

Atomic Structure

Section 7.6

Quantum numbers: Introduction

Distribution of e^- , energy, pathways

Classical mechanics couldn't explain it

- expecting an attraction between proton in nucleus & electron, atom would collapse

Max discovery

- Quantas: smallest packet of energy. ← Quantized approach
- has to be whole number quantity

ex. The \$ in financial transactions

Problem of the atomic structure

Distinction between waves and particles

/
massless

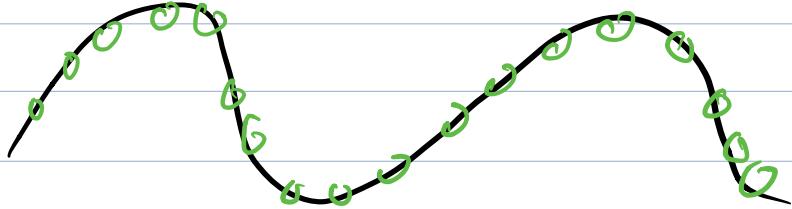
/
mass & position

&
no defn position
(continuous)

Einstein: Duality between them

waves behave like particles & particles behave like waves.





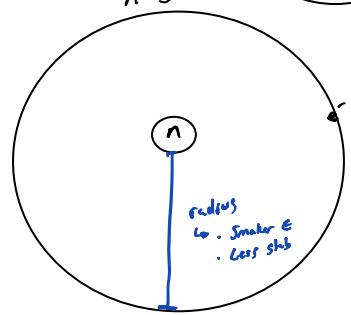
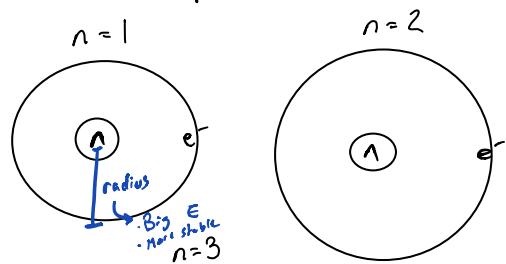
GroSkin: Waves: composed of particles

Definitions

#s that define an orbital:

(1) Principal quantum number, n :

Defines the size and energy of an orbital. It takes values of positive integers ($1, 2, 3, \dots$)



Different sizes

one is positive
 one is negative

$$E = \frac{q_1 \cdot q_2}{r}$$

Charges involved
distance between them

Energy of attraction carries a negative sign

(Energy: capacity to perform work
(physics))

Need an input of energy to separate the attracted electrons

② Azimuthal Quantum number, l :

Defines the shape of an orbital. It takes values from zero to $n-1$

③ Magnetic Quantum number, m_l :

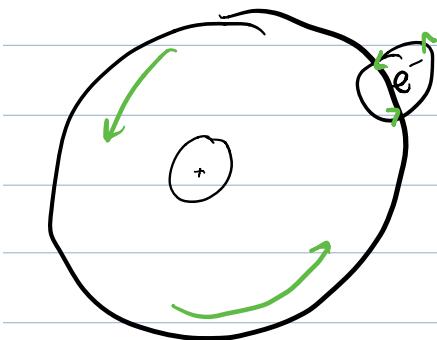
Defines the orientation of an orbital relative to other orbitals in the atom. It takes values from $-l$ to $+l$ (including zero)

Quantum Numbers for the First Four Levels in H-Atom

(n)	(l)	(m_l)	
Principal Quantum No.	Azimuthal l-No.	Orbital Descrip'n (subshells)	Magnetic l-No.
1	0 (s) Sphere = no orientation	1s	0 Ground State
2	0 (s) 1 (p)	2s 2p	0 -1, 0, 1
3	0 (s) 1 (p) 2 (d)	3s 3p 3d	0 -1, 0, 1 -2, -1, 0, 1, 2
4	0 (s) 1 (p) 2 (d) 3 (f)	4s 4p 4d 4f	0 -1, 0, 1 -2, -1, 0, 1, 2 -3, -2, -1, 0, 1, 2, 3

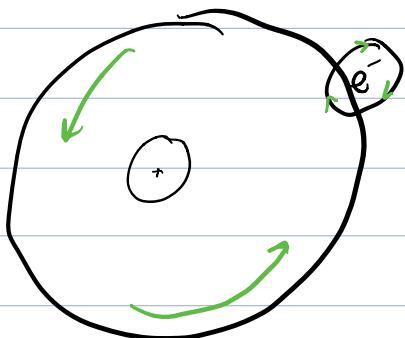
- distance between e^- & nucleus
= more attraction
= least energy [?]

④ Spin Quantum Number, m_s - unrelated to orbital, just a property of an e^-



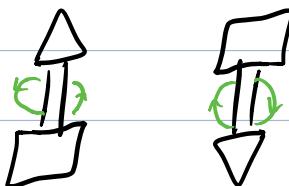
Same direction

OR



Opposite directions

rotating electrons = magnetic field



$$m_s = +\frac{1}{2}$$

$$m_s = -\frac{1}{2}$$

Pauli's Exclusion Principle

No two electrons in the same atom can have four quantum numbers the same.

- i.e. • no more than $2e^-$ in one orbital
- the e^- will have to have opposite spins

Stability: max. attraction vs repulsion

- If we keep adding e^- , the repulsion > attraction
- λ repulsion between 2 magnetic fields
if not opposite

ex. Give a set of 4 quantum no. for each of the e^- in the 3d subshells.

S orbitals	n	l	m_l	m_s	
3	3	2	-2	+1/2 -1/2	10 e^-
	3	2	-1	+1/2 -1/2	
	3	2	0	+1/2 -1/2	
	3	2	1	+1/2 -1/2	
	3	2	2	+1/2 -1/2	

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$n = 2$ SHELL

$2s$ } SUBSHELLS
 $2p$

$2s$
 $2p_x, 2p_y, 2p_z$ } Individual shells

Ex. What is the total number of e^- in an atom that may have the following quantum numbers:

a) $n = 3$

$\ell = 1 (p)$

$m_\ell = -1, 0, 1$

$3p$

$3p_x, 3p_y, 3p_z$

$\ell = 0 (s)$
 $\ell = 1 (p)$
 $\ell = 2 (d)$
 $\ell = 3 (f)$

$3 \text{ OR } 3 \times 2 e^- = 6 e^-$

b) $n = 4$ $\ell = 1$

$4p$

$m_\ell = -1, 0, 1$

$3 \text{ OR } 3 \times 2 e^- = 6 e^-$

c) $n = 3$
 $\ell = 2 (d)$

$m_\ell = -2$

$2e^-$

d) $n=3$

$m_s = +\frac{1}{2}$	ℓ	m_ℓ	<u>ORBITALS</u>
	0 (s)	0	1
	1 (p)	-1, 0, 1	3
	2 (d)	-2, -1, 0, 1, 2	5
			$\frac{5}{9}$ orbitals $\times 10^5$ $= 9 \times 10^5$

ex Which of the following orbital designations are incorrect?

~~1s, 1p, 7d, 9s, 3f, 4f, 2d~~

~~$n=2$~~ $n=7$ ~~$n=3$~~ $n=2$
 ~~$\ell=1$~~ $\ell=2$ ~~$\ell=3$~~ $\ell=2$

BOHR'S MODEL

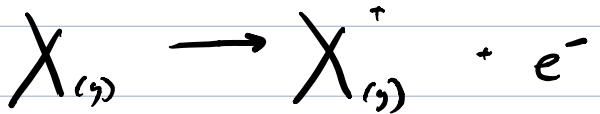
only works
for H⁺

$$E = -2.178 \times 10^{-18} \left(\frac{Z}{n} \right)^2$$

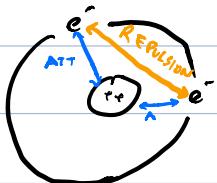
charge of the nucleus
positive integer

IONIZATION ENERGY

Energy required to remove an electron from an atom in the gaseous state:



POLYELECTRONIC ATOMS



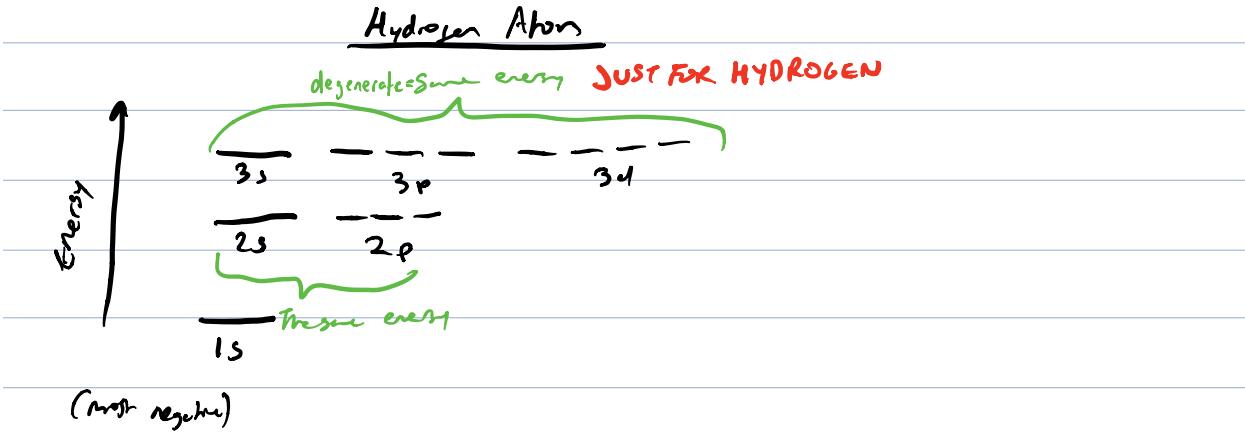
$$E_{i(1\text{st})} = 2372 \text{ kJ/mol} \quad (\text{removal of the first } e^-)$$

IONIZATION ENERGY

$$E_{i(2\text{nd})} = 5248 \text{ kJ/mol} \quad (\text{removal of the 2nd } e^-)$$

$$Z_{\text{effective}} = Z_{\text{actual}} - (e^- \cdot e^-) \text{ Repulsions}$$

