

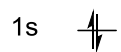
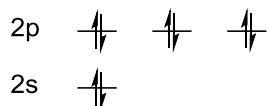
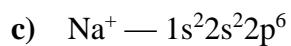
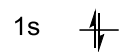
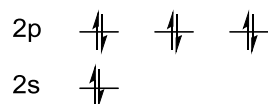
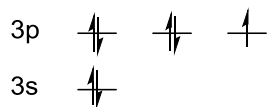
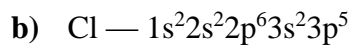
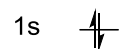
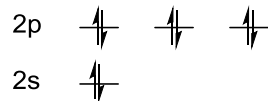
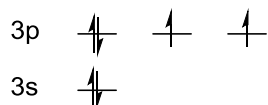
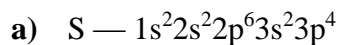
# Chapter 1

## Carbon and Its Compounds

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### CHECKPOINT PROBLEMS

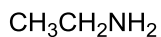
#### Practice Problem 1.1





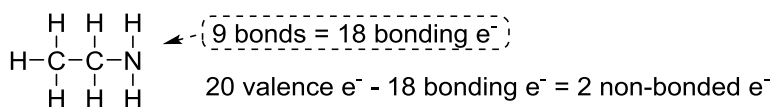
## Practice Problem 1.2

a) Count valence electrons.

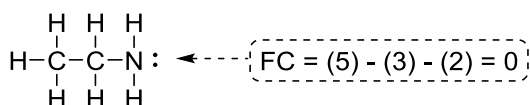


2 carbons (group 4)	$2 \times 4 =$	8
7 hydrogens (group 1)	$7 \times 1 =$	7
1 nitrogen (group 5)	$1 \times 5 =$	5
		<hr/> 20 valence electrons

Build a basic bonding framework and account for electrons used.

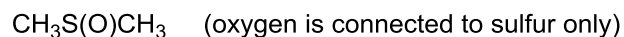


Add remaining electrons and check for formal charges.



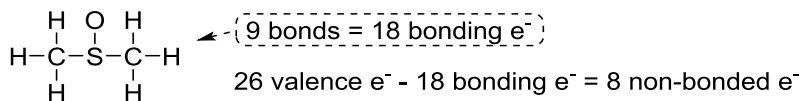
The molecule has a lone pair on the nitrogen. All other electrons are bonding electrons.

b) Count valence electrons.



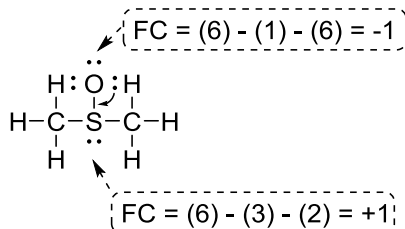
2 carbons (group 4)	$2 \times 4 =$	8
6 hydrogens (group 1)	$6 \times 1 =$	6
1 oxygen (group 6)	$1 \times 6 =$	6
1 sulfur (group 6)	$1 \times 6 =$	6
		<hr/> 26 valence $e^-$

Build a basic bonding framework and account for electrons used.

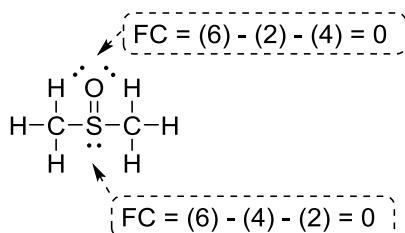




Add remaining electrons and check for formal charges.



There are formal charges on the S and O atoms. They can be removed by making an additional bond between O and S.



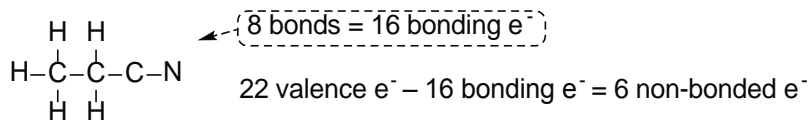
The molecule has a lone pair on the sulfur and two lone pairs on oxygen. All other electrons are bonding electrons.

c) Count valence electrons.

$\text{CH}_3\text{CH}_2\text{CN}$  (nitrogen is connected to one carbon only)

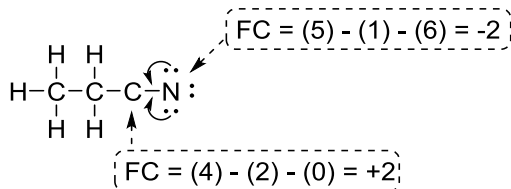
3 carbons (group 4)	$3 \times 4 = 12$
5 hydrogens (group 1)	$5 \times 1 = 5$
1 nitrogen (group 5)	$1 \times 5 = 5$
	<u>22 valence <math>e^-</math></u>

Build a basic bonding framework and account for electrons used.

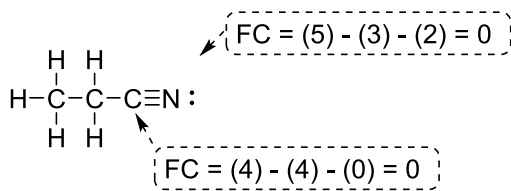




Add remaining electrons and check for formal charges.

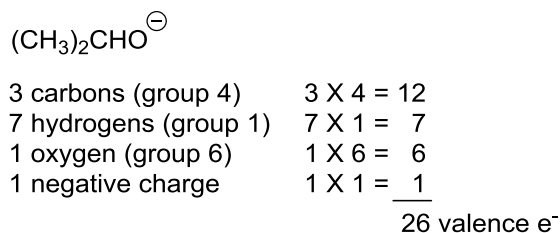


The formal charges on C and N show the carbon needs more electrons and the N has too many. Forming two more bonds between C and N alleviates this problem.

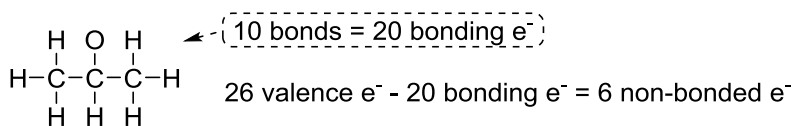


The molecule has a lone pair on the nitrogen. All other electrons are bonding electrons.

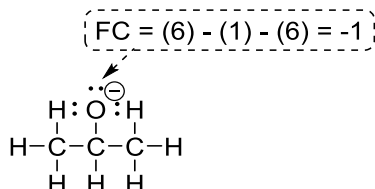
**d)** Count valence electrons.



Build a basic bonding framework and account for electrons used.



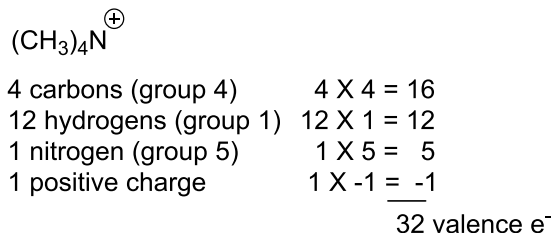
Add remaining electrons and check for formal charges.



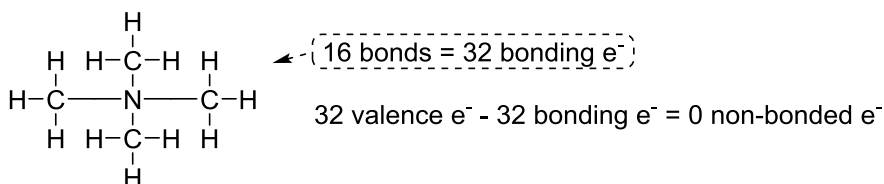
The oxygen atom has three lone pairs and a positive charge. All other electrons are bonding electrons.



