

Name KEY Date _____

☞ Due Test Day!

Pretest: Unit 4 Covalent Bonding

The following is an overview of the concepts, ideas, and problems we have covered in this unit. You are, however, responsible for all material covered, regardless if found here or not! Therefore, be sure to review all your notes, worksheets, assignments, handouts, readings, labs, problems, etc.. On the day of the test you will want to be well-acquainted with the material and organized, you will not want to waste time trying to understand an idea or searching for some needed information. Arrive prepared!

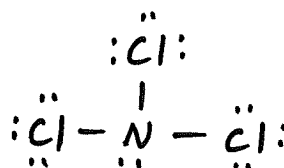
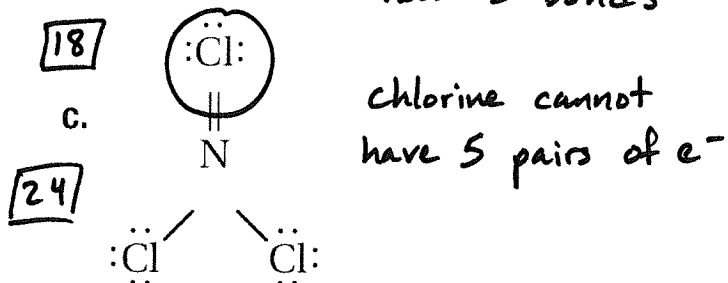
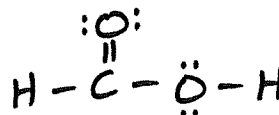
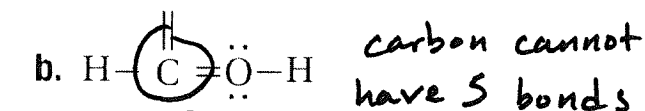
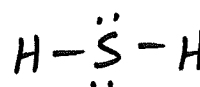
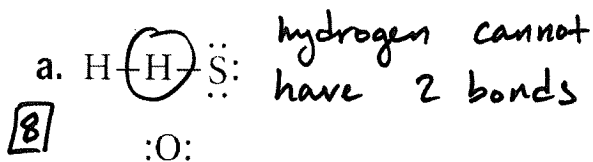
Text References:

- Introduction to Chemical Bonding (6.1)
- Covalent Bonding (6.2)
- Molecular Geometry (6.5)
- Naming Molecular Compounds (7.1, pages 215-219)

Know the following vocabulary terms listed below:

<ul style="list-style-type: none"> • acids • binary compound • bond energy • chemical bond • covalent bonding • dipole 	<ul style="list-style-type: none"> • Formal charge • hybridization • Lewis structure • molecular compound • molecular formula • molecule 	<ul style="list-style-type: none"> • multiple bond • nonpolar covalent bond • polar covalent bond • resonance • single bond • structural formula • VSEPR theory
--	--	--

1. Explain what is wrong with the following Lewis structures, then rewrite each one correctly!



Practice Problems

1.) Draw Lewis dot structures for:

a	H_2O $2(1) + 6 = \boxed{8}$	$\text{H} - \ddot{\text{O}} - \text{H}$
b	CO_2 $4 + 2(6) = \boxed{16}$	$\ddot{\text{O}} = \text{C} = \ddot{\text{O}}$
c	PF_3 $5 + 3(7) = \boxed{26}$	$\begin{array}{c} \text{:}\ddot{\text{F}}\text{ - } \ddot{\text{P}}\text{ - } \ddot{\text{F}}\text{:} \\ \\ \text{:}\ddot{\text{F}}\text{:} \end{array}$
d	H_2S $2(1) + 6 = \boxed{8}$	$\text{H} - \ddot{\text{S}} - \text{H}$
e	CN^- $4 + 5 + 1 = \boxed{10}$	$[\text{:C} \equiv \text{N:}]^-$
f	CF_4 $4 + 4(7) = \boxed{32}$	$\begin{array}{c} \text{:}\ddot{\text{F}}\text{:} \\ \\ \text{:}\ddot{\text{F}}\text{ - } \text{C} \text{ - } \ddot{\text{F}}\text{:} \\ \\ \text{:}\ddot{\text{F}}\text{:} \end{array}$
g	$\text{C}_2\text{H}_3\text{O}_2^-$ $(\text{CH}_3\text{COO}^-)$ <p>(you've had this one before!)</p> $2(4) + 3(1) + 2(6) + 1 = \boxed{24}$	$\left[\begin{array}{c} \text{H} \\ \\ \text{H} - \text{C} - \text{C} - \ddot{\text{O}}\text{:} \\ \quad \\ \text{H} \quad \text{O} \end{array} \right]^-$

