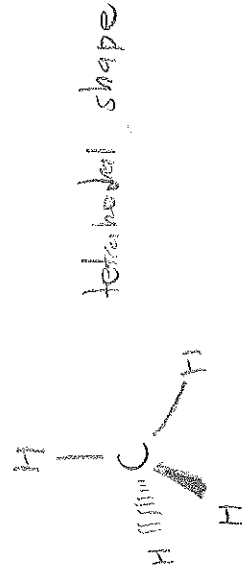


VSEPR

Valence shell electron pair repulsion theory explains about the molecular shape from the electron groups/dominants (bonding/non-bonding pairs). Double/Triple \rightarrow single bonds

- Example - CH_4



- it helps us to get the shape of the molecule + angles between them

- its hard to memorize! ✓

- the electrons try to be furthest away from each other

- Useful to determine the shapes and the angles of a molecule.

- linear $\text{H}-\text{C}-\text{H}$



\rightarrow symmetrical \rightarrow non-polar

- helpful in determine shapes

ex) shapes

- linear

- trigonal planar

Bent

- tetrahedral

Trigonal Pyramid

Bent or angular

- trigonal bipyramid

Seesaw

T-shaped

Linear

- octahedral

Square pyramidal

Square Planar

T-shaped

Linear

~~Seesaw~~

Metallic Bonding

Positive ions

In an e⁻ sea.



- bonding between metals
- has positive metal ions and delocalized electrons
- conducts electricity because delocalized electrons are free to move

non-directional bonds

- and carry a charge
- malleable because ions are free to move
- the electrons are free to move and are not shared.
- free ions so able to conduct electricity
- can be mixed to become alloys?
- malleable / high boiling / soft point

- strong
- very strong
- 1 or more metal

- Positive ions in an electron sea
- increased delocalization of electrons results positive ions in electron sea.

Non directional bond.

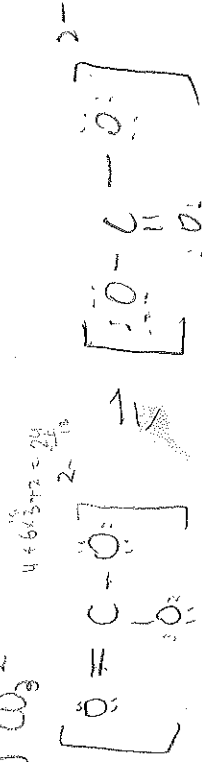
- Hard, malleable, conductive in solid state
- generally high m. and b. point
- exists in both pure metal alloy.

strongly different sized cations in sea of e⁻.

Resonance

- more than one structure

ex) CO_3^{2-}



When drawing resonance structures, the elements do not

move, the valence electrons changes places.
 Substances that have

V Resonance structures are more stable \uparrow eg: CO_3^{2-} ,
 then Subs. without resonance O_3 ,
 C_6H_6

bond length is intermediate between double and single bond.

- can determine best structure by calculating the formal charges
 & closest to 0 for all atoms is best

ie.

Ozone



Hybridization

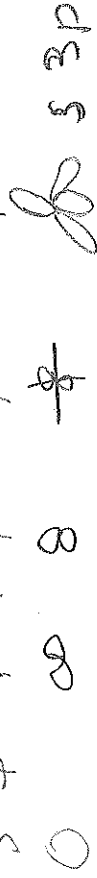
$s + p = sp$ linear.



$s + p + p = sp^2$ trigonal planar



$s + p + p + p = sp^3$ tetrahedral



$s + p + p + p + d = sp^3d$ trigonal bipyramid



$s + p + p + p + d + d = sp^3d^2$ octahedral



- VSEPR can predict shapes of molecules, but doesn't take into account atomic orbitals (s, p, d, f)

- determine shapes of the e^- using VSEPR theory.

- from mixing different types of atomic orbitals on the same atom



σ and π bonds

head-on collision



first bond formed

overlapping

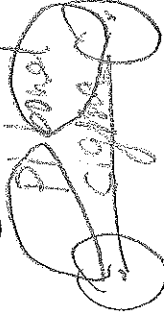
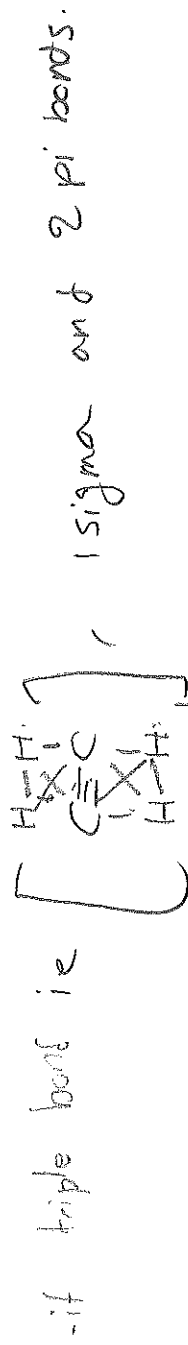
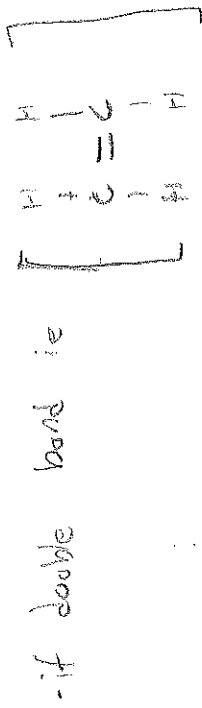
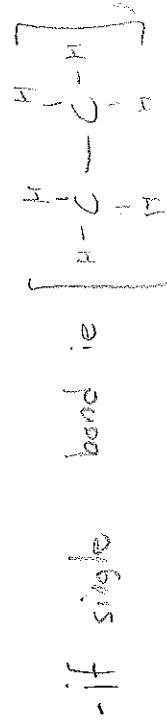


second, third bond formed.

