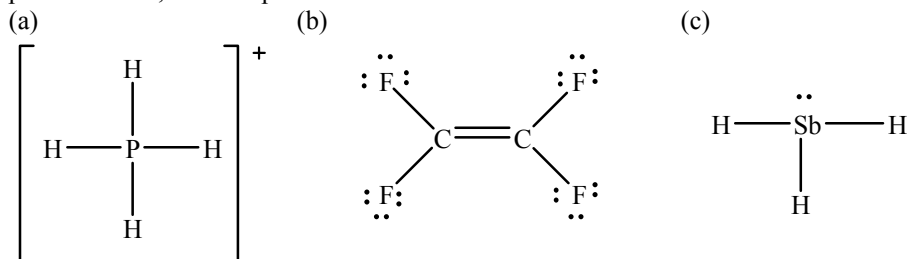




10.6 Plan: Count the valence electrons and draw Lewis structures.

Solution:

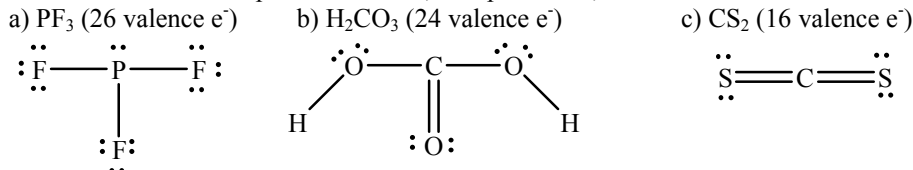
Total valence electrons:  $\text{PH}_4^+$  has 8;  $\text{C}_2\text{F}_4$  has 36; and  $\text{SbH}_3$  has 8. Ignoring H, the atom in the lower group number is central — P, C, and Sb. Added proof: H and F are never central. The two central C's must be adjacent. Place all the other atoms around the central atom. Split the F's 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 C's have 6 electrons each, but can achieve an octet by forming a double bond. The H's need no additional electrons. 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:

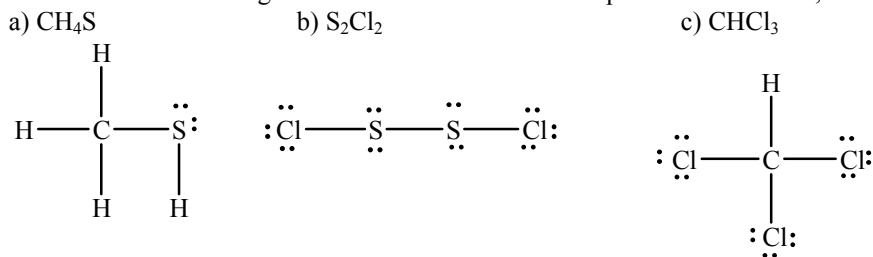
The P and the C's are central. Two of the O's in part b also are central. Place appropriate single bonds between all atoms. The O in part b without a hydrogen needs a double bond as do both the S's in part c. The remaining electrons serve as lone pairs. All atoms, except the H's, have octets.



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

Solution:

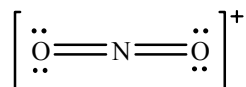
The C's and S's are central. The S in part (a) is attached to an H and the C. All atoms are attached with single bonds and the remaining electrons are divided into lone pairs. All the atoms, except H, have octets.



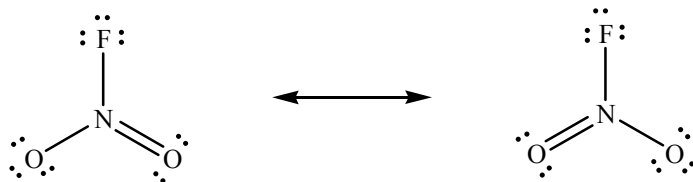
10.9 Plan: The problem asks for resonance structures, so there must be more than one answer for each part.

Solution:

a)  $\text{NO}_2^+$  has 16 valence electrons. No resonance is required as all atoms can achieve an octet with double bonds.



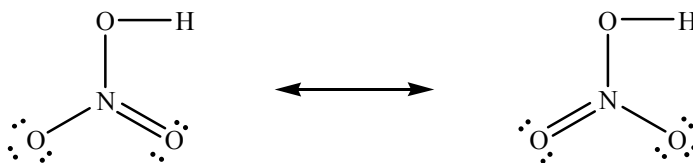
b)  $\text{NO}_2\text{F}$  is like  $\text{NO}_2$ , but has an extra atom (F) to interact with the lone electron on the N. In this structure all atoms have an octet.



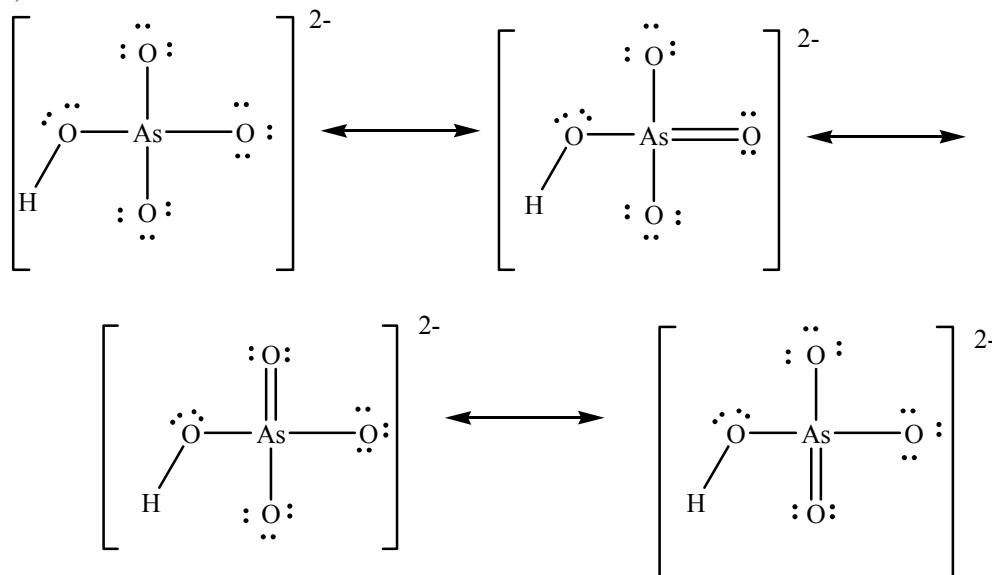
10.10 Plan: Count the valence electrons and draw Lewis structures. Additional structures are needed to show resonance.

Solution:

a)



b)

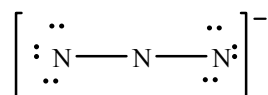


10.11 Plan: Count the valence electrons and draw Lewis structures. Additional structures are needed to show resonance.

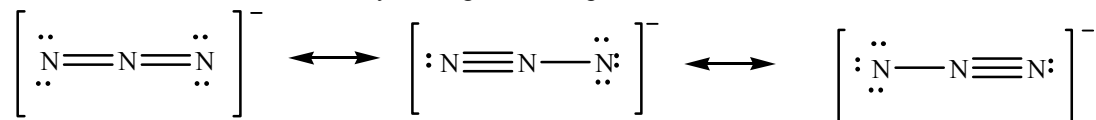
Solution:

a)  $\text{N}_3^-$  has 16 valence electrons.

The nitrogens must all be attached to each other with at least one bond. This uses 4 electrons leaving 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:



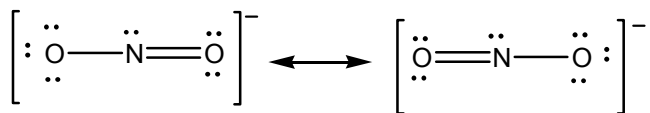
The central N needs 4 electrons. There are three options to do this: (1) one of the end N's could form a triple bond by sharing two of its lone pairs; (2) the other end N could form the triple bond instead; (3) each of the end N's could form a double bond to the central N by sharing one of its pairs.



Formal charges (FC) are shown for comparison.

b)  $\text{NO}_2^-$  has 18 valence electrons.

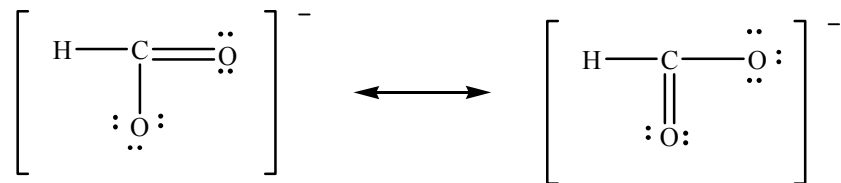
The nitrogen should be the central atom with each of the oxygens attached to it by a single bond ( $2 \times 2 = 4$  electrons). This leaves 14 electrons (7 pairs). If 3 pairs are given to each O and 1 pair is given to the N, then both O's have an octet, but the N only has 6. To complete an octet the N needs to gain a pair of electrons from one O or the other (from a double bond). The resultant structures are:



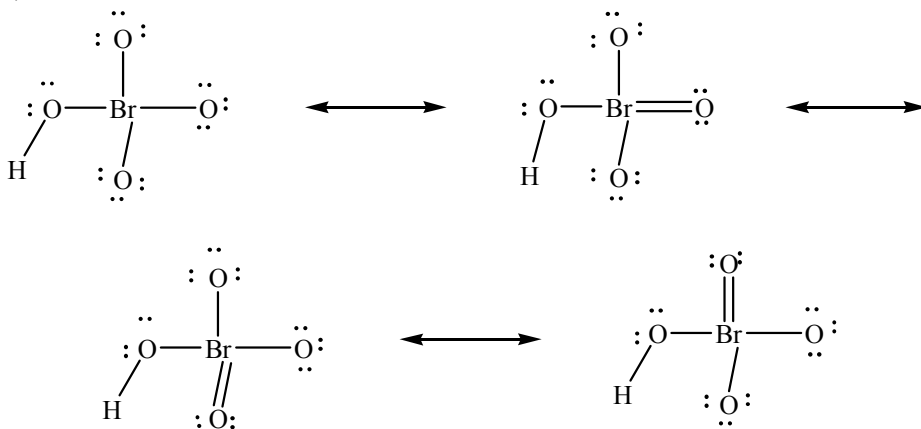
10.12 Plan: Follow the pattern established in the preceding problem.

Solution:

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



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



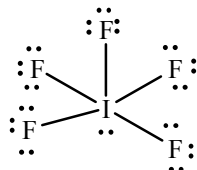
10.13 Plan: Initially, the method used in the preceding problems may be used to establish a Lewis structure. The total of the formal charges must equal the charge on an ion or be equal to 0 for a compound. The FC only needs to be calculated once for a set of identical atoms.

Solution:

a)  $\text{IF}_5$  has 42 valence electrons.

The presence of 5 F's 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 10 electrons leaving 32 electrons (16 pairs). Each F needs 3 pairs to complete an octet. The 5 F's use 15 of the 16 pairs, so there is 1 pair left for the central I.

This gives:



Calculating formal charges:

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

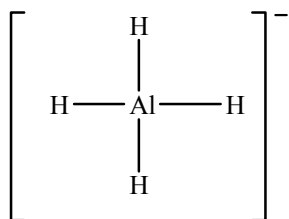
For iodine:  $\text{FC}_I = 7 - (2 + 1/2(10)) = 0$

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

Total formal charge = 0 = charge on the compound.

b)  $\text{AlH}_4^-$  has 8 valence electrons.

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



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

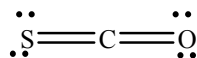
For each hydrogen  $\text{FC}_H = 1 - (0 + 1/2(2)) = 0$

Total formal charge = -1 = charge on the ion.

10.14 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.

Solution:

a) COS has 16 valence electrons.

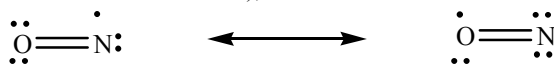


$\text{FC}_S = +6 - (4 + 1/2(4)) = 0$

$\text{FC}_C = +4 - (0 + 1/2(8)) = 0$

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

b) NO (has 11 valence electrons), the odd number means there will be an exception to the octet rule.



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

$\text{FC}_O = 6 - (3 + 1/2(4)) = +1$

$\text{FC}_N = 5 - (3 + 1/2(4)) = 0$

$\text{FC}_N = 5 - (4 + 1/2(4)) = -1$

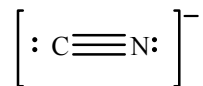
The first resonance structure has a better distribution of formal charges.

10.15 Plan: The same general method as in the preceding problems will be applied.

Solution:

a)  $\text{CN}^-$  has 10 valence electrons.

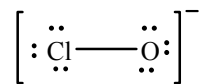
A triple bond between the two atoms plus a lone pair on each atom satisfies the octet rule and uses all 10 electrons.



$$\text{FC}_\text{C} = 4 - (2 + 1/2(6)) = -1 \quad \text{FC}_\text{N} = 5 - (2 + 1/2(6)) = 0$$

Check: The total formal charge equals the charge on the ion (-1).

b)  $\text{ClO}^-$  has 14 valence electrons.



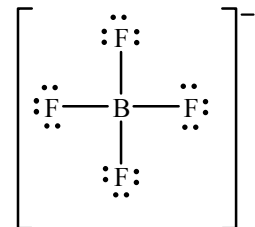
$$\text{FC}_\text{Cl} = 7 - (6 + 1/2(2)) = 0 \quad \text{FC}_\text{O} = 6 - (6 - 1/2(2)) = -1$$

Check: The total formal charge equals the charge on the ion (-1).

10.16 Plan: Use the same procedure as in the preceding problems.

Solution:

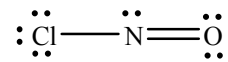
a)  $\text{BF}_4^-$  has 32 valence electrons.



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

$$\text{FC}_\text{B} = 3 - (0 - 1/2(8)) = -1$$

b)  $\text{ClNO}$  has 18 valence electrons.



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

$$\text{FC}_\text{N} = 5 - (2 + 1/2(6)) = 0$$

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

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

Solution:

a)  $\text{BrO}_3^-$  has 26 valence electrons.

