# 7 Covalent Bonding and Electron-Dot Structures

7.1 Only (a) is labeled incorrectly. The bonds in $SiCl_4$ are polar covalent.

7.2 H is positively polarized (blue). O is negatively polarized (red). This is consistent with the electronegativity values for O (3.5) and H (2.1). The more negatively polarized atom should be the one with the larger electronegativity.

7.3 $\mu = Q \times r = (1.60 \times 10^{-19}\ C)(228 \times 10^{-12}\ m)\left(\dfrac{1\ D}{3.336 \times 10^{-30}\ C \cdot m}\right) = 10.94\ D$

% ionic character for AgCl $= \dfrac{6.08\ D}{10.94\ D} \times 100\% = 55.6\%$

7.4 Li–H $\Delta EN = EN(H) - EN(Li) = 2.1 - 1.0 = 1.1$
H–F $\Delta EN = EN(F) - EN(H) = 4.0 - 2.1 = 1.9$
Based on the larger $\Delta EN$ value, you would predict that HF would have the higher % ionic character. The calculations below show the opposite is in fact the case.

LiH; $\mu = Q \times r = (1.60 \times 10^{-19}\ C)(160 \times 10^{-12}\ m)\left(\dfrac{1\ D}{3.336 \times 10^{-30}\ C \cdot m}\right) = 7.67\ D$

% ionic character for LiH $= \dfrac{6.00\ D}{7.67\ D} \times 100\% = 78.2\%$

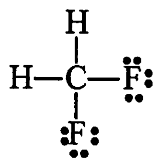
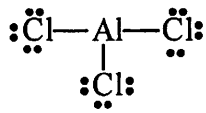
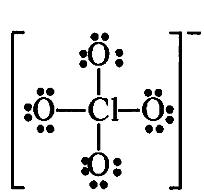
HF; $\mu = Q \times r = (1.60 \times 10^{-19}\ C)(92 \times 10^{-12}\ m)\left(\dfrac{1\ D}{3.336 \times 10^{-30}\ C \cdot m}\right) = 4.41\ D$

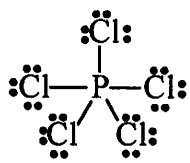
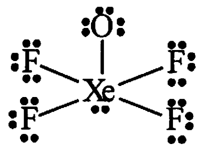
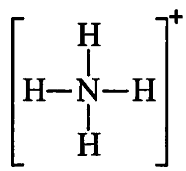
% ionic character for HF $= \dfrac{1.83\ D}{4.41\ D} \times 100\% = 41.4\%$

7.5 (c) $H:\ddot{\underset{\cdot\cdot}{S}}:H$

7.6 (a) $OF_2$ (b) $SiCl_4$

7.7 (a) is incorrect, the least electronegative atom (P) should be placed in the center.
(b) correct
(c) is incorrect, oxygen does not have a complete octet.
(d) is incorrect, there are 2 too many valence electrons.

7.8 (a)  (b)  (c) 

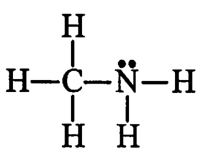
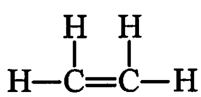
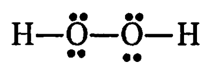
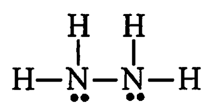
(d)  (e)  (e) 

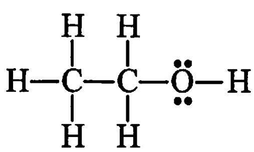
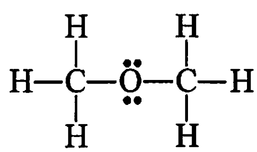
7.9 (d) :C≡O:

7.10 (e) Both (b) and (c) are correct.

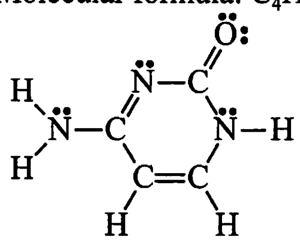
7.11 (c) is the correct structure because it has 19 valence electrons, the two oxygens have complete octets, and the odd electron is on the chlorine.

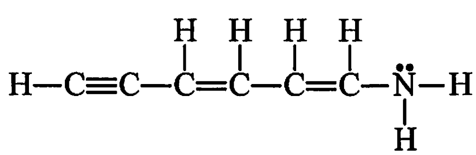
7.12 $O_2^-$ is a radical and the presence of an unpaired electron leads to a highly reactive species.

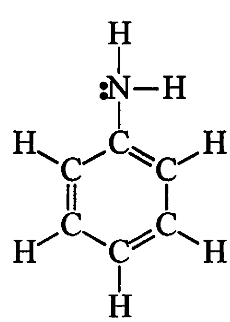
7.13 (a)  (b)  (c)  (d) 

7.14  and 

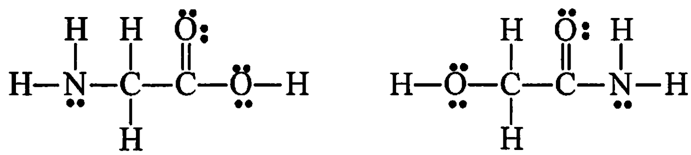
7.15 Molecular formula: $C_4H_5N_3O$



7.16 (a) 

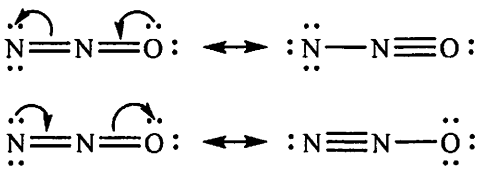


(b)

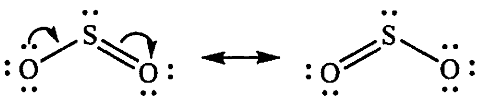


7.17 (a)

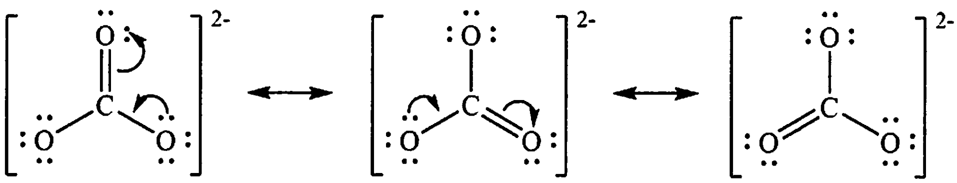
(b)



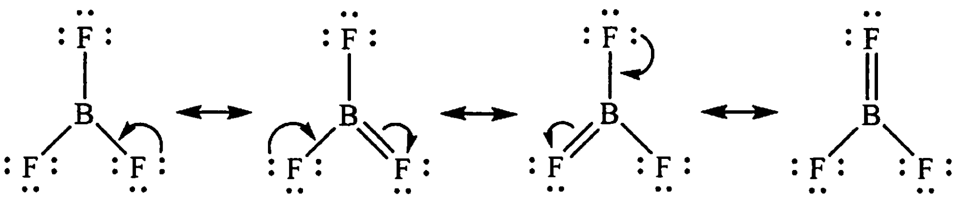
7.18 (a)



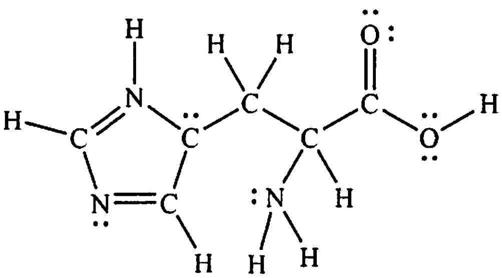
(b)



(c)

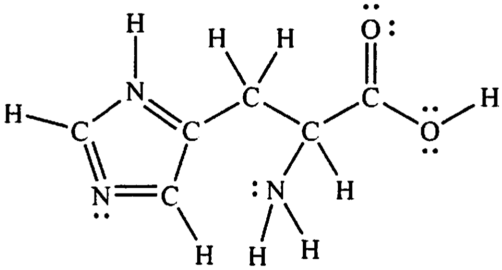


7.19 (a)



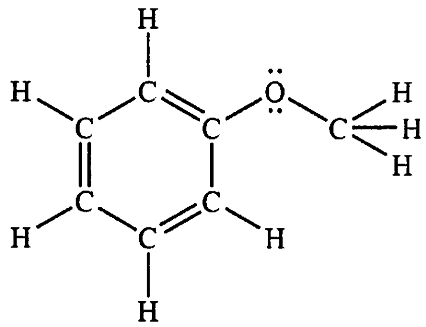
This is a valid resonance structure.

(b)

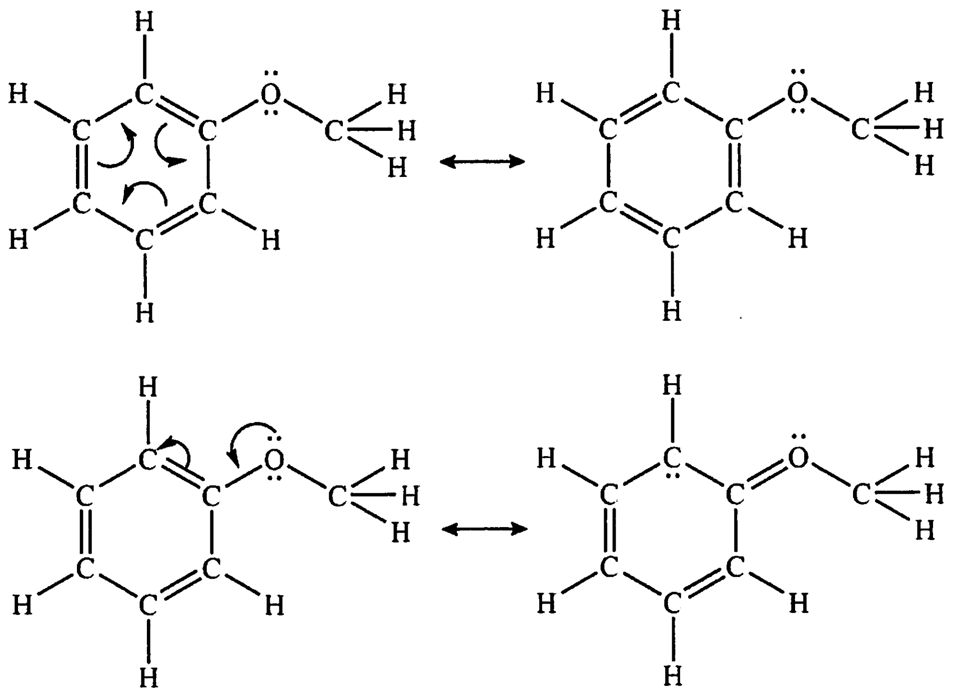


This is not a valid resonance structure. The N with two double bonds has 10 electrons, not 8.

