KEY

GENERAL & ANALYTICAL CHEMISTRY I
CHMG-141
With Dr. Bailey

Name___________

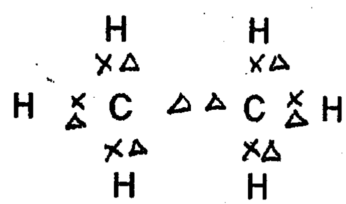
Recitation
Week 7
Covalent Bonding II

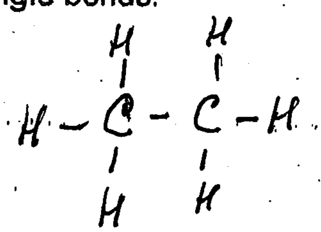
a. C_2H_6

Step 1. # valence e^- available for bonding.

C 2X4=8 e^-
H 6X1=6 e^-
14 e^- (7 pairs)

Step 2. Start Lewis structure using only single bonds:





Step 3. Distribute remaining electrons to lone pair positions on atoms that need an octet. **No remaining e^-**.

Step 4. Identify any atoms that do not yet have an octet. **None**

Step 5. Rearrange electrons from lone pairs of adjacent atoms, forming multiple bonds, so that atoms from Step 4 now have an octet. Count electrons to make sure you have the number that you started with! **No multiple bonds needed.**

Step 6. Calculate formal change on each atom.

C= 4-[0 + ½ (8)] = 0
H= 1-[0+ ½ (2)] = 0

Step 7. Are the bonds in the structure non polar covalent, polar covalent, or ionic? Can you draw resonance structures? If so how many? **All bonds are non polar, there is not a significant difference in electronegativity.** (C – 2.5
H – 2.2)

Answer Key Week 9

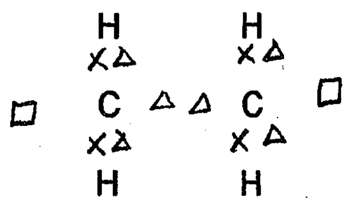
[redacted]

b. C_2H_4

Step 1. # valence e^- available for bonding

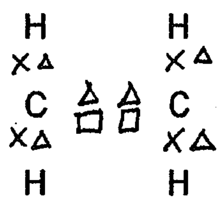
$$\begin{array}{l} C\ \ 2X4=8\ e^- \\ \underline{H\ \ 4X1=8\ e^-} \\ \qquad\quad 12\ e^- \end{array}$$ (6 pairs)

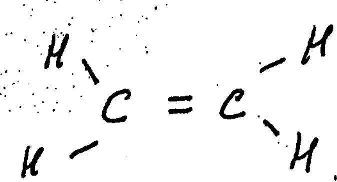
Step 2. Start Lewis structure using only single bonds:



Each C has 1 non bonded e^-

Step 3,4,5.





Step 6. Calculate formal change on each atom.

C= 4-[0 + ½ (8)] = 0

H= 1-[0+ ½ (2)] = 0

Step 7. All bonds are non polar, there is not a significant difference in electronegativity.

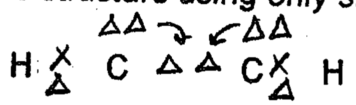
Answer Key Week 9

c. C_2H_2

Step 1. # valence e^- available for bonding

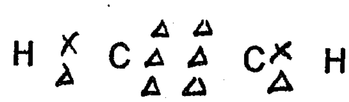
C $2X4=8\ e^-$
H $2X1=2\ e^-$
$10\ e^-$ (5 pairs)

Step 2. Start Lewis structure using only single bonds:



Each C has a lone pair of e^-

Step 3,4,5.



$H-C\equiv C-H$

Step 6. Calculate formal change on each atom.

C= 4-[0 + ½ (8)] = 0
H= 1-[0+ ½ (2)] = 0

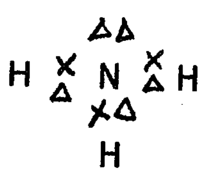
Step 7. All bonds are non polar, there is not a significant difference in electronegativity.

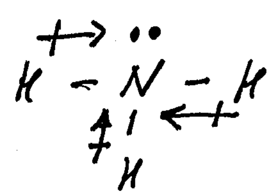
d. NH_3

Step 1. # valence e^- available for bonding

N $1X5=5\ e^-$
H $3X1=3\ e^-$ (4 pairs)
$8\ e^-$ sp^3 hybrid

Step 2. Start Lewis structure using only single bonds:





Step 3,4,5.