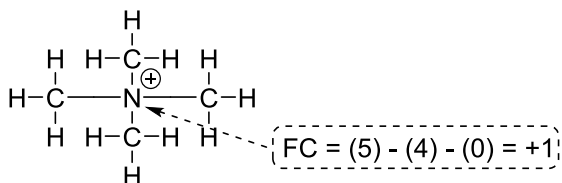
e) Count valence electrons.



Build a basic bonding framework and account for electrons used.

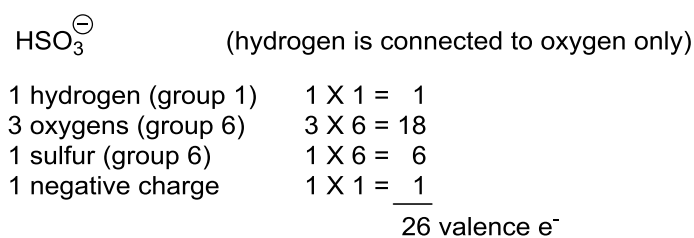


Add remaining electrons and check for formal charges.

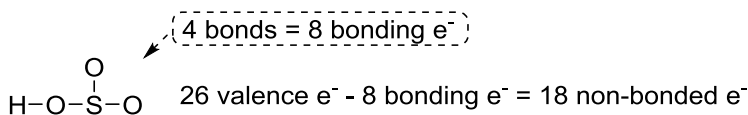


The ion has a formal positive charge on the nitrogen. All electrons are bonding electrons.

f) Count valence electrons.

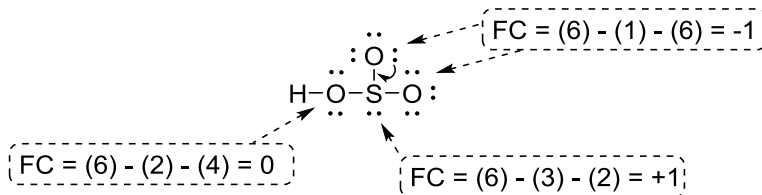


Build a basic bonding framework and account for electrons used.

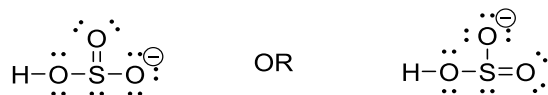




Add remaining electrons and check for formal charges.

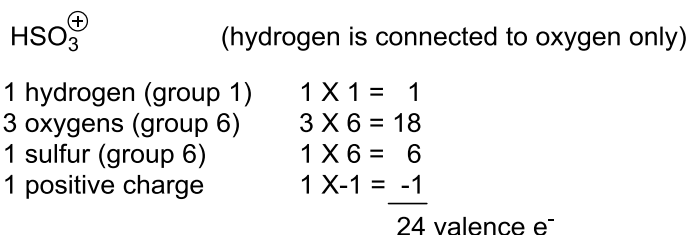


There are formal charges on the sulfur (+1) and two of the oxygens (-1). These can be reduced by forming a double bond between sulfur and either of the oxygen atoms carrying a formal charge.

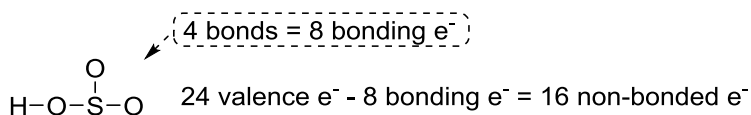


The ion has seven lone pairs on oxygen atoms and one lone pair on sulfur. All other electrons are bonding electrons. One oxygen has a formal negative charge.

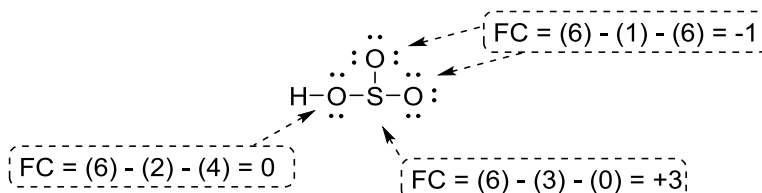
**g)** Count valence electrons.



Build a basic bonding framework and account for electrons used.

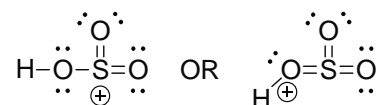


Add remaining electrons (on oxygen first) and check for formal charges.





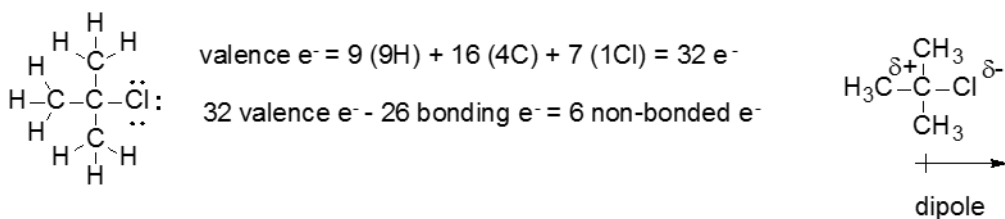
There are formal charges on the sulfur (+3) and two of the oxygens (-1). These can be reduced by forming double bonds between sulfur and both charged oxygen atoms. This expands the octet of the sulfur but, since it is a third-row element, this is allowed.



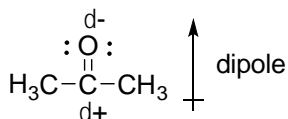
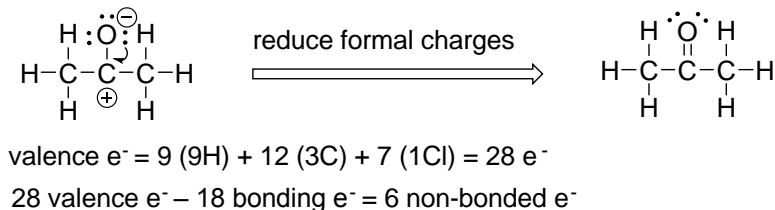
The final ion has six lone pairs on oxygen atoms. The sulfur has a formal positive charge. All other electrons are bonding electrons.

### Practice Problem 1.3

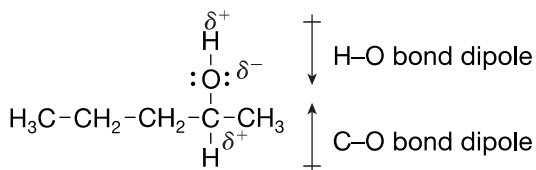
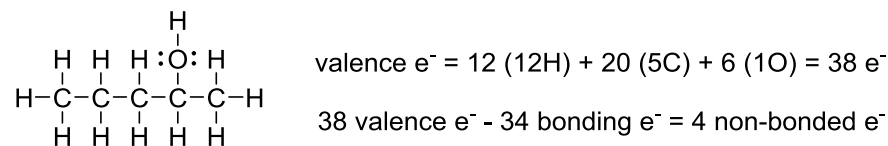
#### a) Lewis Structure



