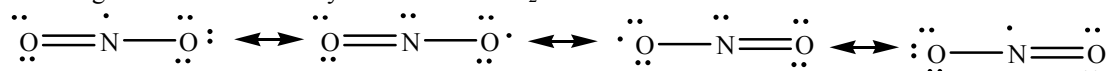


# CHAPTER 10 THE SHAPES OF MOLECULES

10.1 To be the central atom in a compound, the atom must be able to simultaneously bond to at least two other atoms. **He, F, and H** cannot serve as central atoms in a Lewis structure. Helium ( $1s^2$ ) is a noble gas, and as such, it does not need to bond to any other atoms. Hydrogen ( $1s^1$ ) and fluorine ( $1s^2 2s^2 2p^5$ ) only need one electron to complete their valence shells. Thus, they can only bond to one other atom, and they do not have  $d$  orbitals available to expand their valence shells.

10.2 Resonance must be present any time that a single Lewis structure is inadequate in explaining one or more aspects of a molecule or ion. The two N–O bonds in  $\text{NO}_2$  are equivalent; no single Lewis structure shows this. The following Lewis structures may be drawn for  $\text{NO}_2$ :



The average of all of these structures gives equivalent N–O bonds with a bond length that is between N–O and N=O.

10.3 For an element to obey the octet rule it must be surrounded by 8 electrons. To determine the number of electrons present (1) count the individual electrons actually shown adjacent to a particular atom, and (2) add two times the number of bonds to that atom. Using this method the structures shown give: (a)  $0 + 2(4) = 8$ ; (b)  $2 + 2(3) = 8$ ; (c)  $0 + 2(5) = 10$ ; (d)  $2 + 2(3) = 8$ ; (e)  $0 + 2(4) = 8$ ; (f)  $2 + 2(3) = 8$ ; (g)  $0 + 2(3) = 6$ ; (h)  $8 + 2(0) = 8$ . All the structures obey the octet rule except: c and g.

10.4 For an atom to expand its valence shell, it must have readily available  $d$  orbitals. The  $d$  orbitals do not become readily available until the third period or below on the periodic table. For the elements in the problem F, S, H, Al, Se, and Cl, the period numbers are 2, 3, 1, 3, 4, and 3, respectively. All of these elements, except those in the first two periods (H and F), can expand their valence shells.

10.5 Plan: Count the valence electrons and draw Lewis structures.

Solution:

Total valence electrons:  $\text{SiF}_4$ :  $[1 \times \text{Si}(4e^-)] + [4 \times \text{F}(7e^-)] = 32$ ;  $\text{SeCl}_2$ :  $[1 \times \text{Se}(6e^-)] + [2 \times \text{Cl}(7e^-)] = 20$ ;

$\text{COF}_2$ :  $[1 \times \text{C}(4e^-)] + [1 \times \text{O}(6e^-)] + [2 \times \text{F}(7e^-)] = 24$ . The Si, Se, and the C are the central atoms, because these are the elements in their respective compounds with the lower group number (in addition, we are told C is central). Place the other atoms around the central atoms and connect each to the central atom with a single bond.

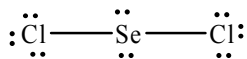
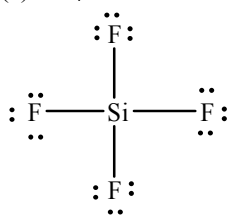
$\text{SiF}_4$ : At this point, 8 electrons ( $2e^-$  in 4 Si–F bonds) have been used with  $32 - 8 = 24$  remaining; the remaining electrons are placed around the fluorines (3 pairs each). All atoms have an octet.

$\text{SeCl}_2$ : The 2 bonds use  $4e^-$  ( $2e^-$  in 2 Se–Cl bonds) leaving  $20 - 4 = 16e^-$ . These  $16e^-$  are used to complete the octets on Se and the Cl atoms.

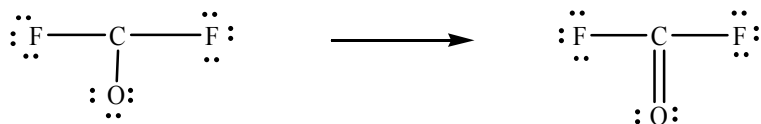
$\text{COF}_2$ : The 3 bonds to the C use  $6e^-$  ( $2e^-$  in 3 bonds) leaving  $24 - 6 = 18e^-$ . These  $18e^-$  are distributed to the surrounding atoms first to complete their octets. After the  $18e^-$  are used, the central C is 2 electrons short of an octet. Forming a double bond to the O (change a lone pair on O to a bonding pair on C) completes the C octet.

(a)  $\text{SiF}_4$

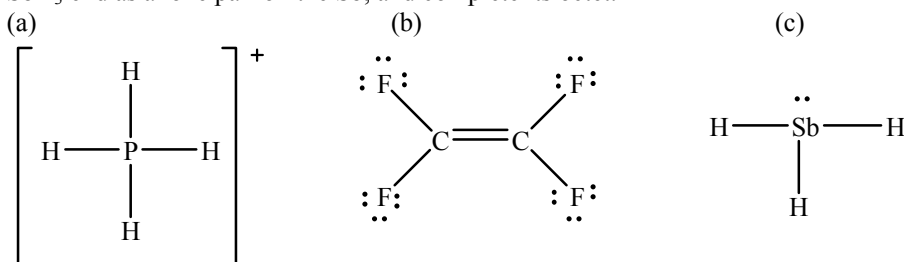
(b)  $\text{SeCl}_2$



(c)  $\text{COF}_2$



- 10.6 Total valence electrons:  $\text{PH}_4^+$  has 8;  $\text{C}_2\text{F}_4$  has 36; and  $\text{SbH}_3$  has 8. Ignoring H, the atom in the lower group number is central: P, C, and Sb. Added proof: H and F are never central. The two central C atoms must be adjacent. Place all the other atoms around the central atom. Split the F atoms so that each C gets 2. Connect all the atoms with single bonds. This uses all the electrons in  $\text{PH}_4^+$ , and gives P an octet. The H atoms need no additional electrons. The C atoms have 6 electrons each, but can achieve an octet by forming a double bond. Splitting the 24 remaining electrons in  $\text{C}_2\text{F}_4$  into 12 pairs and giving 3 pairs to each F leaves each F with an octet. The last 2 electrons in  $\text{SbH}_3$  end as a lone pair on the Sb, and complete its octet.



- 10.7 Plan: Count the valence electrons and draw Lewis structures.

Solution:

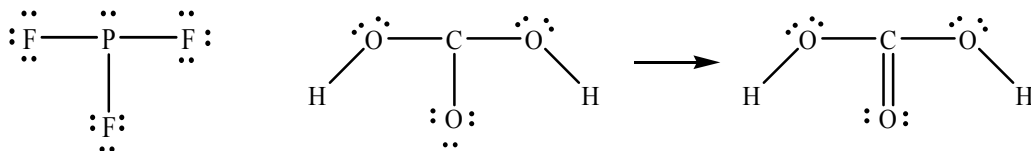
a)  $\text{PF}_3$ :  $[1 \times \text{P}(5e^-)] + [3 \times \text{F}(7e^-)] = 26$  valence electrons. P is the central atom. Draw single bonds from P to the 3 F atoms, using  $2e^- \times 3 \text{ bonds} = 6e^-$ . Remaining  $e^-$ :  $26 - 6 = 20e^-$ . Distribute the  $20e^-$  around the P and F atoms to complete their octets.

b)  $\text{H}_2\text{CO}_3$ :  $[2 \times \text{H}(1e^-)] + [1 \times \text{C}(4e^-)] + 3 \times \text{O}(6e^-) = 24$  valence electrons. C is the central atom with the H atoms attached to the O atoms. Place appropriate single bonds between all atoms using  $2e^- \times 5 \text{ bonds} = 10e^-$  so that  $24 - 10 = 14e^-$  remain. Use these  $14e^-$  to complete the octets of the O atoms (the H atoms already have their two electrons). After the  $14e^-$  are used, the central C is 2 electrons short of an octet. Forming a double bond to the O that does not have an H bonded to it (change a lone pair on O to a bonding pair on C) completes the C octet.

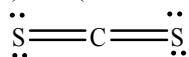
c)  $\text{CS}_2$ :  $[1 \times \text{C}(4e^-)] + [2 \times \text{S}(6e^-)] = 16$  valence electrons. C is the central atom. Draw single bonds from C to the 2 S atoms, using  $2e^- \times 2 \text{ bonds} = 4e^-$ . Remaining  $e^-$ :  $16 - 4 = 12e^-$ . Use these  $12e^-$  to complete the octets of the surrounding S atoms; this leaves C four electrons short of an octet. Form a double bond from each S to the C by changing a lone pair on each S to a bonding pair on C.

a)  $\text{PF}_3$  (26 valence  $e^-$ )

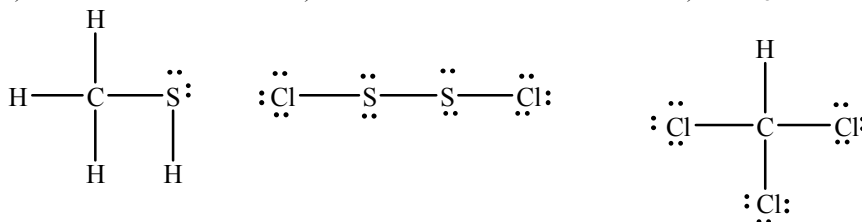
b)  $\text{H}_2\text{CO}_3$  (24 valence  $e^-$ )



c)  $\text{CS}_2$  (16 valence  $e^-$ )



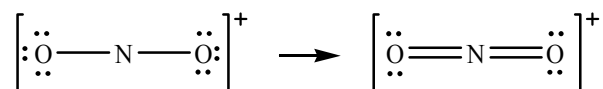
- 10.8 The C and S atoms are central. The S in part (a) is attached to an H and the C. All atoms are attached with single bonds and the remaining electrons are divided into lone pairs. All the atoms, except H, have octets.  
 a)  $\text{CH}_4\text{S}$                       b)  $\text{S}_2\text{Cl}_2$                       c)  $\text{CHCl}_3$



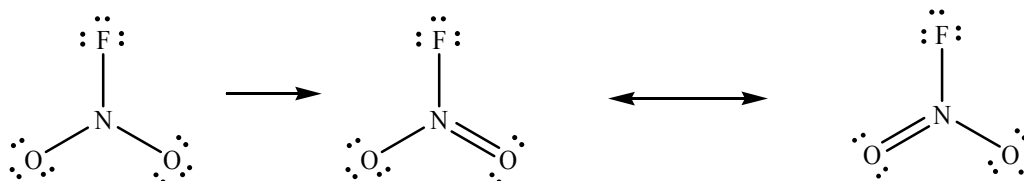
- 10.9 Plan: The problem asks for resonance structures, so there must be more than one answer for each part.

Solution:

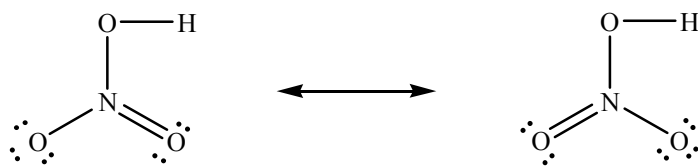
a)  $\text{NO}_2^+$  has  $[1 \times \text{N}(5 e^-)] + [2 \times \text{O}(6 e^-)] - 1 e^- (+ \text{charge}) = 16$  valence electrons. Draw a single bond from N to each O, using  $2e^- \times 2 \text{ bonds} = 4 e^-$ ;  $16 - 4 = 12 e^-$  remain. Distribute these  $12 e^-$  to the O atoms to complete their octets. This leaves N  $4 e^-$  short of an octet. Form a double bond from each O to the N by changing a lone pair on each O to a bonding pair on N. No resonance is required as all atoms can achieve an octet with double bonds.



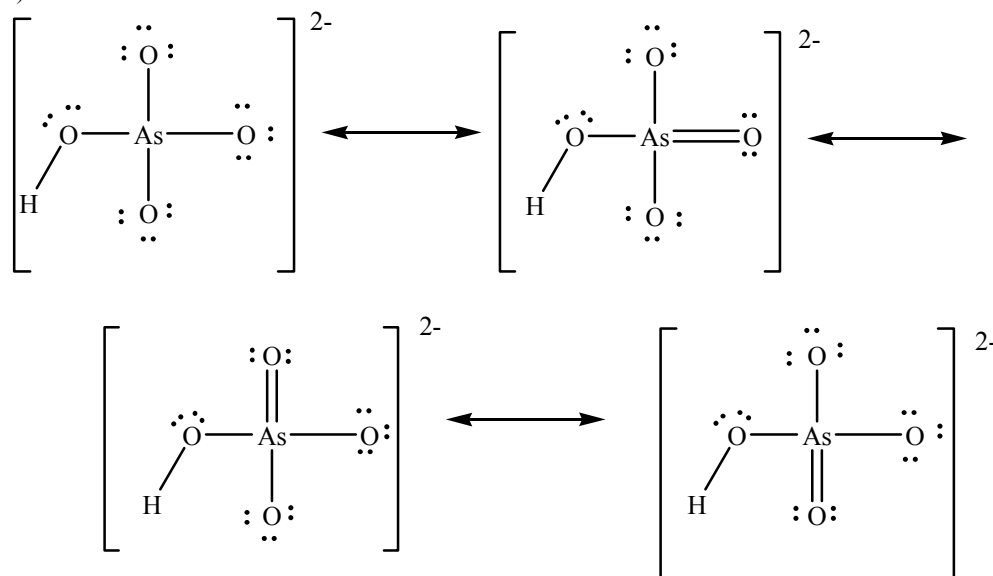
b)  $\text{NO}_2\text{F}$  has  $[1 \times \text{N}(5 e^-)] + [2 \times \text{O}(6 e^-)] + [1 \times \text{F}(7e^-)] = 24$  valence electrons. Draw a single bond from N to each surrounding atom, using  $2e^- \times 3 \text{ bonds} = 6 e^-$ ;  $24 - 6 = 18 e^-$  remain. Distribute these  $18 e^-$  to the O and F atoms to complete their octets. This leaves N  $2 e^-$  short of an octet. Form a double bond from either O to the N by changing a lone pair on O to a bonding pair on N. There are two resonance structures since a lone pair from either of the two O atoms can be moved to a bonding pair with N:



10.10 a)



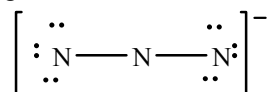
b)



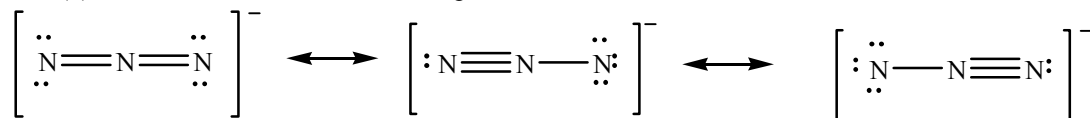
10.11 Plan: Count the valence electrons and draw Lewis structures. Additional structures are needed to show resonance.

Solution:

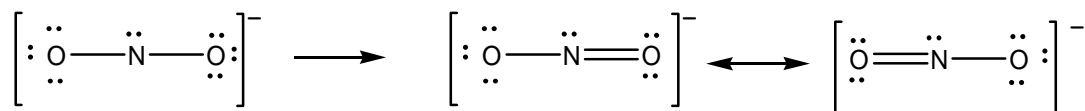
a)  $\text{N}_3^-$  has  $[3 \times \text{N}(5e^-)] + [1 e^- (\text{from charge})] = 16$  valence electrons. Place a single bond between the nitrogen atoms. This uses  $2e^- \times 2 \text{ bonds} = 4$  electrons, leaving  $16 - 4 = 12$  electrons (6 pairs). Giving 3 pairs on each end, nitrogen gives them an octet, but leaves the central N with only 4 electrons as shown below:



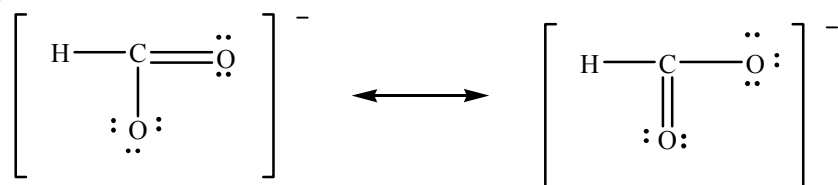
The central N needs 4 electrons. There are three options to do this: (1) each of the end N's could form a double bond to the central N by sharing one of its pairs; (2) one of the end N's could form a triple bond by sharing two of its lone pairs; (3) the other end N could form the triple bond instead.



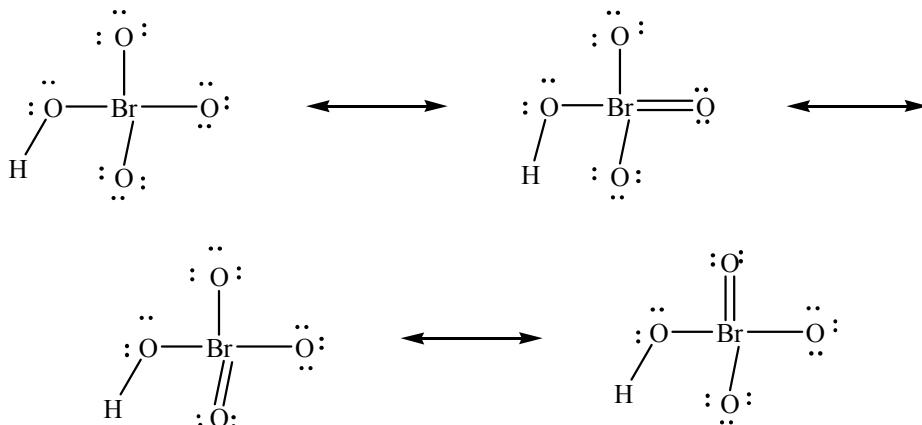
b)  $\text{NO}_2^-$  has  $[1 \times \text{N}(5e^-)] + [2 \times \text{O}(6 e^-)] + [1 e^- (\text{from charge})] = 18$  valence electrons. The nitrogen should be the central atom with each of the oxygen atoms attached to it by a single bond ( $2e^- \times 2 \text{ bonds} = 4$  electrons). This leaves  $18 - 4 = 14$  electrons (7 pairs). If 3 pairs are given to each O and 1 pair is given to the N, then both O's have an octet, but the N only has 6. To complete an octet the N needs to gain a pair of electrons from one O or the other (from a double bond). The resonance structures are:



10.12 a)  $\text{HCO}_2^-$  has 18 valence electrons.



b)  $\text{HBrO}_4$  has 32 valence electrons.

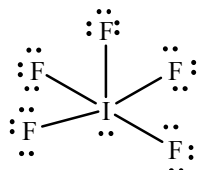


10.13 Plan: Initially, the method used in the preceding problems may be used to establish a Lewis structure. The total of the formal charges must equal the charge on an ion or be equal to 0 for a compound. The FC only needs to be calculated once for a set of identical atoms.

Solution:

a)  $\text{IF}_5$  has  $[1 \times \text{I}(7 e^-)] + [5 \times \text{F}(7 e^-)] = 42$  valence electrons.

The presence of 5 F atoms around the central I means that the I will have a minimum of 10 electrons; thus, this is an exception to the octet rule. The 5 I-F bonds use  $2e^- \times 5 \text{ bonds} = 10$  electrons leaving  $42 - 10 = 32$  electrons (16 pairs). Each F needs 3 pairs to complete an octet. The 5 F atoms use 15 of the 16 pairs, so there is 1 pair left for the central I. This gives:



Calculating formal charges:

$\text{FC} = \text{valence electrons} - (\text{lone electrons} + 1/2 (\text{bonded electrons}))$

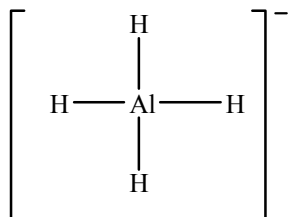
For iodine:  $\text{FC}_\text{I} = 7 - (2 + 1/2(10)) = 0$

For each fluorine:  $\text{FC}_\text{F} = 7 - (6 + 1/2(2)) = 0$

Total formal charge = 0 = charge on the compound.

b)  $\text{AlH}_4^-$  has  $[1 \times \text{Al}(3 e^-)] + [4 \times \text{H}(1 e^-)] + [1 e^- (\text{from charge})] = 8$  valence electrons.

