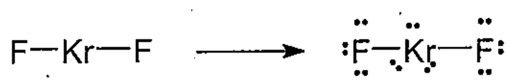
Chapter 7: Covalent Bonding

2. *Refer to Section 7.1 and Examples 7.1 and 7.2.*

Add up the total number of valence electrons. Draw the skeletal structure, then add the electrons (remember that each bond represents two electrons).

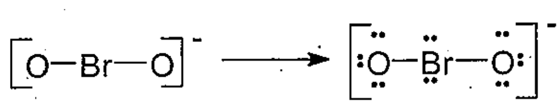
(a) N: 5 valence electrons
3H: 3 x 1 valence electrons
total: 8 electrons

(b) Kr: 8 valence electrons
2F: 2 x 7 valence electrons
total: 22 electrons



(c) N: 5 valence electrons
O: 6 valence electrons
+1: -1 valence electron
total: 10 electrons

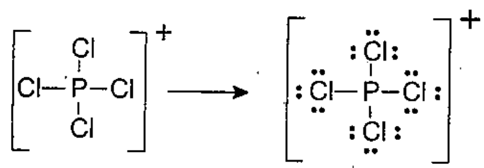
(d) Br: 7 valence electrons
2O: 2 x 6 valence electrons
-1: 1 valence electron
total: 20 electrons



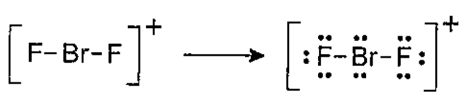
4. *Refer to Section 7.1 and Examples 7.1, 7.2, and 7.4.*

Add up the total number of valence electrons. Draw the skeletal structure, then add the electrons (remember that each bond represents two electrons).

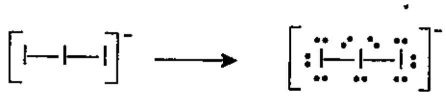
(a) P: 5 valence electrons
4Cl: 4 x 7 valence electrons
+1: -1 valence electron
total: 32 electrons



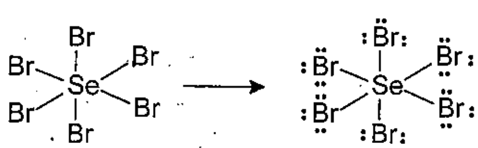
(b) Br: 7 valence electrons
2F: 2 x 7 valence electrons
+1: -1 valence electron
total: 20 electrons



(c) 3I: 3 x 7 valence electrons
-1: 1 valence electron
total: 22 electrons



(d) Se: 6 valence electrons
6Br: 6 x 7 valence electrons
total: 48 electrons



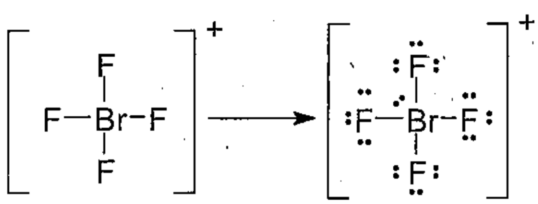
6. *Refer to Section 7.1 and Examples 7.1, 7.2, and 7.4.*

Add up the total number of valence electrons. Draw the skeletal structure, then add the electrons (remember that each bond represents two electrons). Recall the exceptions to the octet rule (for H, I and B).

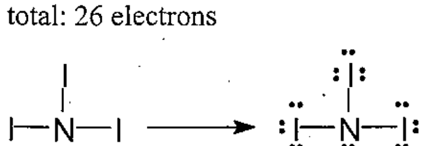
(a) 2C: 2 x 4 valence electrons
-2: 2 x 1 valence electrons
total: 10 electrons

(b) N: 5 valence electrons
F: 7 valence electrons
O: 6 valence electrons
total: 18 electrons

(c) Br: 7 valence electrons
4F: 4 x 7 valence electrons
+1: -1 valence electron
total: 34 electrons



(d) N: 5 valence electrons
3I: 3 x 7 valence electrons
total: 26 electrons



8. *Refer to Section 7.1 and Example 7.1.*

Add up the total number of valence electrons. Draw the skeleton structure, then add the electrons (remember that each bond represents two electrons). Subtract from the number of valence electrons those used for the skeleton structure. Then give each atom an octet (except H) by assigning to it an unshared pair or a multiple bond.

C: 4 valence electrons
O: 6 valence electrons
H: 1 valence electron
+1: -1 valence electron
total: 10 electrons

$[H - C - O]^{+}$ Valence electrons left: $10 - 4 = 6$

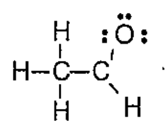
In order to give C and O octets, 5 unshared pairs of electrons (10 electrons) are needed. Since there are only 6 electrons, the octets must be filled using multiple bonds between C and O (recall that H has all the electrons it needs around it).

