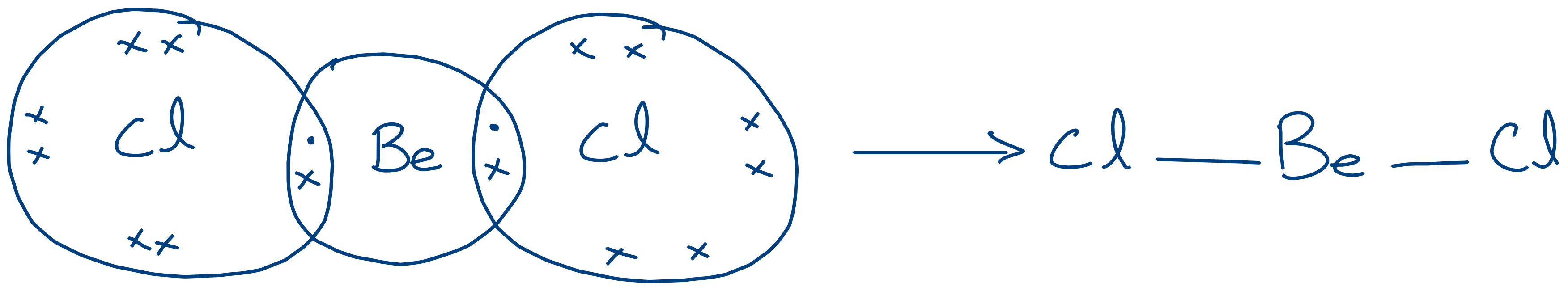
No of e^- pairs	arrangement	diagram	Extra details
1) 2 bonding	linear		- In one plane - Also for single bonds
2) 3 bonding	trigonal planar		- In one plane
3) 4 bonding	tetrahedral		- 3d - rotational symmetry
4) 5 bonding	trigonal bipyramidal		- 3d - 90° between equator and axis - 120° in equator plane
6) 6 bonding	Octahedral		- 3d - All angles 90°

No of e^- pairs	arrangement	diagram	Extra details
7) 3 bonding + 1 lone	trigonal pyramidal		- 3d
8) 2 bonding + 2 lone	V-shaped / bent		- In one plane only.

Drawing structures of molecules

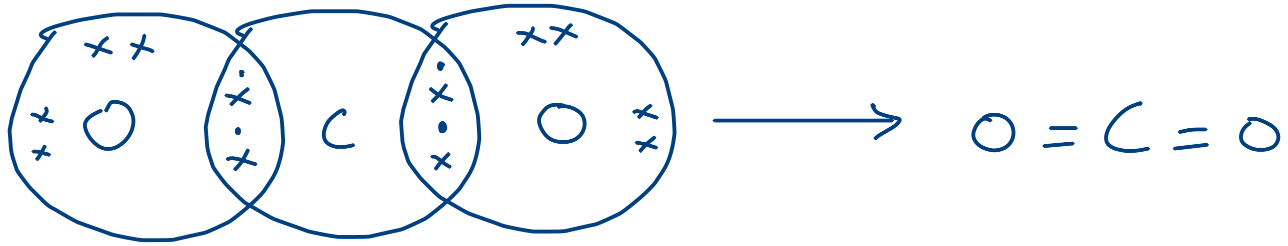
1) Identify central atom
2) Draw dot and cross diagram
3) Deduce structure.
4) Consider double and triple bonds as single bonds

