

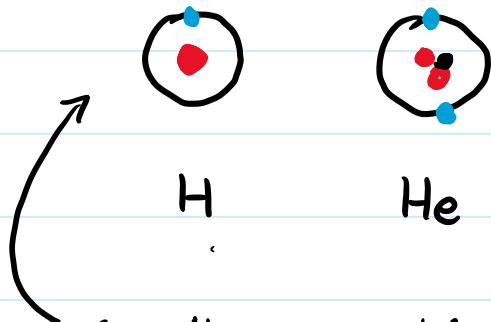
Class Notes

Thursday, August 24, 2017 10:55 PM

This is the original notes makes on class

Relative atomic mass

- different atoms have different masses

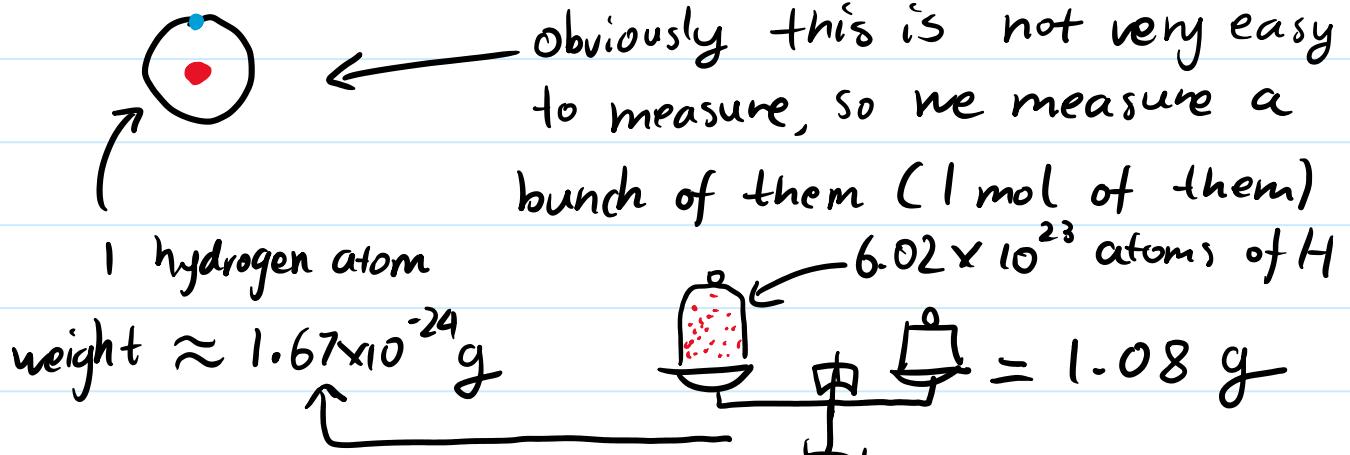


the atom mass mainly comes from protons and neutrons.

Different atom have different configuration of them

so they would have different mass

- one atoms are so small that we can not measure how much is the weight exactly, so we decide to measure a bunch of them



but the problem is that the atom is still writing in a long chain of number, which is not very good.

The deeper reason is laying on what we can do to measure the weight, it's much more easier to just compare one weight with another (discuss later)

- we compare a number of given atom's mass with

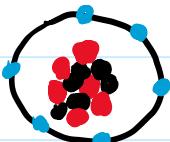
the same number of a standarized atom, and get a relative atomic mass (Ar)

For measurement accuracy we choose C-12,

which has 6 protons and 6 neutrons

actually one unit mass of Ar is

$\frac{1}{12}$ of C¹², which give 1 (unit)



a C-12 atom

↑

exactly 12 unit

as 1 proton and neutron all have 1 unit

so for H it is just unit 1 as it only have 1 proton

So we would obtain the expression for Ar:

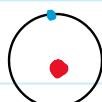
$$\text{Ar [Element Y]} = \frac{\text{average mass of one atom of Y} \times 12}{\text{mass of one atom of C-12}}$$

But you would notice that for most of the case, the number is not a integer

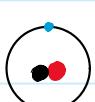
e.g. Ar (H) = 1.0079, which wasn't the same compare with the ${}^1\text{H}$'s proton mass $\rightarrow 1$

- it turns out that the average mass of the atom of a particular element is the actual definition for it.

H has three isotopes



${}^1\text{H}$



${}^2\text{H}$



${}^3\text{H}$

and all of them should takes some part of the Ar depends on the abundance

$$\text{Ar[H]} = {}^1\text{H} \times ({}^1\text{H} \text{ abundance}) + {}^2\text{H} \times ({}^2\text{H} \text{ abundance}) + {}^3\text{H} \times ({}^3\text{H} \text{ abundance}), \text{ and give us} \rightarrow 1.0079$$

- generally speaking the small fraction can be ignore as $1.0079 \rightarrow 1.0$, but in some case, like Cl, it can not be ignore, giving 35.5

$$\text{Ar[Cl]} = 35.5 \text{ as there is } 75\% \text{ of } {}^{35}\text{Cl} \text{ and } 25\% \text{ of } {}^{37}\text{Cl}$$

* How do we actually measure the mass of atoms?

- by using mass spectrometry:

Atomic Structure

Thursday, November 30, 2017 8:07 AM

How can an orbital hold two electrons with negative charges as they tend to repel each other?

- spin (自旋)

$\uparrow \downarrow$ a circulated charged particles will have a property (magnetic) which call spin (canceling the electrical repulsion)

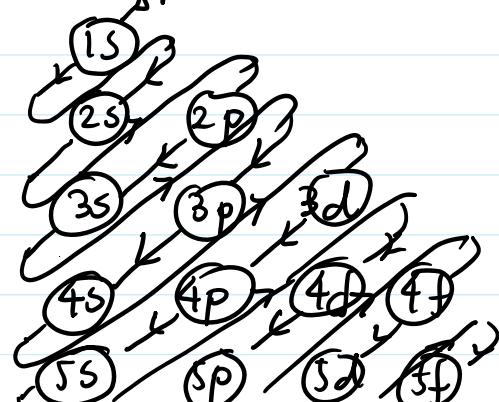
$\uparrow \downarrow$ the two electrons will have spin in opposite direction
 (opposite Spin pair)
Pauli inclusive principle

$1\downarrow$

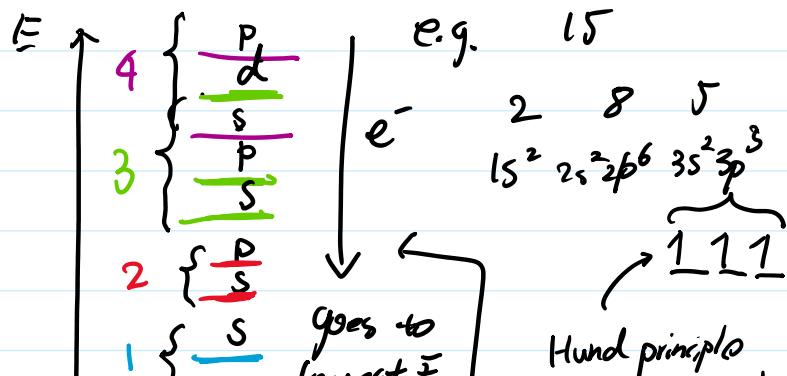
show in chemistry

Complete electronic configuration:

- Three main rules which determine the way in which electrons fill the orbitals.
- 1. Aufbau / building principle: electrons always fill the lowest energy orbitals first



Aufbau rules



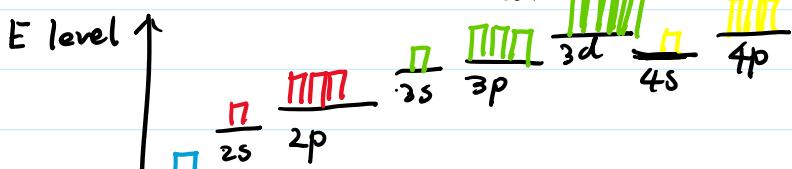
e.g.

$_{19} \text{K}$

$\frac{1s^2}{2} \quad \frac{2s^2 2p^6}{8} \quad \frac{3s^2 3p^6}{8} \quad \frac{4s^1}{1} \quad \checkmark$

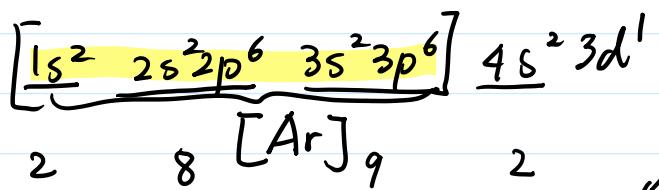
$_{21} \text{Sc}$ (d-block element)

-



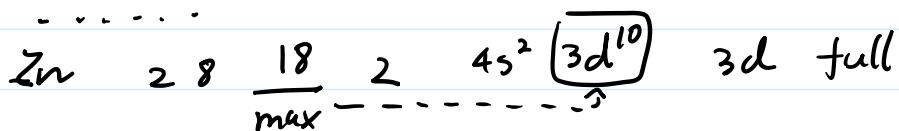
however the things
are not that simple,

21 Sc (d-block element)



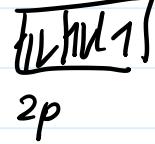
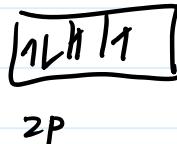
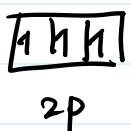
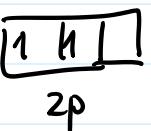
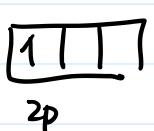
now... are not that simple, there is something called penetration effect

will see that later



2. Hund's rule: electrons never pair up in the same orbital until all orbitals of the same energy are singly occupied

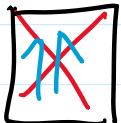
the order of electrons filling in:



→ E level

the reason is that singly occupied electron is slightly lower energy than the paired one (because of the repulsion of each other)
different orbitals has little effect than this (but they still have)

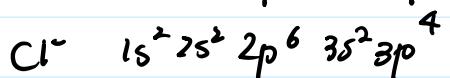
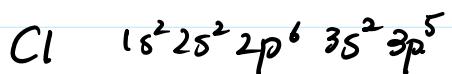
3. Pauli exclusion principle: only two electrons may occupy the same orbital, and they must do so with opposite spin.



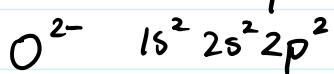
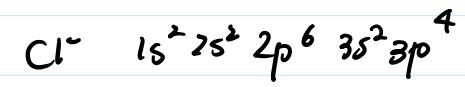
the reason is that the repulsion between two same spin electrons are very high so that it's unstable.

If CIE want's you to draw arrows, they will ask for arrowbox notation.

Electronic structures in ions



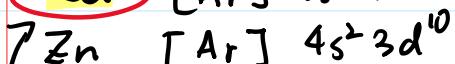
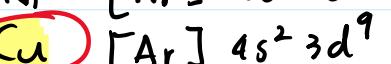
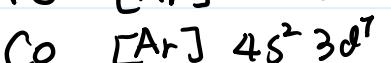
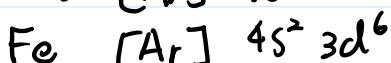
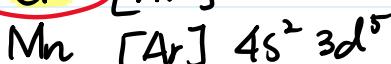
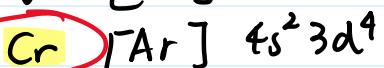
the electron remove from the outermost shell (not the



The electron remove from the outermost shell (not the highest energy level first)

Electronic configuration of transition elements

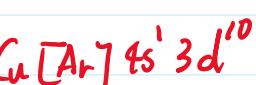
SC $1s^2 2s^2 2p^6 3s^2 3p^6 4s^2 3d^1$ \rightarrow noble gas notation (replace part of the structure) \hookrightarrow find noble gas above the period and replace some



There are chemical evidence showing that the Cr has 5

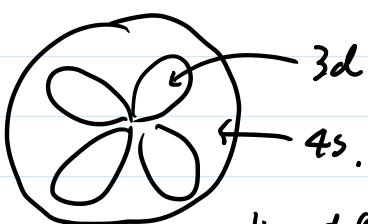
e^- in d. and Cu has 10

e^- in d. \rightarrow these two aren't matching matching the principle

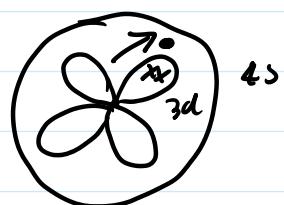


The reason is that the original configuration in d orbit: $3d^5$, is almost full, so the electron will fall into the $3d^6$ to make both $4s$ and $3d$ singly occupied to stable.

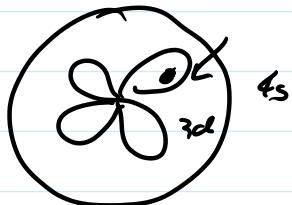
In Cu it's similar (the energy level as $3d$ closer to than $4s$ is smaller, so much stable)



the difference is minor in $3d$ and $4s$

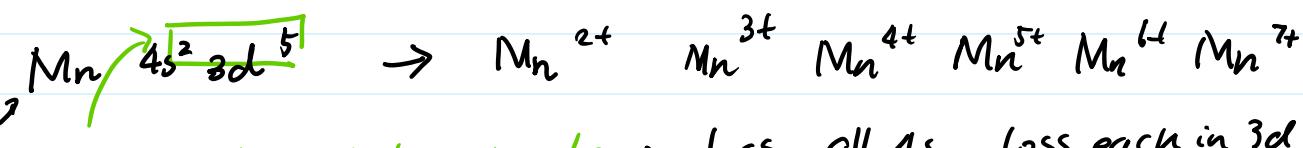


if already occupied repulsion $\rightarrow 4s$



if no electron occupied no repulsion $\leftarrow 3d$
(ions lose at $4s$ first) (half full principle)

$\checkmark \quad 4s^2 3d^3 \rightarrow$ how can we predict the oxidation state



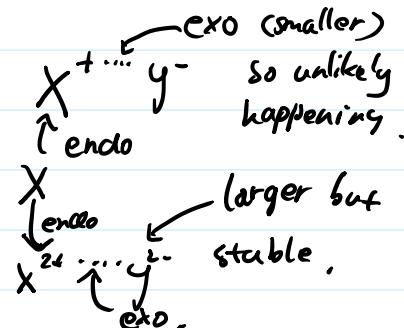
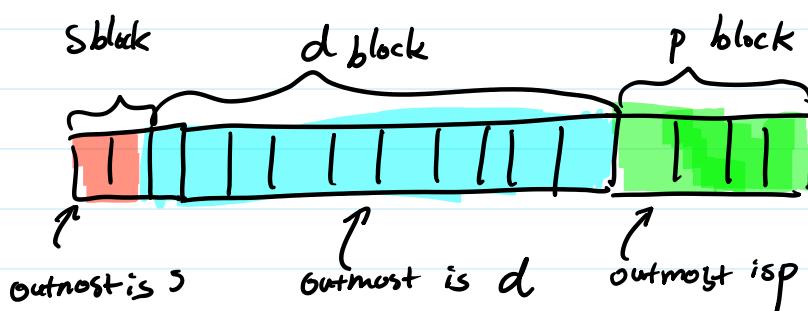


Count the top electron number: loss all 4s, loss each in 3d.

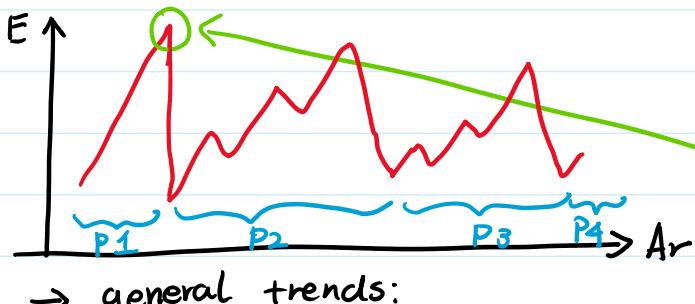
because 4s and 3d are so close that the difference is so small.

Mg is difference: the Mg: $2p^6 \underbrace{3s^2}_{\text{large gap}} \rightarrow \text{loss 2}$. then big jump.