$\sigma \rightarrow$ weak, long

$\sigma + \pi \rightarrow$ stronger, shorter

$\sigma + 2\pi \rightarrow$ strongest shortest



- pi don't rotate

Intramolecular Forces

strongest
hydrogen bonds: hydrogen usually bond with E, O, N
to form strong Dipole Dipole bonds.

London forces: force between molecules in a non-polar
(temporary dipole)

dipole-dipole forces: forces between polar molecules
(permanent dipole)

hydrogen bond is stronger than
dipole-dipole force which
is stronger than London force.

London force exist in all ~~polar~~
~~molecules~~ substances

hydrogen and dipole dipole exist in only polar
substances.

H-bonds are due to H bonding with highly
electronegative elements, O, F, N leading to
significant dipoles.

- London forces occur when a temporary dipole is created through the
transient electron repulsion

- weaker than chemical bonds

- ~~covalent~~

Ionic Bonding

- Between non-metals and metals
- stronger than ~~ionic~~ ^{covalent}
- NaCl
- They are able to conduct electricity only ~~not~~ when molten or in solutions
- crystal
- Lattice structure forms from electrostatic attractions between atoms that have lost electron atoms
- electron apt able to move in solid state / have gained electrons.
- Intramolecular force - ionic compounds do not form molecules ✓
- (ionic compounds) can dissolve in water

www-hard

- melting point and boiling point are higher
- Volatility low.
- formed from e^- being transferred.
- One ~~electron~~ ^{electron} forms e^- (oxidized) and the gains e^- (reduced).
- brittle because lattice structures are unable to change

ie. NaCl lattice structure



Covalent Bonding

1. between non-metals and non-metals
2. the e^- are shared between the atoms
3. the e^- are shared, e^- are nearest to greater Z .
4. may be volatile
5. intramolecular force
6. weaker bonds than ionic bonding
7. molecular elements are non-polar.
8. molecular compounds can be non-polar or polar depending on the bonding structures.
9. Electrostatic attraction b/w shared pair of e^- and + charged nuclei.

(1) consists of molecules not always!

(1) not insoluble in water \in no ions present to carry charge.

(2) don't conduct electricity as no ions are present to carry charge. where atoms are covalently bonded together in a massive network

Polarity

The e^- are slightly more attracted to one atom than another in a compound making the slight positive (non-polar) e^- as the 6^{th} .

the difference between electronegativity

- > 0.5 1.8 ionic (non-polar)
- < 0.5 1.8 polar covalent
- $= 0$ pure covalent

polar substance is soluble in polar solvent

non-polar dissolves in non-polar

ionic also dissolves in ~~non-polar~~ polar

- polar bonding is typically stronger than non-polar bonding

(Hydrogen bonding or dipole-dipole vs London Forces)

- same dissolves same

- polar v.s. non-polar

- the more polar a molecule is, the stronger the intermolecular force, so higher boiling point and less volatile.

- polar mix with polar

- usually ~~same~~ structure w/o lone pairs are polar

↑ not always!

usually
non-polar
intermolecular

Intermolecular

2-3

Forces

→ determines solubility

Polarity, electronegativity

Hydrogen bonding, surface tension, ~~viscosity~~

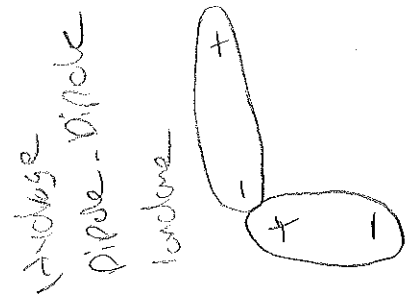
London forces; dipole-dipole attraction,
Van der Waals forces



London forces → Temporary
dipole dipole → permanent

London Force Every compound has London force non-polar
 Dipole-Dipole polar - interactions between dipoles of different molecules
 Hydrogen bond polar only for H-N, H-F, H-O

Between molecules \leftarrow more polar
less polar



(London dispersion forces)
 London force < Dipole < Hydrogen Dipole
 ↑ Polar

Non-polar molecule
 but because we can have π electrons in a certain area temporarily it allows the molecule to be temporarily polar in order to bond

London: temporary dipoles made

greater π e⁻ / protons
 = greater London Force

Pi and Sigma Bonds

single bond | Sigma \Rightarrow can spin on horizontal axis
 double bond | Sigma + 1 Pi cannot spin
~~triple bond~~ | Sigma + 2 Pi

(z-axis)

spin

rotation
 symmetry

Pi: π
 Sigma: σ

Sigma - head to head - s-orbitals
 Pi - tip to tip - p-orbitals
 side to side

e.g.