1) $BeCl_2$



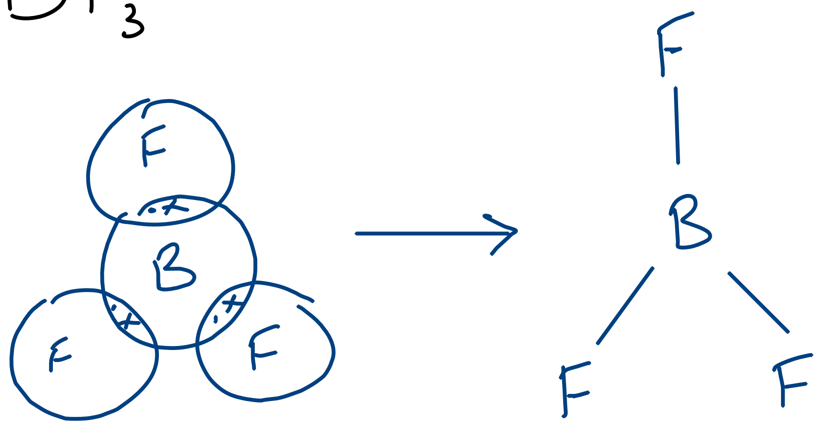
Shape: linear Bond angle: 180°

2) CO_2



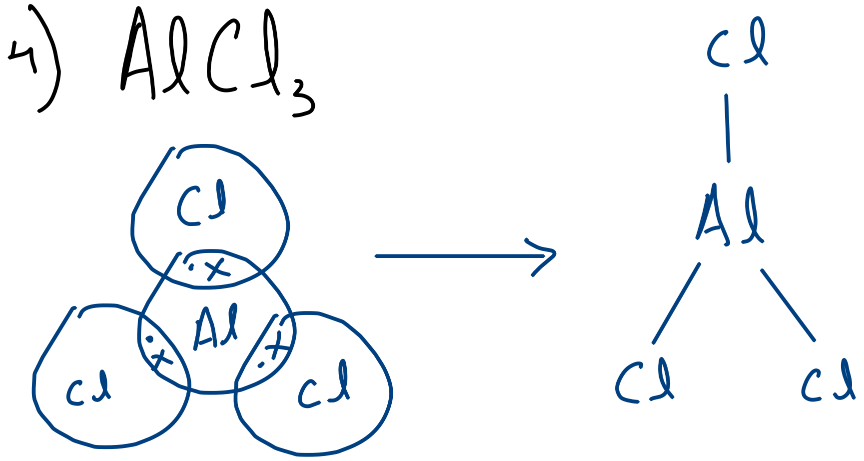
Shape: linear Bond angle: 180°

3) BF_3



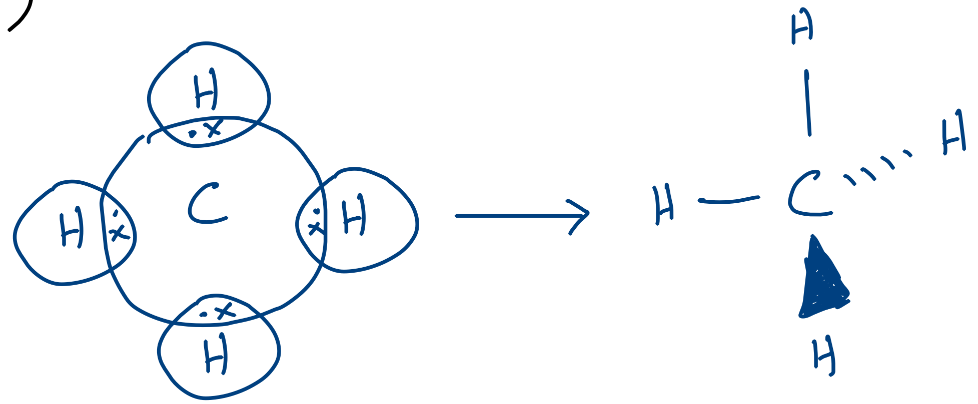
Shape: trigonal planar Bond angle: 120°

4) $AlCl_3$



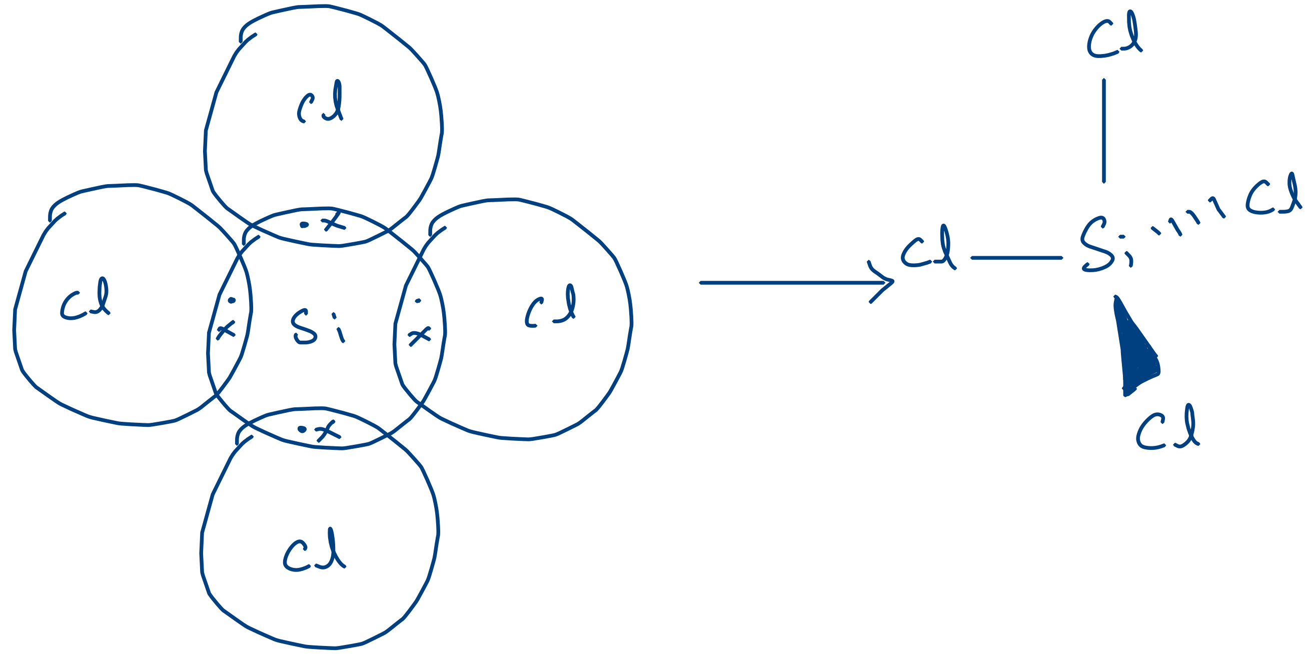
Shape: trigonal planar Bond angle: 120°

5) CH_4



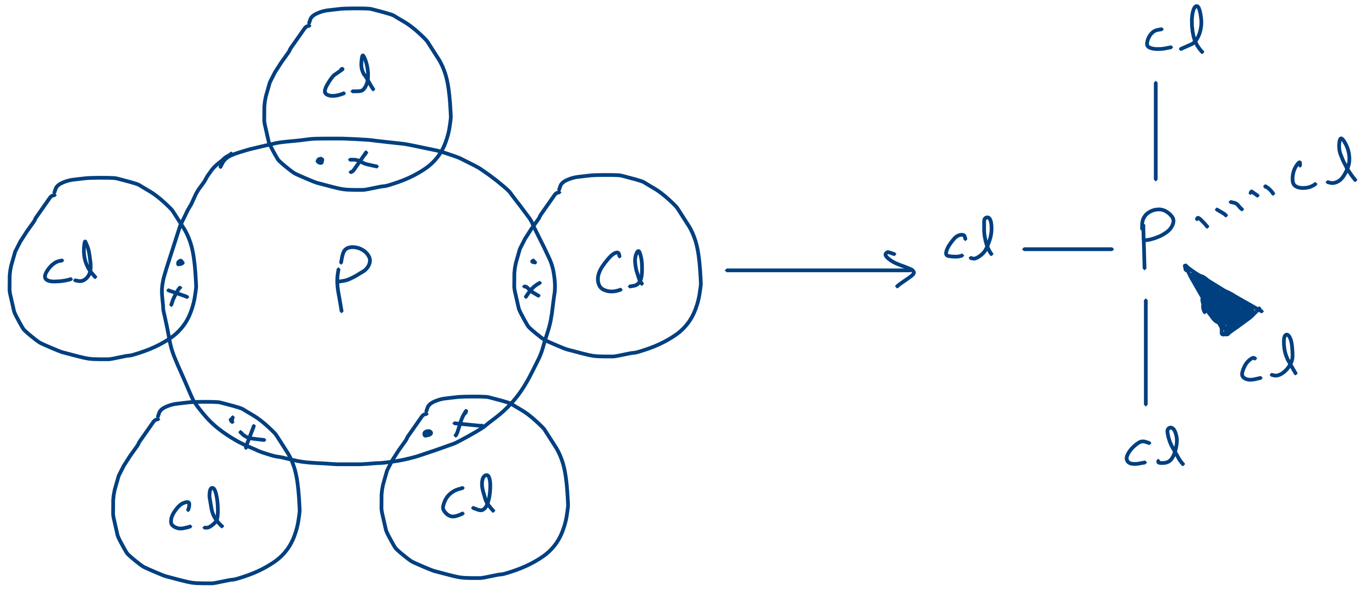
Shape: tetrahedral   Bond angle: 109.5°

6) $SiCl_4$



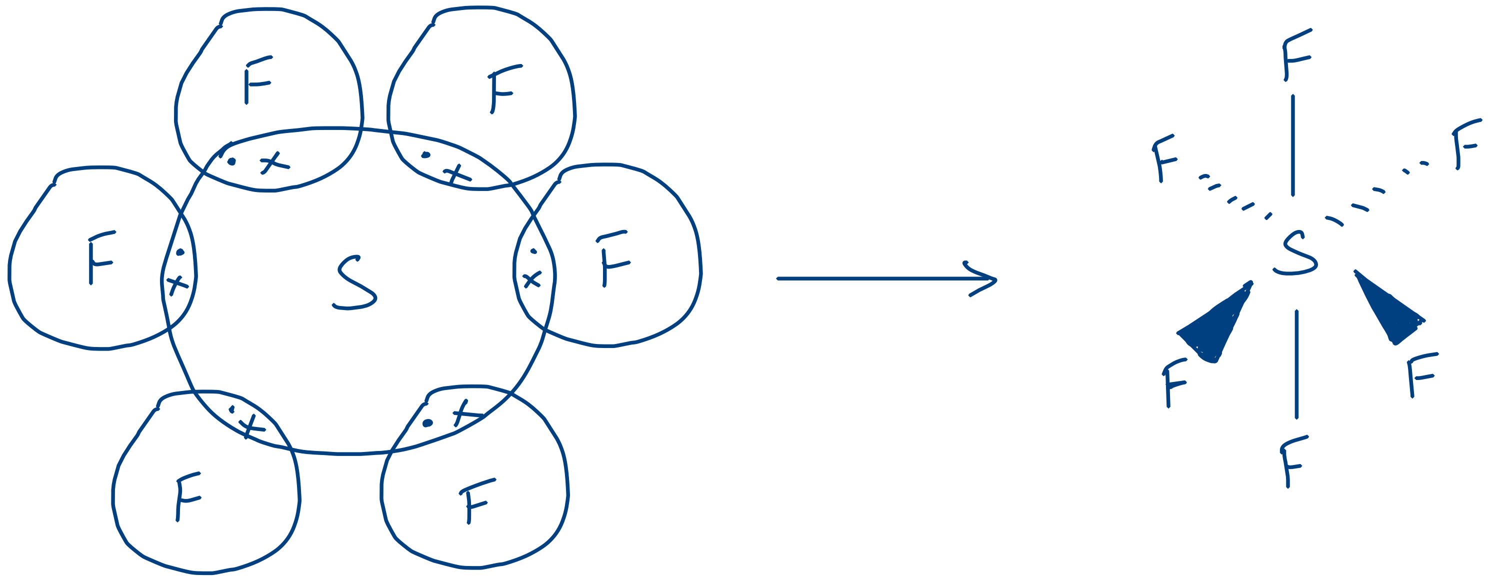
Shape: tetrahedral   Bond angle: 109.5°

7) PCl_5



Shape: Trigonal bipyramidal   Bond angle: 120° & 90°

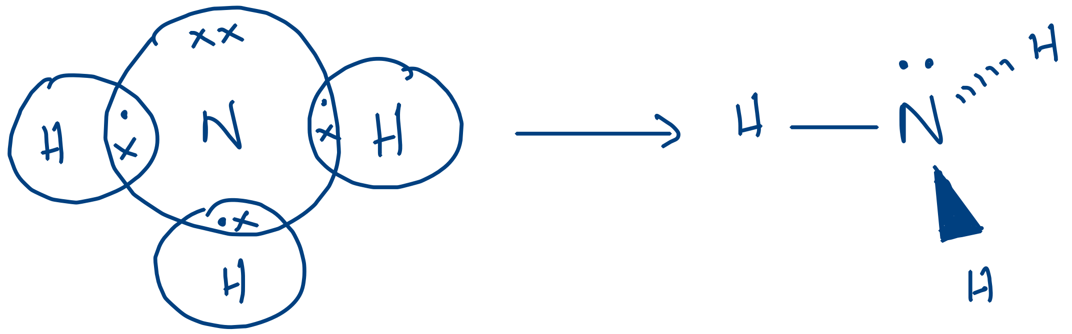
8) SF_6



Shape: Octahedral   Bond angle: 90°

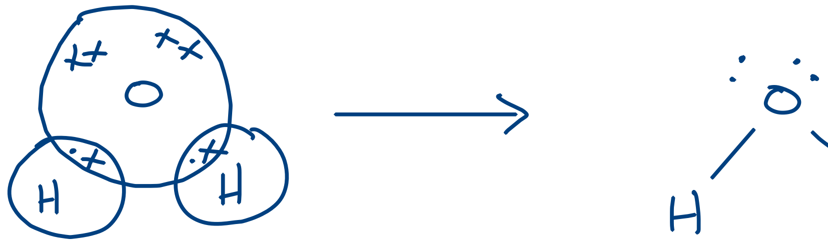