The 4 Al-H bonds use all the electrons and Al has an octet.



$\text{FC} = \text{valence electrons} - (\text{lone electrons} + 1/2 (\text{bonded electrons}))$

For aluminum:  $\text{FC}_\text{Al} = 3 - (0 + 1/2(8)) = -1$

For each hydrogen  $\text{FC}_\text{H} = 1 - (0 + 1/2(2)) = 0$



10.17 Plan: The general procedure is similar to the preceding problems, plus the oxidation number determination.

Solution:

a)  $\text{BrO}_3^-$  has  $[1 \times \text{Br}(7 e^-)] + 3 \times \text{O}(6 e^-) + [1e^- \text{ (from charge)}] = 26$  valence electrons.

Placing the O atoms around the central Br and forming 3 Br–O bonds uses  $2e^- \times 3 \text{ bonds} = 6$  electrons and leaves  $26 - 6 = 20$  electrons (10 pairs). Placing 3 pairs on each O ( $3 \times 3 = 9$  total pairs) leaves 1 pair for the Br and yields structure I below. In structure I, all the atoms have a complete octet. Calculating formal charges:

$$\text{FC}_{\text{Br}} = 7 - (2 + 1/2(6)) = +2 \quad \text{FC}_{\text{O}} = 6 - (6 + 1/2(2)) = -1$$

The  $\text{FC}_{\text{O}}$  is acceptable, but  $\text{FC}_{\text{Br}}$  is larger than is usually acceptable. Forming a double bond between any one of the O atoms gives Structure II. Calculating formal charges:

$$\text{FC}_{\text{Br}} = 7 - (2 + 1/2(8)) = +1 \quad \text{FC}_{\text{O}} = 6 - (6 + 1/2(2)) = -1 \quad \text{FC}_{\text{O}} = 6 - (4 + 1/2(4)) = 0$$

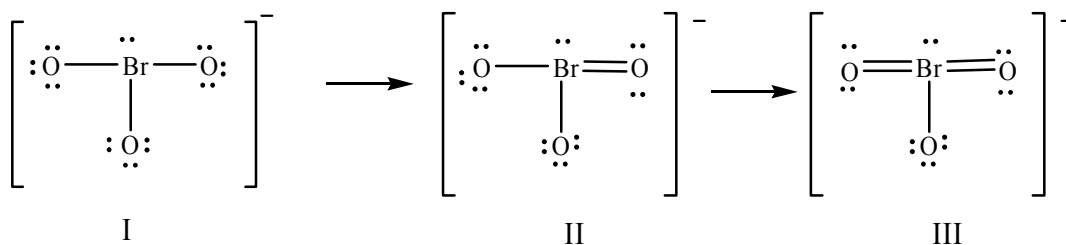
Double bonded O

The  $\text{FC}_{\text{Br}}$  can be improved further by forming a second double bond to one of the other O atoms (structure III).

$$\text{FC}_{\text{Br}} = 7 - (2 + 1/2(10)) = 0 \quad \text{FC}_{\text{O}} = 6 - (6 + 1/2(2)) = -1 \quad \text{FC}_{\text{O}} = 6 - (4 + 1/2(4)) = 0$$

Double bonded O atoms

Structure III has the most reasonable distribution of formal charges.



The oxidation numbers (O.N.) are:  $\text{O.N.}_{\text{Br}} = +5$  and  $\text{O.N.}_{\text{O}} = -2$ . -6  
+5 -2

Check: The total formal charge equals the charge on the ion (-1).  $\text{BrO}_3^-$

b)  $\text{SO}_3^{2-}$  has  $[1 \times \text{S}(6 e^-)] + [3 \times \text{O}(6 e^-)] + [2e^- \text{ (from charge)}] = 26$  valence electrons.

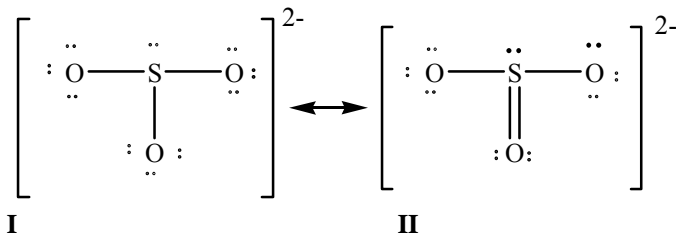
Placing the O atoms around the central S and forming 3 S–O bonds uses  $2e^- \times 3 \text{ bonds} = 6$  electrons and leaves  $26 - 6 = 20$  electrons (10 pairs). Placing 3 pairs on each O ( $3 \times 3 = 9$  total pairs) leaves 1 pair for the S and yields structure I below. In structure I all the atoms have a complete octet. Calculating formal charges:

$$\text{FC}_{\text{S}} = 6 - (2 + 1/2(6)) = +1; \text{FC}_{\text{O}} = 6 - (6 + 1/2(2)) = -1$$

The  $\text{FC}_{\text{O}}$  is acceptable, but  $\text{FC}_{\text{S}}$  is larger than is usually acceptable. Forming a double bond between any one of the O atoms (Structure II) gives:

$$\text{FC}_{\text{S}} = 6 - (2 + 1/2(8)) = 0 \quad \text{FC}_{\text{O}} = 6 - (6 + 1/2(2)) = -1 \quad \text{FC}_{\text{O}} = 6 - (4 + 1/2(4)) = 0$$

Double bonded O

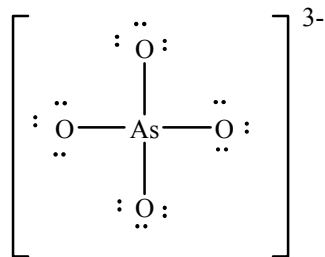


Structure II has the more reasonable distribution of formal charges. -6  
+4 -2

The oxidation numbers (O.N.) are:  $\text{O.N.}_{\text{S}} = +4$  and  $\text{O.N.}_{\text{O}} = -2$ .  $\text{SO}_3^{2-}$

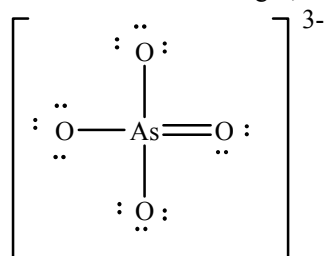
Check: The total formal charge equals the charge on the ion (-2).

- 10.18 a)  $\text{AsO}_4^{3-}$  has 32 valence electrons. See Structure I.  
 $\text{FC}_{\text{As}} = 5 - (0 + 1/2(8)) = +1$        $\text{FC}_{\text{O}} = 6 - (6 + 1/2(2)) = -1$  each  
 Net formal charge  $(+1 - 4) = -3$       The octet rule is followed by all atoms.



I

For more reasonable formal charges, move a lone pair from an O to a bonded pair on As (Structure II):



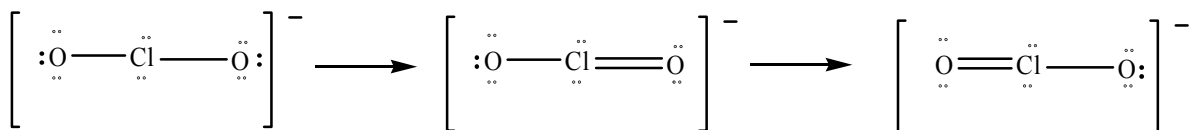
II

$\text{FC}_{\text{As}} = 5 - (0 + 1/2(10)) = 0$        $\text{FC}_{\text{O}} = 6 - (6 + 1/2(2)) = -1$  single bonds  
 $\text{FC}_{\text{O}} = 6 - (4 + 1/2(4)) = 0$  double bonded  
 Net formal charge:  $(0 + 3(-1)) + 0 = -3$       Improved formal charge distribution  
 O.N.: O  $-2$  each  $\times 4 = -8$  total; As  $+5$

b)  $\text{ClO}_2^-$  has 20 valence electrons. For Structure I in which all atoms have an octet:

$\text{FC}_{\text{Cl}} = 7 - (4 + 1/2(4)) = +1$        $\text{FC}_{\text{O}} = 6 - (6 + 1/2(2)) = -1$

For more reasonable formal charges, see Structure II:



I

II

Formal charges in Structure II:

$\text{FC}_{\text{Cl}} = 7 - (4 + 1/2(6)) = 0$

$\text{FC}_{\text{O}(\text{double bond})} = 6 - (4 + 1/2(4)) = 0$

$\text{FC}_{\text{O}(\text{single bond})} = 6 - (6 + 1/2(2)) = -1$

O.N.: O  $-2$  each  $\times 2 = -4$  total; Cl  $+3$

- 10.19 Plan: We know that each of these compounds does not obey the octet rule.

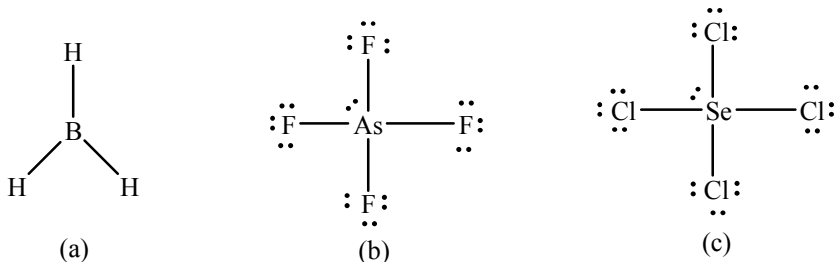
Solution:

a)  $\text{BH}_3$  has  $[1 \times \text{B}(3e^-)] + [3 \times \text{H}(1e^-)] = 6$  valence electrons. These are used in 3 B–H bonds. The B has 6 electrons instead of an octet; this molecule is electron deficient.

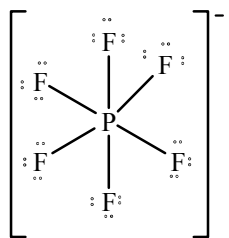
b)  $\text{AsF}_4^-$  has  $[1 \times \text{As}(5e^-)] + [4 \times \text{F}(7e^-)] + [1e^- \text{ (from charge)}] = 34$  valence electrons. Four As–F bonds use 8 electrons leaving  $34 - 8 = 26$  electrons (13 pairs). Each F needs 3 pairs to complete its octet and the remaining pair goes to the As. The As has an expanded octet with 10 electrons. The F cannot expand its octet.



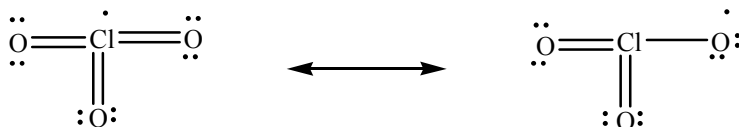
c)  $\text{SeCl}_4$  has  $[1 \times \text{Se}(6e^-)] + 4 \times \text{Cl}(7e^-) = 34$  valence electrons. The  $\text{SeCl}_4$  is isoelectronic (has the same electron structure) as  $\text{AsF}_4^-$ , and so its Lewis structure looks the same. Se has an expanded octet of 10 electrons.



10.20 a)  $\text{PF}_6^-$  has 48 valence electrons. P has an expanded outer shell to 12  $e^-$ .

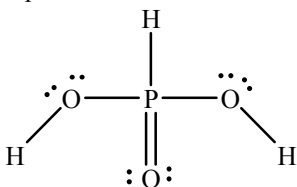


b)  $\text{ClO}_3$  has 25 valence electrons. The odd number means that there will be an exception. This is a radical: the chlorine or one of the oxygens will lack an  $e^-$  to complete its octet.



There are 2 additional resonance structures where the other O atoms are the ones lacking the octet. The FC predicts that Cl will end with the odd electron.

c)  $\text{H}_3\text{PO}_3$  has 26 valence electrons. To balance the formal charges; the O lacking an H will form a double bond to the P. This compound is an exception in that one of the H atoms is attached to the central P. P has an expanded outer shell to 10  $e^-$ .



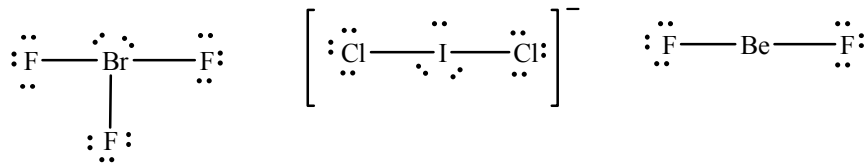
10.21 Plan: We know that each of these compounds does not obey the octet rule. Draw the Lewis structure.

Solution:

a)  $\text{BrF}_3$  has  $[1 \times \text{Br}(7e^-)] + [3 \times \text{F}(7e^-)] = 28$  valence electrons. Placing a single bond between Br and each F uses  $2e^- \times 3 \text{ bonds} = 6e^-$ , leaving  $28 - 6 = 22$  electrons (11 pairs). After the F atoms complete their octets with 3 pairs each, the Br gets the last 2 lone pairs. The Br expands its octet to 10 electrons.

b)  $\text{ICl}_2^-$  has  $[1 \times \text{I}(7e^-)] + [2 \times \text{Cl}(7e^-)] + [1e^- \text{ (from charge)}] = 22$  valence electrons. Placing a single bond between I and each Cl uses  $2e^- \times 2 \text{ bond} = 4e^-$ , leaving  $22 - 4 = 18$  electrons (9 pairs). After the Cl atoms complete their octets with 3 pairs each, the iodine finishes with the last 3 lone pairs. The iodine has an expanded octet of 10 electrons.

c)  $\text{BeF}_2$ ,  $[1 \times \text{Be}(2e^-)] + [2 \times \text{F}(7e^-)] =$  has 16 valence electrons. Placing a single bond between Be and each of the F atoms uses  $2e^- \times 2 \text{ bonds} = 4e^-$ , leaving  $16 - 4 = 12$  electrons (6 pairs). The F atoms complete their octets with 3 pairs each, and there are no electrons left for the Be. Formal charges work against the formation of double bonds. Be, with only 4 electrons, is electron deficient.

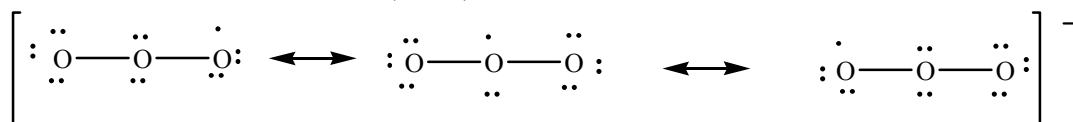
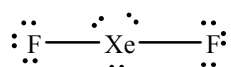
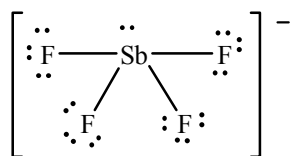


(a)

(b)

(c)

- 10.22 a)  $\text{O}_3^-$  has 19 valence electrons (note the odd number). There are several resonance structures possible, only one is necessary for the answer. One of the O's has the odd electron (7 total).

b)  $\text{XeF}_2$  has 22 valence electrons.Xe has an expanded outer shell of  $10 e^-$ .c)  $\text{SbF}_4^-$  has 34 valence electrons.Sb has an expanded outer shell of  $10 e^-$ .

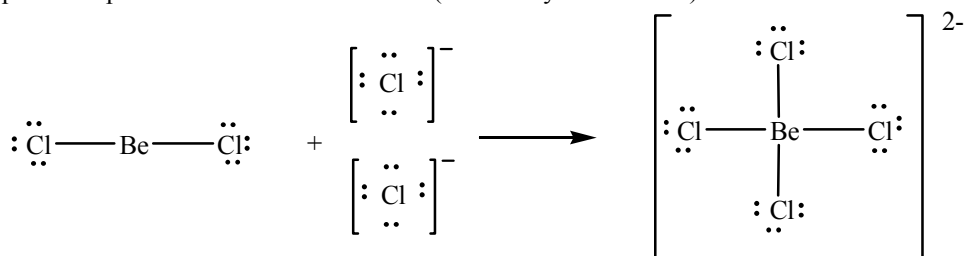
- 10.23 Plan: Draw Lewis structures for the reactants and products.

Solution:

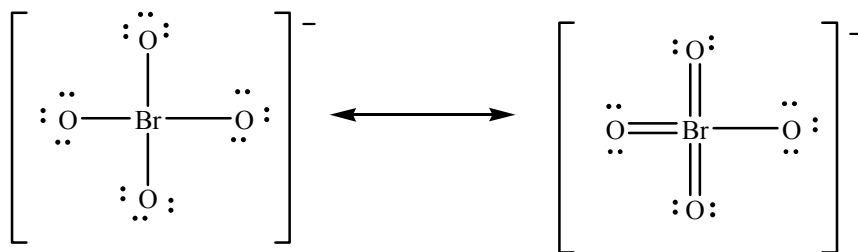
Beryllium chloride has the formula  $\text{BeCl}_2$ .  $\text{BeCl}_2$  has  $[1 \times \text{Be}(2e^-)] + [2 \times \text{Cl}(7e^-)] = 16$  valence electrons. 4 of these electrons are used to place a single bond between Be and each of the Cl atoms, leaving  $16 - 4 = 12$  electrons (6 pairs). These 6 pairs are used to complete the octets of the Cl atoms, but Be does not have an octet – it is electron deficient.

Chloride ion has the formula  $\text{Cl}^-$  with an octet of electrons.

$\text{BeCl}_4^{2-}$  has  $[1 \times \text{Be}(2e^-)] + [4 \times \text{Cl}(7e^-)] + [2e^- \text{ (from charge)}] = 32$  valence electrons. 8 of these electrons are used to place a single bond between Be and each Cl atom, leaving  $32 - 8 = 24$  electrons (12 pairs). These 12 pairs complete the octet of the Cl atoms (Be already has an octet).



- 10.24 Draw a Lewis structure. If the formal charges are not ideal, a second structure may be needed.  $\text{BrO}_4^-$  has 32 valence electrons.



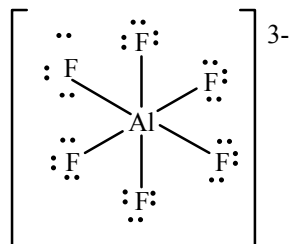
In the structure on the left, all atoms have octets. The formal charges are:

$$FC_{\text{Br}} = 7 - (0 + 1/2(8)) = +3 \quad FC_{\text{O}} = 6 - (6 + 1/2(2)) = -1$$

The structure on the right expands the valence shell of the Br to give more favorable formal charges.

$$FC_{\text{Br}} = 7 - (0 + 1/2(14)) = 0 \quad FC_{\text{O}} = 6 - (6 + 1/2(2)) = -1 \quad FC_{\text{O}(\text{double bonded})} = 6 - (4 + 1/2(4)) = 0$$

- 10.25 Count the total valence electrons and draw a Lewis structure.  $\text{AlF}_6^{3-}$  has 48 valence electrons.



- 10.26 Plan: Use the structures in the text to determine the formal charges.

$$FC = \text{valence electrons} - (\text{lone electrons} + 1/2 (\text{bonded electrons}))$$

Solution:

$$\text{Structure A: } FC_{\text{C}} = 4 - (0 + 1/2(8)) = 0; FC_{\text{O}} = 6 - (4 + 1/2(4)) = 0; FC_{\text{Cl}} = 7 - (6 + 1/2(2)) = 0$$

$$\text{Total FC} = 0$$

$$\text{Structure B: } FC_{\text{C}} = 4 - (0 + 1/2(8)) = 0; FC_{\text{O}} = 6 - (6 + 1/2(2)) = -1;$$

$$FC_{\text{Cl}} = 7 - (4 + 1/2(4)) = +1 (\text{double bonded}); FC_{\text{Cl}} = 7 - (6 + 1/2(2)) = 0 (\text{single bonded})$$

$$\text{Total FC} = 0$$

$$\text{Structure C: } FC_{\text{C}} = 4 - (0 + 1/2(8)) = 0; FC_{\text{O}} = 6 - (6 + 1/2(2)) = -1;$$

$$FC_{\text{Cl}} = 7 - (4 + 1/2(4)) = +1 (\text{double bonded}); FC_{\text{Cl}} = 7 - (6 + 1/2(2)) = 0 (\text{single bonded})$$

$$\text{Total FC} = 0$$

Structure A has the most reasonable set of formal charges.

