UNIT 2 – CHAPTERS 8-9 STUDENT NOTES: BONDING AND HYBRIDIZATION

Bond: a force that holds atoms together

Ionic bond: a transfer of electrons to produce stability – typically from a metal to a nonmetal

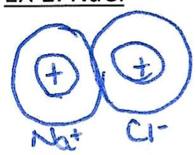
11+ Na – $1s^2 2s^2 2p^6 3s^1$

17+ Cl – $1s^2 2s^2 2p^6 3p^5$

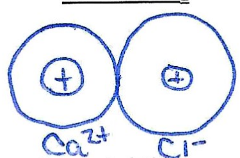
Bond energy: energy needed to break a bond (also referred to as **lattice energy**)

Coulomb's law: $E = 2.37 \times 10^{-19}\ \text{J – nm}\ \dfrac{(Q_1 \cdot Q_2)}{r^2}$

EX 1: NaCl



EX 2: $CaCl_2$



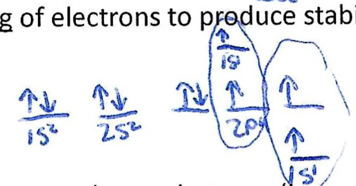
Covalent bonding: sharing of electrons to produce stability

EX 3: H_2O

HYDROGEN:

OXYGEN:

HYDROGEN:



Electronegativity: the ability to attract valence electrons (how much they like to take e^- in)

The type of bond formed is calculated using a percent ionic character

% ionic character = $\dfrac{A-B}{A}$ X 100%

Pauling's Electronegativity Values

H 2.1																	He N/A
Li 1.0	Be 1.5											B 1.5	C 2.5	N 3.0	O 3.5	F 4.0	Ne N/A
Na 0.9	Mg 1.2											Al 1.5	Si 1.8	P 2.1	S 2.5	Cl 3.0	Ar N/A
K 0.8	Ca 1.0	Sc 1.3	Ti 1.5	V 1.6	Cr 1.6	Mn 1.5	Fe 1.8	Co 1.9	Ni 1.8	Cu 1.9	Zn 1.6	Ga 1.6	Ge 1.8	As 2.0	Se 2.4	Br 2.8	Kr N/A*
Rb 0.8	Sr 1.0	Y 1.2	Zr 1.4	Nb 1.6	Mo 1.8	Tc 1.9	Ru 2.2	Rh 2.2	Pd 2.2	Ag 1.9	Cd 1.7	In 1.7	Sn 1.8	Sb 1.9	Te 2.1	I 2.5	Xe N/A*
Cs 0.7	Ba 0.9		Hf 1.3	Ta 1.5	W 1.7	Re 1.9	Os 2.2	Ir 2.2	Pt 2.2	Au 2.4	Hg 1.9	Tl 1.8	Pb 1.9	Bi 1.9	Po 2.0	At 2.2	Rn N/A
Fr 0.7	Ra 0.9																

Ionic	**50% - 100%**
Covalent	**0% - 50%**
Polar Covalent	**5% - 50%**
Nonpolar Covalent	**0% - 5%**

COVALENT BONDING

The idea of the covalent bond was first suggested by the American physical chemist G.N. Lewis in 1916. He pointed out that the electron configuration of the nobel gases appears to be a particularly stable one. Lewis suggested that nonmetal atoms, by sharing electrons to form an electron pair bond, can acquire a stable noble gas structure.

TERMS:

Valence electrons: ELECTRONS IN THE OUTERMOST PRINCIPLE ENERGY LEVEL

Lewis structure: MODEL OF A COVALENT BOND

Bonding pairs of electrons: SHARED PAIRS

Unshared pairs of electrons: NONBONDING PAIRS

Rules for writing Lewis structures

1. Sum all the valence electrons in the molecule.
2. Use a pair of electrons to form a bond between each pair of bonded atoms.
3. Arrange all the other electrons to satisfy the octet (or duet) rule.

Draw the Lewis structure for the following

H_2O

CO_2

CN^-

C_2H_6

Resonance forms: In certain cases, the Lewis structure does not adequately describe the properties of the ion or molecule that it represents. Sometimes bonds in molecules alternate between single and double bonds. This gives more than one correct Lewis structure that are referred to as **resonance forms**.

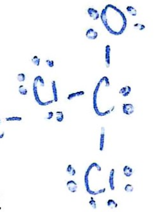
Draw the Lewis structure of the following and any resonance structures

SO_2

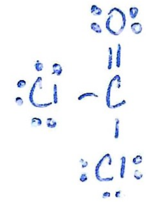
NO_3^-

$COCl_2$

$4+6+2(7) = 24e^-$



RESONANCE?



NO RESONANCE
Cl CANNOT DO
DOUBLE BONDS

Properties of Resonance

1. Resonance forms do not imply different kinds of molecules with electrons shifting eternally between them. There is only one type of SO_2 molecule; its structure is intermediate between those of the two resonance forms drawn for sulfur dioxide.