1st bond is sigma bond!
 \rightarrow 1st bonds follow

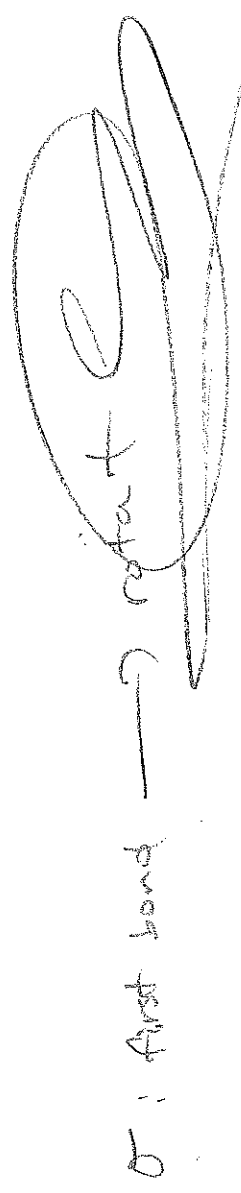
double bonds
 both = u/b

and increase accordingly

o mathematical functions explaining bonds:

Pi: π
 Sigma: σ

$\sigma \rightarrow$ head on \rightarrow the first bond
 $\pi \rightarrow$ sideways \rightarrow all the bonds after the first one



π : 2nd, 3rd bond.
 σ bonds can rotate
 π bonds can't

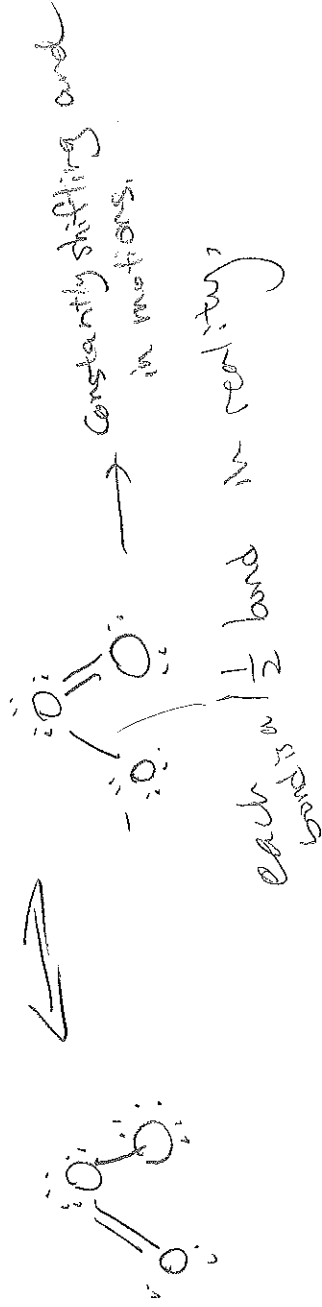
pi bonds are vertical
 $\sigma \Rightarrow \sigma$

Resonance

obtaining multiple Lewis structures for one molecule or ion such as O_3 which contains 2 different Lewis structures

Probability of electron location

example



FC = valance e^- - ($\frac{1}{2}$ bonded e^- - lone pair)

→ FOR MAX CHARGE is one of the more likely shapes/structures a molecule can

inherent structures with resonance

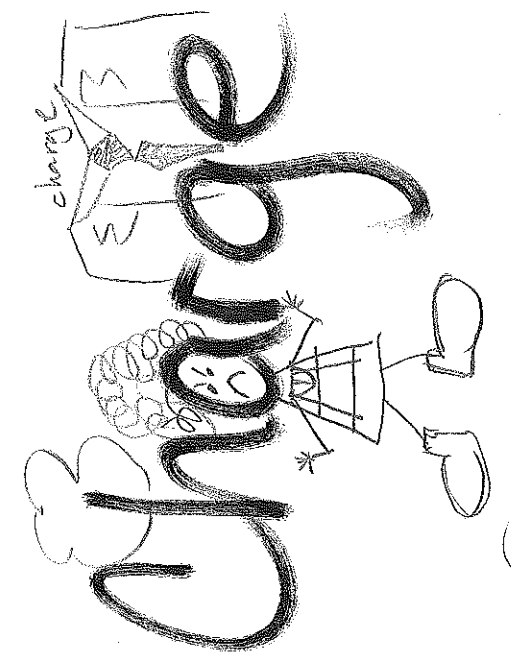
- unusual stability compared to structures without

i.e. Benzene

- different structure with the same molecule

They "resonate" with me 😊

Formal Charge



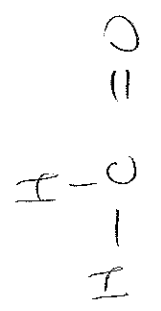
- It's the charge.
- original charge of an element.

- # valence electrons - $\frac{1}{2}$ # of bonds - # lone pairs
- which structure is most likely
- big
- it's not informal

GM also counts things touching element and subtract that from valence

They're radical :P

- Used to determine the best structure of a molecule
- helps determine likelihood of a structure
- Ozone
- Nitrate anion
- O is the best

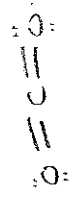


- Nitrate anion \rightarrow best structure
- O is the best
- preferred structure depends on FC close to 0
- if FC cannot be avoided, the (-) FC should reside with the more electronegative element
- use to choose the perfect structure.

Lewis Structures

- Valence electrons
- ~~VSEPR~~
- hybridization
- Bonds

- 1) Valence electrons for each element
 - 2) each bonding pair of electrons is written as a line
- $\text{CO}_2 \rightarrow \text{C} \rightarrow 4$
 $\text{O} \rightarrow 6 \times 2 \rightarrow 12$



- Way of showing bonds
- Shared pair of electrons = line (-)

~~tetrahedral, bent, octahedral, linear, trigonal bipyramidal, square pyramidal~~

- Lewis dot diagram

- ~~VSEPR~~
- ~~Octet Rule~~
- ~~Bonds~~
- ~~hybridization~~
- ~~Linear, tetrahedral, trigonal bipyramidal~~

- invented by Lewis.
- best ever. ~~Worse~~

- helps to visualize bonding shapes/structures

Electron Dot Structure.