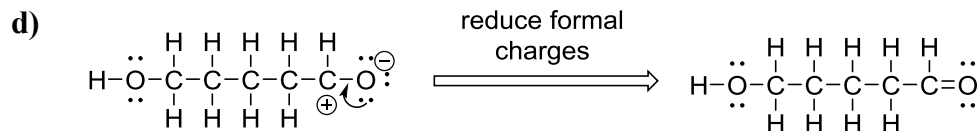
#### b) Lewis Structure



#### c)

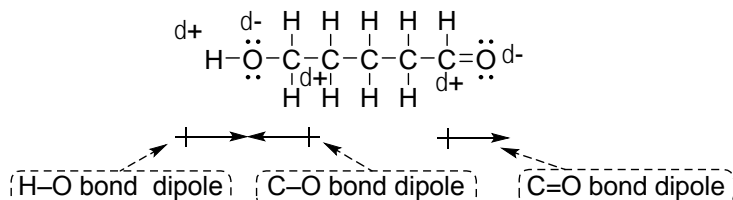






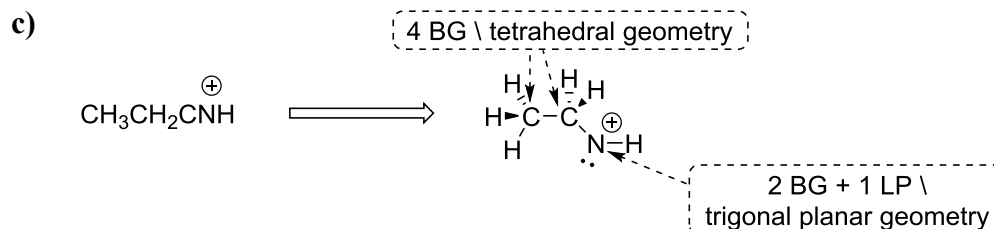
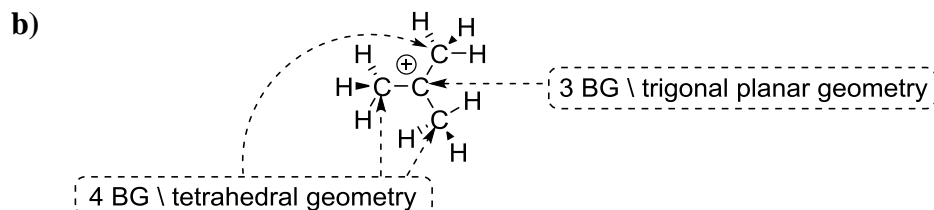
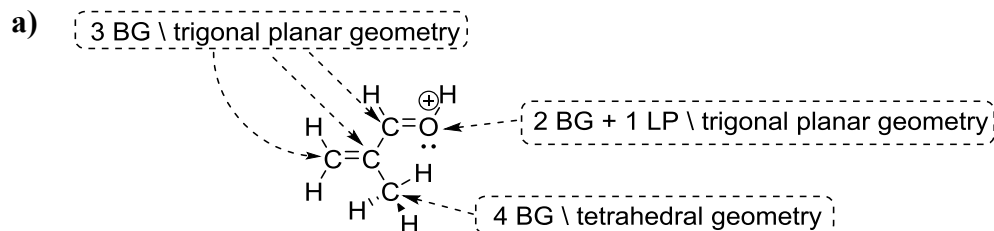
$$\text{valence } e^- = 10 (10H) + 20 (5C) + 12 (2O) = 42 e^-$$

$$42 \text{ valence } e^- - 32 \text{ bonding } e^- = 10 \text{ non-bonded } e^-$$



### Practice Problem 1.4

In the solutions, “BG” is used as an abbreviation for “bond group” and “LP” is used as an abbreviation for “lone pair.”



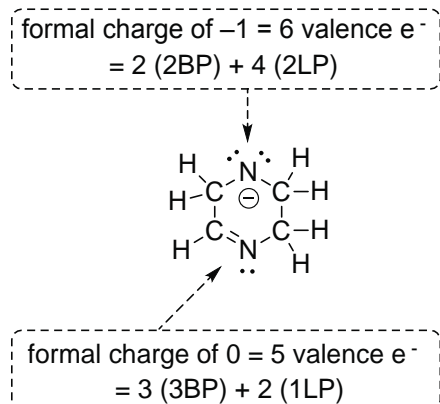
$$\text{valence } e^- = 6 (6H) + 8 (2C) + 5 (1N) - 1 (+ \text{ charge}) = 18 e^-$$

$$18 \text{ valence } e^- - 16 \text{ bonding } e^- = 2 \text{ non-bonded } e^-$$

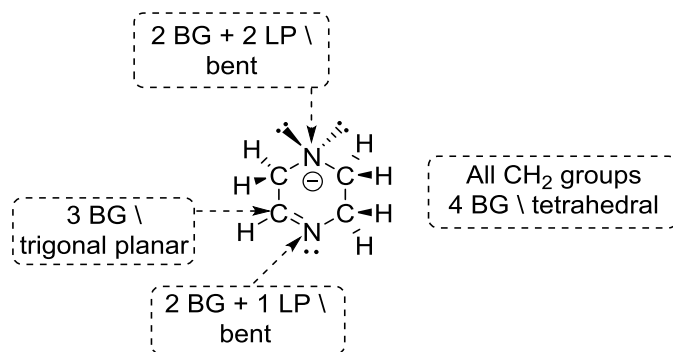


### Integrate the Skill 1.5

All carbons have four bonds and so will not have lone pairs. Lone pairs are added to the nitrogen atoms according to the formal charges indicated. “BP” refers to shared pairs of electrons in bonds between atoms where each atom formally has one of the electrons.

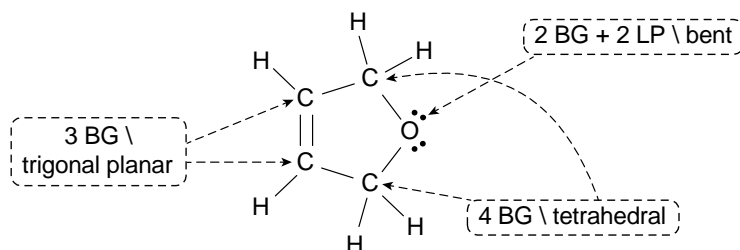


The geometry of the atoms can then be established for all of the atoms.

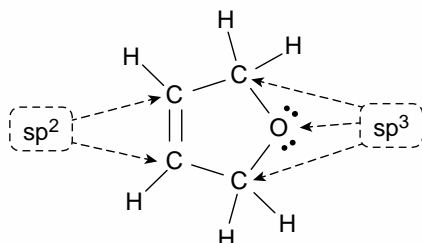


### Practice Problem 1.6

a)



b)



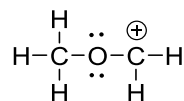


## Integrate the Skill 1.7

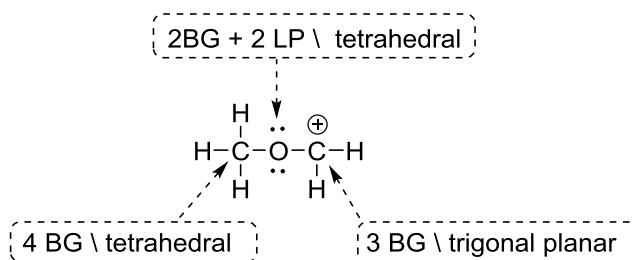
- a) Accounting for all of the electrons leaves four non-bonded electrons to add as lone pairs. These are added to the oxygen first (most electronegative), leaving the carbon with a formal positive charge.

$$\text{valence } e^- = 5 (5H) + 8 (2C) + 6 (1O) - 1 (+ \text{ charge}) = 18 e^-$$

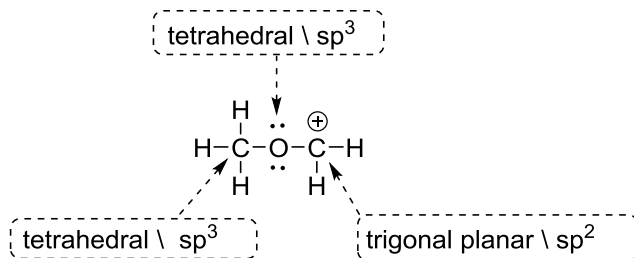
$$18 \text{ valence } e^- - 14 \text{ bonding } e^- = 4 \text{ non-bonded } e^-$$



- b) Electron geometry:



- c) Hybridization:

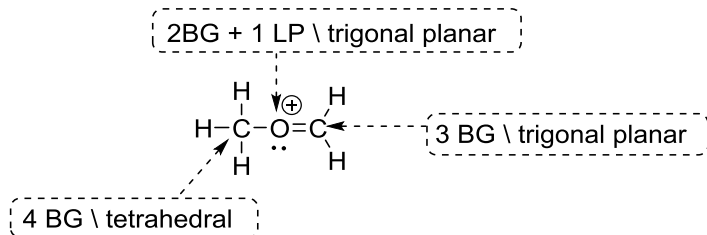


- d) The charged carbon is not saturated, so a second bond to the oxygen can be formed. This moves the formal charge to the oxygen atom.