2.) Name each of the following compounds:

a.) N_2O_3 dinitrogen trioxide

b.) FeP iron(III) (or ferric) phosphide

~~c.) H_3PO_4 phosphoric acid~~

d.) PBr_3 phosphorus tribromide

~~e.) HCl hydrochloric acid~~

f.) $MgCl_2$ magnesium chloride

g.) OF_2 oxygen difluoride

3.) Write formulas for the following compounds:

a.) aluminum hydroxide $Al(OH)_3$

b.) chromium(II) chloride $CrCl_2$

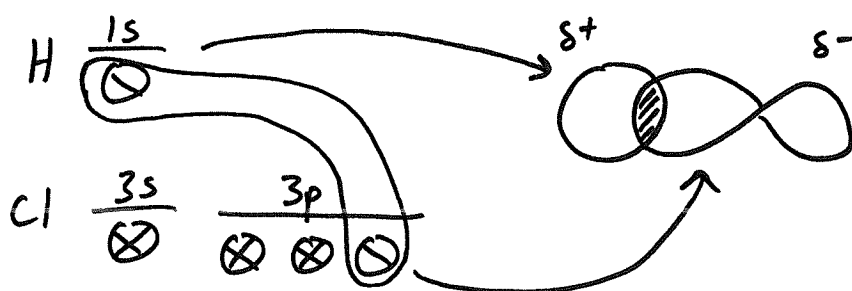
c.) phosphorus pentoxide PO_5

~~d.) hydrobromic acid HBr~~

e.) tetraphosphorus decoxide P_4O_{10}

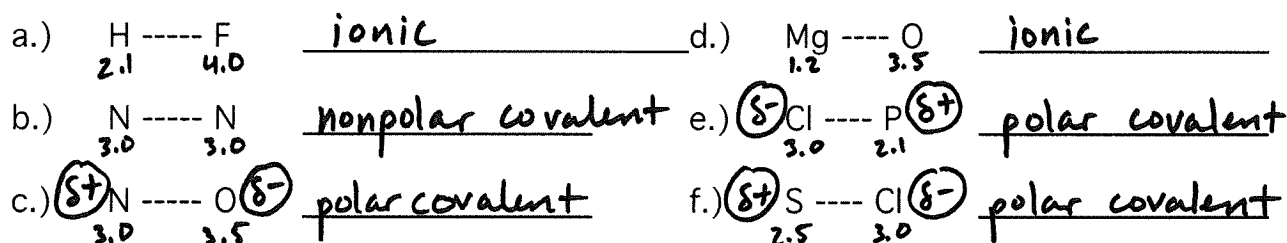
~~f.) sulfurous acid H_2SO_3~~

4.) With words and a diagram, describe the bond that occurs between chlorine and hydrogen in HCl. (Note a similar example was done in class... Check your notes!)



hydrogen's half-filled 1s orbital overlaps with chlorine's half-filled 3p orbital. forms a polar covalent bond
 $H-Cl$

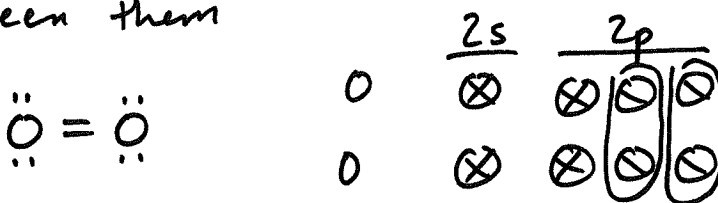
5.) Using the electronegativity values found on the last page of this packet, label each bond below as nonpolar covalent, polar covalent, or ionic. For polar covalent bonds, correctly add $\delta+$ and $\delta-$ to indicate the bond's polarity.