- " - means 2 e⁻ "
- " • means 1 e⁻ "
- single bond
- = double bond
- ≡ triple bond

can show ions with parenthesis



formal charge helps determine shape of Lewis structure.

$$\text{FC of sth} = \text{valence } e^- - \left(\frac{1}{2} (\text{bonding electrons}) + \text{loose } e^- \right)$$

length: $- > = > \equiv$
strength: $\equiv > = > -$

- structures named by Lewis

- least subscripts before

- will be on the test for sure

Covalent Bonding

- Shared pair of electrons
- Attraction between shared pair of electrons and nucleus
- Always ~~Always~~ between non-metals, non-metal
- Intramolecular
- between non-metals and non-metals
- ~~London force~~
- polarity
- acid/~~base~~
- HCl H₂SO₄
- shared pair.
- non metal - ~~metal~~ non-metal
- share electrons.
- gas. ?
- low melting point
- doesn't conduct electricity what about alcohol?
- not soluble in water sugar?
- polar covalent / pure covalent
- low boiling
- may be volatile
- share electron clouds
- consists of molecule
- repulsion of nuclei & attraction of e⁻
- Another bond different from ionic and metallic



Polarity

- electronegativity → creates δ^+ side and δ^- side
- polar / non-polar molecules

- $\Delta \chi_p \rightarrow$ polarizability \propto net polarity and # of $e^- \rightarrow$ m.p. & b.p.
- S^2, S^- A.O.G.
- $CH_4 \rightarrow$ non-polar



$\ominus \rightarrow$ low volatility \rightarrow point

must use shape of molecule (VSEPR)

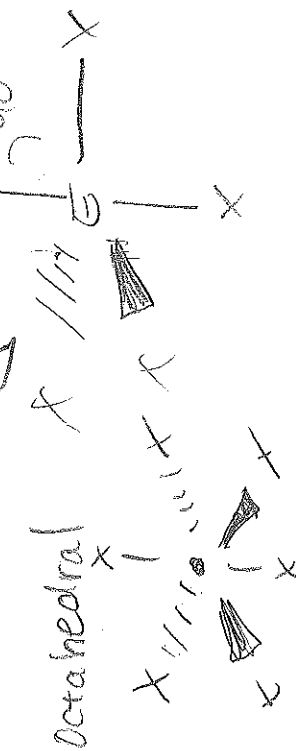
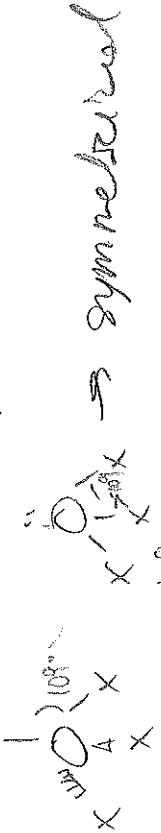
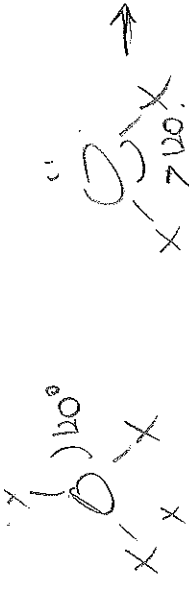
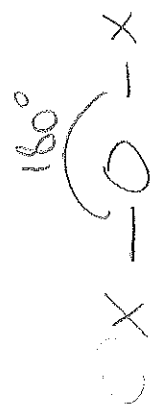
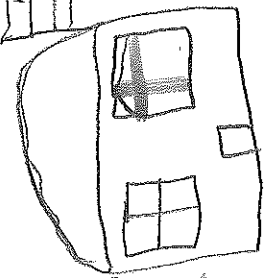
- \rightarrow create dipole - dipole bonds.
- symmetry makes for non-polar
- arrow points at more electronegative value. - more difference = more polar
- polar dissolves polar
- non-polar dissolves non-polar
- dipole-dipole forces between polar molecules
- London forces between non-polar
- its not polar bears - polar can bond with polar \rightarrow
- Polar = $0 < x < 1.8$ - we can use χ to find polarity
- $\Delta \chi_p > 1.8$ ionic
- $0 < \Delta \chi_p \leq 1.8$ polar covalent
- $\Delta \chi_p = 0$ pure covalent at each side
- Non-polar balance will go to the stronger one. stronger bonds causes bonds
- Polar - the most electron
- polar / non-polar. - no ionic bond
- electronegativity

strongest = hydrogen bonding

More polar \rightarrow

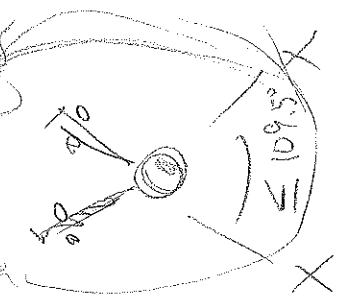
VSEPR Shapes

MIT
green



Bonds = lines
 Double bonds still lines

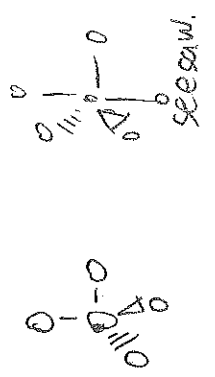
- explains molecule theoretical shapes (areas the electrons may be)
- helps determine polar (non polar)



