

Manual of Possible Solutions

Possible Solutions for Chapter 1 Problems

Problem 1.1

List the following orbitals in order of their relative size: 2s, 2p_y, 1s, 2p_z, 3s

Problem 1.1 Solution

$$1s < 2s < 2p_y \sim 2p_z < 3s$$

Problem 1.2

With the aid of your periodic table, give the electronic configuration for each of the following atoms: F, B, and Na?

Problem 1.2 Solution

Fluorine (⁹F): 1s²2s²2p_x²2p_y²2p_z¹ 3s⁰

Boron (⁵B): 1s²2s²2p_x¹2p_y⁰2p_z⁰ 3s⁰

Sodium (¹¹Na): 1s²2s²2p_x²2p_y²2p_z² 3s¹

Problem 1.3

With the aid of your periodic table, give the electronic configuration and Lewis dot structures for the following atoms:

Li, Be, F, Ne, Na, and Mg.

Problem 1.3 Solution

Orbitals showing the valence electrons are highlighted in blue

Lithium (³Li): 1s²2s¹2p_x⁰2p_y⁰2p_z⁰ 3s⁰ Li •

Beryllium (⁴Be): 1s²2s²2p_x⁰2p_y⁰2p_z⁰ 3s⁰ • Be •

Fluorine (⁹F): 1s²2s²2p_x²2p_y²2p_z¹ 3s⁰ : $\overset{\cdot}{\underset{\cdot}{\text{F}}}$:

Neon (¹⁰Ne): 1s²2s²2p_x²2p_y²2p_z² 3s⁰ : $\overset{\cdot\cdot}{\underset{\cdot\cdot}{\text{Ne}}}$:

Magnesium (¹²Mg): 1s²2s²2p_x²2p_y²2p_z² 3s² • Mg •

Problem 1.4

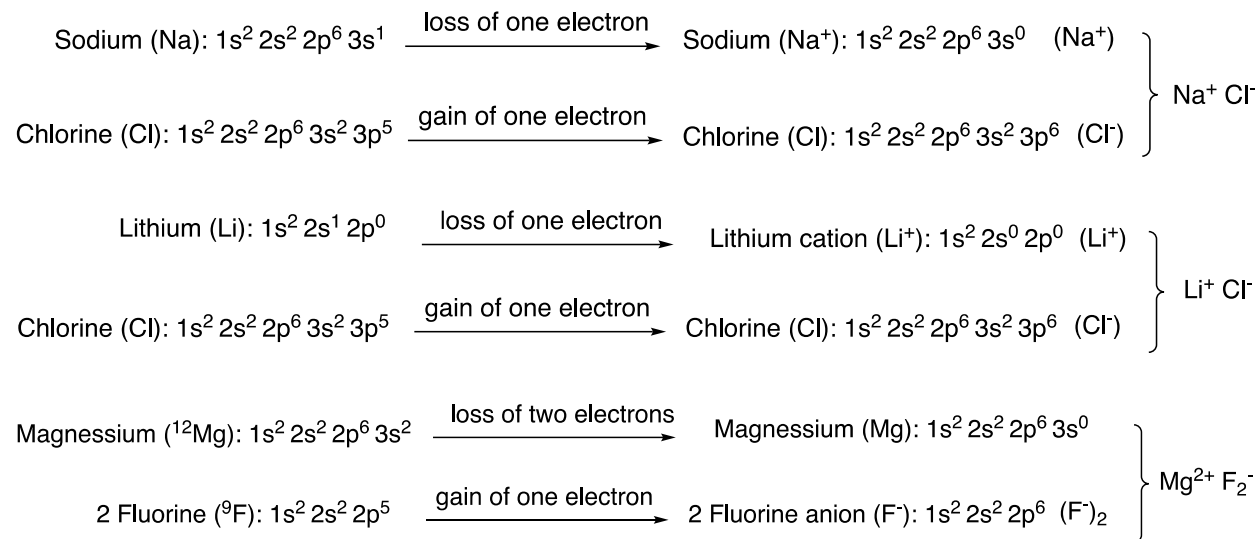
Using a formulism as that shown in Figure 1.8, illustrate how ionic bonds are formed between the following pairs of atoms.

