Chemistry Private Lessons 10/10/24

Filling of orbitals

The shorthand method:

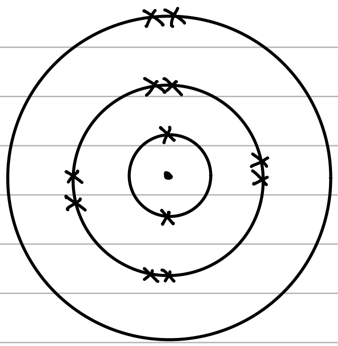
$_{26}Fe \rightarrow 1s^2, 2s^2, 2p^6, 3s^2, 3p^6, 4s^2, 3d^6$ → $[Ar]\ 4s^2, 3d^6$

$_{25}Mn \rightarrow 1s^2, 2s^2, 2p^6, 3s^2, 3p^6, 4s^2, 3d^5$ → $[Ar]\ 4s^2, 3d^5$

$_{30}Zn \rightarrow 1s^2, 2s^2, 2p^6, 3s^2, 3p^6, 4s^2, 3d^{10}$ → $[Ar]\ 4s^2, 3d^{10}$

Size of atoms

Along a period:
↳ nuclear pull increases
↳ shielding remains the same
↳ effective nuclear pull increases ∴ size decreases



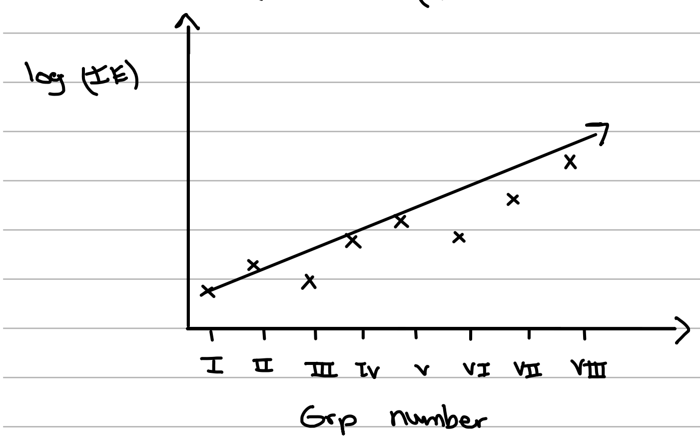
Shielding Defenition: the effect of the inner shells on the nuclear pull.

Affective nuclear pull Defenition: nuclear pull - shielding (the pull on the outermost e^-).

Down the group:
↳ nuclear pull increases
↳ shielding increases
↳ effective nuclear pull remains the same
∴ size is dependant on number of shells
∴ size increases.

Ionisation Energies Defenition: the energy required for one mole of gaseous atoms to lose one mole of electrons to form one mole of unipositive ions at STP.

$X_{(g)} \rightarrow X^{+}_{(g)} + e^{-}$



∴ The overall trend of ionisation energies is that it increases along a period. This is due to both having a smaller radius and a higher effective nuclear pull.

<u>Exceptions in ionisation energies:</u>

1. Grp II vs Grp III

Be vs B

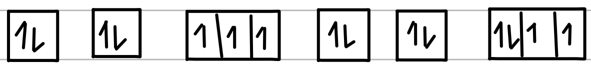
$1s^2, 2s^2, 2p^0$ $1s^2, 2s^2, 2p^1$

↳ Since in Be all the electrons will fill the s orbitals. In B the the final electron will go into the p orbital which is shielded by the full $2s^2$ orbital. This means that the effective nuclear pull will be reduced a little bit.

2. Grp V vs Grp VI

N vs O

$1s^2, 2s^2, 2p^3$ $1s^2, 2s^2, 2p^4$



↳ In N (Grp V) all the electrons are singled unpaired electrons as soon as an extra electron is added (ex: oxygen) the electrons will start to pair up. This means that the electrons will repel and less energy will be required to lose an electron.

<u>Radioactivity</u>

<u>Definition of half lives</u>: the time taken for half the atomic nuclei to decay radioactivity.

$N = No \cdot 2^{(\frac{1}{n})}$

100 -> 50 -> 25 -> 12.5

1st ½ life, 2nd ½ life, 3rd ½ life

Examples:

1. The half life of ^{71}Zn is 2.4 minutes if one had 100g at the beginning how many grams would be left after 7.2 minutes?

$7.2 \div 2.4 = 3$ half lives   3 half lives ∴ 100 -> 50 -> 25 -> 12.5 ∴ 12.5g

2. After 24 days 2mg of an original 128g sample remain, what is the half live of the sample?

128 -> 64 -> 32 -> 16 -> 8 -> 4 -> 2

∴ 6 half lives

$$Ans = \frac{24 \text{ days}}{6} = 4 \text{ days}$$

3. How much time will be required for a sample of ^{3}H to lose 75% of its radioactivity? The half life of ^{3}H is 12.26 years.