$[\mathrm{H} - \mathrm{C} \equiv \mathrm{O}:]^{+}$

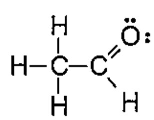
Note that a double bond between the C and O atoms would not allow C to have an octet.

10. Refer to Section 7.1 and Problem 8 above.

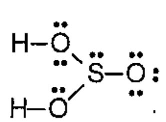
(a) 2C: 2 x 4 valence electrons
4H: 4 x 1 valence electrons
O: 6 valence electrons
total: 18 electrons



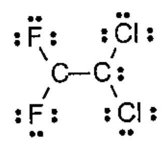
Note that this structure does not have complete octets around the carbon atoms. To complete the octets, move a pair of electrons from the oxygen to form a (double) bond between the carbon and oxygen.



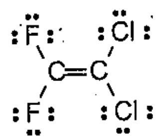
(b) S: 6 valence electrons
2H: 2 x 1 valence electrons
3O: 3 x 6 valence electrons
total: 26 electrons



(c) 2C: 2 x 4 valence electrons
2F: 2 x 7 valence electrons
2Cl: 2 x 7 valence electrons
total: 36 electrons

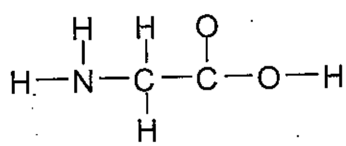


Note that this structure does not have complete octets around both carbon atoms. To complete the octets, move a pair of electrons from one carbon to form a (double) bond between the two carbon atoms.

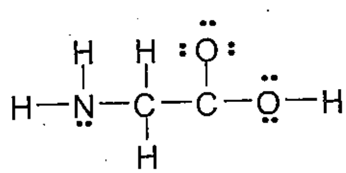


12. Refer to Section 7.1 and Example 7.1.

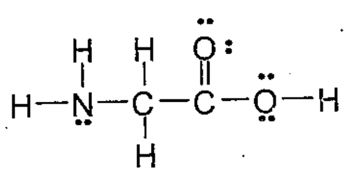
2C: 2 x 4 valence electrons
5H: 5 x 1 valence electrons
2O: 2 x 6 valence electrons
N: 5 valence electrons
total: 30 electrons



From the information given, the structure must be that on the left.



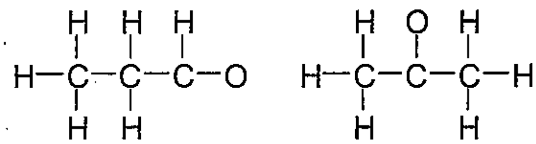
Note that this structure does not have a complete octet around one of the carbon atoms. To complete the octet, move a pair of electrons from the terminal oxygen to form a C=O double bond.



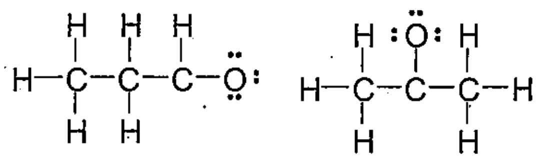
This structure provides octets for all the atoms and provides that all formal charges are zero.

14. Refer to Section 7.1.

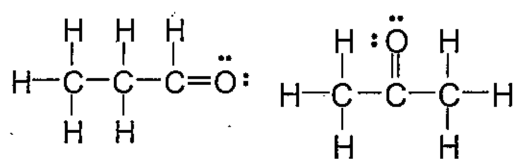
3C: 3 x 4 valence electrons
2H: 2 x 1 valence electrons
O: 6 valence electrons
total: 24 electrons



From the information given, the two structures must be those given to the left.



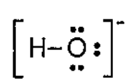
In each of the structures, one of the C's does not have a full octet. Move a pair of electrons from one oxygen to form a (double) bond between the carbon and the oxygen.

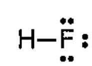


16. Refer to Section 7.1.

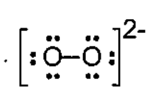
Draw the Lewis structure of the ion, then draw a molecule with the same Lewis structure. This can be facilitated by increasing the atomic number of the atom(s) by the same amount as the negative charge (recall that the negative charge results from extra electrons, and that molecules have equal numbers of electrons and protons.