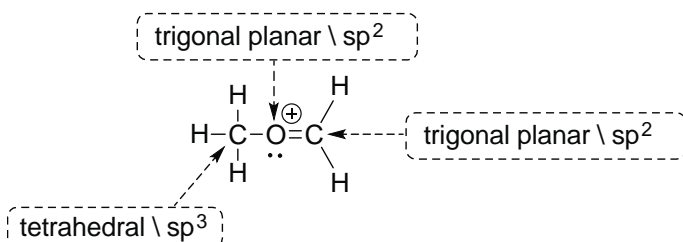




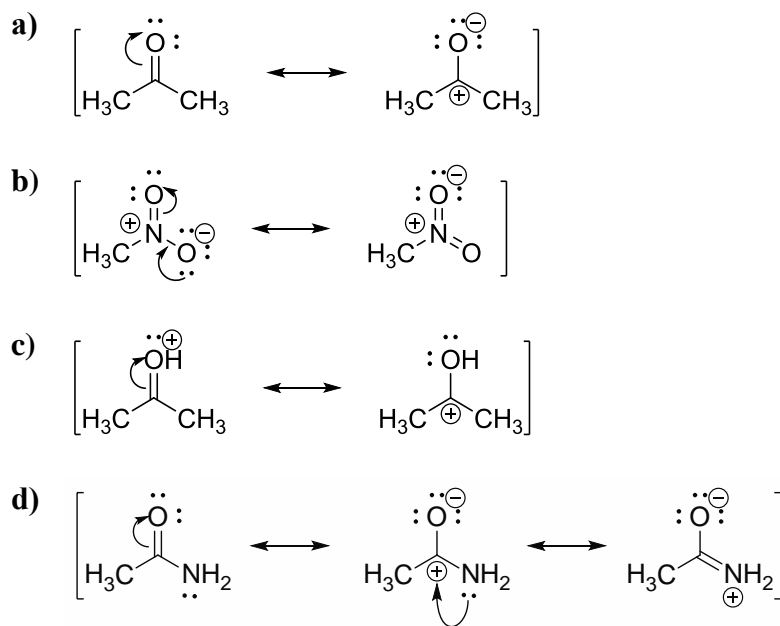
This leads to the following geometries for the new structure.



The corresponding hybridizations would then be



### Practice Problem 1.8



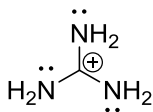


## Integrate the Skill 1.9

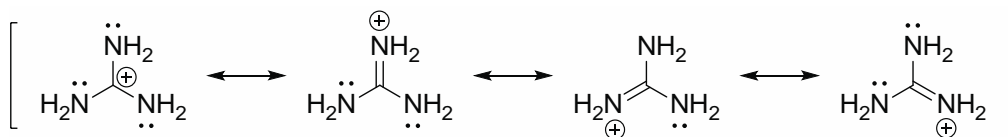
Drawing the basic structure leaves a formal positive charge on the carbon atom and lone pairs on each nitrogen atom.

$$\text{valence } e^- = 6 (6H) + 4 (1C) + 15 (3N) - 1 (+ \text{ charge}) = 24 e^-$$

$$24 \text{ valence } e^- - 18 \text{ bonding } e^- = 6 \text{ non-bonded } e^-$$

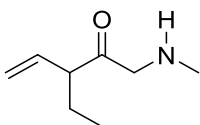


Three more resonance forms can be produced by forming a double bond between each of the nitrogen atoms and the central carbon atom. This leaves the formal charge on a nitrogen atom for each of these new forms.

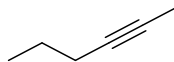


## Practice Problem 1.10

a) i)



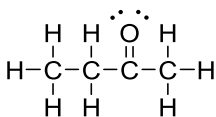
ii)



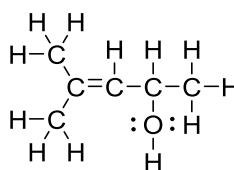
b) i)  $\text{CH}_2\text{CHCH}(\text{CH}_2\text{CH}_3)\text{COCH}_2\text{NHCH}_3$

ii)  $\text{CH}_3\text{CH}_2\text{CH}_2\text{CCCH}_3$

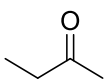
c) i)



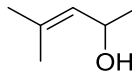
ii)



d) i)



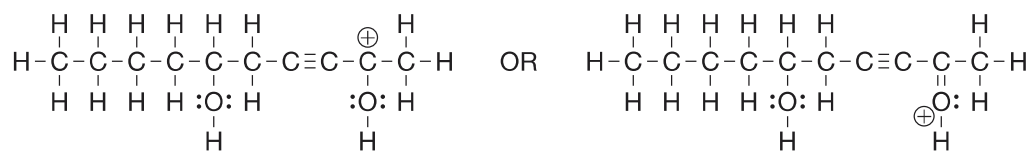
ii)



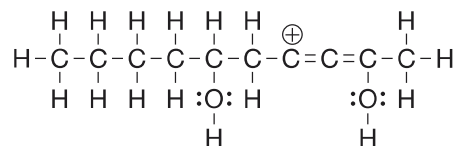


### Integrate the Skill 1.11

The formal charge can be on the carbon or the oxygen. Both are acceptable Lewis structures.



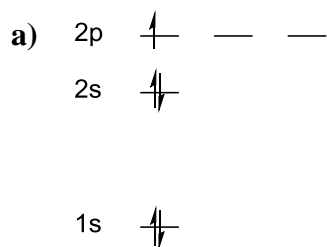
OR





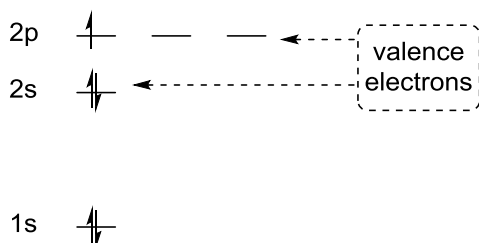
## PROBLEMS

### 1.12



b) The ground-state configuration is  $1s^2 2s^2 2p^1$ .

c) Boron has four valence orbitals.