- doesn't show characteristics of s, p, d, f

- you can use it to determine if something is polar

or not



Hybridization

- mixing of atomic orbitals: s, p_x, p_y, p_z

2	linear	sp	
3	trigonal	sp ²	
4	tetrahedral	sp ³	
5	Trigonal Bipyramidal	sp ³ d	
6	Octahedral	sp ³ d ²	

e.g.

s, sp
sp², sp³

Clarifies VSEPR

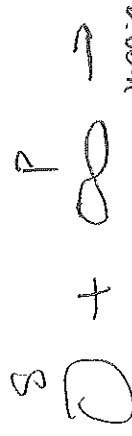
↳ explains the molecular shapes by adding

in pi π -sigma bonds



Shows % of orbital characteristic

Mix of orbitals.



sp hybrid

① Find # of bonds around central e⁻

② Find # in table

Yes

Shows characteristics of sp, d, f.

It sure does!

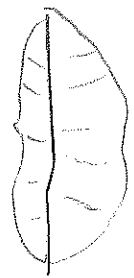
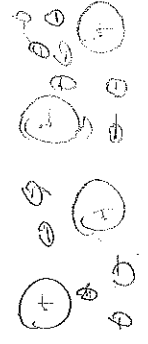
A theory to explain two contradictory theory

- linear, trigonal, tetrahedral, trigonal bipyramidal, octahedral
(sp) (sp²) (sp³) (sp³d) (sp³d₂)
= combination of orbits.

Metallic Bonding

↳ mix of metals (alloy)

ELECTRON SEA MODEL:



→ delocalized e^- → can move
→ malleable

Alloy = mixture of metals conducts electricity

→ high melting / boiling points

→ conductive → works better than copper

→ in metal: $NaCl$???

Metal

- not lattice structure — stronger metallic bonds

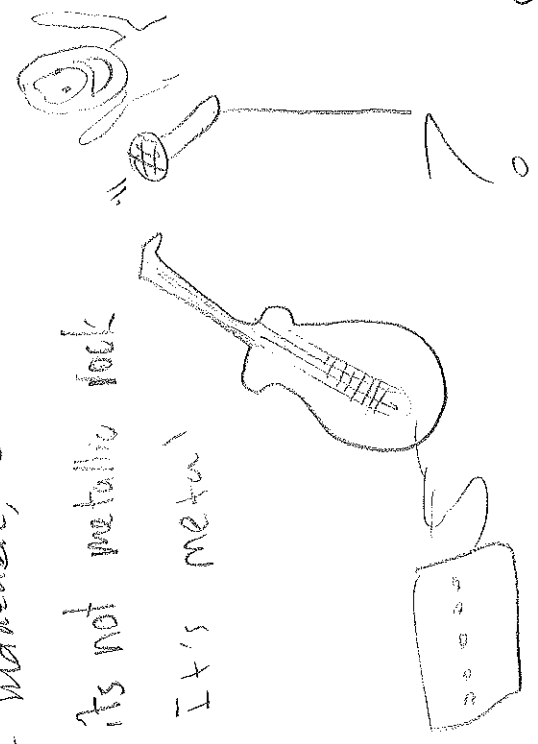
- "sea" of delocalized electrons

- alloys are stronger, because different sized atoms break up sheets

- malleable, ductile, conductive: • High M.P and B.P — in the solid state

- it's not metallic rock

- It's metal!



strength depends on → charge of metals ions
↳ radius of metal

attraction of positive ions and sea of

Aluminum Bonding

recess with both polar and non polar — strength of metal.
— decide melt and boiling bond of metal.



Resonance



can be written as

- Same compound but different placing of the bonds



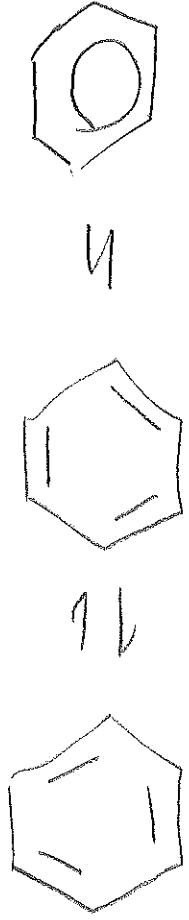
= molecules with different Lewis structures

- Formal charge can be used to detect the best structure.

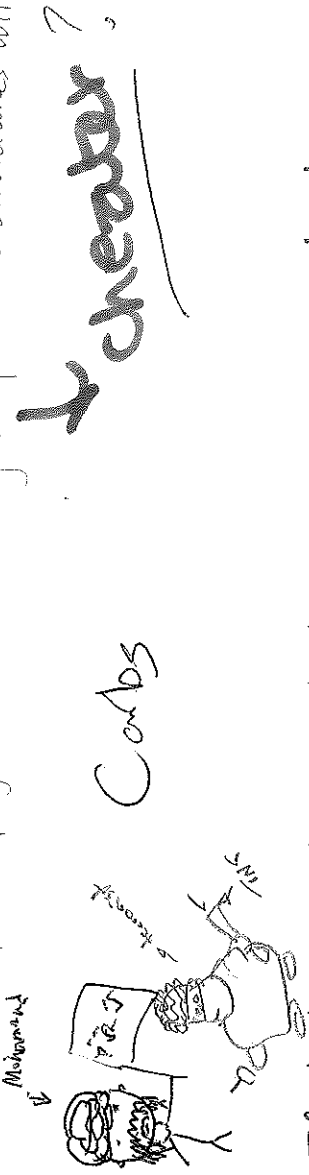
Examples:

Benzene

BVB
The best
Kagawa



• Resonance as a way of depicting a bonding mode that is in between two (or more) extremes ~~to~~ ^{often} displays unusual stability compared to structures without resonance.