Placing the O's around the central Br and forming 3 Br-O bonds uses 6 electrons and leaves 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's gives:

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

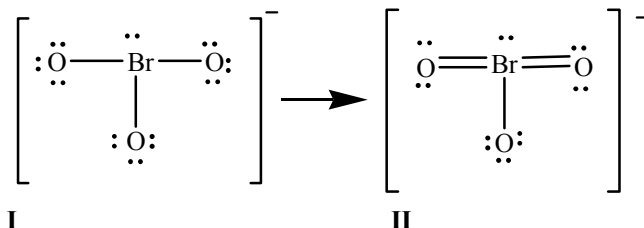
Double bonded O

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

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

Double bonded O's

Structure II has the more reasonable distribution of formal charges.  
 The oxidation numbers (O.N.) are:  $O.N._{Br} = +5$  and  $O.N._{O} = -2$ .



Check: The total formal charge equals the charge on the ion (-1).

b)  $SO_3^{2-}$  has 26 valence electrons.

Placing the O's around the central S and forming 3 S-O bonds uses 6 electrons and leaves 20 electrons (10 pairs).

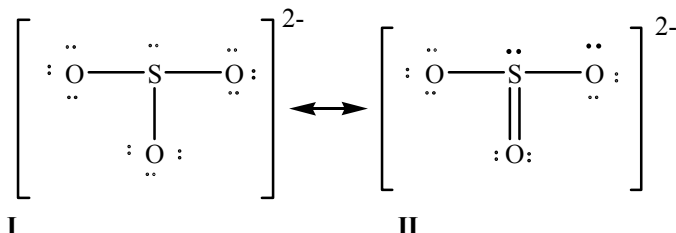
Placing 3 pairs on each O ( $3 \times 3 = 9$  total pairs) leaves 1 pair for the S and yields structure I below. In structure I all the atoms have a complete octet. Calculating formal charges:

$$FC_S = 6 - (2 + 1/2(6)) = +1 \quad FC_O = 6 - (6 + 1/2(2)) = -1$$

The  $FC_O$  is acceptable, but  $FC_S$  is larger than is usually acceptable. Forming a double bond between any one of the O's gives:

$$FC_S = 6 - (2 + 1/2(8)) = 0 \quad FC_O = 6 - (6 + 1/2(2)) = -1 \quad FC_O = 6 - (4 + 1/2(4)) = 0$$

Double bonded O



Structure II has the more reasonable distribution of formal charges.

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

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

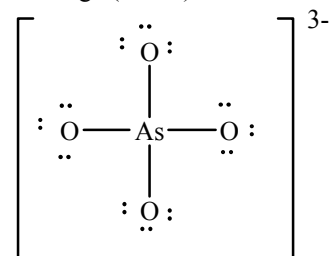
10.18 Plan: Follow the general procedure in problem 10.17.

Solution:

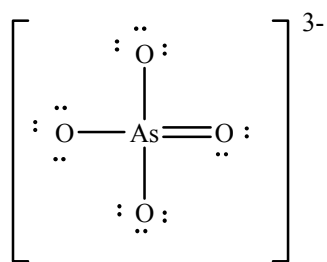
a)  $AsO_4^{3-}$  has 32 valence electrons.

$$FC_{As} = 5 - (0 + 1/2(8)) = +1 \quad FC_O = 6 - (6 + 1/2(2)) = -1 \text{ each}$$

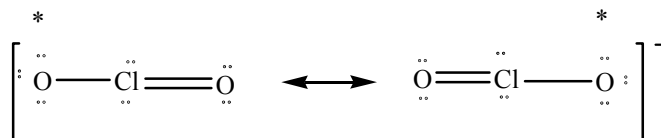
Net formal charge  $(+1 - 4) = -3$  The octet rule is followed by all atoms.



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



b)  $\text{ClO}_2^-$  has 20 valence electrons.

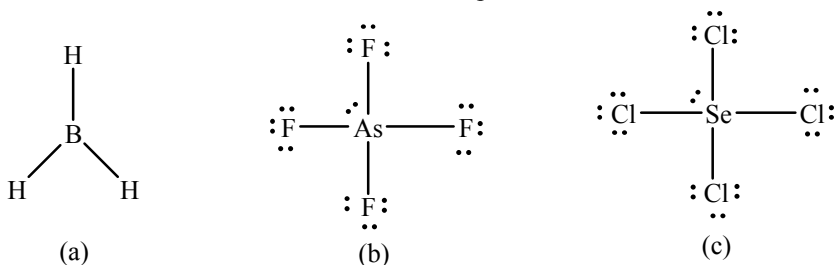


$FC_{Cl} = 7 - (4 - 1/2(6)) = 0$   
 $FC_O = 6 - (4 - 1/2(4)) = 0$   
 Only one answer is needed       $FC_{O^*} = 6 - (6 - 1/2(2)) = -1$   
 O.N.: O -2 each x 2 = -4 total; Cl +3

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

Solution:

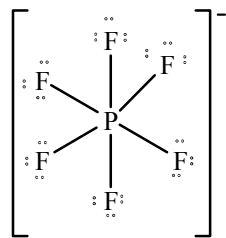
- a)  $\text{BH}_3$  has 6 valence electrons. These are used in 3 B-H bonds. The B has 6 electrons instead of an octet.  
 b)  $\text{AsF}_4^-$  has 34 valence electrons. Four As-F bonds use 8 electrons leaving 26 electrons (13 pairs). Each F needs 3 pairs to complete its octet and the remaining pair goes to the As. The As has an expanded octet with 10 electrons. The F cannot expand its octet.  
 c)  $\text{SeCl}_4$  has 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 Plan: We know that each of these substances does not obey the octet rule.

Solution:

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



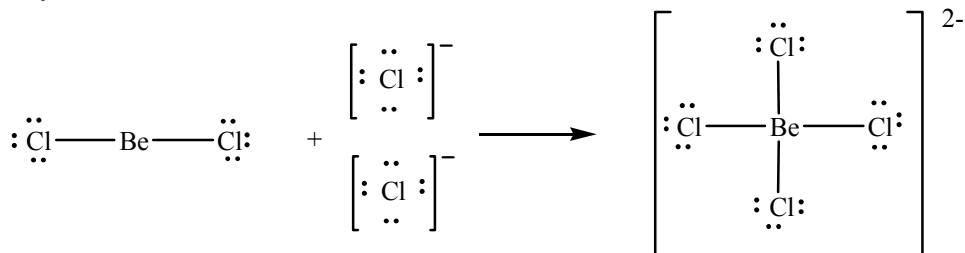




10.23 Plan: Draw Lewis structures for the reactants and products.

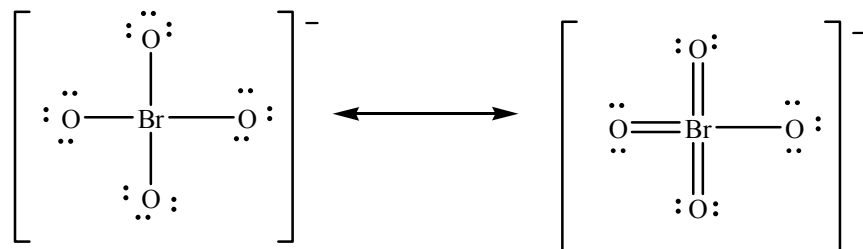
Solution:

Beryllium chloride has the formula  $\text{BeCl}_2$ . Chloride ion has the formula  $\text{Cl}^-$ .



10.24 Plan: Draw a Lewis structure. If the formal charges are not ideal, a second structure may be needed.

Solution:



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

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

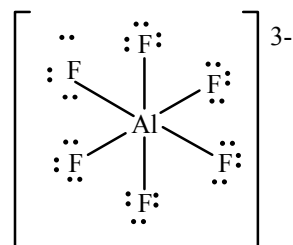
The structure on the right expands the valence shell of the Br to give more favorable formal charges.

$$\text{FC}_{\text{Br}} = 7 - (0 + 1/2(17)) = 0 \quad \text{FC}_{\text{O}} = 6 - (6 + 1/2(2)) = -1 \quad \text{FC}_{\text{O}} = 6 - (4 + 1/2(4)) = 0$$

Doubled bonded O's

10.25 Plan: Count the total valence electrons and draw a Lewis structure.

Solution:



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

Solution:

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

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

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

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

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

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

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

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

Structure A has the most reasonable set of formal charges.

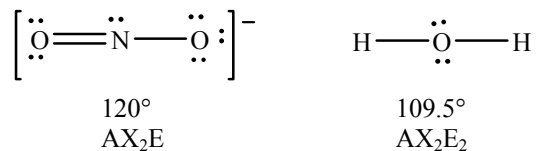
10.27 Plan: Determine the total number of valence electrons present. Next, draw a Lewis structure. Finally, use VSEPR or Valence-Bond theory to predict the shape.

10.28 The molecular shape and the electron-group arrangement are the same when no lone pairs are present.

10.29 Plan: A V-shaped molecule will have the stoichiometry  $AX_2$ , so only  $AX_2E_n$  geometries are needed.

Solution:

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  $AX_2E$  or  $AX_2E_2$ . Examples are:  $NO_2^-$  and  $H_2O$ .



10.30 Plan: Examine a list of all possible structures, and choose the ones needed.

Solution:

Tetrahedral	$AX_4$
Trigonal pyramidal	$AX_3E$
Bent or V-shaped	$AX_2E_2$

10.31 Solution:

The number of electron pairs governs the overall arrangement of the electrons. The superposition of the atoms on this arrangement gives rise to the molecular shape. The actual molecular shape reflects the positions of the atoms.

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.

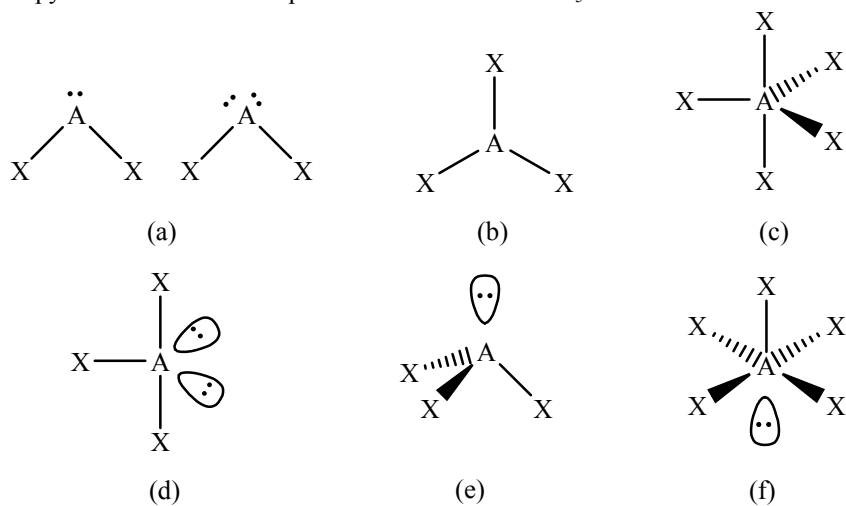
b) A trigonal planar molecule follows the formula  $AX_3$  and has no lone electron pairs.

c) A trigonal bipyramidal molecule contains five electron groups (single bonds) and no lone pairs ( $AX_5$ ).

d) A T-shape molecule has three bonding groups and two lone pairs ( $AX_3E_2$ ).

e) A trigonal pyramidal molecule follows the formula  $AX_3E$ .

f) A square pyramidal molecule shape follows the formula  $AX_5E$ .



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

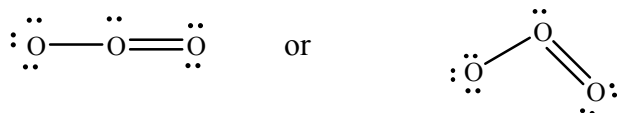
Solution:

a) AX <sub>3</sub> E	tetrahedral	109.5°	smaller
b) AX <sub>2</sub>	linear	180°	none
c) AX <sub>3</sub>	trigonal planar	120°	none
d) AX <sub>2</sub> E <sub>2</sub>	tetrahedral	109.5°	smaller
e) AX <sub>2</sub>	linear	180°	none
f) AX <sub>4</sub> E	trigonal bipyramidal	120°, 90°	smaller

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

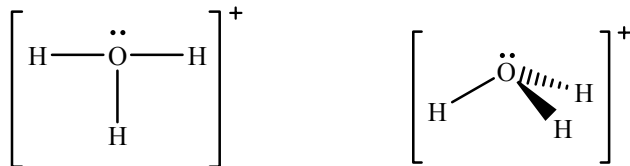
Solution:

a) The molecule has 18 electrons. This gives the following Lewis structure:



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

b) This ion has 8 electrons. This gives the following Lewis structure:



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

c) The molecule has 26 electrons. This gives the following Lewis structure:

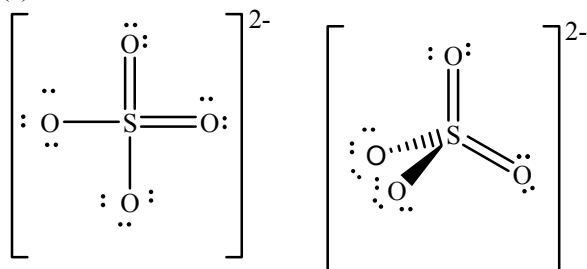


There are four groups around the N, one of which is a lone pair. This gives a **tetrahedral** electron-group arrangement, a **trigonal pyramidal** molecular shape, and an ideal bond angle of **109.5°**.

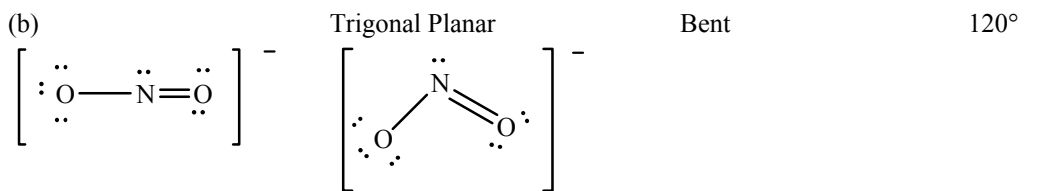
10.35 Plan: First, draw a Lewis structure, and then apply VSEPR.