7.20 (a)



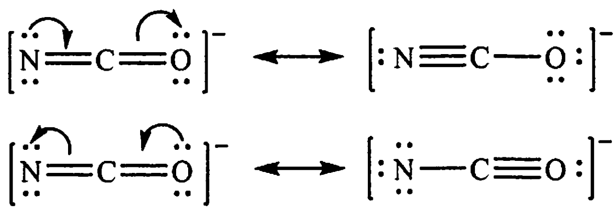
(b)



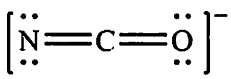
7.21 $\left[:\ddot{N}=C=\ddot{O}:\right]^-$

| | | |
|---|---|---|
| For nitrogen: | Isolated nitrogen valence electrons | 5 |
| | Bound nitrogen bonding electrons | 4 |
| | Bound nitrogen nonbonding electrons | 4 |
| | Formal charge = 5 – ½(4) – 4 = –1 | |
| For carbon: | Isolated carbon valence electrons | 4 |
| | Bound carbon bonding electrons | 8 |
| | Bound carbon nonbonding electrons | 0 |
| | Formal charge = 4 – ½ (8) – 0 = 0 | |
| For oxygen: | Isolated oxygen valence electrons | 6 |
| | Bound oxygen bonding electrons | 4 |
| | Bound oxygen nonbonding electrons | 4 |
| | Formal charge = 6 – ½(4) – 4 = 0 | |

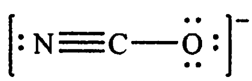
7.22 (a)



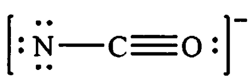
(b)



| | | |
|---|---|---|
| For nitrogen: | Isolated nitrogen valence electrons | 5 |
| | Bound nitrogen bonding electrons | 4 |
| | Bound nitrogen nonbonding electrons | 4 |
| | Formal charge = 5 – ½(4) – 4 = –1 | |
| For carbon: | Isolated carbon valence electrons | 4 |
| | Bound carbon bonding electrons | 8 |
| | Bound carbon nonbonding electrons | 0 |
| | Formal charge = 4 – ½(8) – 0 = 0 | |
| For oxygen: | Isolated oxygen valence electrons | 6 |
| | Bound oxygen bonding electrons | 4 |
| | Bound oxygen nonbonding electrons | 4 |
| | Formal charge = 6 – ½(4) – 4 = 0 | |



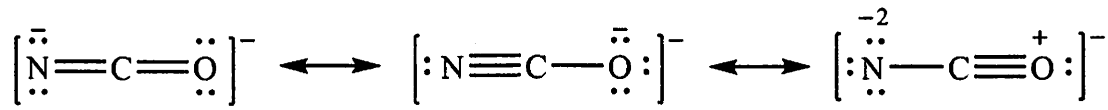
| | | |
|---|---|---|
| For nitrogen: | Isolated nitrogen valence electrons | 5 |
| | Bound nitrogen bonding electrons | 6 |
| | Bound nitrogen nonbonding electrons | 2 |
| | Formal charge = 5 – ½(6) – 2 = 0 | |
| For carbon: | Isolated carbon valence electrons | 4 |
| | Bound carbon bonding electrons | 8 |
| | Bound carbon nonbonding electrons | 0 |
| | Formal charge = 4 – ½(8) – 0 = 0 | |
| For oxygen: | Isolated oxygen valence electrons | 6 |
| | Bound oxygen bonding electrons | 2 |
| | Bound oxygen nonbonding electrons | 6 |
| | Formal charge = 6 – ½(2) – 6 = –1 | |



| | | |
|---|---|---|
| For nitrogen: | Isolated nitrogen valence electrons | 5 |
| | Bound nitrogen bonding electrons | 2 |
| | Bound nitrogen nonbonding electrons | 6 |
| | Formal charge = 5 – ½(2) – 6 = –2 | |

| | | |
|---|---|---|
| For carbon: | Isolated carbon valence electrons | 4 |
| | Bound carbon bonding electrons | 8 |
| | Bound carbon nonbonding electrons | 0 |
| | Formal charge = 4 – ½(8) – 0 = 0 | |

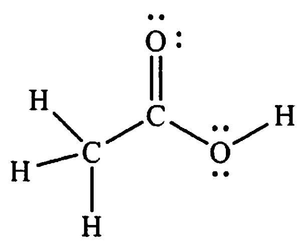
| | | |
|---|---|---|
| For oxygen: | Isolated oxygen valence electrons | 6 |
| | Bound oxygen bonding electrons | 6 |
| | Bound oxygen nonbonding electrons | 2 |
| | Formal charge = 6 – ½(6) – 2 = +1 | |



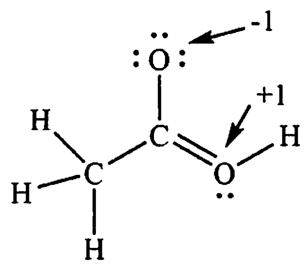
The first two structures make the largest contribution to the resonance hybrid because the –1 formal charge is on either of the electronegative N or O. The third structure has a +1 formal charge on O and does not significantly contribute to the resonance hybrid.

(c) Carbon–nitrogen because of the triple bond contribution to the resonance hybrid.

7.23



All atoms in this structure have 0 formal charge.



| | | |
|---|---|---|
| For top oxygen: | Isolated oxygen valence electrons | 6 |
| | Bound oxygen bonding electrons | 2 |
| | Bound oxygen nonbonding electrons | 6 |
| | Formal charge = 6 – ½(2) – 6 = –1 | |

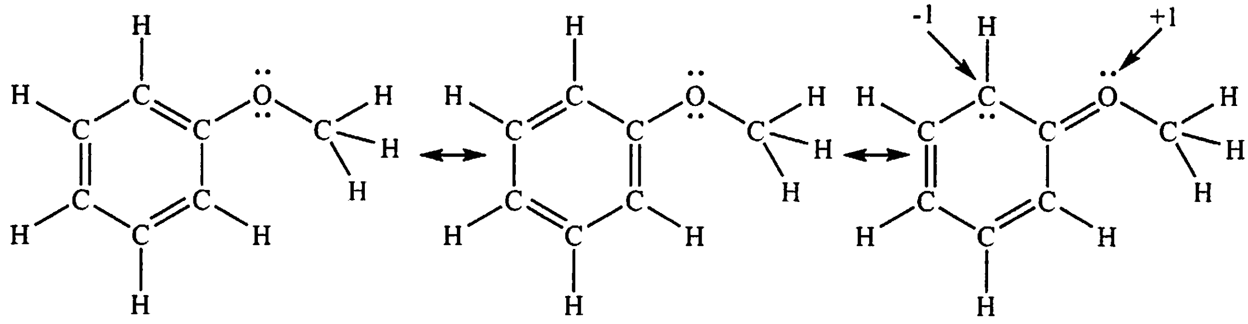
| | | |
|---|---|---|
| For right oxygen: | Isolated oxygen valence electrons | 6 |
| | Bound oxygen bonding electrons | 6 |
| | Bound oxygen nonbonding electrons | 2 |
| | Formal charge = 6 – ½(6) – 2 = +1 | |

All the other atoms in this structure have 0 formal charge.

The structure without formal charges makes a larger contribution to the resonance hybrid because energy is required to separate + and – charges. Thus, the actual electronic structure of acetic acid is closer to that of the more favorable, lower energy structure.

7.24



1 2 3

All atoms in structures 1 and 2 have 0 formal charge.

In structure 3:

| | | |
|---|---|---|
| For carbon: | Isolated carbon valence electrons | 4 |
| | Bound carbon bonding electrons | 6 |
| | Bound carbon nonbonding electrons | 2 |
| | Formal charge = 4 – ½(6) – 2 = –1 | |
| For oxygen: | Isolated oxygen valence electrons | 6 |
| | Bound oxygen bonding electrons | 6 |
| | Bound oxygen nonbonding electrons | 2 |
| | Formal charge = 6 – ½(6) – 2 = +1 | |

All the other atoms in structure 3 have 0 formal charge.

The structures 1 and 2 without formal charges make the larger contribution to the resonance hybrid because energy is required to separate + and – charges.

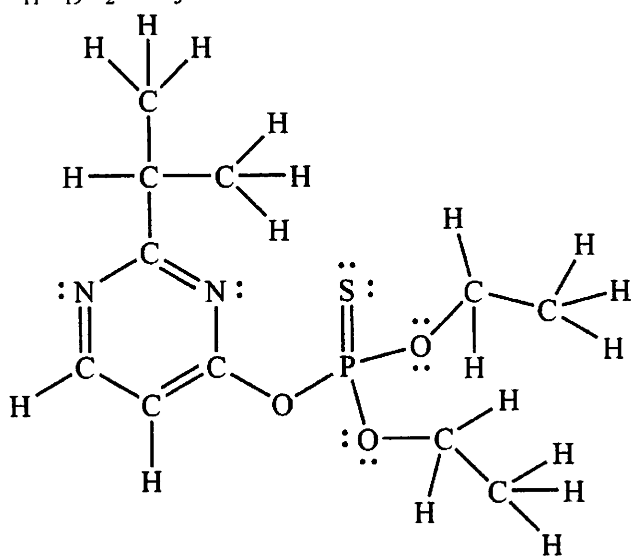
7.25 (a) ΔEN(P=O) = 1.4 and ΔEN(P=S) = 0.4. Both bonds are polar covalent, but the phosphorus-oxygen bond is more polar.
(b) For the reaction between the organophosphate insecticide and the enzyme to occur, the phosphorus atom must bear a positive charge. Greater positive charge leads to increased rate of reaction and increased toxicity of the insecticide. Since oxygen is more electronegative than sulfur it has greater ability to pull shared electrons toward itself. Therefore, the molecule with the P=O bond is more toxic due to the greater positive charge on phosphorus.

7.26 For the reaction between an organophosphate insecticide and the enzyme to occur, the phosphorus atom must bear a positive charge. Greater positive charge leads to increased rate of reaction and increased toxicity of the insecticide. The more electronegative the X group, the more positive the phosphorus. (a) Cl (b) $CF_3$

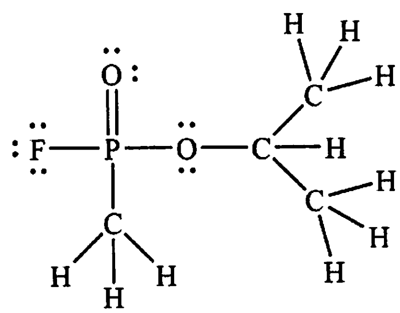
7.27 Phosphorus is in the third row of the periodic table and can utilize d orbitals to hold extra electrons, and therefore form more than four bonds.