- 10.27 Determine the total number of valence electrons present. Next, draw a Lewis structure. Finally, use VSEPR or Valence-Bond theory to predict the shape.

- 10.28 The molecular shape and the electron-group arrangement are the same when there are no lone pairs on the central atom.

- 10.29 A bent (V-shaped) molecule will have the stoichiometry  $\text{AX}_2$ , so only  $\text{AX}_2\text{E}_n$  geometries result in a bent molecule. The presence of one or two lone pairs in the **three and four electron group arrangements** can produce a bent (V-shaped) molecule as either  $\text{AX}_2\text{E}$  or  $\text{AX}_2\text{E}_2$ . Examples are:  $\text{NO}_2^-$  and  $\text{H}_2\text{O}$ .



$$120^\circ$$

$$\text{AX}_2\text{E}$$

$$109.5^\circ$$

$$\text{AX}_2\text{E}_2$$

- 10.30 Plan: Examine a list of all possible structures, and choose the ones with four electron groups since the tetrahedral electron group arrangement has four electron groups.

Solution:

Tetrahedral  $\text{AX}_4$

Trigonal pyramidal  $\text{AX}_3\text{E}$

Bent or V-shaped  $\text{AX}_2\text{E}_2$

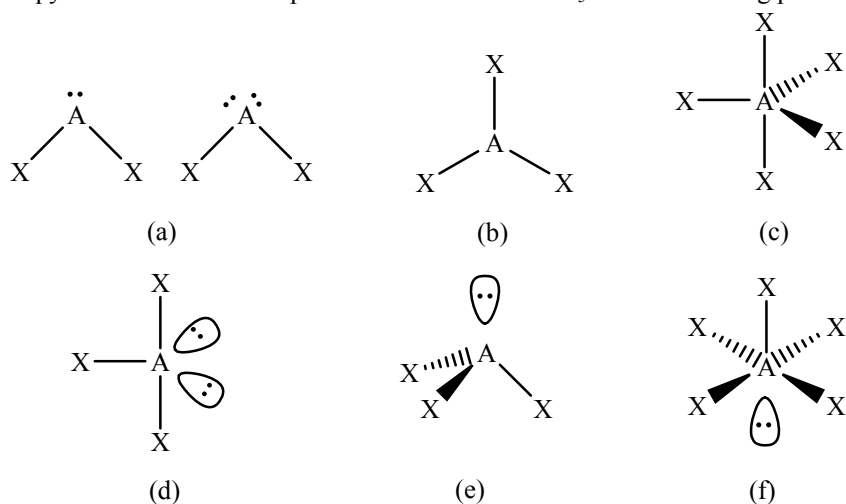
- 10.31 a) **A**, which has a square planar molecular geometry, has the most electron pairs. There are 4 shared pairs and two unshared pairs for a total of six pairs of electrons. The six electron pairs are arranged in an octahedral arrangement with the four bonds in a square planar geometry. B and C have five electron pairs and D has four electron pairs.

- b) **A** has the most unshared pairs around the central atom with two unshared pairs. **B** has only one unshared pair on the central atom and **C** and **D** have no unshared pairs on the central atom.  
 c) **C and D** have only shared pairs around the central atom.

10.32 Plan: Begin with the basic structures and redraw them.

Solution:

- a) A molecule that is V-shaped has two bonds and generally has either one ( $AX_2E$ ) or two ( $AX_2E_2$ ) lone electron pairs.  
 b) A trigonal planar molecule follows the formula  $AX_3$  with three bonds and no lone electron pairs.  
 c) A trigonal bipyramidal molecule contains five bonding pairs (single bonds) and no lone pairs ( $AX_5$ ).  
 d) A T-shaped molecule has three bonding groups and two lone pairs ( $AX_3E_2$ ).  
 e) A trigonal pyramidal molecule follows the formula  $AX_3E$  with three bonding pairs and one lone pair.  
 f) A square pyramidal molecule shape follows the formula  $AX_5E$  with 5 bonding pairs and one lone pair.



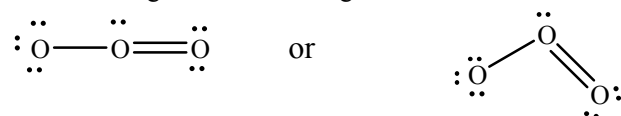
10.33 Determine the geometry from the lone pairs and the number of groups attached to the central atom.

|              |                      |                                  |         |
|--------------|----------------------|----------------------------------|---------|
| a) $AX_3E$   | tetrahedral          | $109.5^\circ$                    | smaller |
| b) $AX_2$    | linear               | $180^\circ$                      | none    |
| c) $AX_3$    | trigonal planar      | $120^\circ$                      | none    |
| d) $AX_2E_2$ | tetrahedral          | $109.5^\circ$                    | smaller |
| e) $AX_2$    | linear               | $180^\circ$                      | none    |
| f) $AX_4E$   | trigonal bipyramidal | $180^\circ, 120^\circ, 90^\circ$ | smaller |

10.34 Plan: First, draw a Lewis structure, and then apply VSEPR.

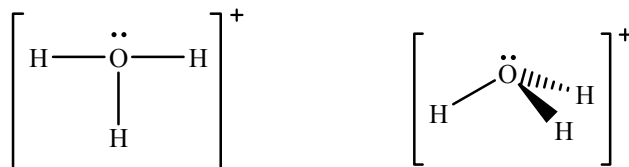
Solution:

- a)  $O_3$ : The molecule has  $[3 \times O(6e^-)] = 18$  valence electrons. 4 electrons are used to place a single bond between the oxygen atoms, leaving  $18 - 4 = 14e^-$  (7 pairs). 6 pairs are required to give the end oxygen atoms an octet; the last pair is distributed to the central oxygen, leaving this atom 2 electrons short of an octet. Form a double bond from one of the end O atoms to the central O by changing a lone pair on the end O to a bonding pair on the central O. This gives the following Lewis structure:



There are three electron groups around the central O, one of which is a lone pair. This gives a **trigonal planar** electron-group arrangement ( $AX_2E$ ), a **bent** molecular shape, and an ideal bond angle of  $120^\circ$ .

- b)  $H_3O^+$ : This ion has  $[3 \times H(1e^-)] + [1 \times O(6e^-)] - [1e^- \text{ (due to + charge)}] = 8$  valence electrons. 6 electrons are used to place a single bond between O and each H, leaving  $8 - 6 = 2e^-$  (1 pair). Distribute this pair to the O atom, giving it an octet (the H atoms only get 2 electrons). This gives the following Lewis structure:



There are four electron groups around the O, one of which is a lone pair. This gives a **tetrahedral** electron-group arrangement ( $AX_3E$ ), a **trigonal pyramidal** molecular shape, and an ideal bond angle of **109.5°**.

c)  $NF_3$ : The molecule has  $[1 \times N(5e^-)] + [3 \times F(7e^-)] = 26$  valence electrons. 6 electrons are used to place a single bond between N and each F, leaving  $26 - 6 = 20 e^-$  (10 pairs). These 10 pairs are distributed to all of the F atoms and the N atoms to give each atom an octet. This gives the following Lewis structure:



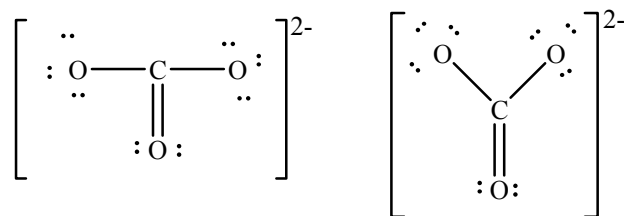
There are four electron groups around the N, one of which is a lone pair. This gives a **tetrahedral** electron-group arrangement ( $AX_3E$ ), a **trigonal pyramidal** molecular shape, and an ideal bond angle of **109.5°**.

| 10.35                                         | Lewis Structure | Electron-group Arrangement | Molecular Shape    | Ideal Bond Angle |
|-----------------------------------------------|-----------------|----------------------------|--------------------|------------------|
| (a)                                           |                 | Tetrahedral                | Tetrahedral        | 109.5°           |
| (b)                                           |                 | Trigonal Planar            | Bent               | 120°             |
| In addition, there are other resonance forms. |                 |                            |                    |                  |
| (c)                                           |                 | Tetrahedral                | Trigonal Pyramidal | 109.5°           |

10.36 **Plan:** First, draw a Lewis structure, and then apply VSEPR.

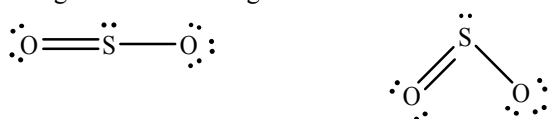
**Solution:**

(a)  $CO_3^{2-}$ : This ion has  $[1 \times C(4e^-)] + [3 \times O(6e^-)] + [2e^- \text{ (from charge)}] = 24$  valence electrons. 6 electrons are used to place single bonds between C and each O atom, leaving  $24 - 6 = 18 e^-$  (9 pairs). These 9 pairs are used to complete the octets of the three O atoms, leaving C 2 electrons short of an octet. Form a double bond from one of the O atoms to C by changing a lone pair on an O to a bonding pair on C. This gives the following Lewis structure:



+ 2 additional resonance forms. There are three groups of electrons around the C, none of which are lone pairs. This gives a **trigonal planar** electron-group arrangement ( $AX_3$ ), a **trigonal planar** molecular shape, and an ideal bond angle of  $120^\circ$ .

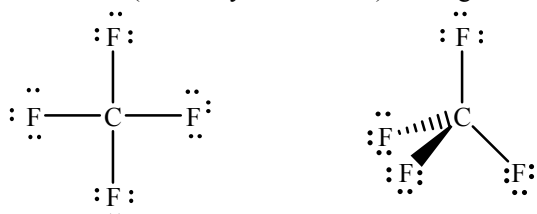
(b)  $SO_2$ : This molecule has  $[1 \times S(6 e^-)] + [2 \times S(6 e^-)] = 18$  valence electrons. 4 electrons are used to place a single bond between S and each O atom, leaving  $18 - 4 = 14 e^-$  (7 pairs). 6 pairs are needed to complete the octets of the O atoms, leaving a pair of electrons for S. S needs one more pair to complete its octet. Form a double bond from one of the end O atoms to the S by changing a lone pair on the O to a bonding pair on the S. This gives the following Lewis structure:



There are three groups of electrons around the C, one of which is a lone pair.

This gives a **trigonal planar** electron-group arrangement ( $AX_2E$ ), a **bent (V-shaped)** molecular shape, and an ideal bond angle of  $120^\circ$ .

(c)  $CF_4$ : This molecule has  $[1 \times C(4 e^-)] + [4 \times F(7 e^-)] = 32$  valence electrons. 8 electrons are used to place a single bond between C and each F, leaving  $32 - 8 = 24 e^-$  (12 pairs). Use these 12 pairs to complete the octets of the F atoms (C already has an octet). This gives the following Lewis structure:

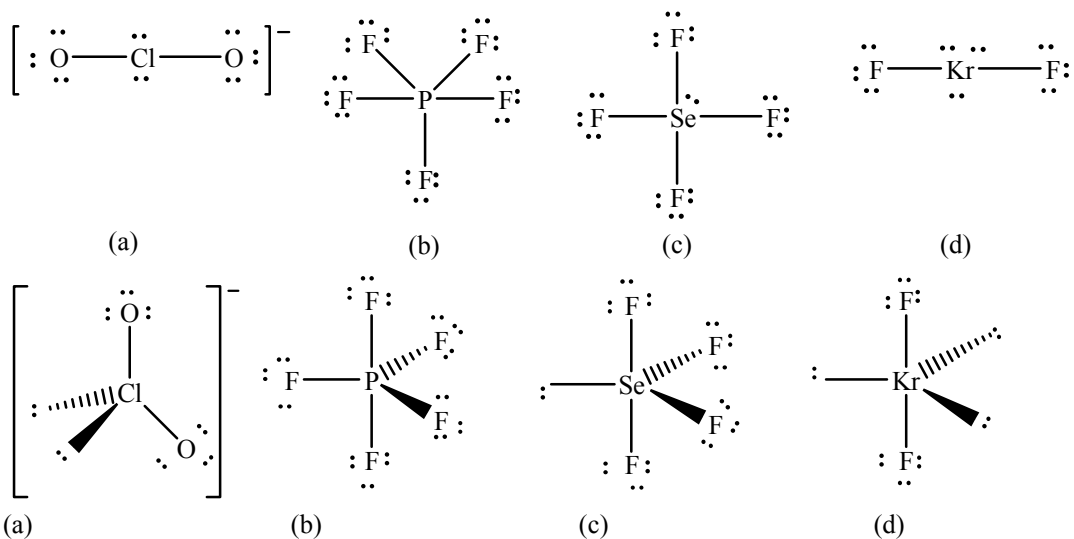


There are four groups of electrons around the C, none of which is a lone pair.

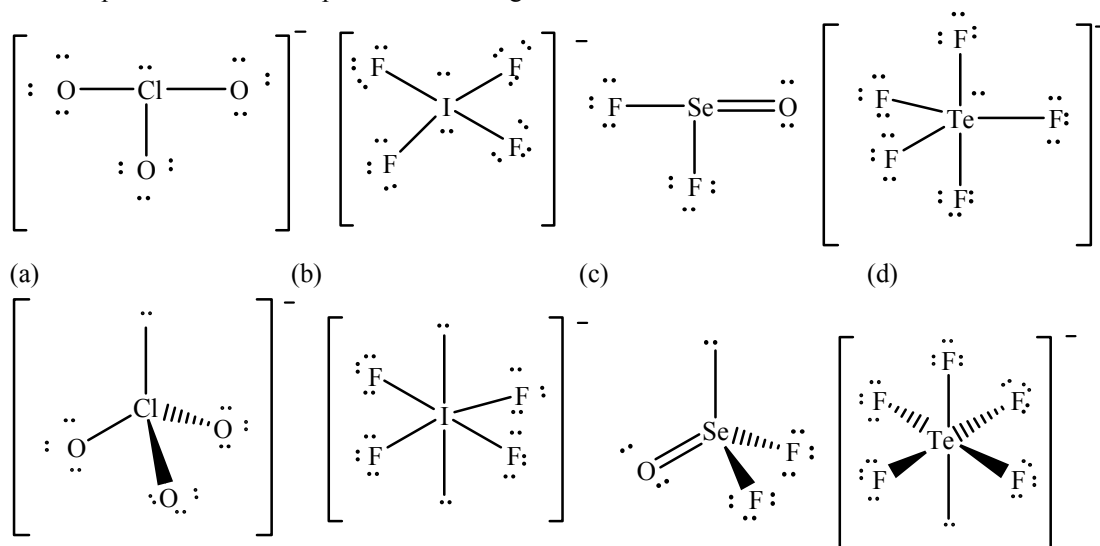
This gives a **tetrahedral** electron-group arrangement ( $AX_4$ ), a **tetrahedral** molecular shape, and an ideal bond angle of  $109.5^\circ$ .

| 10.37 | Lewis Structure | Electron-group Arrangement | Molecular Shape | Ideal Bond Angle |
|-------|-----------------|----------------------------|-----------------|------------------|
|       |                 | Trigonal Planar            | Trigonal Planar | $120^\circ$      |
|       |                 | Linear                     | Linear          | $180^\circ$      |
|       |                 | Tetrahedral                | Tetrahedral     | $109.5^\circ$    |

- 10.38 Plan: Examine the structure shown, and then apply VSEPR.  
Solution:  
 a) This structure shows three electron groups with three bonds around the central atom. There appears to be no distortion of the bond angles so the shape is **trigonal planar**, the classification is **AX<sub>3</sub>**, with an ideal bond angle of **120°**.  
 b) This structure shows three electron groups with three bonds around the central atom. The bonds are distorted down indicating the presence of a lone pair. The shape of the molecule is **trigonal pyramidal** and the classification is **AX<sub>3</sub>E**, with an ideal bond angle of **109.5°**.  
 c) This structure shows five electron groups with five bonds around the central atom. There appears to be no distortion of the bond angles so the shape is **trigonal bipyramidal** and the classification is **AX<sub>5</sub>**, with ideal bond angles of **90°** and **120°**.
- 10.39 a) This structure shows five electron groups with five bonds around the central atom. There appears to be no distortion of the bond angles so the shape is **square pyramidal**, (in reality square pyramidal structures have a slight distortion of the bond angles because there is a lone pair across from the atom at the apex of the pyramid). The classification is **AX<sub>5</sub>E**, with an ideal bond angle of **90°**.  
 b) This structure shows three electron groups with three bonds around the central atom. There appears to be no distortion of the bond angles so the shape is **T-shaped**, (in reality T-shaped structures have a slight distortion of the bond angles to the apical bonds because there are two equatorial lone pairs). The classification is **AX<sub>3</sub>E<sub>2</sub>**, with an ideal bond angle of **90°**.  
 c) This structure shows four electron groups with four bonds around the central atom. There appears to be no distortion of the bond angles so the shape is **tetrahedral** and the classification is **AX<sub>4</sub>**, with ideal bond angles of **109.5°**.
- 10.40 Plan: The Lewis structures must be drawn, and VSEPR applied to the structures.  
Solution:  
 a) The ClO<sub>2</sub><sup>-</sup> ion has [1 x Cl(7 e<sup>-</sup>) + 2 x O(6 e<sup>-</sup>) + 1 e<sup>-</sup> (from charge)] = 20 valence electrons. 4 electrons are used to place a single bond between the Cl and each O, leaving 20 - 4 = 16 electrons (8 pairs). All 8 pairs are used to complete the octets of the Cl and O atoms. There are two bonds (to the O's) and two lone pairs on the Cl for a total of 4 electron groups (AX<sub>2</sub>E<sub>2</sub>). The structure is based on a tetrahedral electron-group arrangement with an ideal bond angle of **109.5°**. The shape is **bent** (or V-shaped). The presence of the lone pairs will cause the remaining angles to be **less than 109.5°**.  
 b) The PF<sub>5</sub> molecule has [1 x P(5 e<sup>-</sup>) + 5 x F(7 e<sup>-</sup>)] = 40 valence electrons. 10 electrons are used to place single bonds between P and each F atom, leaving 40 - 10 = 30 e<sup>-</sup> (15 pairs). The 15 electrons are used to complete the octets of the F atoms. There are 5 bonds to the P and no lone pairs (AX<sub>5</sub>). The electron-group arrangement and the shape is **trigonal bipyramidal**. The ideal bond angles are **90° and 120°**. The absence of lone pairs means the **angles are ideal**.  
 c) The SeF<sub>4</sub> molecule has [1 x Se(6 e<sup>-</sup>) + 4 x F(7 e<sup>-</sup>)] = 34 valence electrons. 8 electrons are used to place single bonds between Se and each F atom, leaving 34 - 8 = 26 e<sup>-</sup> (13 pairs). 12 pairs are used to complete the octets of the F atoms which leaves 1 pair of electrons. This pair is placed on the central Se atom. There are 4 bonds to the Se which also has a lone pair (AX<sub>4</sub>E). The structure is based on a trigonal bipyramidal structure with ideal angles of **90° and 120°**. The shape is **see-saw**. The presence of the lone pairs means the angles are **less than ideal**.  
 d) The KrF<sub>2</sub> molecule has [1 x Kr(8 e<sup>-</sup>) + 2 x F(7 e<sup>-</sup>)] = 22 valence electrons. 4 electrons are used to place a single bond between the Kr atom and each F atom, leaving 22 - 4 = 18 e<sup>-</sup> (9 pairs). 6 pairs are used to complete the octets of the F atoms. The remaining 3 pairs of electrons are placed on the Kr atom. The Kr is the central atom. There are 2 bonds to the Kr and 3 lone pairs (AX<sub>2</sub>E<sub>3</sub>). The structure is based on a trigonal bipyramidal structure with ideal angles of 90° and 120°. The shape is **linear**. The placement of the F atoms makes their ideal bond angle to be 2 x 90° = **180°**. The placement of the lone pairs is such that they cancel each other's repulsion, thus the actual **bond angle is ideal**.