75% decay or 25% left

100 -> 50 -> 25 ∴ 2 half lives

0% decay -> 50% decay -> 75% decay

1 half life = 12.26

2 half lives = 12.26 × 2

= 24.52 years

Carbon Dating

$$^{14}_{6}C \rightarrow ^{14}_{7}N + ^{0}_{-1}e^{-}$$

$\frac{1}{2}$ life = 5730 years

The ammount of carbon in a living organism varies depending on the CO_2 and food intake of the organism. Than once an organism dies the CO_2 start to decompose without the possibility of replenishing it resulting in the decay of Carbon 14.

Nuclear Equations

1. α particle bombardment

$$^{14}_{7}N + ^{4}_{2}\alpha \rightarrow ^{17}_{8}O + ^{1}_{1}P$$

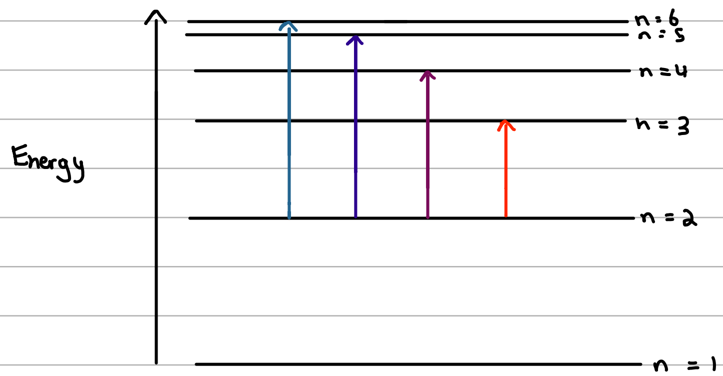
$$^{27}_{13}Al + ^{4}_{2}\alpha \rightarrow ^{30}_{15}P + ^{1}_{0}n$$

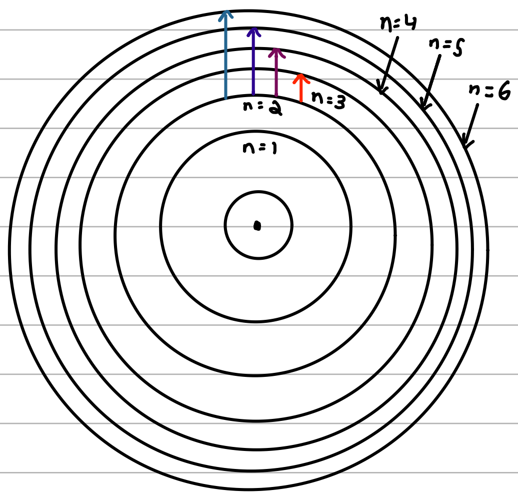
2. n bombardment

$$^{235}_{92}U + ^{1}_{0}n \rightarrow ^{141}_{56}Ba + ^{92}_{36}Kr + 3\,^{1}_{0}n$$