7.28 $C_{11}H_{19}N_2PSO_3$



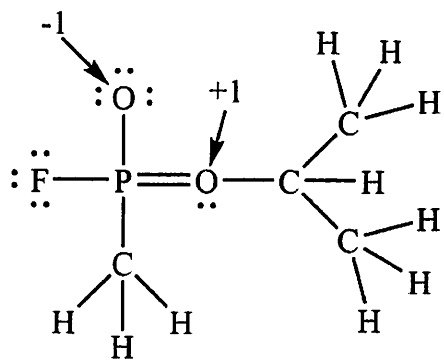
7.29



| | | |
|---|---|---|
| For phosphorus: | Isolated phosphorus valence electrons | 5 |
| | Bound phosphorus bonding electrons | 10 |
| | Bound phosphorus nonbonding electrons | 0 |
| | Formal charge = 5 – ½(10) – 0 = 0 | |
| For top oxygen: | Isolated oxygen valence electrons | 6 |
| | Bound oxygen bonding electrons | 4 |
| | Bound oxygen nonbonding electrons | 4 |
| | Formal charge = 6 – ½(4) – 4 = 0 | |
| For right oxygen: | Isolated oxygen valence electrons | 6 |
| | Bound oxygen bonding electrons | 4 |
| | Bound oxygen nonbonding electrons | 4 |
| | Formal charge = 6 – ½(4) – 4 = 0 | |
| For fluorine: | Isolated fluorine valence electrons | 7 |
| | Bound fluorine bonding electrons | 2 |
| | Bound fluorine nonbonding electrons | 6 |
| | Formal charge = 7 – ½(2) – 6 = 0 | |

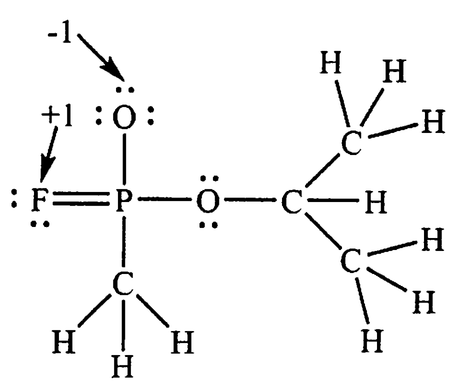
| | | |
|---|---|---|
| For carbon: | Isolated carbon valence electrons | 4 |
| | Bound carbon bonding electrons | 8 |
| | Bound carbon nonbonding electrons | 0 |
| | Formal charge = 4 – ½(8) – 0 = 0 | |

7.30 (a)




| | | |
|---|---|---|
| For fluorine: | Isolated fluorine valence electrons | 7 |
| | Bound fluorine bonding electrons | 2 |
| | Bound fluorine nonbonding electrons | 6 |
| | Formal charge = 7 – ½(2) – 6 = 0 | |
| For phosphorus: | Isolated phosphorus valence electrons | 5 |
| | Bound phosphorus bonding electrons | 10 |
| | Bound phosphorus nonbonding electrons | 0 |
| | Formal charge = 5 – ½(10) – 0 = 0 | |
| For top oxygen: | Isolated oxygen valence electrons | 6 |
| | Bound oxygen bonding electrons | 2 |
| | Bound oxygen nonbonding electrons | 6 |
| | Formal charge = 6 – ½(2) – 6 = –1 | |
| For right oxygen: | Isolated oxygen valence electrons | 6 |
| | Bound oxygen bonding electrons | 6 |
| | Bound oxygen nonbonding electrons | 2 |
| | Formal charge = 6 – ½(6) – 2 = +1 | |

(b)



For fluorine:

| | | |
|---|---|---|
| | Isolated fluorine valence electrons | 7 |
| | Bound fluorine bonding electrons | 4 |
| | Bound fluorine nonbonding electrons | 4 |
| | Formal charge = 7 – ½(4) – 4 = +1 | |

For phosphorus:

| | | |
|---|---|---|
| | Isolated phosphorus valence electrons | 5 |
| | Bound phosphorus bonding electrons | 10 |
| | Bound phosphorus nonbonding electrons | 0 |
| | Formal charge = 5 – ½(10) – 0 = 0 | |

For top oxygen:

| | | |
|---|---|---|
| | Isolated oxygen valence electrons | 6 |
| | Bound oxygen bonding electrons | 2 |
| | Bound oxygen nonbonding electrons | 6 |
| | Formal charge = 6 – ½(2) – 6 = –1 | |

For right oxygen:

| | | |
|---|---|---|
| | Isolated oxygen valence electrons | 6 |
| | Bound oxygen bonding electrons | 4 |
| | Bound oxygen nonbonding electrons | 4 |
| | Formal charge = 6 – ½(4) – 4 = 0 | |

The problem 7.29 structure without formal charges makes the largest contribution to the resonance hybrid because energy is required to separate + and – charges.

## Conceptual Problems

7.31 (a) A (b) D (c) B (d) C

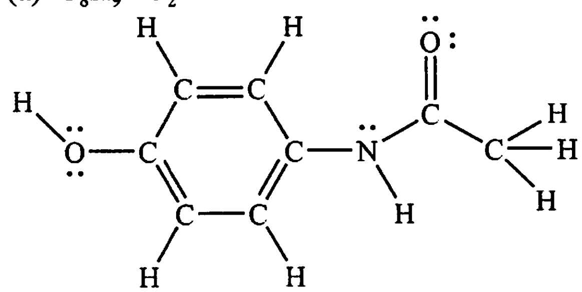
7.32 C–D is the stronger bond. A–B is the longer bond.

7.33 As the electrostatic potential maps are drawn, the Li and Cl are at the tops of each map. The red area is for a negatively polarized region (associated with Cl). The blue area is for a positively polarized region (associated with Li). Map (a) is for $CH_3Cl$ and Map (b) is for $CH_3Li$.

7.34 (a) fluoroethane (b) ethane (c) ethanol (d) acetaldehyde

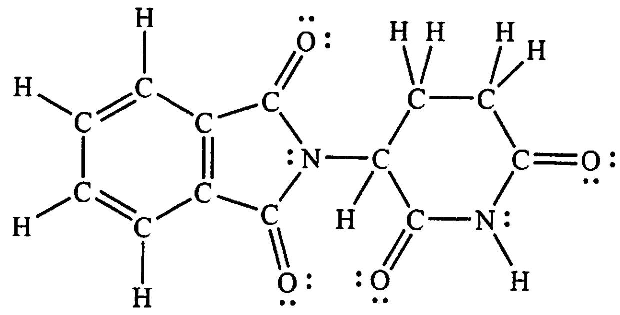
7.35 (a) is ionic; (b) and (c) are covalent.

7.36 (a) $C_8H_9NO_2$

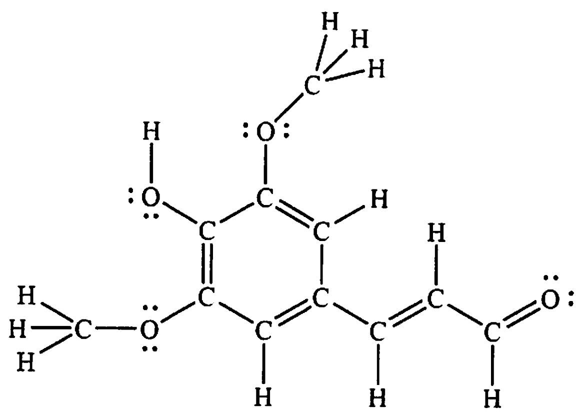


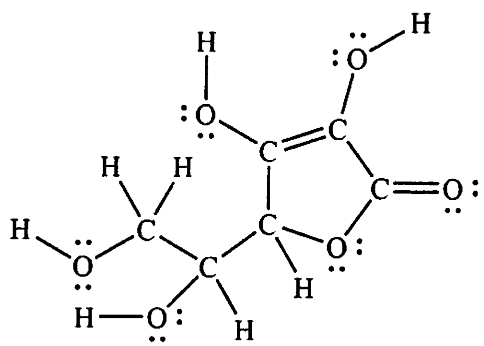
7.37 (a) $C_{13}H_{10}N_2O_4$



7.38



7.39



**Section Problems**
**Covalent Bonds (Sections 7.1)**

7.40 (a) ionic (b) nonpolar covalent (c) covalent

7.41 (a) Attractive forces between the positively charged nuclei and the electrons in both atoms occur when the atoms are close together and a covalent bond forms.

**Strengths of Covalent Bonds (Section 7.2)**

7.42 C–F 450 kJ/mol $\Delta EN = EN(F) - EN(C) = 4.0 - 2.5 = 1.5$
N–F 270 kJ/mol $\Delta EN = EN(F) - EN(N) = 4.0 - 3.0 = 1.0$
O–F 180 kJ/mol $\Delta EN = EN(F) - EN(O) = 4.0 - 3.5 = 0.5$

F–F 159 kJ/mol $\Delta EN = EN(F) - EN(F) = 4.0 - 4.0 = 0$
In general, increased bond polarity leads to increased bond strength.

7.43 C–F 450 kJ/mol $\Delta EN = EN(F) - EN(C) = 4.0 - 2.5 = 1.5$
C–Cl 330 kJ/mol $\Delta EN = EN(Cl) - EN(C) = 3.0 - 2.5 = 0.5$
C–Br 270 kJ/mol $\Delta EN = EN(Br) - EN(C) = 2.8 - 2.5 = 0.3$
C–I 240 kJ/mol $\Delta EN = EN(I) - EN(C) = 2.5 - 2.5 = 0$
In general, increased bond polarity leads to increased bond strength.

7.44 N–H $\Delta EN = EN(N) - EN(H) = 3.0 - 2.1 = 0.9$
O–H $\Delta EN = EN(O) - EN(H) = 3.5 - 2.1 = 1.4$
S–H $\Delta EN = EN(S) - EN(H) = 2.5 - 2.1 = 0.4$
In general, increased bond polarity leads to increased bond strength. The most polar bond is the O–H bond and should be the strongest of the three.

7.45 All three bonds are nonpolar. In general, the longer the bond length, the weaker the bond. I > Br > Cl. The I–I bond is the longest and should be the weakest of the three.

**Polar Covalent Bonds: Electronegativity (Section 7.3)**

7.46 Electronegativity increases from left to right across a period and decreases down a group.

7.47 Z = 119 would be below francium and have a very low electronegativity.

7.48 K < Li < Mg < Pb < C < Br

7.49 Cl > C > Cu > Ca > Cs

7.50 (a) HF

| | | | |
|---|---|---|---|
| fluorine | EN = 4.0 | | |
| hydrogen | EN = 2.1 | | |
| | $\Delta EN = 1.9$ | HF is polar covalent. | |

(b) HI

| | | |
|---|---|---|
| iodine | EN = 2.5 | |
| hydrogen | EN = 2.1 | |
| | $\Delta EN = 0.4$ | HI is polar covalent. |

(c) $PdCl_2$

| | | |
|---|---|---|
| chlorine | EN = 3.0 | |
| palladium | EN = 2.2 | |
| | $\Delta EN = 0.8$ | $PdCl_2$ is polar covalent. |

(d) $BBr_3$

| | | |
|---|---|---|
| bromine | EN = 2.8 | |
| boron | EN = 2.0 | |
| | $\Delta EN = 0.8$ | $BBr_3$ is polar covalent. |

(e) NaOH — $Na^+ - OH^-$ is ionic
$OH^-$
oxygen EN = 3.5
hydrogen EN = 2.1
ΔEN = 1.4 — $OH^-$ is polar covalent.