- 10.41 a) The  $\text{ClO}_3^-$  ion has 26 valence electrons. The Cl is the central atom. There are 3 bonds (to the O atoms) and 1 lone pair on the Cl ( $\text{AX}_3\text{E}$ ). The shape is **trigonal pyramidal**. The structure is based on a tetrahedral electron-group arrangement with an ideal bond angle of  $109.5^\circ$ . The presence of the lone pair will cause the remaining angles to be **less than  $109.5^\circ$** .
- b) The  $\text{IF}_4^-$  ion has 36 valence electrons. The I is the central atom. There are 4 bonds to the I and 2 lone pairs ( $\text{AX}_4\text{E}_2$ ). The shape is **square planar**. The structure is based on an octahedral electron-group arrangement with ideal bond angles of  $90^\circ$ . The repulsion from the two lone pairs cancels so the **angles are ideal**.
- c) The  $\text{SeOF}_2$  molecule has 26 valence electrons. The Se is the central atom. There are 3 bonds to the Se which also has a lone pair ( $\text{AX}_3\text{E}$ ). The shape is **trigonal pyramidal**. The structure is based on a tetrahedral structure with ideal angles of  $109.5^\circ$ . The presence of the lone pair means the angles are **less than ideal**.
- d) The  $\text{TeF}_5^-$  ion has 42 valence electrons. The Te is the central atom. There are 5 bonds to the Te which also has 1 lone pair ( $\text{AX}_5\text{E}$ ). The shape is **square pyramidal**. The structure is based on an octahedral with ideal angles of  $90^\circ$ . The presence of the lone pair means the angles are **less than ideal**.

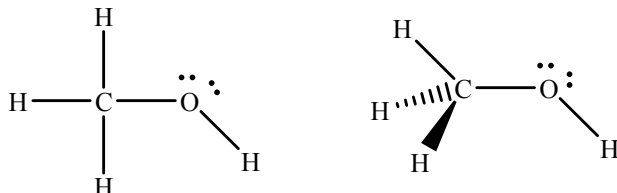




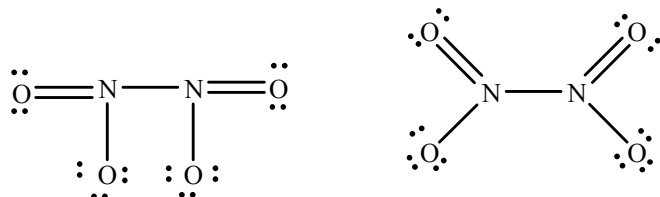
10.42 Plan: The Lewis structures must be drawn, and VSEPR applied to the structures.

Solution:

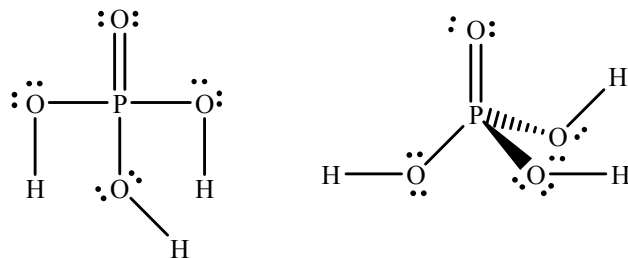
a)  $\text{CH}_3\text{OH}$ : This molecule has  $[1 \times \text{C}(4 e^-)] + [4 \times \text{H}(1 e^-)] + [1 \times \text{O}(6 e^-)] = 14$  valence electrons. In the  $\text{CH}_3\text{OH}$  molecule, both carbon and oxygen serve as central atoms. (H can never be central.) Use 8 electrons to place a single bond between the C and the O atom and 3 of the H atoms and another 2 electrons to place a single bond between the O and the last H atom. This leaves  $14 - 10 = 4 e^-$  (2 pairs). Use these two pairs to complete the octet of the O atom. C already has an octet and each H only gets 2 electrons. The carbon has 4 bonds and no lone pairs ( $\text{AX}_4$ ), so it is **tetrahedral** with **no deviation** (no lone pairs) from the ideal angle of  $109.5^\circ$ . The oxygen has 2 bonds and 2 lone pairs ( $\text{AX}_2\text{E}_2$ ), so it is **V-shaped** or **bent** with the angles **less than the ideal** angle of  $109.5^\circ$ .



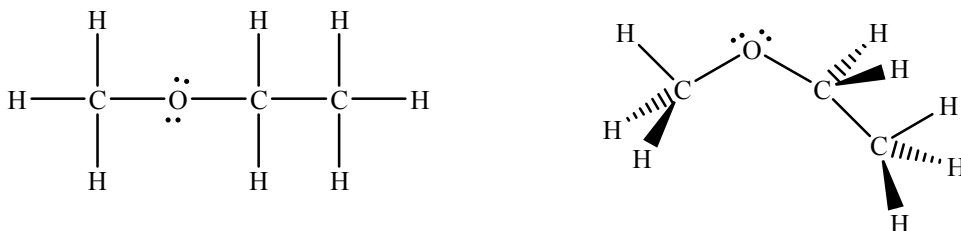
b)  $\text{N}_2\text{O}_4$ : This molecule has  $[2 \times \text{N}(5 e^-)] + [4 \times \text{O}(6 e^-)] = 34$  valence electrons. Use 10 electrons to place a single bond between the two N atoms and between each N and two of the O atoms. This leaves  $34 - 10 = 24 e^-$  (12 pairs). Use the 12 pairs to complete the octets of the oxygen atoms. Neither N atom has an octet however. Form a double bond from one O atom to one N atom by changing a lone pair on the O to a bonding pair on the N. Do this for the other N atom as well. In the  $\text{N}_2\text{O}_4$  molecule, both nitrogens serve as central atoms. This is the arrangement given in the problem. Both nitrogens are equivalent with 3 groups and no lone pairs ( $\text{AX}_3$ ), so the arrangement is **trigonal planar** with **no deviation** (no lone pairs) from the ideal angle of  $120^\circ$ . The same results arise from the other resonance structures.



10.43 a) In the  $\text{H}_3\text{PO}_4$  molecule the P and each of the O atoms with an H attached serve as central atoms. The P has 4 groups and no lone pairs ( $\text{AX}_4$ ), so it is **tetrahedral** with **no deviation** from the ideal angle of  $109.5^\circ$ . The H bearing O atoms have 2 bonds and 2 lone pairs ( $\text{AX}_2\text{E}_2$ ), so the arrangement is **V-shaped** or **bent** with angles **less than the ideal** values of  $109.5^\circ$ .



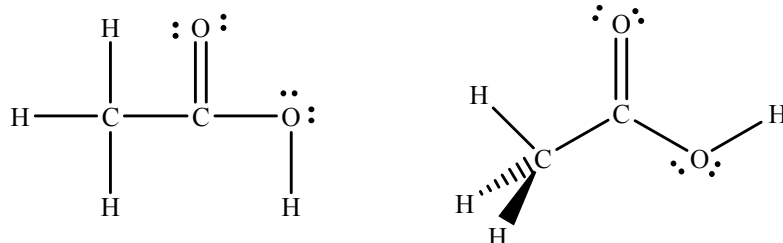
b) In the  $\text{CH}_3\text{OCH}_2\text{CH}_3$  molecule, all atoms except the hydrogens serve as central atoms. All the carbons have 4 bonds and no lone pairs ( $\text{AX}_4$ ), so they are tetrahedral with no deviation from the ideal bond angle of  $109.5^\circ$ . The oxygen has 2 bonds and 2 lone pairs ( $\text{AX}_2\text{E}_2$ ), so the arrangement is **V-shaped** or **bent** with angles **less than the ideal** values of  $109.5^\circ$ .



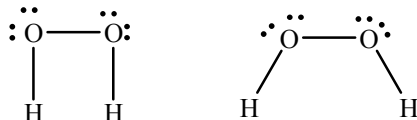
10.44 Plan: The Lewis structures must be drawn, and VSEPR applied to the structures.

Solution:

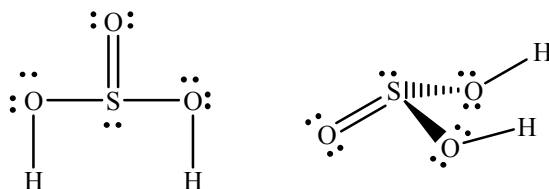
a)  $\text{CH}_3\text{COOH}$  has  $[2 \times \text{C}(4 e^-)] + [4 \times \text{H}(1 e^-)] + [2 \times \text{O}(6 e^-)] = 24$  valence electrons. Use 14 electrons to place a single bond between all of the atoms. This leaves  $24 - 14 = 10 e^-$  (5 pairs). Use these 5 pairs to complete the octets of the O atoms; the C atom bonded to the H atoms has an octet but the other C atom does not have a complete octet. Form a double bond from the O atom (not bonded to H) to the C by changing a lone pair on the O to a bonding pair on the C. In the  $\text{CH}_3\text{COOH}$  molecule, the carbons and the O with H attached serve as central atoms. The carbon bonded to the H atoms has 4 groups and no lone pairs ( $\text{AX}_4$ ), so it is **tetrahedral** with **no deviation** from the ideal angle of  $109.5^\circ$ . The carbon bonded to the O atoms has 3 groups and no lone pairs ( $\text{AX}_3$ ), so it is **trigonal planar** with **no deviation** from the ideal angle of  $120^\circ$ . The H bearing O has 2 bonds and 2 lone pairs ( $\text{AX}_2\text{E}_2$ ), so the arrangement is **V-shaped** or **bent** with an angle **less than the ideal** values of  $109.5^\circ$ .



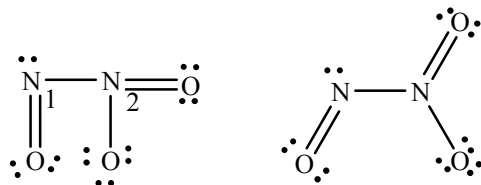
b)  $\text{H}_2\text{O}_2$  has  $[2 \times \text{H}(1 e^-)] + [2 \times \text{O}(6 e^-)] = 14$  valence electrons. Use 6 electrons to place single bonds between the O atoms and between each O atom and an H atom. This leaves  $14 - 6 = 8 e^-$  (4 pairs). Use these 4 pairs to complete the octets of the O atoms. In the  $\text{H}_2\text{O}_2$  molecule, both oxygens serve as central atoms. Both O's have 2 bonds and 2 lone pairs ( $\text{AX}_2\text{E}_2$ ), so they are **V-shaped** or **bent** with angles **less than the ideal** values of  $109.5^\circ$ .



10.45 a) In the  $\text{H}_2\text{SO}_3$  molecule, the S and the O atoms with an H attached serve as central atoms. The S has 3 groups and one lone pair ( $\text{AX}_3\text{E}$ ), so it is **trigonal pyramidal** with angles **less than** from the ideal angle of  $109.5^\circ$ . The H bearing O atoms each have 2 bonds and 2 lone pairs ( $\text{AX}_2\text{E}_2$ ), so the arrangement is **V-shaped** or **bent** with an angle **less than the ideal** values of  $109.5^\circ$ .

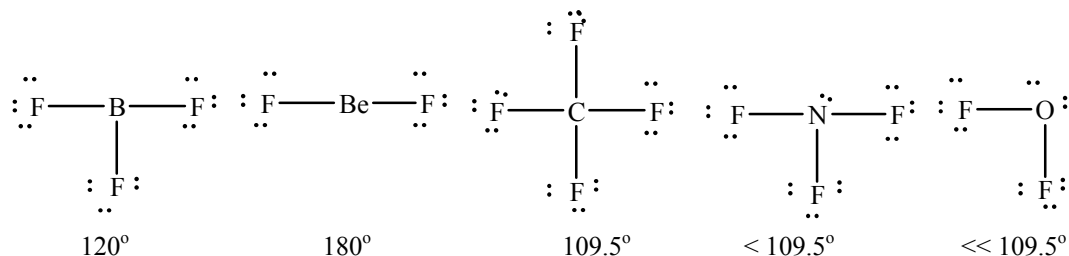


b) The  $\text{N}_2\text{O}_3$  molecule has the structure indicated in the problem with the N atoms serving as central atoms. The nitrogen labeled  $\text{N}_1$  has 2 groups and a lone pair ( $\text{AX}_2\text{E}$ ), so it is **V-shaped** or **bent** with angles **less than the ideal** values of  $120^\circ$ . The nitrogen labeled  $\text{N}_2$  has 3 bonds and no lone pairs ( $\text{AX}_3$ ), so it is **trigonal planar** with **no deviation** from the ideal angle of  $120^\circ$ .



10.46 Plan: First, draw a Lewis structure, and then apply VSEPR.

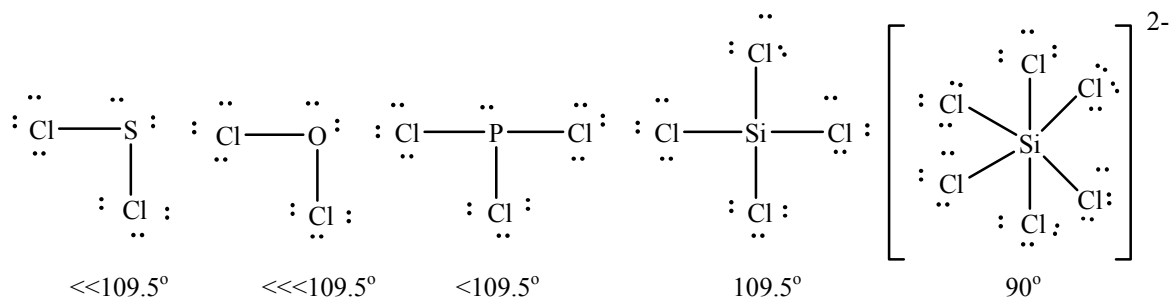
Solution:



Bond angles:  $\text{OF}_2 < \text{NF}_3 < \text{CF}_4 < \text{BF}_3 < \text{BeF}_2$

$\text{BeF}_2$  is an  $\text{AX}_2$  type molecule, so the angle is the ideal  $180^\circ$ .  $\text{BF}_3$  is an  $\text{AX}_3$  molecule, so the angle is the ideal  $120^\circ$ .  $\text{CF}_4$ ,  $\text{NF}_3$ , and  $\text{OF}_2$  all have tetrahedral electron-group arrangements of the following types:  $\text{AX}_4$ ,  $\text{AX}_3\text{E}$ , and  $\text{AX}_2\text{E}_2$ , respectively. The ideal tetrahedral bond angle is  $109.5^\circ$ , which is present in  $\text{CF}_4$ . The one lone pair in  $\text{NF}_3$  decreases the angle a little. The two lone pairs in  $\text{OF}_2$  decrease the angle even more.

10.47



Bond angles:  $\text{SiCl}_4 > \text{PCl}_3 > \text{SCl}_2 > \text{OCl}_2 > \text{SiCl}_6^{2-}$

All the species except  $\text{SiCl}_6^{2-}$  are based on a tetrahedral electron group arrangement.  $\text{SiCl}_6^{2-}$  has an octahedral electron arrangement with an ideal angle of  $90^\circ$ . The tetrahedral arrangement has an ideal bond angle of  $109.5^\circ$ , which is present in  $\text{AX}_4$  species like  $\text{SiCl}_4$ . The ideal tetrahedral bond angle is reduced slightly by the lone pair in  $\text{AX}_3\text{E}$  species such as  $\text{PCl}_3$ . A greater reduction in the ideal tetrahedral bond angle is present in  $\text{AX}_2\text{E}_2$  species such as  $\text{SCl}_2$  and  $\text{OCl}_2$  with two lone pairs. The angle is reduced less around the larger S atom.

10.48 Plan: The ideal bond angles depend on the electron-group arrangement. Deviations depend on lone pairs.

Solution:

a) The C and N have 3 groups, so they are **ideally  $120^\circ$** , and the O has 4 groups, so **ideally the angle is  $109.5^\circ$** . The N and O have lone pairs, so the **angles are less than ideal**.

b) All central atoms have 4 pairs, so ideally all the angles are  **$109.5^\circ$** . The lone pairs on the O **reduce** this value.

c) The B has 3 groups (no lone pairs) leading to an **ideal bond angle of  $120^\circ$** . All the O's have 4 pairs (**ideally  $109.5^\circ$** ), 2 of which are lone, and **reduce the angle**.

10.49 a) The N has 3 groups, no lone pairs, so the angle is **ideal**, and equal to  **$120^\circ$** . The O, attached to the H, has 4 groups (**ideally  $109.5^\circ$** ), the lone pairs **reduce the bond angle from ideal**.

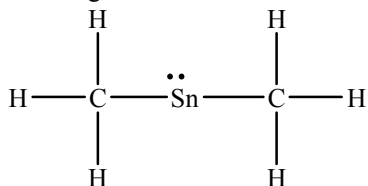
b) The C, attached to the O, has 3 groups and no lone pairs so the angle will be the **ideal  $120^\circ$** . The remaining C has 4 groups, and with no lone pairs the angle will be ideal and equal to  **$109.5^\circ$** .

c) The C with 3 groups will have angles that are **ideal ( $120^\circ$ )**. The O, with the H attached, has 4 groups. The presence of 4 groups gives an **ideal angle of  $109.5^\circ$** , which is **reduced** by the lone pairs.

10.50

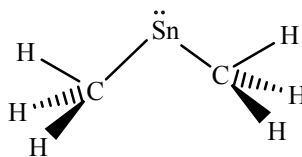
a) Type: AX<sub>2</sub>E

Ideal angle: 120°

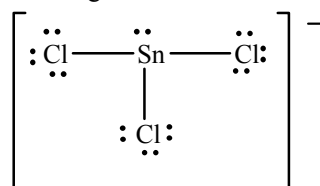


Shape: bent

Actual angle: &lt;120° (because of the lone pair)

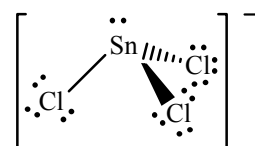
b) Type: AX<sub>3</sub>E

Ideal angle: 109.5°

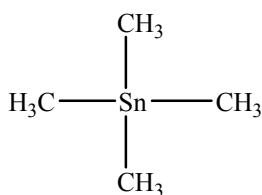


Shape: trigonal pyramidal

Actual angle: &lt;109.5° (because of the lone pair)

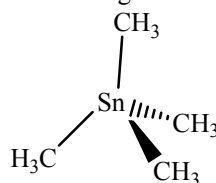
c) Type: AX<sub>4</sub>

Ideal angle: 109.5°

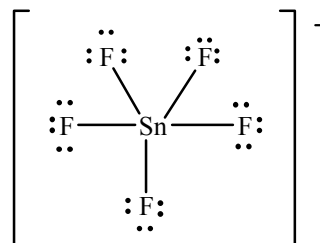


Shape: tetrahedral

Actual angle: 109.5° (there are no lone pairs)

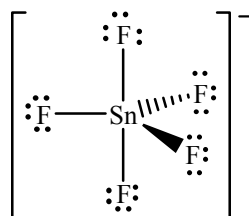
d) Type: AX<sub>5</sub>

Ideal angles: 120° and 90°

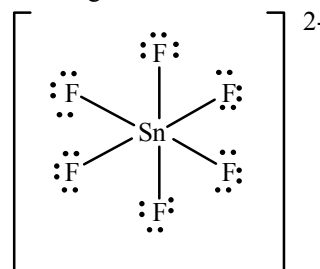


Shape: trigonal bipyramidal

Actual angle: 120° and 90° (there are no lone pairs)

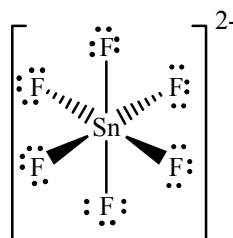
e) Type: AX<sub>6</sub>

Ideal angles: 90°



Shape: octahedral

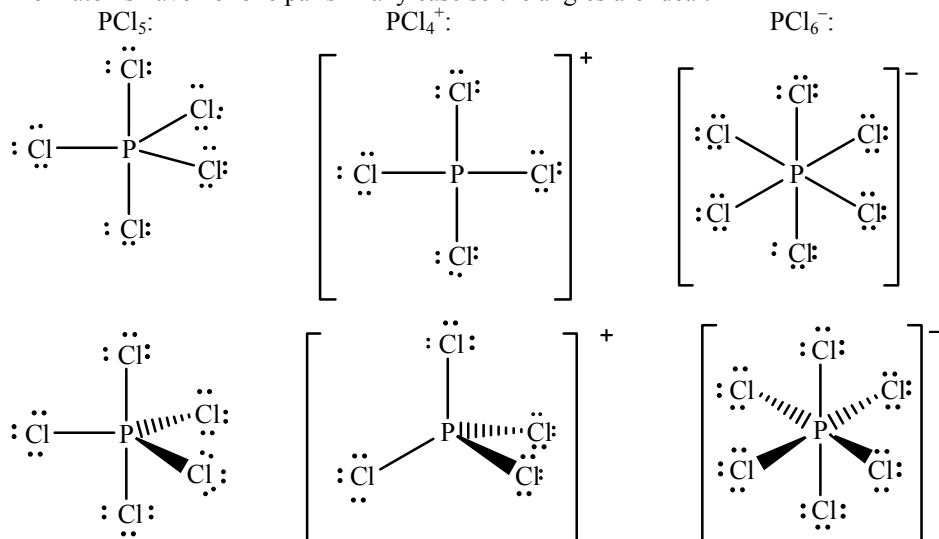
Actual angle: 90° (there are no lone pairs)



10.51 Plan: The Lewis structures are needed to predict the ideal bond angles.

Solution:

The P atoms have no lone pairs in any case so the angles are ideal.