2. Resonance can be anticipated when it is possible to write two or more Lewis structures that are about equally plausible. In the case of the nitrate ion, the three structures are equivalent. One could, in principle, write many other structures, but none would put an octet around each atom.

3. Resonance forms differ only in the distribution of electrons, not in the arrangement of the atoms.

EX: A – B – C   B – A – C (not the same)

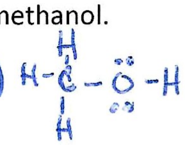
Formal charge: Often it is possible to write two different Lewis structures for a molecule, differing in the arrangement of the atoms, that is A – A – B or A – B – A. Correct arrangement can be determined by the formal charge – which follows the following equation:

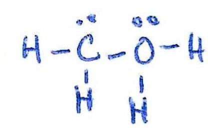
$$\mathbf{C_f = X - \frac{(Y + Z)}{2}}$$

X = the number of valence e^- in the free atom, which is equal to the last digit of the group number in the periodic table.
Y = The number of unshared e^- owned by the atom in the Lewis structure.
Z = The number of bonding e^- shared by the atom in the Lewis structure.

EX 4: Methanol (CH_3OH) has two possible Lewis structures. Calculate the formal charge on the C and O and determine the most probable Lewis structure of methanol.

$$\underset{\substack{C\\X}}{4} + \underset{\substack{H\\Y}}{3(1)} + \underset{\substack{O\\Z}}{6} + \underset{\substack{H\\Y}}{1} = 14e^-$$

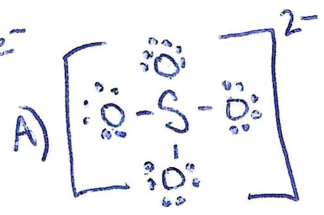
A) 

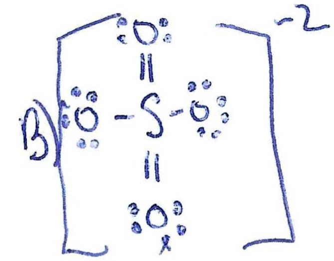
B) 

Ordinarily the more likely Lewis structure is the one which 1) the formal charges are as close to zero as possible, and 2) any negative formal charge is located on the most electronegative atom.

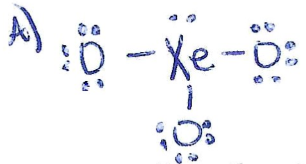
EX: Use formal charge to determine the correct Lewis structure of SO_4^{2-}

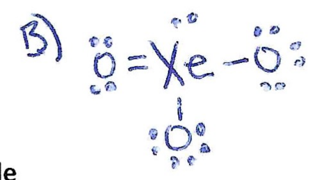
$6+4(6)+2=32e^-$

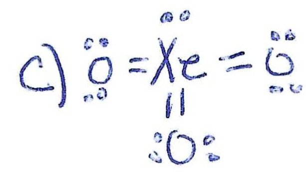




EX: Given the three Lewis structures for XeO_3 (an explosive compound), which is the most appropriate according to the formal charges? $8+3(6)=26e^-$







Exceptions to the octet rule

1. The second row elements C, N, O, and F should always be assumed to obey the octet rule.

2. The second row elements B and Be often have fewer than eight electrons around them in their compounds. These electron deficient compounds are very reactive.

EX: BF_3

$BeCl_2$

3. The second row elements never exceed the octet rule, since their valence orbitals (2s and 2p) can accommodate only 8 electrons.

4. Third row and heavier elements often satisfy the octet rule but can exceed the octet rule by using their empty valence orbitals.

EX: ClF_3

ICl_4^-

I_3^- (triiodide ion)

5. When writing the Lewis structure for a molecule, satisfy the octet rule for the atom first. If electrons remain after the octet rule has been satisfied, then place them on the elements having available d orbitals (elements in Period 3 or beyond).

Molecular Geometry – Valence Shell Electron Pair Repulsion Theory (VSEPR)

- Says the valence electron pairs surrounding an atom repel one another. Consequently, the orbitals containing those electron pairs are orientated to be as far apart as possible.
- Affects molecule characteristics like
  - Polarity
  - Bond angles

EX: For the following determine: a) the Lewis structure b) molecular geometry c) bond angles d) polarity

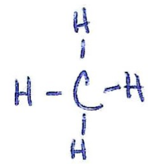
$BeCl_2$ $2+2(7)=16e^-$ 

BF_3

$3+3(7)=24e^-$

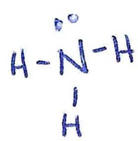
CH_4

$4+4(1)=8e^-$



NH_3

$5+3(1)=8e^-$

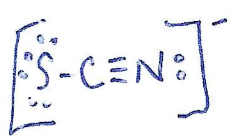


H_2O

$2(1)+6=8e^-$

SCN^-

$6+4+5+1=16e^-$



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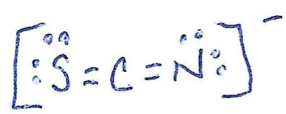
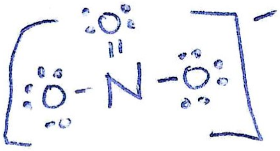