Step 6. Calculate formal change on each atom.

N=5-[2+ ½ (6)] = 0
H=1-[0+ ½ (2)] = 0

Step 7. The bonds in the structure are polar covalent.

Electroneg. differ. = 3.0 − 2.2 = 0.8
N H

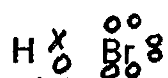
Answer Key Week 9

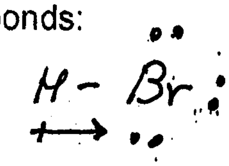
e. HBr

Step 1. # valence e^- available for bonding

Br 1X7=7 e^-
H 1X1=1 e^- (4 pairs)
8 e^- sp^3 hybrid

Step 2-5. Start Lewis structure using only single bonds:





Step 6. Calculate formal change on each atom.

Br=7-[6+ ½ (2)] = 0
H=1-[0+ ½ (2)] = 0

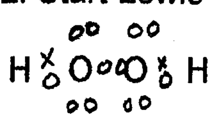
Step 7. All bonds are polar. The bond is polar

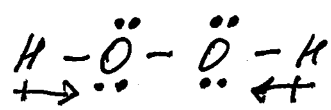
f. H_2O_2

Step 1. # valence e^- available for bonding

H 2X1= 2 e^-
O 2X6= 12 e^-
14 e^- (7 pairs)

Step 2. Start Lewis structure using only single bonds:





Step 3,4,5. Each Oxygen has 2 lone pair of electrons. Oxygen has an octet. There are no multiple bonds

Step 6. Calculate formal change on each atom.

H=1-[0+ ½ (2)] = 0
O=6-[4+ ½ (4)] = 0

Step 7. O-H ~~The bonds~~ in the structure are [redacted] polar covalent.

Memorize H_2O_2

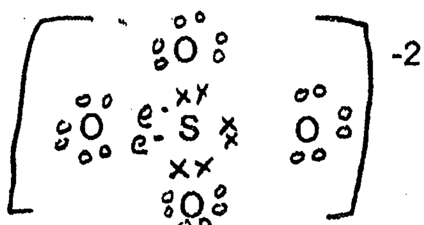
Answer Key Week ~~9~~

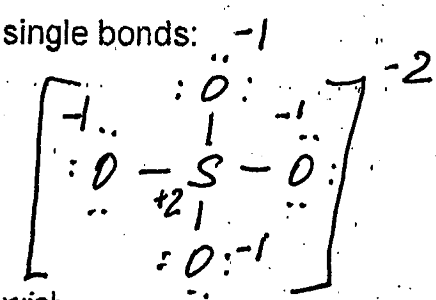
g. SO_4^{-2}

Step 1. # valence e^- available for bonding

S 1X6= 6e^-
O 4X6= 24 e^-
Charge = 2e^- (16 pairs)
32 e^- sp^3 hybrid

Step 2. Start Lewis structure using only single bonds:





~~Coordinate Covalent Bonding Gilbert Lewis~~

Step 6. Calculate formal change on each atom.

$S \quad 6-\frac{1}{2}\cdot 8-0=+2$

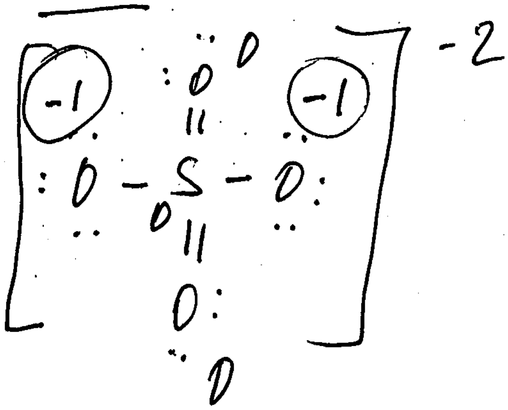
$O \quad 6-\frac{1}{2}\cdot 2-6=-1$

Step 7. The bonds are polar covalent forming (later) an ionic structure.