The original PCl<sub>5</sub> is AX<sub>5</sub>, so the shape is trigonal bipyramidal, and the angles are 120° and 90°.

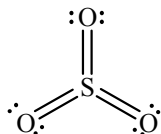
The PCl<sub>4</sub><sup>+</sup> is AX<sub>4</sub>, so the shape is tetrahedral, and the angles are 109.5°.

The PCl<sub>6</sub><sup>-</sup> is AX<sub>6</sub>, so the shape is octahedral, and the angles are 90°.

Half the PCl<sub>5</sub> (trigonal bipyramidal, 120° and 90°) become tetrahedral PCl<sub>4</sub><sup>+</sup> (tetrahedral, 109.5°), and the other half become octahedral PCl<sub>6</sub><sup>-</sup> (octahedral, 90°).

10.52 Molecules are polar if they have polar bonds that are not arranged to cancel each other. A polar bond is present any time there is a bond between elements with differing electronegativities.

10.53 Molecules are polar if they have polar bonds that are not arranged to cancel each other. If the polar covalent bonds are arranged in such a way as to cancel each other, the molecule will be nonpolar. An example of a molecule with polar covalent bonds that is not polar is SO<sub>3</sub>. The trigonal planar shape causes the three polar S–O bonds to cancel.



10.54 Molecules must come together to react. This becomes difficult for large molecules. Biomolecules are generally large molecules and have difficulty reacting if the shapes of the molecules are not compatible.

10.55 Plan: To determine if a bond is polar, determine the electronegativity difference of the atoms participating in the bond. To determine if a molecule is polar (has a dipole moment), it must have polar bonds, and a certain shape determined by VSEPR.

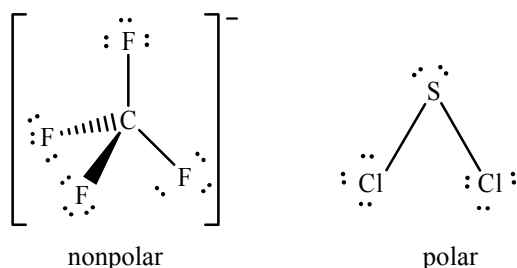
Solution:

a) The greater the difference in electronegativity the more polar the bond:

| Molecule         | Bond  | Electronegativities | Electronegativity difference |
|------------------|-------|---------------------|------------------------------|
| SCl <sub>2</sub> | S–Cl  | S = 2.5 Cl = 3.0    | 3.0 – 2.5 = 0.5              |
| F <sub>2</sub>   | F–F   | F = 4.0 F = 4.0     | 4.0 – 4.0 = 0.0              |
| CS <sub>2</sub>  | C–S   | C = 2.5 S = 2.5     | 2.5 – 2.5 = 0.0              |
| CF <sub>4</sub>  | C–F   | C = 2.5 F = 4.0     | 4.0 – 2.5 = 1.5              |
| BrCl             | Br–Cl | Br = 2.8 Cl = 3.0   | 3.0 – 2.8 = 0.2              |

The polarities of the bonds increases in the order: F–F = C–S < Br–Cl < S–Cl < C–F. Thus, **CF<sub>4</sub>** has the most polar bonds.

b) The  $F_2$  and  $CS_2$  cannot be polar since they do not have polar bonds.  $CF_4$  is a  $AX_4$  molecule, so it is tetrahedral with the 4 polar C–F bonds arranged to cancel each other giving an overall nonpolar molecule. **BrCl has a dipole moment** since there are no other bonds to cancel the polar Br–Cl bond.  **$SCl_2$  has a dipole moment** (is polar) because it is a bent molecule,  $AX_2E_2$ , and the S–Cl bonds both pull to one side.

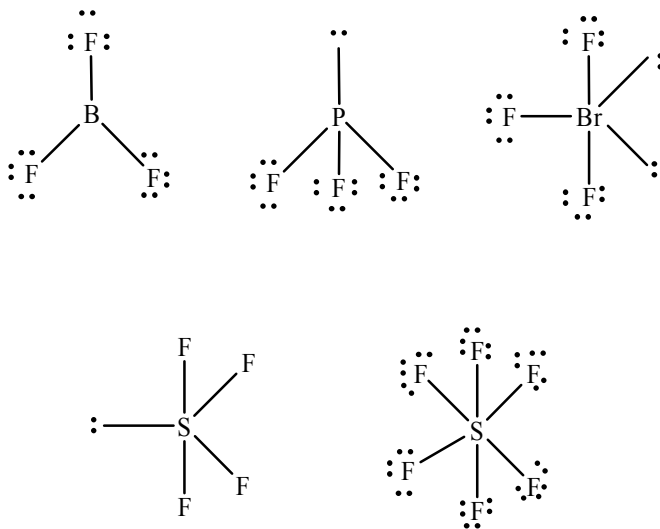


10.56 a) The greater the difference in electronegativity the more polar the bond:

| Molecule | Bond | Electronegativities | Electronegativity difference |
|----------|------|---------------------|------------------------------|
| $BF_3$   | B–F  | B = 2.0 F = 4.0     | $4.0 - 2.0 = 2.0$            |
| $PF_3$   | P–F  | P = 2.1 F = 4.0     | $4.0 - 2.1 = 1.9$            |
| $BrF_3$  | Br–F | Br = 2.8 F = 4.0    | $4.0 - 2.8 = 1.2$            |
| $SF_4$   | S–F  | S = 2.5 F = 4.0     | $4.0 - 2.5 = 1.5$            |
| $SF_6$   | S–F  | S = 2.5 Cl = 4.0    | $4.0 - 2.5 = 1.5$            |

The polarities of the bonds are increasing in the order: Br–F < S–F < P–F < B–F. Thus,  **$BF_3$**  has the most polar bonds.

b) All the molecules meet the requirement of having polar bonds. The arrangement of the bonds must be considered in each case.  $BF_3$  is trigonal planar,  $AX_3$ , so it is nonpolar because the polarities of the bonds cancel.  **$PF_3$  has a dipole moment** (is polar) because it has a trigonal pyramidal geometry,  $AX_3E$ .  **$BrF_3$  has a dipole moment** because it has a T-shaped geometry,  $AX_3E_2$ .  **$SF_4$  has a dipole moment** because it has a see-saw geometry,  $AX_4E$ .  $SF_6$  is nonpolar because it is octahedral,  $AX_6$ , and the bonds are arranged so they cancel.



10.57 Plan: If only 2 atoms are involved, only an electronegativity difference is needed. If there are more than 2 atoms, the structure must be determined.

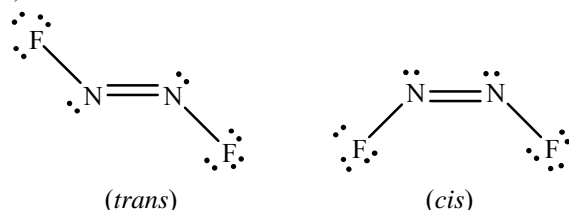
Solution:

a) All the bonds are polar covalent. The  $SO_3$  molecule is trigonal planar,  $AX_3$ , so the bond dipoles cancel leading to a nonpolar molecule (no dipole moment). The  $SO_2$  molecule is bent,  $AX_2E$ , so the polar bonds both pull to one side.  **$SO_2$  has a greater dipole moment** because it is the only one of the pair that is polar.



10.60 The possible structures for the compounds differ only in the positions of the N-F bonds. These structures are “fixed” because the N=N bond does not allow rotation. The N-F bonds are polar.

a)



b) In the *trans*-form the N-F bonds pull equally in opposite directions, thus, they cancel and the molecule is nonpolar. The N-F bonds in the *cis*-form pull in the same general direction resulting in a polar molecule.

10.61 Plan: The Lewis structures are needed to do this problem. For part (b), obtain bond energy values from Table 9.2.

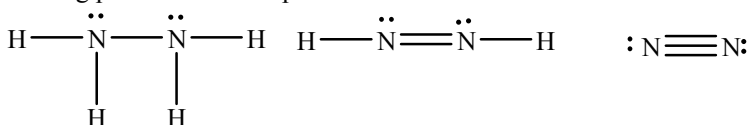
Solution:

a) The H atoms cannot be central, and they are evenly distributed on the N's.

$\text{N}_2\text{H}_4$  has  $[2 \times \text{N}(5 e^-)] + [4 \times \text{H}(1 e^-)] = 14$  valence electrons, 10 of which are used in the bonds between the atoms. The remaining two pairs are used to complete the octets of the N atoms.

$\text{N}_2\text{H}_2$  has  $[2 \times \text{N}(5 e^-)] + (2 \times \text{H}(1 e^-)) = 12$  valence electrons, 6 of which are used in the bonds between the atoms. The remaining three pairs of electrons are not enough to complete the octets of both N atoms, so one lone pair is moved to a bonding pair between the N atoms.

$\text{N}_2$  has  $[2 \times \text{N}(5 e^-)] = 10$  valence electrons, 2 of which are used to place a single bond between the two N atoms. Since only 4 pairs of electrons remain and 6 pairs are required to complete the octets, two lone pairs become bonding pairs to form a triple bond.



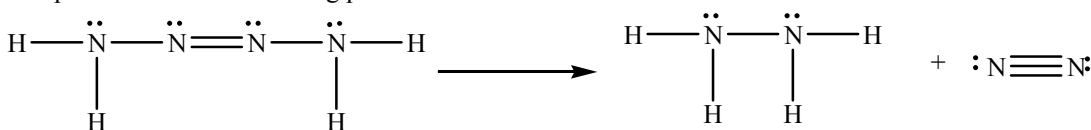
Hydrazine

Diazene

Nitrogen

The single (bond order = 1) **N-N bond is weaker and longer** than any of the others are. The **triple bond (bond order = 3) is stronger and shorter** than any of the others. The **double bond (bond order = 2) has an intermediate strength and length.**

b)  $\text{N}_4\text{H}_4$  has  $[4 \times \text{N}(5 e^-)] + [4 \times \text{H}(1 e^-)] = 24$  valence electrons, 14 of which are used for single bonds between the atoms. When the remaining 5 pairs are distributed to complete the octets, one N atom lacks two electrons. A lone pair is moved to a bonding pair for a double bond.



$$\Delta H^\circ_{\text{bonds broken}} = 4 \text{ N-H} = 4 \text{ mol (391 kJ/mol)}$$

$$2 \text{ N-N} = 2 \text{ mol (160 kJ/mol)}$$

$$1 \text{ N=N} = 1 \text{ mol (418 kJ/mol)}$$

$$= 2302 \text{ kJ}$$

$$\Delta H^\circ_{\text{bonds formed}} = 4 \text{ N-H} = 4 \text{ mol (-391 kJ/mol)}$$

$$1 \text{ N-N} = 1 \text{ mol (-160 kJ/mol)}$$

$$1 \text{ N}\equiv\text{N} = 1 \text{ mol (-945 kJ/mol)}$$

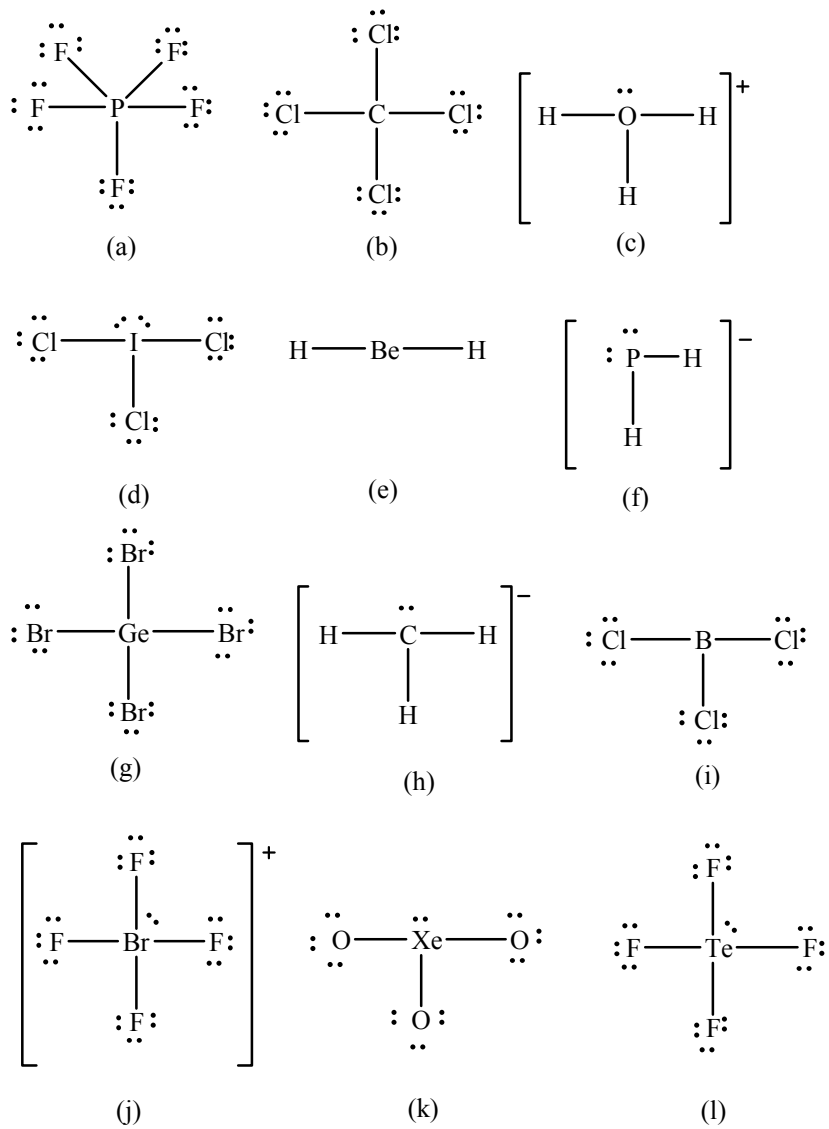
$$= -2669 \text{ kJ}$$

$$\Delta H^\circ_{\text{rxn}} = \Delta H^\circ_{\text{bonds broken}} + \Delta H^\circ_{\text{bonds formed}} = 2302 \text{ kJ} + (-2669 \text{ kJ}) = \mathbf{-367 \text{ kJ}}$$

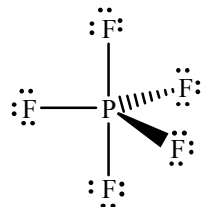
Note: It is correct to report the answer in kJ or kJ/mol as long as the value refers to a reactant or product with a molar coefficient of 1.



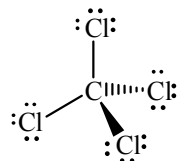
10.62



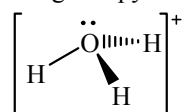
10.63 a)  $AX_5$  = trigonal bipyramidal



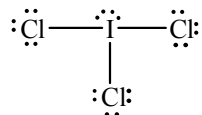
b)  $AX_4$  = tetrahedral



c)  $AX_3E$  = trigonal pyramidal



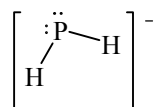
d)  $AX_3E_2$  = T-shaped



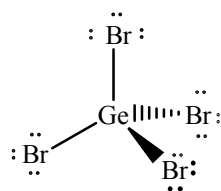
e)  $AX_2$  = linear



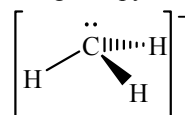
f)  $AX_2E_2$  = bent



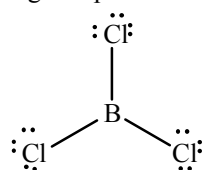
g)  $AX_4$  = tetrahedral



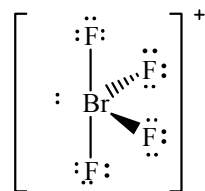
h)  $AX_3E$  = trigonal pyramidal



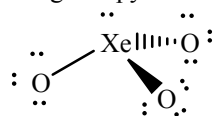
i)  $AX_3$  = trigonal planar



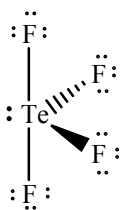
j)  $AX_4E$  = see-saw



k)  $AX_3E$  = trigonal pyramidal



l) AX<sub>4</sub>E = see-saw



10.64 a) SiF<sub>4</sub> with its 32 valence electrons is an AX<sub>4</sub> molecule and has a tetrahedral molecular shape. SiF<sub>5</sub><sup>-</sup> with its 40 valence electrons is an AX<sub>5</sub> ion and has a trigonal bipyramid molecular shape. **B** best represents the change in molecular shape from tetrahedral to trigonal bipyramid.

b) **SiF<sub>4</sub>: tetrahedral, AX<sub>4</sub>; SiF<sub>5</sub><sup>-</sup>: trigonal bipyramid, AX<sub>5</sub>.**

10.65 Plan: Use the Lewis structures shown in the text.

FC = valence electrons - [lone electrons + 1/2 (bonded electrons)]

Solution:

a) Formal charges for Al<sub>2</sub>Cl<sub>6</sub>:

$$FC_{Al} = 3 - (0 + 1/2(8)) = -1$$

$$FC_{Cl, ends} = 7 - (6 + 1/2(2)) = 0$$

$$FC_{Cl, bridging} = 7 - (4 + 1/2(4)) = +1$$

(Check: Formal charges add to zero, the charge on the compound.)

Formal charges for I<sub>2</sub>Cl<sub>6</sub>:

$$FC_I = 7 - (4 + 1/2(8)) = -1$$

$$FC_{Cl, ends} = 7 - (6 + 1/2(2)) = 0$$

$$FC_{Cl, bridging} = 7 - (4 + 1/2(4)) = +1$$

(Check: Formal charges add to zero, the charge on the compound.)

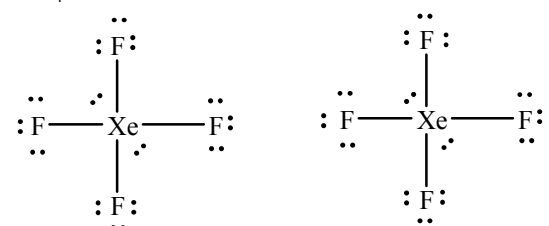
b) The aluminum atoms have no lone pairs and are AX<sub>4</sub>, so they are tetrahedral. The 2 tetrahedral Al's cannot give a planar structure. The iodine atoms have two lone pairs each and are AX<sub>4</sub>E<sub>2</sub> so they are square planar. Placing the square planar I's adjacent can give a planar molecule.