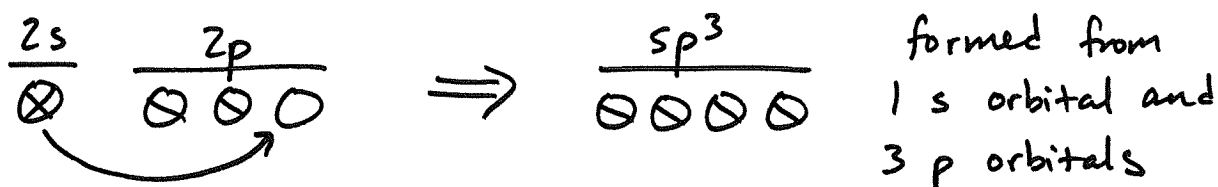
6.) In terms of bonding, why does the element oxygen always appear as a diatomic element (O_2 not just O)? Explain thoroughly, using labeled orbital diagrams in your discussion.

Each oxygen atom has 2 half-filled orbitals.

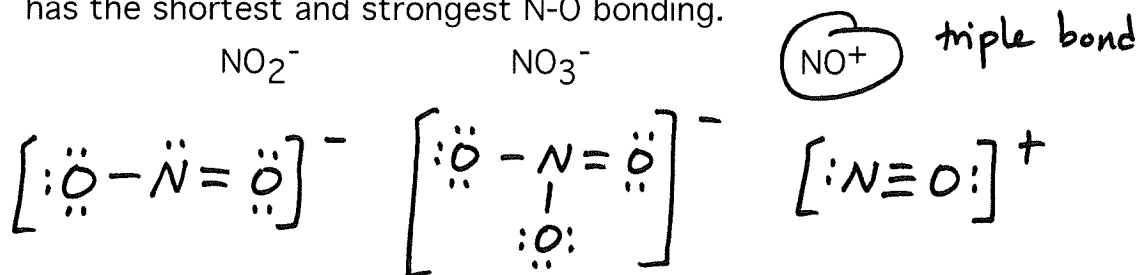
By the 2 atoms sharing their unpaired electrons, each completes its valence shell - results in a double bond between them



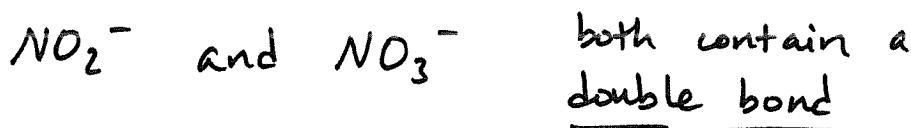
7.) Use an orbital diagram to show how carbon creates four identical hybrid orbitals. (What are these orbitals called? Why are they called this?)



8.) i. Draw a proper Lewis structure for each of the following ions, then circle the one that has the shortest and strongest N-O bonding.



ii. Which compound(s) from part i has a N-O bond that is composed of 1 pi bond and 1 sigma bond?



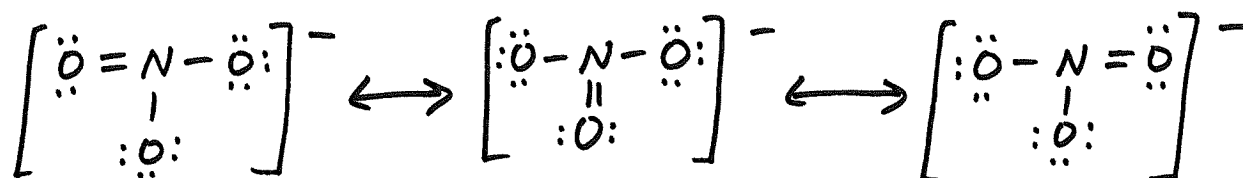
9.)

Valence e ⁻ Total	Lewis Dot Structure	Molecular Shape & Bond Angle	Name of Shape	Molecular Polarity
NO ₃ ⁻ [24]			trigonal planar	nonpolar
BrCN [16]		Br - C ≡ N 180°	linear	polar
SO ₃ [24]			trigonal planar	nonpolar
C ₂ HF [16]		H - C ≡ C - F 180°	linear	polar
HNO ₂ (oxyacid) [18]			bent 119° (with respect to N) bent 104.5° (with respect to O)	polar
BF ₃ [24]			trigonal planar	nonpolar
HCOOH [18]			trigonal planar (with respect to C) bent (with respect to the O with 2 bonds)	polar