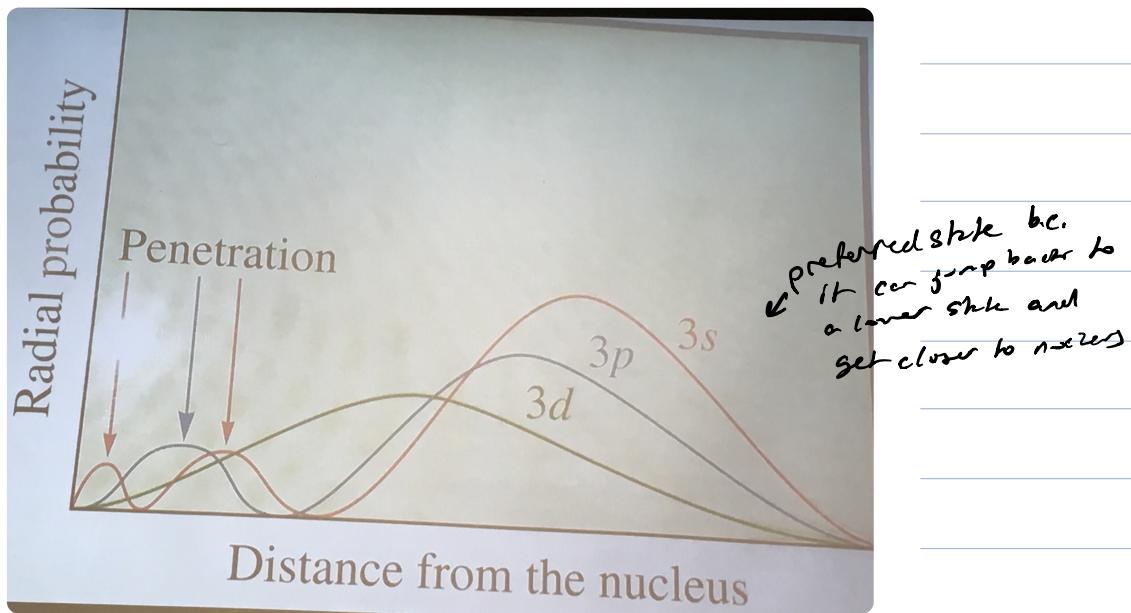
$$E = -2.178 \times 10^{-18} \left(\frac{Z_{\text{eff}}}{n} \right)^2$$



DEGENERATE ORBITALS (orbitals with same energy)

HYDROGEN: orbitals with same value for the principal quantum no.

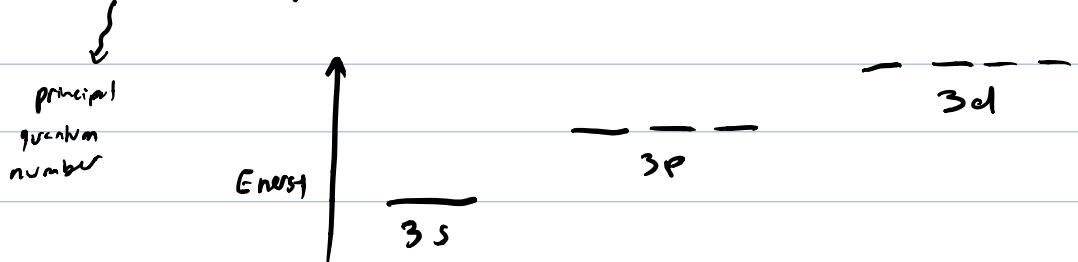
POLYELECTRONICS: orbitals with the same principal and azimuthal quantum number



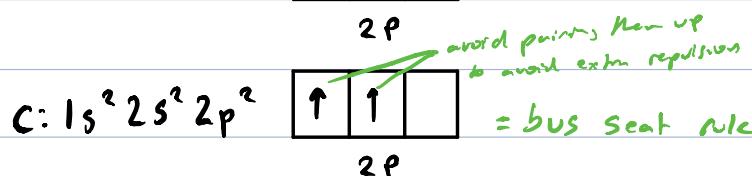
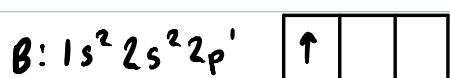
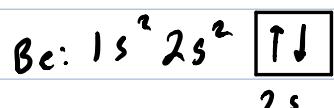
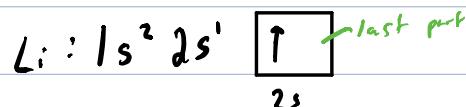
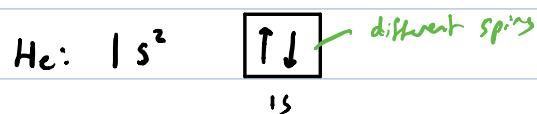
ORDER OF FILLING OF ORBITS

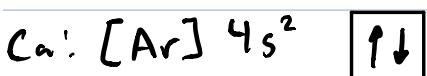
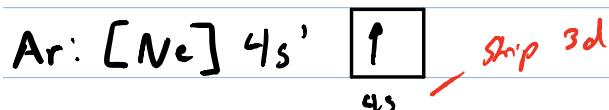
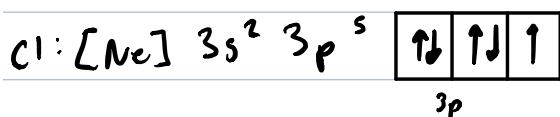
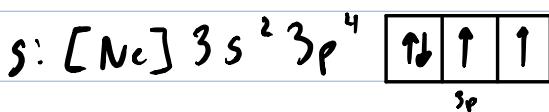
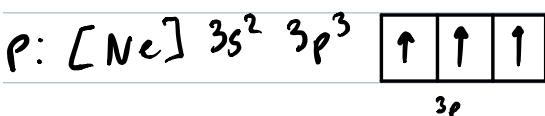
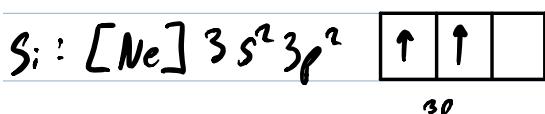
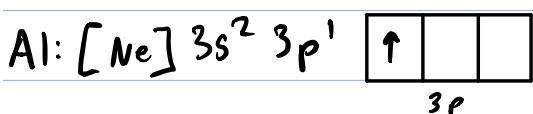
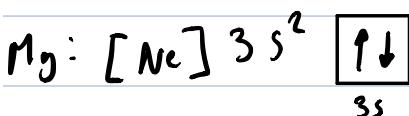
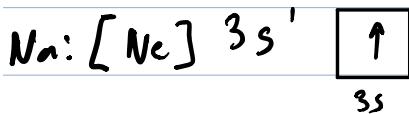
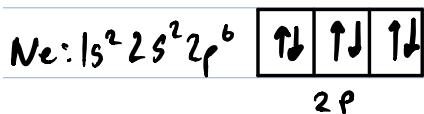
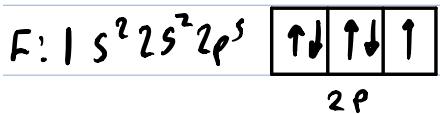
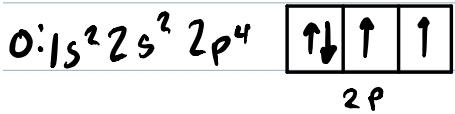
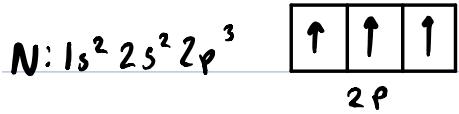
(or order of stability of orbitals)

$$ns > np > nd$$

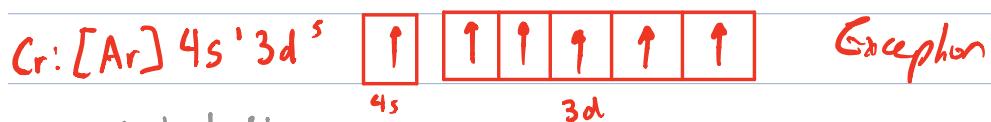
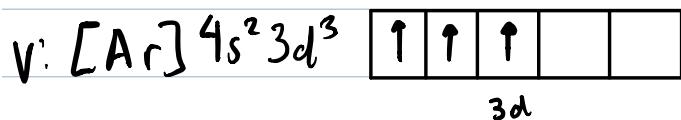
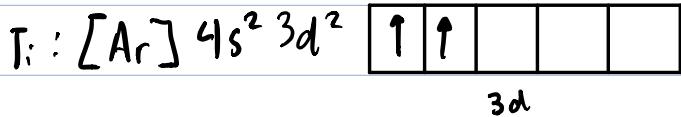
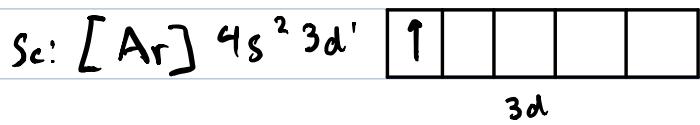