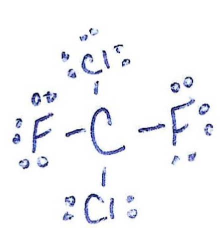
NO_3^-

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



CF_2Cl_2

$4+2(7)+2(7)=32e^-$



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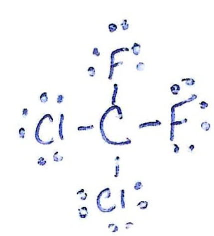


Atomic Orbitals – Hybridization

For a covalent bond to form, unpaired electrons must be matched up. When this doesn't happen, the atom must hybridize orbitals to form bonds.

Be

ATOMIC ORBITALS

2S 2P

B ↑↓ ↑↓ ↑ _ _ BF_3

1S 2S 2P

C ↑↓ ↑↓ ↑ ↑ _ CH_4

1S 2S 2P

BONDING ORBITALS

2S 2P

(ALL HAVE SAME ENERGY)

Types of hybridization

sp^3 hybridization

sp^2 hybridization

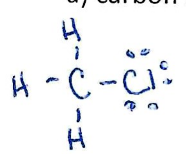
sp hybridization

dsp^3 hybridization

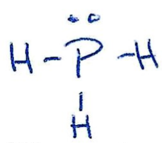
d^2sp^3 hybridization

EX 10: Give the hybridization of

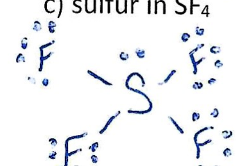
a) carbon in CH_3Cl $4+3(1)+7=14e^-$



b) phosphorous in PH_3 $5+3(1)=8e^-$

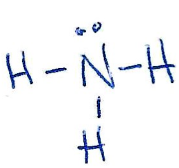


c) sulfur in SF_4 $6+4(7)=34e^-$

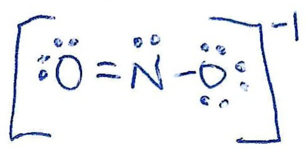


EX 11: State the hybridization of nitrogen and the number of sigma and pi bonds in the following

a) NH_3 $5+3(1)=8e^-$



b) NO_2^- $5+2(6)+1=18e^-$



c) N_2 $2(5)=10e^-$

<u>Molecular orbital (M.O.) Model</u>: describes the location of e^- in molecules

Atomic orbital model:

M.O. model:

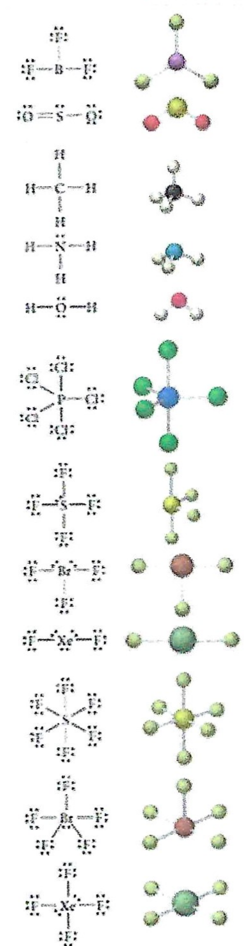
<u>EX 12</u>: Draw the M.O. model of the following molecules and give the bond order of each

NOT INCLUDED IN AP CHEMISTRY CURRICULUM AT THIS TIME.

H_2	Li_2
H_2^-	Be_2
He_2	B_2

TABLE 10.1 Electron and Molecular Geometries

Electron Groups*	Bonding Groups	Lone Pairs	Electron Geometry	Molecular Geometry	Approximate Bond Angles	Example
2	2	0	Linear	Linear	180°	
3	3	0	Trigonal planar	Trigonal planar	120°	
3	2	1	Trigonal planar	Bent	104.5°	
4	4	0	Tetrahedral	Tetrahedral	109.5°	
4	3	1	Tetrahedral	Trigonal pyramidal	107°	
4	2	2	Tetrahedral	Bent	<109.5°	
5	5	0	Trigonal bipyramidal	Trigonal bipyramidal	120° (equatorial) 90° (axial)	
5	4	1	Trigonal bipyramidal	Seesaw	<120° (equatorial) <90° (axial)	
5	3	2	Trigonal bipyramidal	T-shaped	<90°	
5	2	3	Trigonal bipyramidal	Linear	180°	
6	6	0	Octahedral	Octahedral	90°	
6	5	1	Octahedral	Square pyramidal	<90°	
6	4	2	Octahedral	Square planar	90°	



*Count only electron groups around the central atom. Each of the following is considered one electron group: a lone pair, a single bond, a double bond, a triple bond, or a single electron.