* Phosphorus and Sulphur can accept more electrons into empty d orbitals and sustain more than an octet configuration.

9) NH_3



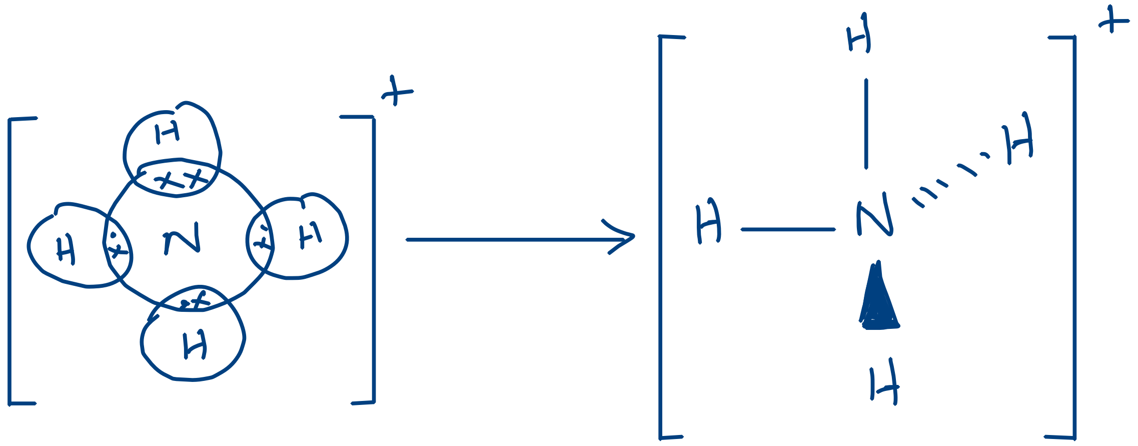
Shape: trigonal pyramidal   Bond angle: 107°

10) H_2O



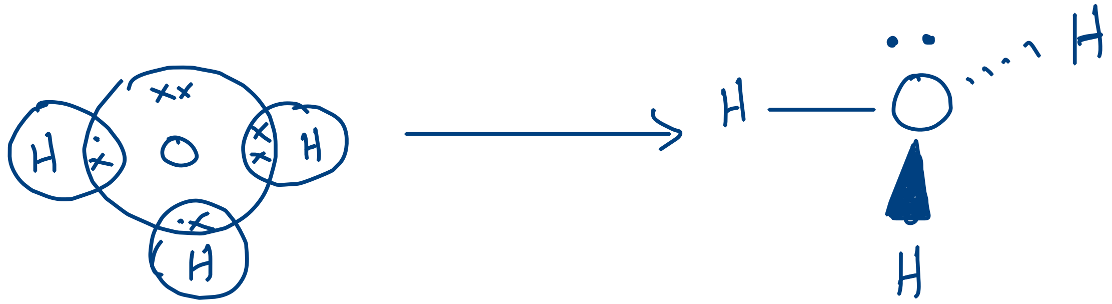
Shape: V shaped / bent   Bond angle: 104.5°

11) NH_4^+



Shape: tetrahedral   Bond angle: 109.5°

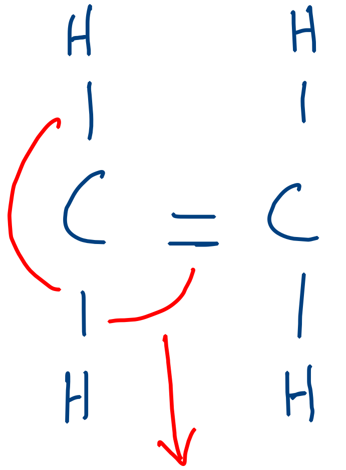
12) H_3O^+



Shape: trigonal pyramidal   Bond angle: 107°

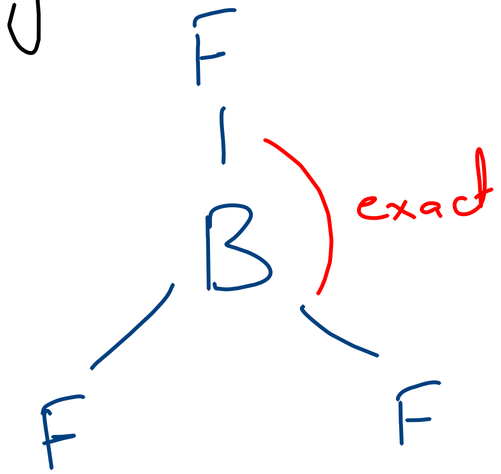
Multiple bonds vs. Single bonds

little less than 120°



little more than 120°

vs.



exactly 120°

Order of repulsion :-

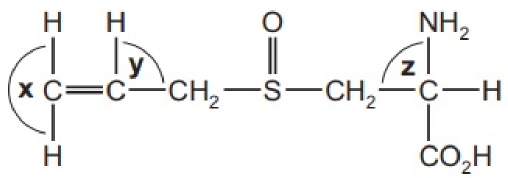
least ... highest

bonding - bonding < bonding - lone < lone - lone

single bond < double bond < triple bond.

