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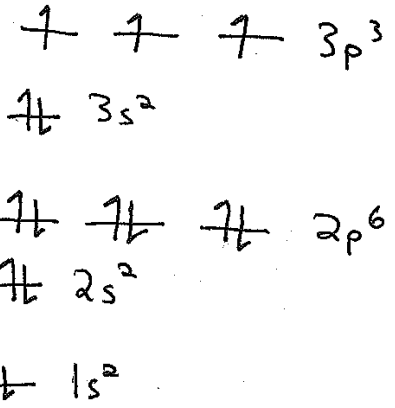
CH 235 SI

9/5/13

Session 1

Electron Configurations

Draw the energy level diagram for phosphorus. This will help to visualize and put meaning into quantum numbers n, ℓ, m_ℓ, m_s.



Write the electronic configuration of the following:

P $1s^2 2s^2 2p^6 3s^2 3p^3$

As $1s^2 2s^2 2p^6 3s^2 3p^6 4s^2 3d^{10} 4p^3$

*Cr $1s^2 2s^2 2p^6 3s^2 3p^6 4s^1 3d^5$

*Cu $1s^2 2s^2 2p^6 3s^2 3p^6 4s^1 3d^{10}$

* Test taking tip: Half-filled or completely filled shells are MORE STABLE

*As for tests, only worry about Cr + Cu unless given an ion.

Lewis Dot Structures

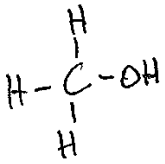
Use the following rules to draw Lewis structures:

1) Put the least electronegative atom that can have multiple bonds as the central atom.
2) Sum total valence e^- + track throughout.
3) Add single bonds to any atoms that require 2 additional e^- to fulfill its octet.
   - Treat OH as single bond
4) Add double bonds as necessary after valence e^- run out.
5) Add triple bonds as necessary.

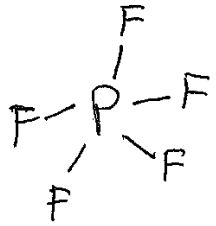
Remember to check that each atom has an octet.

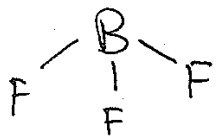
* Test taking tip: Boron + aluminum will have 6 valence e⁻; Any 3rd row p block element can expand its octet.

Draw Lewis Dot Structures for:

CH_3OH 

CO_2 $O=C=O$

PF_5 

BF_3 

Formal Charge

F.C. = # valence − # bonds − # lone e⁻

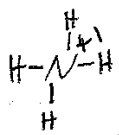
* Test tip → Shortcut: # valence − # balls − # sticks

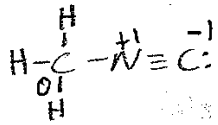
= # valence − # things

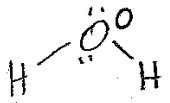
How to make each atom as happy as possible with its formal charge:

1) Keep FCs as low as possible.
2) Satisfy the nature of the atom; e.g. negative FC for electronegative atoms,
3) Keep nonzero FCs close as possible.
4) Sum of FCs should equal charge of molecule.

Assign formal charges to the following:







Bonds + Hybridization

σ bonds → single bonds

π bonds → any additional bond, e.g. double has 1 π and triple has 2 π

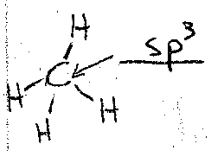
Regions of e^- density → any area with <u>any</u> amount of electrons

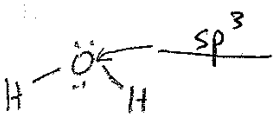
Hybridization	# areas e^- density
s	1
sp	2
sp^2	3
sp^3	4
sp^3d	5
sp^3d^2	6

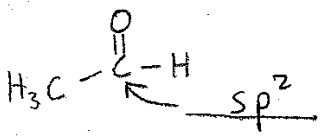
* Test taking tip: Hybridization "goes up" by one for each area of e^- density

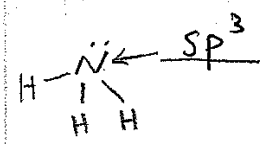
NEVER forget that not all bonds/ electrons are shown in some structures

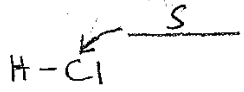
Assign a hybridization to each atom indicated:

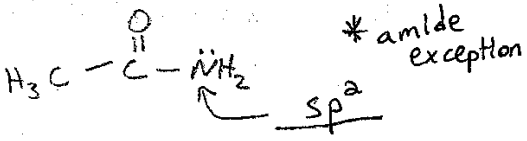




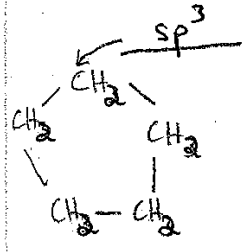


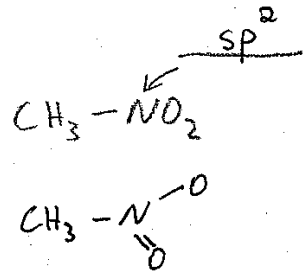


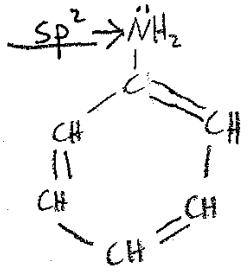




* amide exception







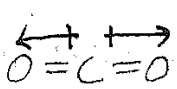
a mide exception

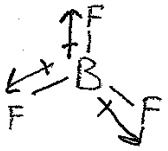
Bond Polarity

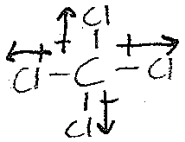
Bond polarity exists when... there is a difference in electronegativity.

* Inductive Effect: movement of e^- within the bonds of a molecule due to the electronegativity of nearby atoms

Draw any dipoles within the following molecules:

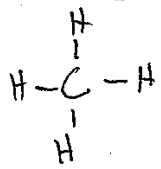




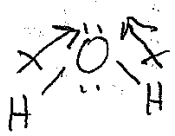


Determine if the following have net dipole moments:

CH_4 NO

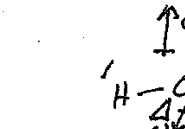


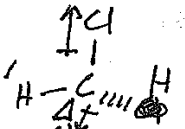
H_2O YES



CH_2Cl_2 YES

Wrong! You must look at the 3D molecular geometry!





Doesn't cancel.

*Test taking tip: Always look at the molecular geometry to determine net dipoles!