Tl is ignored because the new ions formed is so weakly bond



Sc and Zn are not transition metal: form a positive ions and still have a incomplete subshells



ionisation energy of the first 20 element

→ general trends:

increase of ionisation energy in same period

- nuclear charge increase
- shielding effect remain the same
- same distance to the nucleus

→ a decrease at G2-3

- in P subshell rather than S subshell so distance slightly increase
- shielding effect of the inner shell increase
- outweigh the increase in nucleus charge

→ a decrease at G5-6

- Spain - pair repulsion

→ repulsion of pair of electron within a P orbital

- shielding the same

- distance the same.

however, the ionization energy will not get even lower here even it has two Spain pair → only removing one

distance closest to nucleus
He has the highest first ionisation energy because in P1 there is no inner electrons

remove from higher energy level
p has a higher energy than S so easier to remove



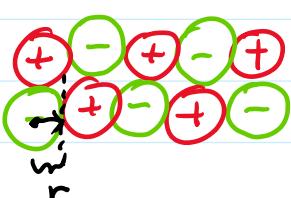
write electronic configuration first



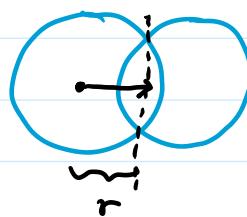
one of this

Atomic and Ionic Radius

ionic radius



covalent radius

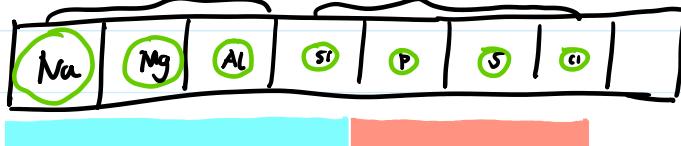


metallic radius Van der waals



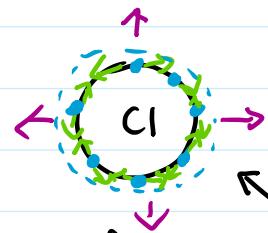
Atomic radius across the period → decrease.

nuclear charge ↑ outer electron closer
metallic radius covalent radius

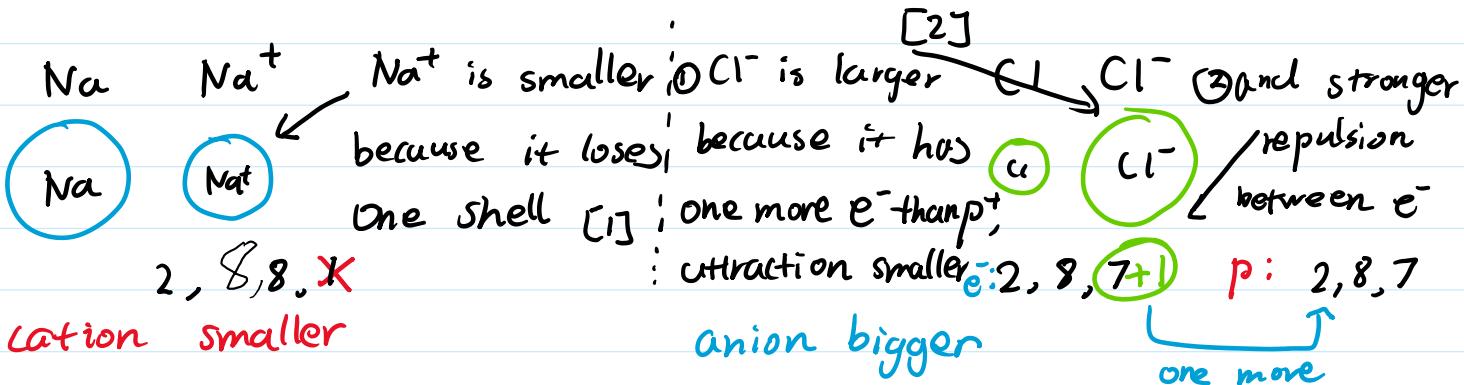


significant decrease; not significant decrease

the electrons are repulsion a bit further from the nucleus
(expansion)



Ionic radius



Compare the size of ions:



2,8 2,8 2,8

more polarizing



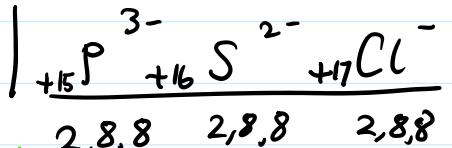
slightly decrease

as the nuclear charge increase

put Na^+ in water
it stays in it
pH 6.5~6.7

put Al^{3+} in water,
it won't move free
pH around 3

charge high and small



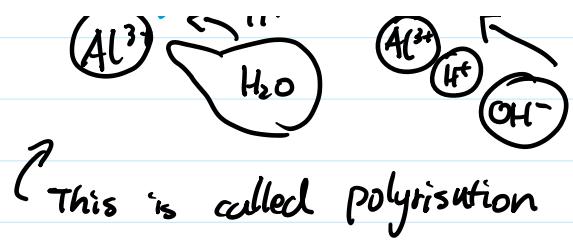
2,8,8 2,8,8 2,8,8

3 more 2 more 1 more



slightly decrease

as the electrons more than protons is lesser
(and nuclear charge increasing)?



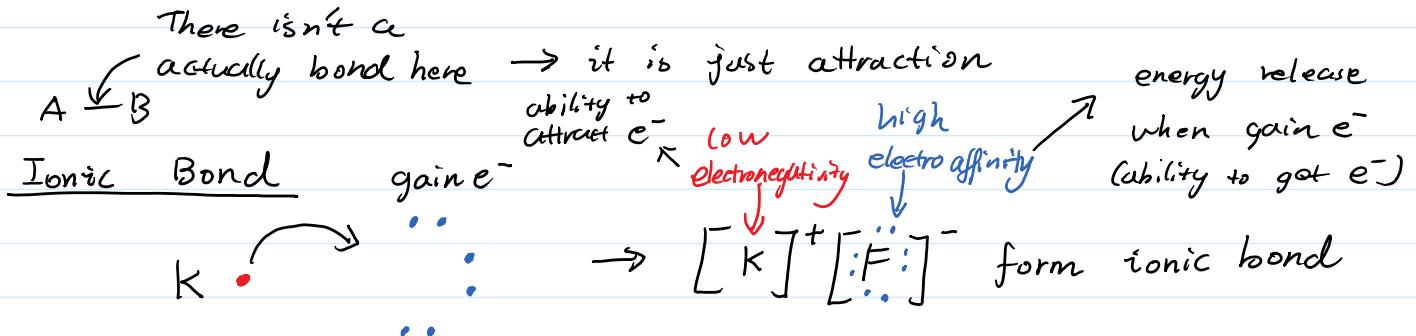
Chemical Bonding

Thursday, November 30, 2017 8:06 AM

bond formation \rightarrow exothermic \rightarrow get chemical stability

- so atoms tends to form bond together rather than to be free

chemical bonds are electrostatic attraction between particles in a molecule



But why is K giving e^- to F, not vice versa?

because of electronegativity \rightarrow the ability to attract e^- from others.

or low first ionisation energy in K and high first ionisation energy in F

electronegativity: ability to attract a pair of e^- in a bond.

Δ electronegativity Low
covalent bond

Δ electronegativity high
ionic bond

electropositivity: metal are electropositive, non-metal are electronegative
lose e^- attract e^-

ionic bond formation: metal lose e^- , non-metal gain e^- (98% situation)
exception $\text{NH}_4^+ \text{Cl}^-$

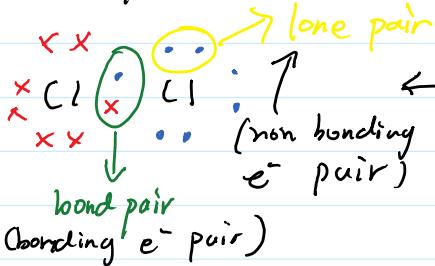
ionic bonding is a strong electrostatic attraction between oppositely charged ions

Covalent Bond

atoms still want to achieve a noble gas configuration (octet rule)

they don't lose or gain e^- , they share e^-

the shared pair e^- are called bond pair / shared pair



Notice for the element below period 2, you can have more than 4 bond!

e.g. PCl_5

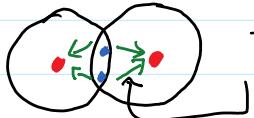
$\text{P}^{+5} 10 e^-$ expansion of octet

C only made octet as 2s and 2p but p made expansion beyond $8e^-$ as 3s, 3p and 3d

This is not used, but is possible

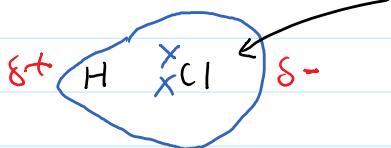
The element beyond period 3 make this possible as they have d/l what hold the covalent bond?

It's also electrostatic force (bond pair attract by the both side nucleus)

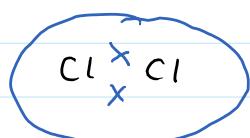


This electrostatic attraction holds the covalent bond together

higher electronegativity

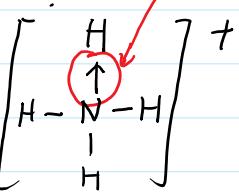
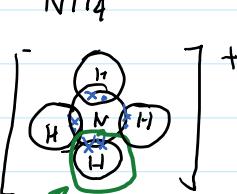
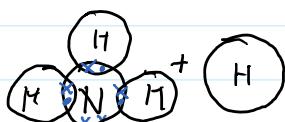
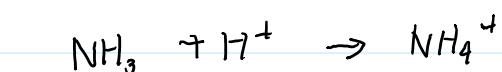


shared pair not equal



shared pair equal!

Dative Covalent Bond / coordinate Bond

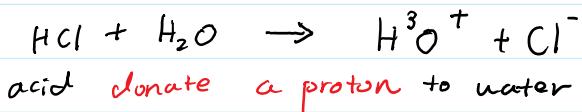
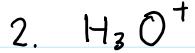


The arrow shows from O^- donor to acceptor

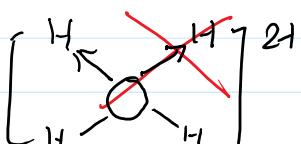
This is the dative covalent bond

Both e^- in a covalent bond are donated by only one of the atoms the bond formed is called Dative covalent bond

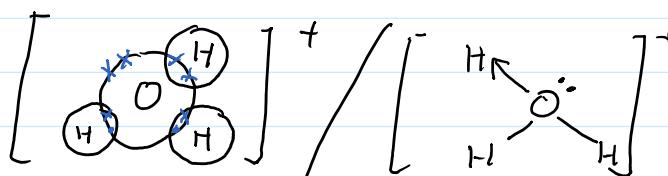
(No bond energy different)



but there is no H_4O^+



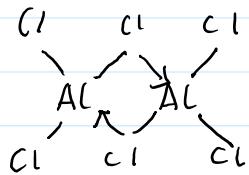
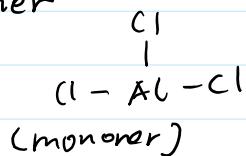
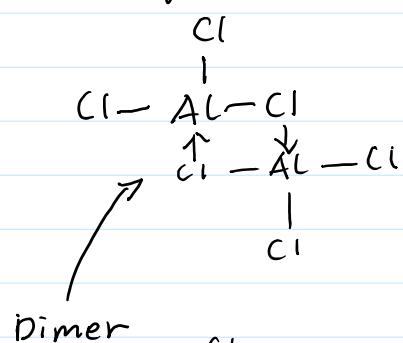
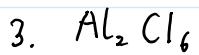
This is not allow!



The positive charge repulsion (high activation energy)

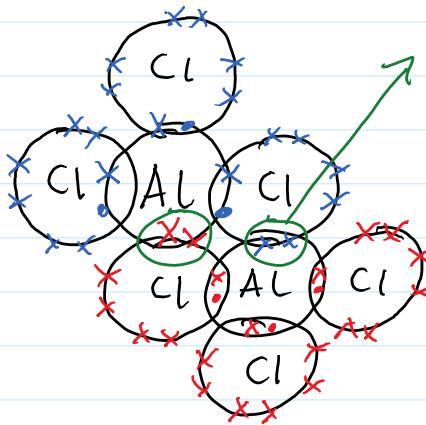
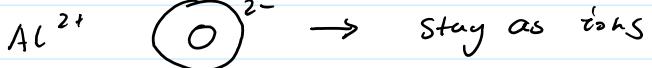
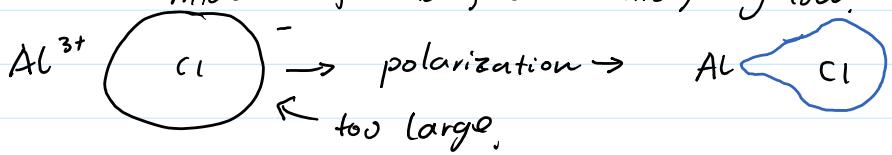


$AlCl_6$ is covalent bond (simple molecule)



AlCl_3 is covalent bond (simple molecule)
not an ionic bond \leftarrow the b.p. is low

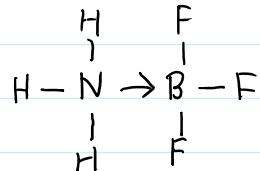
Al at middle of table, so Δ electronegativity low.



AL is not full (electron deficiency)
It's going to find someone else
and drag e^- from.

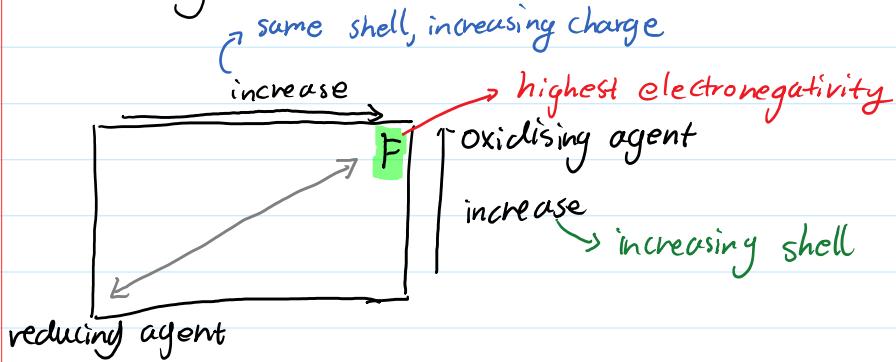
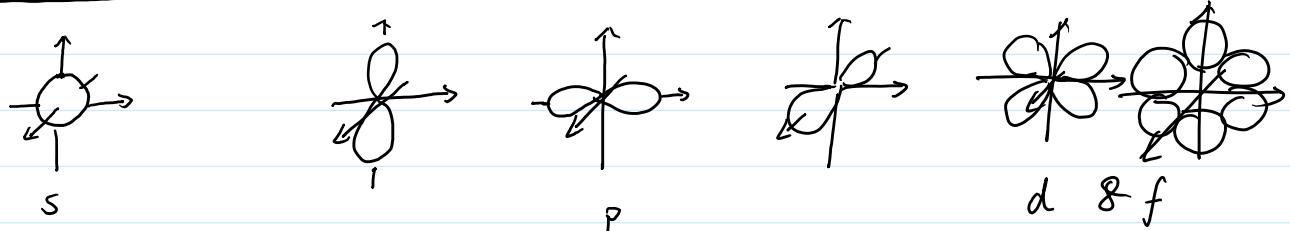
In pure sample, it
just drag from another e^-

BF_3 and NH_3



Electronegativity

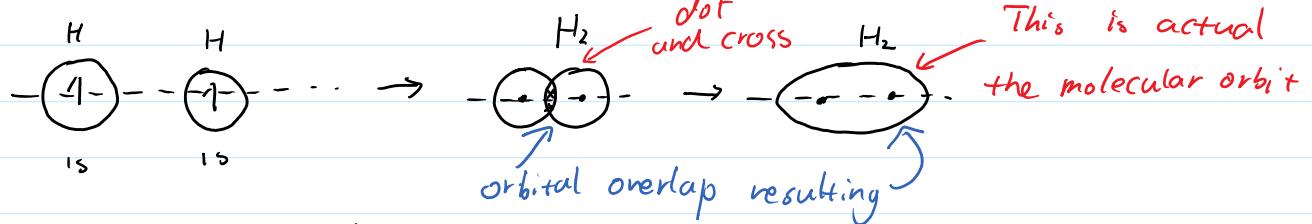
Electronegativity of an atom is the measure of the ability of an atom attracting electrons