(f) $CH_3Li$
lithium EN = 1.0
carbon EN = 2.5
ΔEN = 1.5 — $CH_3Li$ is polar covalent.

7.51 The electronegativity for each element is shown in parentheses.
(a) C (2.5), H (2.1), Cl (3.0): The C–Cl bond is more polar than the C–H bond because of the larger electronegativity difference between the bonded atoms.
(b) Si (1.8), Li (1.0), Cl (3.0): The Si–Cl bond is more polar than the Si–Li bond because of the larger electronegativity difference between the bonded atoms.
(c) N (3.0), Cl (3.0), Mg (1.2): The N–Mg bond is more polar than the N–Cl bond because of the larger electronegativity difference between the bonded atoms.

7.52 (a) $\overset{\delta-}{C} - \overset{\delta+}{H}$ $\quad \overset{\delta+}{C} - \overset{\delta-}{Cl}$ (b) $\overset{\delta-}{Si} - \overset{\delta+}{Li}$ $\quad \overset{\delta+}{Si} - \overset{\delta-}{Cl}$

(c) $N - Cl$ $\quad \overset{\delta-}{N} - \overset{\delta+}{Mg}$

7.53 (a) $\overset{\delta-}{F} - \overset{\delta+}{H}$ (b) $\overset{\delta-}{I} - \overset{\delta+}{H}$ (c) $\overset{\delta-}{Cl} - \overset{\delta+}{Pd}$

(d) $\overset{\delta-}{Br} - \overset{\delta+}{B}$ (e) $\overset{\delta-}{O} - \overset{\delta+}{H}$ (f) $\overset{\delta-}{C} - \overset{\delta+}{Li}$

7.54 (a) MgO, $BaCl_2$ (b) $P_4$ (c) $CdBr_2$, $BrF_3$, $NF_3$, $POCl_3$, LiBr

7.55 (a) $CaCl_2$, NaF, LiF (b) $S_8$ (c) $SOCl_2$, $CBr_4$, BrCl, $AsH_3$

7.56 (a) $CCl_4$
chlorine EN = 3.0
carbon EN = 2.5
ΔEN = 0.5

(b) $BaCl_2$
chlorine EN = 3.0
barium EN = 0.9
ΔEN = 2.1

(c) $TiCl_3$
chlorine EN = 3.0
titanium EN = 1.5
ΔEN = 1.5

(d) $Cl_2O$ oxygen $EN = 3.5$
chlorine $\underline{EN = 3.0}$
$\Delta EN = 0.5$

Increasing ionic character: $CCl_4 \sim ClO_2 < TiCl_3 < BaCl_2$

7.57 (a) $NH_3$ nitrogen $EN = 3.0$
hydrogen $\underline{EN = 2.1}$
$\Delta EN = 0.9$

(b) $NCl_3$ nitrogen $EN = 3.0$
chlorine $\underline{EN = 3.0}$
$\Delta EN = 0.0$

(c) $Na_3N$ nitrogen $EN = 3.0$
sodium $\underline{EN = 0.9}$
$\Delta EN = 2.1$

(d) $NO_2$ oxygen $EN = 3.5$
nitrogen $\underline{EN = 3.0}$
$\Delta EN = 0.5$

Increasing ionic character: $NCl_3 < NO_2 < NH_3 < Na_3N$

7.58 (a) $MgBr_2$ (b) $PBr_3$

7.59 (a) $CaCl_2$ (b) $SiCl_4$

7.60 $$\mu = Q \times r = (1.60 \times 10^{-19}\ C)(213.9 \times 10^{-12}\ m)\left(\frac{1\ D}{3.336 \times 10^{-30}\ C\cdot m}\right) = 10.26\ D$$

$$\%\ \text{ionic character for BrCl} = \frac{0.518\ D}{10.26\ D} \times 100\% = 5.05\%$$

7.61 $$\mu = Q \times r = (1.60 \times 10^{-19}\ C)(162.8 \times 10^{-12}\ m)\left(\frac{1\ D}{3.336 \times 10^{-30}\ C\cdot m}\right) = 7.81\ D$$

$$\%\ \text{ionic character for ClF} = \frac{0.887\ D}{7.81\ D} \times 100\% = 11.4\%$$

## A Comparison of Ionic and Covalent Compounds (Section 7.4)

7.62 (b) and (c) are ionic compounds; (a) is a covalent compound and is most likely a gas at room temperature.

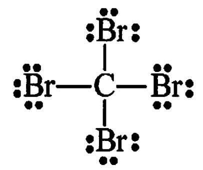
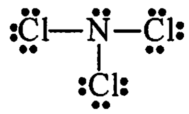
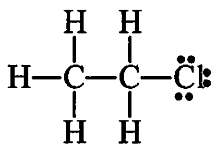
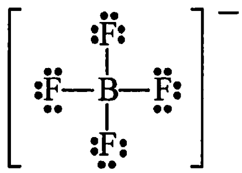
7.63 (a) and (b) are covalent compounds; (c) is an ionic compound and is most likely a solid at room temperature.

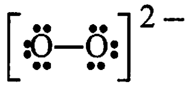
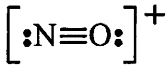
 

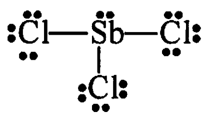
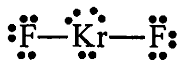
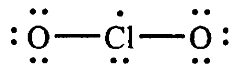
## Electron-Dot Structures and Resonance (Sections 7.5–7.7)

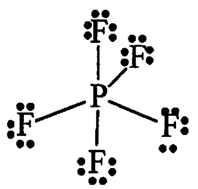
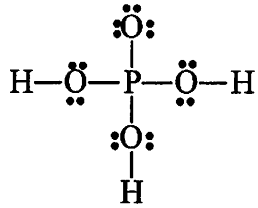
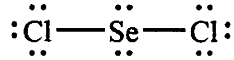
7.64 The octet rule states that main-group elements tend to react so that they attain a noble gas electron configuration with filled s and p sublevels (8 electrons) in their valence electron shells. The transition metals are characterized by partially filled d orbitals that can be used to expand their valence shell beyond the normal octet of electrons.

7.65 (a) $AlCl_3$ Al has only 6 electrons around it. (c) $PCl_5$ P has 10 electrons around it.

7.66 (a)  (b)  (c)  (d) 

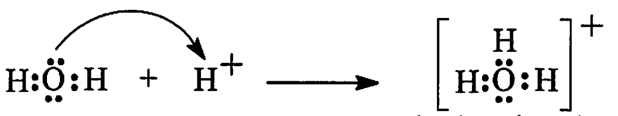
(e)  (f) 

7.67 (a)  (b)  (c) 

(d)  (e)  (f) 

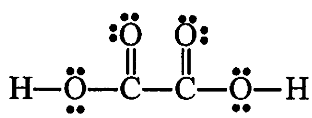
7.68 (c) is the correct electron-dot for $XeF_5^+$ because it accounts for the required number of bonding (10) and lone pair (32) electrons.

7.69



hydronium ion

7.70



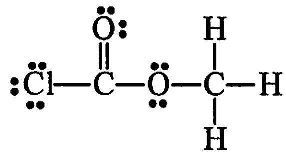
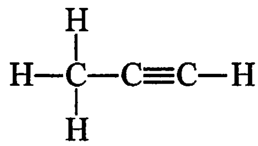
7.71 $:\ddot{S}=C=\ddot{S}:$ ; $CS_2$ has two double bonds.

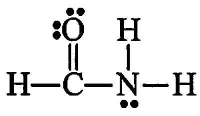
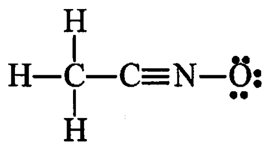
7.72 (a) The anion has 32 valence electrons. Each Cl has seven valence electrons (28 total). The minus one charge on the anion accounts for one valence electron. This leaves three valence electrons for X. X is Al.

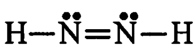
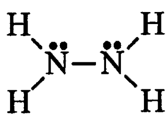
 

(b) The cation has eight valence electrons. Each H has one valence electron (4 total). X is left with four valence electrons. Since this is a cation, one valence electron was removed from X. X has five valence electrons. X is P.

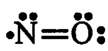
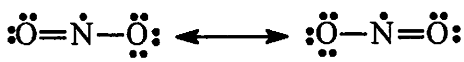
7.73 (a) This fourth-row element has six valence electrons. It is Se.
(b) This fourth-row element has eight valence electrons. It is Kr.

7.74 (a)  (b) 

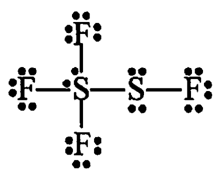
7.75 (a)  (b) 

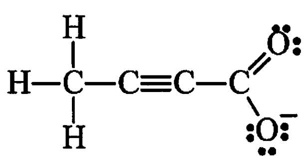
7.76  

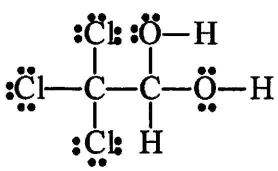
$N_2H_2$ has the stronger N–N bond because of the higher bond order.

7.77  

NO has the stronger N–O bond because of the higher bond order.

7.78 (a) 

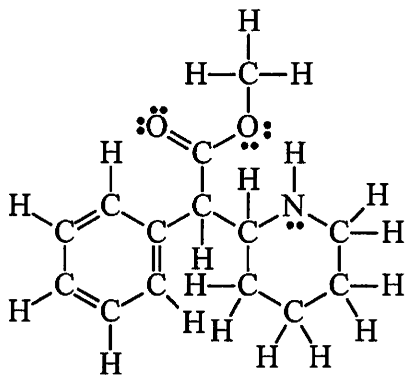
(b) 