AUFBAU PRINCIPE

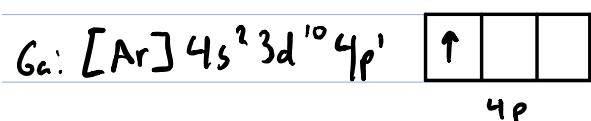
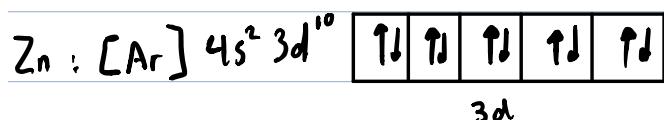
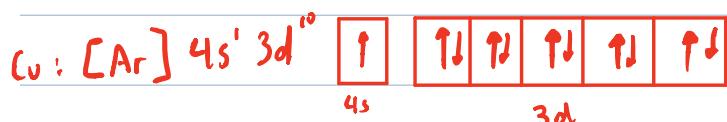
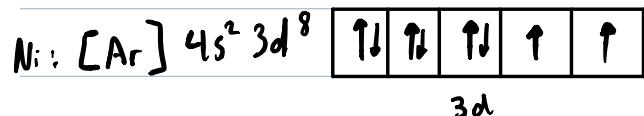
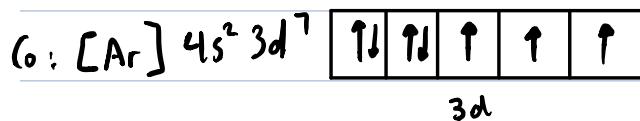
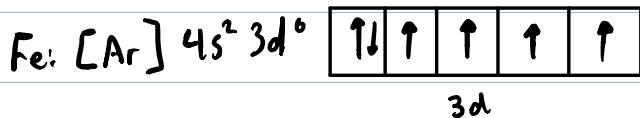
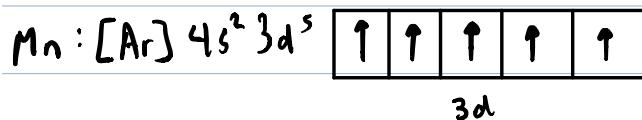
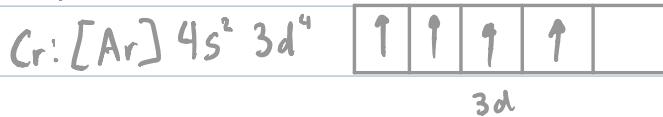


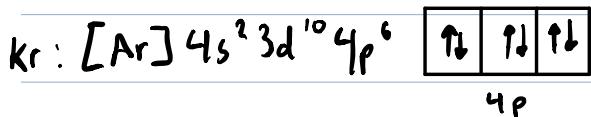
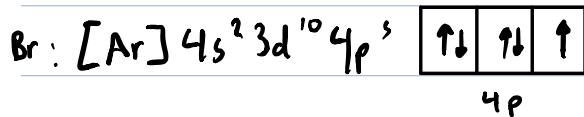
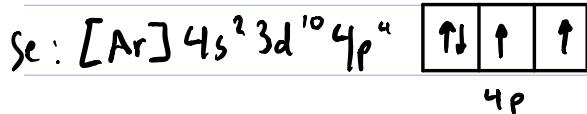
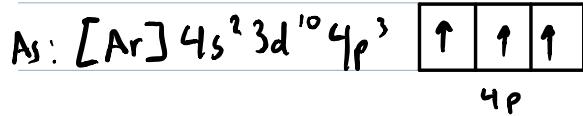
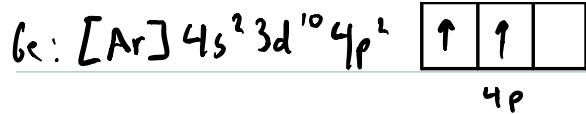


4s



Instead of:





$4f^{1 \rightarrow 14}$ lanthanides

$5f^{1 \rightarrow 14}$ actinides

VALENCE ELECTRONS

= Electrons in the outermost shell (highest value for the principal quantum number).

Exercises

ex. Give the electron configuration for Sulphur (S), Cadmium (Cd), Hafnium (Hf) and Radium (Ra)



ex. Identify the following elements:

a) An excited state of this element has the electron configurations:

excited $1s^2 2s^2 2p^5 3s^1 = ^{+e^- : 2+2+5+1=10}$
 $\hookrightarrow \text{Atomic no.} = 10 \therefore \text{Ne } [\text{Ne}]$

ground $1s^2 2s^2 2p^6$

b) An excited state of the element has the electron configurations:

exc. $[\text{Kr}] 5s^2 4d^6 5p^2 6s^1$ (A_1)

ground: $[\text{Kr}] 5s^2 4d^5$

c) The ground state ... is:

$[\text{Ne}] 3s^2 3p^4 - 6 \text{ valence } e^- = \text{Group } (\text{S})$

d) The ground state ... contains 3 unpaired $6p$ e^-

Two



Cast to one \Rightarrow the last to go

ex. Cl: $[\text{Ne}] 3s^2 3p^5$

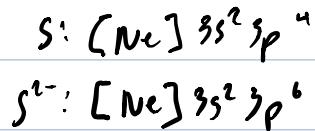
Ca: $[\text{Ar}] 4s^2$

Al: $[\text{Ne}] 3s^2 3p^1$

Cr: $[\text{Ne}] 3s^2 3p^6$

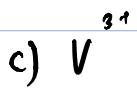
Ca²⁺: $[\text{Ar}]$

Al³⁺: $[\text{Ne}] 3s^2$



Transition elements: "S" e⁻ are the first to be removed

ex. Which of the following ions has five unpaired e⁻?



*No unpaired e⁻



$$\# \text{ of } e^- = 2n^2$$

\uparrow \nwarrow not orbitals

2e⁻ / orbitals

PERIODIC TRENDS

October 7, 2016

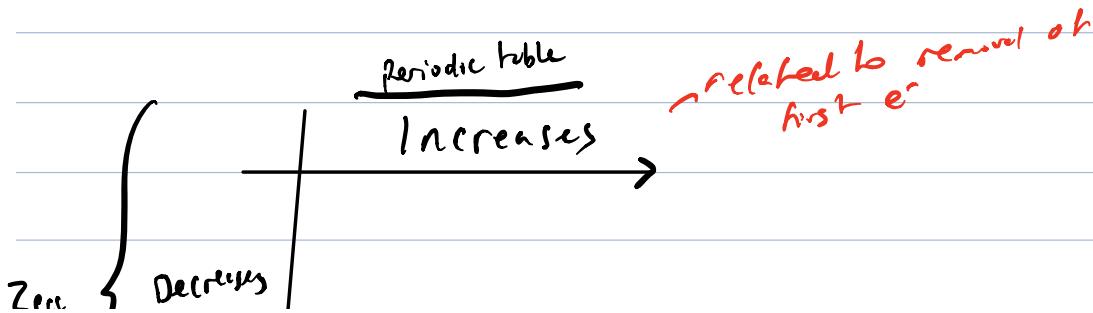
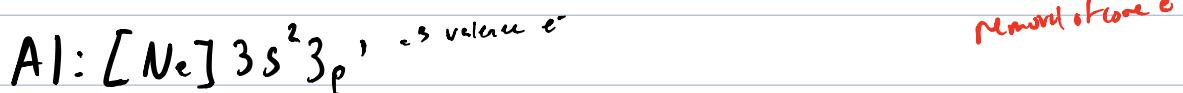
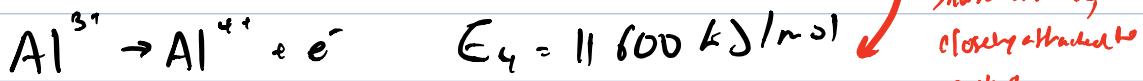
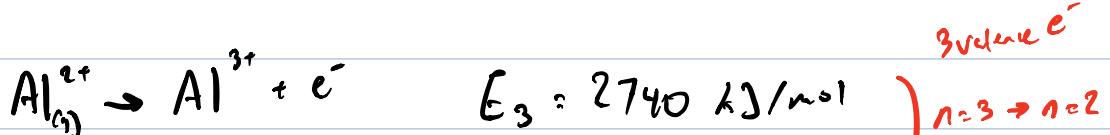
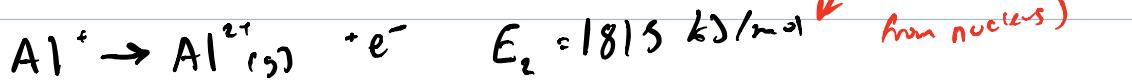
$$Z = \# \text{ of } p^+$$

$$Z_{\text{eff}} = Z + (e^- - e^+) \text{ Repulsions}$$

Z_{eff} goes slightly down (due to shielding effect) when a new orbital appears

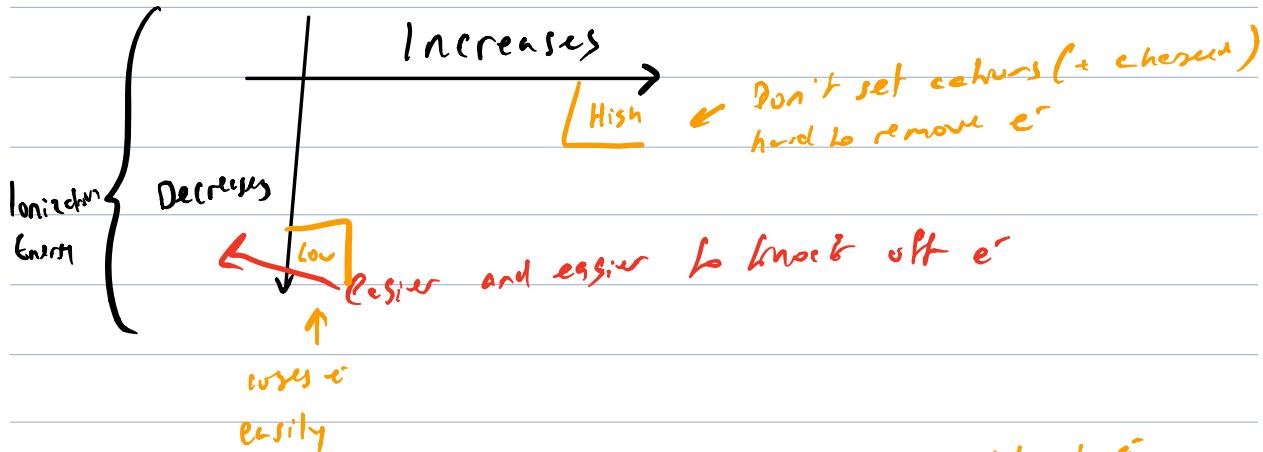
New period: - Attraction force is inversely proportional to the square of the distance
• New period = further apart
• Z_{eff} increases along period

Ionization energy



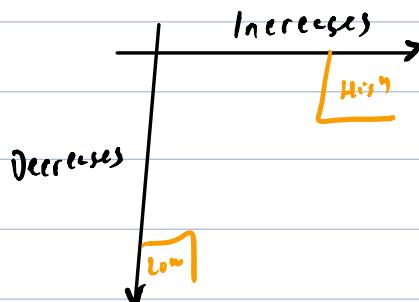
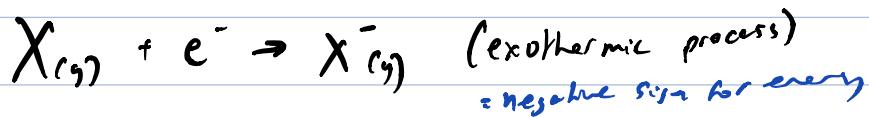
(↓)

↑ Similar trend



Electron Affinity - affinity for additional e^-

Energy associated with accepting additional electrons.