Solution:

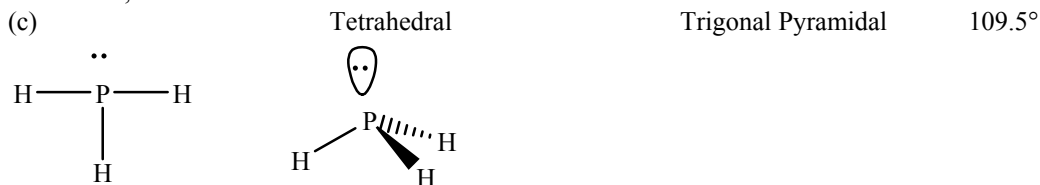
Lewis Structure	Electron-group Arrangement	Molecular Shape	Ideal Bond Angle
(a)	Tetrahedral	Tetrahedral	109.5°



In addition, there are other resonance forms.



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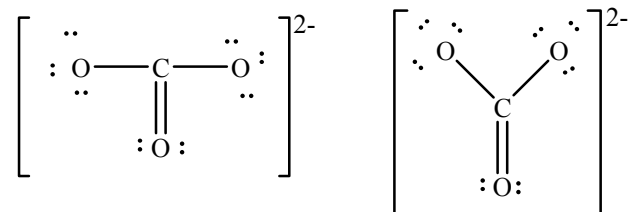


10.36 Plan: First, draw a Lewis structure, and then apply VSEPR.

Solution:

Lewis Structure	Electron-group Arrangement	Molecular Shape	Ideal Bond Angle
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(a) (All resonance forms)	Trigonal Planar	Trigonal Planar	120°
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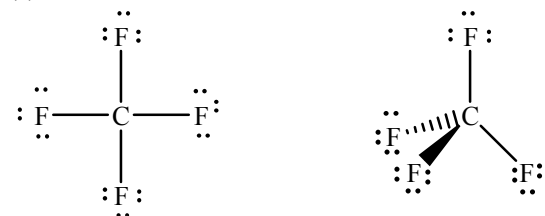


+ 2 additional resonance forms

(b)	Trigonal Planar	Bent (V-shaped)	120°
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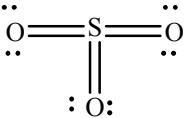
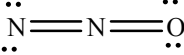
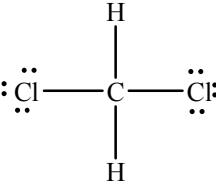


(c)	Tetrahedral	Tetrahedral	109.5°
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10.37 Plan: First, draw a Lewis structure; then apply VSEPR.

Solution:

Lewis Structure	Electron-group Arrangement	Molecular Shape	Ideal Bond Angle
	Trigonal Planar	Trigonal Planar	120°
	Linear	Linear	180°
	Tetrahedral	Tetrahedral	109.5°

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 Plan: Examine the structure shown, and then apply VSEPR.

Solution:

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. The structures are shown after the discussion.

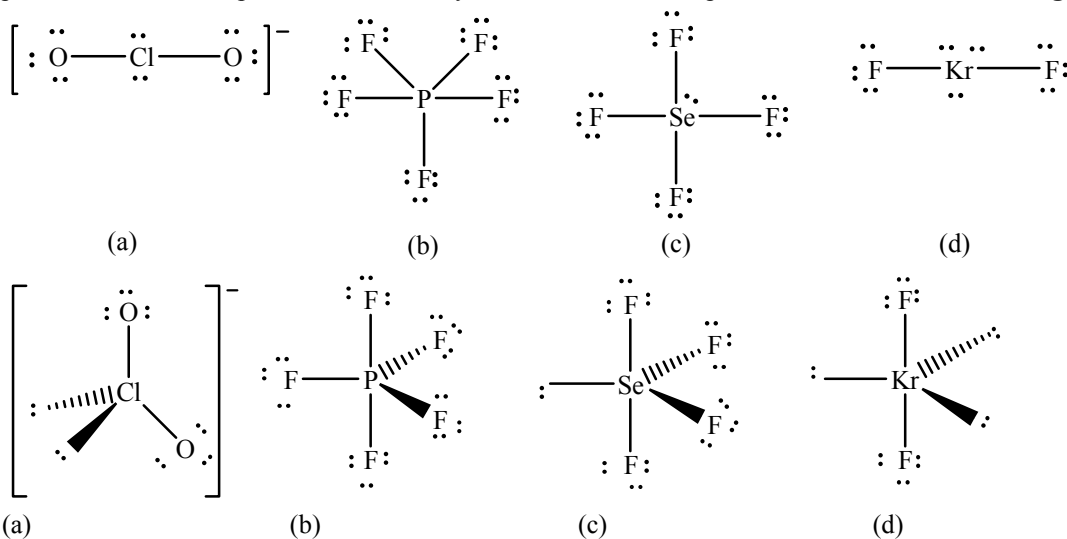
Solution:

a) The  $\text{ClO}_2^-$  ion has 20 valence electrons. The Cl is the central atom. There are two bonds (to the O's) and two lone pairs on the Cl ( $\text{AX}_2\text{E}_2$ ). The shape is **bent** (or V-shape). The structure is based on a tetrahedral electron-group arrangement with an ideal bond angle of  $109.5^\circ$ . The presence of the lone pairs will cause the remaining angles to be **less than  $109.5^\circ$** .

b) The  $\text{PF}_5$  molecule has 40 valence electrons. The P is the central atom. There are 5 bonds to the P and no lone pairs ( $\text{AX}_5$ ). The shape is **trigonal bipyramidal**. The ideal bond angles are  $90^\circ$  and  $120^\circ$ . The absence of lone pairs means the **angles are ideal**.

c) The  $\text{SeF}_4$  molecule has 34 valence electrons. The Se is the central atom. There are 4 bonds to the Se which also has a lone pair ( $\text{AX}_4\text{E}$ ). The shape is **see-saw**. The structure is based on a trigonal bipyramidal structure with ideal angles of  $90^\circ$  and  $120^\circ$ . The presence of the lone pairs means the angles are **less than ideal**.

d) The  $\text{KrF}_2$  molecule has 22 valence electrons. The Kr is the central atom. There are 2 bonds to the Kr which also has 3 lone pairs ( $\text{AX}_2\text{E}_3$ ). The shape is **linear**. The structure is based on a trigonal bipyramidal structure with ideal angles of  $90^\circ$  and  $120^\circ$ . The placement of the F's makes their ideal bond angle to be  $2 \times 90^\circ = 180^\circ$ . The placement of the lone pairs is such that they cancel each other's repulsion, thus the actual **bond angle is ideal**.



10.41 Plan: The Lewis structures must be drawn, and VSEPR applied to the structures. The structures are shown after the discussion.

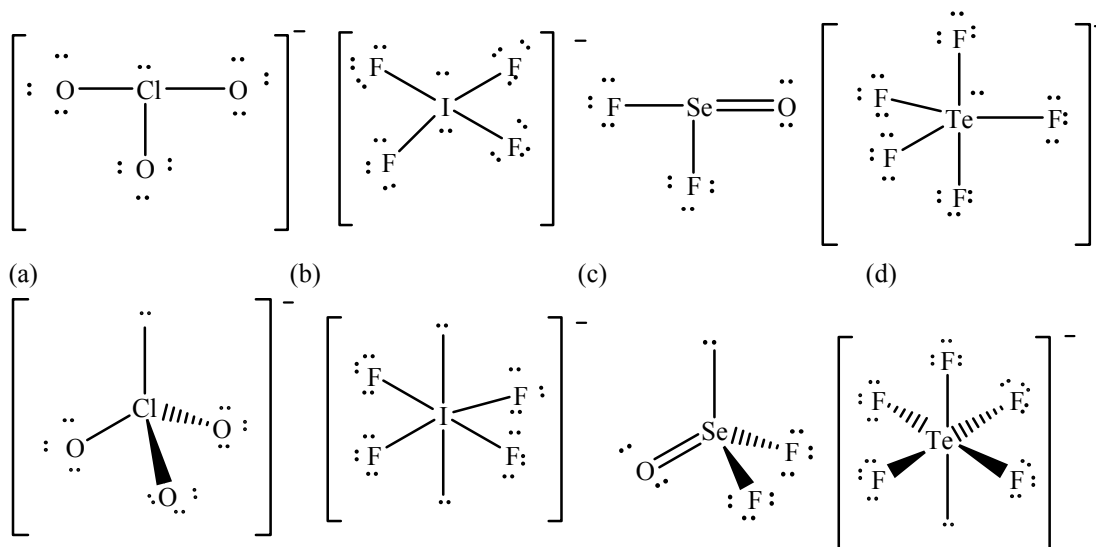
Solution:

a) The  $\text{ClO}_3^-$  ion has 26 valence electrons. The Cl is the central atom. There are 3 bonds (to the O's) 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**.

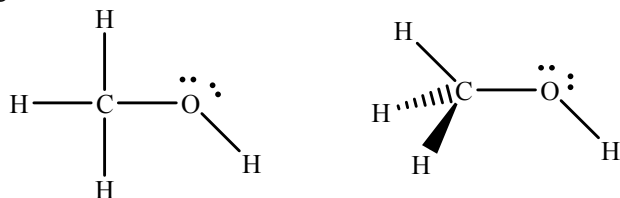
d) The  $\text{TeF}_5^-$  ion has 42 valence electrons. The Te is the central atom. There are 5 bonds to the Te which also has 1 lone pair ( $\text{AX}_5\text{E}$ ). The shape is **square pyramidal**. The structure is based on an octahedral with ideal angles of  $90^\circ$ . The presence of the lone pair means the angles are **less than ideal**.



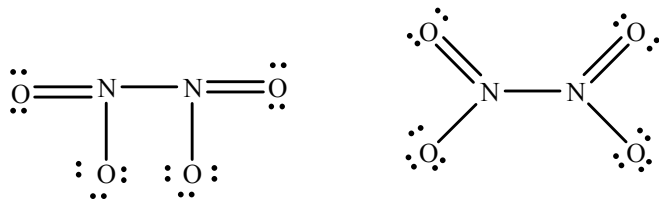
10.42 Plan: The Lewis structures must be drawn, and VSEPR applied to the structures.

Solution:

a) In the  $\text{CH}_3\text{OH}$  molecule, both carbon and oxygen serve as central atoms. (H can never be central.) 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) 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 bonds 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.

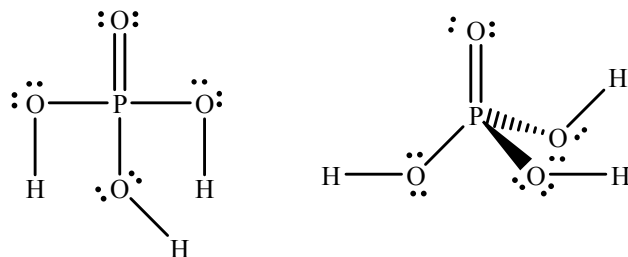




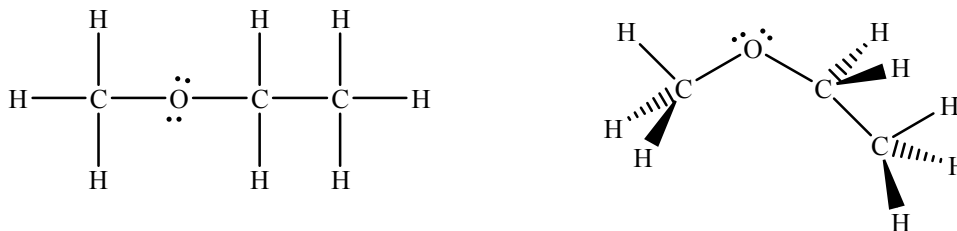
10.43 Plan: The Lewis structures must be drawn, and VSEPR applied to the structures.

Solution:

a) In the  $\text{H}_3\text{PO}_4$  molecule the P and each of the O's 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's 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  $109.5^\circ$ .



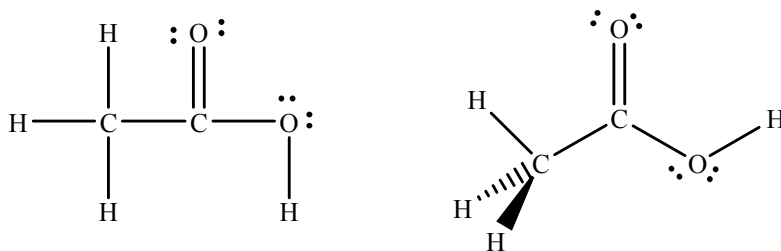
b) In the  $\text{CH}_3\text{OCH}_2\text{CH}_3$  molecule, all atoms except the hydrogens serve as central atoms. All the carbons have 4 bonds and no lone pairs ( $\text{AX}_4$ ), so they are tetrahedral with no deviation from the ideal bond angle of  $109.5^\circ$ . The oxygen has 2 bonds and 2 lone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



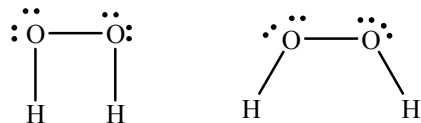
10.44 Plan: The Lewis structures must be drawn, and VSEPR applied to the structures.

Solution:

a) In the  $\text{CH}_3\text{COOH}$  molecule, the carbons and the O with H attached serve as central atoms. The carbon labeled  $\text{C}_1$  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 labeled  $\text{C}_2$  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  $109.5^\circ$ .



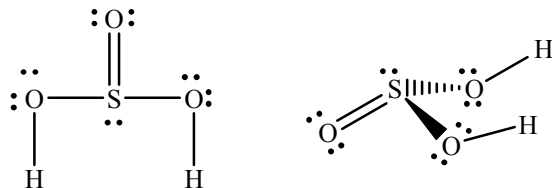
b) In the  $\text{H}_2\text{O}_2$  molecule, all oxygens serve as central atoms. Both O's have 2 bonds and 2 lone pairs ( $\text{AX}_2\text{E}_2$ ), so they are **V-shaped** or **bent** with angles **less than the ideal** values of  $109.5^\circ$ .