(a) Na and Cl

(b) Li and Cl

(c) Mg and F

Problem 1.4 Solution

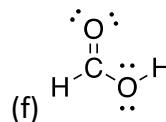
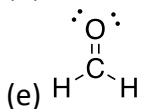
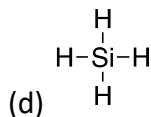
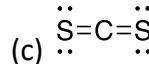
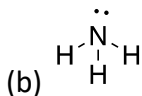
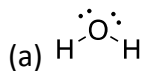


Problem 1.5

Draw the Lewis dot structure for the following molecules and clearly show all nonbonding electrons.

(a) H_2O , (b) NH_3 , (c) CS_2 , (d) SiH_4 , (e) CH_2O , (f) CH_2O_2

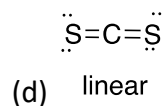
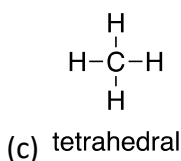
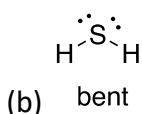
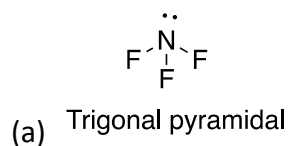
Problem 1.5 Solution

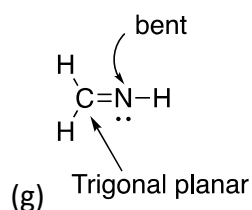
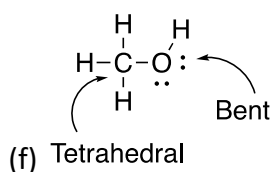
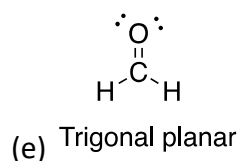


Problem 1.6

Give the Lewis dot structure for the following molecules and predict the geometry about each atom of your structure that is bonded to at least two other atoms: (a) NF_3 , (b) H_2S , (c) CH_4 , (d) CS_2 , (e) CH_2O , (f) CH_3OH , (g) CH_3N .

Problem 1.6 Solution





Problem 1.7

Of the following pairs of atoms, look at the periodic table and determine which is more electronegative.

(a) K and Br; (b) Cl and Br; (c) N and C; (d) Mg and C.

Problem 1.7 Solution

(a) Br; (b) Cl; (c) N; (d) C

Problem 1.8

Classify the following covalent bonds as polar or nonpolar.

(a) O-H; (b) H-H; (c) HN=NH; (d) C-Mg; (e) C-O; (f) C-Cl

Problem 1.8 Solution

(a) O-H (polar); (b) H-H (nonpolar); (c) HN=NH (nonpolar); (d) C-Mg (polar); (e) C-O (polar); (f) C-Cl (polar)

Problem 1.9

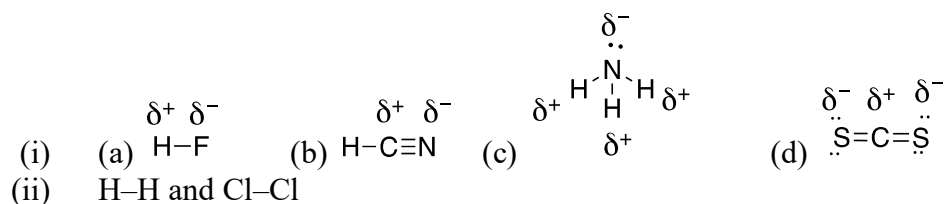
(i) Give the Lewis dot structures of the molecules shown below and use the δ^+ and δ^- representations to indicate the polarity of the covalent bonds.

(a) HF; (b) HCN; (c) NH₃; (d) CS₂.

(ii) Which of the following molecules have nonpolar covalent bonds?

H₂, Cl₂, CH₃Cl, CO₂, COS, and H₂O.