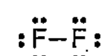
(a) O: 6 valence electrons
H: 1 valence electron
-1: 1 valence electron
total: 8 electrons



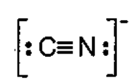


(b) 2O: 2 x 6 valence electrons
-2: 2 x 1 valence electrons
total: 14 electrons

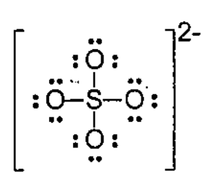


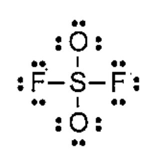


(c) C: 4 valence electrons
N: 5 valence electrons
-1: 1 valence electron
total: 10 electrons



(d) S: 6 valence electrons
4O: 4 x 6 valence electrons
-2: 2 x 1 valence electrons
total: 32 electrons

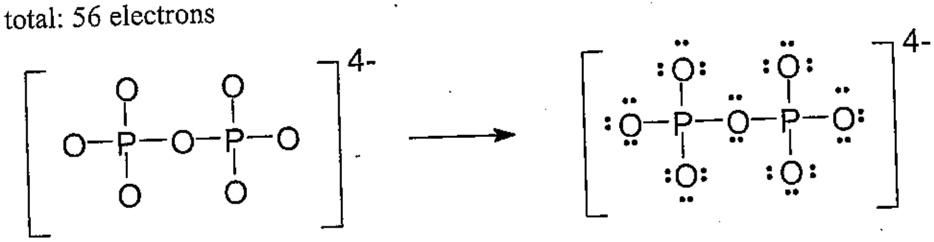




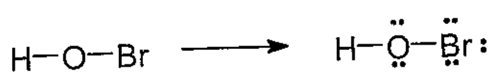
18. Refer to Section 7.1 and Examples 7.1, 7.2, and 7.4.

Add up the total number of valence electrons. Draw the skeletal structure, then add the electrons (remember that each bond represents two electrons). Recall the exceptions to the octet rule (for H, I and B).

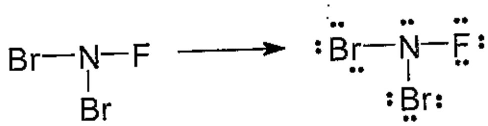
(a) 2P: 2 x 5 valence electrons
7O: 7 x 6 valence electrons
-4: 4 x 1 valence electrons
total: 56 electrons



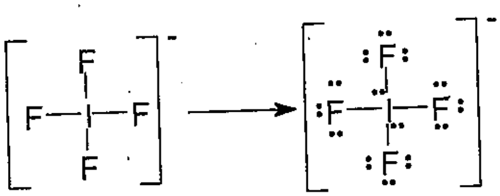
(b) H: 1 valence electron
O: 6 valence electrons
Br: 7 valence electrons
total: 14 electrons



(c) N: 5 valence electrons
F: 7 valence electrons
2Br: 2 x 7 valence electrons
total: 26 electrons



(d) I: 7 valence electrons
4F: 4 x 7 valence electrons
-1: 1 valence electron
total: 36 electrons



20. Refer to Section 7.1.

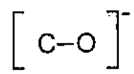
Parts b-d have odd numbers of electrons. Consequently, the final structures will have an unpaired electron.

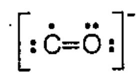
(a) Be: 2 valence electrons
2H: 2 x 1 valence electrons
total: 4 electrons

H-Be-H

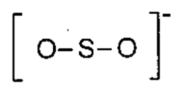
The skeletal structure is the final Lewis structure.

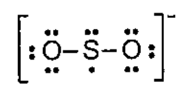
(b) C: 4 valence electrons
O: 6 valence electrons
-1: 1 valence electron
total: 11 electrons



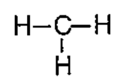


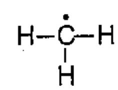
(c) S: 6 valence electrons
O: 2 x 6 valence electrons
-1: 1 valence electron
total: 19 electrons





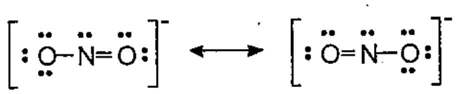
(d) C: 4 valence electrons
3H: 3 x 1 valence electrons
total: 7 electrons

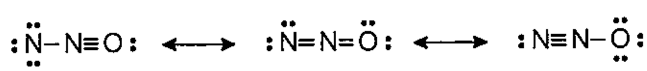


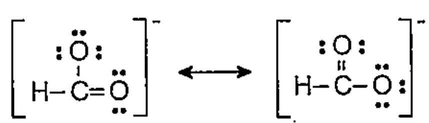


22. Refer to Section 7.1 and Example 7.3.

Write the Lewis structure for the compound, then draw resonance forms by changing the positions of electron pairs. Remember that the skeletal structure cannot change, and the new structures must also abide by the rules for Lewis structures.

(a) 

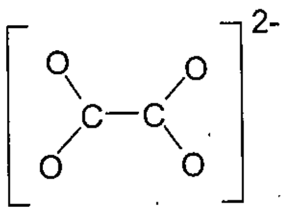
(b) 

(c) 

24. *Refer to Section 7.1 and Example 7.3.*

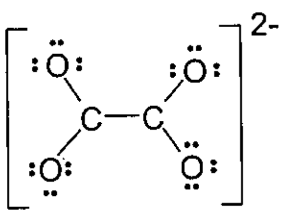
(a) Add up the total number of valence electrons. Draw the skeleton structure, then add the electrons (remember that each bond represents two electrons). Subtract from the number of valence electrons those used for the skeleton structure. Then give each atom an octet (except H) by assigning to it an unshared pair or a multiple bond.