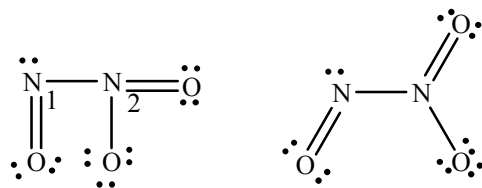
10.45 Plan: The Lewis structures must be drawn, and VSEPR applied to the structures.

Solution:

a) In the  $\text{H}_2\text{SO}_3$  molecule, the S and the O's with an H attached serve as central atoms. The S has 3 groups and no lone pairs ( $\text{AX}_3$ ), so it is **trigonal pyramidal** with **no deviation** from the ideal angle of  $120^\circ$ . The H bearing O's 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  $109.5^\circ$ .

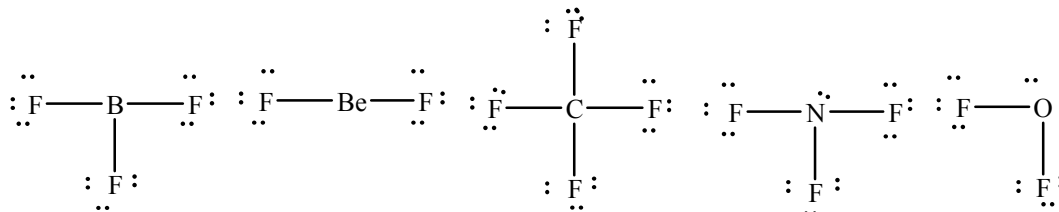


b) The  $\text{N}_2\text{O}_3$  molecule has the structure indicated in the problem with the N's serving as central atoms. The nitrogen labeled  $\text{N}_1$  has 2 bonds and a lone pair ( $\text{AX}_2\text{E}$ ), so it is **V-shaped** or **bent** with angles **less than the ideal** values of  $120^\circ$ . 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:

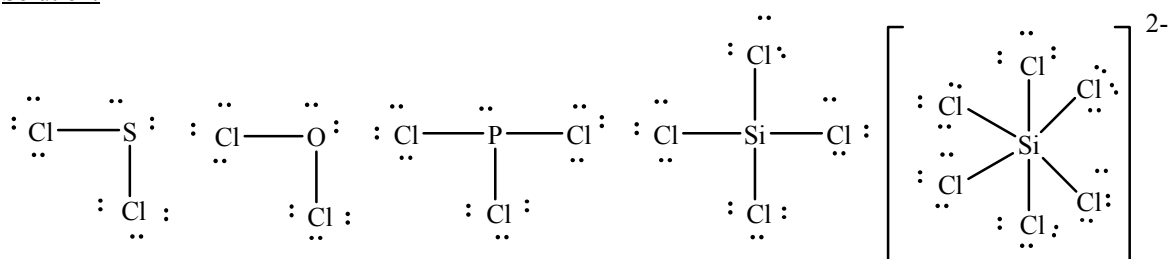


Bond angles:  $\text{OF}_2 < \text{NF}_3 < \text{CF}_4 < \text{BF}_3 < \text{BeF}_2$

$\text{BeF}_2$  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 Plan: First, draw a Lewis structure, and then apply VSEPR.

Solution:



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

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 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  $\text{SiCl}_2$  and  $\text{OCl}_2$ . 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 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 Plan: The ideal bond angles depend on the electron-group arrangement. Deviations depend on lone pairs.

Solution:

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 3 groups, and with no lone pairs the angle will be ideal and equal to  $120^\circ$ .

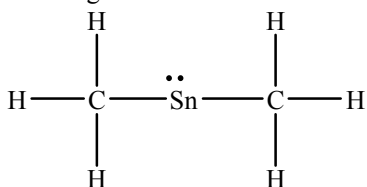
c) The C with 3 groups will have angles that are ideal ( $120^\circ$ ). The O, with the H attached, has 4 groups. The presence of 4 groups gives an ideal angle of  $109.5^\circ$ , which is reduced by the lone pairs.

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

Solution:

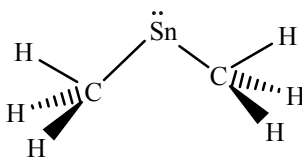
a) Type:  $\text{AX}_2\text{E}$

Ideal angle:  $120^\circ$



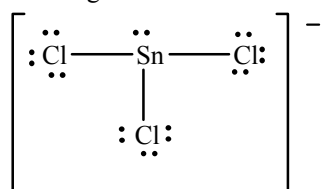
Shape: bent

Actual angle:  $<120^\circ$  (because of the lone pair)



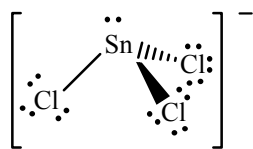
b) Type:  $\text{AX}_3\text{E}$

Ideal angle:  $109.5^\circ$

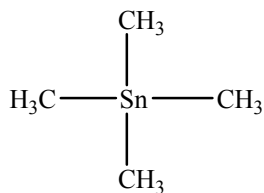


Shape: trigonal pyramidal

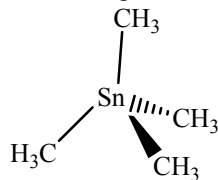
Actual angle:  $<109.5^\circ$  (because of the lone pair)



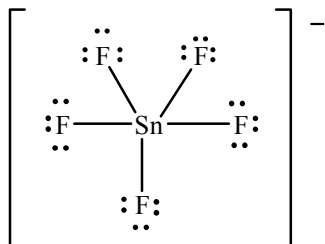
c) Type:  $AX_4$   
Ideal angle:  $109.5^\circ$



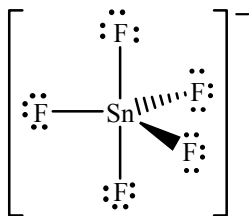
Shape: tetrahedral  
Actual angle:  $<109.5^\circ$  (because of the lone pair)



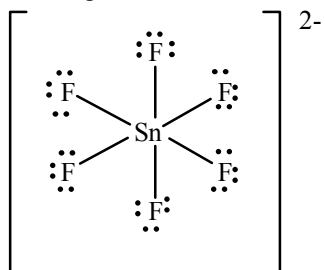
d) Type:  $AX_5$   
Ideal angles:  $120^\circ$  and  $90^\circ$



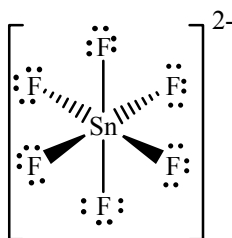
Shape: trigonal pyramidal  
Actual angle:  $120^\circ$  and  $90^\circ$  (there are no lone pairs)



e) Type:  $AX_6$   
Ideal angles:  $90^\circ$



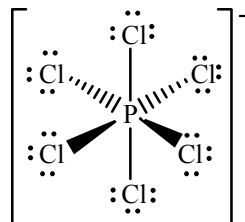
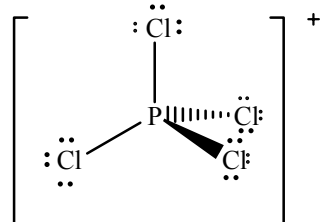
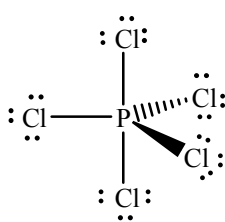
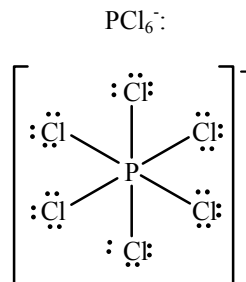
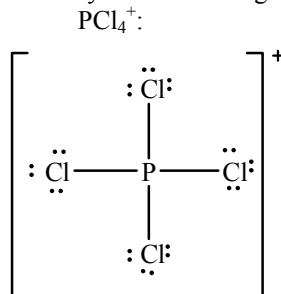
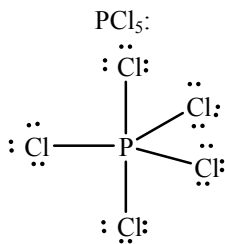
Shape: octahedral  
Actual angle:  $90^\circ$  (there are no lone pairs)



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  $\text{PCl}_5$  is  $\text{AX}_5$ , so the shape is trigonal bipyramidal, and the angles are  $120^\circ$  and  $90^\circ$ .  
 The  $\text{PCl}_4^+$  is  $\text{AX}_4$ , so the shape is tetrahedral, and the angles are  $109.5^\circ$ .  
 The  $\text{PCl}_6^-$  is  $\text{AX}_6$ , so the shape is octahedral, and the angles are  $90^\circ$ .  
 Half the  $\text{PCl}_5$  (trigonal bipyramidal,  $120^\circ$  and  $90^\circ$ ) become tetrahedral  $\text{PCl}_4^+$  (tetrahedral,  $109.5^\circ$ ), and the other half become octahedral  $\text{PCl}_6^-$  (octahedral,  $90^\circ$ ).

10.52 Plan: Apply VSEPR.

Solution:

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 Plan: Apply VSEPR.

Solution:

When polar bonds are arranged so that their polarities cancel, the molecule is nonpolar. An example is  $\text{BF}_3$  where the 3 polar B-F bonds are pointed towards the corners of an equilateral triangle. The polarities of the 3 B-F bonds cancel.

10.54 Plan: Apply VSEPR.

Solution:

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, 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
$\text{SCl}_2$	S-Cl	S = 2.5 Cl = 3.0	$3.0 - 2.5 = 0.5$
$\text{F}_2$	F-F	F = 4.0 F = 4.0	$4.0 - 4.0 = 0.0$
$\text{CS}_2$	C-S	C = 2.5 S = 2.5	$2.5 - 2.5 = 0.0$
$\text{CF}_4$	C-F	C = 2.5 F = 4.0	$4.0 - 2.5 = 1.5$
$\text{BrCl}$	Br-Cl	Br = 2.8 Cl = 3.0	$3.0 - 2.8 = 0.2$

The polarities of the bonds increases in the order:  $\text{F-F} = \text{C-S} < \text{Br-Cl} < \text{S-Cl} < \text{C-F}$ . Thus,  $\text{CF}_4$  has the most polar bonds.

b) The  $\text{F}_2$  and  $\text{CS}_2$  cannot be polar since they do not have polar bonds.  $\text{CF}_4$  is a  $\text{AX}_4$  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.  **$\text{SCl}_2$  has a dipole moment** (is polar) because it is a bent molecule,  $\text{AX}_2\text{E}_2$ , and the S-Cl bonds both pull to one side.

10.56 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, it must have polar bonds, and a certain shape determined by VSEPR.

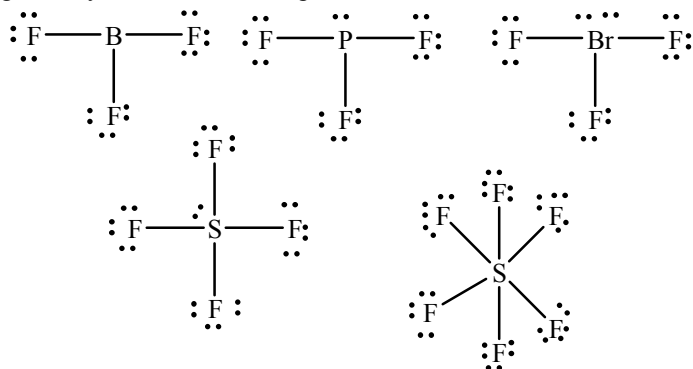
Solution:

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

Molecule	Bond	Electronegativities	Electronegativity difference
$\text{BF}_3$	B-F	B = 2.0 F = 4.0	$4.0 - 2.0 = 2.0$
$\text{PF}_3$	P-F	P = 2.1 F = 4.0	$4.0 - 2.1 = 1.9$
$\text{BrF}_3$	Br-F	Br = 2.8 F = 4.0	$4.0 - 2.8 = 1.2$
$\text{SF}_4$	S-F	S = 2.5 F = 4.0	$4.0 - 2.5 = 1.5$
$\text{SF}_6$	S-F	S = 2.5 Cl = 4.0	$4.0 - 2.5 = 1.5$

The polarities of the bonds are increasing in the order:  $\text{Br-F} < \text{S-F} < \text{P-F} < \text{B-F}$ . Thus,  **$\text{BF}_3$**  has the most polar bonds.

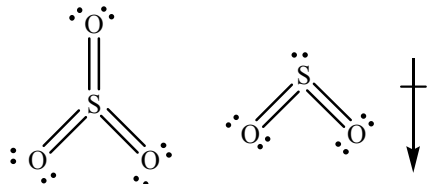
b) All the molecules meet the requirement of having polar bonds. The arrangement of the bonds must be considered in each case.  $\text{BF}_3$  is trigonal planar,  $\text{AX}_3$ , so it is nonpolar because the polarities of the bonds cancel.  **$\text{PF}_3$  has a dipole moment** (is polar) because it has a trigonal pyramidal geometry,  $\text{AX}_3\text{E}$ .  **$\text{BrF}_3$  has a dipole moment** because it has a T-shaped geometry,  $\text{AX}_3\text{E}_2$ .  **$\text{SF}_4$  has a dipole moment** because it has a see-saw geometry,  $\text{AX}_4\text{E}$ .  $\text{SF}_6$  is nonpolar because it is octahedral,  $\text{AX}_6$ , 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.