7.79 

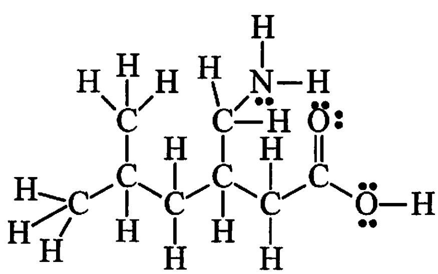
 

## Electron-Dot Structures for Molecules with Second-Row Elements (Section 7.8)

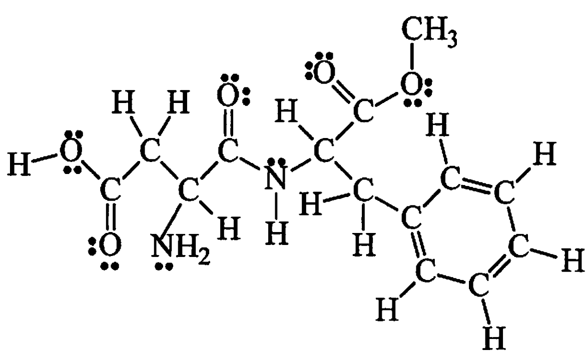
7.80



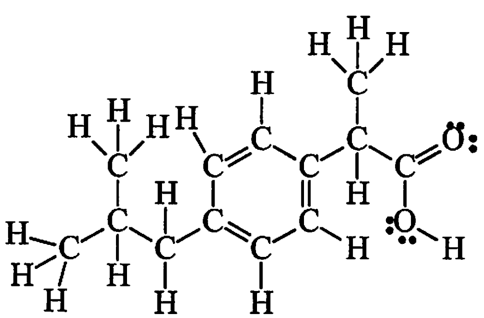
7.81



7.82

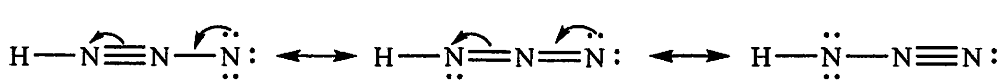


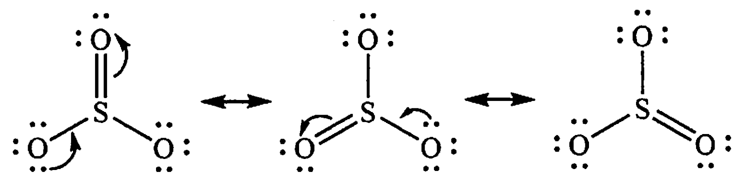
7.83

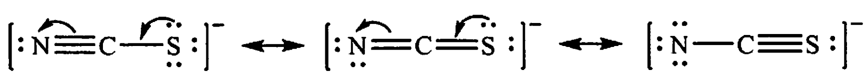


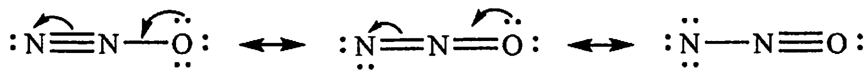
## Resonance (Section 7.9)

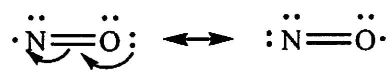
7.84 (a)

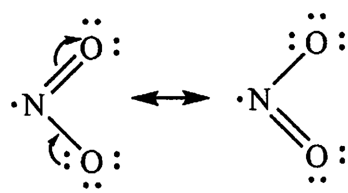


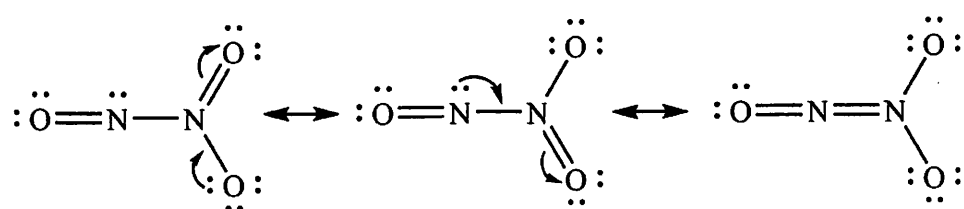
(b) 

(c) 

7.85 (a) 

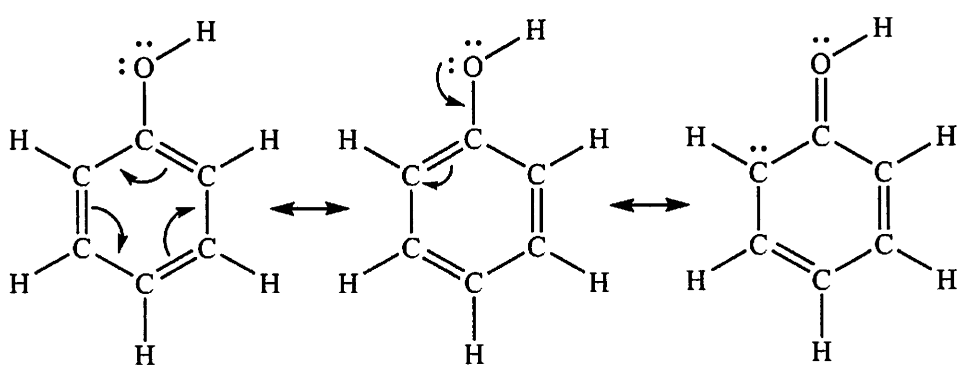
(b) 

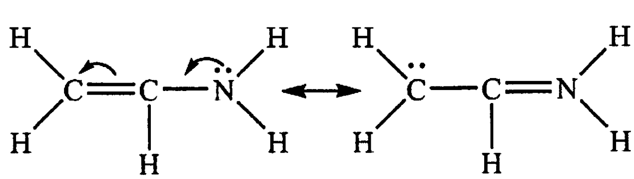
(c) 

(d) 

7.86 (a) yes (b) no (c) yes (d) yes

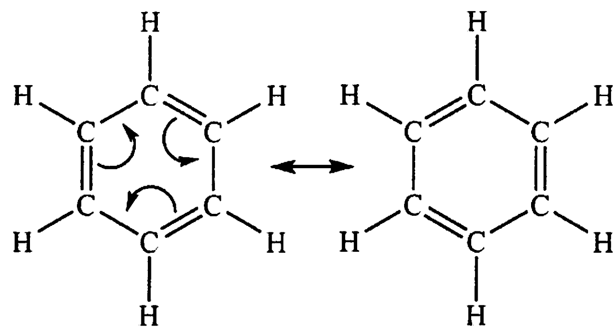
7.87 (a) yes (b) no (c) yes

7.88 

7.89 

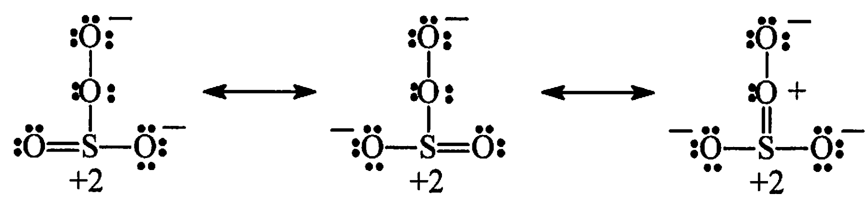
 

7.90 (a)

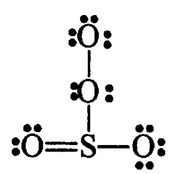


(b) ii)

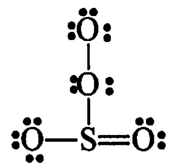
7.91



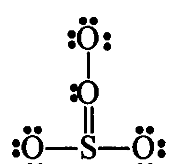
For



| | | |
|---|---|---|
| For oxygen: (top) | Isolated oxygen valence electrons | 6 |
| | Bound oxygen bonding electrons | 2 |
| | Bound oxygen nonbonding electrons | 6 |
| | Formal charge = 6 – ½(2) – 6 = –1 | |
| For oxygen: (middle) | Isolated oxygen valence electrons | 6 |
| | Bound oxygen bonding electrons | 4 |
| | Bound oxygen nonbonding electrons | 4 |
| | Formal charge = 6 – ½(4) – 4 = 0 | |
| For oxygen: (left) | Isolated oxygen valence electrons | 6 |
| | Bound oxygen bonding electrons | 4 |
| | Bound oxygen nonbonding electrons | 4 |
| | Formal charge = 6 – ½(4) – 4 = 0 | |
| For oxygen: (right) | Isolated oxygen valence electrons | 6 |
| | Bound oxygen bonding electrons | 2 |
| | Bound oxygen nonbonding electrons | 6 |
| | Formal charge = 6 – ½(2) – 6 = –1 | |
| For sulfur: | Isolated sulfur valence electrons | 6 |
| | Bound sulfur bonding electrons | 8 |
| | Bound sulfur nonbonding electrons | 0 |
| | Formal charge = 6 – ½(8) – 0 = +2 | |

For 

| | | |
|---|---|---|
| For oxygen: (top) | Isolated oxygen valence electrons | 6 |
| | Bound oxygen bonding electrons | 2 |
| | Bound oxygen nonbonding electrons | 6 |
| | Formal charge = 6 – ½(2) – 6 = –1 | |
| For oxygen: (middle) | Isolated oxygen valence electrons | 6 |
| | Bound oxygen bonding electrons | 4 |
| | Bound oxygen nonbonding electrons | 4 |
| | Formal charge = 6 – ½(4) – 4 = 0 | |
| For oxygen: (left) | Isolated oxygen valence electrons | 6 |
| | Bound oxygen bonding electrons | 2 |
| | Bound oxygen nonbonding electrons | 6 |
| | Formal charge = 6 – ½(2) – 6 = –1 | |
| For oxygen: (right) | Isolated oxygen valence electrons | 6 |
| | Bound oxygen bonding electrons | 4 |
| | Bound oxygen nonbonding electrons | 4 |
| | Formal charge = 6 – ½(4) – 4 = 0 | |
| For sulfur: | Isolated sulfur valence electrons | 6 |
| | Bound sulfur bonding electrons | 8 |
| | Bound sulfur nonbonding electrons | 0 |
| | Formal charge = 6 – ½(8) – 0 = +2 | |

For 

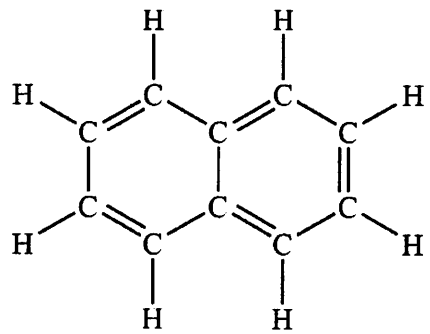
| | | |
|---|---|---|
| For oxygen: (top) | Isolated oxygen valence electrons | 6 |
| | Bound oxygen bonding electrons | 2 |
| | Bound oxygen nonbonding electrons | 6 |
| | Formal charge = 6 – ½(2) – 6 = –1 | |
| For oxygen: (middle) | Isolated oxygen valence electrons | 6 |
| | Bound oxygen bonding electrons | 6 |
| | Bound oxygen nonbonding electrons | 2 |
| | Formal charge = 6 – ½(6) – 2 = +1 | |

| | | |
|---|---|---|
| For oxygen: (left) | Isolated oxygen valence electrons | 6 |
| | Bound oxygen bonding electrons | 2 |
| | Bound oxygen nonbonding electrons | 6 |
| | Formal charge = 6 – ½(2) – 6 = –1 | |