2C: 2 x 4 valence electrons
4O: 4 x 6 valence electrons
-2: 2 x 1 valence electrons
total: 34 electrons



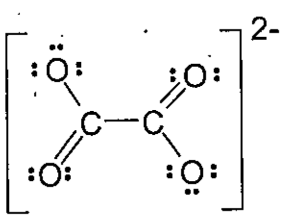
Valence electrons left: 34 – 10 = 24

Try to fill the octet of each atom using unshared pairs. There are 12 unshared pairs available.

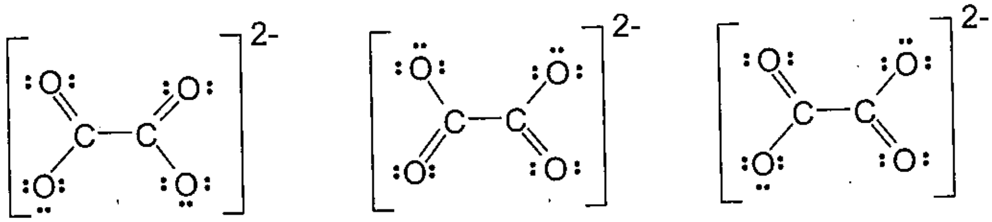


Notice that there are not enough unshared pairs to give carbon an octet.

To complete the octets, move a pair of electrons from each of two oxygens to form (double) bonds between the carbon and oxygen.



(b) The three resonance structures are drawn by moving the double bonds and the unshared pairs so that each atom still has an octet.

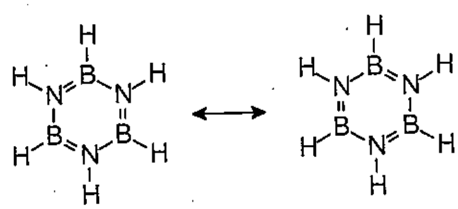


(c) Resonance forms differ only in distribution of electrons, not in arrangement of atoms. Thus, the structure given in c is **not** a resonance form.

26. Refer to Section 7.1 and Example 7.3.

Write the Lewis structure for borazine, then draw resonance forms by changing the positions of electron pairs. Remember that the skeletal structure cannot change, and the new structures must also abide by the rules for Lewis structures.

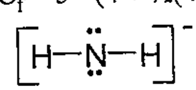
3N: 3 x 5 valence electrons
3B: 3 x 3 valence electrons
6H: 6 x 1 valence electrons
total: 30 electrons



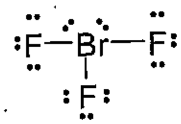
28. Refer to Section 7.1 and Table 7.2.

Draw the Lewis structure. Then apply the formula: $C_f = e_{valence} - (e_{unshared} + \frac{1}{2}(e_{bonding}))$
[formal charge = valence electrons - (unshared electrons + ½(bonding electrons))]
Note that this is the formula given in the text as: $C_f = X - (Y + Z/2)$.

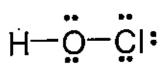
(a) N: 5 valence electrons
2H: 2 x 1 valence electrons
-1: 1 valence electron
total: 8 electrons
$C_f = 5 - (4 + \frac{1}{2}(4)) = -1$



(b) Br: 7 valence electrons
3F: 3 x 7 valence electrons
total: 28 electrons
$C_f = 7 - (4 + ½(6)) = 0$



(c) H: 1 valence electron
O: 6 valence electrons
Cl: 7 valence electrons
total: 14 electrons
$C_f = 6 - (4 + ½(4)) = 0$



30. *Refer to Section 7.2, Examples 7.5 and 7.6, Figure 7.5, and Table 7.3.*

In both structures, all 3 oxygen atoms have a formal charge of -1. $C_f = 6 - (6 + ½(2)) = -1$. Therefore, the formal charges on the sulfur atoms will determine the better structure.

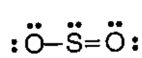
Structure I
Central S: $C_f = 6 - (2 + ½(6)) = +1$
Chain S: $C_f = 6 - (4 + ½(4)) = 0$

Structure II
Central S: $C_f = 6 - (0 + ½(8)) = +2$
Terminal S: $C_f = 6 - (6 + ½(2)) = -1$

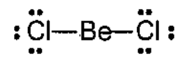
Structure I is the better choice of structures since the formal charges on those sulfur atoms are as close to zero as possible.

32. *Refer to Section 7.2, Examples 7.5 and 7.6, Figure 7.5, and Table 7.3.*

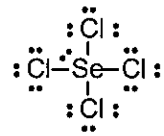
Draw the Lewis structure of the compound and determine the number of bonded groups and the number of electron pairs. Then use Table 7.3 to assign the geometry.