Ex: $CaSO_4$

Na_2SO_4

OR



$S \quad 6-\frac{1}{2}\cdot 12-0=0$

$O \quad 6-\frac{1}{2}\cdot 4-4=0$

$O \quad 6-\frac{1}{2}\cdot 2-6=-1$

Answer Key Week ~~8~~ 7

h. N_2O [use N-N-O arrangement]

Step 1. # valence e^- available for bonding

N 2X5= 10 e^-

O 1X6= 6 e^- ← (8 pairs)

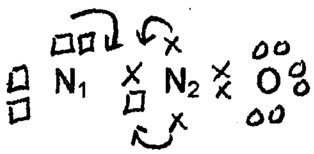
16 e^- sp hybrid

$N\ 5-\frac{1}{2}6-2=0$

$N\ 5-\frac{1}{2}\cdot8-0=+1$

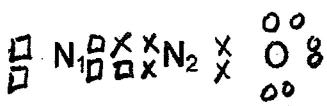
$O\ 6-\frac{1}{2}\cdot2-6=-1$

Step 2. Start Lewis structure using only single bonds:



Both N's do not have an octet

Multiple bonds



$\overset{0}{:N}\equiv\overset{+1}{N}-\overset{-1}{\ddot{\underset{\cdot\cdot}{O}}}:$

$\overset{-2}{:\ddot{\underset{\cdot\cdot}{N}}}-\overset{+1}{N}\equiv\overset{+1}{O}:$ Doesn't WORK

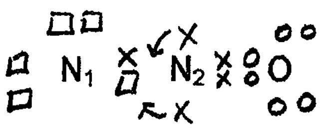
$N\ 5-\frac{1}{2}\cdot2-6=-2$

$N\ 5-\frac{1}{2}\cdot8-0=+1$

$O\ 6-\frac{1}{2}\cdot6-2=+1$

~~Coordinate covalent bond between central N, & O~~.

Triple bond between each N.



$\overset{-1}{:N}=\overset{+1}{N}=\overset{0}{\ddot{O}}:$

$N\ 5-\frac{1}{2}\cdot4-4=-1$

$N\ 5-\frac{1}{2}\cdot8-0=+1$

$O\ 6-\frac{1}{2}\cdot4-4=0$

Double Bond between N_1 & N_2, and between N_2 & O.

Very Messy! ~~Tax paid~~ Nitrous oxide AKA laughing gas.

Step 6. First Structure:

$N_1=5-[2+\frac{1}{2}(6)]=0$	$N_1=5-[4+\frac{1}{2}(4)]=-1$
$N_2=5-[0+\frac{1}{2}(8)]=+1$	$N_2=5-[0+\frac{1}{2}(8)]=+1$
$O=6-[6+\frac{1}{2}(2)]=-1$	$O=6-[4+\frac{1}{2}(2)]=0$

Step 7. The bonds in the structure are non polar covalent. There are two resonance structures.

$:\ddot{N}=N=\ddot{O}: \longleftrightarrow :N\equiv N-\ddot{\underset{\cdot\cdot}{O}}: \longleftrightarrow \cancel{:\ddot{\underset{\cdot\cdot}{N}}-N\equiv O:}$

Answer Key Week ~~9~~ 7

i. NO_3^-

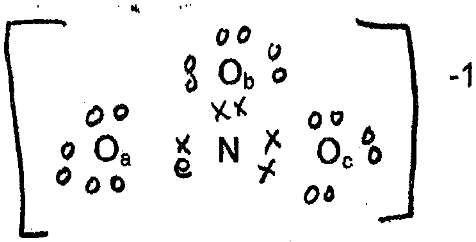
Step 1. # valence e^- available for bonding

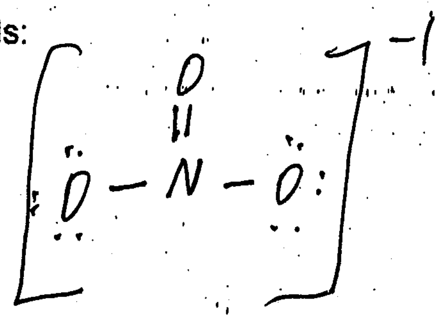
$$\begin{array}{ll} N\ 1X5= & 5e^- \\ O\ 3X6= & 18e^- \\ \underline{Charge =} & \underline{1e^-} \\ & 24e^- \end{array}$$

(12 pairs)

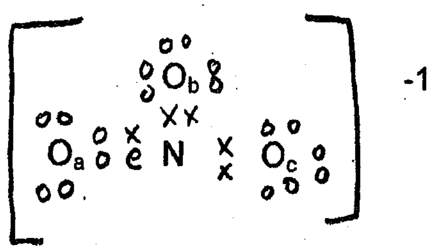
sp^2 hybrid

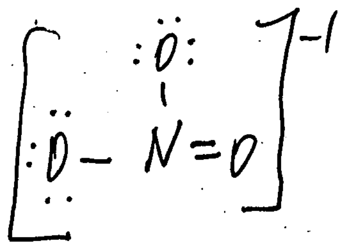
Step 2. Start Lewis structure using only single bonds:



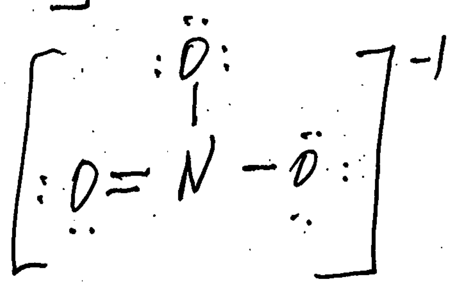


Nitrogen has only 6 electrons, needs 2 more → double bond





or



Coordinate Covalent Bonding *Gilbert Lewis*!

Step 6. Calculate formal change on each atom.

$N=5-[0+ \frac{1}{2}\ (8)] = +1$

$O_a=6-[4+ \frac{1}{2}\ (4)] = 0$

$O_b=6-[6+ \frac{1}{2}\ (2)] = -1$

$O_c=6-[6+ \frac{1}{2}\ (2)] = -1$

Step 7.

Why can't the double bond be located between O_b and Nitrogen?

Resonance structures. There are 3 resonance structures possible.

The bonds are polar covalent forming an ionic structure.

= hybrid!

Answer Key Week ~~8~~ 7

j. CO_3^{-2}

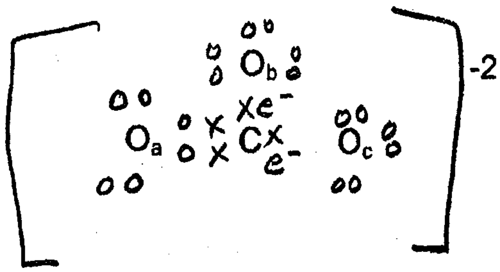
Step 1. # valence e^- available for bonding

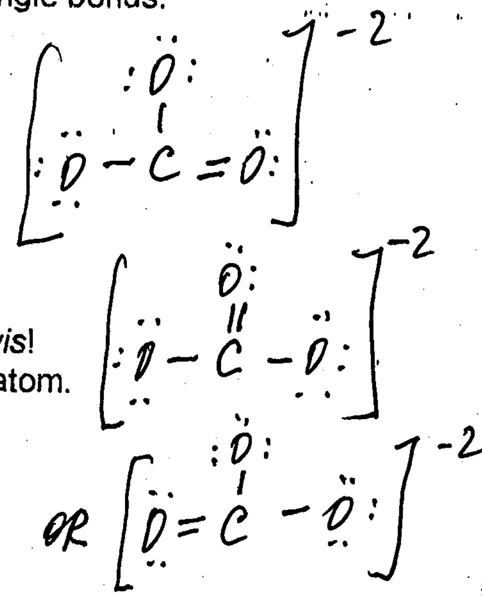
C	1X4=	$4e^-$
O	3X6=	$18e^-$
Charge =		$2e^-$
		$24e^-$

(12 pairs) sp^2 hybrid

Looks very similar to the last problem!

Step 2. Start Lewis structure using only single bonds:





Coordinate Covalent Bonding *Gilbert Lewis*!

Step 6. Calculate formal change on each atom.

$C = 4-[0+ \frac{1}{2}(8)] = 0$

$O_a = 6-[4+ \frac{1}{2}(4)] = 0$

$O_b = 6-[6+ \frac{1}{2}(2)] = -1$

$O_c = 6-[6+ \frac{1}{2}(2)] = -1$

Step 7. Why can't the double bond be located between O_b and C?

Resonance structures. There are 3 resonance structures possible. The bonds are polar covalent forming an ionic structure.

= hybrid.

Answer Key Week ~~8~~ 7

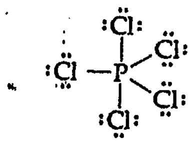
~~10/29/2008~~

k. PCl_5

Step 1. # valence e^- available for bonding

P 1X5= 5 e^-
<u>Cl 5X7= 35</u> e^-
40 e^- sp^3d hybrid **Exception to octet rule**

Step 2. Start Lewis structure using only single bonds:



Step 6. Calculate formal change on each atom.