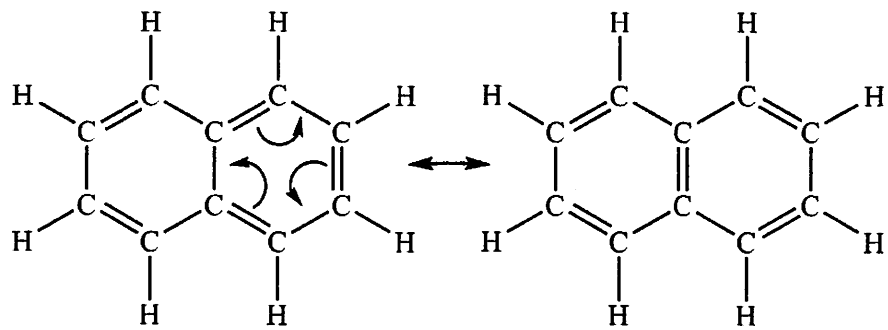
| | | |
|---|---|---|
| For oxygen: (right) | Isolated oxygen valence electrons | 6 |
| | Bound oxygen bonding electrons | 2 |
| | Bound oxygen nonbonding electrons | 6 |
| | Formal charge = 6 – ½(2) – 6 = –1 | |

| | | |
|---|---|---|
| For sulfur: | Isolated sulfur valence electrons | 6 |
| | Bound sulfur bonding electrons | 8 |
| | Bound sulfur nonbonding electrons | 0 |
| | Formal charge = 6 – ½(8) – 0 = +2 | |

7.92 (a)



(b)



7.93 Structures (a) and (b) are resonance hybrids of the same molecule.

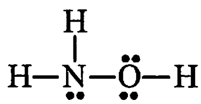
**Formal Charges (Section 7.10)**

7.94 :C≡O:

| | | |
|---|---|---|
| For carbon: | Isolated carbon valence electrons | 4 |
| | Bound carbon bonding electrons | 6 |
| | Bound carbon nonbonding electrons | 2 |
| | Formal charge = 4 – ½(6) – 2 = –1 | |

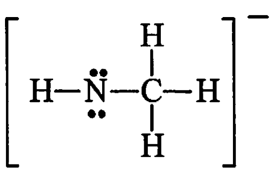
| | | |
|---|---|---|
| For oxygen: | Isolated oxygen valence electrons | 6 |
| | Bound oxygen bonding electrons | 6 |
| | Bound oxygen nonbonding electrons | 2 |
| | Formal charge = 6 – ½(6) – 2 = +1 | |

7.95 (a)

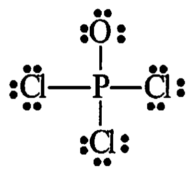


| | | |
|---|---|---|
| For hydrogen: | Isolated hydrogen valence electrons | 1 |
| | Bound hydrogen bonding electrons | 2 |
| | Bound hydrogen nonbonding electrons | 0 |
| | Formal charge = 1 – ½(2) – 0 = 0 | |
| For nitrogen: | Isolated nitrogen valence electrons | 5 |
| | Bound nitrogen bonding electrons | 6 |
| | Bound nitrogen nonbonding electrons | 2 |
| | Formal charge = 5 – ½(6) – 2 = 0 | |
| For oxygen: | Isolated oxygen valence electrons | 6 |
| | Bound oxygen bonding electrons | 4 |
| | Bound oxygen nonbonding electrons | 4 |
| | Formal charge = 6 – ½(4) – 4 = 0 | |

(b)



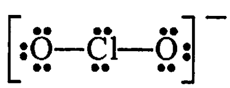
| | | |
|---|---|---|
| For hydrogen: | Isolated hydrogen valence electrons | 1 |
| | Bound hydrogen bonding electrons | 2 |
| | Bound hydrogen nonbonding electrons | 0 |
| | Formal charge = 1 – ½(2) – 0 = 0 | |
| For nitrogen: | Isolated nitrogen valence electrons | 5 |
| | Bound nitrogen bonding electrons | 4 |
| | Bound nitrogen nonbonding electrons | 4 |
| | Formal charge = 5 – ½(4) – 4 = –1 | |
| For carbon: | Isolated carbon valence electrons | 4 |
| | Bound carbon bonding electrons | 8 |
| | Bound carbon nonbonding electrons | 0 |
| | Formal charge = 4 – ½(8) – 0 = 0 | |

(c) 

For chlorine:
Isolated chlorine valence electrons 7
Bound chlorine bonding electrons 2
Bound chlorine nonbonding electrons 6
Formal charge = 7 – ½(2) – 6 = 0

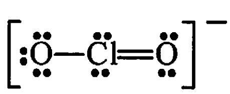
For oxygen:
Isolated oxygen valence electrons 6
Bound oxygen bonding electrons 2
Bound oxygen nonbonding electrons 6
Formal charge = 6 – ½(2) – 6 = –1

For phosphorus:
Isolated phosphorus valence electrons 5
Bound phosphorus bonding electrons 8
Bound phosphorus nonbonding electrons 0
Formal charge = 5 – ½(8) – 0 = +1

7.96 

For both oxygens:
Isolated oxygen valence electrons 6
Bound oxygen bonding electrons 2
Bound oxygen nonbonding electrons 6
Formal charge = 6 – ½(2) – 6 = –1

For chlorine:
Isolated chlorine valence electrons 7
Bound chlorine bonding electrons 4
Bound chlorine nonbonding electrons 4
Formal charge = 7 – ½(4) – 4 = +1

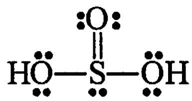


For left oxygen:
Isolated oxygen valence electrons 6
Bound oxygen bonding electrons 2
Bound oxygen nonbonding electrons 6
Formal charge = 6 – ½(2) – 6 = –1

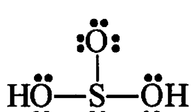
For right oxygen:
Isolated oxygen valence electrons 6
Bound oxygen bonding electrons 4
Bound oxygen nonbonding electrons 4
Formal charge = 6 – ½(4) – 4 = 0

| | | |
|---|---|---|
| For chlorine: | Isolated chlorine valence electrons | 7 |
| | Bound chlorine bonding electrons | 6 |
| | Bound chlorine nonbonding electrons | 4 |
| | Formal charge = 7 – ½(6) – 4 = 0 | |

7.97



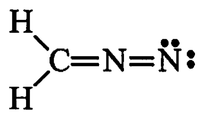
| | | |
|---|---|---|
| For sulfur: | Isolated sulfur valence electrons | 6 |
| | Bound sulfur bonding electrons | 8 |
| | Bound sulfur nonbonding electrons | 2 |
| | Formal charge = 6 – ½(8) – 2 = 0 | |
| For doubly bound oxygen: | Isolated oxygen valence electrons | 6 |
| | Bound oxygen bonding electrons | 4 |
| | Bound oxygen nonbonding electrons | 4 |
| | Formal charge = 6 – ½(4) – 4 = 0 | |
| For oxygen bound to hydrogen: | Isolated oxygen valence electrons | 6 |
| | Bound oxygen bonding electrons | 4 |
| | Bound oxygen nonbonding electrons | 4 |
| | Formal charge = 6 – ½(4) – 4 = 0 | |
| For hydrogen: | Isolated hydrogen valence electrons | 1 |
| | Bound hydrogen bonding electrons | 2 |
| | Bound hydrogen nonbonding electrons | 0 |
| | Formal charge = 1 – ½(2) – 0 = 0 | |



| | | |
|---|---|---|
| For sulfur: | Isolated sulfur valence electrons | 6 |
| | Bound sulfur bonding electrons | 6 |
| | Bound sulfur nonbonding electrons | 2 |
| | Formal charge = 6 – ½(6) – 2 = +1 | |
| For oxygen not bound to hydrogen: | Isolated oxygen valence electrons | 6 |
| | Bound oxygen bonding electrons | 2 |
| | Bound oxygen nonbonding electrons | 6 |
| | Formal charge = 6 – ½(2) – 6 = – 1 | |
| For oxygen bound to hydrogen: | Isolated oxygen valence electrons | 6 |
| | Bound oxygen bonding electrons | 4 |
| | Bound oxygen nonbonding electrons | 4 |
| | Formal charge = 6 – ½(4) – 4 = 0 | |

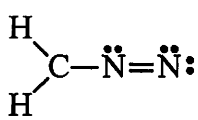
| | | |
|---|---|---|
| For hydrogen: | Isolated hydrogen valence electrons | 1 |
| | Bound hydrogen bonding electrons | 2 |
| | Bound hydrogen nonbonding electrons | 0 |
| | Formal charge = 1 – ½(2) – 0 = 0 | |

7.98 (a)



| | | |
|---|---|---|
| For hydrogen: | Isolated hydrogen valence electrons | 1 |
| | Bound hydrogen bonding electrons | 2 |
| | Bound hydrogen nonbonding electrons | 0 |
| | Formal charge = 1 – ½(2) – 0 = 0 | |
| For nitrogen: (central) | Isolated nitrogen valence electrons | 5 |
| | Bound nitrogen bonding electrons | 8 |
| | Bound nitrogen nonbonding electrons | 0 |
| | Formal charge = 5 – ½(8) – 0 = +1 | |
| For nitrogen: (terminal) | Isolated nitrogen valence electrons | 5 |
| | Bound nitrogen bonding electrons | 4 |
| | Bound nitrogen nonbonding electrons | 4 |
| | Formal charge = 5 – ½(4) – 4 = –1 | |
| For carbon: | Isolated carbon valence electrons | 4 |
| | Bound carbon bonding electrons | 8 |
| | Bound carbon nonbonding electrons | 0 |
| | Formal charge = 4 – ½(8) – 0 = 0 | |

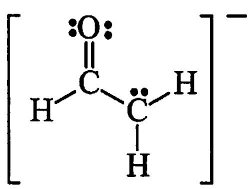
(b)



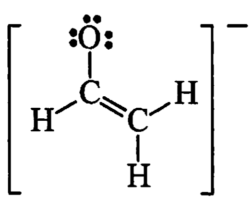
| | | |
|---|---|---|
| For hydrogen: | Isolated hydrogen valence electrons | 1 |
| | Bound hydrogen bonding electrons | 2 |
| | Bound hydrogen nonbonding electrons | 0 |
| | Formal charge = 1 – ½(2) – 0 = 0 | |
| For nitrogen: (central) | Isolated nitrogen valence electrons | 5 |
| | Bound nitrogen bonding electrons | 6 |
| | Bound nitrogen nonbonding electrons | 2 |
| | Formal charge = 5 – ½(6) – 2 = 0 | |
| For nitrogen: (terminal) | Isolated nitrogen valence electrons | 5 |
| | Bound nitrogen bonding electrons | 4 |
| | Bound nitrogen nonbonding electrons | 4 |
| | Formal charge = 5 – ½(4) – 4 = –1 | |

| | | |
|---|---|---|
| For carbon: | Isolated carbon valence electrons | 4 |
| | Bound carbon bonding electrons | 6 |
| | Bound carbon nonbonding electrons | 0 |
| | Formal charge = 4 – ½(6) – 0 = +1 | |

Structure (a) is more important because of the octet of electrons around carbon.