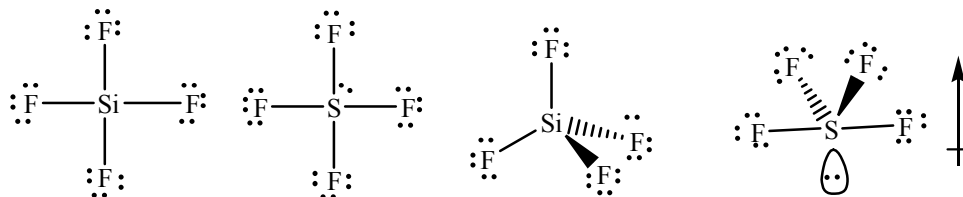
Solution:

a) All the bonds are polar covalent. The  $\text{SO}_3$  molecule is trigonal planar,  $\text{AX}_3$ , so the bond dipoles cancel leading to a nonpolar molecule (no dipole moment). The  $\text{SO}_2$  molecule is bent,  $\text{AX}_2\text{E}$ , so the polar bonds both pull to one side.  **$\text{SO}_2$  has a greater dipole moment** because it is the only one of the pair that is polar.



b)  $\text{ICl}$  and  $\text{IF}$  are polar, as are all diatomic molecules composed of atoms with differing electronegativities. The electronegativity difference for  $\text{ICl}$  is less than that for  $\text{IF}$ . The greater difference means that  **$\text{IF}$  has a greater dipole moment**.

c) All the bonds are polar covalent. The  $\text{SiF}_4$  molecule is nonpolar (has no dipole moment) because the bonds are arranged tetrahedrally,  $\text{AX}_4$ .  $\text{SF}_4$  is  $\text{AX}_4\text{E}$ , so it has a see-saw shape, where the bond dipoles do not cancel.  **$\text{SF}_4$  has the greater dipole moment**.



d)  $\text{H}_2\text{O}$  and  $\text{H}_2\text{S}$  have the same basic structure. They are both bent molecules,  $\text{AX}_2\text{E}_2$ , and as such, they are polar. The electronegativity difference in  $\text{H}_2\text{O}$  is greater so  **$\text{H}_2\text{O}$  has a greater dipole moment**.

10.58 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. Resonance structures are not required.

Solution:

a) All the bonds are polar covalent. Both the molecules are bent ( $\text{SO}_2$  and  $\text{ClO}_2$  are  $\text{AX}_2\text{E}_2$ ) and, therefore, polar.  $\text{SO}_2$  has a greater electronegativity difference so  **$\text{SO}_2$  has a greater dipole moment**.

b) HBr and HCl are polar, as are all diatomic molecules composed of atoms with differing electronegativities. The electronegativity difference for HBr is less than that for HCl. The greater difference means that **HCl has a greater dipole moment**.

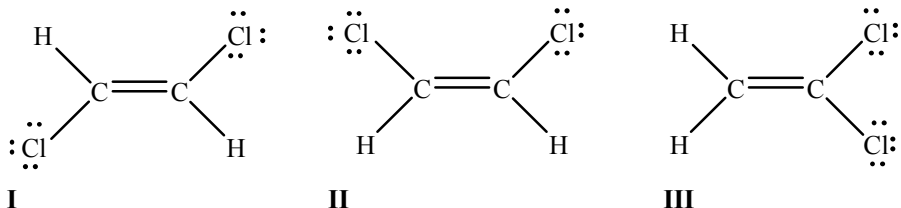
c) All the bonds are polar covalent. The  $\text{BeCl}_2$  molecule is nonpolar (has no dipole moment) because the bonds are arranged linearly,  $\text{AX}_2$ .  $\text{SCl}_2$  is  $\text{AX}_2\text{E}_2$ , so it has a bent shape, where the bond dipoles do not cancel.  **$\text{SF}_4$  has the greater dipole moment**.

d) All the bonds are polar covalent.  $\text{AsF}_5$  is  $\text{AX}_5$ , so it is trigonal bipyramidal and nonpolar.  $\text{AsF}_3$  is  $\text{AX}_3\text{E}$ , so it is trigonal pyramidal and polar.  **$\text{AsF}_3$  has a greater dipole moment**.

10.59 Plan: Draw Lewis structures, and then apply VSEPR.

Solution:

There are 3 possible structures for the compound  $\text{C}_2\text{H}_2\text{Cl}_2$ :



The presence of the double bond prevents rotation about the  $\text{C}=\text{C}$  bond, so the structures are “fixed.” The  $\text{C}-\text{Cl}$  bonds are more polar than the  $\text{C}-\text{H}$  bonds, so the key to predicting the polarity is the positioning of the  $\text{C}-\text{Cl}$  bonds. Compound I has the  $\text{C}-\text{Cl}$  bonds arranged so that they cancel leaving I as a nonpolar molecule. Both II and III have  $\text{C}-\text{Cl}$  bonds on the same side so the bonds work together making both molecules polar. Both I and II will react with  $\text{H}_2$  to give a compound with a  $\text{Cl}$  attached to each  $\text{C}$  (same product). Compound III will react with  $\text{H}_2$  to give a compound with 2  $\text{Cl}$ 's on one  $\text{C}$  and none on the other (different product). **Compound I must be X** as it is the only one that is nonpolar (has no dipole moment). **Compound II must be Z** because it is polar and gives the same product as compound X. This means that compound III must be the remaining compound — Y.

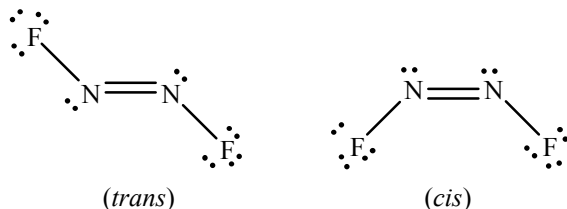
**Compound Y (III) has a dipole moment.**

10.60 Plan: Draw Lewis structures, and then apply VSEPR.

Solution:

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

a)

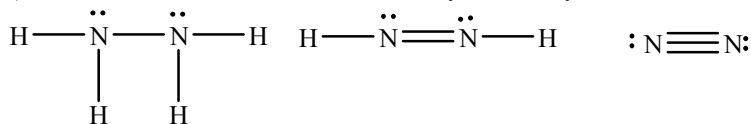


b) In the *trans*-form the  $\text{N}-\text{F}$  bonds pull equally in opposite directions, thus, they cancel and the molecule is nonpolar. The  $\text{N}-\text{F}$  bonds in the *cis*-form pull in the same general direction resulting in a polar molecule.

10.61 Plan: The Lewis structures are needed to do this problem.

Solution:

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



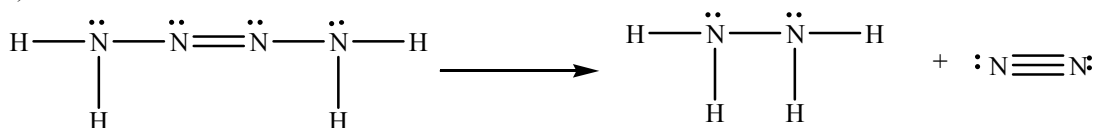
Hydrazine

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)



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

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

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

$$= 2302 \text{ kJ}$$

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

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

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

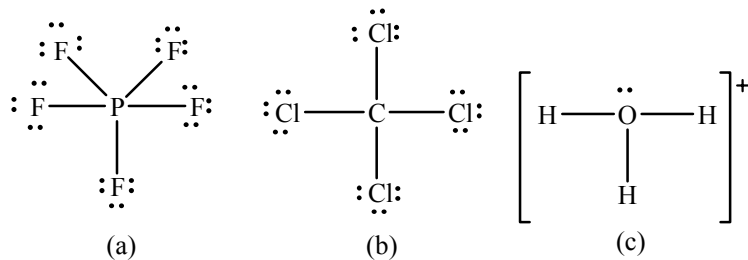
$$= 2669 \text{ kJ}$$

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

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 Plan: Count the total number of valence electrons, then arrange the atoms for bonds, and finally try to follow the octet rule.

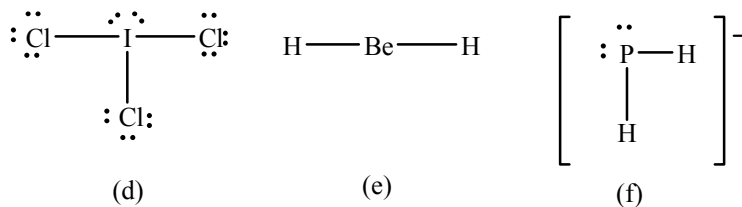
Solution:



(a)

(b)

(c)

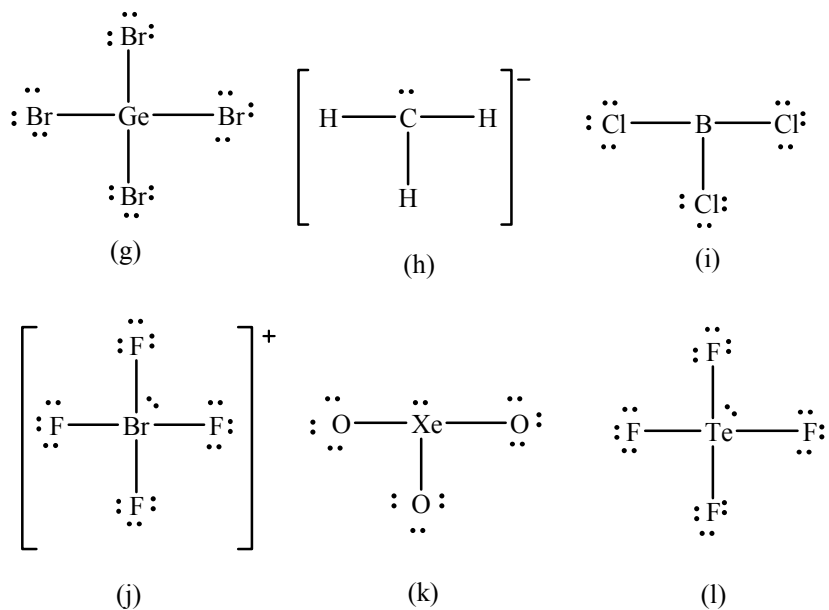


(d)

(e)

(f)

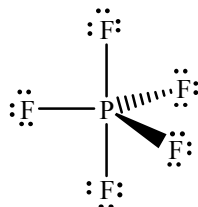




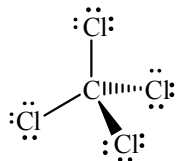
10.63 Plan: Apply VSEPR to the structures shown in problem 10.62.

Solution:

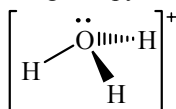
a)  $AX_5$  = trigonal bipyramidal



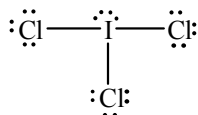
b)  $AX_4$  = tetrahedral



c)  $AX_3E$  = trigonal pyramidal



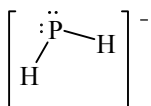
d)  $AX_3E_2$  = T-shaped



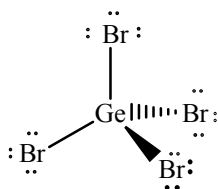
e)  $AX_2$  = linear



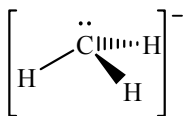
f)  $AX_2E_2$  = bent



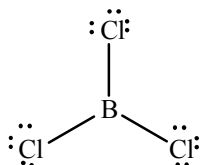
g)  $AX_4 = \text{tetrahedral}$



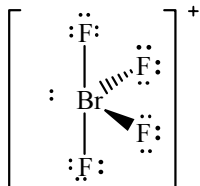
h)  $AX_3E = \text{trigonal pyramidal}$



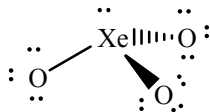
i)  $AX_3 = \text{trigonal planar}$



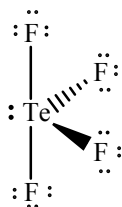
j)  $AX_4E = \text{see-saw}$



k)  $AX_3E = \text{trigonal pyramidal}$



l)  $AX_4E = \text{see-saw}$



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

Solution:

a) Formal charges for  $Al_2Cl_6$ :

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

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

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

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

Formal charges for  $I_2Cl_6$ :

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

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

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

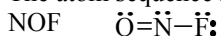
(Check: Formal charges add to zero, the charge on the compound.)

b) The aluminum atoms are  $AX_4$ , so they are tetrahedral. The 2 tetrahedral Al's cannot give a planar structure. The iodine atoms are  $AX_4E_2$  are square planar. Placing the square planar I's adjacent can give a planar molecule.

10.65 Plan: Draw Lewis structures, and then determine the formal charges.

Solution:

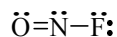
The atom sequence may be ONF, NOF, or NFO. F is never central so the structure cannot be NFO.



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

$$\text{FC}_{\text{N}} = +5 - (2 + 1/2(6)) = 0$$

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



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

$$\text{FC}_{\text{N}} = +5 - (4 + 1/2(4)) = -1$$

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

Thus, the structure on the left is the correct structure.

10.66 Plan: Draw Lewis structures, and then apply VSEPR.

Solution:

The Lewis structure for each is required.

Compound      Pre-VSEPR      Lewis Structure

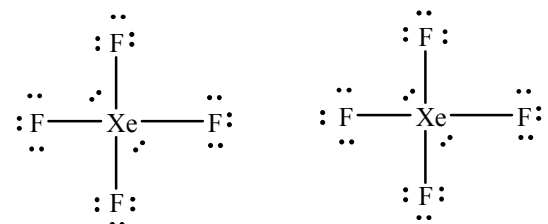
XeF<sub>2</sub>              Linear



VSEPR

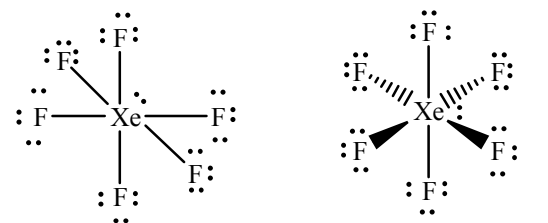
Linear (AX<sub>2</sub>E<sub>3</sub>)

XeF<sub>4</sub>              Tetrahedral



Square planar (AX<sub>4</sub>E<sub>2</sub>)

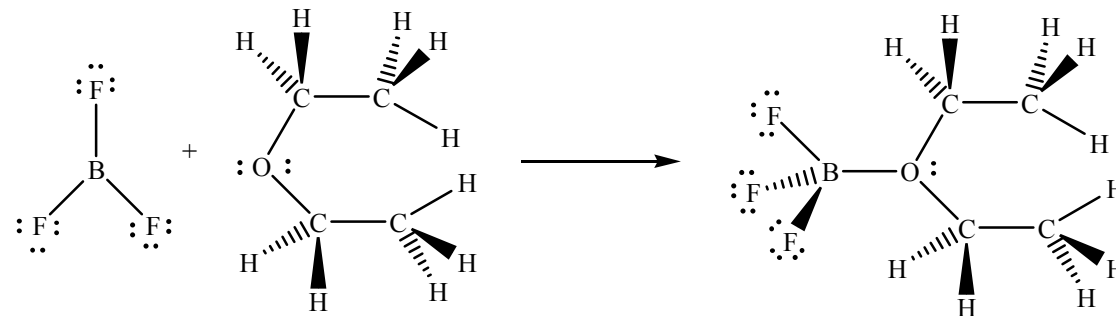
XeF<sub>6</sub>              Octahedral



Distorted octahedral (AX<sub>6</sub>E)

10.67 Plan: Draw Lewis structures, and then apply VSEPR.

Solution:

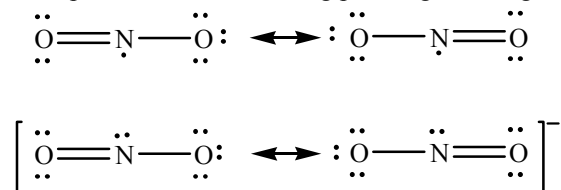


The BF<sub>3</sub> (AX<sub>3</sub>) is **trigonal planar** (bond angle = 120°). The O (AX<sub>2</sub>E<sub>2</sub>) is **bent** in the ether. In the product B is AX<sub>4</sub>, thus, it is **tetrahedral**. The O in the product is now AX<sub>3</sub>E and so it is **trigonal pyramidal**.