Atomic Orbitals

the probability of electrons appear on certain area

Molecular orbitals

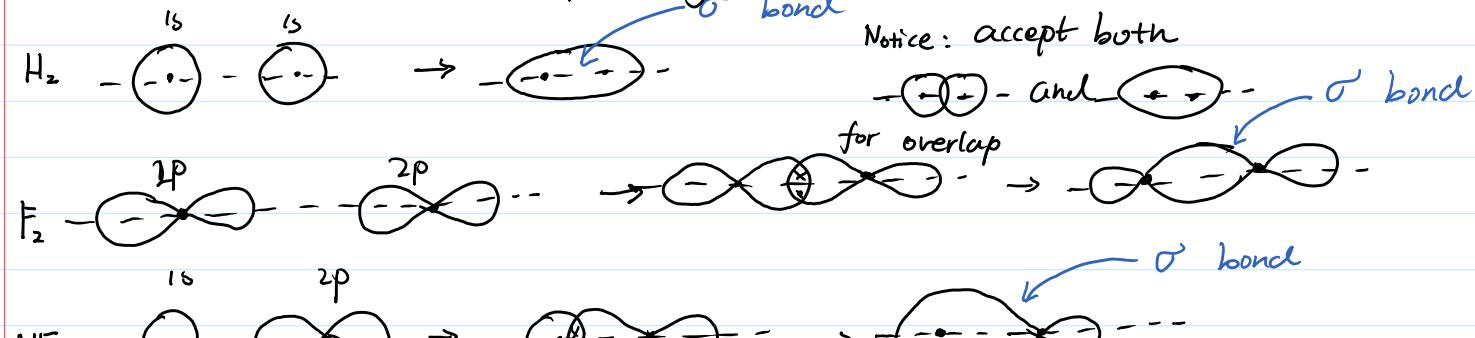
molecular orbitals form when atomic orbitals overlap resulting in bonds between atoms

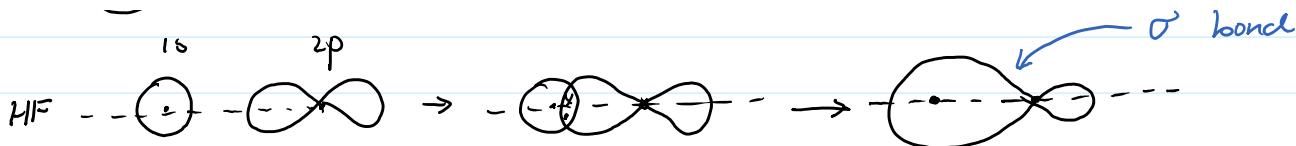


σ bond and π bond

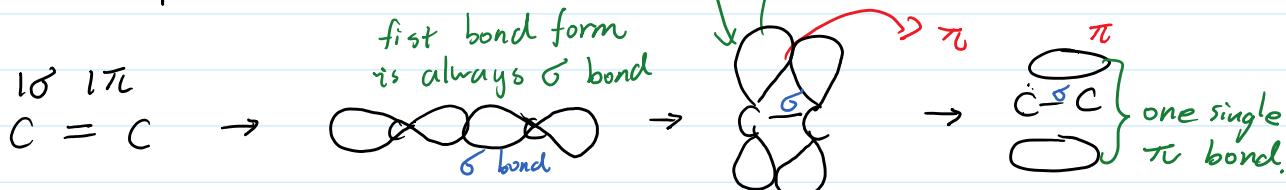
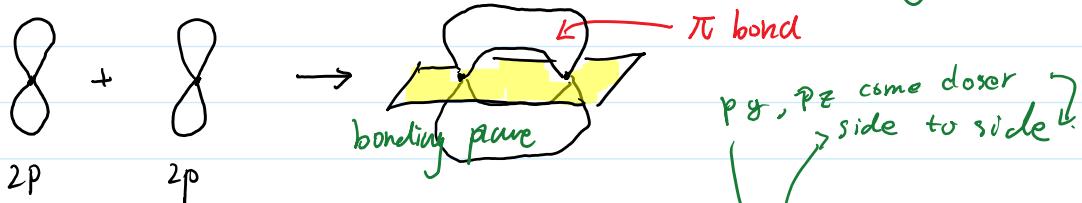
These bonds are results from overlap of atomic orbitals

π bond: 'head-to-head' overlap along the bond axis

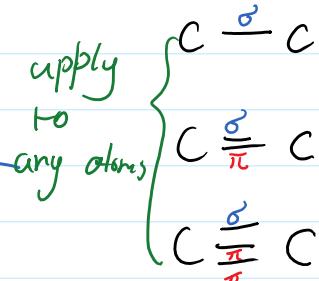
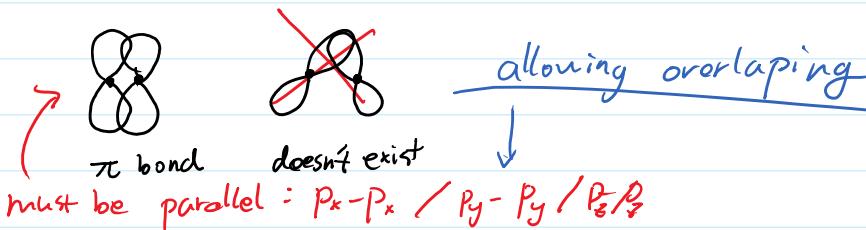
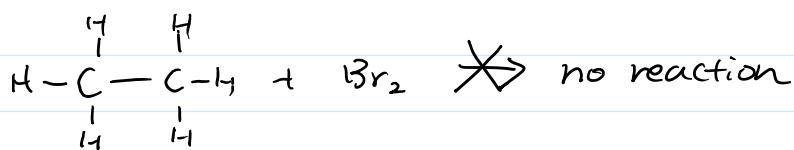
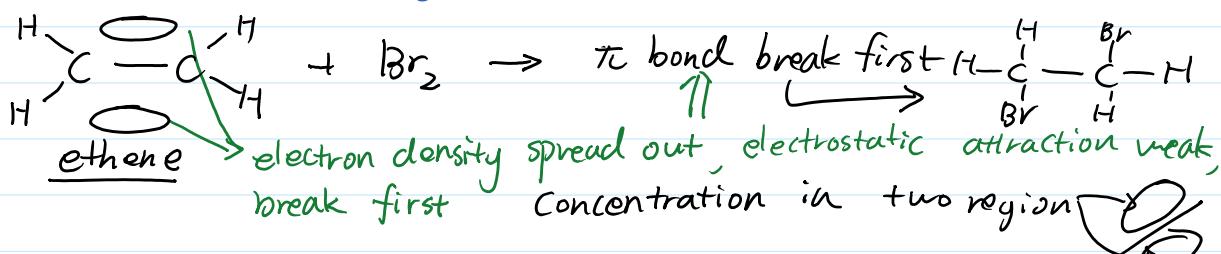
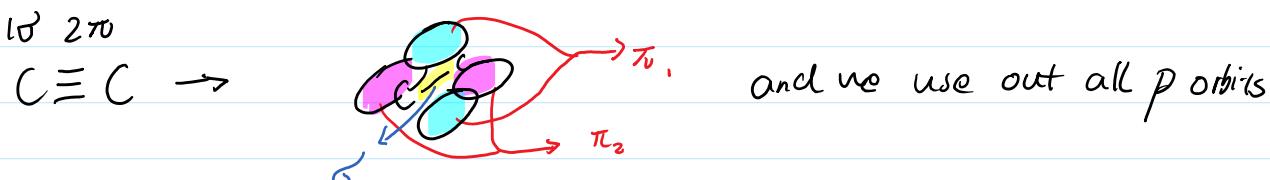




π bond: side to side/lateral overlap of **p orbitals** only possible for p orbitals!



bromine water test for unsaturated carbon chain \rightarrow decolorized



Shapes Of Molecules

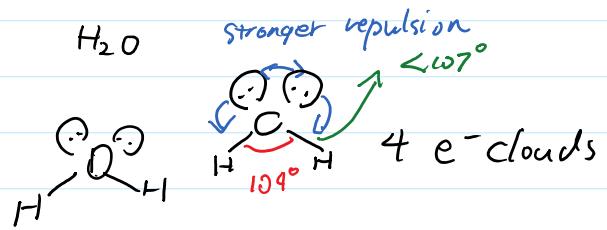
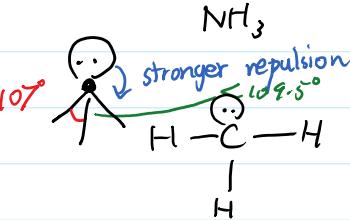
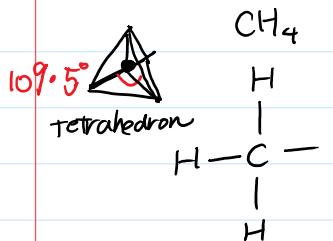
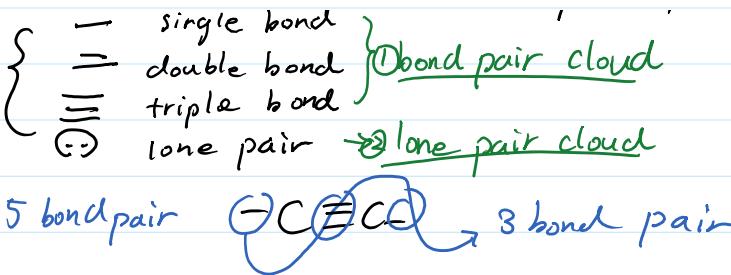
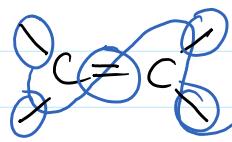
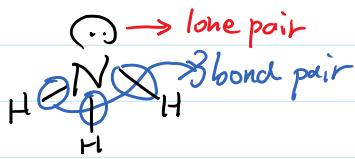
It's a description of the way the atoms in the molecule occupy space

Electron cloud Repulsion Theory \rightarrow VSEPR (valence shell electron pair repulsion)

negative charge which repel each other

an electron cloud could be { $\begin{cases} \text{--- single bond} \\ \text{= double bond} \\ \text{= triple bond} \end{cases}$ } ① bond pair cloud

(negative charge which repel each other)
an electron cloud could be



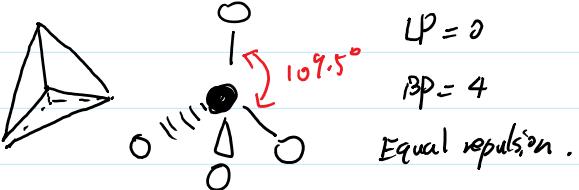
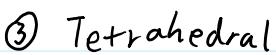
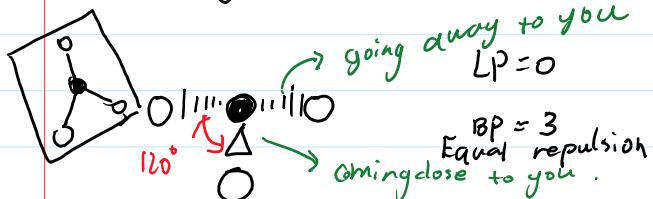
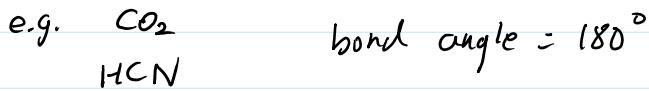
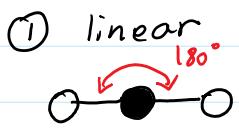
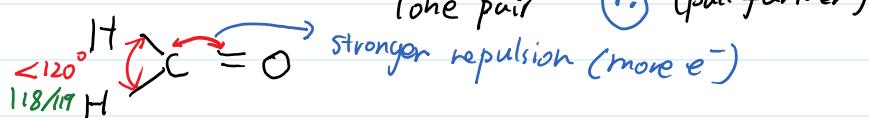
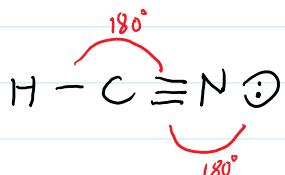
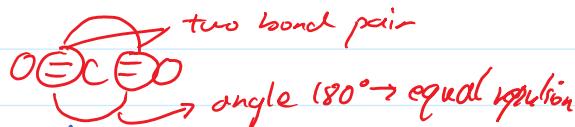
The point is that to minimize repulsion between electron clouds to stabilize them

↑ The atoms stay as far as possible

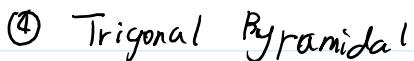
lone pairs of electrons exert a greater repelling effect than bond pairs do

lone pair - lone pair > lone pair - bond pair > bond pair - bond pair

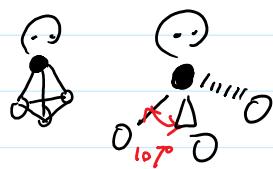
lone pair closer to nucleus → repel each other more



3 atoms in same plane.



④ Trigonal Pyramidal

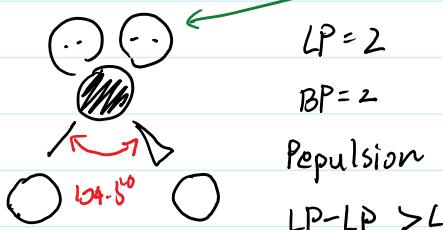


$$\begin{array}{l} LP = 1 \\ BP = 3 \\ LP - BP \geq BP - BP \end{array}$$

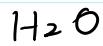
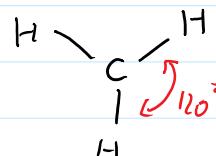
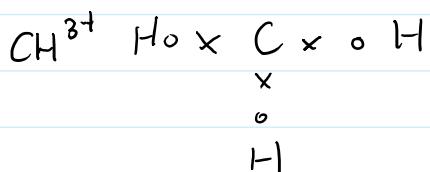
e.g. NH_3
 PCl_3

bond angle = 107°

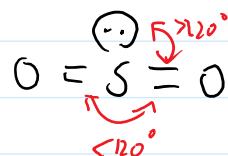
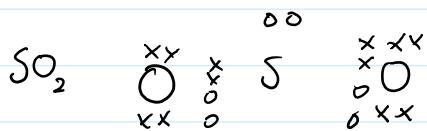
⑤ Bent [non-linear] lone pair



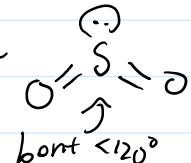
e.g.