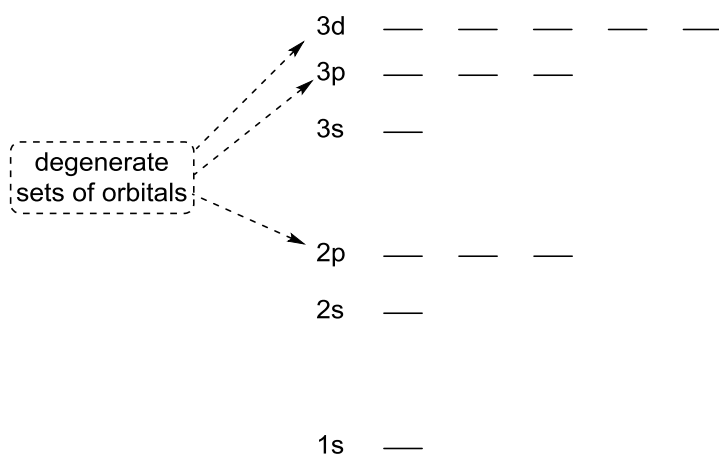
### 1.13

a) Hydrogen has one electron. In the ground state, it is contained in the 1s orbital. An electron in the 2s orbital would have to be an excited state of the hydrogen atom.

b) Carbon has six electrons. Its electron configuration is  $1s^2 2s^2 2p^2$ . Therefore, the valence electrons are in 2s and 2p orbitals. The 1s orbital is a core orbital.

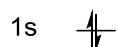
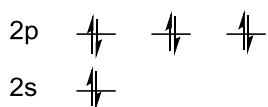
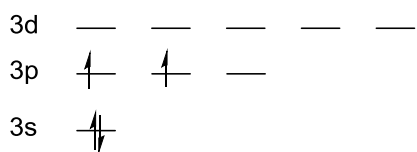
### 1.14

a) The degenerate sets of atomic orbitals are indicated below.





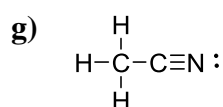
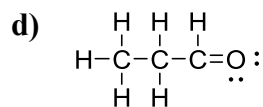
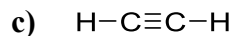
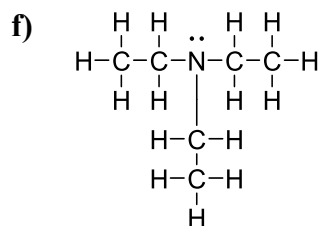
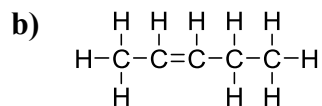
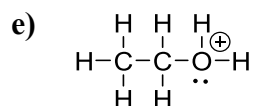
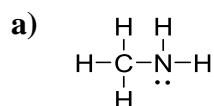
- b) The 3p orbitals are lower in energy than the 3d.
- c) The 3s orbital should be spherical, as are the 1s and 2s orbitals, but larger.
- d) The 3p orbitals should be similar in shape to the 2p orbitals, but larger.
- e) The ground-state electron configuration of silicon:



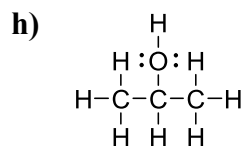
- f) The ground-state electron configuration of silicon is  $1s^2 2s^2 2p^6 3s^2 3p^2$ .
- g) Silicon has four valence electrons ( $3s^2 3p^2$ ), the same as carbon. The Lewis dot diagram of silicon would therefore be



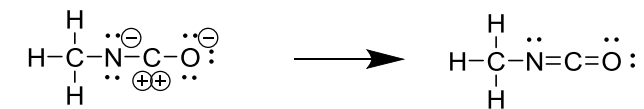
## 1.15



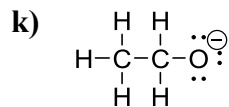
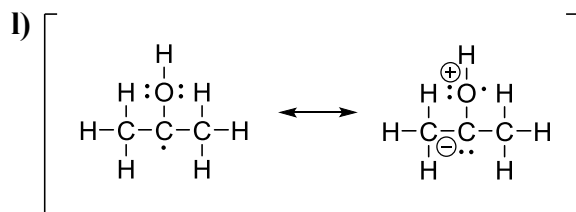
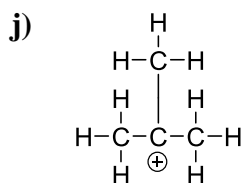




i) An intermediate structure is useful in determining the final structure.

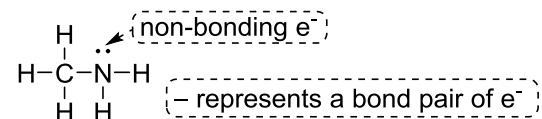


$$\begin{aligned}
 3(3\text{H}) + 8(2\text{C}) + 5(\text{N}) + 6(\text{O}) &= 22 \text{ e}^- \\
 22 \text{ e}^- - 12 \text{ bonding e}^- &= 10 \text{ LP e}^-
 \end{aligned}$$



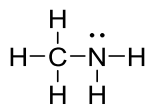
## 1.16

Bond pairs are shown as dashes (—), as indicated.

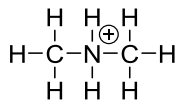


## 1.17

a) A molecule of formula C<sub>2</sub>H<sub>5</sub>N having no formal charge on any atom:

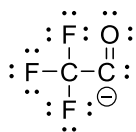


b) A cation of formula C<sub>2</sub>H<sub>8</sub>N<sup>+</sup>:





c) An anion of formula  $\text{C}_2\text{F}_3\text{O}^-$  having a  $\text{C}=\text{O}$  bond:

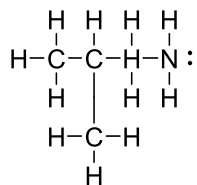


d) Two neutral molecules of formula  $\text{C}_2\text{H}_3\text{N}$ , both having a  $\text{C}\text{--}\text{N}$  triple bond:

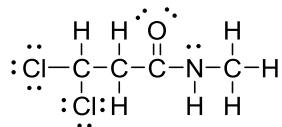


## 1.18

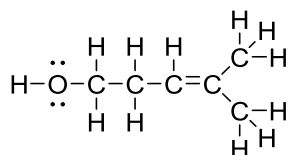
a)  $(\text{CH}_3)_2\text{CHCH}_2\text{NH}_2$



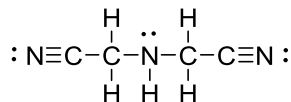
c)  $\text{Cl}_2\text{CHCH}_2\text{CONHCH}_3$



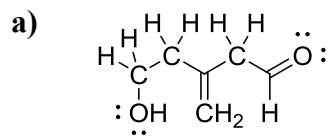
b)  $\text{HO}(\text{CH}_2)_2\text{CH}=\text{C}(\text{CH}_2\text{CH}_3)_2$



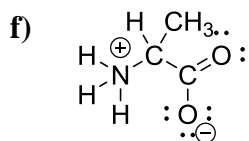
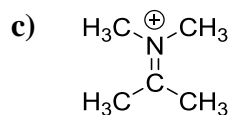
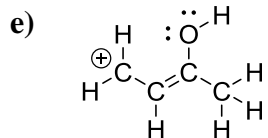
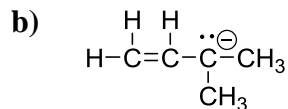
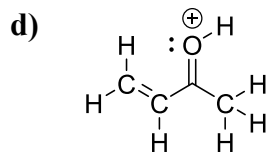
d)  $\text{NH}(\text{CH}_2\text{CN})_2$



## 1.19



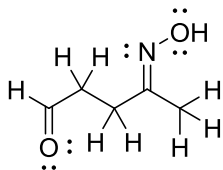
no formal charges



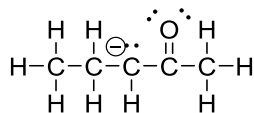


## 1.20

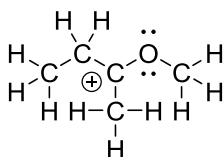
a)



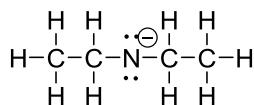
c)



b)



d)



## 1.21

Electronegativities are provided in Figure 1.6.