4 The structural formula of alliin is shown.

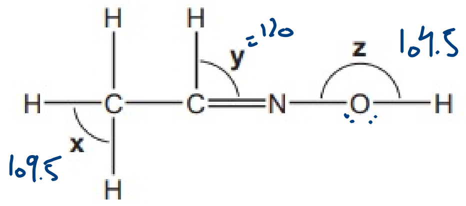
alliin



What are the approximate bond angles **x**, **y** and **z** in a molecule of alliin?

	x	y	z
A	90°	90°	109°
B	120° ✓	109°	90°
C	120° ✓	120° ✓	109° ✓
D	180°	109°	109°

4 Ethanal reacts with hydroxylamine, NH_2OH, to form the molecule shown.



What is the order of **increasing** bond angle in this structure from smallest to largest?

A z, x, y B y, z, x C x, z, y D z, y, x

4 In which set do all the molecules have all their atoms arranged in one plane?

A $AlCl_3$, BF_3, PH_3

B $AlCl_3$, CO_2, NH_3

C BF_3, C_2H_4, C_3H_6

D C_2H_4, CO_2, H_2O

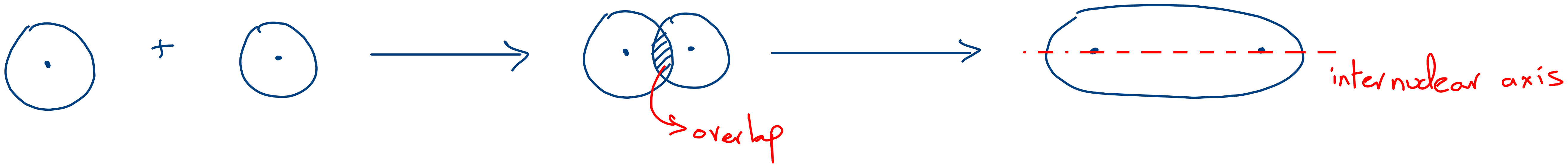
1 In which of the following molecules are the bond angles equal to 120°?

1 BF_3 ✓ 2 NCl_3 ✗ 3 H_3O^+ ✗

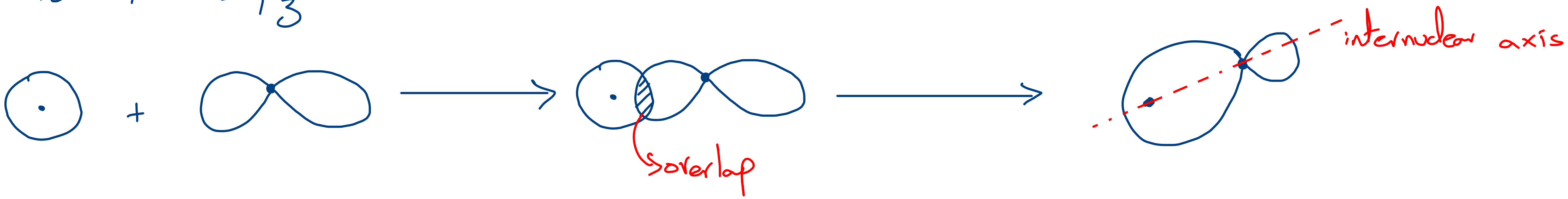
Orbital Overlap :-

1) A single bond is made when 2 orbitals overlap end to end —— this is known as a σ bond

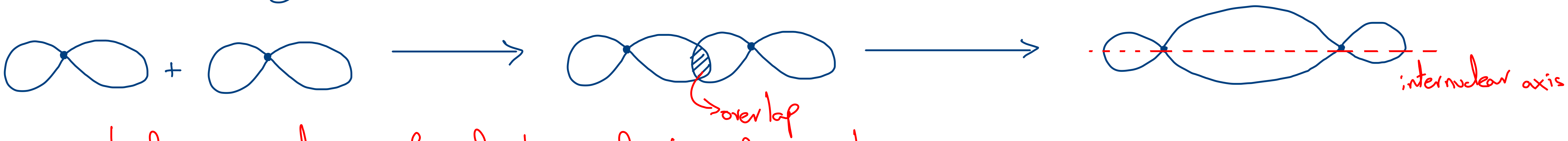
a) 1s + 1s



b) 1s + $2p_z$



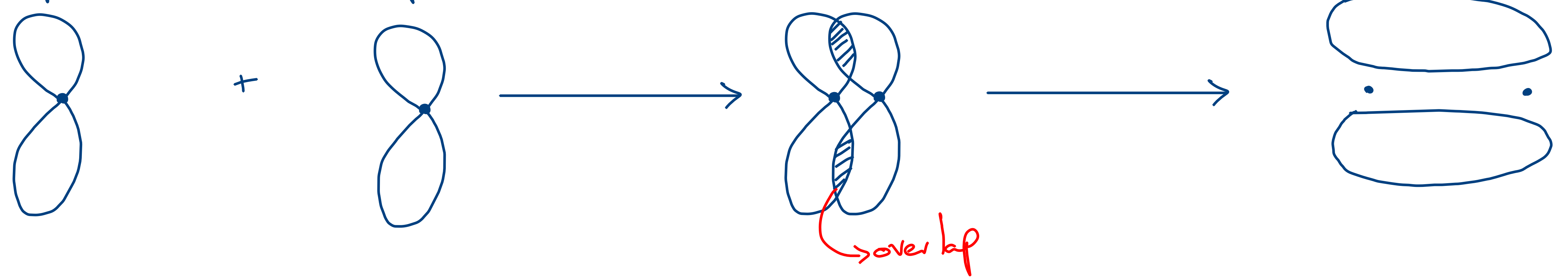
c) $2p_z$ + $2p_z$



- σ bonds are always formed by end to end overlap
- e^- density lies along internuclear axis.
- 1 region of overlap
- stronger than π bonds

2) A pair of parallel p orbitals can overlap sideways to produce a π bond

a) $2p_x$ + $2p_x$

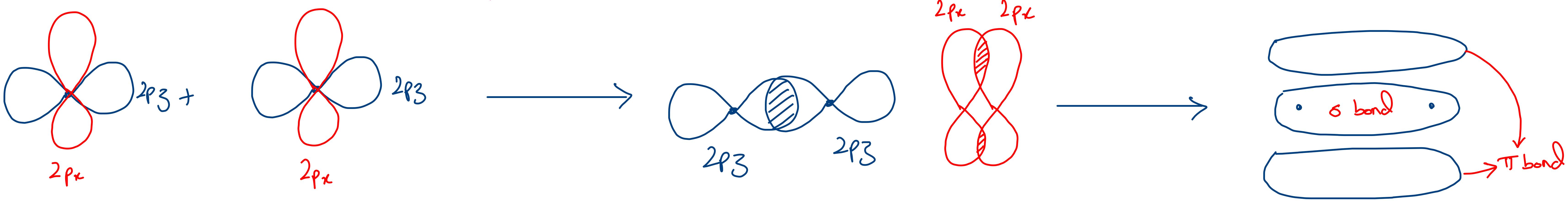


- π bonds are always formed by sideways overlap
- e^- density lies above and below internuclear axis (that is why they are weaker than σ bonds)
- 2 region's of overlap
- only present in double and triple bonds (a σ bond has to be made before a π bond)
- prevents rotation about itself so gives rise to isomerism in alkenes.