10.68 Plan: Draw Lewis structures, and then apply VSEPR.

Solution:

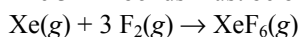
From the Lewis structures, both are AX<sub>2</sub>E, but the “pair” in NO<sub>2</sub> is only half a pair, so it only exerts “half” the repulsion. This allows the bond angle to open to a larger than normal bond angle. The “complete” lone pair in NO<sub>2</sub><sup>-</sup>, like other lone pairs, forces the bonding pairs together to give a smaller than normal bond angle.



10.69 Plan: Use the heat of formation given in the problem and the values in the bond energy table.

Solution:

The 3 F-F bonds must be broken, and 6 Xe-F bonds are formed.



$$\Delta H_f = 3 \text{BE}_{\text{F-F}} - 6 \text{BE}_{\text{Xe-F}}$$

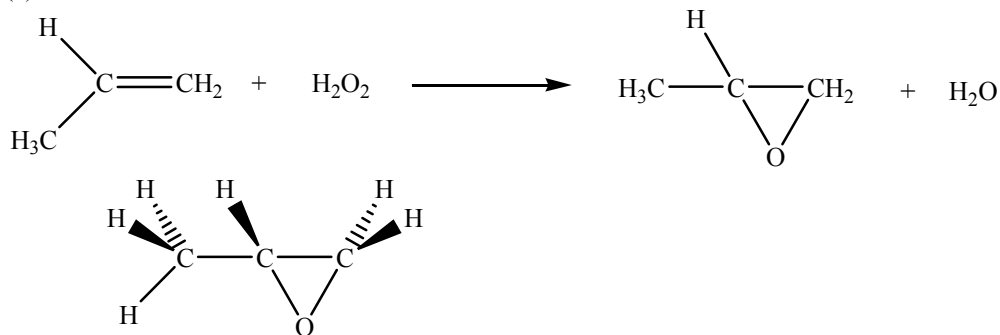
$$-402 \text{ kJ/mol} = (3)(159 \text{ kJ/mol}) - 6 \text{BE}_{\text{Xe-F}}$$

$$\text{BE}_{\text{Xe-F}} = 146.5 = \mathbf{146 \text{ kJ/mol}}$$

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

Solution:

(a)

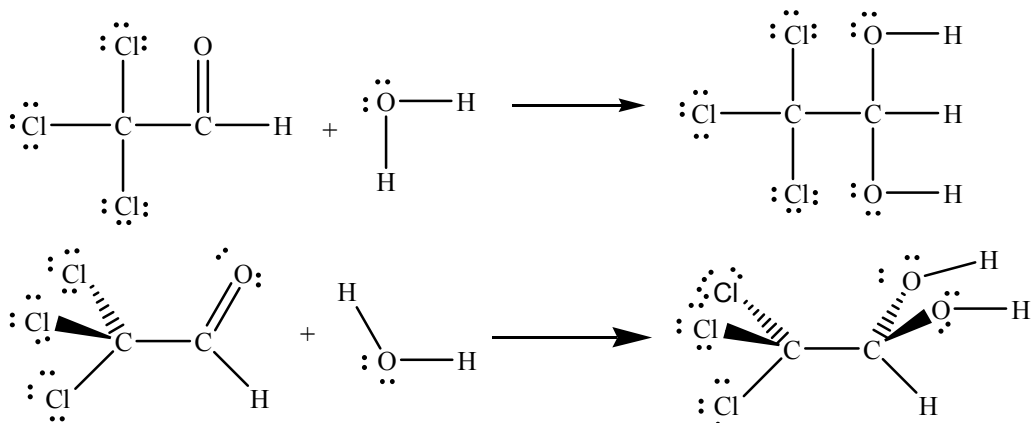


a) In propylene oxide, the C's are all AX<sub>4</sub>. All the ideal bond angles for the C's in propylene oxide are 109.5°.

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°. The angles around the 2 carbons in the rings are reduced from the ideal 109.5° to 60°.

10.71 Plan: Draw the Lewis structures, and then use VSEPR to predict the shapes.

Solution:

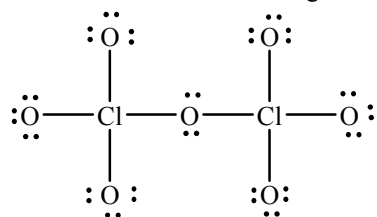


The C with the chlorines attached does not change shape. The other C changes from trigonal planar (AX<sub>3</sub>) to tetrahedral (AX<sub>4</sub>).

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

Solution:

The molecule has 56 electrons. Putting the molecule together as described in the problem gives:



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

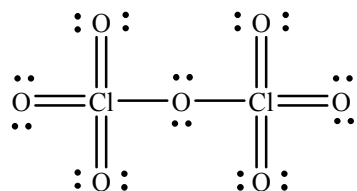
$$FC_{Cl} = 7 - (0 + 1/2(8)) = +3$$

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

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

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

The formal charges in the Cl's 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:



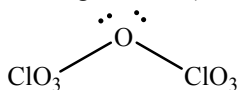
The formal charge need to be recalculated:

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

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

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

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

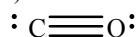


The central O is AX<sub>2</sub>E<sub>2</sub>, which has an ideal bond angle of 109.5°, but the lone pairs make this less than ideal.

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

Solution:

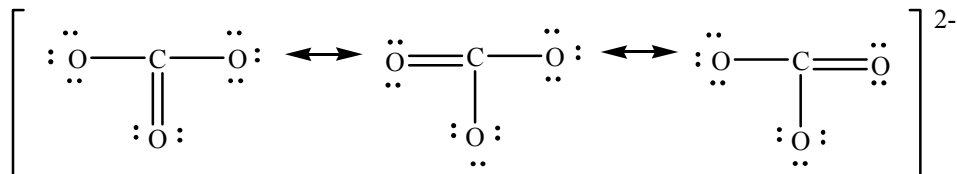
a)



Bond Order (avg.)

3.0

b)

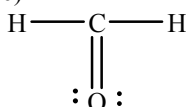


Each C-O bond is a single bond two-thirds of the time and a double bond the rest of the time.

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

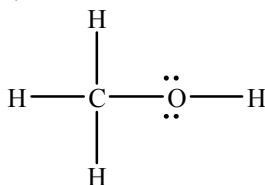
1.33

c)



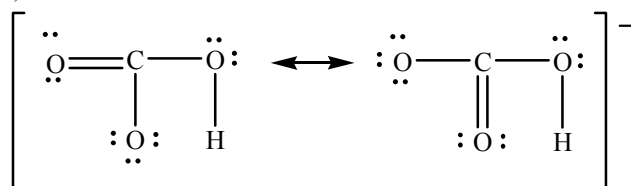
2.0

d)



1.0

e)



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

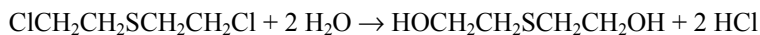
Bond strength **d < b < e < c < a** ignoring O attached to H

Bond length **a < c < e < b < d**

10.74 Plan: Write and balance the equation. Use the bond energy table to obtain the appropriate bond energies.

Solution:

The reaction is balanced as usual:



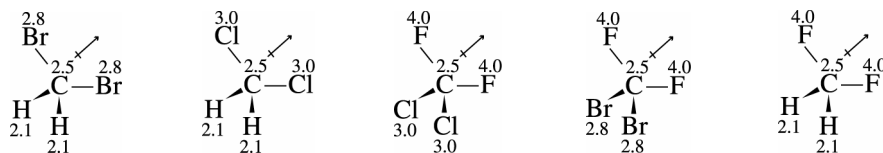
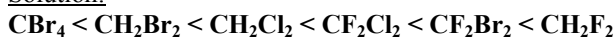
Most of the molecule remains the same.

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)
	2 x 2 O-H	4(467 kJ/mol)		2 O-H	2(467 kJ/mol)
	2 C-O	2(358 kJ/mol)			
	Totals:	7062 kJ			7020 kJ

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

10.75 Plan: Draw Lewis structures, and then apply VSEPR.

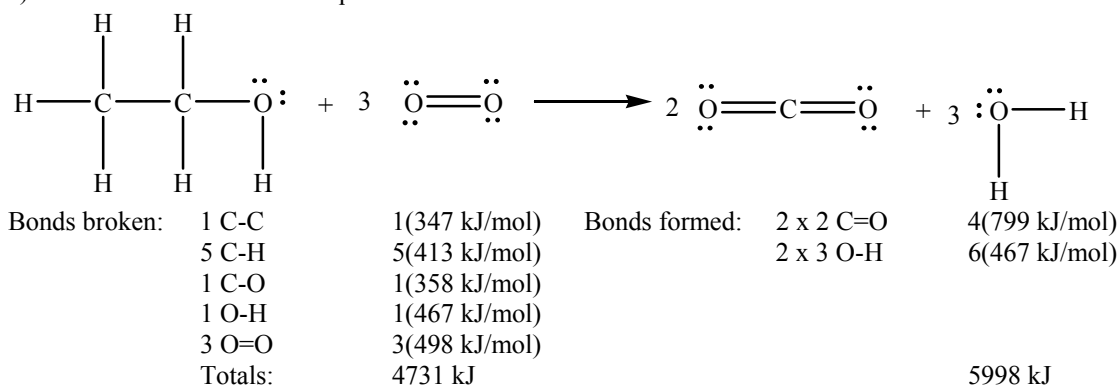
Solution:



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  $\text{H}_2\text{O}$ .

Solution:

a) To save time the balanced equation includes the Lewis structures:



$$\Delta H = \text{Bonds broken} - \text{Bonds formed}$$

$$= 4731 \text{ kJ} - 5998 \text{ kJ} = \mathbf{-1267 \text{ kJ}}$$
 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:

$$\Delta H_{\text{comb(liquid)}} = -1267 \text{ kJ} + (1 \text{ mol})(40.5 \text{ kJ/mol}) = -1226.5 = \mathbf{-1226 \text{ kJ per mol}}$$
 of liquid ethanol burned

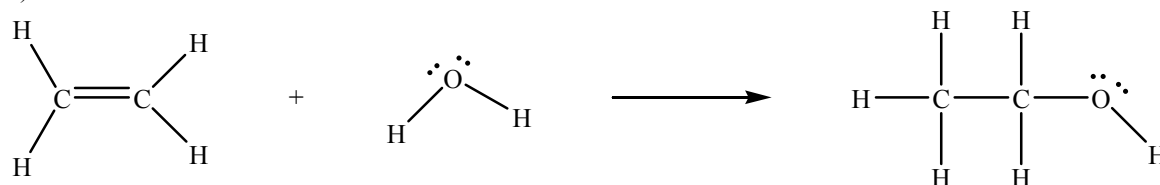
$$\text{c) } \Delta H_{\text{comb(liquid)}} = \sum n\Delta H_f^\circ (\text{products}) - \sum m\Delta H_f^\circ (\text{reactants})$$

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

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

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.

d) The Lewis structures for the reaction are:



$$\begin{aligned} \text{The energy to break the reactant bonds is: } & \text{BE}(\text{C}=\text{C}) + 4 \text{ BE}(\text{C}-\text{H}) + 2 \text{ BE}(\text{O}-\text{H}) \\ & = (1 \text{ mol})(614 \text{ kJ/mol}) + (4 \text{ mol})(413 \text{ kJ/mol}) + (2 \text{ mol})(467 \text{ kJ/mol}) \\ & = 3200. \text{ kJ} \end{aligned}$$

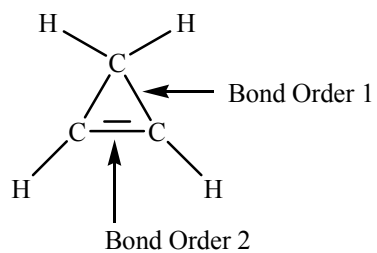
$$\begin{aligned} \text{The energy from forming the product bonds is: } & \text{BE}(\text{C}-\text{C}) + 5 \text{ BE}(\text{C}-\text{H}) + \text{BE}(\text{C}-\text{O}) + \text{BE}(\text{O}-\text{H}) \\ & = (1 \text{ mol})(347 \text{ kJ/mol}) + (5 \text{ mol})(413 \text{ kJ/mol}) + (1 \text{ mol})(358 \text{ kJ/mol}) + (1 \text{ mol})(467 \text{ kJ/mol}) \\ & = 3237 \text{ kJ} \end{aligned}$$

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

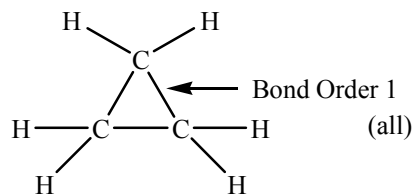
10.77 Plan: Draw the appropriate Lewis structures and observe the bonds.