- a) The N-H bond would be more polar due to the larger electronegativity difference between N(3.0) and H(2.2) than between B(2.0) and H(2.2). The distinction between them is that the N-H bond is polarized toward the nitrogen, while the B-H bond is polarized toward the hydrogen.

b) i)



ii)



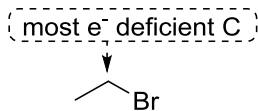
iii)



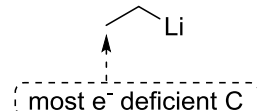
iv)



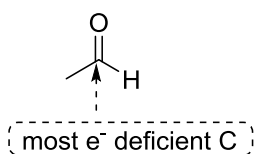
c) i)



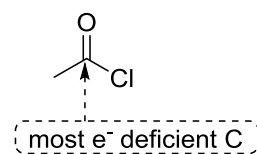
iii)



ii)



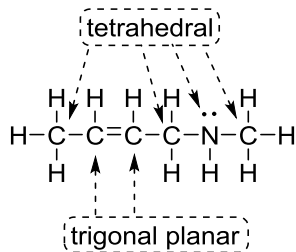
iv)



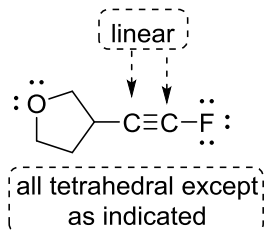


## 1.22

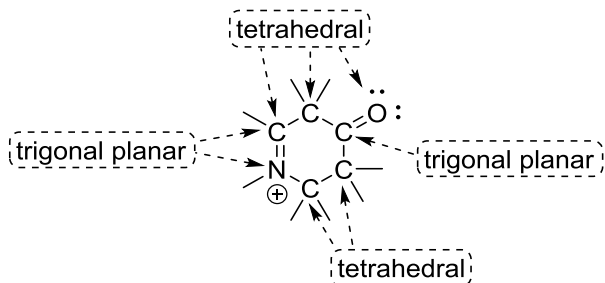
a)



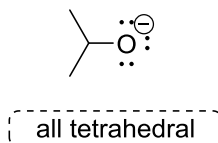
c)



b)

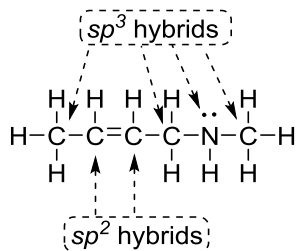


d)

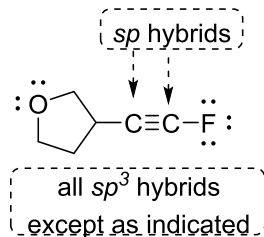


## 1.23

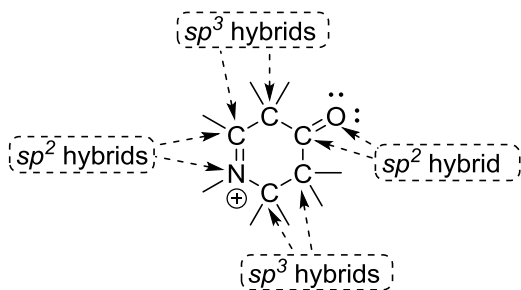
a)



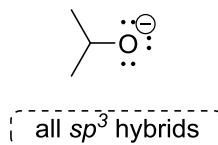
c)



b)

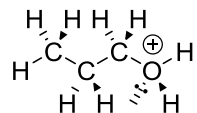


d)

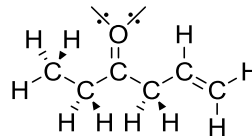


## 1.24

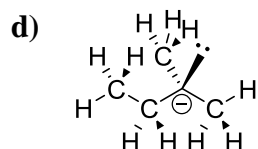
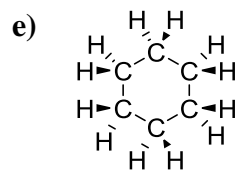
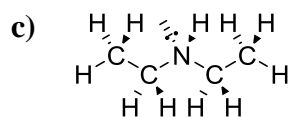
a)



b)

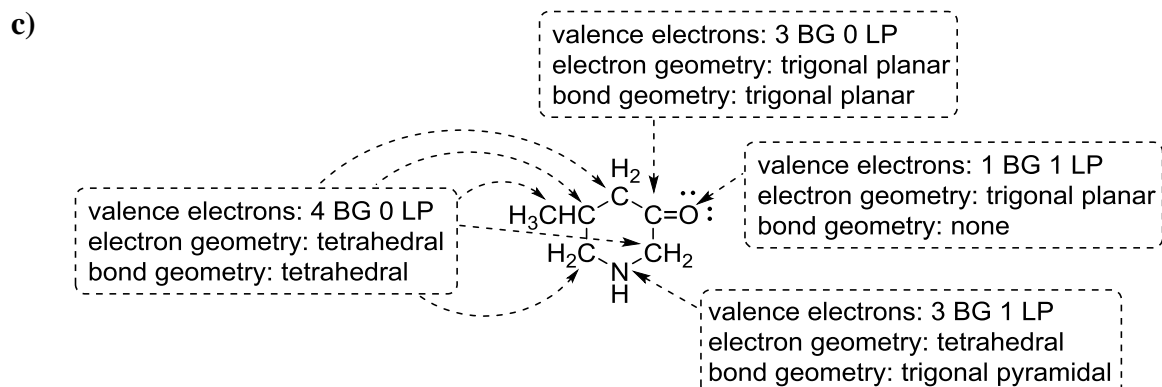
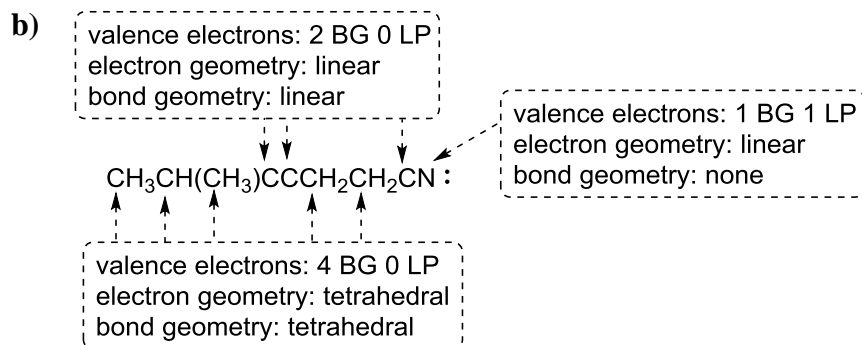
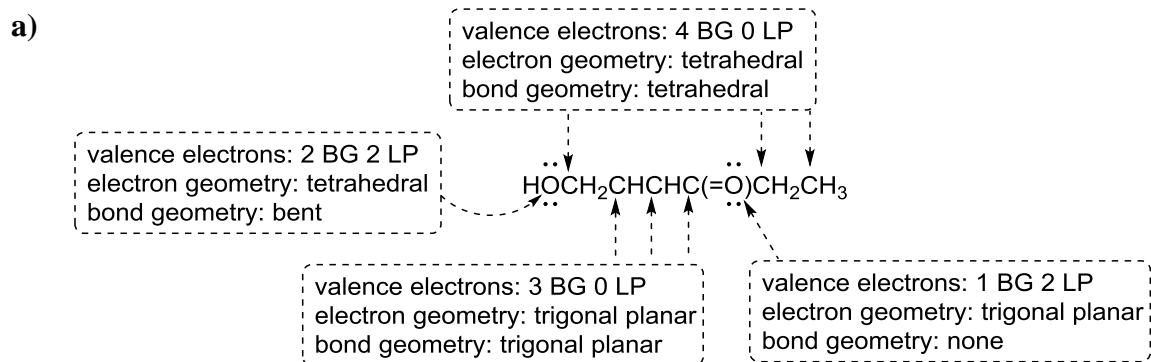