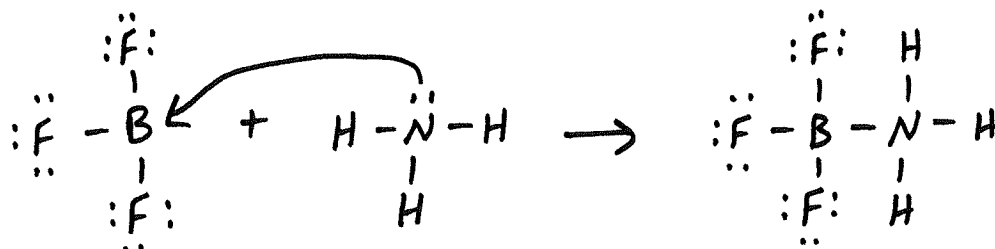
10.) Which compound(s) in problem #9 (above) contain polar bonds, but is/are nonpolar due to symmetry?

NO₃⁻, SO₃, and BF₃

11.) Draw the three resonance structures for the nitrate ion, NO_3^- .

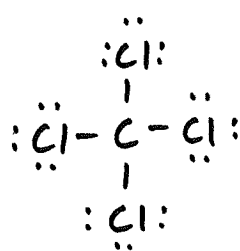


12.) Use Lewis structures and arrows to show the coordinate covalent bonding that forms between BF_3 and NH_3 .



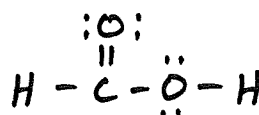
13.) Use the table below to calculate the energy need to break all the bonds in the following compounds:

a. CCl_4



$$\begin{array}{l} \text{C-Cl} \\ 4(327) = \boxed{1308 \text{ kJ}} \end{array}$$

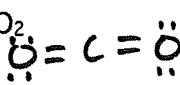
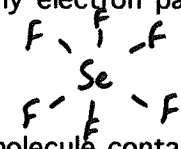
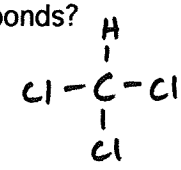
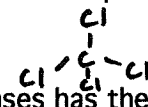
b. HCOOH



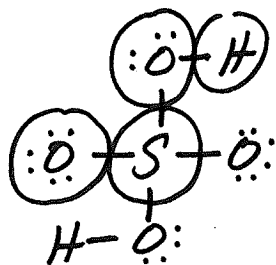
$$\begin{array}{l} \text{C-H} \quad \text{C=O} \quad \text{C-O} \quad \text{O-H} \\ (418) + (732) + (358) + (459) = \boxed{1967 \text{ kJ}} \end{array}$$

Bond	Average bond energy (kJ/mol)	Bond	Average bond energy (kJ/mol)	Bond	Average bond energy (kJ/mol)	Bond	Average bond energy (kJ/mol)
H-H	436	C-C	346	C-C	346	C-O	358
F-F	159	C-N	305	C=C	612	C=O	732
Cl-Cl	243	C-O	358	C≡C	835	C≡O	1072
Br-Br	193	C-H	418	C-N	305	N-N	163
I-I	151	C-Cl	327	C=N	615	N=N	418
H-F	569	C-Br	285	C≡N	887	N≡N	945
H-Cl	432	N-N	163				
H-Br	366	N-H	386				
H-I	299	O-H	459				

14.) Circle the correct answers.

- The common name of SiI_4 is tetraiodosilane. What is its molecular compound name?
 - silane tetraiodide
 - silane tetraiodine
 - silicon iodide
 - silicon tetraiodide
- Which of the following compounds contains at least one pi bond?
 - CO_2 
 - CHCl_3
 - AsI_3
 - BeF_2
- The Lewis structure for silicon disulfide is _____. $\text{Si}_2\text{S}_2 = 16e^-$
 - $\text{Si} = \ddot{\text{S}}:$
 - $:\ddot{\text{S}} = \text{Si} = \ddot{\text{S}}:$
 - $:\ddot{\text{S}} : \text{Si} : \ddot{\text{S}}:$
 - $:\text{S} = \text{Si} = \text{S}:$
- The central selenium atom in selenium hexafluoride forms an expanded octet. How many electron pairs surround the central Se atom?
 - 4
 - 5
 - 6
 - 7
- Chloroform (CHCl_3) was one of the first anesthetics used in medicine. The chloroform molecule contains 26 valence electrons in total. How many of these valence electrons take part in covalent bonds?
 - 26
 - 13
 - 8
 - 4
- All of the following compounds have bent molecular shapes EXCEPT _____.
 - $\text{BeH}_2 = 4e^-$ $\text{H}-\text{Be}-\text{H}$
 - H_2S
 - H_2O
 - SeH_2
- Which of the following compounds is NOT polar?
 - H_2S
 - CCl_4 
 - SiH_3Cl
 - AsH_3
- Which of the following diatomic gases has the shortest bond between its two atoms?
 - HI
 - O_2
 - Cl_2
 - N_2 $:\text{N} \equiv \text{N}:$