Atomic Radius



Increases

Decreases

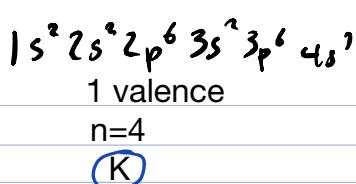
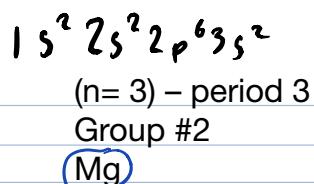
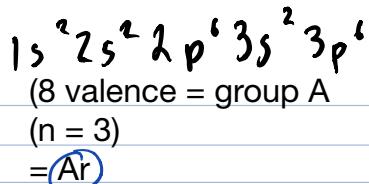
weaker attraction = further

ex. The first ionization energy for Phosphorus is 1060 kJ/mol
and for sulfur it is 1005 kJ/mol. Why? Explain.

Goes against the periodic trend of Zeff increasing across period.

From P to S, you are adding new electrons in the same orbital, so slight loss of e-

Ex. Three elements have the following electron configurations:



Their ionization energies (not in the same order) are: 0.4189, 0.7377, 1.5205.



Their atomic radii (not in the same order) are: 1.60 Å, 0.94 Å, 1.97 Å.

Identify the elements and match the appropriate values for E. (ionization energy) and r.

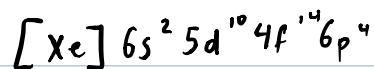
For radius

$$\text{Å} = 1.0 \times 10^{-8} \text{ cm}$$

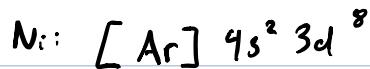
Angstrom

ex. Give the ground state electron configurations for the following:

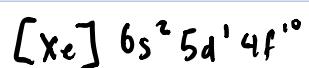
a) Po



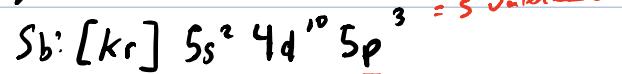
b) N_i^{3+}



c) Ho Lanthanide



d) Sb²⁺



Ex. Give the values of the four quantum numbers for each of the specified e^- :

a) $n=5$, Sb



n	ℓ	m_ℓ	m_s	
5	0	0	+1/2	
5	0	0	-1/2	
5	1	-1	+1/2	
5	1	0	-1/2	parallel same spin
5	1	1	-1/2	

b) $\ell=2$, Zr



(all 3d and 4d e^-)

n	ℓ	m_ℓ	m_s
3d	3	2	-2 $+\frac{1}{2}$ -1/2 $+\frac{1}{2}$
	3	2	-1 -1/2
	3	2	0 -1/2
	3	2	1 -1/2
	3	2	2 -1/2
4d	4	2	-2 $+\frac{1}{2}$
	4	2	-1 $+\frac{1}{2}$

could have been -2, -1, 0, 1, 2, but have to be different

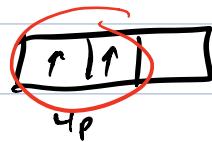
ex. Heaviest Halogen without d electrons: Cl

↑
Group 7

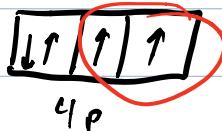
ex. Heaviest element with half filled p orbital: Bi

2 are half filled

ex. 2 elements with 2 half filled 4p orbitals



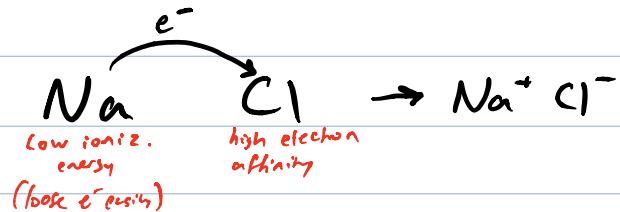
Sn & Te



October 11, 2016

BONDING ENERGY

Ionic Bond:

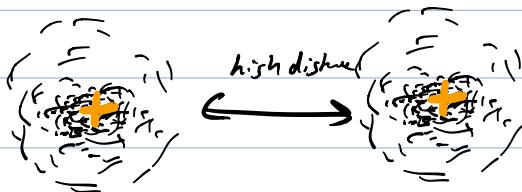


$$E = 2.31 \times 10^{-19} \frac{q_1 q_2}{r} \text{ charges of ions}$$

Not repulsive
for eq.

$$\text{Ex NaCl: } E_{\text{NaCl}} = 2.31 \times 10^{-19} \frac{(+1)(-1)}{0.276 \text{ nm}} = - 8.37 \times 10^{-9} \text{ J}$$

Covalent Bond:



Can't be too close: due to repulsions between e⁻ & between p⁺

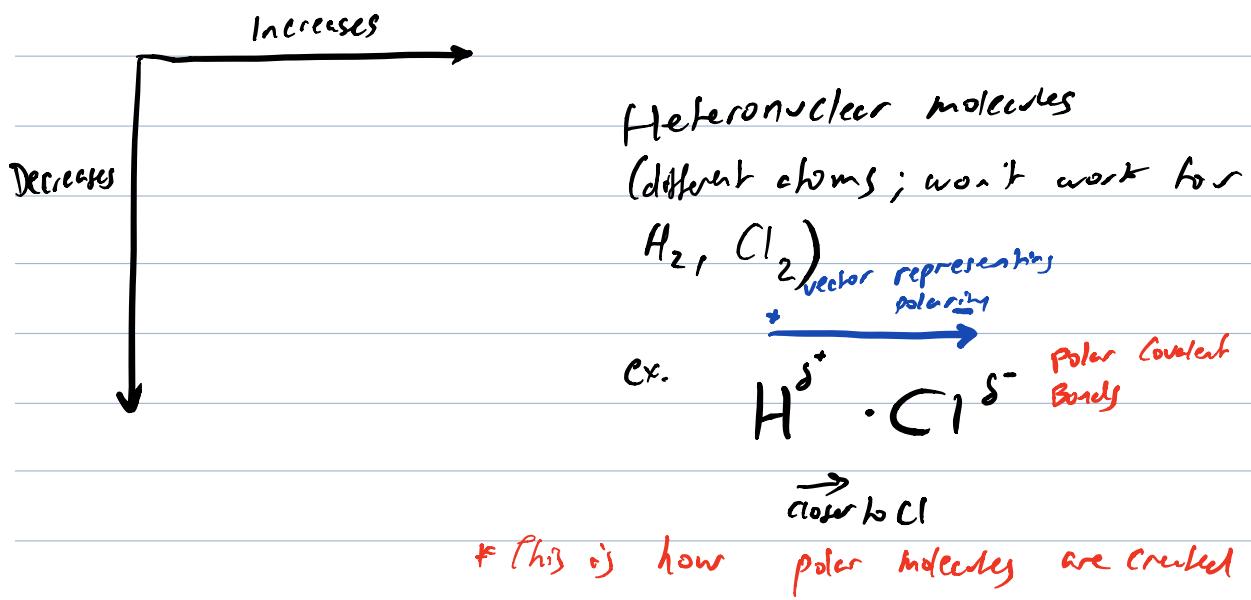
It reaches an optimal distance and formulates to get lowest energy (stronger bonds)

Electronegativity

The ability to attract electrons in a covalent bond

Very similar to electron affinity

Trend will be similar to electron affinity



electron affinity vs. electronegativity



doesn't take over the other e^- ,
just pulls it closer; shift part of
the other atom's shell

- Can assign relative electronegativity values between different elements in the periodic table

ex.



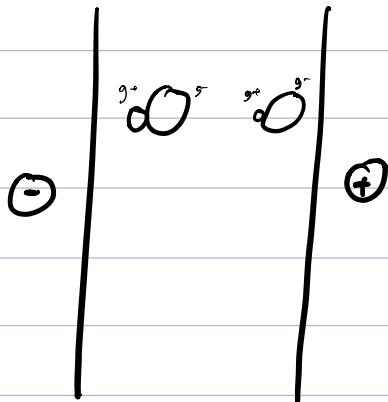
dipole = dipolar = polar



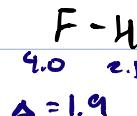
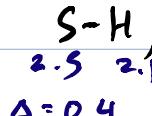
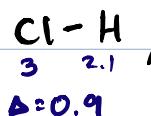
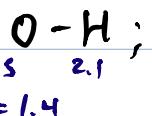
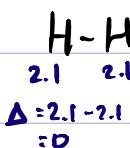
This is a dipolar molecule

If you place a polar molecule in an electric field

HF molecule

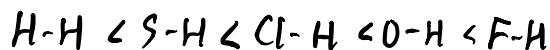


ex. Order the following bonds according to polarity

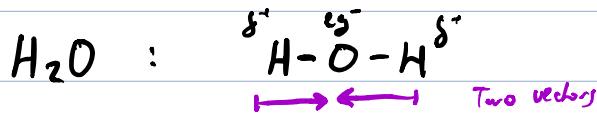


Not polar

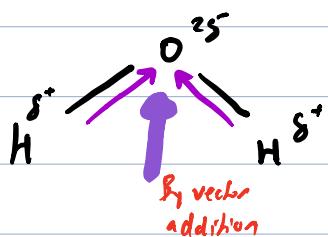
These δ s won't be given;
will have to refer to periodic
table