Solution:

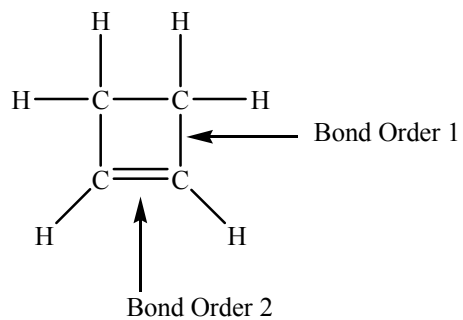
a)



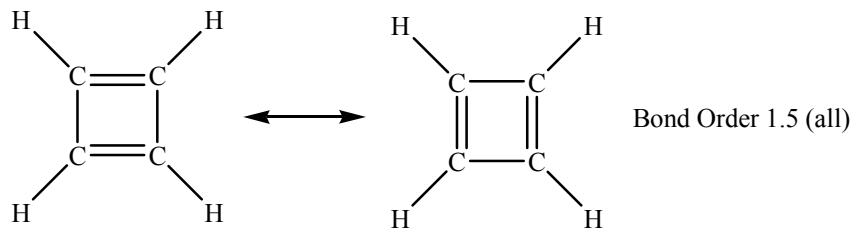
b)



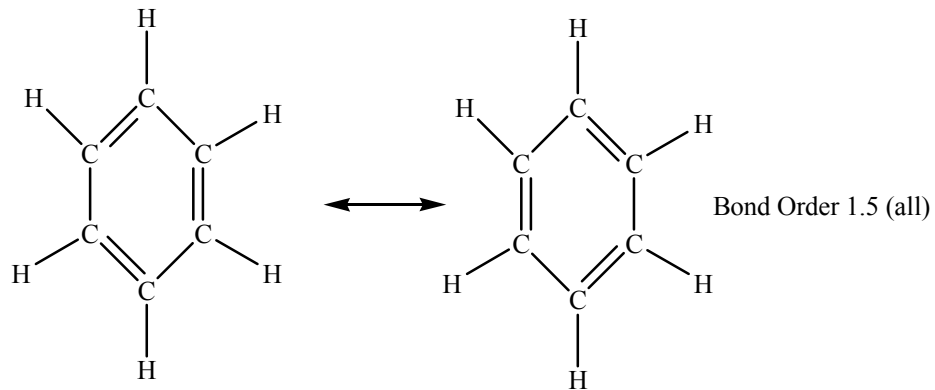
c)



d)



e)





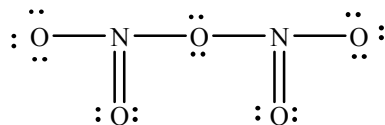
- 10.78 Plan: Determine the empirical formula from the percent composition (assuming 100 grams of compound). The empirical formula and the molar mass may then be used to determine the molecular formula. Count the valence electrons in the empirical formula and then construct the Lewis structure. Name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

Solution:

$$\begin{array}{ll} \text{N} & 25.9 \text{ g} / (14.01 \text{ g/mol}) = 1.849 \text{ mol} / 1.849 \text{ mol} = 1.00 \quad \text{Only round at the end.} \\ \text{O} & (100.0 - 25.9) \text{ g} / (16.00 \text{ g/mol}) = 4.631 \text{ mol} / 1.849 \text{ mol} = 2.50 \end{array}$$

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:

$$\begin{array}{ll} \text{H} & 2.24 \text{ g} / (1.008 \text{ g/mol}) = 2.222 \text{ mol} / 2.222 \text{ mol} = 1.00 \quad \text{Only round at the end.} \\ \text{C} & 26.7 \text{ g} / (12.01 \text{ g/mol}) = 2.223 \text{ mol} / 2.222 \text{ mol} = 1.00 \\ \text{O} & 71.1 \text{ g} / (16.00 \text{ g/mol}) = 4.443 \text{ mol} / 2.222 \text{ mol} = 2.00 \end{array}$$

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

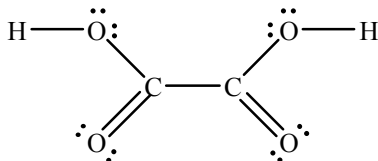
The titration required:

$$(0.040 \text{ mol NaOH} / \text{L}) (50.0 \text{ mL}) (1 \text{ L} / 1000 \text{ mL}) (1 \text{ mmole} / 0.001 \text{ mol}) = 2.0 \text{ mmole NaOH}$$

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>.

There are 34 valence electrons to be used in the Lewis structure. A C-C bond attaches the two empirical formula units:



- 10.80 Plan: 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. Try to draw other empirical formulas.

Solution:

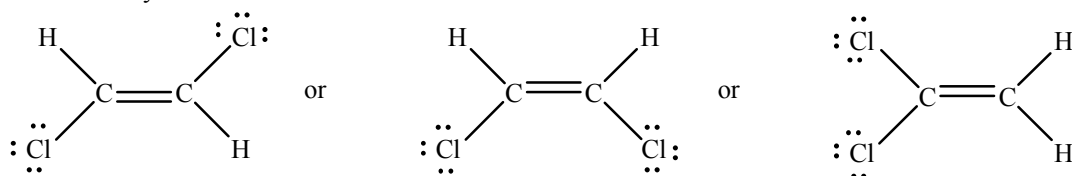
$$\begin{array}{ll} \text{C} & 24.8 \text{ g} / (12.01 \text{ g/mol}) = 2.065 \text{ mol} / 2.006 \text{ mol} = 1.00 \quad \text{Only round at the end.} \\ \text{H} & 2.08 \text{ g} / (1.008 \text{ g/mol}) = 2.063 \text{ mol} / 2.006 \text{ mol} = 1.00 \\ \text{Cl} & 73.1 \text{ g} / (35.45 \text{ g/mol}) = 2.006 \text{ mol} / 2.006 \text{ mol} = 1.00 \end{array}$$

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

The density is 4.3 g/L at STP:

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

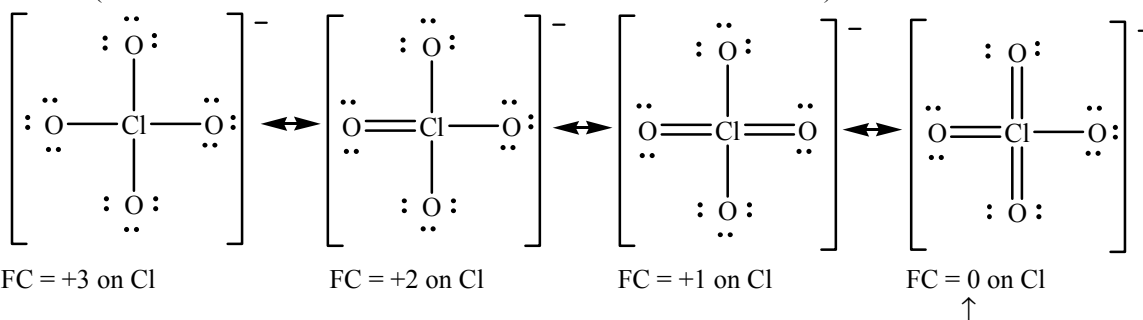
The molar mass (96 g/mol) is double the empirical formula mass (48.47 g/mol), so the empirical formula must be doubled to get the molecular formula:  $C_2H_2Cl_2$ . The formula contains 24 valence electrons. A variety of structures may be drawn:



10.81 Plan: Count the valence electrons and begin four Lewis structures. Place bonds as described in the problem.

Solution:

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.)



FC = +3 on Cl

FC = +2 on Cl

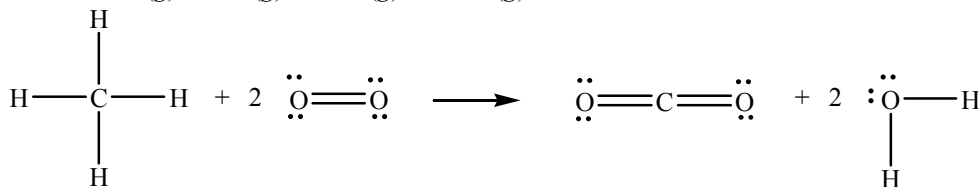
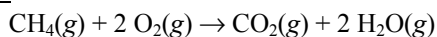
FC = +1 on Cl

FC = 0 on Cl

$$\text{Average Bond order} = \frac{(3 \times 2) + (1 \times 1)}{4} = 1.75 \quad (\text{most important})$$

10.82 Plan: Write the balanced chemical equations for the reactions. Draw the Lewis structures and look for potential problems. 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.

Solution:

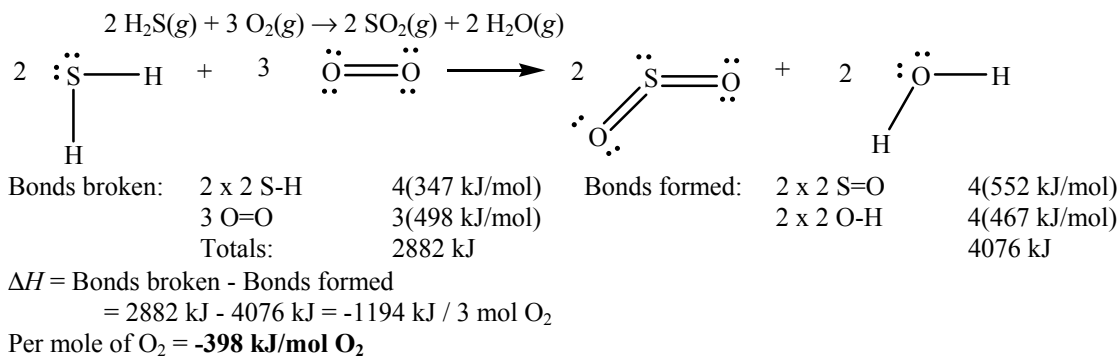


Bonds broken:	4 C-H	4(413 kJ/mol)	Bonds formed:	2 C=O	2(799 kJ/mol)
	O=O	2(498 kJ/mol)		2 x 2 O-H	4(467 kJ/mol)
	Totals:	2648 kJ			3466 kJ

$$\Delta H = \text{Bonds broken} - \text{Bonds formed}$$

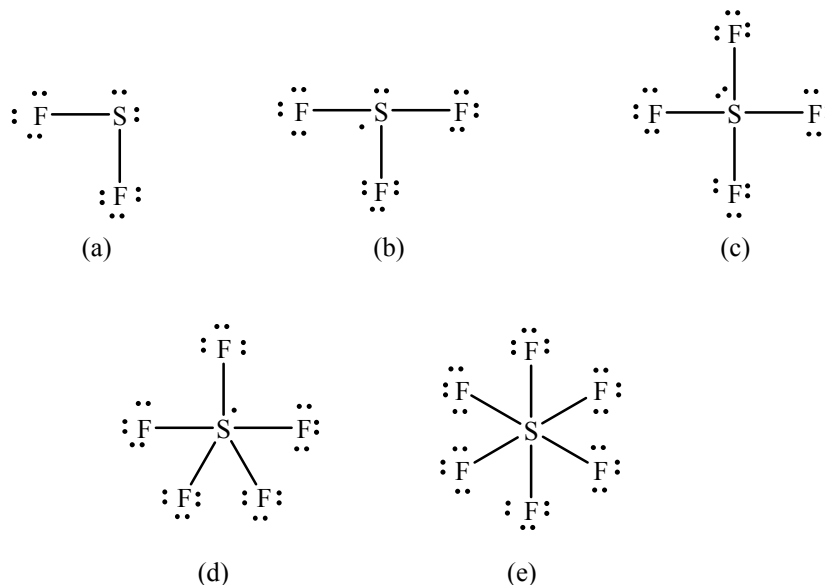
$$= 2648 \text{ kJ} - 3466 \text{ kJ} = -818 \text{ kJ} / 2 \text{ mol } O_2$$

Per mole of  $O_2 = -409 \text{ kJ/mol } O_2$



10.83 Plan: Draw the Lewis structures and look for potential problems.

Solution:



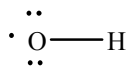
Stable: a, c, and e

Unstable radicals: b and d

10.84 Plan: Draw the Lewis structure of the OH species. Find the additional information from the bond energy table.

Solution:

a) The OH molecule has 7 valence electrons. Thus, 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:

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

$$[\text{BE}_{\text{O}=\text{O}} + \text{BE}_{\text{H}-\text{H}}] - [\text{BE}_{\text{O}-\text{H}}] = 39.0 \text{ kJ}$$

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

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

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

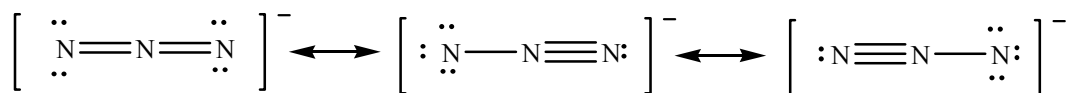
$$934 \text{ kJ} - 426 \text{ kJ} = \mathbf{508 \text{ kJ}}$$

10.85 Plan: Determine the total valence electrons and draw resonance forms for the two species.

Solution:

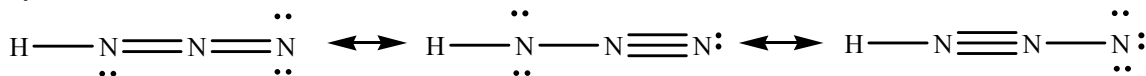
Both  $\text{N}_3^-$  and  $\text{HN}_3$  have 16 valence electrons.

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 N on the other end is 0. The first resonance structure is more important, the structure should have 2 equal bonds with a bond order of 2.