(a) SO_2 ⇒ 

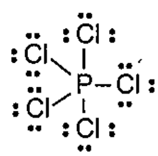
2 bonded groups, one electron pair, thus AX_2E and **bent**.

(b) $BeCl_2 \Rightarrow$ 

2 bonded groups, no electron pairs, thus AX_2 and **linear**.

(c) $SeCl_4 \Rightarrow$ 

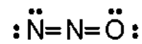
4 bonded groups, one electron pair, thus AX_4E and **seesaw**.

(d) $PCl_5 \Rightarrow$ 

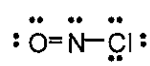
5 bonded groups, no electron pairs, thus AX_5 and **trigonal bipyramid**.

34. Refer to Section 7.2, Examples 7.5 and 7.6, Figures 7.4 and 7.5, and Table 7.3.

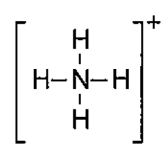
Draw the Lewis structure of the compound and determine the number of bonded groups and the number of electron pairs. Then use Table 7.3 to assign the geometry.

(a) NNO $\Rightarrow$ 

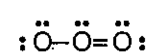
2 bonded groups, no electron pairs, thus AX_2 and **linear**.

(b) ONCl $\Rightarrow$ 

2 bonded groups, one electron pair, thus AX_2E and **bent**.

(c) $NH_4^+ \Rightarrow$ 

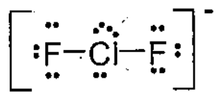
4 bonded groups, no electron pairs, thus AX_4 and **tetrahedron**.

(d) $O_3 \Rightarrow$ 

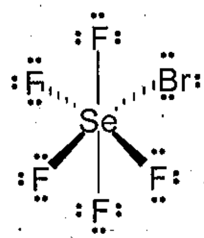
2 bonded groups, one electron pair, thus AX_2E and **bent**.

36. Refer to Section 7.2, Examples 7.5 and 7.6, Figures 7.4 and 7.5, and Table 7.3.

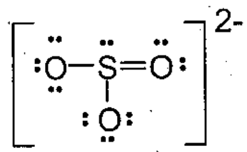
Draw the Lewis structure of the compound and determine the number of bonded groups and the number of electron pairs. Then use Table 7.3 to assign the geometry.

(a) ClF_2^- ⇒ 

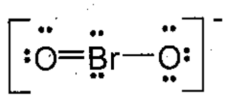
2 bonded groups, three electron pairs, thus AX_2E_3 and **linear.**

(b) SeF_5Br ⇒ 

6 bonded groups, no electron pairs, thus AX_6 and **octahedral.**

(c) SO_3^{2-} ⇒ 

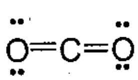
3 bonded groups, one electron pair, thus AX_3E and **trigonal pyramid.**

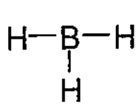
(d) BrO_2^- ⇒ 

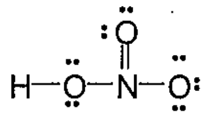
2 bonded groups, two electron pairs, thus AX_2E_2 and **bent.**

38. Refer to Section 7.2, Figure 7.5, and Table 7.3.

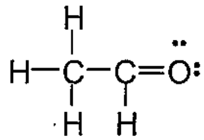
Draw the Lewis structure, determine the geometry and, from that, the ideal bond angles.

(a)  AX_2, linear **180°**

(b)  AX_3, trigonal planar **120°**

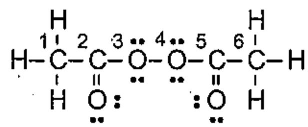
(c) 

O(left): AX_2E_2, bent **109.5°**
N: AX_3, trigonal planar **120°**

(d) 

C(left): AX_4, tetrahedron **109.5°**
C(right): AX_3, trigonal planar **120°**

40. Refer to Section 7.2, Figure 7.5, and Table 7.3.

(a) 

(b)
C^1, C^6: AX_4, tetrahedron **109.5°**
C^2, C^5: AX_3, trigonal planar **120°**
O^3, O^4: AX_2E_2, bent **109.5°**

42. Refer to Section 7.2.

Draw the Lewis structure for the molecule. Then determine the geometry of the indicated atom and from that, the bond angles.

1	AX_3	bent	**120°**
2	AX_2E_2	bent	**109.5°**
3	AX_3	trigonal planar	**120°**

44. Refer to Sections 6.8 and 7.3, and Example 7.7.

Consider the electronegativity of the bonded atoms to determine if there are dipoles. If there is a dipole, consider if it is canceled by a dipole in the opposite direction. Then consider the geometry of the molecule to determine if there is a net dipole.
The molecules with bonded atoms which are not symmetrical about the central atom are **(a) SO_2**, and **(c) $SeCl_4$**. These molecules have net dipoles.