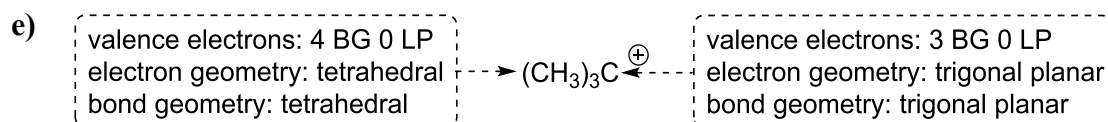
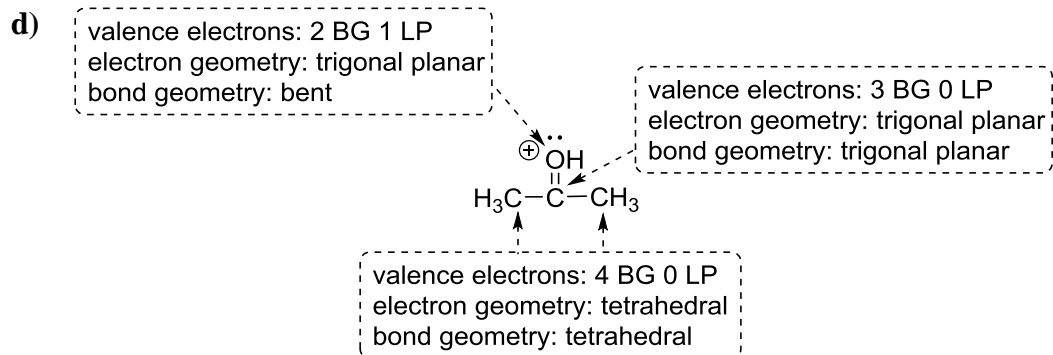


1.25

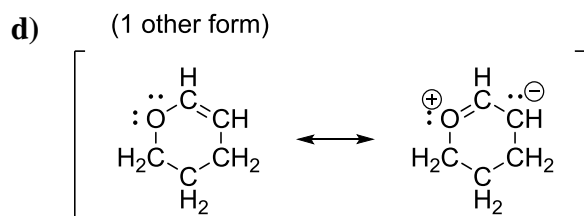
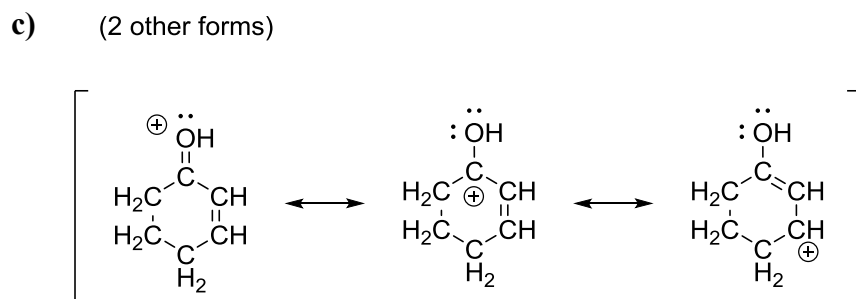
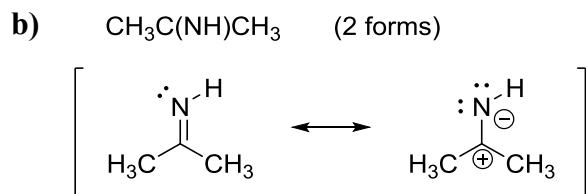
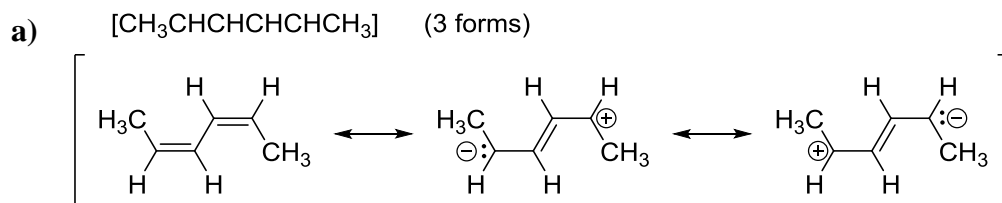
In each diagram, the hybridization, electron pair geometry, and bond geometry for the non-hydrogen atoms are indicated.





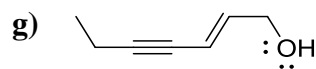
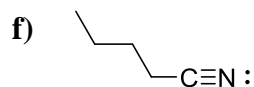
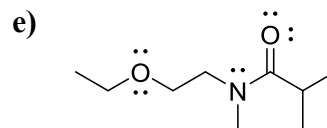
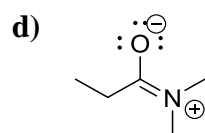
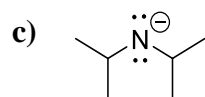
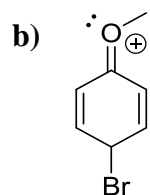
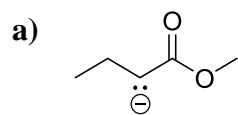


## 1.26

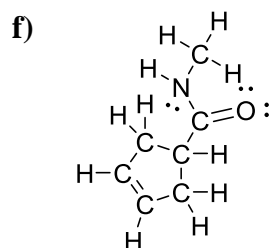
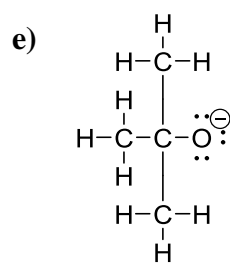
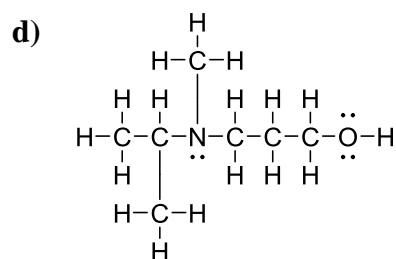
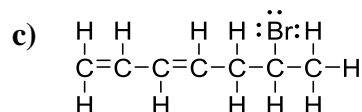
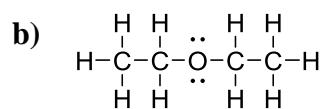
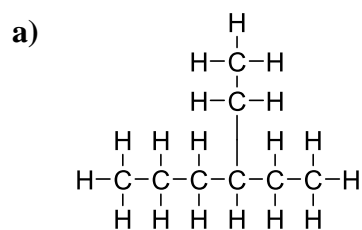




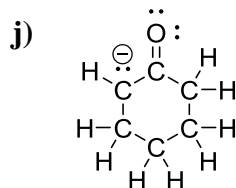
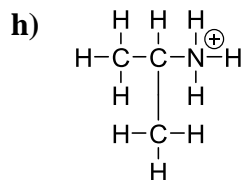
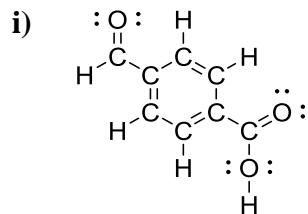
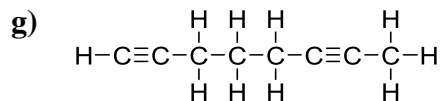
### 1.27