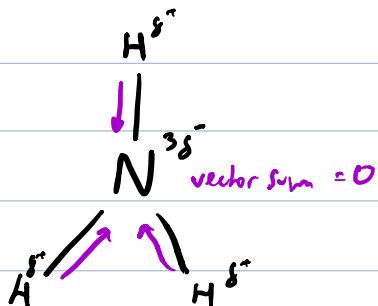
Polarity and Polyatomic Molecules



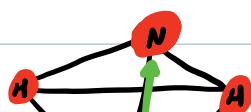
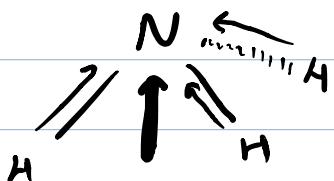
Not polar, but if you look at the geometry:

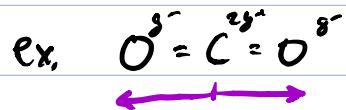
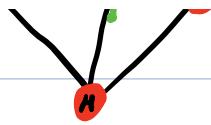


ex. NH_3 - Ammonia

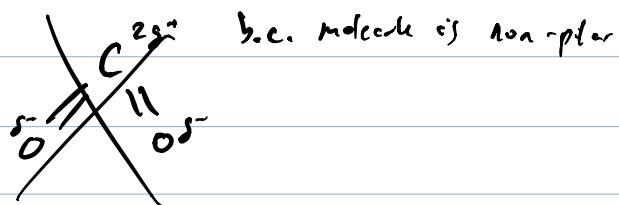


Achiral geometry:



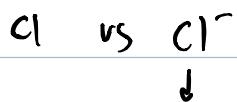


Alkenes geometry:

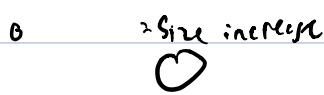


Ion Size

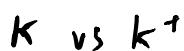
Atomic Radius



Weaker attraction



Only for ions
in the same
group



ISOELECTRONIC GROUPS:

Series of a couple of ions that have the same no. of e⁻

e) Arrange the ions Sc^{2+} , Sr^{2+} , Br^- , Rb^+ in order of decreasing size



ex) Arrange the ions Cs^+ , I^- , Te^{2-} , $(\text{Ca}^{2+}, \text{P}_\text{O}^{2-})$, Be^{2+}
 hot part of it



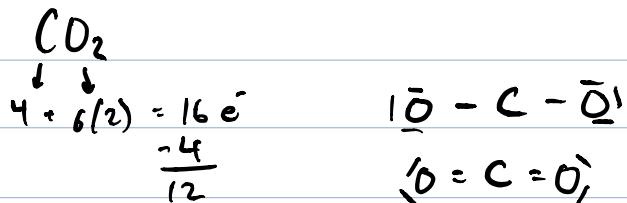
LOCALIZED ELECTRON BONDING MODEL



LEWIS STRUCTURES

Br²⁺: How many valence δ^- ?

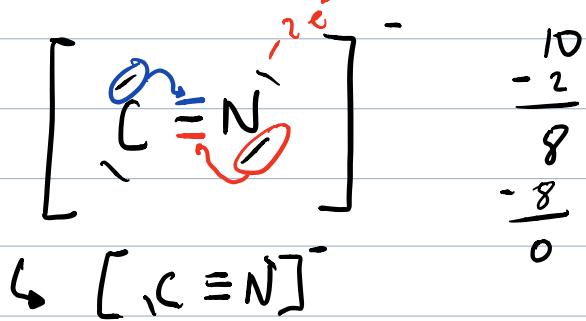
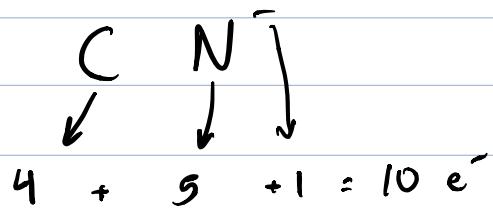
- 1) Add up the valence δ^- of all the atoms involved
- 2) form bonds using two δ^- for each
- 3) Distribute the remaining δ^- so that each atom achieves the noble gas configuration



Ions:

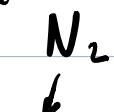
October 19, 2016

ex.

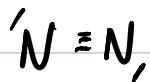


Lewis Structure
 Based on the assumption that e^- aren't moving
 Just a model
 (in reality - they move)

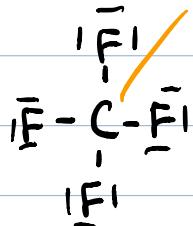
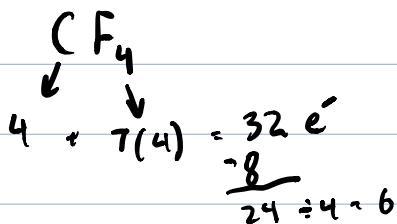
ex.



$$5(2) = 10 e^- \text{ (same as previous ex.)}$$



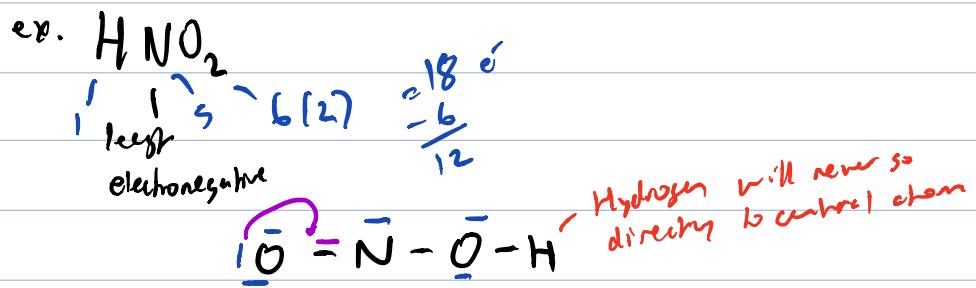
ex.



Central atom is the least electronegative (except hydrogen)
 can only bond to share greatest no. e^-
 Geometry is irrelevant

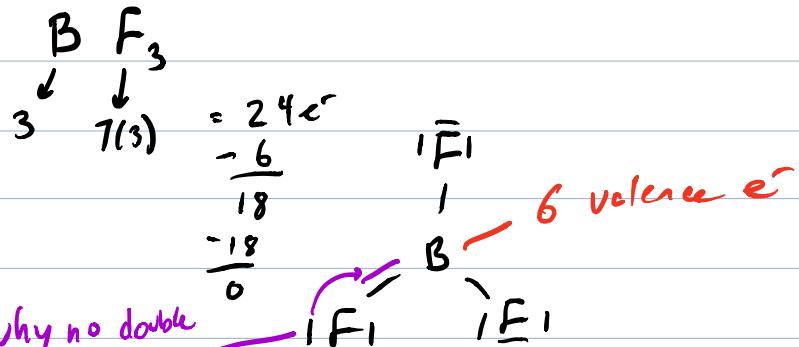
If it's highly electronegative,
 it will be selfish & want to share
 e^-

To Pick central Atom, pick atom most electronegative



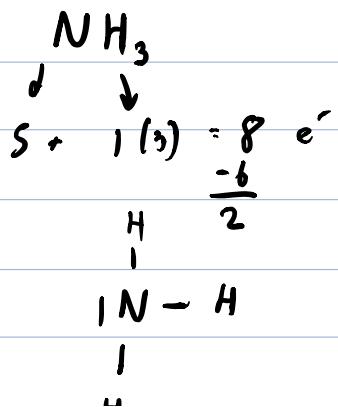
EXCEPTIONS TO THE OCTATE RULE

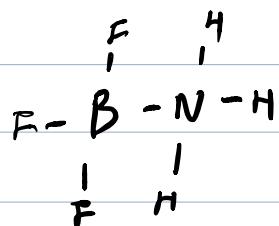
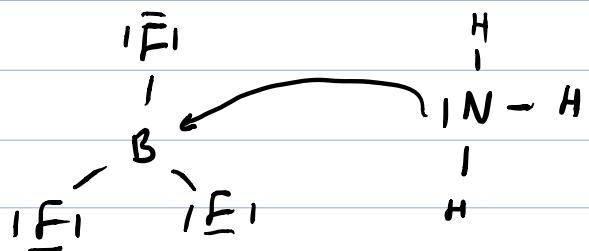
Boron:



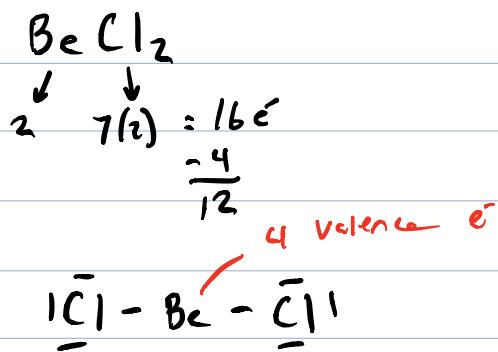
Why no double bond? F is very electronegative, doesn't want to share

Boron can be relatively stable with 6 valence e^- , but if given the option, it will go for the extra e^-

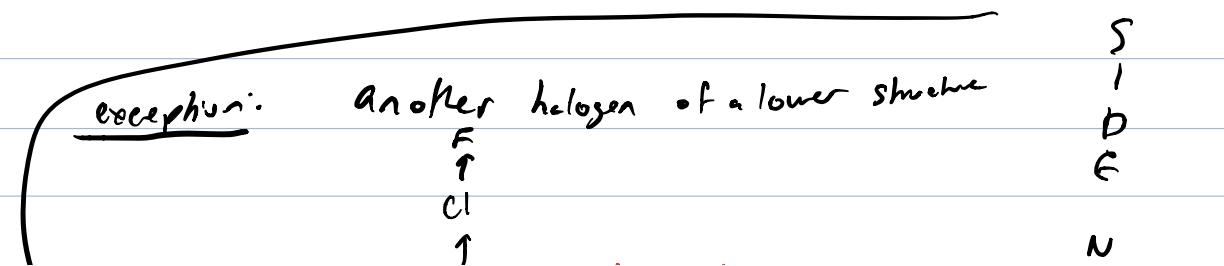




Boronium:



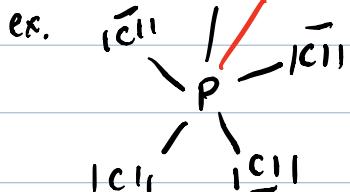
So 2 exceptions: Be & B



Br - central atom

$\frac{0}{T}$
 E

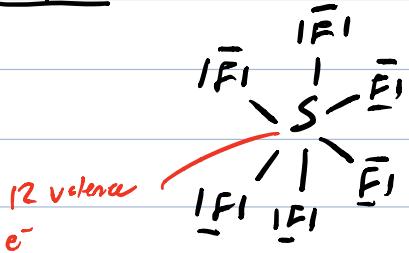
ex. PCl_5 above $8e^-$



exception: phosphorus

$$\begin{array}{c} \text{PCl}_5 \\ \downarrow \quad \downarrow \\ \text{S} \quad 7(5) = 40 e^- \\ \frac{-10}{30} \\ \frac{-30}{0} \end{array}$$

exception: sulfur



$$\begin{array}{c} \text{SF}_2 \\ \downarrow \quad \downarrow \\ 6 \quad 7(6) = 48 \\ \frac{-12}{36} \\ \frac{-36}{6} \end{array}$$

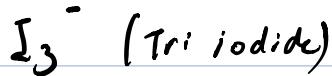
P & S both in 3rd period

extra $e^- \rightarrow$ into d orbitals

* won't work for 2nd period because no d orbitals available

+ elements in periods 3 or lower can accommodate more

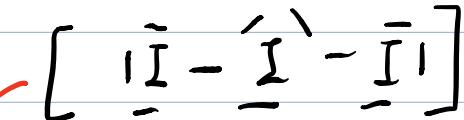
exception:



$$\downarrow \\ \text{I}(s) + 1 = 22 e^-$$

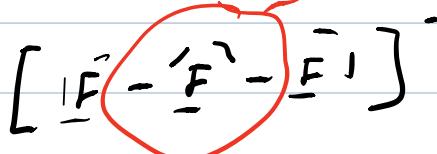
$$\begin{array}{r} -4 \\ -18 \\ \hline -22 \end{array}$$

Need
"E" for
ions



$$\downarrow \\ \text{F}(s) + 1 = 22$$

Doesn't exist b.c. F is in the 2nd period



* All examples were even numbers *

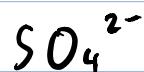
ODD ELECTRON SPECIES (FREE RADICALS)



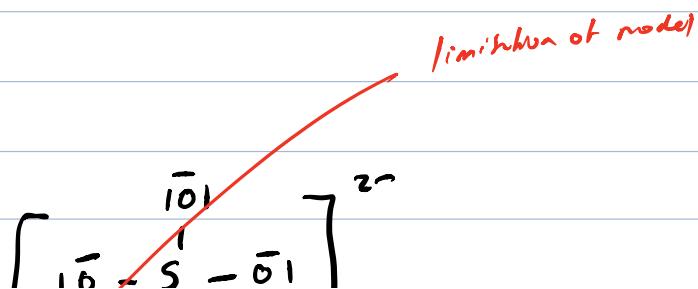
$$\begin{array}{r} s + 6 = 11 e^- \\ -2 \\ \hline 9 \end{array}$$



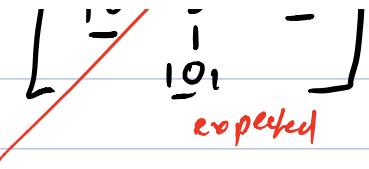
- involved in cancers, aging
- they are very reactive



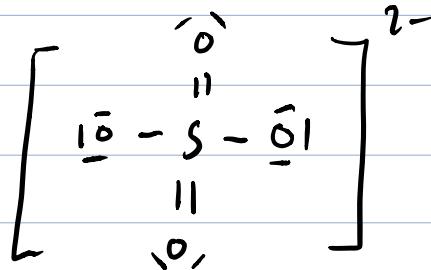
$$\begin{array}{r} 6 + 6(4) = 2 = 32 e^- \\ -8 \\ \hline 24 \end{array}$$



$$\frac{-24}{6}$$



expected

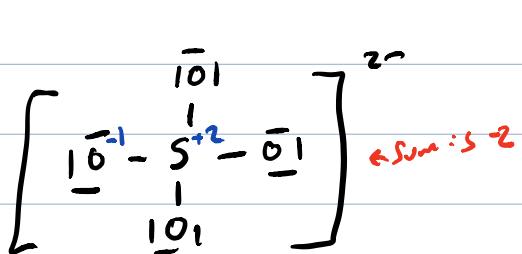


OBSERVED

Formal Charge Concept

$$= \left(\begin{array}{l} \text{# of valence e}^- \text{ on a free, neutral} \\ \text{atom (group #)} \end{array} \right) - \left(\begin{array}{l} \text{# of valence e}^- \text{ assigned to} \\ \text{the atom in the molecule or} \\ \text{ion} \end{array} \right)$$

↓ ↓
BONDING PAIRS: LONE PAIRS:
 Count with $\frac{1}{2}$ Count fully

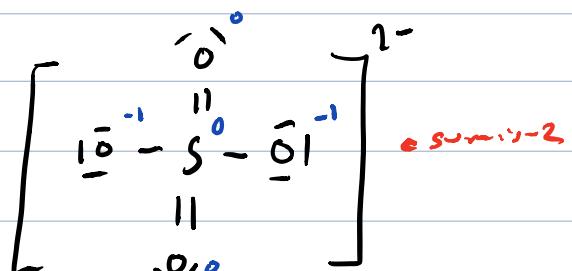


expected

$$\text{S: } 6 - 4 = +2$$

$$\text{O: } 6 - 7 = -1$$

Group



observed

$$\text{S: } 6 - 6 = 0$$

$$\underline{1} \quad \text{O: } 6 - 7 = -1$$

$$\underline{2} \quad \text{O: } 6 - 6 = 0$$

First Rule: Sum of all charges in ion
is equal to the charge of the ion
• In neutral \rightarrow sum is 0

BALANCING FORMAL ATOMS AS CLOSE TO 0

(Can only be done if central atom is 3rd + period)

/
b.c. it's acquiring
more than 8

Other rule: If there's an excessive negative charge should only be based on the most electronegative atom

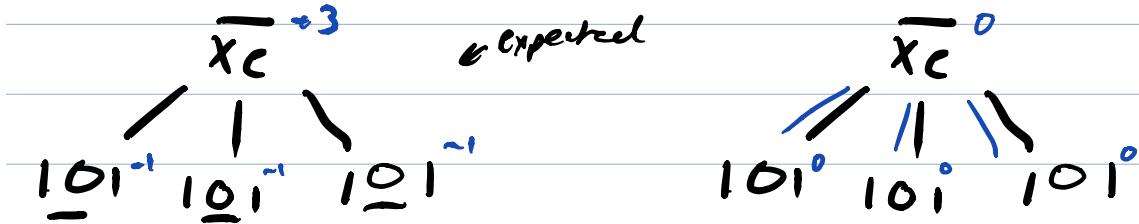
Should do this process for 3rd period & beyond



exists even though it's a nozzle

$$\begin{array}{r} \text{XeO}_3 \\ \downarrow \quad \downarrow \\ 8 + 6(3) = 26e^- \\ - \frac{6}{20} \\ - 18 \end{array}$$

$\frac{1}{2}$



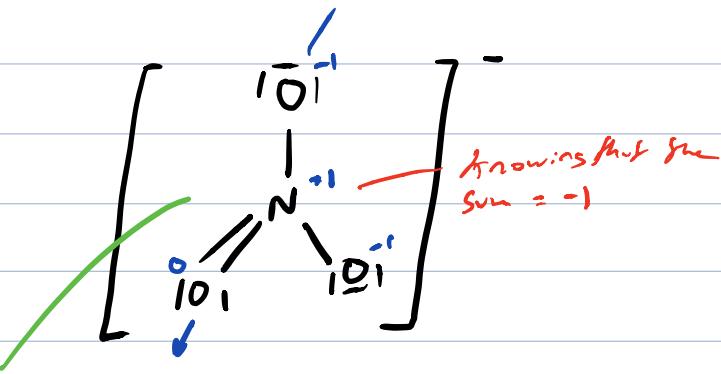
$$Xe: 8 - 5 = 3$$

$e.g. $NO_3^-$$

$$5 + 6(3) + 1 = 24 e^-$$

$$\begin{array}{r} 6 \\ \hline 18 \\ -18 \\ \hline 0 \end{array}$$

$$6-7 = -1$$

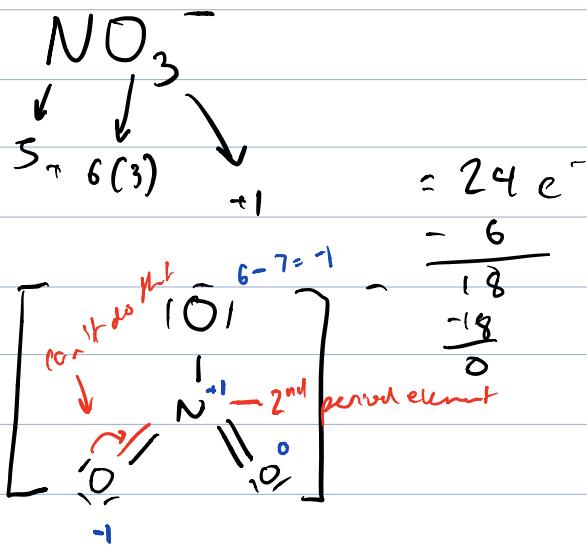


2nd period element,
can't do anything

Most common bonding schemes will all equal to 0 ← most stable



October 21, 2016



What if we move the double bond to another O?

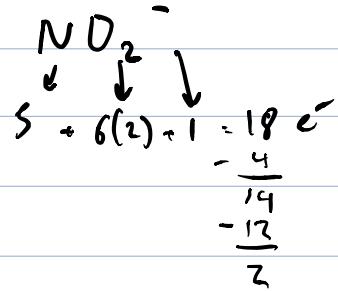
Still correct via Lewis

AL & AREN'T CORRECT EXPERIMENTALLY

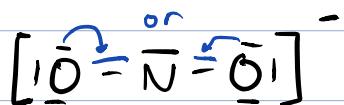
The bonds we see experimentally are a single/double bond hybrid ($\frac{2}{3}$ single, $\frac{1}{3}$ double)

All 3 theoretical structures are called:
Resonance Structure

Ex.