If it is an ion, put brackets around the compound/molecule and the charge.

↳ highly stable

Examples
DDT - TNT - styrene

Polarity

- Caused by different electronegativity between two bonding atoms

e^-

Polar — electronegativity

Vector — addition & subtraction

VSEPR \rightarrow shape dependent

NR - Symmetrical P - Asymmetrical
bent or angular

tetrahedral

- Creates dipoles in the molecule

- Polar substances dissolve in polar solvents, ionic bonds dissolve in polar solvents, Non-polar substances only dissolve in non-polar solvents



Molecular polarity: determine the shape of molecule (VSEPR)

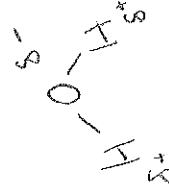
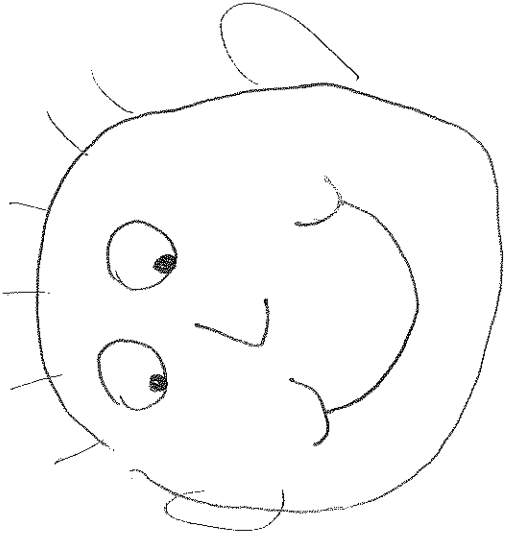
determine polarity of individual bonds, represent this with arrows.

use vector addition to add up polarity and find out if — is polar

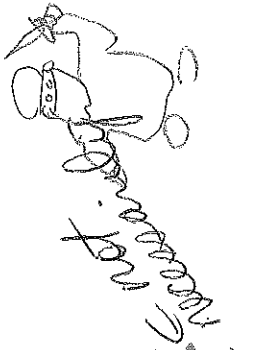
bond polarity: determined by the difference of electronegativity

if difference is greater than 1.7 it's polar, non-polar
Polar bonds \rightarrow caused by difference in electronegativity.

Polar molecules \rightarrow caused by overall charge of molecules. + shape



Hybridization



It is the mix of orbitals to create larger orbitals with properties of the previous orbitals combined.

Linear \rightarrow $sp \rightarrow 50\% s, 50\% p$ 2 e^- domains

Trig-Planar $\rightarrow sp^2 \rightarrow 33\% s, 67\% p$ 3 e^- domains

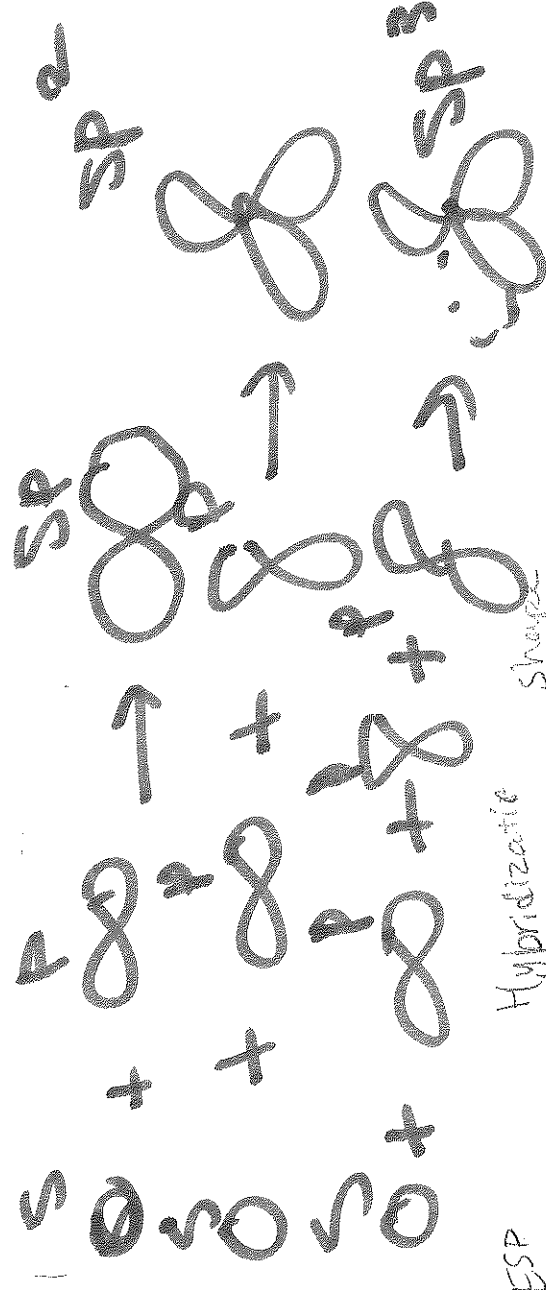
Tetrahedral $\rightarrow sp^3 \rightarrow 75\% s, 25\% p$ 4 e^- domains

Trig-bipyramidal $\rightarrow sp^3d \rightarrow 20\% s, 60\% p, 20\% d$

Octahedral $\rightarrow sp^3d^2 \rightarrow 6 e^-$ domains

deduced from the electron domain geometry rather than the molecular geometry.