- A single bond is always a σ bond
- A double bond is a σ bond and π bond
- A triple bond is a σ bond and 2π bonds

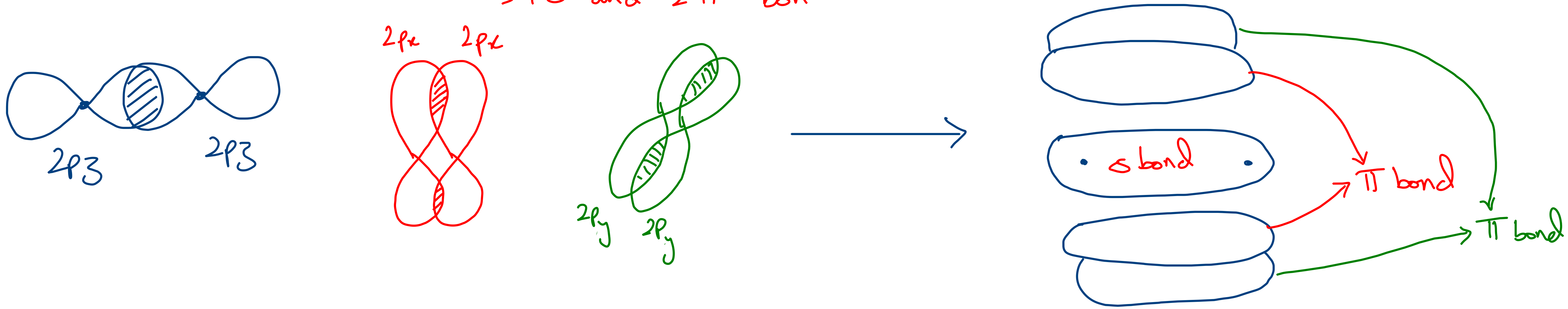
1) Orbital overlap in O_2 $(O=O)$ [2p orbitals overlap) ↳ 1 σ and 1 π bond

2p: $\boxed{\uparrow\downarrow}\boxed{\uparrow}\boxed{\uparrow}$



2) Orbital overlap in N_2 $(N \equiv N)$ [2p orbitals overlap) ↳ 1 σ and 2 π bon

2p: $\boxed{\uparrow}\boxed{\uparrow}\boxed{\uparrow}$



Hybridization:-

In the previous examples, there wer enough electrons available in the orbitals overlapping to create the bonds

But for example carbon, it never has enough electrons to form bonds, so how does it form bonds?

Atoms that donot have enough electrons in their orbitals overlapping, go under hybridization.

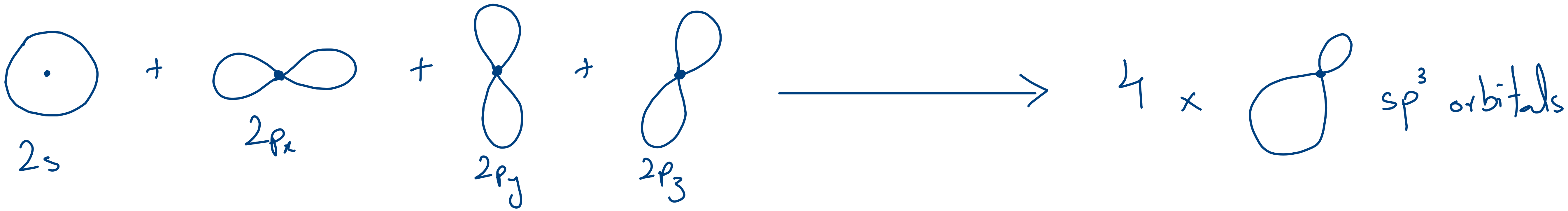
Sp³ hybridization:

The carbon in methane forms 4 bonds but

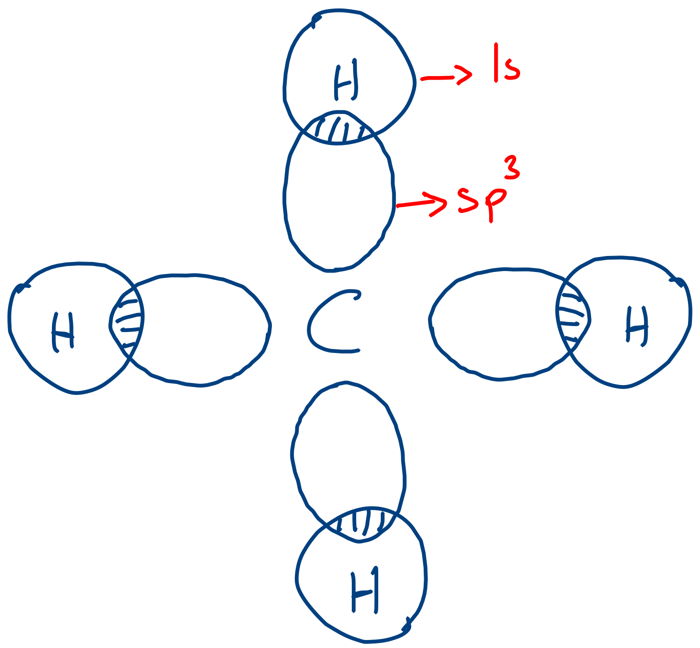
C: $\underset{1s^2}{\boxed{\uparrow\downarrow}}$ $\underset{2s^2}{\boxed{\uparrow\downarrow}}$ $\underset{2p^2}{\boxed{\uparrow}\boxed{\uparrow}\boxed{}}$ → The 4 electrons aren't present! $\xrightarrow[\text{We need 4 electrons in outermost subshell}]{\text{hybridizes}}$ 2s & 3x 2p combine

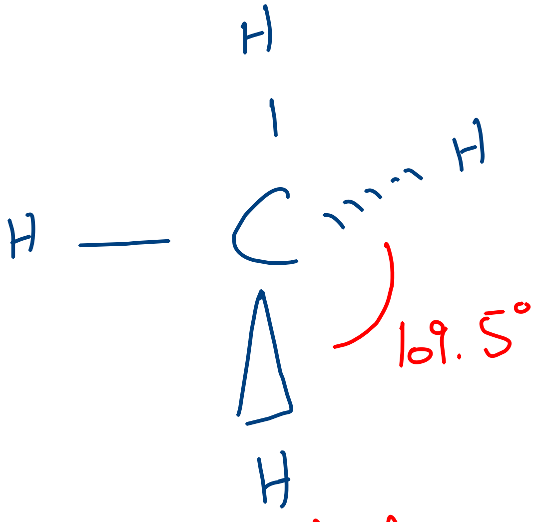
4 sp³ hybridized orbitals

$\underbrace{\boxed{\uparrow}}_{1\times 2s \text{ orbital}}\underbrace{\boxed{\uparrow}\boxed{\uparrow}\boxed{\uparrow}}_{3\times 2p \text{ orbitals}}$



Now carbon can bond to 4 hydrogen atoms via σ bonds.