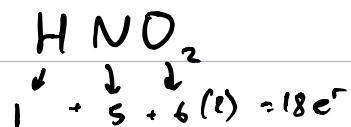
*Refer to notes in
textbook



Resonance hybrid

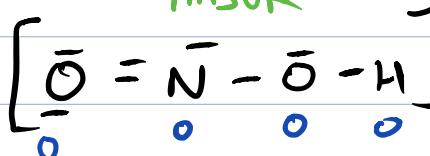


Ex.

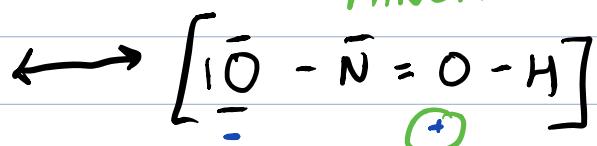


Not equivalent resonance, but still resonance

MAJOR

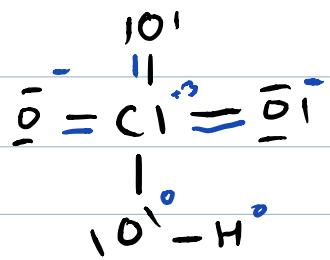
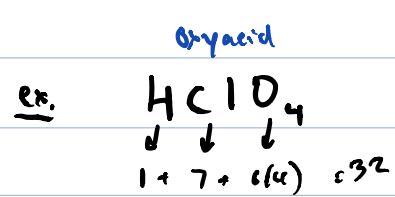


MINOR

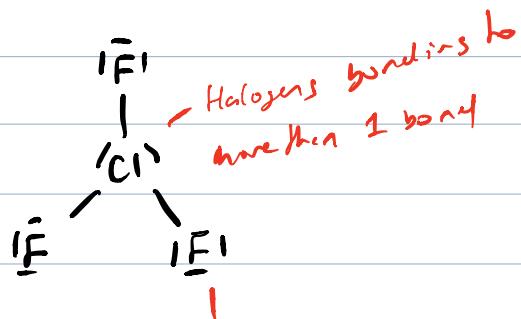
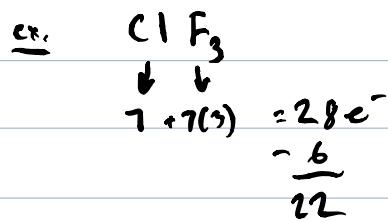


(+)

Excess of negative charge = most electronegative



Exception of Halogen forming more than 1 bond



higher up in the orbital

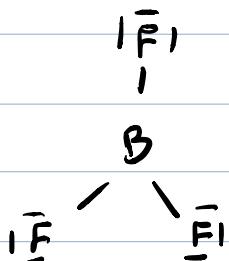
VSEPR - MODEL



Linear structure



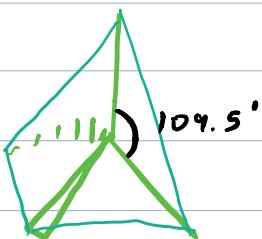
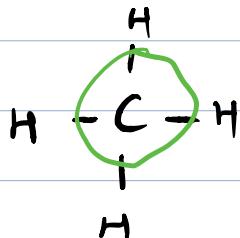
180°



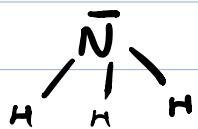
120°



Trigonal Planar

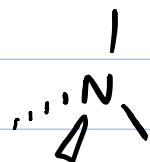


Tetrahedral



Arrangement of σ pairs

Molecular Geometry



Tetrahedral



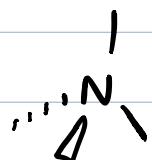
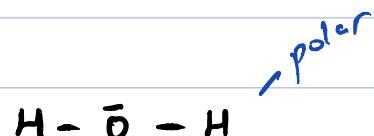
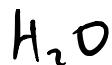
Trigonal Pyramidal

< 109.5

position of lone pair

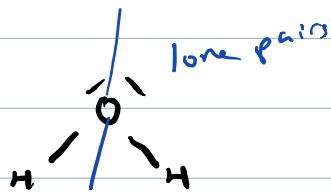
taking into account lone pair

lone pair squeezes on the others



Tetrahedral

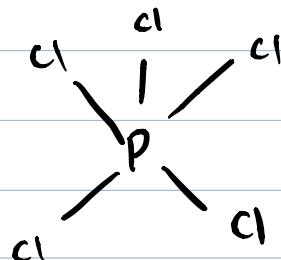
* Not responsible
for angles



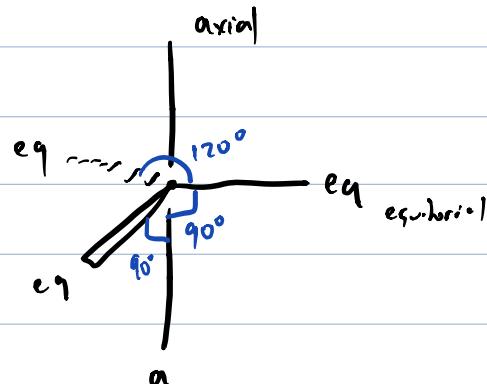
V-shaped or Bent

<109.5

October 26, 2012

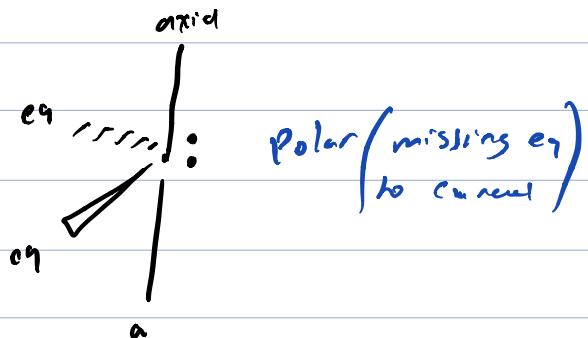
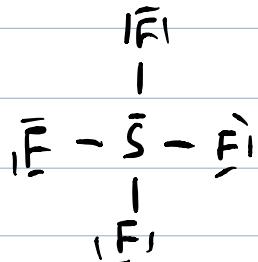
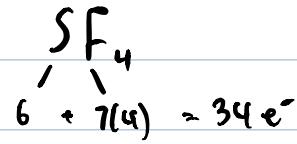


Non Polar
(cancel out)



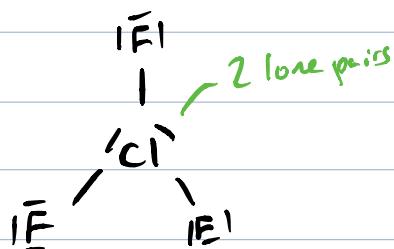
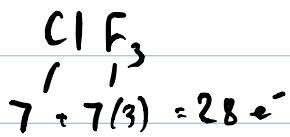
Trigonal Bipyramidal (Same for molecular geometry)

$90^\circ, 120^\circ$



Arrangement: Trigonal Bipyramidal
of e⁻ pairs

Molecular Geometry: Distorted tetrahedral or see saw
 $< 90^\circ$ and $< 120^\circ$
Lone pairs will always be placed in equatorial positions



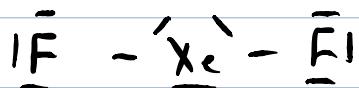
Arrangement: Trigonal Bipyramidal

M. Geometry: T-Shaped POLAR

$< 90^\circ$

XeF_2

$$8 - 7(2) = 22e^-$$

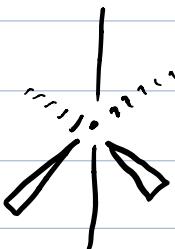
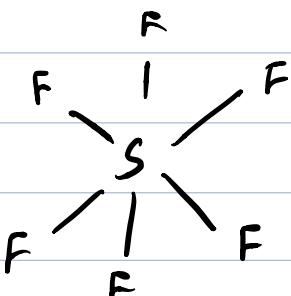


Arrangement : Trigonal Bipyramidal
of e⁻ pairs

M. Geometry : Linear Non Polar

180°

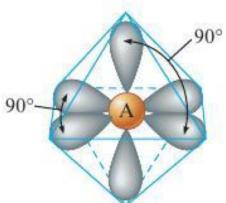
SF_6



Octahedral Non Polar

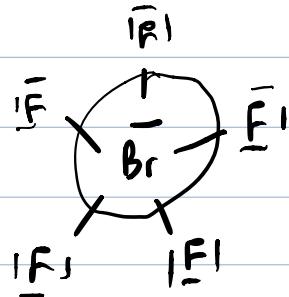
90°

No axial, no equatorial



BrF_3

$$7 + 7(3) = 42 \text{ e}^-$$



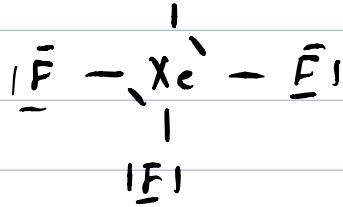
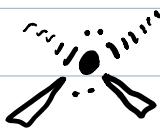
σ -pair arrangement: Octahedral

M. geo: Square Pyramidal Polar
 $< 90^\circ$

XeF_4

$$8 + 7(4) = 36 \text{ e}^-$$

$$\frac{8}{28} \text{ e}^-$$



Arrangement: Octahedral
or σ -pairs

M. geo: Lone pairs as far apart
as possible

Square Planar

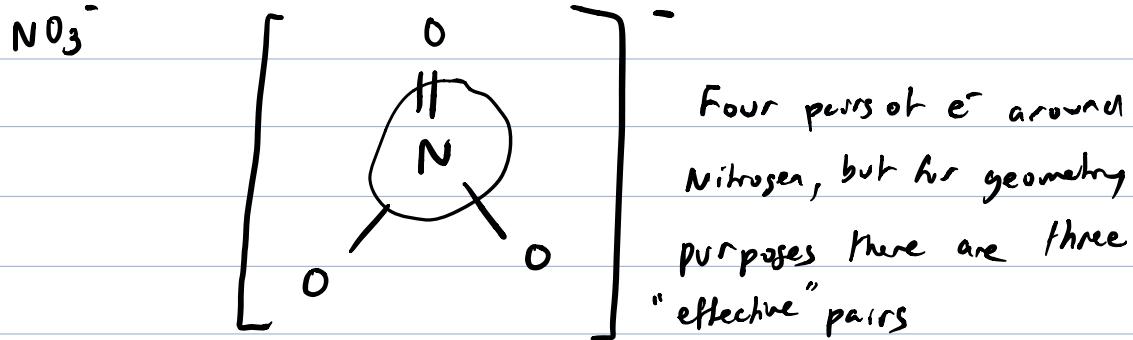
90° Non-polar

ALL SINGLE BONDS

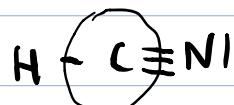
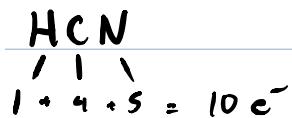
Now with double bonds...