7.99 

| | | |
|---|---|---|
| For oxygen: | Isolated oxygen valence electrons | 6 |
| | Bound oxygen bonding electrons | 4 |
| | Bound oxygen nonbonding electrons | 4 |
| | Formal charge = 6 – ½(4) – 4 = 0 | |
| For left carbon: | Isolated carbon valence electrons | 4 |
| | Bound carbon bonding electrons | 8 |
| | Bound carbon nonbonding electrons | 0 |
| | Formal charge = 4 – ½(8) – 0 = 0 | |
| For right carbon: | Isolated carbon valence electrons | 4 |
| | Bound carbon bonding electrons | 6 |
| | Bound carbon nonbonding electrons | 2 |
| | Formal charge = 4 – ½(6) – 2 = –1 | |

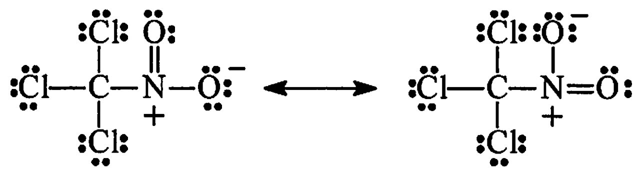


| | | |
|---|---|---|
| For oxygen: | Isolated oxygen valence electrons | 6 |
| | Bound oxygen bonding electrons | 2 |
| | Bound oxygen nonbonding electrons | 6 |
| | Formal charge = 6 – ½(2) – 6 = –1 | |
| For left carbon: | Isolated carbon valence electrons | 4 |
| | Bound carbon bonding electrons | 8 |
| | Bound carbon nonbonding electrons | 0 |
| | Formal charge = 4 – ½(8) – 0 = 0 | |
| For right carbon: | Isolated carbon valence electrons | 4 |
| | Bound carbon bonding electrons | 8 |
| | Bound carbon nonbonding electrons | 0 |
| | Formal charge = 4 – ½(8) – 0 = 0 | |

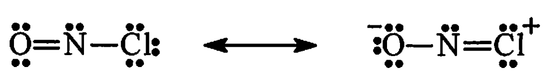
The second structure is more important because of the −1 formal charge on the more electronegative oxygen.

7.100



| | | |
|---|---|---|
| For chlorine: | Isolated chlorine valence electrons | 7 |
| | Bound chlorine bonding electrons | 2 |
| | Bound chlorine nonbonding electrons | 6 |
| | Formal charge = 7 – ½(2) – 6 = 0 | |
| For carbon: | Isolated carbon valence electrons | 4 |
| | Bound carbon bonding electrons | 8 |
| | Bound carbon nonbonding electrons | 0 |
| | Formal charge = 4 – ½(8) – 0 = 0 | |
| For nitrogen: | Isolated nitrogen valence electrons | 5 |
| | Bound nitrogen bonding electrons | 8 |
| | Bound nitrogen nonbonding electrons | 0 |
| | Formal charge = 5 – ½(8) – 0 = +1 | |
| For oxygen: (double bonded) | Isolated oxygen valence electrons | 6 |
| | Bound oxygen bonding electrons | 4 |
| | Bound oxygen nonbonding electrons | 4 |
| | Formal charge = 6 – ½(4) – 4 = 0 | |
| For oxygen: (single bonded) | Isolated oxygen valence electrons | 6 |
| | Bound oxygen bonding electrons | 2 |
| | Bound oxygen nonbonding electrons | 6 |
| | Formal charge = 6 – ½(2) – 6 = –1 | |

7.101



| | | |
|---|---|---|
| For chlorine: (single bonded) | Isolated chlorine valence electrons | 7 |
| | Bound chlorine bonding electrons | 2 |
| | Bound chlorine nonbonding electrons | 6 |
| | Formal charge = 7 – ½(2) – 6 = 0 | |
| For chlorine: (double bonded) | Isolated chlorine valence electrons | 7 |
| | Bound chlorine bonding electrons | 4 |
| | Bound chlorine nonbonding electrons | 4 |
| | Formal charge = 7 – ½(4) – 4 = +1 | |

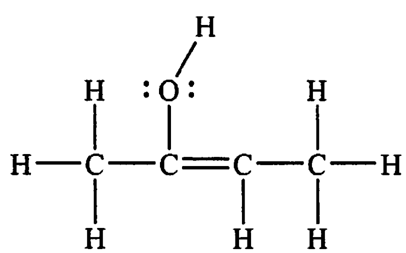
| | | |
|---|---|---|
| For nitrogen: | Isolated nitrogen valence electrons | 5 |
| | Bound nitrogen bonding electrons | 6 |
| | Bound nitrogen nonbonding electrons | 2 |
| | Formal charge = 5 – ½(6) – 2 = 0 | |

| | | |
|---|---|---|
| For oxygen: (double bonded) | Isolated oxygen valence electrons | 6 |
| | Bound oxygen bonding electrons | 4 |
| | Bound oxygen nonbonding electrons | 4 |
| | Formal charge = 6 – ½(4) – 4 = 0 | |

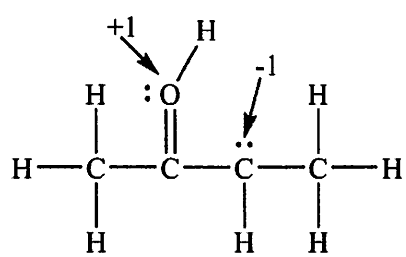
| | | |
|---|---|---|
| For oxygen: (single bonded) | Isolated oxygen valence electrons | 6 |
| | Bound oxygen bonding electrons | 2 |
| | Bound oxygen nonbonding electrons | 6 |
| | Formal charge = 6 – ½(2) – 6 = –1 | |

:Ö=N̈—C̈l: is the larger contributor to the resonance hybrid because it has no formal charges.

7.102



All atoms have 0 formal charge.

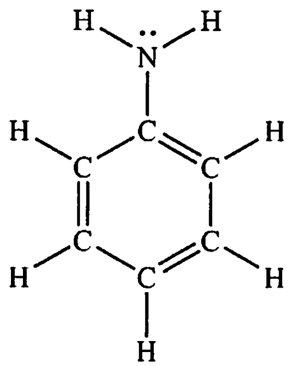



| | | |
|---|---|---|
| For oxygen: | Isolated oxygen valence electrons | 6 |
| | Bound oxygen bonding electrons | 6 |
| | Bound oxygen nonbonding electrons | 2 |
| | Formal charge = 6 – ½(6) – 2 = +1 | |

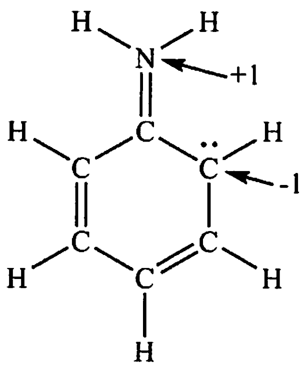
| | | |
|---|---|---|
| For carbon: | Isolated carbon valence electrons | 4 |
| | Bound carbon bonding electrons | 6 |
| | Bound carbon nonbonding electrons | 2 |
| | Formal charge = 4 – ½(6) – 2 = –1 | |

All other atoms have 0 formal charge.
The original structure is the larger contributor.

7.103



All atoms have 0 formal charge.

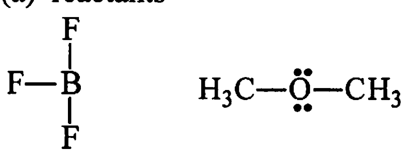


| | | |
|---|---|---|
| For nitrogen: | Isolated nitrogen valence electrons | 5 |
| | Bound nitrogen bonding electrons | 8 |
| | Bound nitrogen nonbonding electrons | 0 |
| | Formal charge = 5 – ½(8) – 0 = +1 | |

| | | |
|---|---|---|
| For carbon: | Isolated carbon valence electrons | 4 |
| | Bound carbon bonding electrons | 6 |
| | Bound carbon nonbonding electrons | 2 |
| | Formal charge = 4 – ½(6) – 2 = –1 | |

All other atoms have 0 formal charge.
The original structure is the larger contributor.

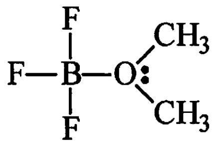
7.104 (a) reactants



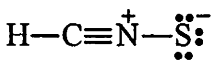
| | | |
|---|---|---|
| For boron: | Isolated boron valence electrons | 3 |
| | Bound boron bonding electrons | 6 |
| | Bound boron nonbonding electrons | 0 |
| | Formal charge = 3 – ½(6) – 0 = 0 | |

| | | |
|---|---|---|
| For oxygen: | Isolated oxygen valence electrons | 6 |
| | Bound oxygen bonding electrons | 4 |
| | Bound oxygen nonbonding electrons | 4 |
| | Formal charge $= 6 - ½(4) - 4 = 0$ | |

product



| | | |
|---|---|---|
| For boron: | Isolated boron valence electrons | 3 |
| | Bound boron bonding electrons | 8 |
| | Bound boron nonbonding electrons | 0 |
| | Formal charge $= 3 - ½(8) - 0 = -1$ | |
| For oxygen: | Isolated oxygen valence electrons | 6 |
| | Bound oxygen bonding electrons | 6 |
| | Bound oxygen nonbonding electrons | 2 |
| | Formal charge $= 6 - ½(6) - 2 = +1$ | |

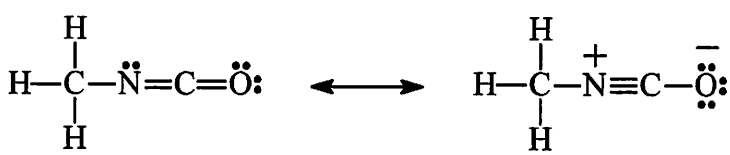
7.105 (a) 

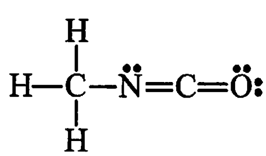
| | | |
|---|---|---|
| For hydrogen: | Isolated hydrogen valence electrons | 1 |
| | Bound hydrogen bonding electrons | 2 |
| | Bound hydrogen nonbonding electrons | 0 |
| | Formal charge $= 1 - ½(2) - 0 = 0$ | |
| For carbon: | Isolated carbon valence electrons | 4 |
| | Bound carbon bonding electrons | 8 |
| | Bound carbon nonbonding electrons | 0 |
| | Formal charge $= 4 - ½(8) - 0 = 0$ | |
| For nitrogen: | Isolated nitrogen valence electrons | 5 |
| | Bound nitrogen bonding electrons | 8 |
| | Bound nitrogen nonbonding electrons | 0 |
| | Formal charge $= 5 - ½(8) - 0 = +1$ | |
| For sulfur: | Isolated sulfur valence electrons | 6 |
| | Bound sulfur bonding electrons | 2 |
| | Bound sulfur nonbonding electrons | 6 |
| | Formal charge $= 6 - ½(2) - 6 = -1$ | |

(b) H—N̈=C=S̈:

| | | |
|---|---|---|
| For hydrogen: | Isolated hydrogen valence electrons | 1 |
| | Bound hydrogen bonding electrons | 2 |
| | Bound hydrogen nonbonding electrons | 0 |
| | Formal charge = 1 – ½(2) – 0 = 0 | |
| For nitrogen: | Isolated nitrogen valence electrons | 5 |
| | Bound nitrogen bonding electrons | 6 |
| | Bound nitrogen nonbonding electrons | 2 |
| | Formal charge = 5 – ½(6) – 2 = 0 | |
| For carbon: | Isolated carbon valence electrons | 4 |
| | Bound carbon bonding electrons | 8 |
| | Bound carbon nonbonding electrons | 0 |
| | Formal charge = 4 – ½(8) – 0 = 0 | |
| For sulfur: | Isolated sulfur valence electrons | 6 |
| | Bound sulfur bonding electrons | 4 |
| | Bound sulfur nonbonding electrons | 4 |
| | Formal charge = 6 – ½(4) – 4 = 0 | |

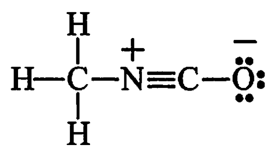
(c) H—N̈=C=S̈: is more stable because it has no formal charges.