Bond Angle = 104.5° 

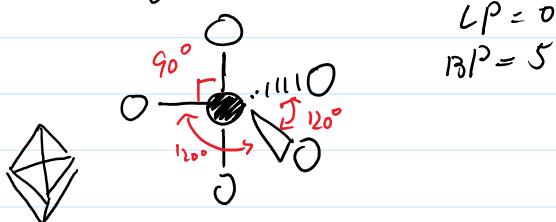
trigonal planar



trigonal planar

bent \rightarrow V shape

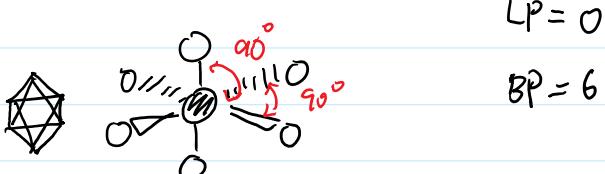
⑥ Trigonal bipyramidal



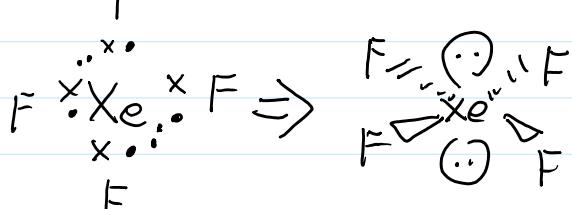
e.g.

bond angle: $120^\circ / 90^\circ$

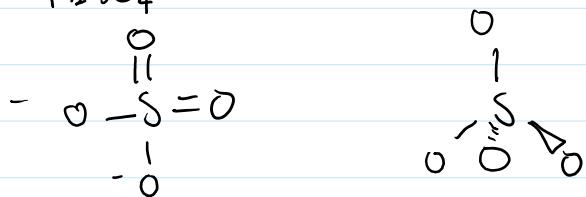
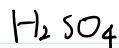
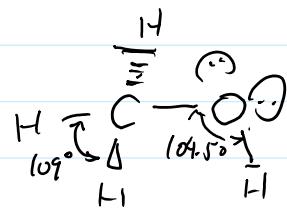
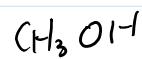
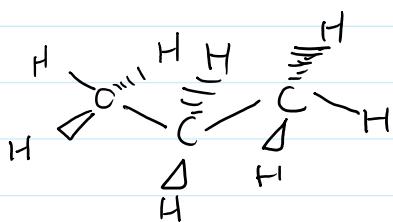
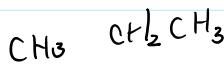
⑦ Octahedral



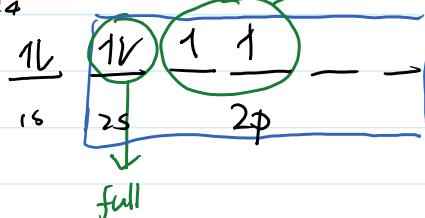
e.g.

bond angle: 90° 

⑧ Square planar



Hybridisation



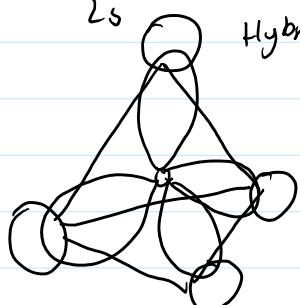
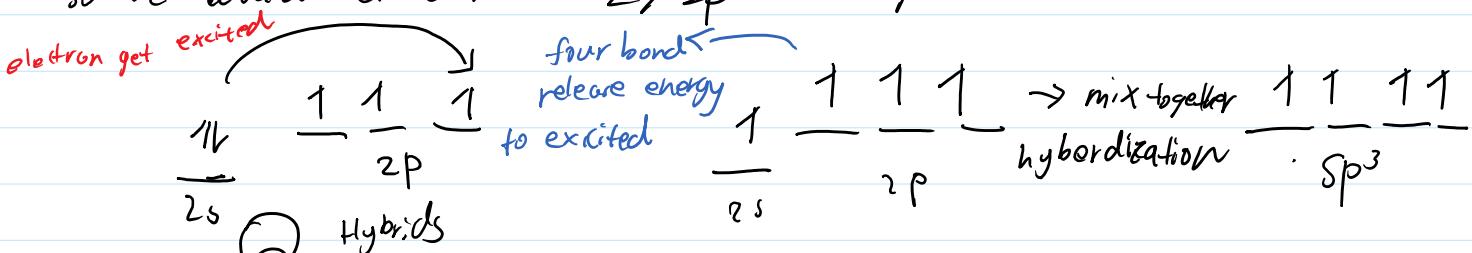
2 orbits that can form bond.

CH_4 where does this come from?

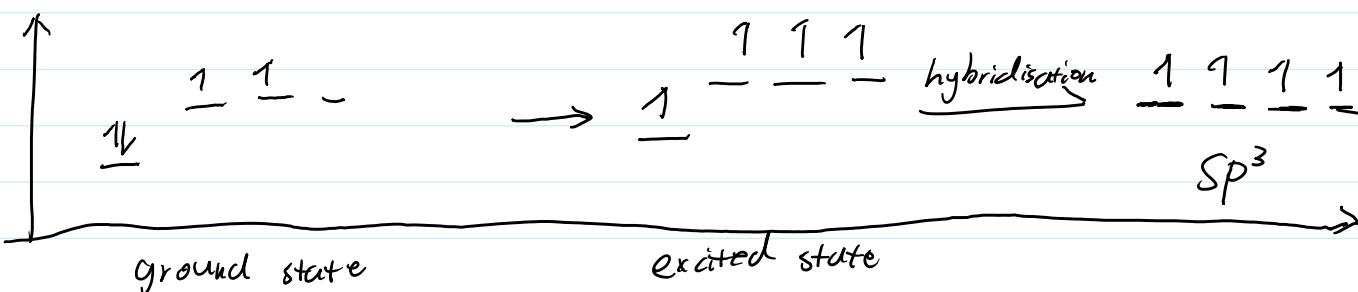
and



so we would consider the 2s/2p mix together



a hybrid orbital results from the mixing of different types of atomic orbitals on the same atom when it forms a covalent compound. The new set of orbitals of same energy are called 'hybrid orbitals'



ground state

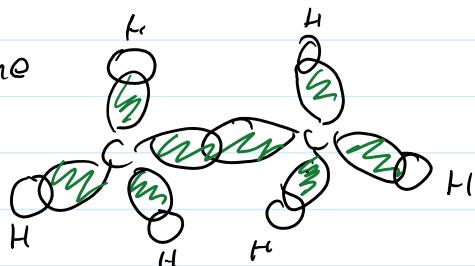
excited state

Sp³ hybrid orbitals:



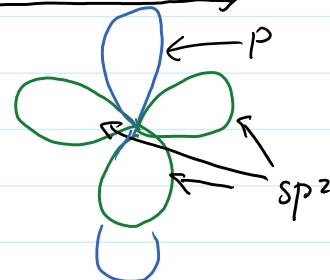
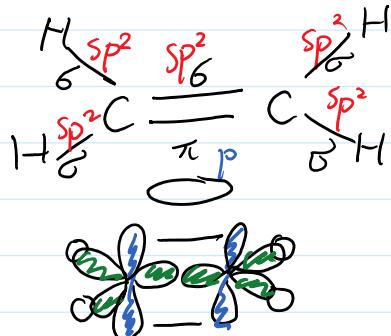
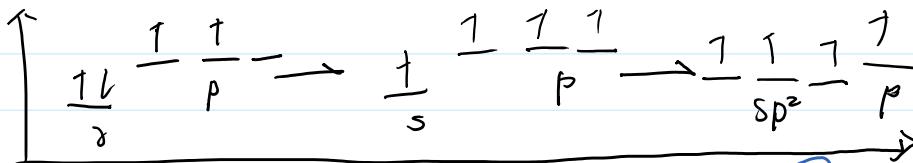
alkane \rightarrow four bond \rightarrow $sp^3(4\sigma)$ (—)

e.g. ethane

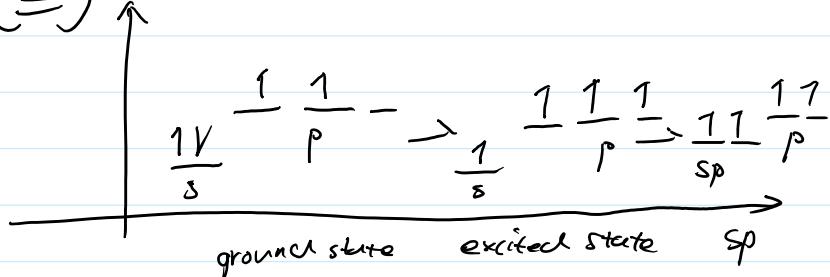
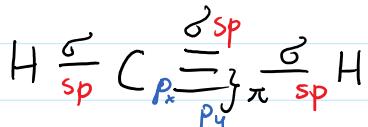


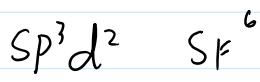
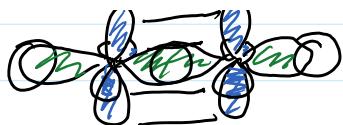
SP_2 hybridisation (3σ) (=)

Ethene



sp hybridisation $(2s)(\equiv)$ ↑



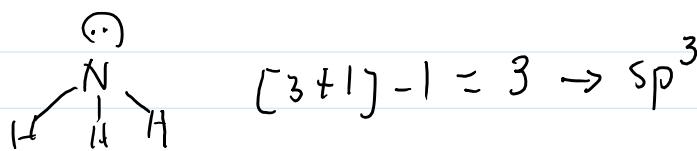
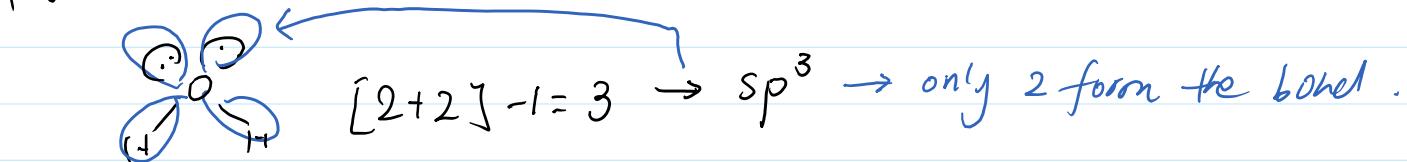
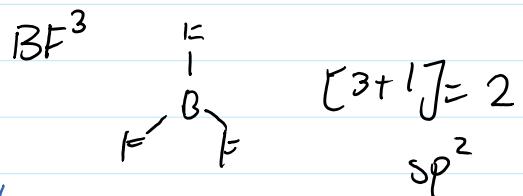
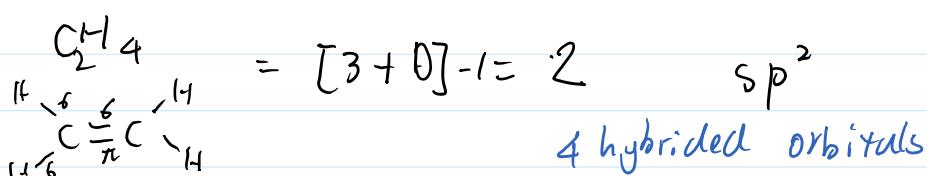


Predicting hybridisation

$$x \text{ in } sp^x = [\text{total } 6 \text{ bonds} + Lp] \quad |$$

$$CH_4 = [4 + 0] - 1 = 3 \quad sp^3$$

we are only interested in carbon atom



Total number of e ⁻ pair	shape			
2	Linear	CO_2, HCl	$O \begin{smallmatrix} \pi \\ 6 \end{smallmatrix} C = \begin{smallmatrix} \pi \\ 6 \end{smallmatrix} O$	sp
3	Trigonal planar 3 BP O 120°	BF_3	$\begin{smallmatrix} 120^\circ \\ \sigma/B \\ 6 \end{smallmatrix}$	sp^2
	Bent 2 BP 1 LP	SO_2	$O \begin{smallmatrix} \pi \\ 120^\circ \\ 180^\circ \end{smallmatrix} S = O$	sp^2
4	Tetrahedral 4 LP	CH_4	-	sp^3
	Pyramidal 1 LP	NH_3	-	sp^3

4

Tetrahedral

 $1s1p$ NH_3 sp^3

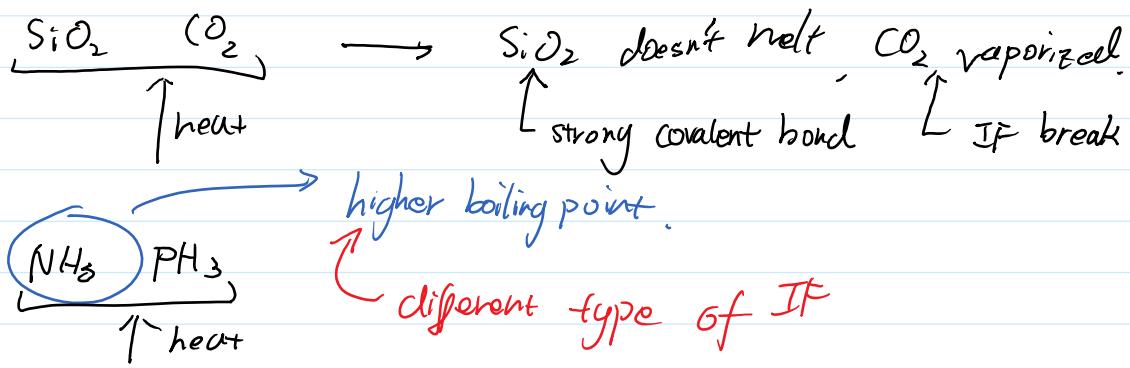
Bent

 $2s1p$ H_2O sp^3

Covalent molecule } CO_2 , H_2O , Cl_2 , CH_4 , C_60 , Carbon nanotube \leftarrow simple molecule
 Covalent molecule } SiO_2 , Diamond, Graphite, Graphene (one layer of graphite) \leftarrow macro molecule
 \uparrow higher melting point
 bonded by covalent bond
 Although look macro... it's simple
 giant covalent structure
 lattice

Some attraction between H_2O , or it will fly over
 but the water remains, so we know there is some
 force attracting them, calling Intermolecular Force (IF)

I_2 The Iodine solid also have IF, but it seems like
 they are different than H_2O 's



Intermolecular Forces

- dealing with simple covalent molecule only

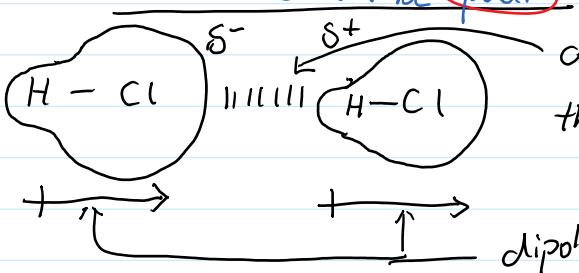
Simple molecule have low MP & BP \leftarrow due to weak IF
 there are three different IF

IF (electrostatic) } Hydrogen Bonding
 permanent dipole - dipole force
 Van der waals force (Temporary dipole dipole force) \leftarrow weak

Permanent Dipole - Dipole force (Polar)

Permenant Dipole - Dipole force (Polar)

the molecule must be polar

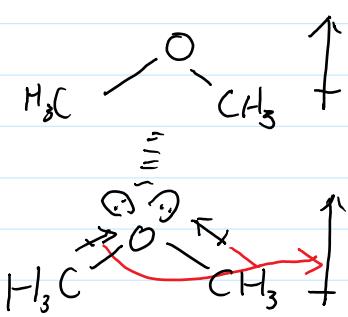


attracting according to charge: $\delta-$ \leftrightarrow $\delta+$

this is called permanent dipole-dipole force.

To Score

- Show Polar
- draw dipole
- draw dotted line



attraction between a slightly positive charge and a slightly negative charge.

||||| ~~X~~ ← this is abond!

draw this do not draw this we want attraction,

Van der Waals Forces (Non Polar)

- The major (maybe only) between non polar molecules.

no charge, but still attraction



$\propto \frac{1}{r^3}$ ← This is called van der Waals forces
 $3r - 1r$ instantaneous/temporary dipole

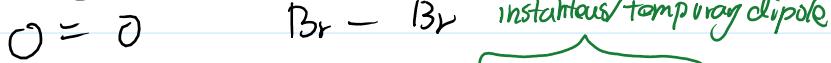
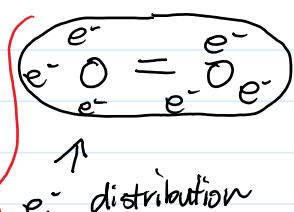
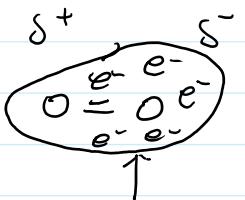


Diagram illustrating the electron distribution in the O=O bond. The oxygen atoms are shown with partial positive charges (δ^+) and partial negative charges (δ^-). The bond is labeled with a plus sign (+) and a minus sign (-) above the line, indicating a net dipole. Red arrows point from the text 'Slightly positive' to the δ^+ labels and from 'Slightly negative' to the δ^- labels. The oxygen atoms are shown with lone pairs of electrons ($\ddot{\text{e}}$ and e^-) and a central double bond ($\text{O}=\text{O}$).

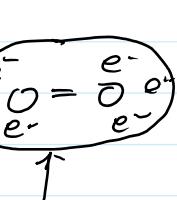
This
happens
fast
(ps)



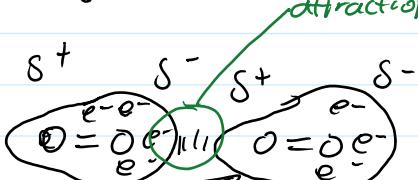
100



when two coming together



more concentrated
randomly ' area of e'



attraction force

$$s^+ \quad s^- \quad s^+$$

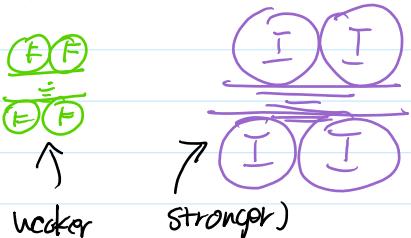
push electrons away

causing induced dipole (repel).