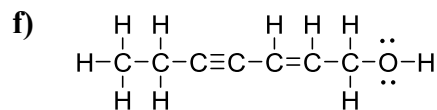
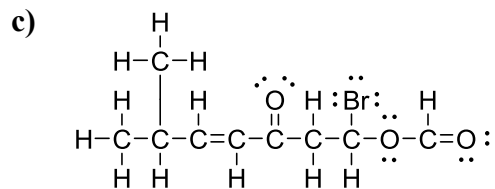
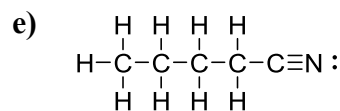
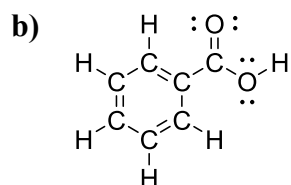
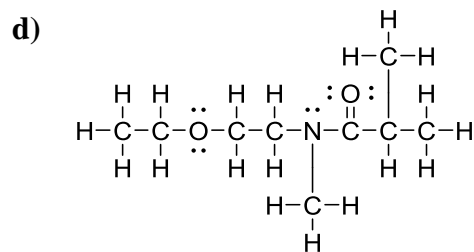
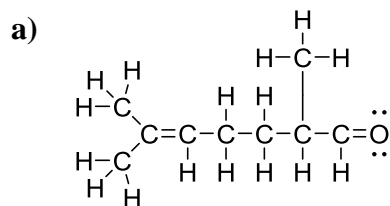
### 1.28



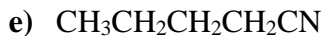
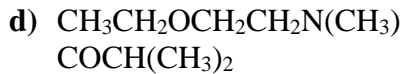
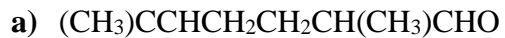




1.29



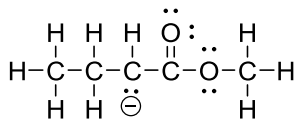
1.30



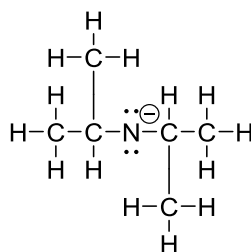


## 1.31

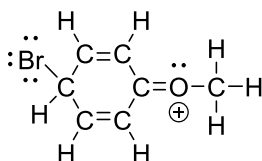
a)



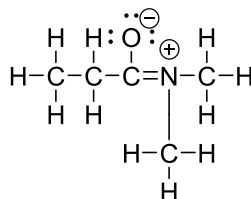
c)



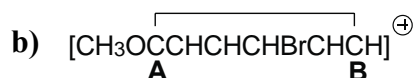
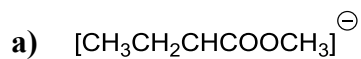
b)



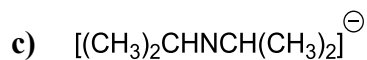
d)



## 1.32

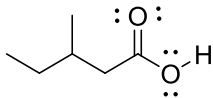


Note the use of the bracket drawn above this condensed structure to indicate the ring in the molecule. The bracket denotes the connection between carbon atoms **A** and **B**, which are actually beside each other in the molecule but are located at opposite ends of its condensed structure. All six carbon atoms in the ring are, therefore, in the bracket.

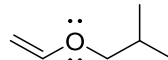


## 1.33

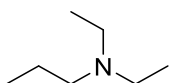
a)



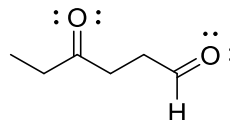
c)



b)



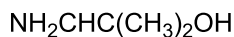
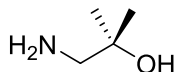
d)



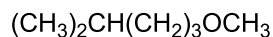
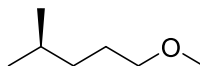


### 1.34

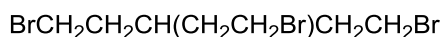
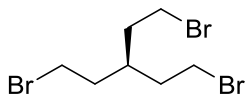
a)



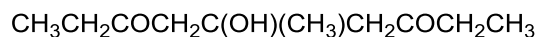
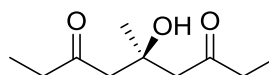
c)



b)

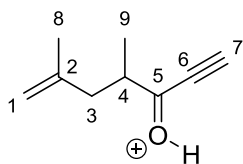


d)



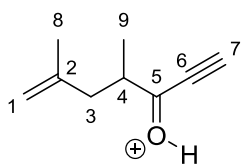
### 1.35

- a) The geometry at each non-H atom in the above molecule appears below. Carbons are numbered in the diagram for clarity.



C1, C2, C5 - trigonal planar  
C3, C4, C8, C9 - tetrahedral  
C6, C7 - linear  
O - bent ( $120^\circ$ )

- b) Hybridizations for each non-H atom:



C1, C2, C5 -  $\text{sp}^2$   
C3, C4, C8, C9 -  $\text{sp}^3$   
C6, C7 -  $\text{sp}$   
O -  $\text{sp}^2$

Bond descriptions:

Bond 1 – C  $\text{sp}^3$ - C  $\text{sp}^2$   $\sigma$ -bond

Bond 2 – C  $\text{sp}^2$ - C  $\text{sp}^2$   $\sigma$ -bond  
+ C p- C p  $\pi$ -bond

Bond 3 – C  $\text{sp}^3$ - C  $\text{sp}^3$   $\sigma$ -bond

Bond 4 – C  $\text{sp}^3$ - C  $\text{sp}^2$   $\sigma$ -bond

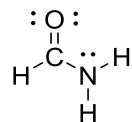
Bond 5 – C  $\text{sp}^2$ - C  $\text{sp}$   $\sigma$ -bond

Bond 6 – O  $\text{sp}^2$ - H s  $\sigma$ -bond



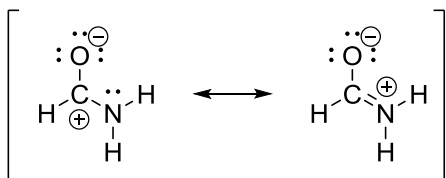
### 1.36

a) Lewis structure of formamide with filled valence atomic orbitals and formal charge:



formamide

b) Formamide resonance forms:



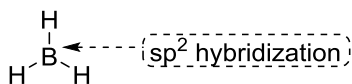
### 1.37

a) Boron has three valence electrons.

b) Lewis structure and geometry of  $\text{BH}_3$ :



c) Hybridization of a boron atom:



Boron has an incomplete octet, since there are only three shared bond pairs with hydrogen atoms.

### 1.38

A bonding  $\pi$  molecular orbital has electron density between the bonding atoms (when occupied) and strengthens the bond between them. An anti-bonding  $\pi$  molecular orbital has reduced electron density between the atoms (when occupied) and weakens the bond between them.



## MCAT Style Problems

1.39

Answer: (a). There is an undrawn H atom implied on the carbon atom in this representation.

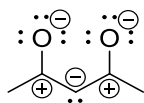
1.40

Answer: (d). In order to have complete octets on all the CH groups, one of them needs a lone pair, giving the carbon atom a formal negative charge.

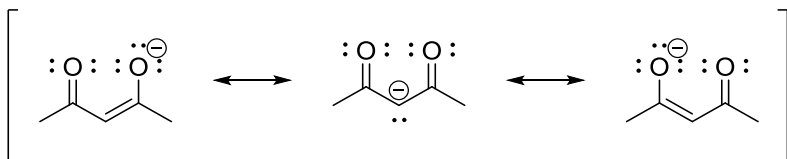
## Challenge Problem

1.41

Assigning the valence electrons as single bonds and lone pairs would lead to the following intermediate Lewis structure:



Forming bonds with the lone pairs to reduce the formal charges would lead to the resonance forms for the molecule, as shown below.



This shows that the two oxygen atoms and the central carbon atom have negative charges in one of the resonance forms. They will be the most likely sites to act as electron donors in reactions.