46. Refer to Section 7.2 and Example 7.7.

First determine if the central atom has an octet. Then consider the electronegativity of the atoms to determine if there are dipoles. Then consider the geometry of the molecule to determine if there is a net dipole.

All 4 molecules have octets about the central atom.

(a) There is only one dipole, between the central N and the O. Thus, there is a net dipole and **the molecule is a dipole.**

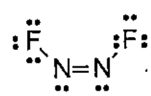
(b) There are 2 dipoles, between N-O and N-Cl. Since this molecule is bent, there is a net dipole, and **the molecule is a dipole.**

(c) There are 4 dipoles, between each of the N-H bonds. Since the molecule is tetrahedron, the 4 dipoles cancel each other out; there is no net dipole, so **the molecule is not a dipole.**

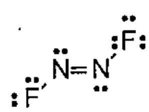
(d) Since all the atoms are identical, there are no dipoles and one would thus assume there is no net dipole. However, a dipole depends only on an unsymmetrical distribution of electrons. Since this molecule is bent, that criterion is met, and **the molecule is a dipole.**

48. Refer to Section 7.3.

Draw the Lewis structures of the molecules. Then consider the electronegativity of the atoms to determine if there are dipoles. Then consider the geometry of the molecule to determine if there is a net dipole.


cis

The *cis* structure has a dipole along each N-F bond. The two dipoles don't cancel, so this molecule has a net dipole and is **polar**.


trans

The *trans* structure also has a dipole along each N-F bond, but these two dipoles do cancel, so this molecule does not have a net dipole and is **not polar**.

50. Refer to Section 7.4, Example 7.8, and Problem 32 (above).

Recall that the total number of groups (bonded atoms and electron pairs) around the central atom is equal to the number of orbitals that hybridized. Furthermore, the sum of the superscripts in the hybrid orbital notation gives the total number of hybrid orbitals.

(a)	SO_2	AX_2E	3 groups	sp^2
(b)	$BeCl_2$	AX_2	2 groups	sp
(c)	$SeCl_4$	AX_4E	5 groups	sp^3d
(d)	PCl_5	AX_5	5 groups	sp^3d

52. *Refer to Section 7.4, Table 7.4, Example 7.9, and Problem 34 (above).*

Recall that the total number of groups (bonded atoms and electron pairs) around the central atom is equal to the number of orbitals that hybridized. Furthermore, the sum of the superscripts in the hybrid orbital notation gives the total number of hybrid orbitals.

(a)	NNO	AX_2	2 groups	sp
(b)	ONCl	AX_2E	3 groups	sp^2
(c)	NH_4^+	AX_4	4 groups	sp^3
(d)	O_3	AX_2E	3 groups	sp^2

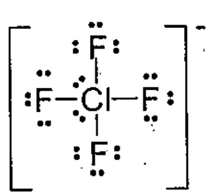
54. *Refer to Section 7.4, Example 7.8, and Problem 36 (above).*

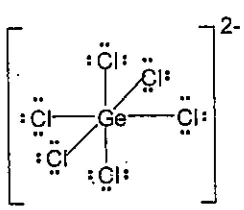
Recall that the total number of groups (bonded atoms and electron pairs) around the central atom is equal to the number of orbitals that hybridized. Furthermore, the sum of the superscripts in the hybrid orbital notation gives the total number of hybrid orbitals.

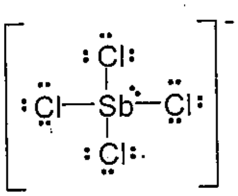
(a)	ClF_2^-	AX_2E_3	5 groups	sp^3d
(b)	SeF_5Br	AX_6	6 groups	sp^3d^2
(c)	SO_3^{2-}	AX_3E	4 groups	sp^3
(d)	BrO_2^-	AX_2E_2	4 groups	sp^3

56. *Refer to Section 7.4 and Table 7.2.*

Draw the Lewis structure for each of the molecules. Then determine the number of electron pairs (whether in a bond or as a lone pair) and the hybridization.

(a)  AX_4E_2 6 e^- pairs sp^3d^2

(b)  AX_6 6 e^- pairs sp^3d^2

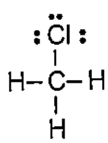
(c)  AX_4E 5 e^- pairs sp^3d

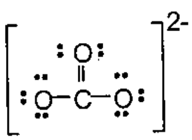
58. *Refer to Section 7.4 and Problem 26 (above).*

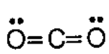
The Lewis structure for borazine was determined in problem 26. Each B and N is bonded to three other atoms and each has no lone pairs (AX_3). Thus, the hybridization is sp^2.

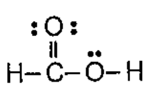
60. *Refer to Section 7.4 and Example 7.9.*

Draw the Lewis structure for each of the molecules. Then determine the number of electron pairs (shared and unshared) and the hybridization.