Two factors

- number of $e^- \approx (M_r)$
the more e^- , the stronger the repel \rightarrow stronger bond.

- Surface area of contact (size of the molecule)

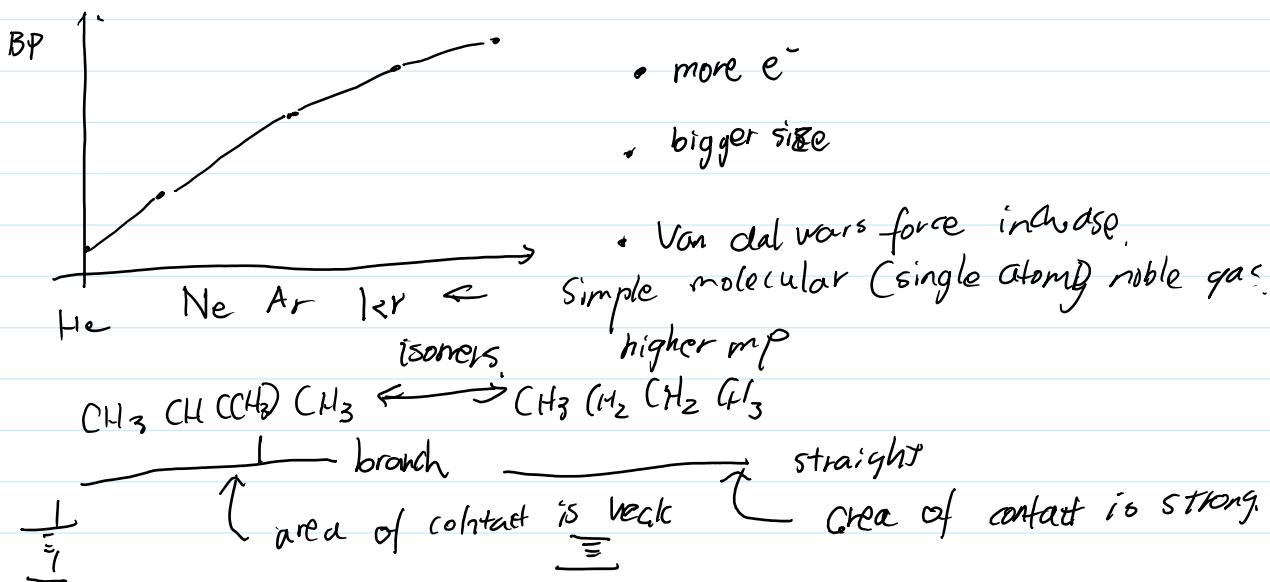


Van der Waals force

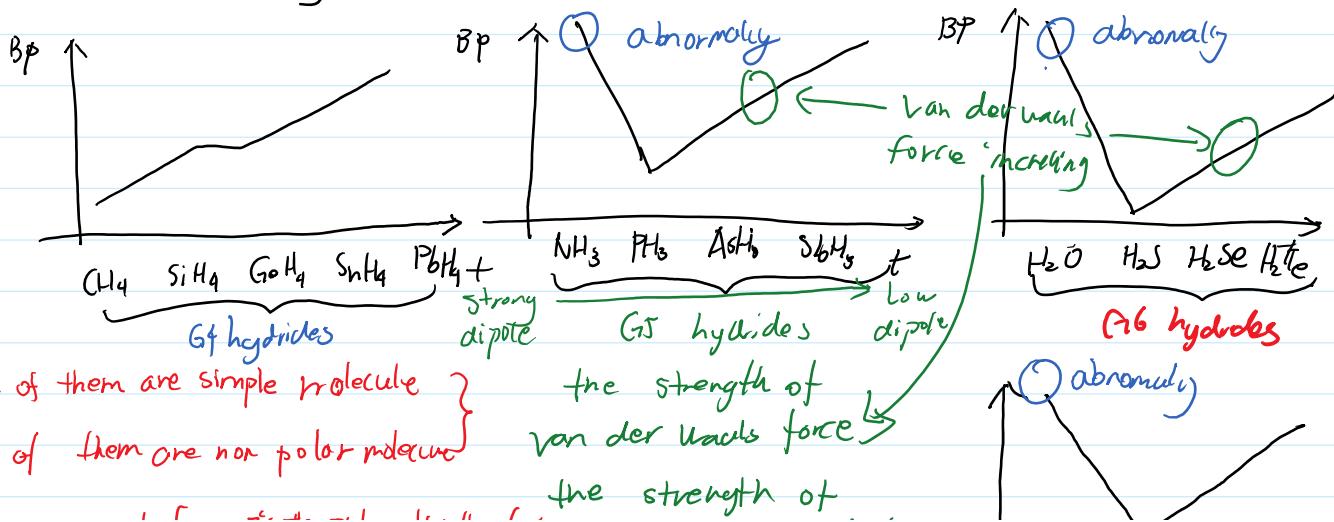
instantaneous dipole-induced dipole force

Notice

every molecule has Van der Waals force! (including PbBr_3)
It's just that it's too weak that we neglect it.

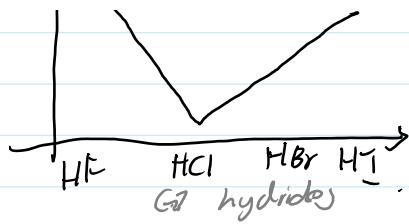


Hydrogen Bonding

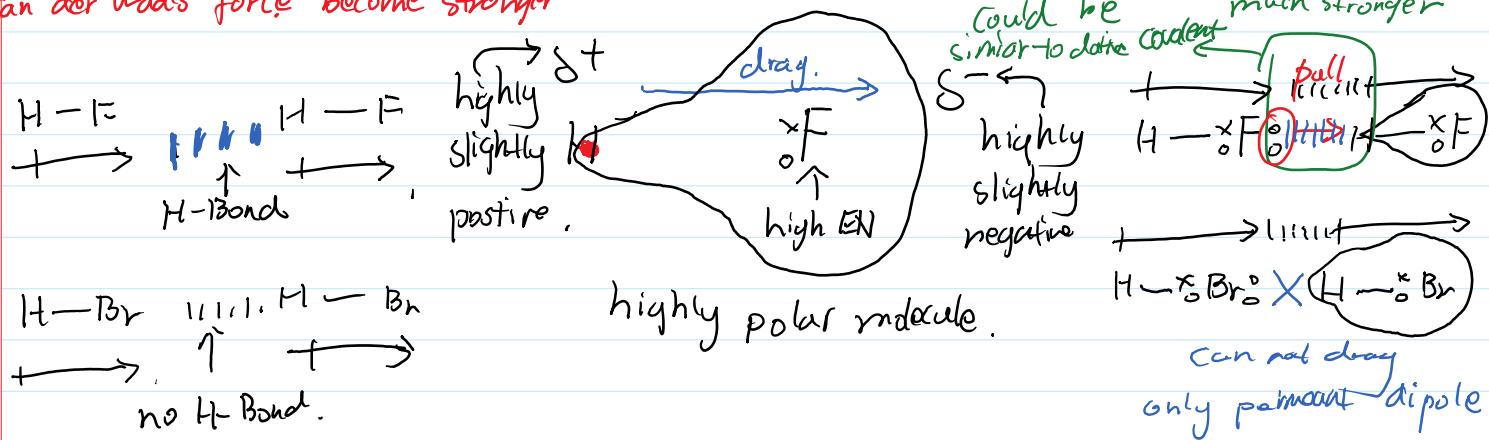


- all of them are simple molecule
- all of them are non-polar molecule

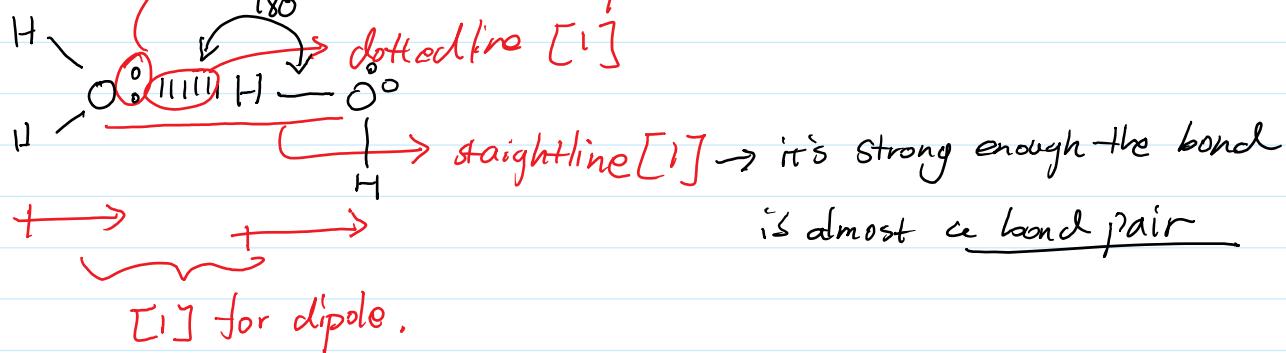
- all of them are non polar molecules van der waals force / the strength of
 - van der waals force is the only attractive force the permanent dipole-dipole
 - number of C^- increase force
 - shape increase.
 - van der waals force become stronger $\text{NH}_3, \text{H}_2\text{O}, \text{H}_2$



NH_3 , H_2O , HF form other bond \rightarrow Hydrogen Bond

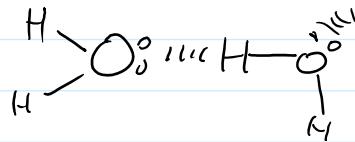


→ the attraction between lone pair and H nucleus. [1]



at 0°C lots of space for straight line H bond.

form ice
(less dense)



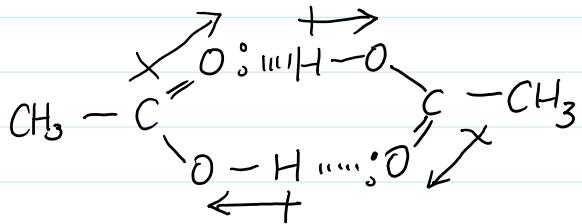
States of Matter

Thursday, November 30, 2017 8:09 AM

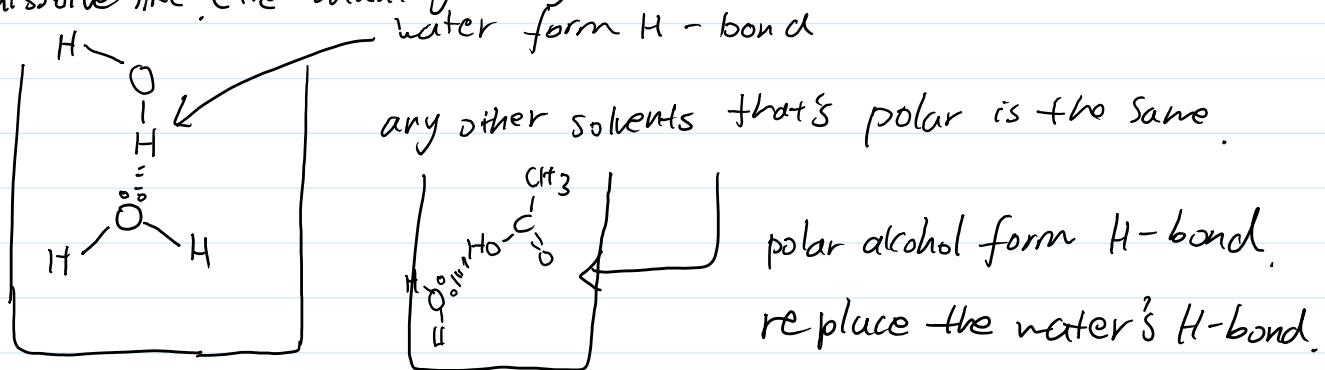
9/18

Monday, September 18, 2017 8:05 AM

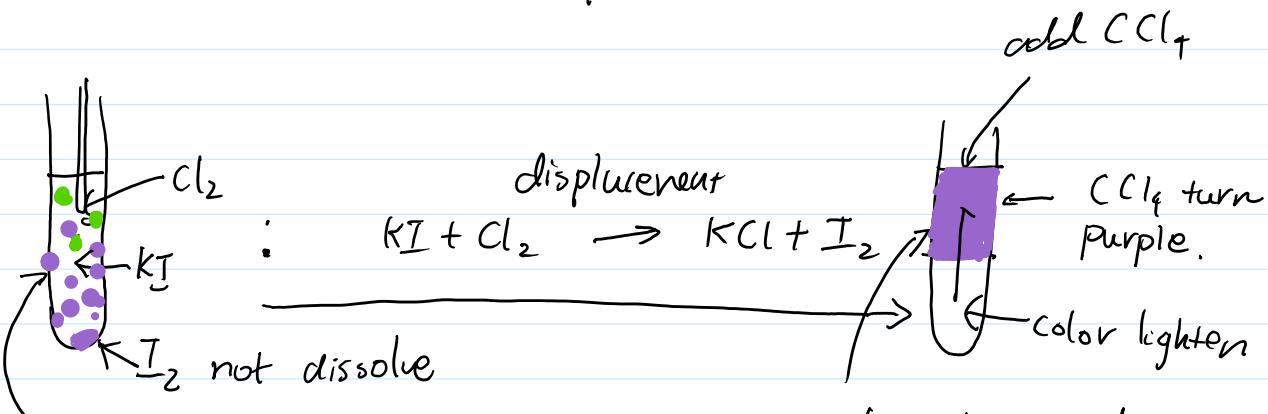
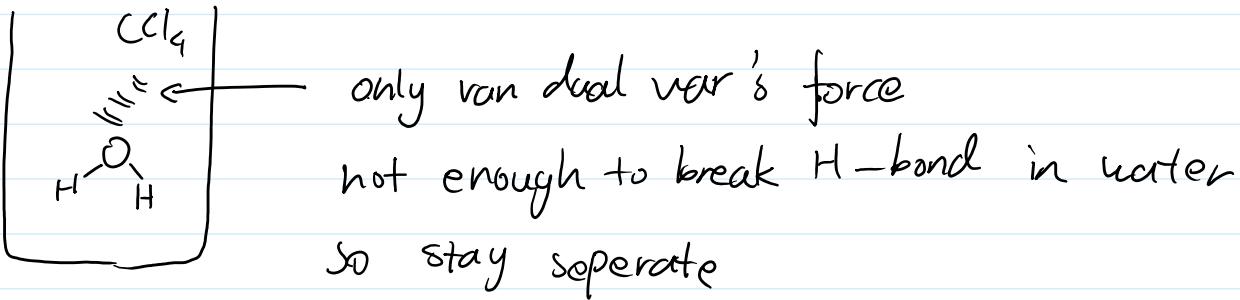
Many organic (carboxylic) acids form H-bond dimers in solid state.



Like dissolve like (the solubility rule)



the non-polar doesn't dissolve?



not dissolve I₂ (brown) in water not form H-bond.

I₂ dissolve better → non-polar.

(I₂ form Van der waal's force with Cl₂)

• notice:

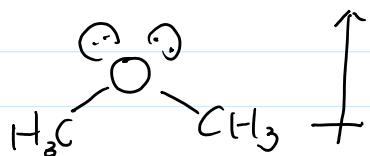
→ dissolves water → non polar.

• notice:

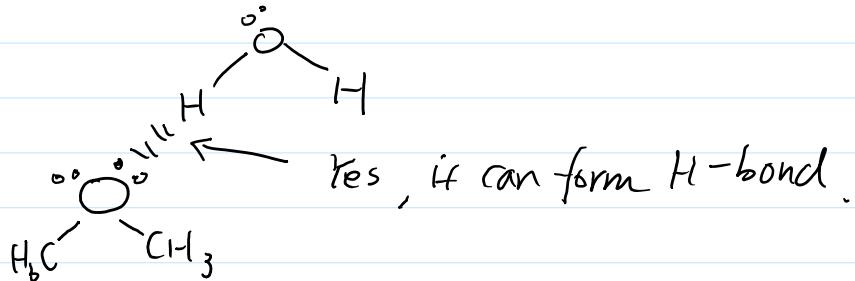
like dissolves like is only for simple molecules!