15.) Use the concept of Formal Charge to determine which of the following is the most likely structure for sulfuric acid, H_2SO_4 .



$$FC(\text{H}) = 1 - 1 = 0$$

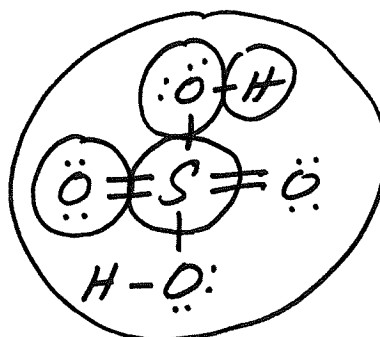
$$FC(-\text{O}-) = 6 - 6 = 0$$

$$FC(\text{S}) = 6 - 4 = +2$$

$$FC(-\text{O}) = 6 - 7 = -1$$

$$\Delta FC = (+2) - (-1) = \boxed{3}$$

or



$$FC(\text{H}) = 1 - 1 = 0$$

$$FC(-\text{O}-) = 6 - 6 = 0$$

$$FC(\text{S}) = 6 - 6 = 0$$

$$FC(=\text{O}) = 6 - 6 = 0$$

$$\Delta FC = (0) - (0) = \boxed{0}$$

		Electronegativity Values																																																																																																					
1	1	H	2.1																	Group 18	1																																																																																		
	2	Group 1																		2																																																																																			
	3	3	Li	1.0	1.5																	10	Ne																																																																																
	4	4	Na	0.9	1.2																	17	Ar																																																																																
	5	5	K	0.8	1.0	Ca	1.0	Group 3	21	Sc	1.3	Group 4	22	Ti	1.3	Group 5	23	V	1.6	Group 6	24	Cr	1.6	Group 7	25	Mn	1.5	Group 8	26	Fe	1.8	Group 9	27	Co	1.8	Group 10	28	Ni	1.8	Group 11	29	Cu	1.9	Group 12	30	Zn	1.6																																																								
	6	6	Rb	0.8	1.0	Sr	0.9	38	Y	1.2	40	Zr	1.4	41	Nb	1.6	42	Mo	1.8	43	Tc	1.9	44	Ru	2.2	45	Rh	2.2	46	Pd	2.2	47	Ag	1.9	48	Cd	1.7	77	Ir	2.2	78	Pt	2.2	79	Au	2.4	80	Hg	1.9																																																						
	7	7	Cs	0.7	0.9	Ba	0.9	56	La	1.1	57	Hf	1.3	58	Ta	1.5	59	W	1.7	60	Re	1.9	61	Os	2.2	62	Ir	2.2	63	Pt	2.2	64	Au	2.4	65	Hg	1.9	81	Tl	1.8	82	Pb	1.8	83	Bi	1.9	84	Po	2.0	85	At	2.2	86	Rn	2.4																																																
			Fr	0.7	0.9	Ra	0.9	87	Ac	1.1	88	Rf	1.3	89	Db	1.5	90	Sg	1.7	91	Bh	1.9	92	Hs	2.2	93	Mt	2.2	94	Ds	2.2	95	Rg	2.4	96	Cn	1.9	97	Uu	1.9	98	Uu	1.9	99	Uu	1.9	100	Uu	1.9	101	Uu	1.9	102	Uu	1.9	103	Uu	1.9	104	Uu	1.9	105	Uu	1.9	106	Uu	1.9	107	Uu	1.9	108	Uu	1.9	109	Uu	1.9	110	Uu	1.9	111	Uu	1.9	112	Uu	1.9	113	Uu	1.9	114	Uu	1.9	115	Uu	1.9	116	Uu	1.9	117	Uu	1.9	118	Uu	1.9