10.66 The Lewis structure for each is required.

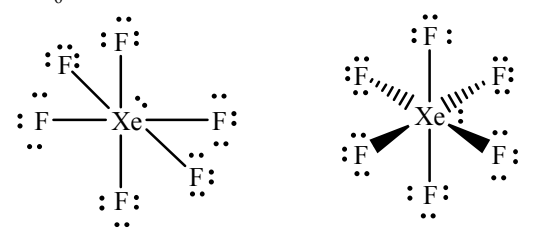
|                  |           |                 |                                                 |
|------------------|-----------|-----------------|-------------------------------------------------|
| Compound         | Pre-VSEPR | Lewis Structure | VSEPR                                           |
| XeF <sub>2</sub> | Linear    |                 | <b>Linear</b> (AX <sub>2</sub> E <sub>3</sub> ) |



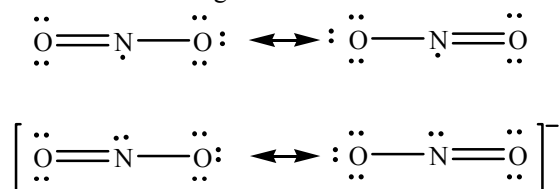
|                  |             |  |                                                        |
|------------------|-------------|--|--------------------------------------------------------|
| XeF <sub>4</sub> | Tetrahedral |  | <b>Square planar</b> (AX <sub>4</sub> E <sub>2</sub> ) |
|------------------|-------------|--|--------------------------------------------------------|



|                  |            |  |                                                 |
|------------------|------------|--|-------------------------------------------------|
| XeF <sub>6</sub> | Octahedral |  | <b>Distorted octahedral</b> (AX <sub>6</sub> E) |
|------------------|------------|--|-------------------------------------------------|



- 10.67 a)  $\text{SO}_3$  is an  $\text{AX}_3$  molecule and has a trigonal planar shape.  $\text{SO}_3^{2-}$  is an  $\text{AX}_3\text{E}$  species and has a trigonal pyramid molecular shape. **C** best illustrates the change in molecular shape from trigonal planar to trigonal pyramid.
- b) **Yes**, there is a change in polarity during the reaction as the nonpolar  $\text{SO}_3$  molecule becomes the polar  $\text{SO}_3^{2-}$  ion.
- 10.68 From the Lewis structures, both are  $\text{AX}_2\text{E}$  which has an ideal bond angle of  $120^\circ$ . But the “lone pair” on N in  $\text{NO}_2$  is only half a pair, so it only exerts “half” the repulsion. This allows the bond angle to open to a larger than normal bond angle. The “complete” lone pair in  $\text{NO}_2^-$ , like other lone pairs, forces the bonding pairs together to give a smaller than normal bond angle.

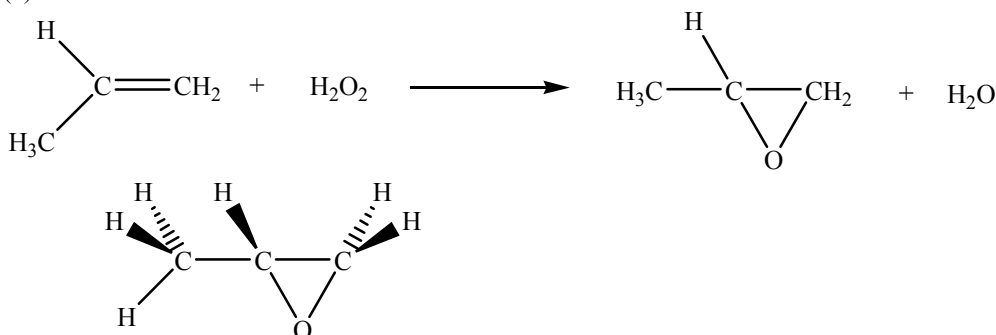


- 10.68  $\text{Xe}(g) + 3 \text{F}_2(g) \rightarrow \text{XeF}_6(g)$   
 The 3 F-F bonds must be broken, and 6 Xe-F bonds are formed.  
 $\Delta H_f = 3 \text{BE}_{\text{F-F}} + 6 \text{BE}_{\text{Xe-F}}$   
 $-402 \text{ kJ/mol} = (3) (159 \text{ kJ/mol}) + 6 \text{BE}_{\text{Xe-F}}$   
 $-879 \text{ kJ/mol} = 6 \text{BE}_{\text{Xe-F}}$   
 $146.5 = \mathbf{146 \text{ kJ/mol}} = \text{BE}_{\text{Xe-F}}$

- 10.70 Plan: Draw the Lewis structures, and then use VSEPR to describe propylene oxide.

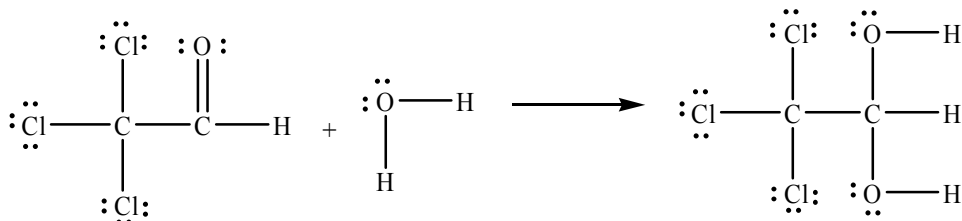
Solution:

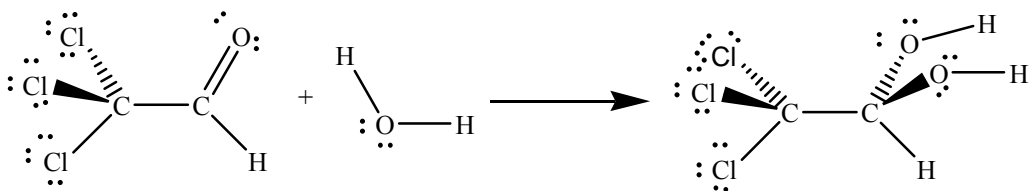
(a)



- a) In propylene oxide, the C atoms are all  $\text{AX}_4$ . The C atoms do not have any unshared (lone) pairs. All the ideal bond angles for the C atoms in propylene oxide are  $109.5^\circ$ .
- b) In propylene oxide, the C that is not part of the three-membered ring should have an ideal angle. The atoms in the ring form an equilateral triangle. The angles in an equilateral triangle are  $60^\circ$ . The angles around the 2 carbons in the rings are reduced from the ideal  $109.5^\circ$  to  $60^\circ$ .

10.71



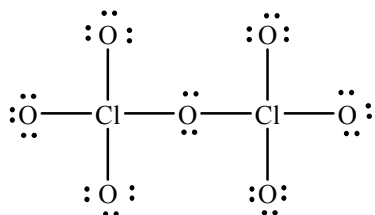


The C with the chlorines attached does not change shape. That C is tetrahedral in both compounds. The other C changes from trigonal planar (AX<sub>3</sub>) to tetrahedral (AX<sub>4</sub>).

10.72 Plan: Draw an acceptable Lewis structure using both the octet rule and the minimization of formal charges.

Solution:

The molecule has [2 x Cl(7 e<sup>-</sup>)] + [7 x O(6 e<sup>-</sup>)] = 56 valence electrons. Use 16 electrons to place single bonds between the O atoms and the Cl atoms. This leaves 56 - 16 = 40 e<sup>-</sup> (20 pairs). Use these 20 pairs to complete the octets of the O and Cl atoms.



The formal charges now need to be calculated for each type of atom:

$$FC = \text{valence electrons} - (\text{lone electrons} + 1/2 (\text{bonded electrons}))$$

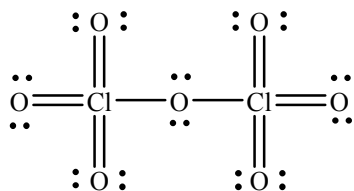
$$FC_{Cl} = 7 - (0 + 1/2(8)) = +3$$

$$FC_O = 6 - (4 + 1/2(4)) = 0 \text{ (bridging O)}$$

$$FC_O = 6 - (6 + 1/2(2)) = -1 \text{ (terminal O atoms)}$$

The formal charges check (the total is 0), but the values for the Cl atoms are high. An attempt should be made to decrease the formal charges.

The formal charges in the Cl atoms may be reduced by the formation of double bonds. To reduce the formal charge from +3 to 0 would require three double bonds. This is allowed since Cl can exceed an octet. This gives:



The formal charge needs to be recalculated:

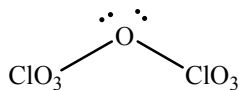
$$FC_{Cl} = 7 - (0 + 1/2(14)) = 0$$

$$FC_O = 6 - (4 + 1/2(4)) = 0 \text{ (bridging O)}$$

$$FC_O = 6 - (4 + 1/2(4)) = 0 \text{ (terminal O atoms)}$$

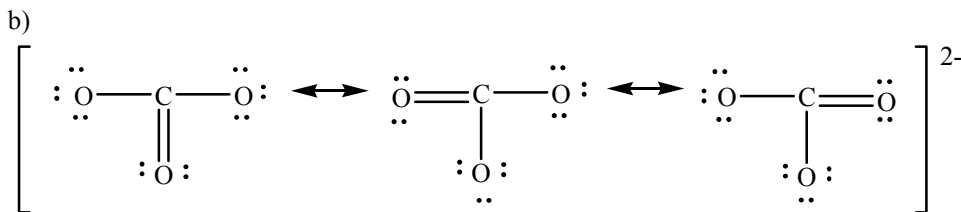
The formal charges check (the total is 0). All the values are reasonable (0).

The central bridging O is AX<sub>2</sub>E<sub>2</sub>, which has an ideal bond angle of 109.5°, but the lone pairs make this less than ideal.



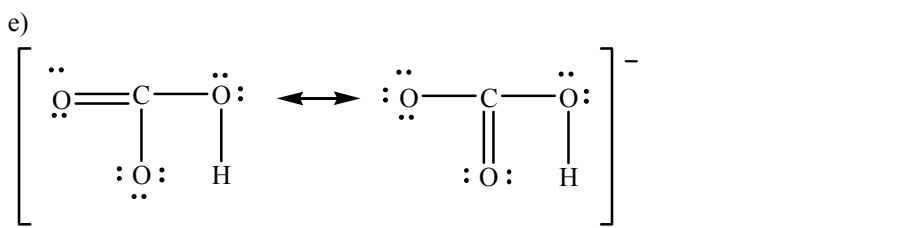
10.73 Plan: Count the valence electrons, and draw the appropriate Lewis structures. Draw resonance structures when necessary.

Solution:



Each C–O bond is a single bond two-thirds of the time and a double bond the rest of the time.

The average is  $[(1 + 1 + 2)/3] = 4/3 = 1.33$



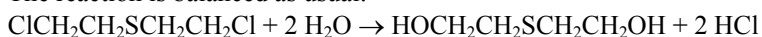
The resonating double bond means the average bond length is  $[(1 + 2)/2] = 1.5$

The C–O bond for the O attached to the H does not resonate and remains 1.0

Bond length  $\mathbf{a < c < e < b < d}$  ignoring O attached to H in e)

Bond strength  $\mathbf{d < b < e < c < a}$

10.74 The reaction is balanced as usual:

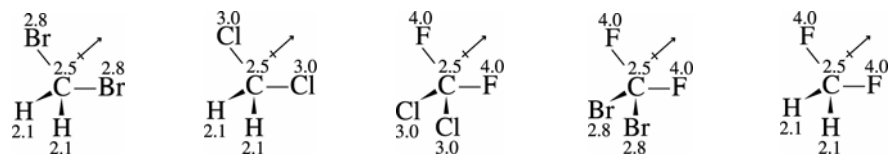


Most of the molecule remains the same.

|               |                  |                      |               |               |                       |
|---------------|------------------|----------------------|---------------|---------------|-----------------------|
| Bonds broken: | 2 C-Cl           | 2(339 kJ/mol)        | Bonds formed: | 2 H-Cl        | 2(-427 kJ/mol)        |
|               | 8 C-H            | 8(413 kJ/mol)        |               | 8 C-H         | 8(-413 kJ/mol)        |
|               | 2 C-S            | 2(259 kJ/mol)        |               | 2 C-S         | 2(-259 kJ/mol)        |
|               | 2 C-C            | 2(347 kJ/mol)        |               | 2 C-C         | 2(-347 kJ/mol)        |
|               | <u>2 x 2 O-H</u> | <u>4(467 kJ/mol)</u> |               | 2 O-H         | 2(-467 kJ/mol)        |
|               | <b>Total:</b>    | <b>7062 kJ</b>       |               | <u>2 C-O</u>  | <u>2(-358 kJ/mol)</u> |
|               |                  |                      |               | <b>Total:</b> | <b>- 7020 kJ</b>      |

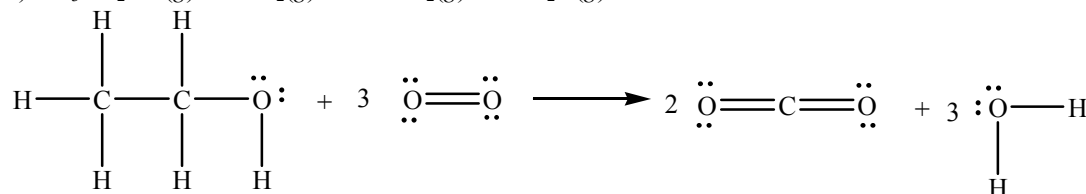
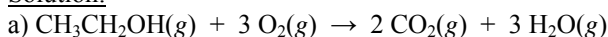
$$\begin{aligned} \Delta H^\circ_{\text{rxn}} &= \Delta H^\circ_{\text{bonds broken}} + \Delta H^\circ_{\text{bonds formed}} \\ &= 7062 \text{ kJ} + (-7020 \text{ kJ}) = \mathbf{42 \text{ kJ}} \end{aligned}$$

10.75  $\text{CBr}_4 < \text{CH}_2\text{Br}_2 < \text{CH}_2\text{Cl}_2 < \text{CF}_2\text{Cl}_2 < \text{CF}_2\text{Br}_2 < \text{CH}_2\text{F}_2$



10.76 Plan: Assume all reactants and products are assumed to be gaseous. Ethanol burns (combusts) with  $\text{O}_2$  to produce  $\text{CO}_2$  and  $\text{H}_2\text{O}$ . Remember that energy is released (-) when bonds are formed.

Solution:



|               |       |                      |               |           |                        |
|---------------|-------|----------------------|---------------|-----------|------------------------|
| Bonds broken: | 1 C-C | 1(347 kJ/mol)        | Bonds formed: | 2 x 2 C=O | 4(- 799 kJ/mol)        |
|               | 5 C-H | 5(413 kJ/mol)        |               | 2 x 3 O-H | <u>6(- 467 kJ/mol)</u> |
|               | 1 C-O | 1(358 kJ/mol)        | Total:        |           | <u>-5998 kJ</u>        |
|               | 1 O-H | 1(467 kJ/mol)        |               |           |                        |
|               | 3 O=O | <u>3(498 kJ/mol)</u> |               |           |                        |
| Total:        |       | 4731 kJ              |               |           |                        |

$\Delta H = \text{Bonds broken} + \text{Bonds formed}$   
 $= 4731 \text{ kJ} + (-5998 \text{ kJ}) = -1267 \text{ kJ}$  for each mole of ethanol burned.

b) If it takes 40.5 kJ/mol to vaporize the ethanol, part of the heat of combustion must be used to convert liquid ethanol to gaseous ethanol. The new value becomes:

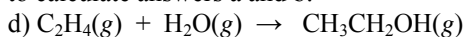
$$\Delta H_{\text{comb(liquid)}} = -1267 \text{ kJ} + (1 \text{ mol}) \left[ \frac{40.5 \text{ kJ}}{1 \text{ mol}} \right] = -1226.5 = -1226 \text{ kJ per mol of liquid ethanol burned}$$

c)  $\Delta H_{\text{comb(liquid)}} = \sum n\Delta H_f^\circ (\text{products}) - \sum m\Delta H_f^\circ (\text{reactants})$

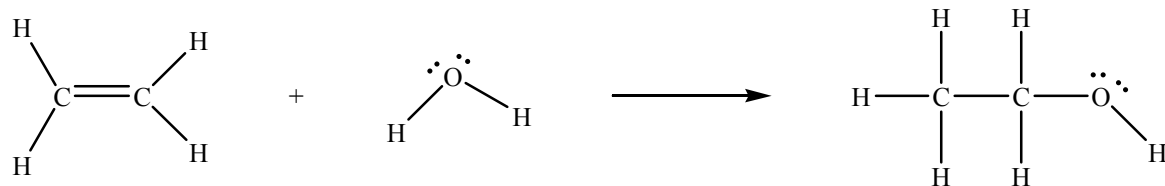
$$\Delta H_{\text{comb(liquid)}} = [2(\Delta H_f^\circ (\text{CO}_2(g))) + 3(\Delta H_f^\circ (\text{H}_2\text{O}(g)))] - [(\Delta H_f^\circ (\text{C}_2\text{H}_5\text{OH}(l))) + 3(\Delta H_f^\circ (\text{O}_2(g)))]$$

$$\Delta H_{\text{comb(liquid)}} = [2(-393.5 \text{ kJ}) + 3(-241.826 \text{ kJ})] - [(-277.63 \text{ kJ}) + 3(0 \text{ kJ})] = -1234.848 = -1234.8 \text{ kJ}$$

The two answers differ by less than 10 kJ. This is a very good agreement since average bond energies were used to calculate answers a and b.



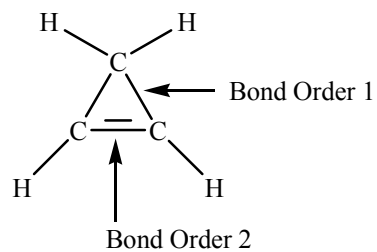
The Lewis structures for the reaction are:



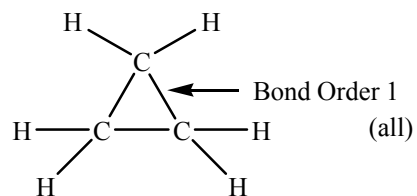
|               |                |                      |               |         |                        |
|---------------|----------------|----------------------|---------------|---------|------------------------|
| Bonds broken: | 1 x C=C        | 1(614 kJ/mol)        | Bonds formed: | 1 x C-C | 1(-347 kJ/mol)         |
|               | 4 x C-H        | 4(413 kJ/mol)        |               | 5 x C-H | 5(- 413 kJ/mol)        |
|               | <u>2 x O-H</u> | <u>2(467 kJ/mol)</u> |               | 1 x C-O | 1(- 358 kJ/mol)        |
| Total:        |                | 3200 kJ              | Total:        | 1 x O-H | <u>1(- 467 kJ/mol)</u> |
|               |                |                      |               |         | <u>-5998 kJ</u>        |

$$\Delta H_{\text{rxn}}^\circ = \Delta H_{\text{bonds broken}}^\circ + \Delta H_{\text{bonds formed}}^\circ = 3200. \text{ kJ} + (-3237 \text{ kJ}) = -37 \text{ kJ}$$

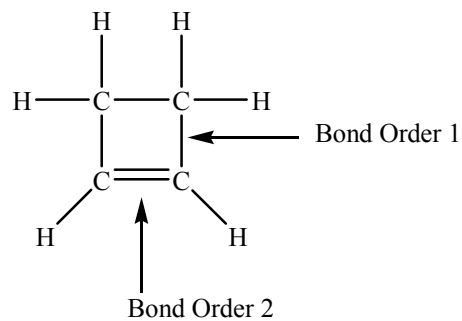
10.77 a)  $C_3H_4$ : 16 valence electrons.



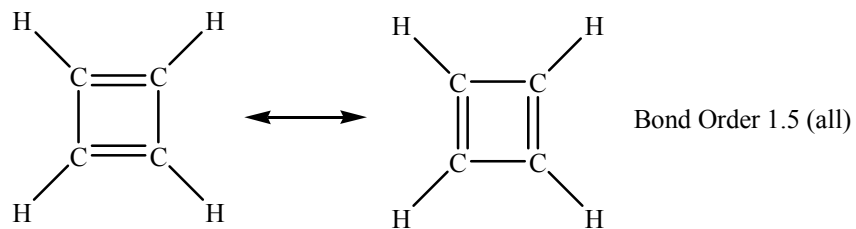
b)  $C_3H_6$ : 18 valence electrons.