Is $\text{CH}_3-\text{O}-\text{CH}_3$ dissolve in water?

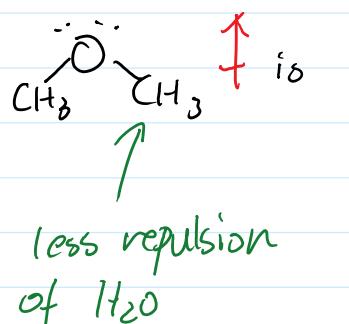
[1] Is it polar / non polar



[1] Can it form H-bond with water?



But C-H is non-polar

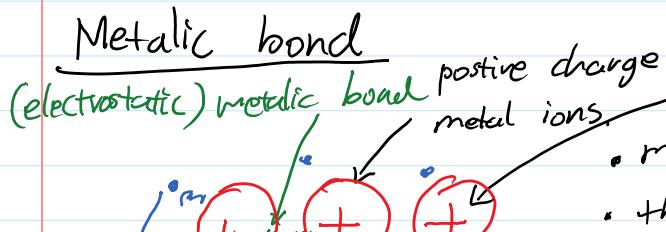


is better soluble than

chain is non polar
only form van der waals force.

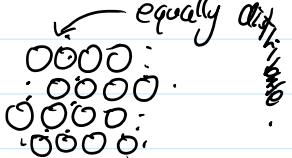
distribute the formation of H-bond.

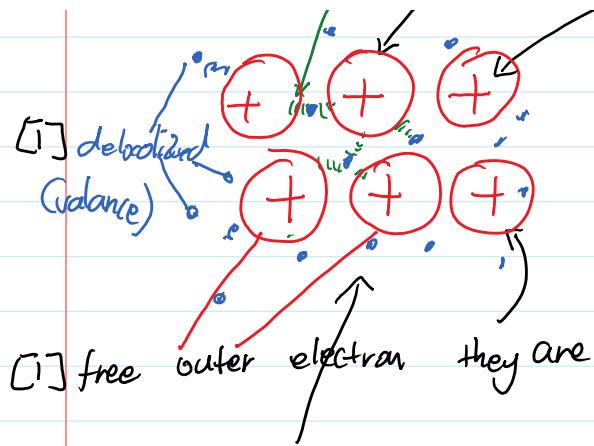
The longer the CH chain, the worse the molecules will be soluble



[1] proper diagram

- metallic bond can be found in all metal.
- they are less electronegative, so the





- metallic bond can be found in all metal.
- they are less electronegative, so the outer electron tends to delocalized
- form attraction between e^- and metal ions
- strong metallic bond is formed

They are called giant metallic lattice

$$Na \text{ (b.p.)} < Al \text{ (b.p.)}$$

- more free e^- delocalized ($Al = 3$, $Na = 1$)
- ion radius smaller

Stronger attraction

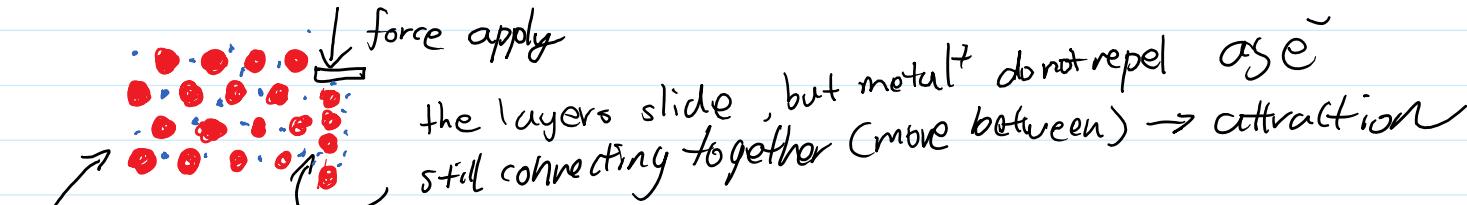


high charge density

also a better conductor (more e^-)

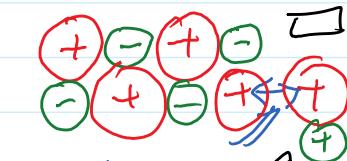
- good conductors as free electron carry charge.

- they are also malleable because the ions can slide past each other



orderly and repeating 3D pattern (lattice)

For ionic structure \downarrow force apply



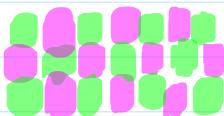
so it's brittle (break easily)

Structure

- giant ionic structure

- Simple molecular structure
- giant covalent structure (Macromolecules)
- giant metallic structure

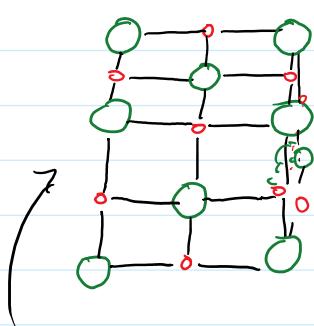
Giant Ionic Structure.



Bonding: Strong ionic bond.

structure: giant ionic lattice

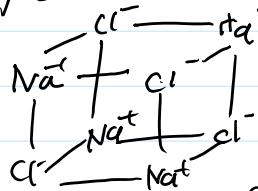
particles in the lattice: +ve and -ve ions



NaCl unit cell

1 Na^+ around is 6 Cl^-

→ too difficult to draw, you can simply draw

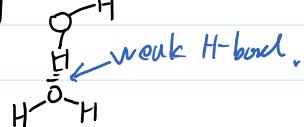


a not complete unit cell.
structure [1]

• high m.p. & b.p.

strong electrostatic between oppositely charged ions [1]
requires a lot of energy to breakdown [1]

NaCl	Diamond	Copper	Ice	Rock salt
high m.p.	high m.p.	high m.p.	low m.p.	lower m.p.
solid	solid	solid	solid	solid
not conductors	not conductors	conductors	not conductors	not conductors.



Compare:

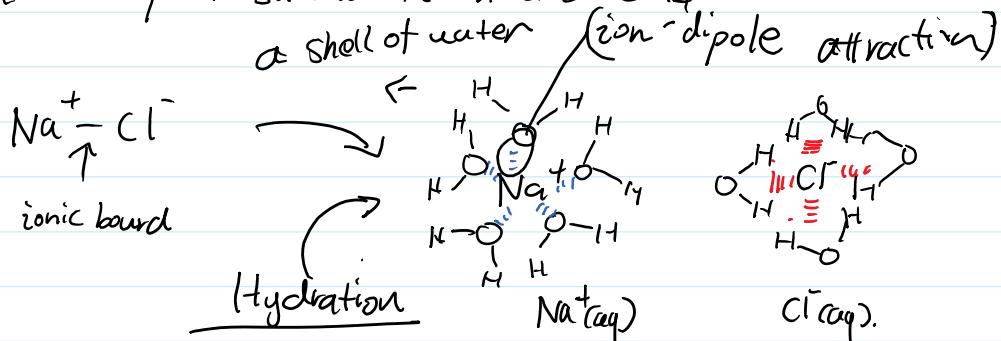
M.p. of NaCl & MgO

the attraction is weak

- higher the charge on ions
- smaller the radius.

Ionic crystals are brittle as discussed before

Most of the ionic compounds are soluble in polar solvents like water but insoluble in non-polar solvents like hexane CCl_4

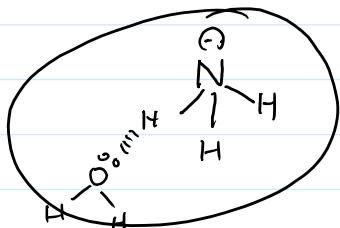


CCl_4 has no dipole and can not attract ions.

Ionic compounds don't conduct electricity in solid state but are good conductors in aqueous and molten states

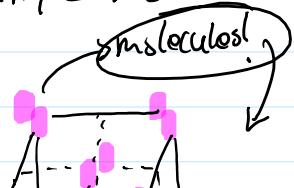
Simple molecules

- Molecular substances tend to be gases / liquid / low mp. solid (IMF is weak).
- Non conductors of electricity, except Carbon Nano tubes ~~graphite~~
- usually insoluble in water unless highly polar capable of forming H-bond with water: e.g. $^3\text{NH}_3$ & CH_3OH

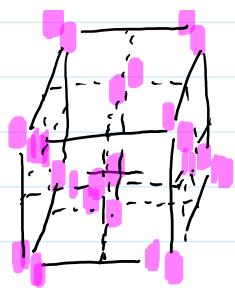


form H-bond with water

Simple molecular lattice



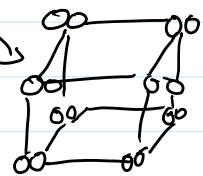
they can still form simple molecular lattice using



they can still form simple molecular lattice using

van der waal force.

OO |||| OO



weak van der waal force



strong
covalent
bond

we are only discussing the van der waal force for ionic

9/21

Thursday, September 21, 2017 9:31 AM

Ice / water

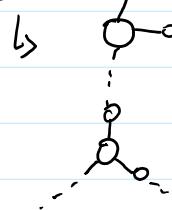
lower density due to H-bond

it's also a simple molecular lattice

ice floats on water

covalent bond between the atoms and hydrogen bond between molecules

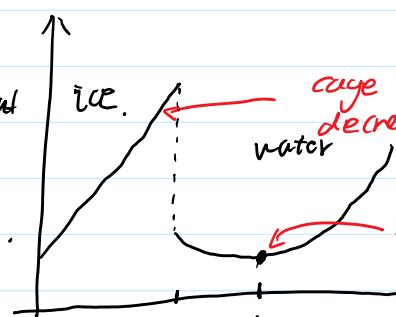
the lattice is not a perfect crystal



rage has different bond length.

many open space

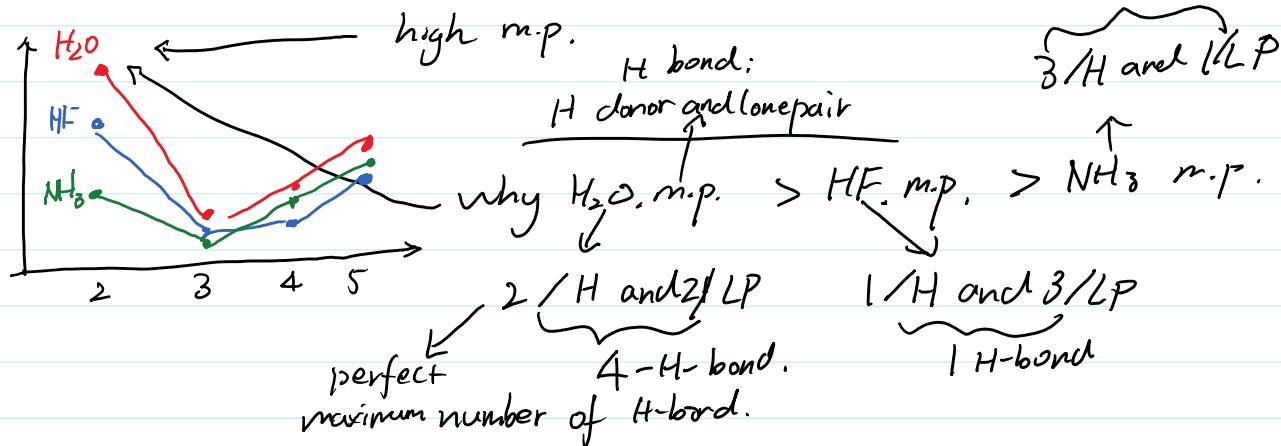
length.



cage cause waste of space
decrease of density

1 H-bond

similar to tetrahedral $\sim 109.5^\circ$



- high heat capacity \rightarrow cooler
- large surface tension \rightarrow H-bond pull water together
- high viscosity \rightarrow strong I.M.F. (difficult to flow)

Fullerene allotropes of carbon

Diamond
graphite
C₆₀
Carbon nanotubes

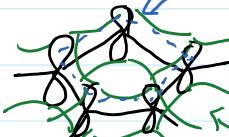
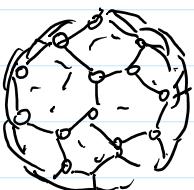
bucky ball (C₆₀)

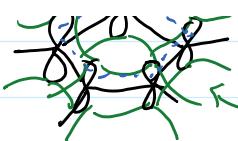
- each carbon forms 3 σ bond \rightarrow sp^2 hybridization

the fourth e⁻ form π bond.

they shift to form delocalized electron (a giant share orbital)

This means that C₆₀ is a good conductor within it





they shift to form delocalized electrons (a giant share orbital)

This means that C_{60} is a good conductor within it



The C_{60} can also be used as lubricant

- C_{60} not polar \rightarrow not soluble in H_2O
- May dissolve in $\text{CCl}_4 \rightarrow$ van der waals

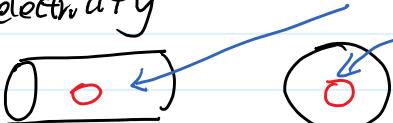


Carbon NanoTube

- $3 \text{C} \equiv \text{C}$ bond with delocalized electrons
- good conductor of electricity

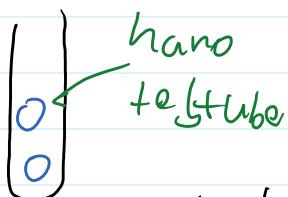
roll up the graphite \rightarrow carbon nano tube

the continuous π bond system

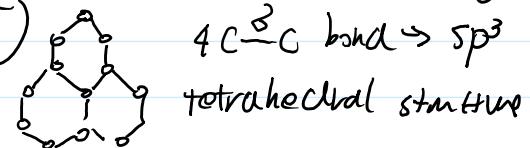
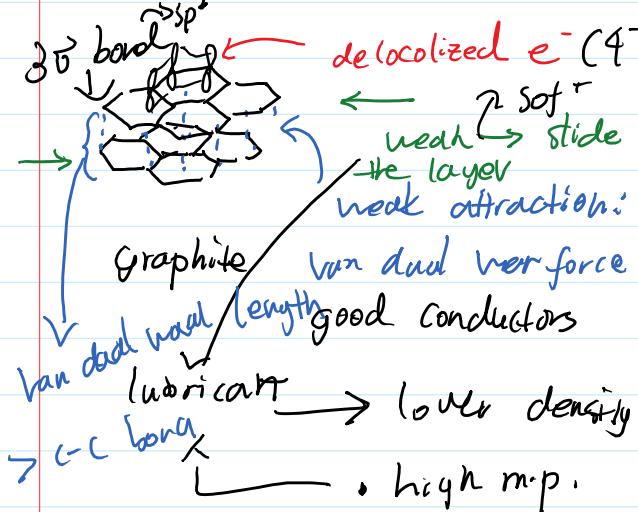


They can carry reactant element in it

\rightarrow control the configuration of direction

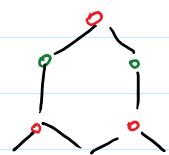


Giant Covalent structure \rightarrow lattice of atoms held together by strong covalent bond



$4 \text{C} \equiv \text{C}$ bond $\rightarrow \text{sp}^3$

tetrahedral structure



Strong-covalent bond



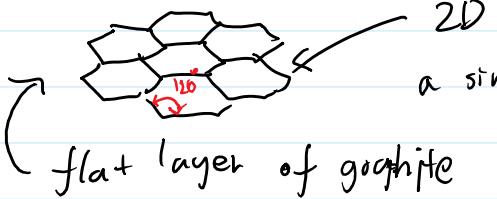
no solubility

not conductor

SiO_2

- SiO_2 strong bond
- diamond like structure

graphene:

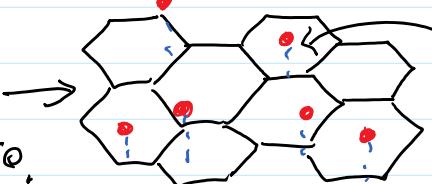


2D structure.
a single layer.

- good conductors \rightarrow wire
 \rightarrow cooling

hydrogen fuel cell storage.

also could be
use as catalyst surface.