(a)  AX_4 4 e^- pairs sp^3

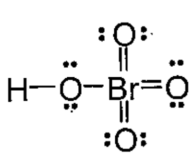
(b)  AX_3 3 e^- pairs sp^2

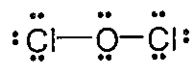
(c)  AX_2 2 e^- pairs sp

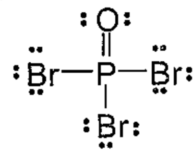
(d)  AX_3 3 e^- pairs sp^2

62. *Refer to Section 7.4 and Examples 7.9 and 7.10.*

Draw the Lewis structure for each of the molecules. Then determine the number of electron pairs (shared and unshared) and the hybridization.

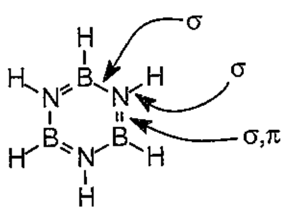
(a)  AX_4 4 e^- pairs sp^3

(b)  AX_2E_2 4 e^- pairs sp^3

(c)  AX_4 4 e^- pairs sp^3

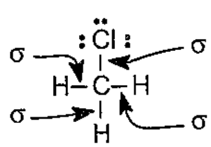
64. Refer to Section 7.4 and Example 7.11.

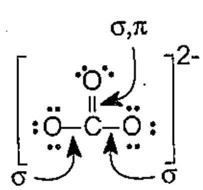
Recall that each single bond is a sigma (σ) bond and each double bond is composed of a sigma (σ) and a pi (π) bond.

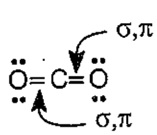


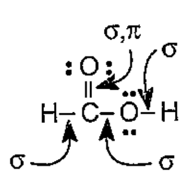
9 single bonds (9 σ) and 3 double bonds (3 σ, and 3 π) adds up to 12 σ-bonds and 3 π-bonds.

66. Refer to Section 7.4, Example 7.10, and Problem 60 (above).

(a)  4 σ-bonds

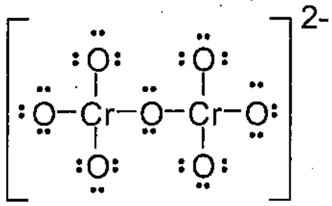
(b)  3 σ-bonds, 1 π-bond

(c)  2 σ-bonds, 2 π-bonds

(d)  4 σ-bonds, 1 π-bond

68. Refer to Section 7.1.

Write the skeleton structure from the information given. If there are no Cr-Cr bonds, then there must an O between the two Cr atoms. Since there are no O-O bonds, all O must be bonded to Cr. To be symmetrical, put three O on each Cr.



70. Refer to Sections 7.2, 7.3, and 7.4, Tables 7.3 and 7.4, and Figure 7.8.

For the formula AX_mE_n, m is the number of atoms around Central Atom A and n is the number of electron pairs around A. The geometry is determined by the "Species" formula, as indicated in Table 7.3 and Figure 7.8. The hybridization is also determined by the number of electron pairs (bonding and unshaired, m+n) as shown in Table 7.4. Finally, recall the atoms, X, must be symmetrically disposed around the central atom, A, for the molecule to be nonpolar.

Species	Atoms Around Central Atom A	Unshared Pairs Around A	Geometry	Hybridization	Polarity
AX_2E_2	2	2	bent	sp^3	polar
AX_3	3	0	trigonal planar	sp^2	nonpolar
AX_4E_2	4	2	square planar	sp^3d^2	nonpolar
AX_5	5	0	trigonal bipyramid	sp^3d	nonpolar

72. Refer to Section 7.1.

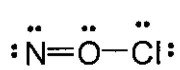
As a rule of thumb, the least electronegative atom will be central atom. This can be confirmed by drawing Lewis structures and determining which is the best (obeys all the rules and has the least formal charges).

(a) For HCN, recall that H only has single bonds, thus there are two possibilities for the Lewis Structure that obey the octet rule. When H has one bond, the formal charge is 0.

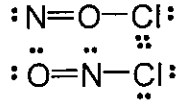
H—N≡C: In this structure, the most electronegative atom, N, has a +1 formal charge and C has a –1 formal charge.

H—C≡N: In this structure, both N and C have formal charges of 0, Thus, it is the best Lewis structure and **C** is the central atom.

(b) For NOCl, there are also two possibilities.



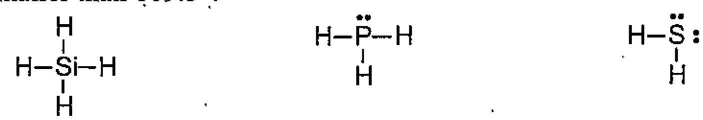
Formal charges: Cl: 0; O: +1; N: -1.



Formal charges: Cl:0; O: 0; N: 0. Thus, this is the best Lewis structure and **N** is the central atom.