$$^{14}_{2}N + ^{1}_{0}n \rightarrow ^{16}_{6}C + ^{1}_{1}p$$

Atomic Spectra



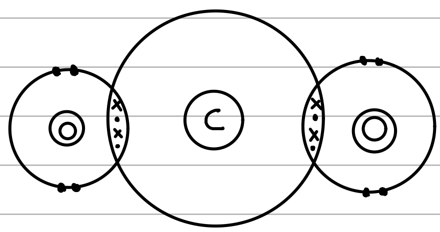


Bonding

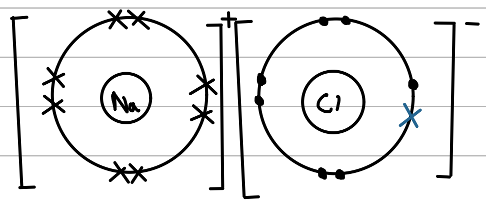
1. CO_2

$_6C$

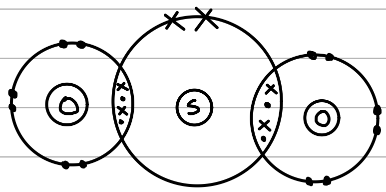
$_8O$



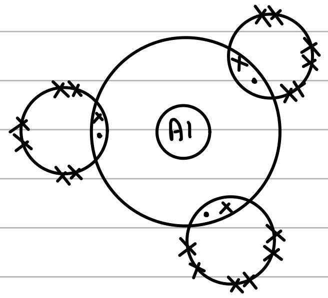
2. NaCl



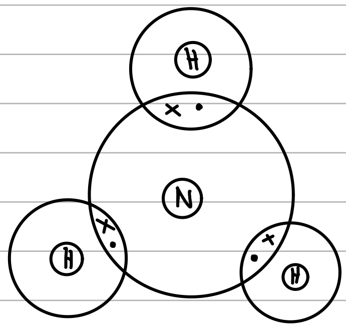
3. SO_2



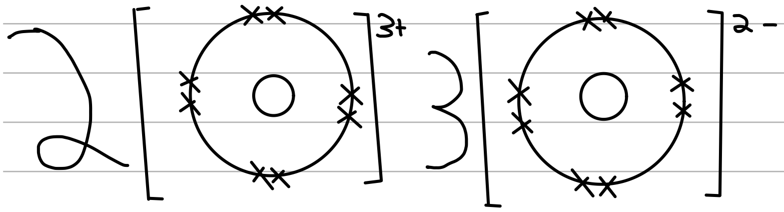
4. $AlCl_3$



5. NH_3



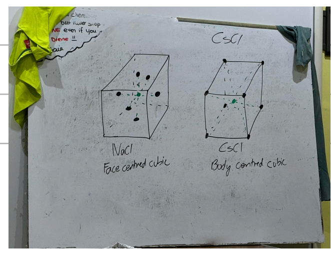
6. Al_2O_3



Ionic Bonding

↳ In ionic bonding there are no real bonds just electrostatic attractions. This means that rather than having one cation and one anion, one would have an array of ions organized in a lattice structure.

In a lattice structure the cations and anions are organized as repeating units, in NaCl each sodium is surrounded by 6 chlorides whereas in CsCl each Cs is surrounded by 8 Cl^- ions. This difference is due to the fact that Cs is much bigger and ∴ it can allow for more chloride ions to surround it.



Electrostatic Forces

There are 3 types of electrostatic forces:

1. Induced dipole - induced dipole
2. Permanent dipole - permanent dipole
3. H - bonding

These forces are dependent on the electronegativity of the atoms.

Electronegativity Definition: a measure of the attraction of an atom for electrons in a covalent bond.

- Flourine is the most electronegative atom, while Francium would be the least electronegative.
- An electronegativity value of less than 1 would produce a covalent bond.
- An electronegativity value of more than 2 would produce an ionic bond.
- An electronegativity value of between 1 and 2 would produce an intermediate bond.

There are 2 main trends in electronegativity:

1. Going down the group electronegativity decreases.
2. Going along a period electronegativity increases.

It is important to note that in reality there is a whole spectrum between covalent and ionic character, with a difference of 0 in the electronegativity being purely 100% covalent and a difference of 3 in the electronegativity being 100% ionic. All other compounds would have either a degree of covalency or of ionic character.

Polar Bonds

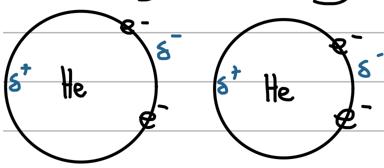
$A^{\delta+} \rightarrow B^{\delta-}$ → When 2 atoms have differen electronegativities than the electrons will not be equally shared. The arrow on the bond shows the direction of the transfer of electrons, resulting in B having a slightly higher density. Which is denoted as a delta minus (δ^-). The other atom must than be denoted as a δ^+.

Polar Molecules

$B^{\delta-} \Leftarrow A^{\delta+} \Rightarrow B^{\delta-}$ → molecules can be either polar or non-polar, a polar molecule is any molecule that is unsymmetrical and contains different electronegativities. Symmetrical molecules are in fact considered as non-polar molecules. In these molecules any polarity will cancel others out such as in $O=C=O$.

Induced dipole - Induced dipole

In the cases where the molecule is symmetrical or the atoms have very similar electronegativities, temporary charges will be formed resulting from the random movement of electrons. This movement of electrons creats partial charges with different opposite charges creating temporary intermolecular forces.



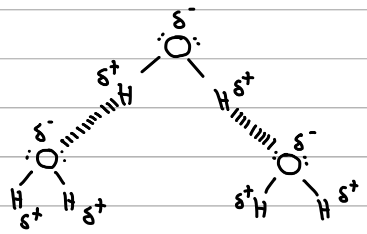
The movement of electrons will induce electrons in other atoms to shift creating more induced dipoles. In fact the more electrons present in an atom the more induced dipoles are produced and ∴ the more temporary bonds are formed.

Permanent dipoles - Permanent dipoles

$H^{\delta+} \rightarrow Cl^{\delta-} - - - - H^{\delta+} \rightarrow Cl^{\delta-}$ —> unsymmetrical molecules which have different electronegative atoms will form permanent polar bonds (permanent dipoles). This means that adjacent molecules can for permanent dipoles - permanent dipole interactions through the charges of the atoms.

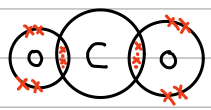
Hydrogen Bonding

A Hydrogen bond is a special permanent dipole, this involves the interaction of a hydrogen with an electronegative atom such as hydrogen, oxygen or fluorine.