Hydrazoic acid:



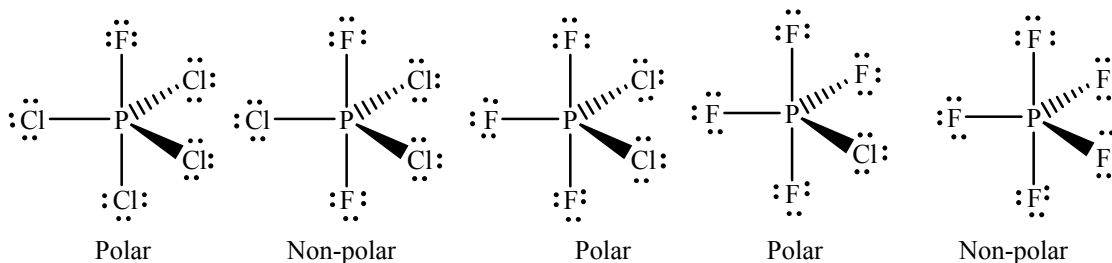
$\text{HN}_3$  also has 3 resonance structures. The formal change for the H is 0 in all the structures. In the structure with 2 double bonds, the formal charges for the N's are, left to right: 0, +1, and -1. The structure where the H is attached to the single bonded N, has N's with the following formal charges: -1, +1, and 0. In the final Lewis structure, the formal charges on the N's 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.

Solution:

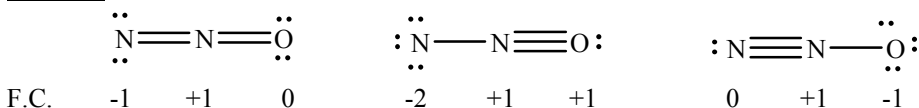
a) The F atoms will occupy the smaller axial positions first. The molecules containing only F or only Cl are nonpolar, as all the polar bonds would cancel. The molecules with one F or one Cl would be polar since there would be no corresponding bond to cancel the polarity. The presence of 2 axial F's means that their polarities will cancel (as would the 3 Cl's) giving a nonpolar molecule. The molecule with 3 F's is also polar.

b)



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

Solution:



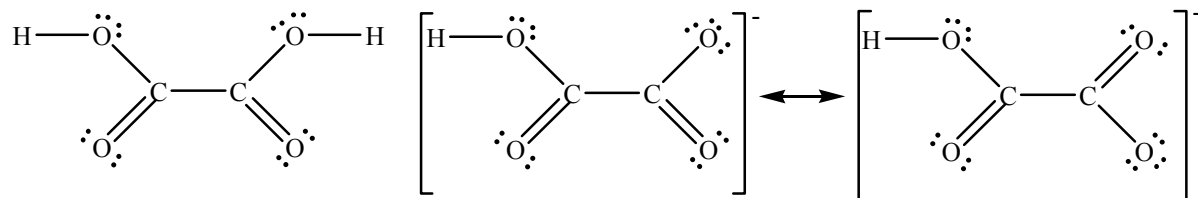
The third structure has a more reasonable distribution of formal charges. The third form has a strong bond between the N's and a weak N-O bond. It is easy to break the N-O bond which is why this compound easily decomposes to support combustion.

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

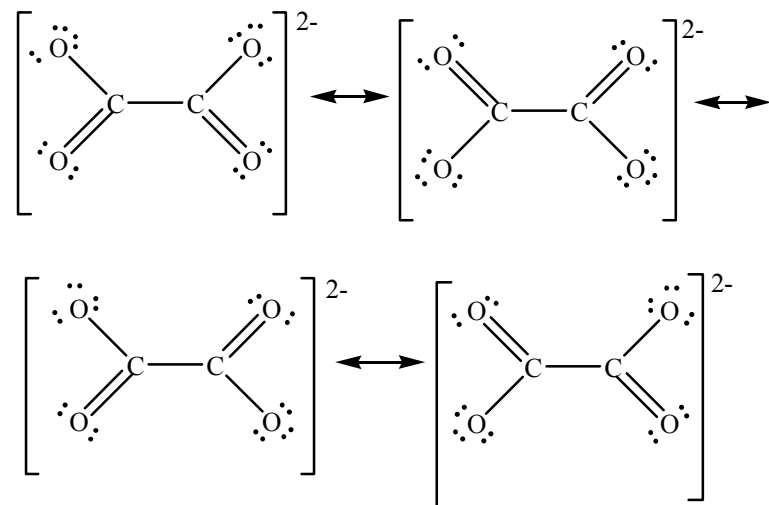
Solution:

$\text{H}_2\text{C}_2\text{O}_4$ :

$\text{HC}_2\text{O}_4^-$ :



$\text{C}_2\text{O}_4^{2-}$ :



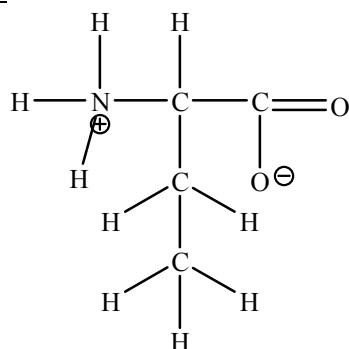
In  $\text{H}_2\text{C}_2\text{O}_4$ , there are 2 shorter  $\text{C}=\text{O}$  bonds and 2 longer  $\text{C}-\text{O}$  bonds.

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 Plan: Count the valence electrons. Thirty electrons are accounted for in the structure shown in the problem.

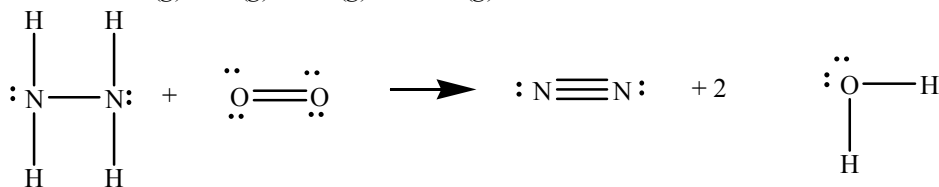
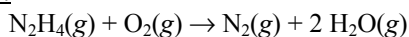
Solution:



All the atoms have 0 formal charge except the  $\text{N}$  ( $\text{FC} = +1$ ), and the single bonded  $\text{O}$  ( $\text{FC} = -1$ ).

10.90 Plan: Write a balanced equation. Find the appropriate bond energies in the table.

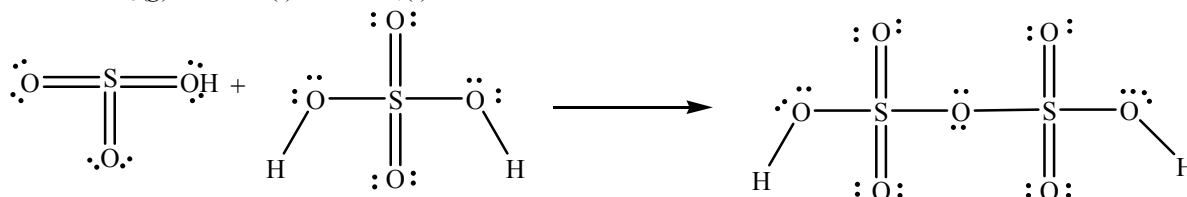
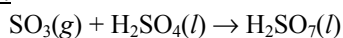
Solution:



$$\begin{aligned} \Delta H &= \Sigma \text{BE}_{\text{broken}} - \Sigma \text{BE}_{\text{formed}} \\ &= [4(391) + 160 + 498] - [945 + 2 \times 2(467)] \text{ kJ} \\ &= 2222 - 2813 \text{ kJ} = -591 \text{ kJ per mole of N}_2\text{H}_4 \text{ reacted} \end{aligned}$$

10.91 Plan: Draw the Lewis structures and look for potential problems. Calculate the heat of reaction from the bond energies.

Solution:

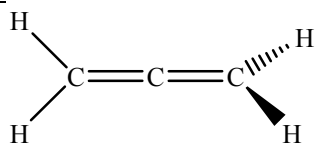


Bonds broken:	5 S=O	5(552 kJ/mol)	Bonds formed:	4 S=O	4(552 kJ/mol)
	2 S-O	2(265 kJ/mol)		4 S-O	4(265 kJ/mol)
	2 O-H	2(467 kJ/mol)		2 O-H	2(467 kJ/mol)
	Totals:	4224 kJ			4202 kJ

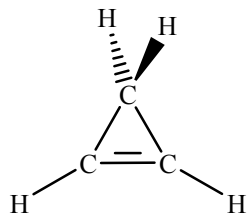
$$\begin{aligned} \Delta H &= \text{Bonds broken} - \text{Bonds formed} \\ &= 4224 \text{ kJ} - 4202 \text{ kJ} = 22 \text{ kJ} \end{aligned}$$

10.92 Plan: Draw the 2 Lewis structures, and then apply VSEPR to predict the angles.

Solution:



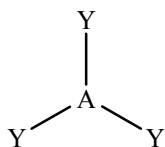
There are no deviations from the ideal angles. The central carbon is linear ( $180^\circ$ ). The end C's are trigonal planar ( $120^\circ$ ).



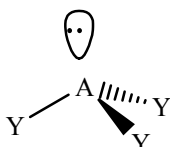
Ideally, the single bonded carbon should be tetrahedral ( $109.5^\circ$ ), and the double bonded carbons are trigonal planar ( $120^\circ$ ). The 3-membered ring will approximate an equilateral triangle with  $60^\circ$  angles. The external bonds are probably close to ideal, but the internal bonds are much less.

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

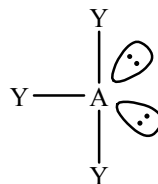
Solution:



(a)  
3 groups  
( $AX_3$ )



(b)  
4 groups  
( $AX_3E$ )

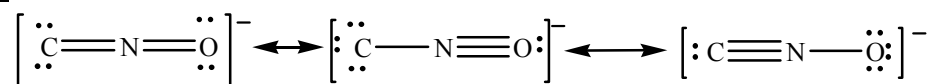


(c)  
5 groups  
( $AX_3E_2$ )

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 Plan: Draw the resonance structures for the fulminate ion, and determine the formal charges:

Solution:



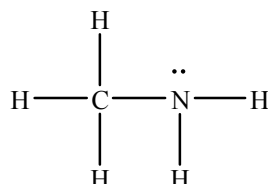
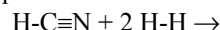
F.C.            -2    +1    0                    -3    +1    +1                    -1    +1    -1

None of the structures has a good distribution of formal charges, thus, none are stable. The best of the choices is the middle structure.

10.95 Plan: First, draw the appropriate Lewis structures, and then follow the procedure from the preceding problem.

Solution:

The simplified Lewis structures for the reaction are:



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

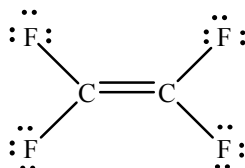
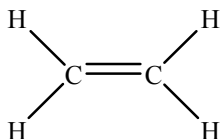
The energy from forming the product bonds is:  $3 BE(\text{C-H}) + BE(\text{C-N}) + 2 BE(\text{N-H})$   
 $= (3 \text{ mol}) (413 \text{ kJ/mol}) + (1 \text{ mol}) (305 \text{ kJ/mol}) + (2 \text{ mol}) (391 \text{ kJ/mol})$   
 $= 2326 \text{ kJ}$

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

10.96 Plan: Draw the appropriate Lewis structures, and use VSEPR to predict the angles.

Solution:

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  $\text{C}=\text{C}$  bond repels better than the single bonds. The larger F's cannot get as close together as the smaller H's, so the angles in tetrafluoroethylene are not reduced as much.

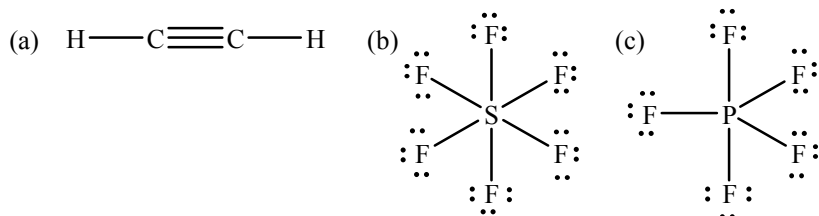
10.97 Plan: Compare the two structures and look for similarities.

Solution:

The tops of the molecules are similar and the tops will interact with biomolecules in a similar manner.

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

Solution:



a) The H's are separated by two carbon-hydrogen bonds (109 pm) and a carbon-carbon triple bond (121 pm).

Total separation = 2 (109 pm) + 121 pm = **339 pm**

b) The fluorines on opposite sides of the S are separated by twice the sulfur-fluorine bond length (158 pm).

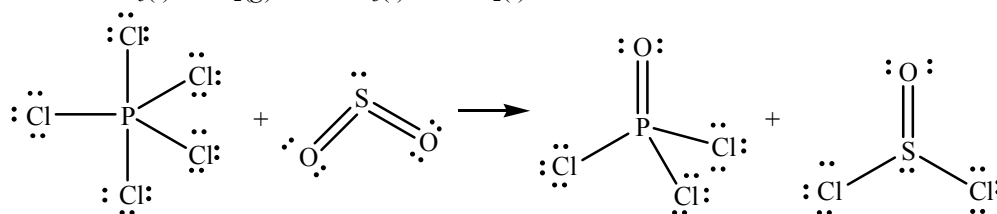
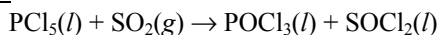
Total separation = 2 (158 pm) = **316 pm**

Adjacent fluorines are at two corners of a right triangle, with the sulfur at the 90° angle. Two sides of the triangle are equal to the sulfur-fluorine bond length (158 pm). The separation of the fluorines 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  $a = b = 158$  pm. Thus,  $c^2 = (158 \text{ pm})^2 + (158 \text{ pm})^2$ , and so  $c = 223.4457 = \mathbf{223 \text{ pm}}$

c) Thus, distance between equatorial fluorines may be determined in a similar manner as between adjacent fluorines in SF<sub>6</sub>. The central angle is 120°. The distance is = 273.66 = **274 pm**.

10.99 Plan: First, write a balanced chemical equation, then draw the appropriate Lewis structures. Finally, apply VSEPR to the Lewis structures.

Solution:



PCl<sub>5</sub>: AX<sub>5</sub> **trigonal bipyramid**

SO<sub>2</sub>: AX<sub>2</sub>E **bent**

POCl<sub>3</sub>: AX<sub>4</sub> **tetrahedral**

SOCl<sub>2</sub>: AX<sub>3</sub>E **trigonal pyramid**