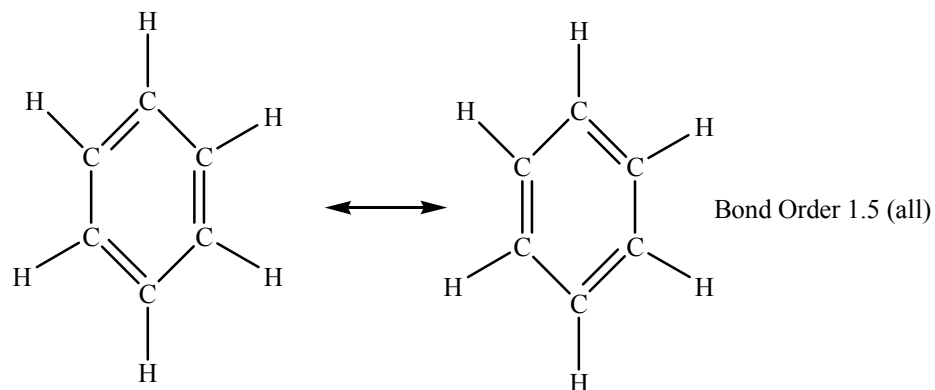
c)  $C_4H_6$ : 22 valence electrons..



d)  $C_4H_4$ : 20 valence electrons. Resonance exists.



e)  $C_6H_6$ : 30 valence electrons. Resonance exists.





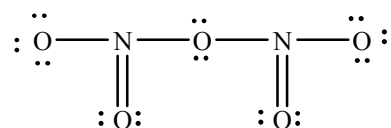
- 10.78 Determine the empirical formula from the percent composition (assuming 100 grams of compound). The empirical formula and the molar mass may then be used to determine the molecular formula. Count the valence electrons in the molecular formula and then construct the Lewis structure. Name the compound from its molecular formula.

$$\text{N} \quad 25.9 \text{ g} / (14.01 \text{ g/mol}) = 1.849 \text{ mol} / 1.849 \text{ mol} = 1.00 \quad \text{Only round at the end.}$$

$$\text{O} \quad (100.0 - 25.9) \text{ g} / (16.00 \text{ g/mol}) = 4.631 \text{ mol} / 1.849 \text{ mol} = 2.50$$

Doubling the ratios gives N = 2 and O = 5 or N<sub>2</sub>O<sub>5</sub> with a molar mass of 108.02 g/mol. Since this is the same as the molar mass given in the problem, the empirical and molecular formulas are both N<sub>2</sub>O<sub>5</sub>.

This formula has 40 valence electrons, and when drawn with no N–N or O–O bonds one gets the following Lewis structure:



The name of this compound is **dinitrogen pentoxide**.

- 10.79 Plan: Determine the empirical formula from the percent composition (assuming 100 grams of compound). Use the titration data to determine the mole ratio of acid to the NaOH. This ratio relates the number of acidic H's to the formula of the acid. Finally, combine this information to construct the Lewis structure.

Solution:

$$\text{H} \quad (2.24 \text{ g H}) \left( \frac{1 \text{ mol}}{1.008 \text{ g H}} \right) = 2.222 \text{ mol H} \quad \frac{2.222 \text{ mol}}{2.222 \text{ mol}} = 1.00$$

$$\text{C} \quad (26.7 \text{ g C}) \left( \frac{1 \text{ mol}}{12.01 \text{ g C}} \right) = 2.223 \text{ mol C} \quad \frac{2.223 \text{ mol}}{2.222 \text{ mol}} = 1.00$$

$$\text{O} \quad (71.1 \text{ g O}) \left( \frac{1 \text{ mol}}{16.00 \text{ g O}} \right) = 4.444 \text{ mol O} \quad \frac{4.444 \text{ mol}}{2.222 \text{ mol}} = 2.00$$

The empirical formula is HCO<sub>2</sub>.

The titration required:

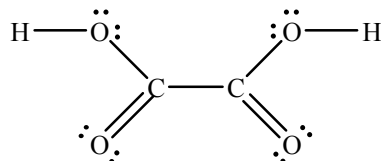
$$\left( \frac{0.040 \text{ mol NaOH}}{\text{L}} \right) \left( \frac{50.0 \text{ mL}}{\text{L}} \right) \left( \frac{1 \text{ L}}{1000 \text{ mL}} \right) \left( \frac{1 \text{ mmol}}{0.001 \text{ mol}} \right) = 2.0 \text{ mmole NaOH}$$

Thus, the ratio is 2.0 mmole base / 1.0 mmole acid, or each acid molecule has two hydrogens to react (diprotic).

The empirical formula indicates a monoprotic acid, so the formula must be doubled to: H<sub>2</sub>C<sub>2</sub>O<sub>4</sub>.

H<sub>2</sub>C<sub>2</sub>O<sub>4</sub> has [2 x H(1 e<sup>-</sup>)] + [2 x C(4 e<sup>-</sup>)] + [4 x O(6 e<sup>-</sup>)] = 34 valence electrons to be used in the Lewis structure.

14 of these electrons are used to bond the atoms with single bonds, leaving 34 – 14 = 20 electrons or 10 pairs of electrons. When these 10 pairs of electrons are distributed to the atoms to complete octets, neither C atom has an octet; a lone pair from the oxygen without hydrogen is changed to a bonding pair on C.



- 10.80 Determine the empirical formula from the percent composition (assuming 100 grams of compound). The molar mass may be determined from the density of the gas. The empirical formula and the molar mass may then be used to determine the molecular formula. Count the valence electrons in the empirical formula and then construct the Lewis structure.

$$\text{C} \quad 24.8 \text{ g} / (12.01 \text{ g/mol}) = 2.065 \text{ mol} / 2.006 \text{ mol} = 1.00$$

$$\text{H} \quad 2.08 \text{ g} / (1.008 \text{ g/mol}) = 2.063 \text{ mol} / 2.006 \text{ mol} = 1.00$$

$$\text{Cl} \quad 73.1 \text{ g} / (35.45 \text{ g/mol}) = 2.006 \text{ mol} / 2.006 \text{ mol} = 1.00$$

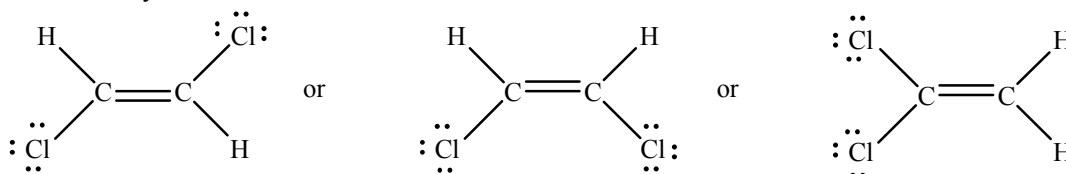
Only round at the end.

Thus, the empirical formula is CHCl, and its molar mass is: 48.47 g/mol.

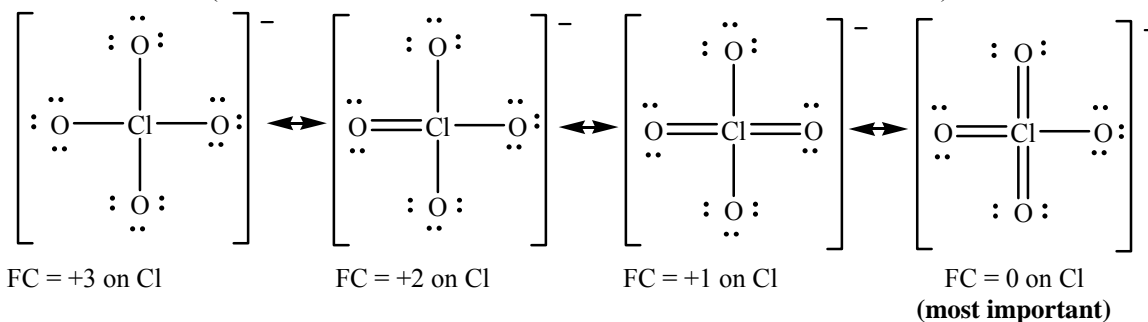
The density is 4.3 g/L at STP:

$$\mathcal{M} = (4.3 \text{ g/L}) (22.4 \text{ L/mol}) = 96.32 = 96 \text{ g/mol}$$

The molar mass (96 g/mol) is double the empirical formula mass (48.47 g/mol), so the empirical formula must be doubled to get the molecular formula:  $\text{C}_2\text{H}_2\text{Cl}_2$ . The formula contains 24 valence electrons. A variety of structures may be drawn:



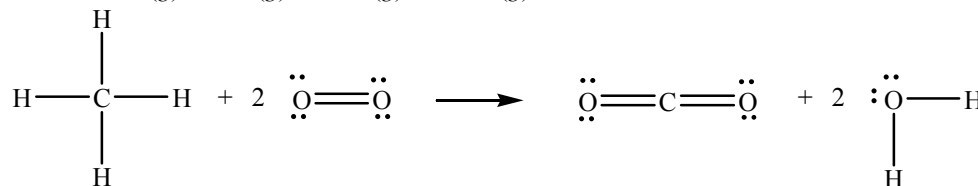
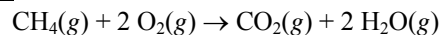
- 10.81 There are 32 valence electrons present. Begin four Lewis structures by placing a Cl in the center and 4 O's around it. Connect all the O's to the central Cl's with single bonds. In the second structure, convert one of the single bonds to a double bond. In the third structure, two of the bonds are double, and in the last, three of the bonds are double. (It does not matter which bonds are chosen to become double bonds.)



Average Bond order of the last structure:  $\frac{(3 \times 2) + (1 \times 1)}{4} = 1.75$

- 10.82 Plan: Write the balanced chemical equations for the reactions. Draw the Lewis structures. Calculate the heat of reaction from the bond energies, and divide the value by the number of moles of oxygen gas appearing in each reaction to get the heat of reaction per mole of oxygen.

Solution:

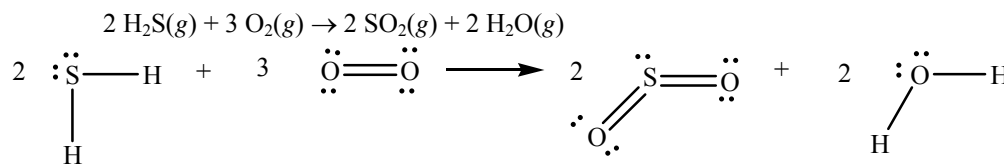


|               |        |                      |               |           |                       |
|---------------|--------|----------------------|---------------|-----------|-----------------------|
| Bonds broken: | 4 C-H  | 4(413 kJ/mol)        | Bonds formed: | 2 C=O     | 2(-799 kJ/mol)        |
|               | 2 O=O  | <u>2(498 kJ/mol)</u> |               | 2 x 2 O-H | <u>4(-467 kJ/mol)</u> |
|               | Total: | 2648 kJ              |               | Total:    | -3466 kJ              |

$$\Delta H = \Delta H^\circ_{\text{rxn}} = \Delta H^\circ_{\text{bonds broken}} + \Delta H^\circ_{\text{bonds formed}}$$

$$= 2648 \text{ kJ} + (-3466 \text{ kJ}) = -818 \text{ kJ} / 2 \text{ mol O}_2$$

Per mole of  $\text{O}_2 = -409 \text{ kJ/mol O}_2$



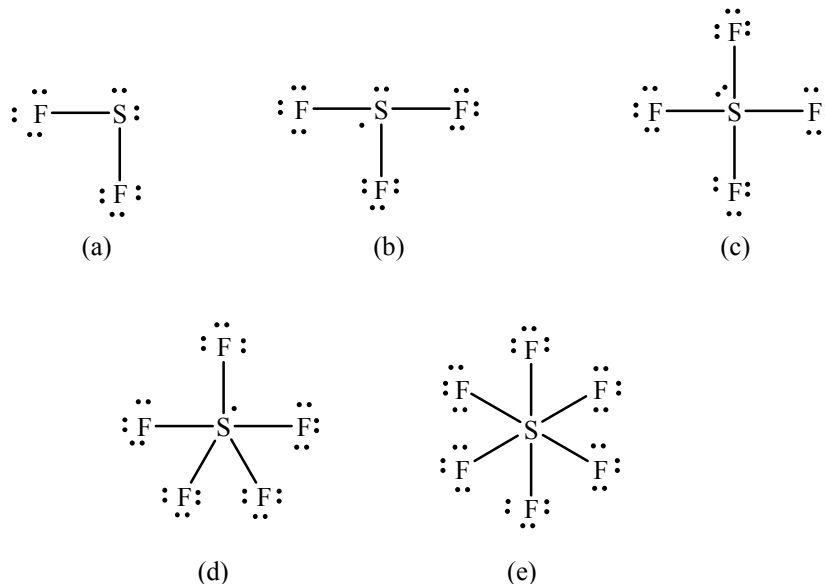
|               |           |                      |               |           |                       |
|---------------|-----------|----------------------|---------------|-----------|-----------------------|
| Bonds broken: | 2 x 2 S-H | 4(347 kJ/mol)        | Bonds formed: | 2 x 2 S=O | 4(-552 kJ/mol)        |
|               | 3 O=O     | <u>3(498 kJ/mol)</u> |               | 2 x 2 O-H | <u>4(-467 kJ/mol)</u> |
|               | Total:    | 2882 kJ              |               | Total:    | -4076 kJ              |

$$\Delta H^\circ_{\text{rxn}} = \Delta H^\circ_{\text{bonds broken}} + \Delta H^\circ_{\text{bonds formed}}$$

$$= 2882 \text{ kJ} + (-4076 \text{ kJ}) = -1194 \text{ kJ} / 3 \text{ mol O}_2$$

Per mole of O<sub>2</sub> = **-398 kJ/mol O<sub>2</sub>**

10.83



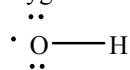
Stable: a, c, and e

**Unstable radicals: b and d**

10.84 Plan: Draw the Lewis structure of the OH species. Find the additional bond energy information in Table 9.2.

Solution:

a) The OH molecule has  $[1 \times \text{O}(6 e^-)] + [1 \times \text{H}(1 e^-)] = 7$  valence electrons to be used in the Lewis structure. 2 of these electrons are used to bond the atoms with a single bond, leaving  $7 - 2 = 5$  electrons. Those 5 electrons are given to oxygen. But no atom can have an octet, and one electron is left unpaired. The Lewis structure is:



b) The formation reaction is:  $1/2 \text{O}_2(g) + 1/2 \text{H}_2(g) \rightarrow \text{OH}(g)$ . The heat of reaction is:

$$\Delta H_{\text{reaction}} = \Delta H_{\text{broken}} + \Delta H_{\text{formed}} = 39.0 \text{ kJ}$$

$$[1/2(\text{BE}_{\text{O=O}}) + 1/2(\text{BE}_{\text{H-H}})] + [\text{BE}_{\text{O-H}}] = 39.0 \text{ kJ}$$

$$[(1/2 \text{ mol})(498 \text{ kJ/mol}) + (1/2 \text{ mol})(432 \text{ kJ/mol})] + [\text{BE}_{\text{O-H}}] = 39.0 \text{ kJ}$$

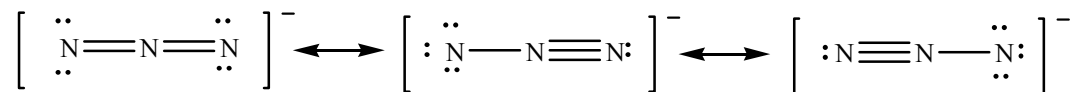
$$\text{BE}_{\text{O-H}} = \mathbf{-426 \text{ kJ or } 426 \text{ kJ}}$$

c) The average bond energy (from the bond energy table) is 467 kJ/mol. There are 2 O-H bonds in water for a total of  $2 \times 467 \text{ kJ/mol} = 934 \text{ kJ}$ . The answer to part b) accounts for 426 kJ of this, leaving:

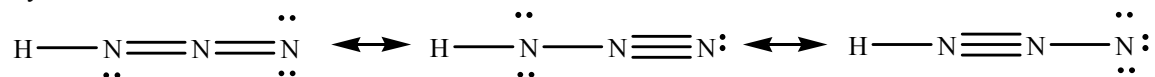
$$934 \text{ kJ} - 426 \text{ kJ} = \mathbf{508 \text{ kJ}}$$

10.85 Both N<sub>3</sub><sup>-</sup> and HN<sub>3</sub> have 16 valence electrons.

Azide ion:



There are three resonance structures for the  $\text{N}_3^-$  ion. The formal charges in the first structure are, from left to right,  $-1$ ,  $+1$ , and  $-1$ . In the other 2 Lewis structures the single bonded N has a formal charge of  $-2$ , making both of these less stable than the first structure. The central N is  $+1$  and the triple bonded N is  $0$ . The first resonance structure is more important, the structure should have 2 equal bonds with a bond order of 2.  
Hydrazoic acid:

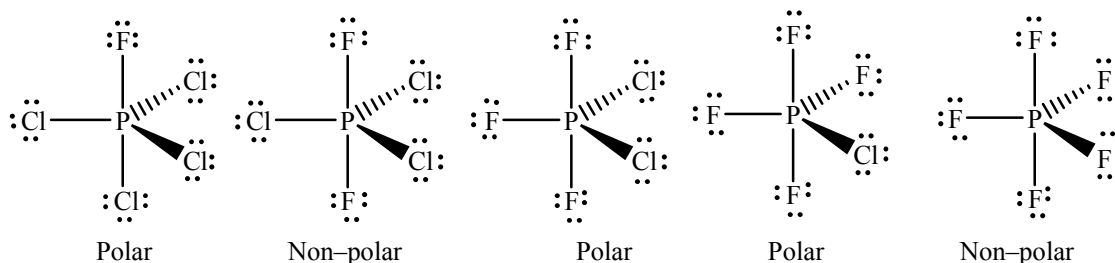


$\text{HN}_3$  also has 3 resonance structures. The formal change for the H is 0 in all the structures. In the structure with 2 double bonds, the formal charges for the N atoms are, left to right:  $0$ ,  $+1$ , and  $-1$ . The structure where the H is attached to the single bonded N, has N atoms with the following formal charges:  $-1$ ,  $+1$ , and  $0$ . In the final Lewis structure, the formal charges on the N atoms are:  $+1$ ,  $+1$ , and  $-2$ . The third structure is clearly not as good as the other two. The first two structures should be averaged to give, starting at the H-end, a bond order of 1.5 then a bond order of 2.5. Thus, the two bonds are unequal.

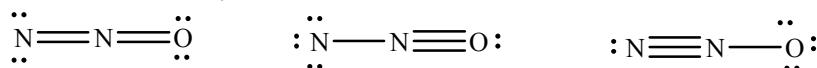
- 10.86 Plan: The basic Lewis structure will be the same for all species. The Cl atoms are larger than the F atoms. All of the molecules are of the type  $\text{AX}_5$  and have trigonal bipyramid molecular shape. The equatorial positions are in the plane of the triangle and the axial positions above and below the plane of the triangle. In this molecular shape, there is more room in the equatorial positions.

Solution:

- a) The F atoms will occupy the smaller axial positions first so that the larger Cl atoms can occupy the equatorial positions which are less crowded.  
b) The molecules containing only F or only Cl are nonpolar (have no dipole moment), as all the polar bonds would cancel. The molecules with one F or one Cl would be polar since there would be no corresponding bond to cancel the polarity. The presence of 2 axial F atoms means that their polarities will cancel (as would the 3 Cl atoms) giving a nonpolar molecule. The molecule with 3 F atoms is also polar.



- 10.87  $\text{N}_2\text{O}$  has 16 valence electrons; there are three resonance structures.