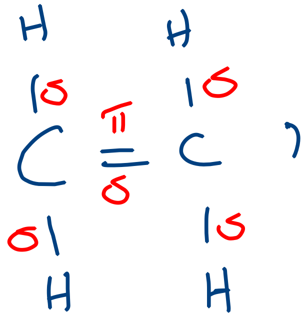


tetrahedral

- sp^3 hybridization is done when 4 single bonds are to be made (4 σ bonds)
- Usually in alkanes

Sp^2 hybridization :-

In ethene (C_2H_4)



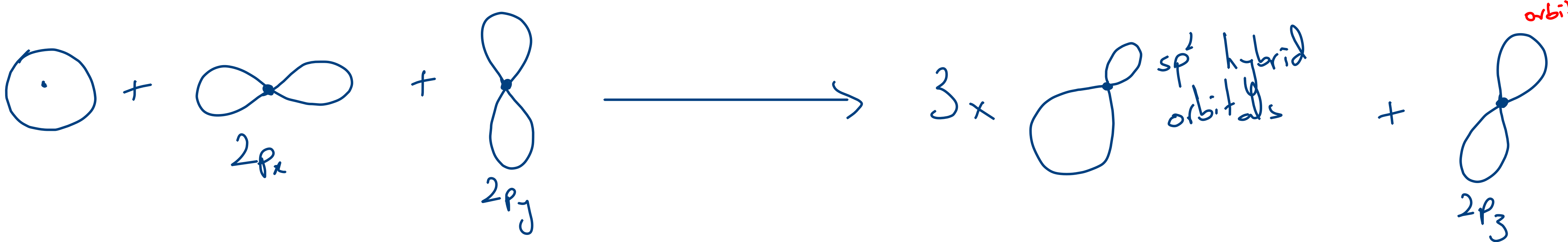
, the carbon atom needs to form 3 single (σ bonds) (need 3 orbitals for overlap) and 1 π bond (need 1p orbital to sideways overlap)

C: [↑↓] $1s^2$ [↑↓] $2s^2$ [↑][↑][ ] $2p^2$

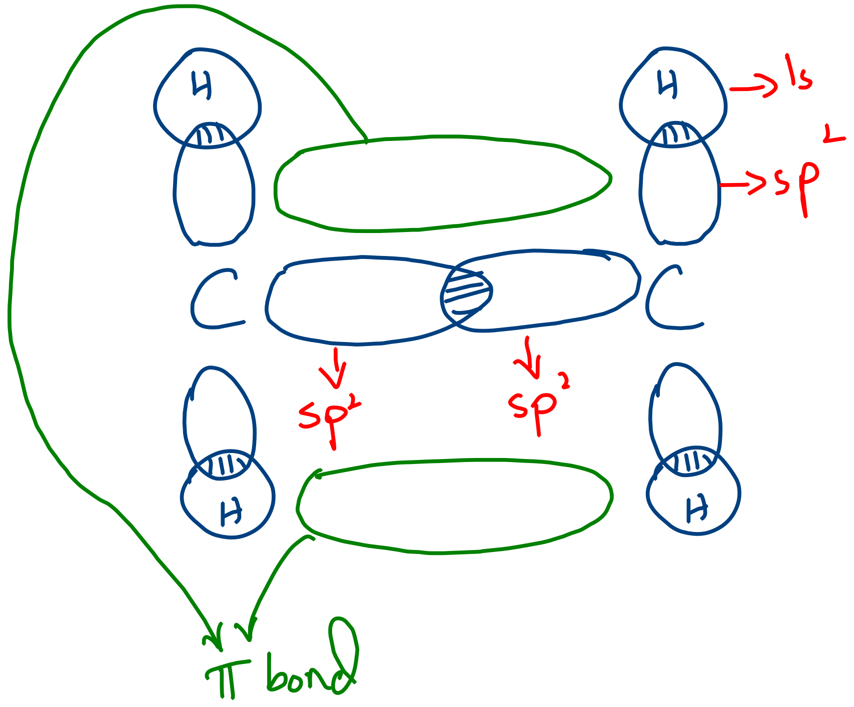
$\xrightarrow[\text{we need 1p orbital for sideways overlap}]{\text{we need 3 orbitals for end to end overlap}}$ 2s & 2x 2p combine

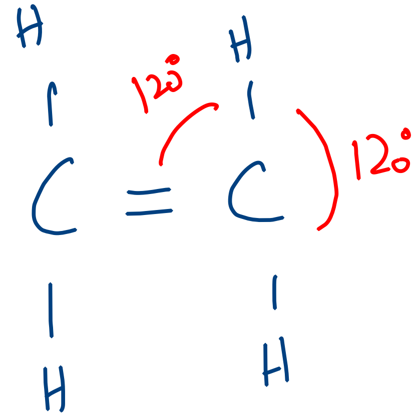
[↑][↑][↑] (1x 2s orbital, 2x 2p orbitals) ↳ 3 sp^2

[↑] 1x 2p orbital ⟶ form π bond



Ethene can now form :-





trigonal planar

- sp^2 hybridization is done when 3 single bonds and 1 π bond is to be made
- Usually in alKenes.

Sp hybridization :-

In ethyne (C_2H_2) $H \overset{\sigma}{-} C \overset{\pi}{\underset{\pi}{\equiv}}{}^{\sigma} C \overset{\sigma}{-} H$, the carbon atom needs to form 2 σ bonds and 2π bonds

↳ 2 end to end overlaps

↳ 2 sideways overlaps of 2p orbitals

C: [↑↓] [↑↓] [↑ | ↑ |]

$1s^2$ $2s^2$ $2p^2$

we need 2 orbitals for end to end overlap

we need 2p orbitals for sideways overlap

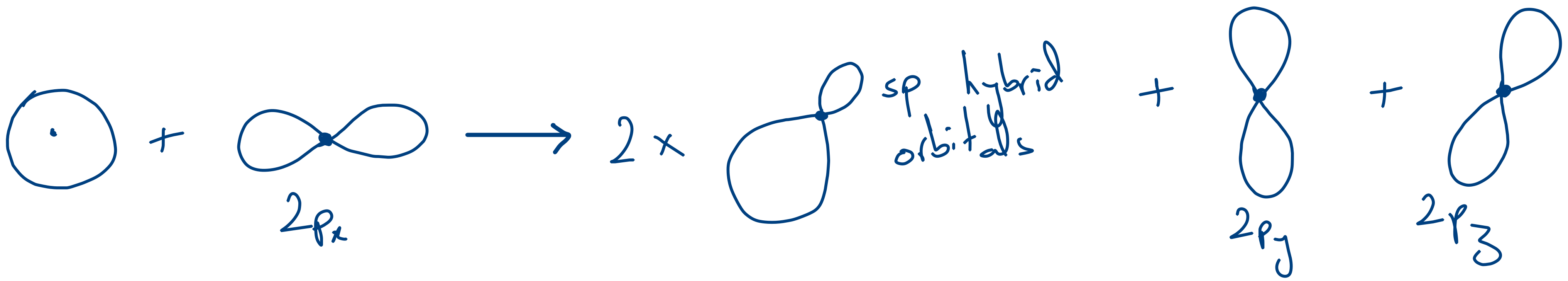
2s & 1x2p combine

[↑ | ↑] 2 sp hybridized orbitals + [↑ | ↑]