Combine VSEPR model with ~~sp~~ orbitals



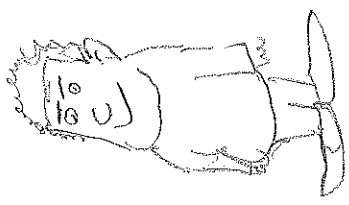
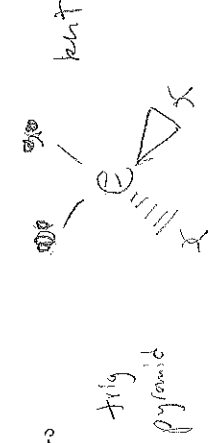
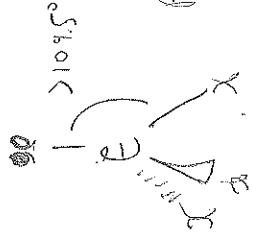
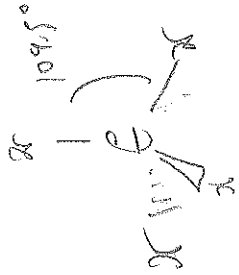
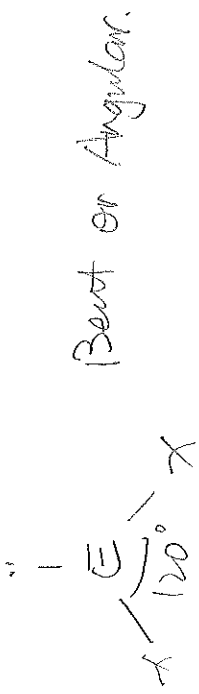
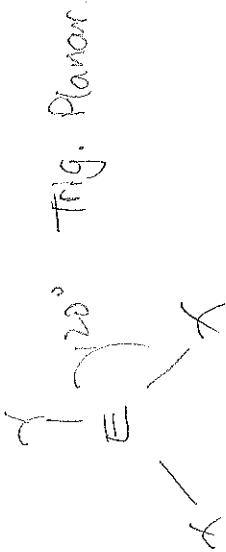
VSEP	Hybridization	Shape
2	sp	Linear
3	sp^2	Trig-planar
4	sp^3	Tetrahedral
5	sp^3d	Trig-bipyramidal
6	sp^3d^2	Octahedral



VSEPR + shapes

- Way that we arrange atoms around the central atom based on their bonding and non-bonding electrons.

- Shapes are determined by the number of bonds and non-bonding electrons. **Electrons arrange themselves in a way that they are furthest away from one another.** lone pairs = polar



e domains	lone pairs	1 lone pair	2 lone pairs	3 lone pairs	4 lone pairs
2	Linear				
3	trigonal planar	bent or angular			
4	tetrahedral	trigonal pyramidal	bent		
5	Trig bipyramidal	seesaw	T-shape	linear	
6	octahedral	Square pyramidal	square planar	T-shape	linear

⇒ lone electrons occupy equatorial positions

~~same~~

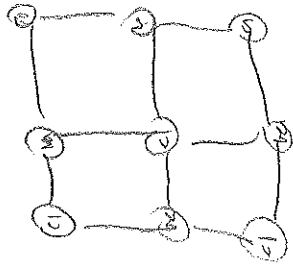
is considered M hybridization

- double & triple bonds are seen as one as single bonds.

Ionic bonds

Metal + ~~nonmetals~~ (nonmetals) ^{occurs if difference} e⁻ negativity is

Strong Bond ABOVE 1.8 (EXP)



Cat ion = +

anion = -

Can ~~not~~ conduct electricity in water and in molten form.

Lattice structure

ex. salt, FeO₃, MgO

Dissociates when dissolved in water \rightarrow or any polar solvent

~~Weaker~~ stronger than covalent bonds

Conduct electricity in liquid state, ^{dissolved} & heat

They don't form molecules. yes.
High boiling and melting points.

Cation + anion.
(+) (metal) (-) (non-metal)