F.C.  $-1$   $+1$   $0$   $-2$   $+1$   $+1$   $0$   $+1$   $-1$

The third structure has a more reasonable distribution of formal charges. The third form has a strong triple bond between the N's and a weak N-O bond. It is easy to break the N-O bond which is why this compound easily decomposes to support combustion.

- 10.88 Plan: Count the valence electrons and draw Lewis structures for the resonance forms.

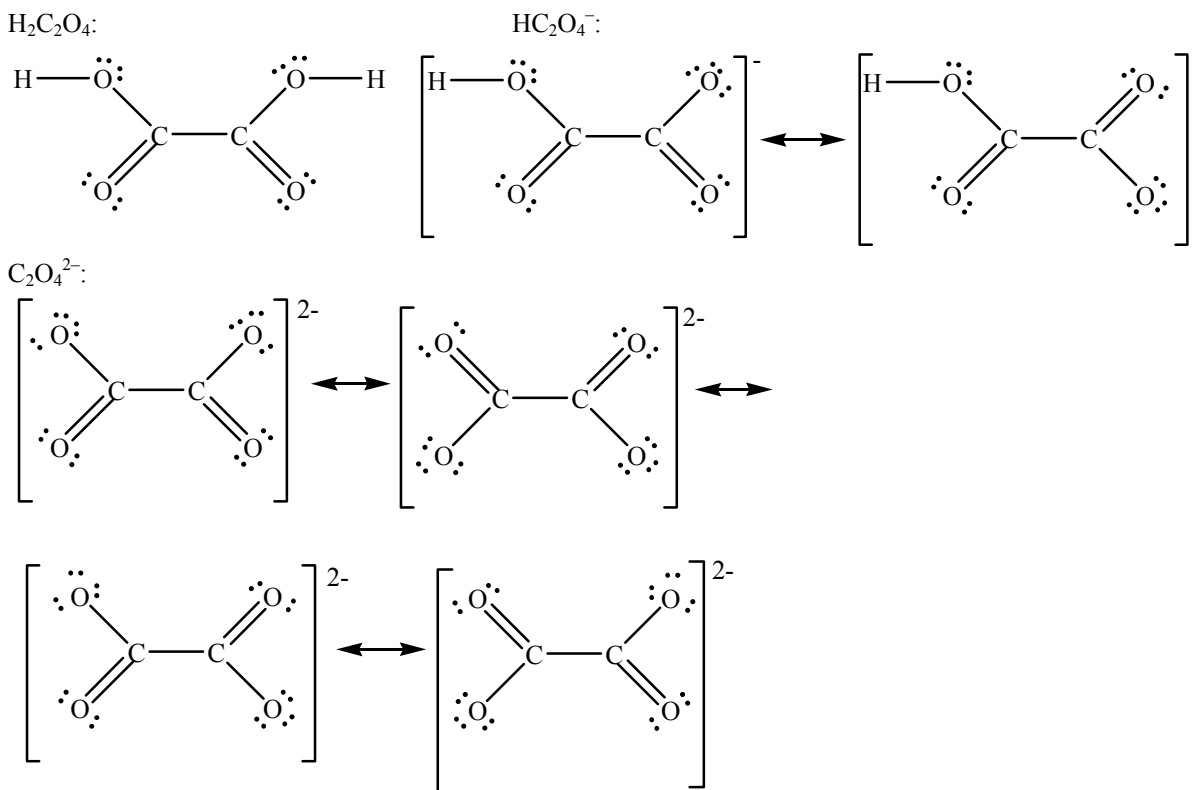
Solution:

The  $\text{H}_2\text{C}_2\text{O}_4$  molecule has  $[2 \times \text{H}(1 \text{ e}^-)] + [2 \times \text{C}(4 \text{ e}^-)] + [4 \times \text{O}(6 \text{ e}^-)] = 34$  valence electrons to be used in the Lewis structure. 14 of these electrons are used to bond the atoms with a single bond, leaving  $34 - 14 = 20$  electrons. If these 20 electrons are given to the oxygen atoms to complete their octet, the carbon atoms do not have octets. A lone pair from each of the oxygen atoms without hydrogen is changed to a bonding pair on C.

The  $\text{HC}_2\text{O}_4^-$  ion has  $[1 \times \text{H}(1 \text{ e}^-)] + [2 \times \text{C}(4 \text{ e}^-)] + [4 \times \text{O}(6 \text{ e}^-)] + [1 \text{ e}^- \text{ (from the charge)}] = 34$  valence electrons to be used in the Lewis structure. 12 of these electrons are used to bond the atoms with a single bond, leaving  $34 - 12 = 22$  electrons. If these 22 electrons are given to the oxygen atoms to complete their octet, the carbon atoms do not have octets. A lone pair from two of the oxygen atoms without hydrogen is changed to a bonding pair on C.

There are two resonance structures.

The  $\text{C}_2\text{O}_4^{2-}$  ion has  $[2 \times \text{C}(4 e^-)] + [4 \times \text{O}(6 e^-)] + [2 e^- \text{ (from the charge)}] = 34$  valence electrons to be used in the Lewis structure. 10 of these electrons are used to bond the atoms with a single bond, leaving  $34 - 10 = 24$  electrons. If these 24 electrons are given to the oxygen atoms to complete their octets, the carbon atoms do not have octets. A lone pair from each two oxygen atoms is changed to a bonding pair on C. There are four resonance structures.

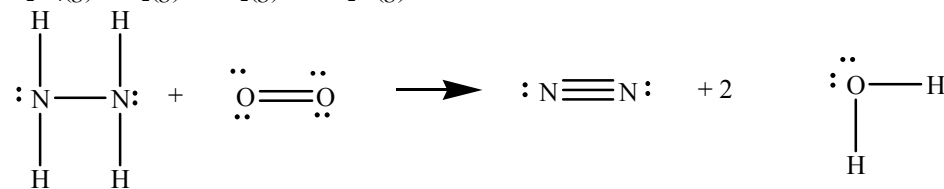
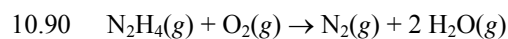
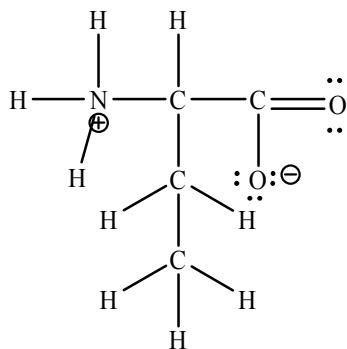


In  $\text{H}_2\text{C}_2\text{O}_4$ , there are 2 shorter  $\text{C}=\text{O}$  bonds and 2 longer  $\text{C}-\text{O}$  bonds.

In  $\text{HC}_2\text{O}_4^-$ , the  $\text{C}-\text{O}$  bonds on the side retaining the  $\text{H}$  remain as 1 long  $\text{C}-\text{O}$  and 1 shorter  $\text{C}=\text{O}$ . The  $\text{C}-\text{O}$  bonds on the other side of the molecule have resonance forms with an average bond order of 1.5, so they are intermediate in length and strength.

In  $\text{C}_2\text{O}_4^{2-}$ , all the carbon to oxygen bonds are resonating and have an average bond order of 1.5.

- 10.89 The molecule has 42 valence electrons. 30 electrons are already accounted for in the skeleton structure in the bonds.  $42 - 30 = 12$  valence electrons remain. If these 12 electrons are given to the two oxygen atoms to complete their octets, the carbon atom that is bonded to the two oxygen atoms does not have an octet. A lone pair from one of the oxygen atoms is changed to a bonding pair on the C. All the atoms have 0 formal charge except the N ( $\text{FC} = +1$ ), and the single bonded O ( $\text{FC} = -1$ )



|               |                                         |                                                                |               |                                |                                                     |
|---------------|-----------------------------------------|----------------------------------------------------------------|---------------|--------------------------------|-----------------------------------------------------|
| Bonds broken: | 4 x N-H<br>1 x N-N<br>1 x O=O<br>Total: | 4(391 kJ/mol)<br>1(160 kJ/mol)<br><u>498 kJ/mol</u><br>2222 kJ | Bonds formed: | 1 x N≡N<br>2 x 2 O-H<br>Total: | 1(-945 kJ/mol)<br>4(-467 kJ/mol)<br><u>-2813 kJ</u> |
|---------------|-----------------------------------------|----------------------------------------------------------------|---------------|--------------------------------|-----------------------------------------------------|

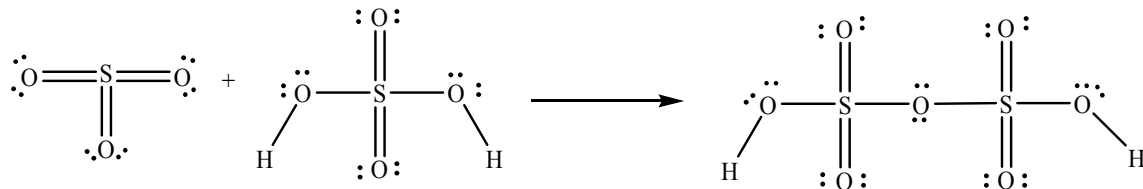
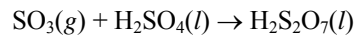
$$\Delta H^\circ_{\text{rxn}} = \Delta H^\circ_{\text{bonds broken}} + \Delta H^\circ_{\text{bonds formed}}$$

$$= 2222 + (-2813 \text{ kJ}) = -591 \text{ kJ per mole of } \text{N}_2\text{H}_4 \text{ reacted}$$

- 10.91 a) **Yes**, the black sphere can represent selenium.  $\text{SeF}_4$  has 34 valence electrons. Eight of these electrons are used in the four Se-F single bonds and 24 electrons are used to complete the octets of the F atoms. The remaining electron pair goes to selenium and the molecule is  $\text{AX}_4\text{E}$ . The molecular geometry is the see-saw molecular shape shown.
- b) **Yes**, the black sphere can represent nitrogen if the species is an anion with a  $-1$  charge. The  $\text{NF}_4^-$  ion has 34 valence electrons and would have the see-saw molecular shape as an  $\text{AX}_4\text{E}$  species.
- c) For  $\text{BrF}_4$  to have the 34 valence electrons needed for this see-saw molecular geometry, the charge of the species must be **+1**.  $\text{BrF}_4^+$  would have  $[1 \times \text{Br}(7e^-)] + [4 \times \text{F}(7e^-)] - [1 e^- \text{ from } + \text{ charge}] = 34$  valence electrons.

10.92 Plan: Draw the Lewis structures. Calculate the heat of reaction using the bond energies in Table 9.2.

Solution:



|               |                                    |                                                                   |               |                                    |                                                                       |
|---------------|------------------------------------|-------------------------------------------------------------------|---------------|------------------------------------|-----------------------------------------------------------------------|
| Bonds broken: | 5 S=O<br>2 S-O<br>2 O-H<br>Totals: | 5(552 kJ/mol)<br>2(265 kJ/mol)<br><u>2(467 kJ/mol)</u><br>4224 kJ | Bonds formed: | 4 S=O<br>4 S-O<br>2 O-H<br>Totals: | 4(-552 kJ/mol)<br>4(-265 kJ/mol)<br><u>2(-467 kJ/mol)</u><br>-4202 kJ |
|---------------|------------------------------------|-------------------------------------------------------------------|---------------|------------------------------------|-----------------------------------------------------------------------|

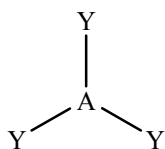
$$\Delta H^\circ_{\text{rxn}} = \Delta H^\circ_{\text{bonds broken}} + \Delta H^\circ_{\text{bonds formed}}$$

$$= 4224 \text{ kJ} + (-4202 \text{ kJ}) = 22 \text{ kJ}$$

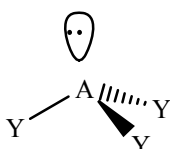
10.93 Plan: Pick the VSEPR structures for  $AY_3$  substances. Then determine which are polar.

Solution:

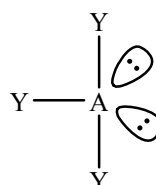
The molecular shapes that have a central atom bonded to three other atoms are trigonal planar, trigonal pyramid and T-shaped:



(a)  
3 groups  
( $AX_3$ )  
trigonal planar



(b)  
4 groups  
( $AX_3E$ )  
trigonal pyramid



(c)  
5 groups  
( $AX_3E_2$ )  
T-shaped

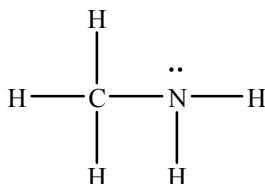
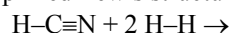
The trigonal planar molecules, such as (a), are nonpolar, so it cannot be  $AY_3$ . Trigonal pyramidal molecules (b) and T-shaped molecules (c) are polar, so either could represent  $AY_3$ .

10.94 a) Shape A is T-shaped ( $AX_3E_2$ ); Shape B is trigonal planar ( $AX_3$ ); Shape C is trigonal pyramid ( $AX_3E$ ).  $XeF_3^+$ , with 28 valence electrons, has two unshared pairs on Xe and is  $AX_3E_2$  and is the T-shaped molecular shape in A.  $SbBr_3$ , with 26 valence electrons, has one unshared pair on Sb; thus it is  $AX_3E$  and is the trigonal pyramidal molecular shape in C.  $GaCl_3$  with 24 valence electrons, has no unshared pairs on Ga; thus it is  $AX_3$  and is the trigonal planar shape in B.

b) Shapes **A and C** are polar.

c) Shape **A**, which is T-shaped, has the most valence electrons (10) around the central atom.

10.95 The simplified Lewis structures for the reaction are:

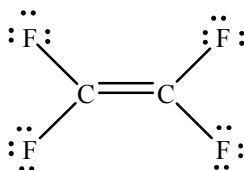
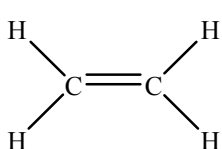


The energy to break the reactant bonds is:  $BE(C-H) + BE(C\equiv N) + 2 BE(H-H)$   
 $= (1 \text{ mol}) (413 \text{ kJ/mol}) + (1 \text{ mol}) (891 \text{ kJ/mol}) + (2 \text{ mol}) (432 \text{ kJ/mol})$   
 $= 2168 \text{ kJ}$

The energy from forming the product bonds is:  $3 BE(C-H) + BE(C-N) + 2 BE(N-H)$   
 $= (3 \text{ mol}) (-413 \text{ kJ/mol}) + (1 \text{ mol}) (-305 \text{ kJ/mol}) + (2 \text{ mol}) (-391 \text{ kJ/mol})$   
 $= -2326 \text{ kJ}$

$\Delta H^\circ_{\text{rxn}} = \Delta H^\circ_{\text{bonds broken}} - \Delta H^\circ_{\text{bonds formed}} = 2168 \text{ kJ} + (-2326 \text{ kJ}) = \mathbf{-158 \text{ kJ}}$

10.96 a)



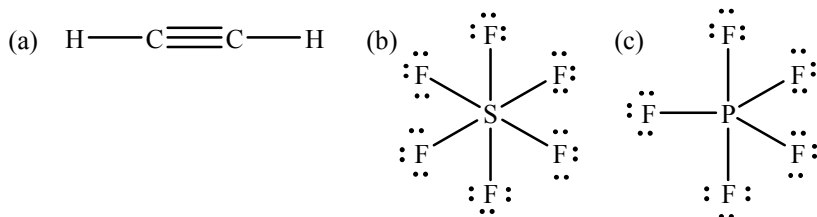
All the carbons are trigonal planar so the ideal angles should all be  $120^\circ$ .

b) The observed angles are slightly less than ideal because the  $C=C$  bond repels better than the single bonds. The larger F atoms cannot get as close together as the smaller H atoms, so the angles in tetrafluoroethylene are not reduced as much.

10.97 The top portion of the molecules are similar and the top portions will interact with biomolecules in a similar manner.

10.98 Plan: Draw the Lewis structure of each compound. Atoms  $180^\circ$  apart are separated by the sum of the bond's length. Atoms not at  $180^\circ$  apart must have their distances determined by geometrical relationships.

Solution:



a)  $\text{C}_2\text{H}_2$  has  $[2 \times \text{C}(4 e^-)] + [2 \times \text{H}(1 e^-)] = 10$  valence electrons to be used in the Lewis structure. 6 of these electrons are used to bond the atoms with a single bond, leaving  $10 - 6 = 4$  electrons. Giving one carbon atom the 4 electrons to complete its octet results in the other carbon atom not having an octet. The two lone pairs from the carbon with an octet are changed to two bonding pairs for a triple bond between the two carbon atoms. The molecular shape is linear. The H atoms are separated by two carbon-hydrogen bonds (109 pm) and a carbon-carbon triple bond (121 pm).

Total separation =  $2(109 \text{ pm}) + 121 \text{ pm} = \mathbf{339 \text{ pm}}$

b)  $\text{SF}_6$  has  $[1 \times \text{S}(6 e^-)] + [6 \times \text{F}(7 e^-)] = 48$  valence electrons to be used in the Lewis structure. 12 of these electrons are used to bond the atoms with a single bond, leaving  $48 - 12 = 36$  electrons. These 36 electrons are given to the fluorine atoms to complete their octets. The molecular shape is octahedral. The fluorine atoms on opposite sides of the S are separated by twice the sulfur-fluorine bond length (158 pm).

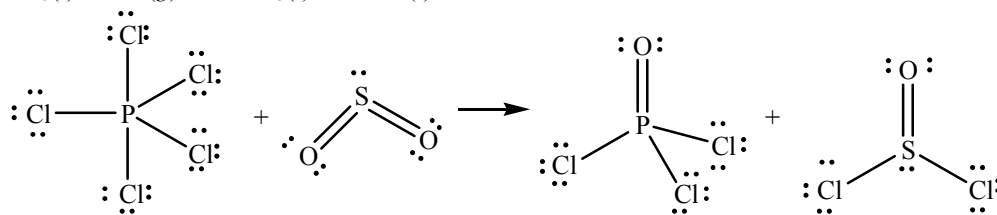
Total separation =  $2(158 \text{ pm}) = \mathbf{316 \text{ pm}}$

Adjacent fluorines are at two corners of a right triangle, with the sulfur at the  $90^\circ$  angle. Two sides of the triangle are equal to the sulfur-fluorine bond length (158 pm). The separation of the fluorine atoms is at a distance equal to the hypotenuse of this triangle. This length of the hypotenuse may be found using the Pythagorean Theorem ( $a^2 + b^2 = c^2$ ). In this case  $a = b = 158 \text{ pm}$ . Thus,  $c^2 = (158 \text{ pm})^2 + (158 \text{ pm})^2$ , and so  $c = 223.4457 = \mathbf{223 \text{ pm}}$

c)  $\text{PF}_5$  has  $[1 \times \text{P}(5 e^-)] + [5 \times \text{F}(7 e^-)] = 40$  valence electrons to be used in the Lewis structure. 10 of these electrons are used to bond the atoms with a single bond, leaving  $40 - 10 = 30$  electrons. These 30 electrons are given to the fluorine atoms to complete their octets. The molecular shape is trigonal bipyramid. Adjacent equatorial fluorine atoms are at two corners of a triangle with an F-P-F bond angle of  $120^\circ$ . The length of the P-F bond is 156 pm. If the  $120^\circ$  bond angle is A, then the F-F bond distance is a and the P-F bond distances are b and c. The F-F bond distance can be found using the Law of Cosines:  $a^2 = b^2 + c^2 - 2bc(\cos A)$ .

$a^2 = (156)^2 + (156)^2 - 2(156)(156)\cos 120^\circ$ .  $a = 270.1999 = \mathbf{271 \text{ pm}}$ .

10.99  $\text{PCl}_5(l) + \text{SO}_2(g) \rightarrow \text{POCl}_3(l) + \text{SOCl}_2(l)$



$\text{PCl}_5$ :  $\text{AX}_5$       **trigonal bipyramid**  
 $\text{SO}_2$ :  $\text{AX}_2\text{E}$       **bent**  
 $\text{POCl}_3$ :  $\text{AX}_4$       **tetrahedral**  
 $\text{SOCl}_2$ :  $\text{AX}_3\text{E}$       **trigonal pyramid**