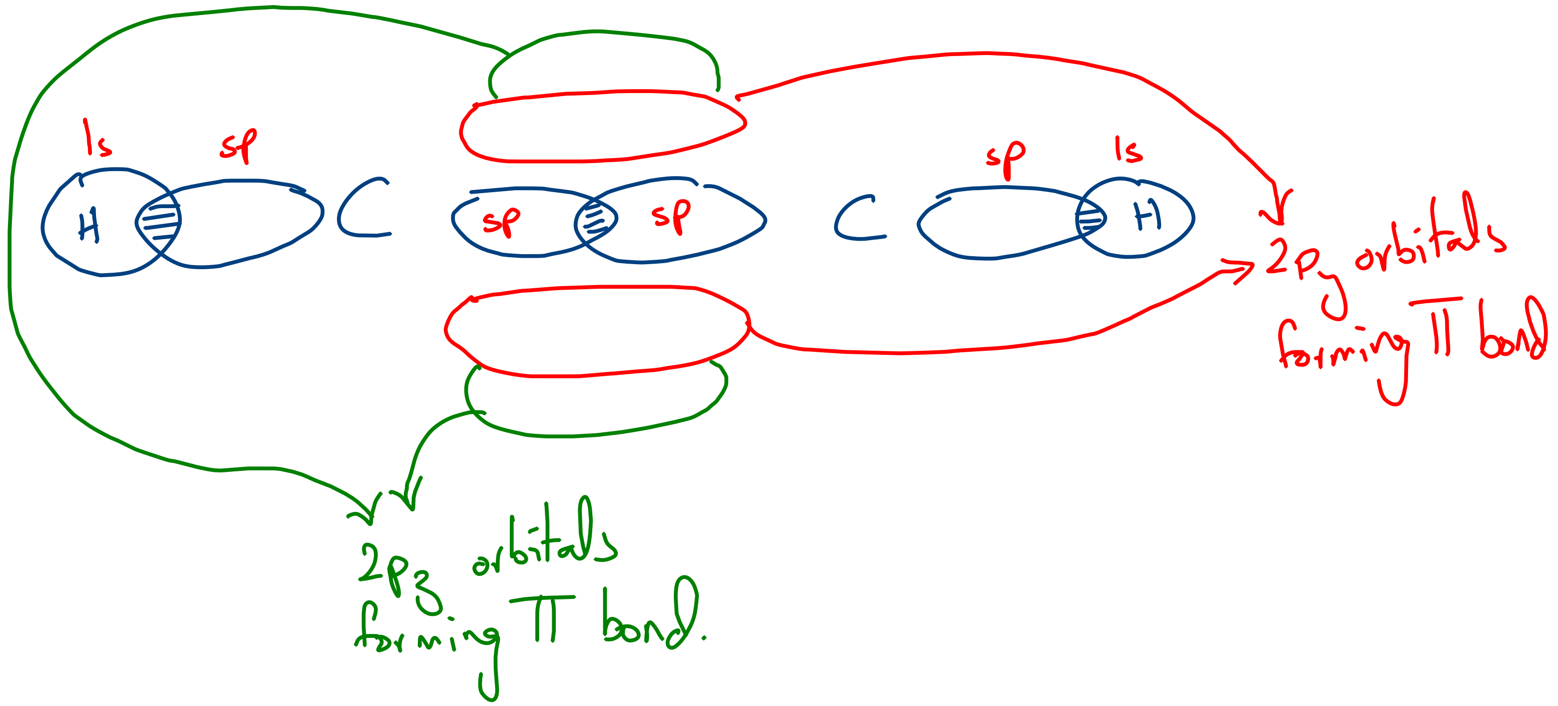
1x2s orbital, 1x2p orbital

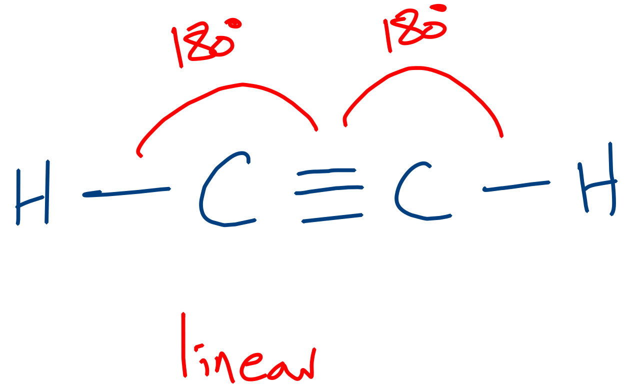
$2p_y$ $2p_z$

2x2p orbitals for 2π bonds



Ethyne can now form:-





- sp hybridization is done when 2 σ bonds and 2 π bonds are to be made
- Usually in alkynes and CO_2

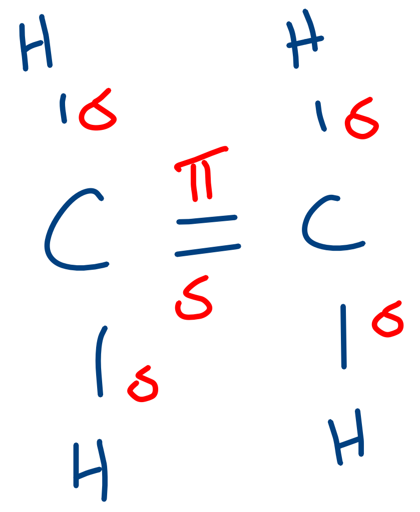
In CO_2, $O \overset{\pi}{\underset{\sigma}{=}} C \overset{\pi}{\underset{\sigma}{=}} O$, the carbon needs to form 2σ and 2π bonds → so sp hybridization

Hybridization in:-

1) Graphite / Graphene / Nanotubes: 1 C bonded to 3 C so 3σ bonds so sp^2 hybridized
2) Diamond: 1 C bonded to 4 C so 4σ bonds so sp^3 hybridized
3) Fullerenes / Buckminsterfullerene: 1 C bonded to 3 C so 3σ bonds so sp^2 hybridization

3 When considering one molecule of ethene, which row describes both the hybridisation of the atomic orbitals in the carbon atoms and the overall bonding?

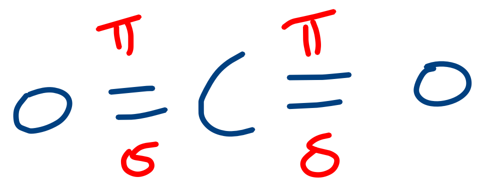
	hybridisation	bonding
A	sp^2	4 σ bonds 1 π bond
B	sp^2	5 σ bonds 1 π bond
C	sp^3	4 σ bonds 1 π bond
D	sp^3	5 σ bonds 1 π bond



C needs 3 σ and 1 π so sp^2

3 How many sigma (σ) and pi (π) bonds are present in the structure of CO_2?

	σ	π
A	0	4
B	2	2
C	3	1
D	4	0



Bond lengths and bond energies:-

Bond length:- The distance between the nuclei of 2 covalently bonded atoms

$\frac{1}{2} \times$ bond length = atomic radius (for homonuclear diatomic e.g H_2, F_2, Cl_2 etc)

Bond energy:- The minimum amount of energy required to break 1 mole of a bond in 1 mole of gaseous molecules.