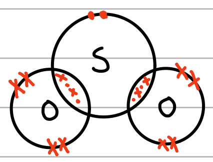


Bonding - Shapes

4 6
CO_2



6 6
SO_2

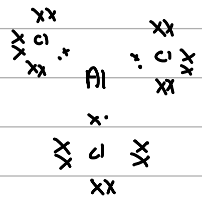


The octet rule (having a full outer shell) is only true for the outermost shell (atom), in fact the central atom can end up having either less than 8 e⁻'s (electron deficient) or more than 8 e⁻'s (d shell expansion)

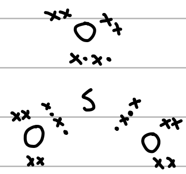
The VSEPR predicted shapes of molecules can be found in a systematic way by using the number of electron pairs to determine the shape of the molecules. To predict the shape of the molecules, first, draw out the Lewis structure of the molecule. In order to be able to draw the correct shape always do the following steps, in order.

- Identify the central atom
- Note the number of bonding electrons.
- Note the number of valence electrons (number of electrons in the outer shell).
- Note the charge of the central atom.
- Find the number of lone pairs: Valence electrons - bonding electrons - charge.
- Steric Number: Atoms joined to Central atom + lone pairs.

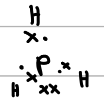
$AlCl_3$



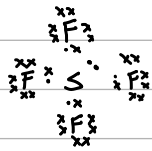
SO_3



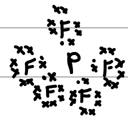
5 1
PH_3



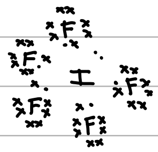
6 7
SF_4



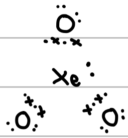
5 7
PF_5



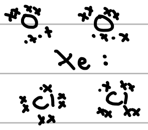
7 7
IF_5



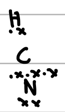
XeO_3
8 6



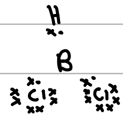
XeO_2Cl_2
8 6 7



HCN
1 4 5



$BHCl_2$
3 1 7



Steric Number = number of surrounding atoms + lone pairs

$AlCl_3$ s.n = 3 ∴ trigonal planar

SO_3 s.n = 3 ∴ trigonal planar

PH_3 s.n = 4 ∴ Angular

SF_4 s.n = 5 ∴ seesaw

PF_5 s.n = 5 ∴ Trigonal Bipyramidal

IF_5 s.n = 6 ∴ square pyramidal

XeO_3 s.n = 4 ∴ Trigonal pyramidal

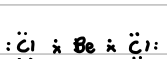
XeO_2Cl_2 s.n = 5 ∴ Seesaw

HCN s.n = 2 ∴ Linear

$BHCl_2$ s.n = 3 ∴ trigonal Planar

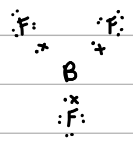
<u>HW Exercise:</u>

1. $BeCl_2$
2 7



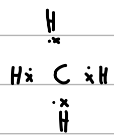
∴ Linear

2. BF_3
3 7



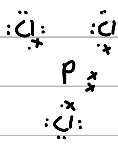
∴ trigonal Planar

3. CH_4
4 1



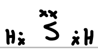
∴ tetrahedral

4. PCl_3
5 7



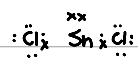
∴ trigonal pyramidal

5. H_2S
1 6



∴ Bent

6. $SnCl_2$
4 7

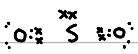


∴ bent

7. CO_2
4 6

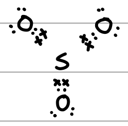
∴ Linear

8. SO_2
6 6



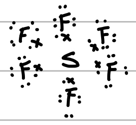
∴ Bent

9. SO_3
6 6



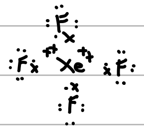
∴ Trigonal Planar

10. SF_6 (6 7 6)



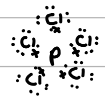
∴ Octahedral

11. XeF_4 (8 7)



∴ Square Planar

12. PCl_5 (5 7 5)



∴ Trigonal bipyramidal

Number of Electron Dense Areas	Electron-Pair Geometry	Molecular Geometry				
		No Lone Pairs	1 lone Pair	2 lone Pairs	3 lone Pairs	4 lone Pairs
2	Linear	Linear				
3	Trigonal planar	Trigonal planar	Bent			
4	Tetrahedral	Tetrahedral	Trigonal pyramidal	Bent		
5	Trigonal bipyramidal	Trigonal bipyramidal	Sawhorse	T-shaped	Linear	
6	Octahedral	Octahedral	Square pyramidal	Square planar	T-shaped	Linear