74. Refer to Section 7.2.

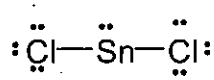
Draw Lewis structures of each molecule. Electron pairs in unshared pairs exert greater repulsion than electron pairs in bonds. Thus, the structures with unshared pairs will have the bond angles smaller than 109.5°.



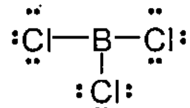
PH_3 and H_2S have at least one unshared pair and will have bond angles less than 109.5°.

76. Refer to Section 7.4.

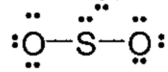
Draw Lewis structures of each molecule. Electron pairs in unshared pairs exert greater repulsion than electron pairs in bonds. Thus, the structures with unshared pairs will have the bond angles smaller than 120°.



One lone pair around Sn. Thus, the bond angle will be **less than 120°.**



No lone pairs around B. Thus, the bond angle will equal 120°.



One lone pair around S. Thus, the bond angle will be **less than 120°.**

78. Refer to Chapter 5 and Sections 7.2 - 7.3.

By calculating the moles of Cl and the total moles of F in the products, one can determine the mole ratio of Cl to F and thus the formula of ClF_x. Then one can proceed to address the questions of geometry and polarity.

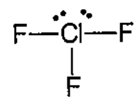
$$n = \frac{PV}{RT} = \frac{(3.00\,\text{atm})(0.457\,\text{L})}{(0.0821\,\text{L}\cdot\text{atm/mol.}\cdot\text{K})(348\,\text{K})} = 0.0480\,\text{mol. Cl}$$

$$5.63\,\text{g UF}_6 \times \frac{1\,\text{mol. UF}_6}{352\,\text{g UF}_6} \times \frac{6\,\text{mol. F}}{1\,\text{mol. UF}_6} = 0.0960\,\text{mol. F}$$

moles Cl = 0.0480 moles.
moles F = 0.0480 + 0.0960 = 0.1440 moles.

$$\frac{0.1440\,\text{mol. F}}{0.0480\,\text{mol. Cl}} = 3\,\text{mol. F/1 mol. Cl}$$

Thus, $x = 3$ ClF_3



Geometry: AX_3E_2 **T-shaped.**

This molecule is **polar**.

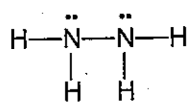
Bond angles are about **90°** and **180°**.

Cl: **sp^3d** hybridized

3 σ bonds, **0 π** bonds.

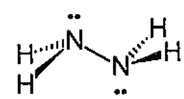
79. Refer to Sections 7.1, 7,2, and 7.4.

Lewis Structure:



Each N is AX_3E, and thus would be **trigonal pyramid.** Due to the lone pairs of electrons on the nitrogens, the bond angles will be slightly less than 109.5°. As drawn, the dipoles are not balanced and thus the molecule is polar.

When one considers rotation about the N-N bond, the dipoles cancel and the molecule is nonpolar.

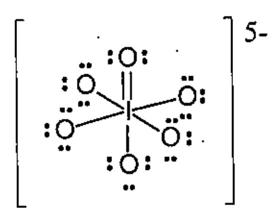


80. Refer to Sections 7.1 - 7.4.

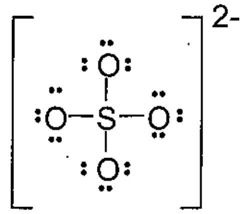
There are 6 groups around the central iodine atom.

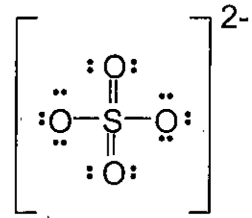
The hybridization is sp^3d^2.

The geometry is octahedron.



81. Refer to Sections 7.1 - 7.4.

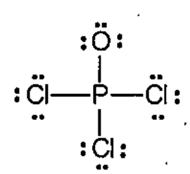


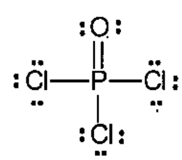


Geometry:	AX_4: tetrahedron	AX_4: tetrahedron
Hybridization:	sp^3	sp^3
C_f(S):	6 - (0 + ½(8)) = 2	6 - (0 + ½(12)) = 0
C_f(O) (with single bond):	6 - (6 + ½(2)) = -1	6 - (6 + ½(2)) = -1
C_f(O) (with double bond):		6 - (4 + ½(4)) = 0

Since the atoms in the second structure have charges closer to zero, it is the better structure.

82. Refer to Section 7.1.





C_f(P):	5 - (0 + 4) = +1	5 - (0 + 5) = 0
C_f(O):	6 - (6 + 1) = -1	6 - (4 + 2) = 0
C_f(Cl):	7 - (6 + 1) = 0	7 - (6 + 1) = 0

Since the second structure has no formal charges, it is the better structure.