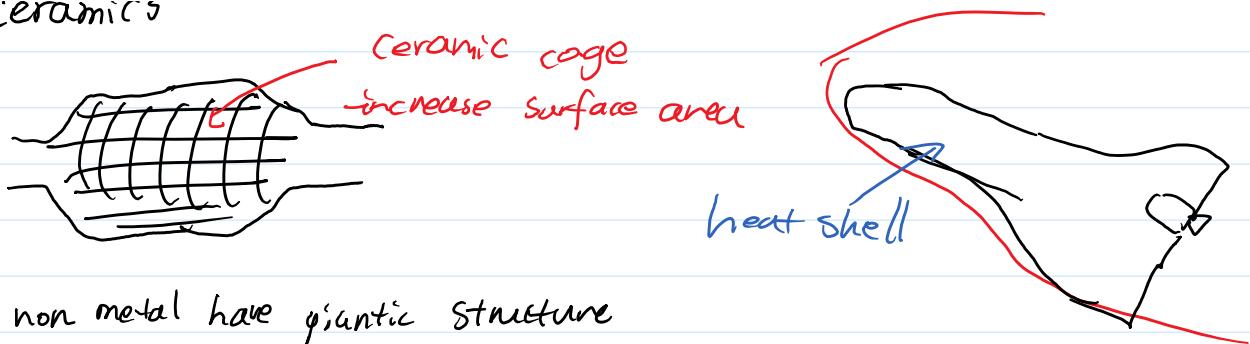
absorption \rightarrow
link by van der waals force.

Ceramics

Ceramic cage



Ceramics



a non metal have giantic structure

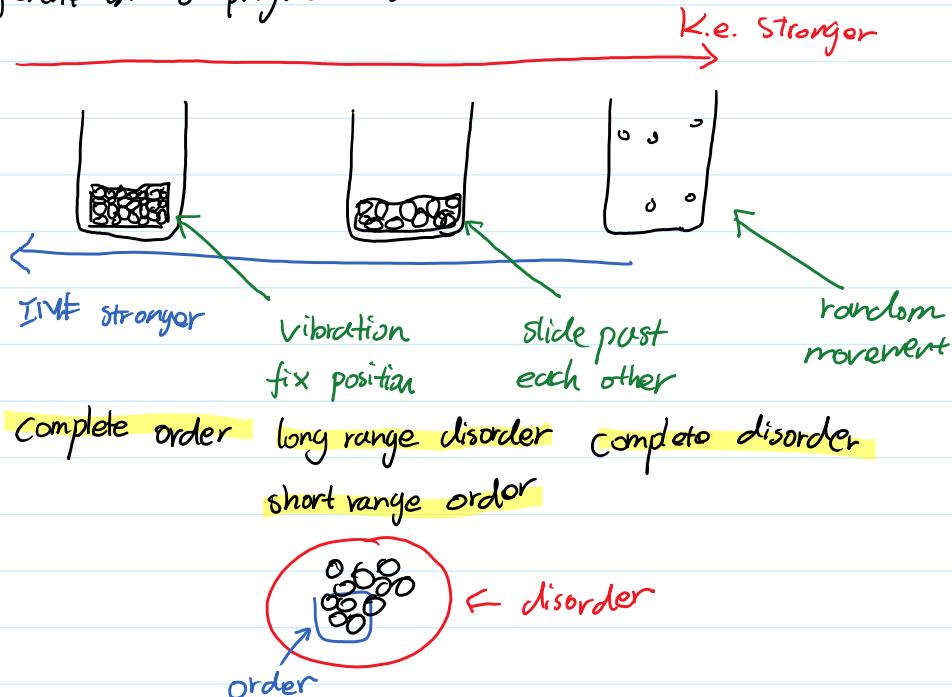
- high thermal stability
- electrical & thermal insulators
- hard and strong but brittle.
- furnace lining
- electrical insulators
- cooking utensils
- biological implants

9/25

Monday, September 25, 2017 8:36 AM

Kinetic particle theory

- matter is made of small particles
- Their arrangement, motion, kinetic energy, IMF and distance between them are different in 3 physical states



Ideal and Real gases

assumption for Ideal gas:

1. There are no IMF between particles
2. They have no volume for these particles (negligible)
3. The gas particles are in constant random motion
4. The collisions between the molecules are perfectly elastic (no kinetic energy is lost during collisions with each other and also with the walls of the container)

No Real gas are under normal conditions

He is the closest one

- London Forces only
 - lowest London forces (smallest ϵ number)
- He H_2 O Cl (lowest contact surface)

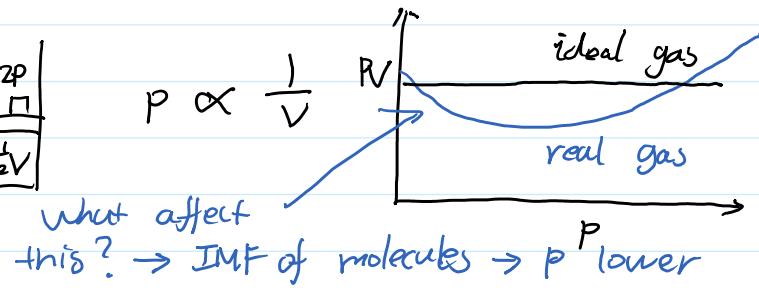
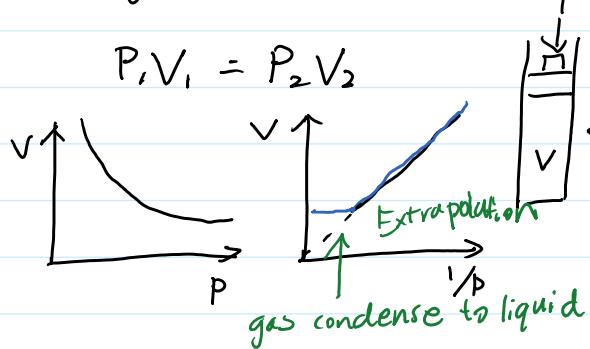


$$PV = nRT$$

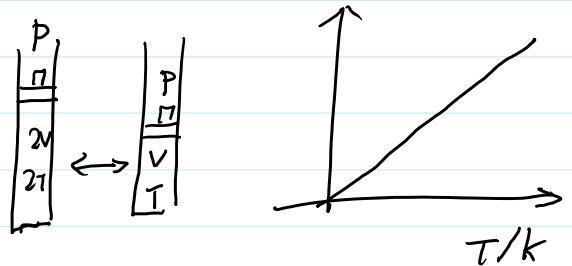
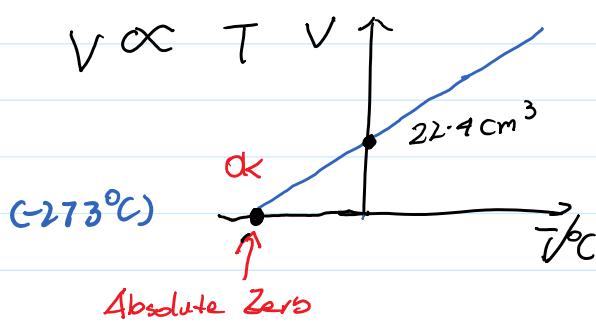
... T is constant

$$PV = \text{constant}$$

- Boyle's law: when T is constant, $PV = \text{constant}$



- Charles's law: when P is constant, $V/T = \text{constant}$



Thermodynamic Temperature

- Avogadro's law: when T and P are constant, $V/n = \text{constant}$ $V \propto n$

$$\therefore PV = \text{constant} \quad \frac{V}{T} = \text{constant} \quad \therefore \frac{PV}{T} = \text{constant}$$

$$\therefore V/n = \text{constant}$$

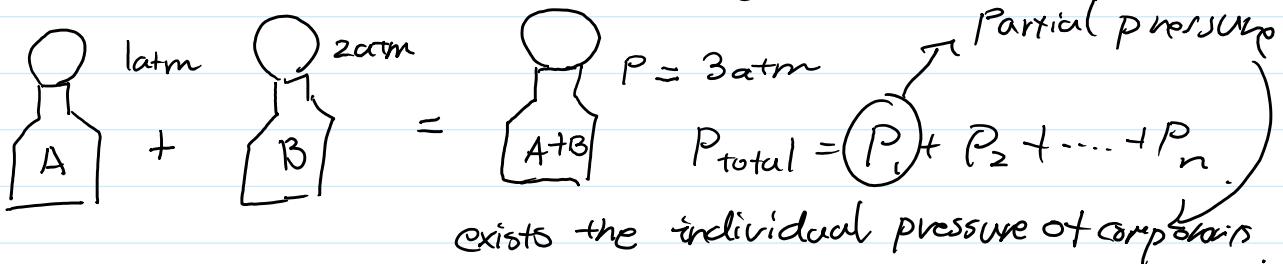
$$\therefore \frac{PV}{nT} = \text{constant} \longrightarrow PV = nT - \text{constant} \rightarrow PV = nRT$$

$$P = Pa \rightarrow 1 \text{ atm} = 10^5 \text{ Pa}$$

$$V = m^3 \rightarrow 1m^3 = 1000 \text{ dm}^3$$

$$n = m/M \quad R = \text{Gas Constant} = 8.314 \text{ J K}^{-1} \text{ mol}^{-1} \quad T = \text{temperature} \quad [K]$$

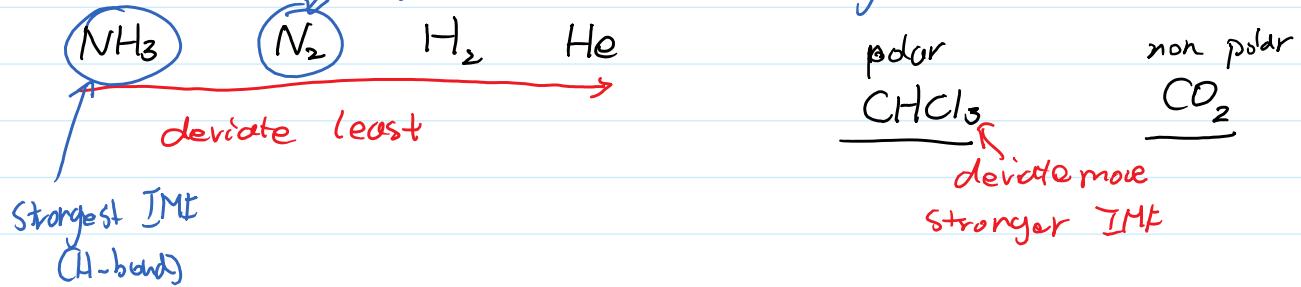
The ideal gas is the one which obey $PV = nRT$ exactly



Why do real gas deviate from ideal behavior?

Why do real gas deviate from ideal behavior?

1. There exists various types of IMF between gas molecules
2. Molecules do have a volume greater than zero
larger surface area/more charge



Real gases are most like an ideal gas at lower pressure and higher temperature

The real gases are most like an ideal gas

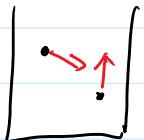
- At lower pressure
- At high temperature

At lower pressure \rightarrow less collision \rightarrow far away

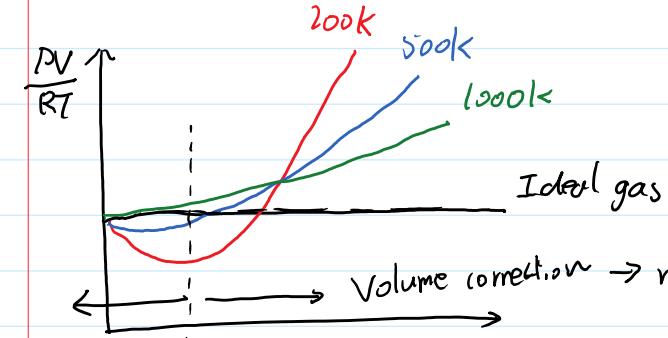


- Volume negligible
- IMF negligible (weak)

At high temperature

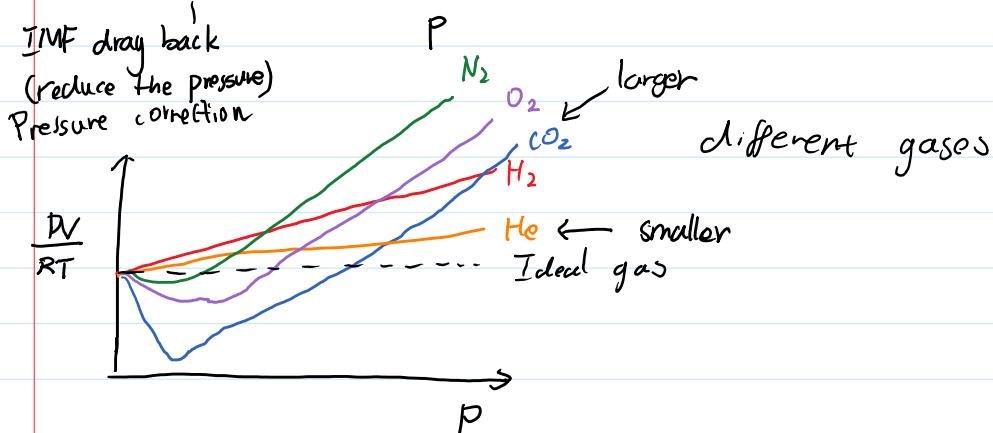


Kinetic energy high \rightarrow IMF negligible



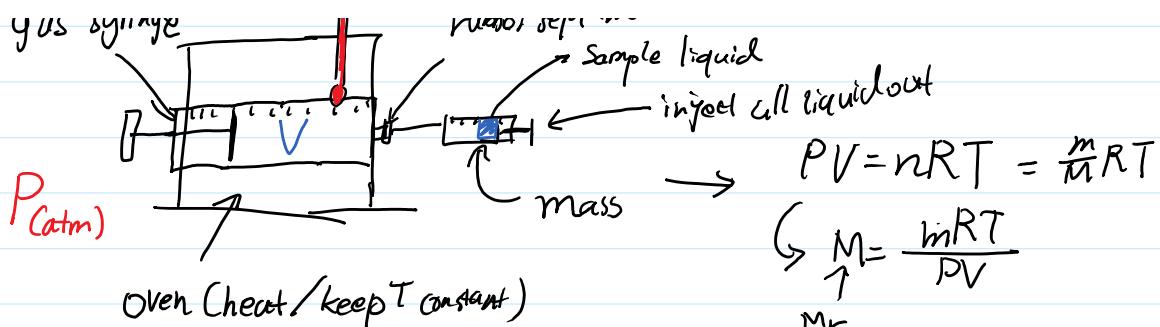
As $T \downarrow$, the real gas deviate from ideal gas further and further

Volume correction \rightarrow not to be proportional due to high pressure



Measure the M_r of a volatile liquid using ideal gas equation





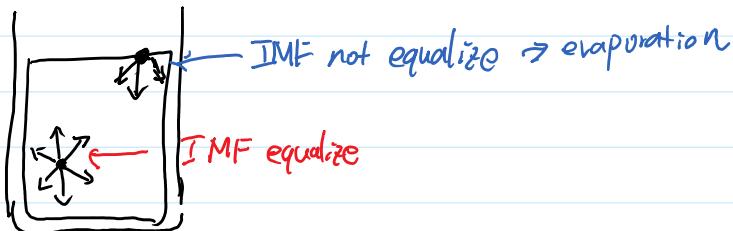
Limitation:

- Only for volatile liquid
- Liquid with low IMF is more suitable

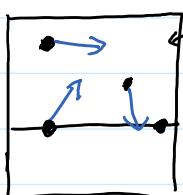
Liquid

- more motion
- different particles have different k_e .
- average $k.c. = T$

Some of the particles at the surface will have energies higher than $T \rightarrow$ overcome IMF \rightarrow evaporate



In a closed container vapour pressure \rightarrow maximum vapour pressure

 keep T constant \rightarrow get state of equilibrium

Rate of Evaporation = Rate of condensation \rightarrow fixed number of gas molecules \rightarrow constant

So we call vapour pressure in this state saturated vapour pressure

\rightarrow It's the pressure exerted by the vapor when it is in equilibrium with its liquid in a closed container at a constant T

$P_s \text{ H}_2\text{O} < P_s \text{ CCl}_4$, weaker IMF
Stronger IMF

Heat the liquid $\rightarrow V_p \uparrow \rightarrow$ stops at boiling point \rightarrow



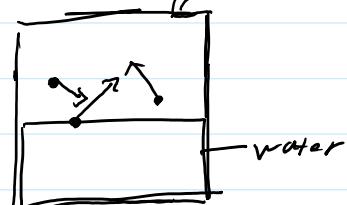
heat the liquid. \uparrow up \uparrow stops at boiling point \uparrow

B.P.