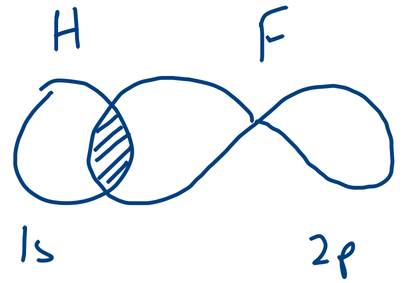
e.g. $H-H_{(g)} \longrightarrow 2H_{(g)}$

bond	bond length/ nm	bond energy/ kJ mol^{-1}
H-F	0.009	+562
H-Cl	0.127	+431
H-Br	0.141	+366
H-I	0.161	+299

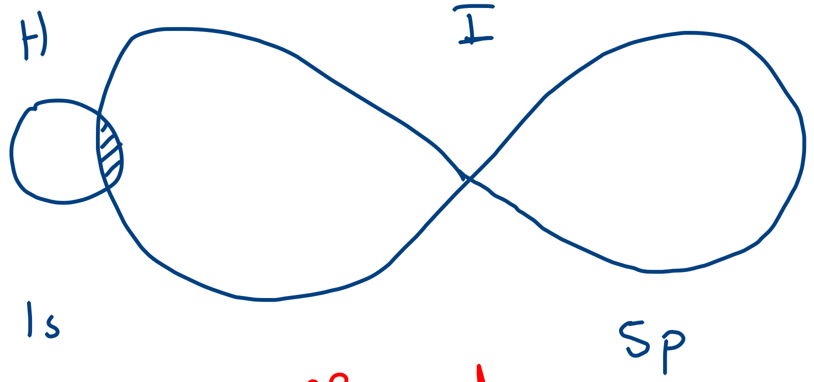
bond length increases

bond energy decreases (bond gets weaker)

The bond energy decreases down the group as the bond gets less polar. Also with bigger nuclei, the orbital overlap becomes less efficient.



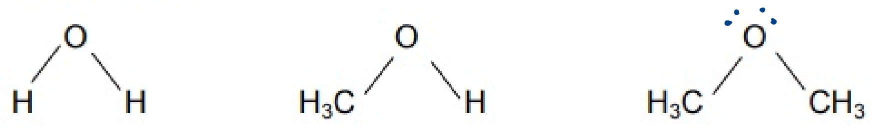
more efficient



less efficient

H-I to H-F: reactivity decreases as it becomes difficult to break existing bond and make new ones

1 The structural formulae of water, methanol and methoxymethane, CH_3OCH_3, are given below.



a i How many lone pairs of electrons are there around the oxygen atom in methoxymethane?

.........2.........

ii Suggest the size of the C-O-C bond angle in methoxymethane.

......104.5°.........

[2]

The physical properties of a covalent compound, such as its melting point, boiling point or solubility, are related to the strength of the attractive forces between the molecules of that compound.

These relatively weak attractive forces are called intermolecular forces. They differ in their strength and include the following.

A interactions involving permanent dipoles

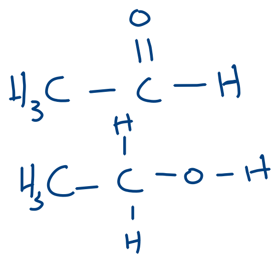
B interactions involving temporary or induced dipoles

C hydrogen bonding

b By using the letters **A**, **B** or **C**, state the **strongest** intermolecular force in **each** of the following compounds.

For each compound, write the answer on the dotted line.

ethanal	CH_3CHO	......A......
ethanol	CH_3CH_2OH	......C......
methoxymethane	CH_3OCH_3	......A......
2-methylpropane	$(CH_3)_2CHCH_3$	......B......



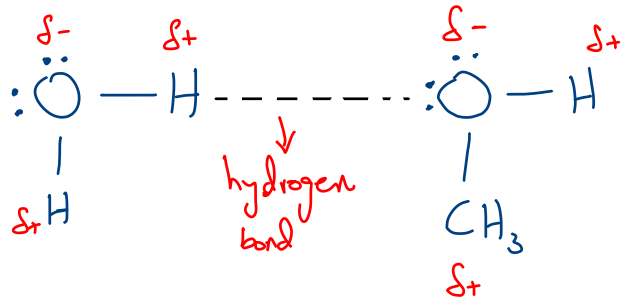
[4]

c Methanol and water are completely soluble in each other.

i Which intermolecular force exists between methanol molecules and water molecules that makes these two liquids soluble in each other?

.........hydrogen bonding.........

ii Draw a diagram that clearly shows this intermolecular force. Your diagram should show any lone pairs or dipoles present on either molecule that you consider important.



[4]

2 a i State what is meant by the term *polar* when applied to a covalent bond.

One nucleus in the molecule pulls on the shared pair of electrons more towards itself.

ii Consider the covalent bonds in molecules of hydrogen and water. State whether the covalent bonds are polar or non-polar. Explain your answers.

Bonds in hydrogen non polar

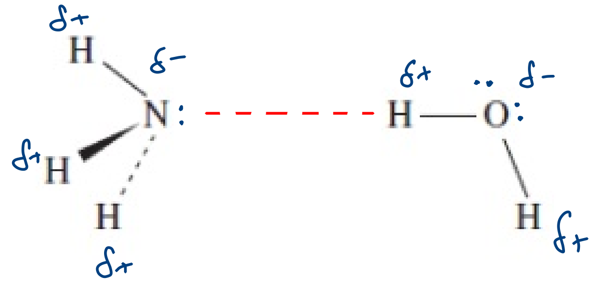
Bonds in water Polar

Explanation Oxygen has a much higher electronegativity than hydrogen. In H_2, the 2 hydrogens pull the shared pair with the same force

[4]

b Ammonia is very soluble in water because it is able to form hydrogen bonds with water molecules.

i Complete the diagram below to show how an ammonia molecule forms a hydrogen bond with a water molecule. Include partial charges and all the lone pairs of electrons.



ii The bond angle in a molecule of water is about 104.5°. State the bond angle in an ammonia molecule and explain why it is different from that in water.

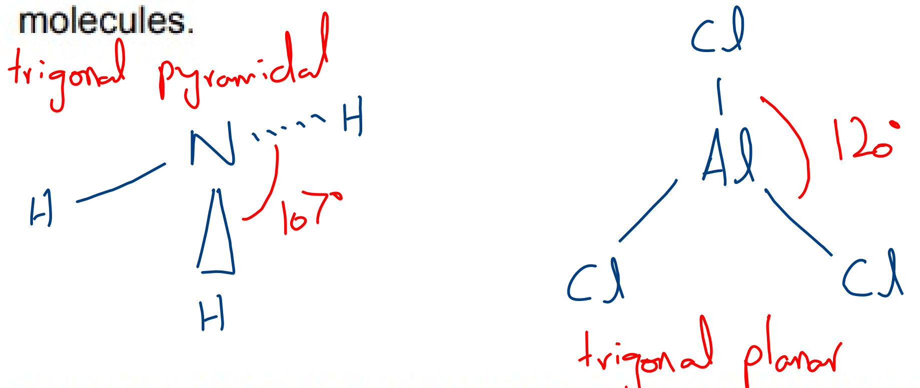
Bond angle in ammonia 107°

Explanation Nitrogen in ammonia has only 1 lone pair while oxygen in water has 2 lone pairs.

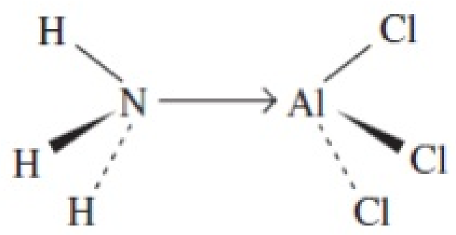
[6]

c Ammonia reacts with aluminium chloride.

i Draw diagrams to show the 3-dimensional shape of the NH_3 and $AlCl_3$ molecules.



ii The molecule formed during the reaction is shown below.



Name the type of bond formed between the nitrogen and aluminium atoms. Explain how this bond is formed.

Type of bond Dative bond

Explanation Aluminium is electron deficient so nitrogen uses its lone pair and completely shares it with aluminium.

[4]