MULTIPLE BONDS

Will count as if they were single bonds (single source time charge)

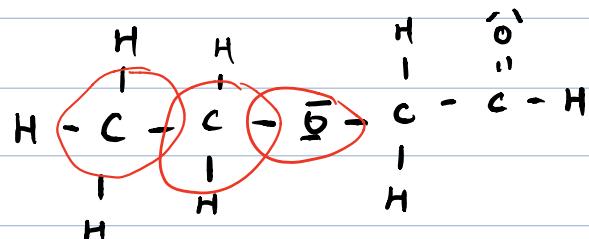


Trigonal Planar



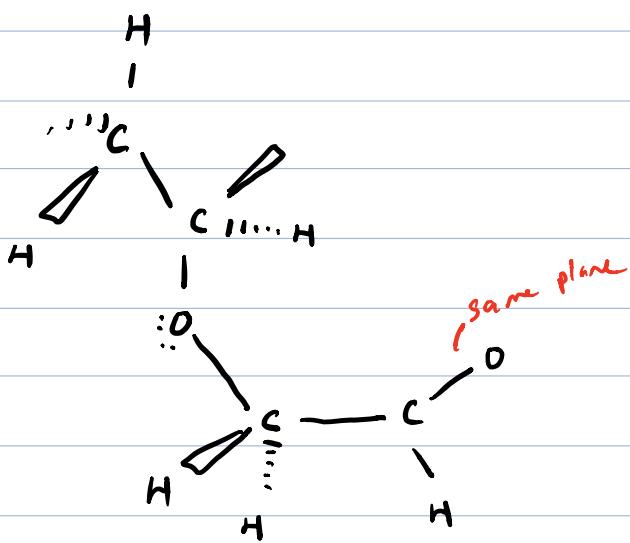
Linear

MOLECULES WITH NO SINGLE CENTRAL ATOM

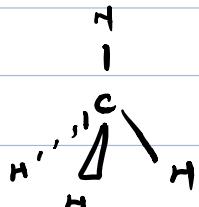
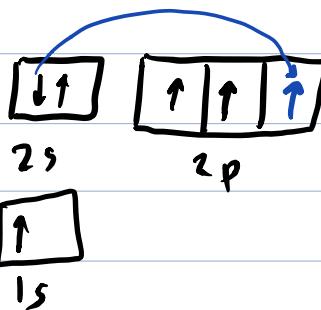
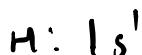
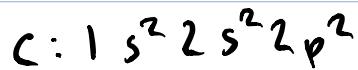
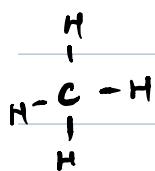


Not responsible for transcribing

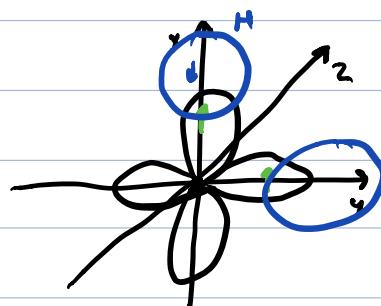
One C-atom structure



HIBRIDIZATION THEORY

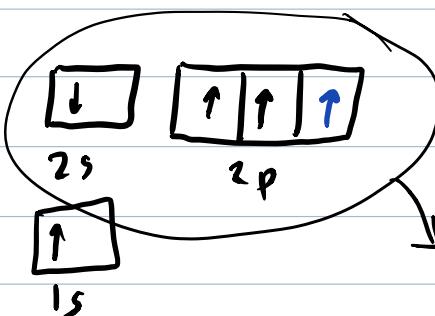
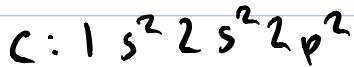


tetrahedral



October 28, 2016

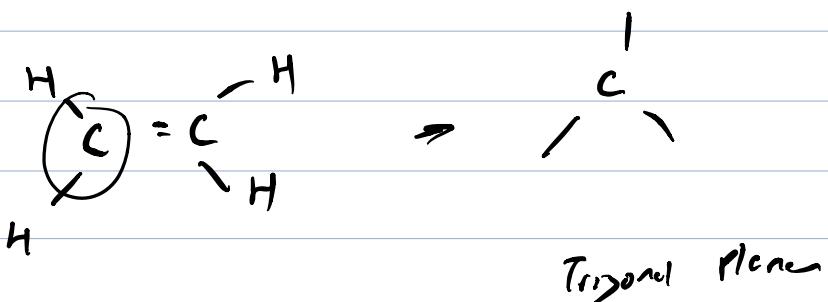
overlap: molecular orbital, opposite spins, same rules



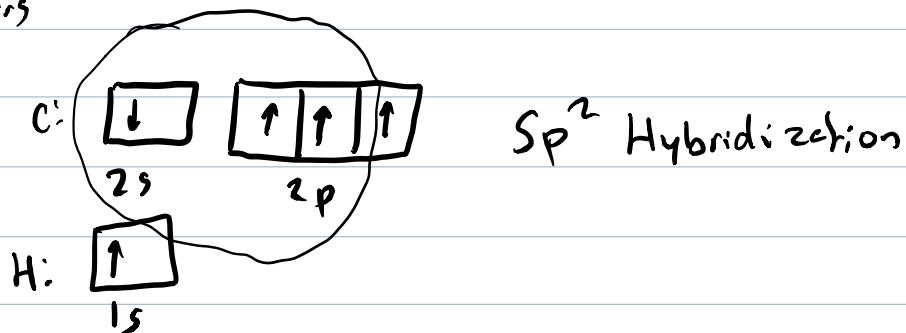
Put in blender:
then you get

Process called hybridization →
the true molecular structure

Central atom with 4 pairs of e⁻ around: sp^3 HYBRIDIZATION
→ 4 sp^3 hybridized orbitals are formed (109.5°)



3 effective pairs



$3sp^2$ Hybridized orbitals are formed (120°)

Sigma bond: formed through a head-on overlap, stronger than π -bond.

(i)

π bond: sideways overlap (of the remaining p's)

double bond: pi bond + Sigma bond

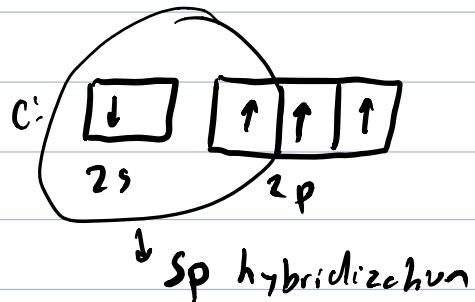
↑
stronger bond

so double bond ^{in energy} > single, but
" ≠ 2 single

* can't have a π -bond by itself, without σ -bond

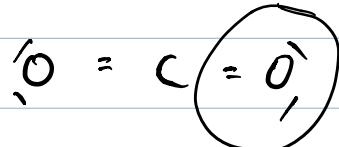


2 effective pairs



You get a linear arrangement

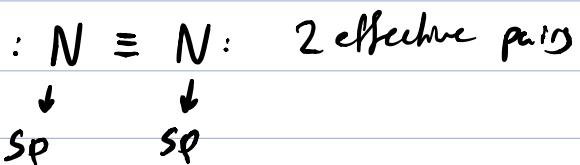
So 2 sp hybridized orbitals are formed (180°)



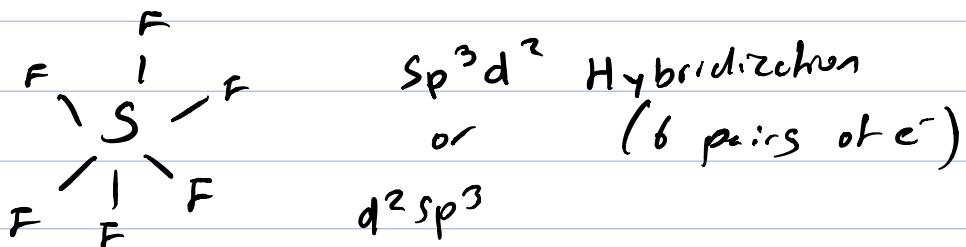
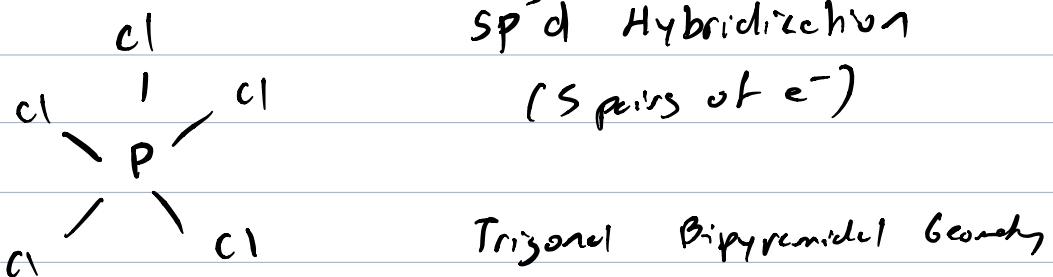
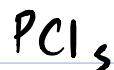
3 effective pairs

so sp^2 hybridized

so 1 p orbital remaining



* No set of bonds can involve more than 1 sigma bond



END OF TEST MATERIAL

* Non polar bonds: + → -

r-

o

,

most electronegative