It's the T that the V_s equals to external pressure

1 atm

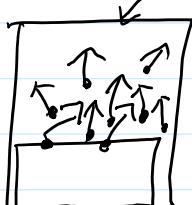
goes out at all
liquids, not only surface



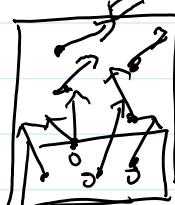
R.T.



V_s



increase



B.P.

1 atm
 V_p

$V_s \rightarrow$ formation of
bubbles

if \downarrow IMF \rightarrow $\uparrow V_s \rightarrow$ CCl_4 b.p. \downarrow
external constant.

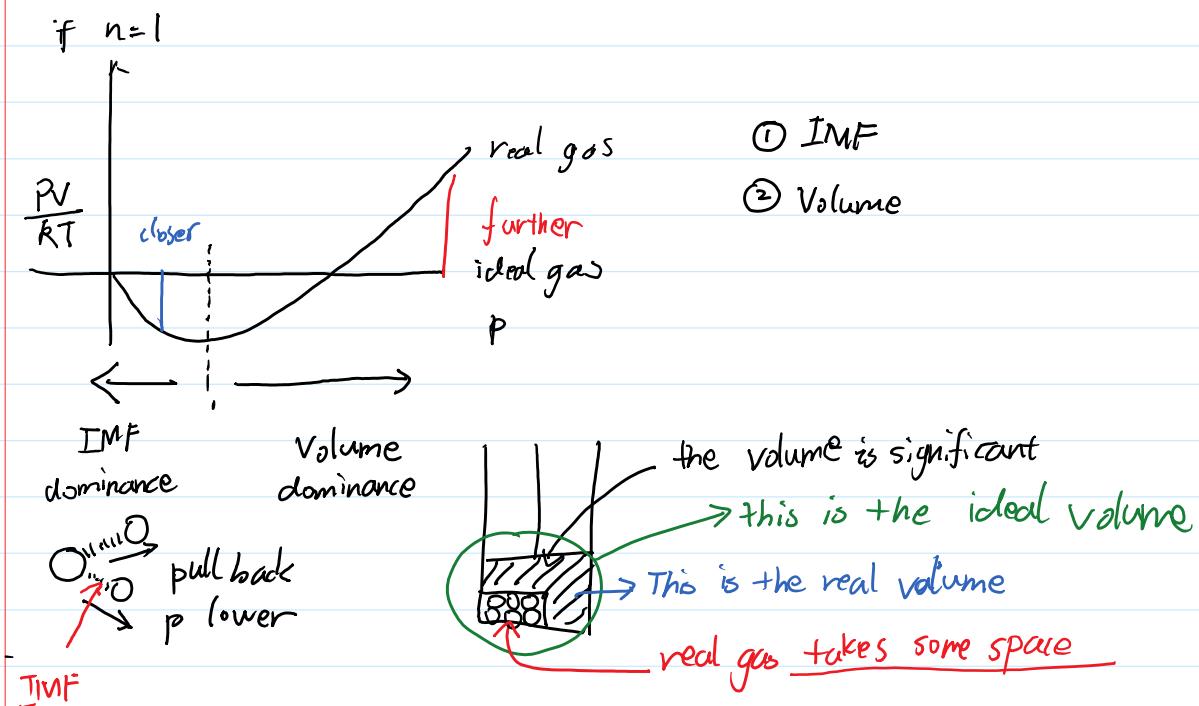
$\uparrow h \rightarrow \downarrow$ external \rightarrow lower V_s needed \rightarrow lower B.p.

pressure cookers \rightarrow \uparrow external \rightarrow higher $V.p.$ needed \rightarrow higher B.P.

Until B.p. IMF holds together \rightarrow cancel out

Enthalpy Change

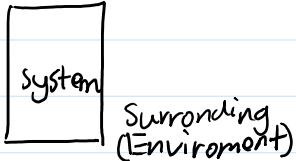
Thursday, November 30, 2017 8:10 AM



Chemical Energetics

Enthalpy change, ΔH

the heat absorbed or released by a system at constant pressure

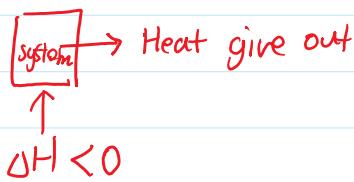


H
Enthalpy (chemical potential energy) \rightarrow Can not measure
We can, however, measure the energy change ΔH

Exothermic

Heat is released

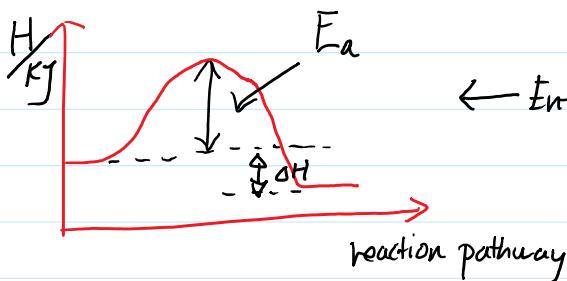
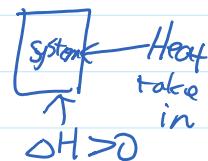
$$\Delta H < 0$$



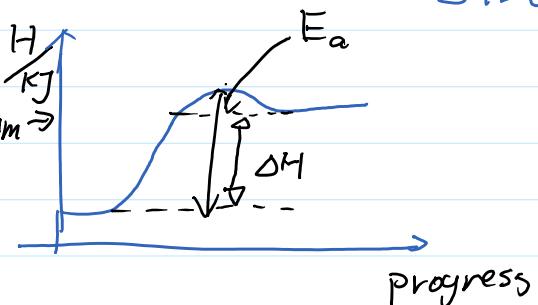
Endothermic

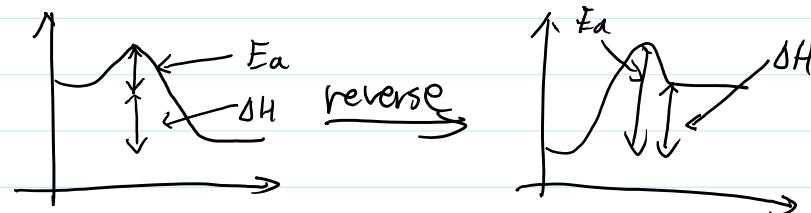
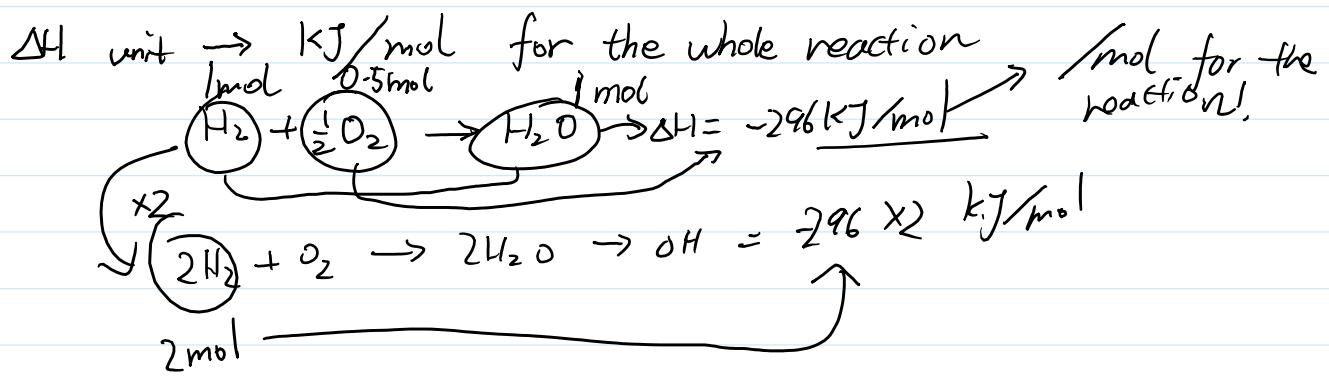
Heat is absorbed

$$\Delta H > 0$$



Enthalpy profile diagram





Enthalpy change is one of the variables involved when predicting whether or not a reaction will occur.

when a reaction is exothermic (ΔH is negative), that is a favorable condition

a favorable condition, but not always
 another factor is entropy \rightarrow this need to be equal for
 temperature

apply the same for here

If a reaction is endothermic (ΔH is positive), then it's an unfavourable

counter example: $\text{NH}_3\text{Cl} + \text{H}_2\text{O} \rightarrow$ continuous but Endo!

Standard Enthalpy Change for reaction (ΔH^\ominus)

• Many factors need to keep constant \rightarrow fair comparison

standard condition \ominus

and we put it ΔH^\ominus This is not enough

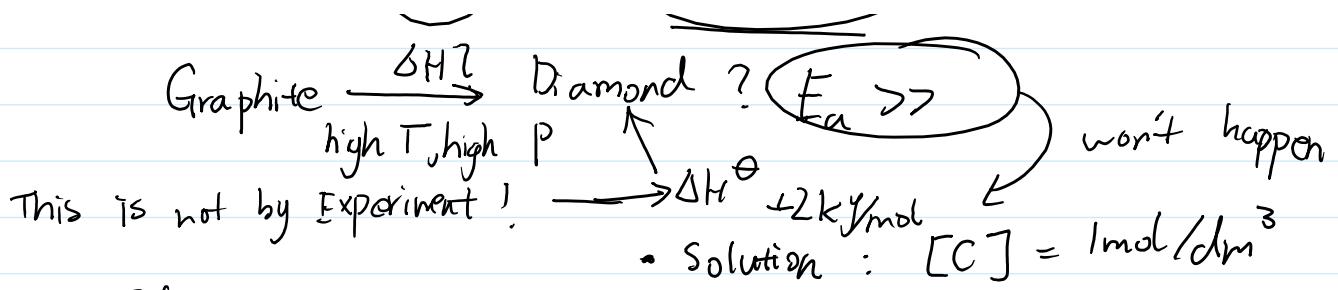
e.g. $\text{I}_2(s)$ $\text{Br}_2(l)$ $\text{Cl}_2(g)$

Cs C graphite This is the most stable

$\text{Graphite} \xrightarrow{\Delta H^\ominus} \text{Diamond}?$ (F \gg)

• P: 1 atm = 100 kPa
 • T: 298 K = 25°C
 • Reaction in most stable form

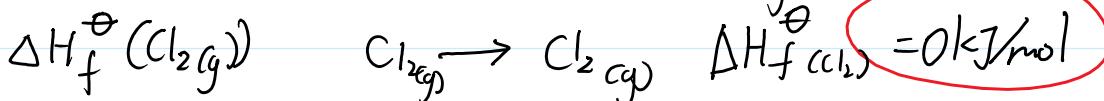
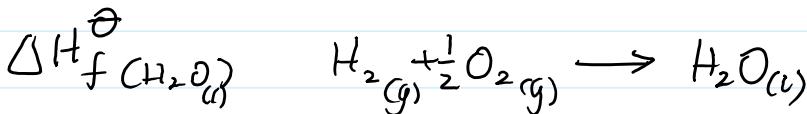
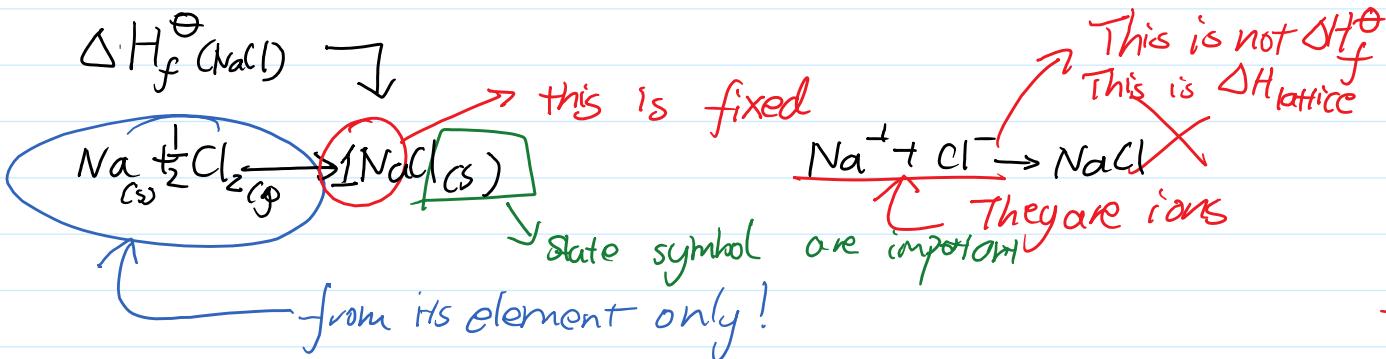
at this condition



ΔH^\ominus (6 topics)
 formation
 Combustion
 Atomisation
 neutralisation
 solution

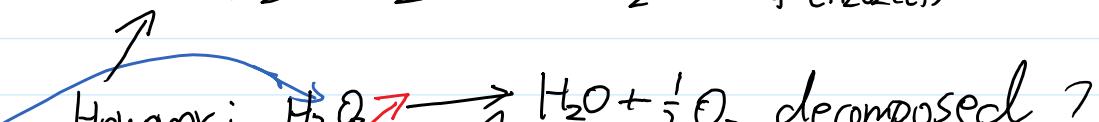
ΔH^\ominus for formation (ΔH_f^\ominus)

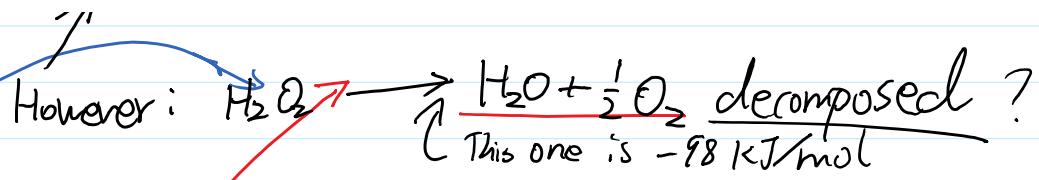
ΔH when 1 mol of a substance is formed from its elements under Θ



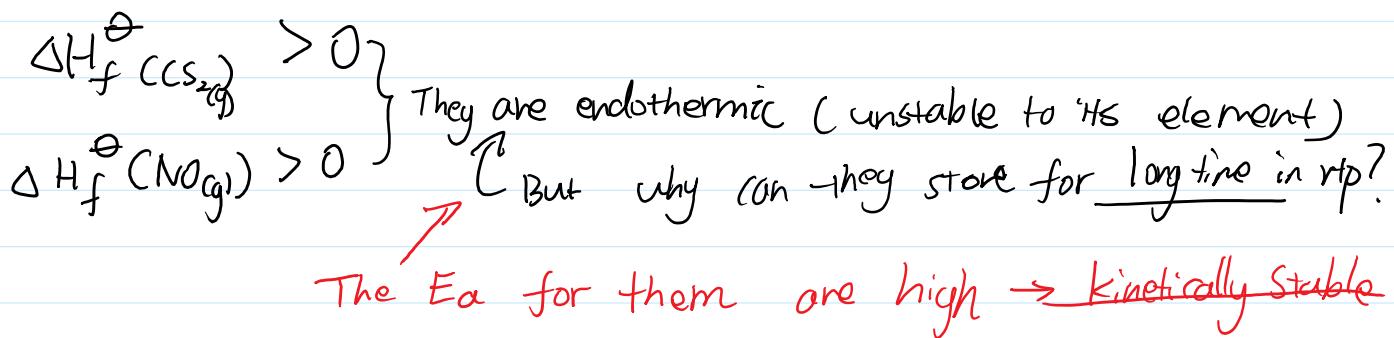
ΔH_f^\ominus for elements is 0, because there is no heat change!

if $\Delta H_f^\ominus < 0$, then it's more energetically stable than their elements





It's not decompose to its elements This is another reaction, so comparing the wrong subject



So we can see that the sign of ΔH is not always reliable.

- ① Reaction conditions not standard/catalyst is used
- ② Says nothing about kinetic stability of P relative to R
- ③ Entropy - another major factor