7.106


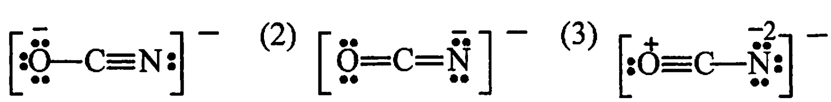
For


| | | |
|---|---|---|
| For hydrogen: | Isolated hydrogen valence electrons | 1 |
| | Bound hydrogen bonding electrons | 2 |
| | Bound hydrogen nonbonding electrons | 0 |
| | Formal charge = 1 – ½(2) – 0 = 0 | |
| For carbon: (left) | Isolated carbon valence electrons | 4 |
| | Bound carbon bonding electrons | 8 |
| | Bound carbon nonbonding electrons | 0 |
| | Formal charge = 4 – ½(8) – 0 = 0 | |
| For nitrogen: | Isolated nitrogen valence electrons | 5 |
| | Bound nitrogen bonding electrons | 6 |
| | Bound nitrogen nonbonding electrons | 2 |
| | Formal charge = 5 – ½(6) – 2 = 0 | |

| | | |
|---|---|---|
| For carbon: (right) | Isolated carbon valence electrons | 4 |
| | Bound carbon bonding electrons | 8 |
| | Bound carbon nonbonding electrons | 0 |
| | Formal charge = 4 – ½(8) – 0 = 0 | |
| For oxygen: | Isolated oxygen valence electrons | 6 |
| | Bound oxygen bonding electrons | 4 |
| | Bound oxygen nonbonding electrons | 4 |
| | Formal charge = 6 – ½(4) – 4 = 0 | |

For 

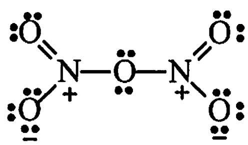
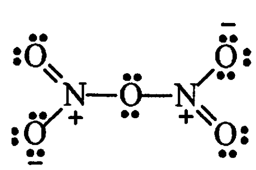
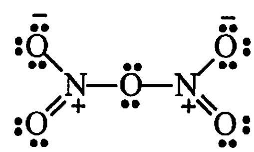
| | | |
|---|---|---|
| For hydrogen: | Isolated hydrogen valence electrons | 1 |
| | Bound hydrogen bonding electrons | 2 |
| | Bound hydrogen nonbonding electrons | 0 |
| | Formal charge = 1 – ½(2) – 0 = 0 | |
| For carbon: (left) | Isolated carbon valence electrons | 4 |
| | Bound carbon bonding electrons | 8 |
| | Bound carbon nonbonding electrons | 0 |
| | Formal charge = 4 – ½(8) – 0 = 0 | |
| For nitrogen: | Isolated nitrogen valence electrons | 5 |
| | Bound nitrogen bonding electrons | 8 |
| | Bound nitrogen nonbonding electrons | 0 |
| | Formal charge = 5 – ½(8) – 0 = +1 | |
| For carbon: (right) | Isolated carbon valence electrons | 4 |
| | Bound carbon bonding electrons | 8 |
| | Bound carbon nonbonding electrons | 0 |
| | Formal charge = 4 – ½(8) – 0 = 0 | |
| For oxygen: | Isolated oxygen valence electrons | 6 |
| | Bound oxygen bonding electrons | 2 |
| | Bound oxygen nonbonding electrons | 6 |
| | Formal charge = 6 – ½(2) – 6 = –1 | |

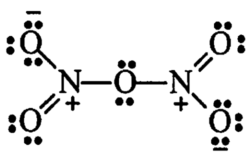
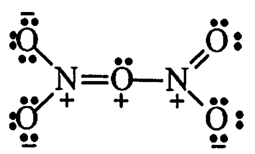
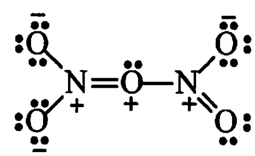
7.107 (a) (1) 

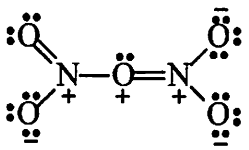
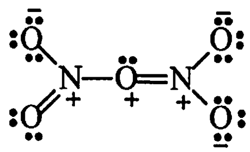
(b) Structure (1) makes the greatest contribution to the resonance hybrid because of the –1 formal charge on the oxygen. Structure (3) makes the least contribution to the resonance hybrid because of the +1 formal charge on the oxygen.

## Multiconcept Problems

7.108 (a)  (b)  (c) 

(d)  (e)  (f) 

(g)  (h) 

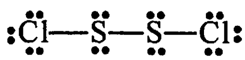
Structures (a) – (d) make more important contributions to the resonance hybrid because of only −1 and 0 formal charges on the oxygens. A +1 formal charge is unlikely.

7.109 (a) Assume a 100.0 g sample. From the percent composition data, a 100.0 g sample contains 47.5 g S and 52.5 g Cl.

$$47.5 \text{ g S} \times \frac{1 \text{ mol S}}{32.065 \text{ g S}} = 1.48 \text{ mol S}$$

$$52.5 \text{ g Cl} \times \frac{1 \text{ mol Cl}}{35.453 \text{ g Cl}} = 1.48 \text{ mol Cl}$$

Because the two mole quantities are the same, the empirical formula is SCl.



| | | |
|---|---|---|
| For sulfur: | Isolated sulfur valence electrons | 6 |
| | Bound sulfur bonding electrons | 4 |
| | Bound sulfur nonbonding electrons | 4 |
| | Formal charge = 6 – ½(4) – 4 = 0 | |
| For chlorine: | Isolated chlorine valence electrons | 7 |
| | Bound chlorine bonding electrons | 2 |
| | Bound chlorine nonbonding electrons | 6 |
| | Formal charge = 7 – ½(2) – 6 = 0 | |

7.110 (a) Assume a 100.0 g sample. From the percent composition data, a 100.0 g sample contains 47.5 g S and 52.5 g Cl.

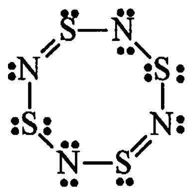
$$69.6 \text{ g S} \times \frac{1 \text{ mol S}}{32.065 \text{ g S}} = 2.17 \text{ mol S}; \qquad 30.4 \text{ g N} \times \frac{1 \text{ mol N}}{14.007 \text{ g N}} = 2.17 \text{ mol N}$$

Because the two mole quantities are the same, the empirical formula is SN.
The empirical formula weight = 47.07

$$\text{Multiple} = \frac{\text{molecular weight}}{\text{empirical formula weight}} = \frac{184.3}{47.07} = 4.00$$

The molecular formula is $S_4N_4$.

(b)

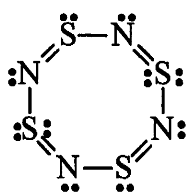


| | | |
|---|---|---|
| For sulfur: (1 lone pair) | Isolated sulfur valence electrons | 6 |
| | Bound sulfur bonding electrons | 6 |
| | Bound sulfur nonbonding electrons | 2 |
| | Formal charge = 6 – ½(6) – 2 = +1 | |

| | | |
|---|---|---|
| For sulfur: (2 lone pairs) | Isolated sulfur valence electrons | 6 |
| | Bound sulfur bonding electrons | 4 |
| | Bound sulfur nonbonding electrons | 4 |
| | Formal charge = 6 – ½(4) – 4 = 0 | |

| | | |
|---|---|---|
| For nitrogen: (1 lone pair) | Isolated nitrogen valence electrons | 5 |
| | Bound nitrogen bonding electrons | 6 |
| | Bound nitrogen nonbonding electrons | 2 |
| | Formal charge = 5 – ½(6) – 2 = 0 | |

| | | |
|---|---|---|
| For nitrogen: (2 lone pairs) | Isolated nitrogen valence electrons | 5 |
| | Bound nitrogen bonding electrons | 4 |
| | Bound nitrogen nonbonding electrons | 4 |
| | Formal charge = 5 – ½(4) – 4 = –1 | |



| | | |
|---|---|---|
| For sulfur: (1 lone pair) | Isolated sulfur valence electrons | 6 |
| | Bound sulfur bonding electrons | 6 |
| | Bound sulfur nonbonding electrons | 2 |
| | Formal charge = 6 – ½(6) – 2 = +1 | |

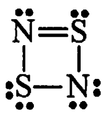
| | | |
|---|---|---|
| For sulfur: (2 lone pairs) | Isolated sulfur valence electrons | 6 |
| | Bound sulfur bonding electrons | 6 |
| | Bound sulfur nonbonding electrons | 4 |
| | Formal charge = 6 – ½(6) – 4 = –1 | |

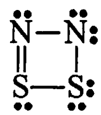
 

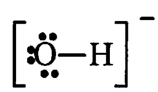
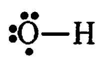
| | | |
|---|---|---|
| For nitrogen: | Isolated nitrogen valence electrons | 5 |
| | Bound nitrogen bonding electrons | 6 |
| | Bound nitrogen nonbonding electrons | 2 |
| | Formal charge = 5 – ½(6) – 2 = 0 | |

(c) $\text{Multiple} = \dfrac{\text{molecular weight}}{\text{empirical formula weight}} = \dfrac{92.2}{47.07} = 1.96 = 2$

The molecular formula is $S_2N_2$. Two possible structures are shown here.





7.111 (a)  

(b) The oxygen in OH has a half-filled 2p orbital that can accept the additional electron. For a 2p orbital, n = 2 and $l$ = 1.

(c) The electron affinity for OH is slightly more negative than for an O atom because when OH gains an additional electron, it achieves an octet configuration.

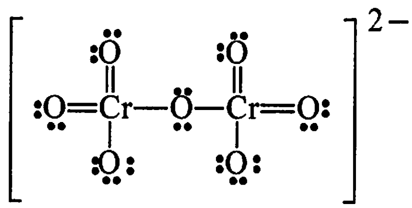
7.112 (a) (4 orbitals)(3 electrons) = 12 outer-shell electrons

(b) 3 electrons

(c) $1s^3\,2s^3\,2p^6$; 

(d) 

7.113 (a)



All formal charges are zero except for the two single bonded oxygens which each have a formal charge of –1.

(b) Each Cr atom has 6 pairs of electrons around it.