- Polar

\rightarrow dissolve in polar solvents - become individual ions in solvent

electro negativity $\Delta > 1.8$, ^{difference}

high m.p., brittle, do not conduct electro

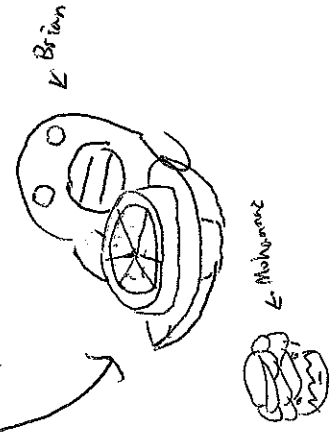
3,1415 9263 399 79 3284

Pi and Sigma bonds

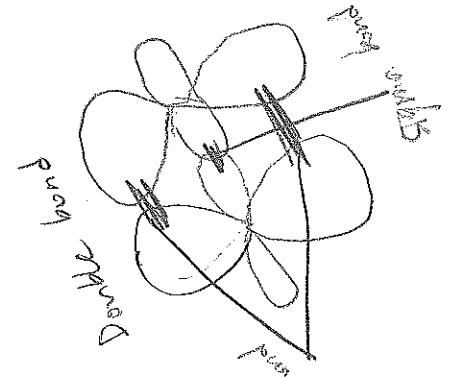
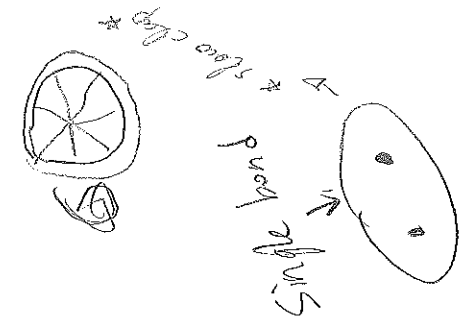
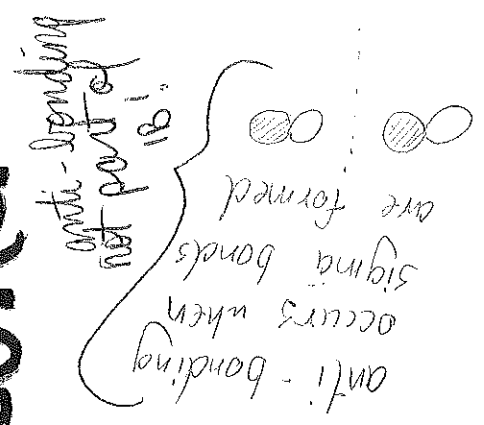
π .

- single bond = 1 sigma bond
- double bond = 1 sigma bond + 1 π bond
- triple bond = 1 sigma bond + 2 π bonds

- Sigma bonds result from head on overlap of two atomic orbitals
- π bonds result from the sideways interaction of two orbitals.



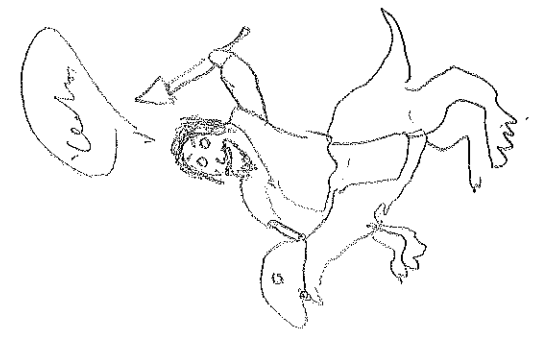
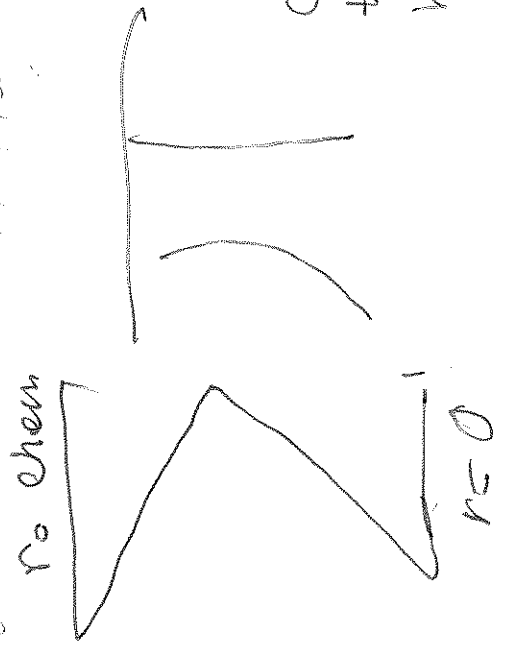
- π bonds cannot be rotated **on axis**



Bonds - higher e^- (probability) density

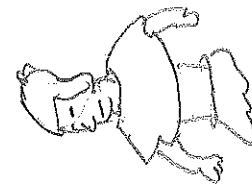
- first bond is always σ and the rest is π bonds. σ is stronger than π bonds. - because of π bonds we have isomers.

# of electron domains	orbitals involved
2.	s, 1p
3.	s, 2p
4.	s, 3p
5.	s, 3p, 1d
6.	s, 3p, 2d



Called " π " (π) bond because π is the greek letter of π and its the hybridization of the p orbitals.

Intermolecular



Forces

Forces in between molecules

— sometimes can't have stronger dipole dipole

Dispersion/London Forces - temporary dipoles

in all compounds

Dipole-Dipole - permanent dipoles

Must be polar

Hydrogen bonds. - between H and N, O, or F

- strongest

- magnified dipole-dipole

Strongest IMFs - higher boiling points
↳ more energy required to decrease

IMFs

IMFs - high surface tension

- Weak IMFs have low volatility

- London Force: generally weak.

: (caused by movement of electrons within the covalent bonds that without permanent dipole. - London forces always present



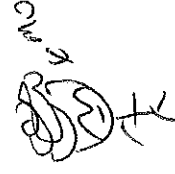
London forces

named after a

man with last name

London.

John Helber



1.

- occurred due to the Fayer Paradox (existence of r^2 & r^6 bond but unpredictable)

② 2-1 ②

Metallic Bonding



② ionic (+)

Mo! Choi!

Alloys - stronger because different sized metal ions
↳ doesn't compact as easily

Metal + Metal

corrosion resistant - different corrosion rates of metals
Strong Bond (+) ②

② delocalized electrons - conducts electricity & heat
b/c e⁻ can move around the charge/energy

- malleable
- ductile
- lustrous

shiny?

↳ because of their lattice structure ②



high boiling points -

- Free e⁻
- positive nucleus

The following has been

② in metals (pure + alloys)

general fairly strong

good generate electricity and heat

② metals in a sea of delocalized electrons. (+)

The strength of metallic bond depends on

- number of valence electrons
- charge of metal ion
- ionic radius of the cation

(+)

② information

②

②

②