P=5-[0+ ½ (10)] = 0
Cl=7-[6+ ½ (2)] = 0

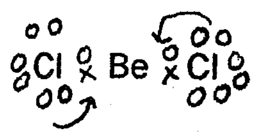
Step 7. The bonds in the structure are non polar covalent.

l. $BeCl_2$

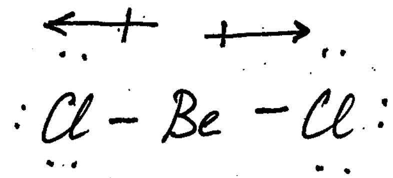
Step 1. # valence e^- available for bonding

Be1X2= 2 e^-
<u>Cl 2X7= 14</u> e^-
16 e^- sp hybrid

Step 2. Start Lewis structure using only single bonds:



Be does not have octet ~~Multiple bonds~~



Be does not have an octet. ~~Need multiple bonds.~~

Step 6. Calculate formal change on each atom.

~~Be=2-[0+½(4)]=2~~
~~Two: Cl=7-[4+½(6)]=-1~~

$Be = 2 - \frac{1}{2} \cdot 4 - 0 = 0$

$Cl = 7 - \frac{1}{2} \cdot 2 - 6 = 0$

Step 7. The bonds are polar covalent.

Answer Key Week ~~9~~ 7

m. C_6H_6 [What if the two ends of the chain connect?]

Benzene

Memorize!

Step 1. # valence e^- available for bonding

C $6X4=24\ e^-$

H $6X1=\ 6\ e^-$

$30\ e^-$ (15 pairs)

Step 2. Start Lewis structure using only single bonds:

$C_1\ C_2\ C_3\ C_4\ C_5\ C_6$ Each C has 7 valence e^-

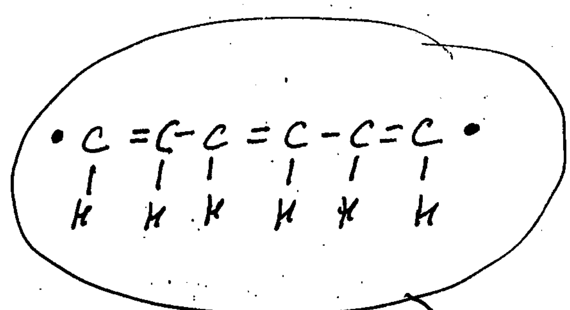
H H H H H H Multiple bonds

$C_1\ C_2\ C_3\ C_4\ C_5\ C_6$

H H H H H H

$C_1 = C_2 - C_3 = C_4 - C_5 = C_6$

H H H H H H

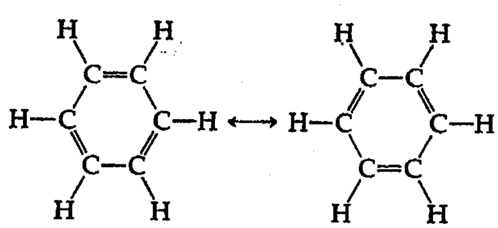


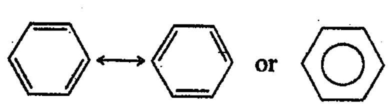
Step 6. Calculate formal change on each atom.

C= 4-[0 + ½ (8)] = 0

H= 1-[0+ ½ (2)] = 0

Step 7. The bonds in the structure are non polar covalent. There are two resonance structures. De-localized electrons.





2. The Structure of Molecules

Each person in the group should choose and complete a raw in the table below by

(a) finding a n example of a molecule or an ion with the given structure
(b) predicting the molecular or ionic geometry
(c) estimating bond angles

"A" represents a central atom; "B" represents a terminal atom, and "E" represents an unshared electron pair on the central atom.

Structure	Example	Molecular Geometry	Bond Angles
AB_2	CO_2 $\ddot{O}=C=\ddot{O}$	Linear	180°
AB_2E	SO_2 $:\ddot{O}=\ddot{S}-\ddot{O}:$	Bent	$< 120^\circ$
AB_3	BF_3 $:\ddot{F}-B-\ddot{F}:$ with $:\ddot{F}:$ on B	Trigonal planar	120°
AB_4	CH_4 NH_4^+ $[H-N(H)(H)-H]^+$	Tetrahedral	109.5°
AB_3E	PH_3 $H-\ddot{P}-H$ with H on P	Pyramidal	$< 109.5^\circ$ (107.5°)
AB_2E_2	H_2O $H-\ddot{O}-H$	Bent	$< 109.5^\circ$ (104.5°)
AB_5	PCl_5	Trigonal-bipyramidal	90°, 120°
AB_6	SF_6	Octahedral	90°