Problem 1.9 Solution

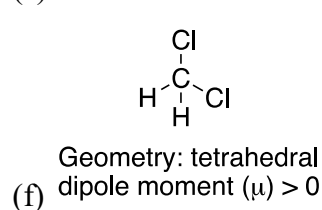
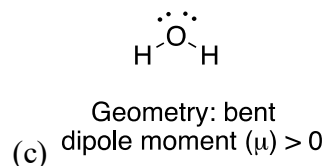
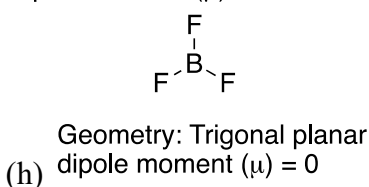
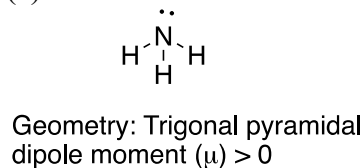
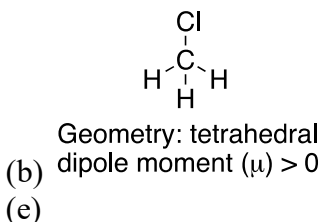
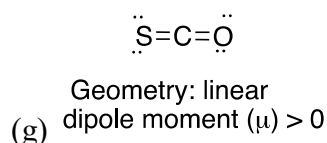
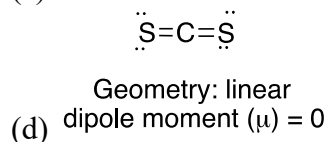
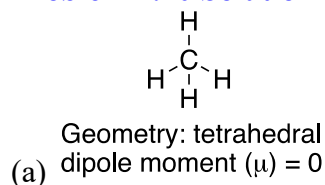


Problem 1.10

Draw the Lewis dot structures of each of the following molecules; determine the geometries and which have dipole moments of zero or nonzero.

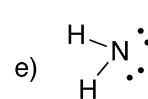
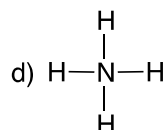
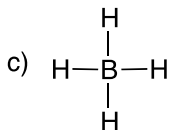
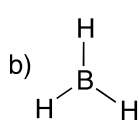
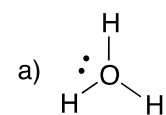
(a) CH₄; (b) CHCl₃; (c) H₂O; (d) CS₂; (e) NH₃; (f) CH₂Cl₂; (g) COS; (h) BF₃

Problem 1.10 Solution



Problem 1.11

Utilize one of the equations above to calculate the formal charge for the central atom of each of the following species.



Problem 1.11 Solution

(a) Formal charge (oxygen) = $6 - 2 - \frac{1}{2}(6) = +1$

(b) Formal charge (boron) = $3 - 0 - \frac{1}{2}(6) = 0$

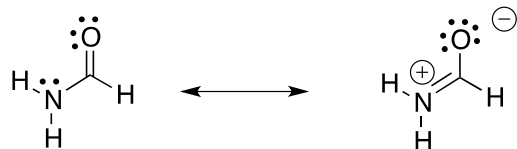
(c) Formal charge (boron) = $3 - 0 - \frac{1}{2}(8) = -1$

(d) Formal charge (nitrogen) = $5 - 0 - \frac{1}{2}(8) = +1$

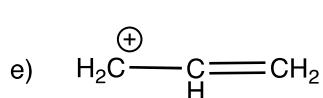
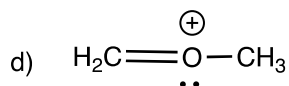
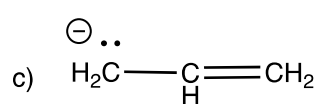
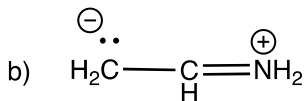
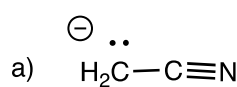
(e) Formal charge (nitrogen) = $5 - 4 - \frac{1}{2}(4) = -1$

Problem 1.12

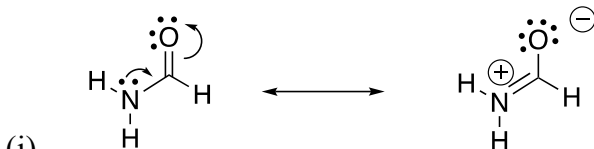
- (i) Use *curved arrow formalism* to show electron movement, show how the resonance structure on the left is transformed to the resonance structures on the right.



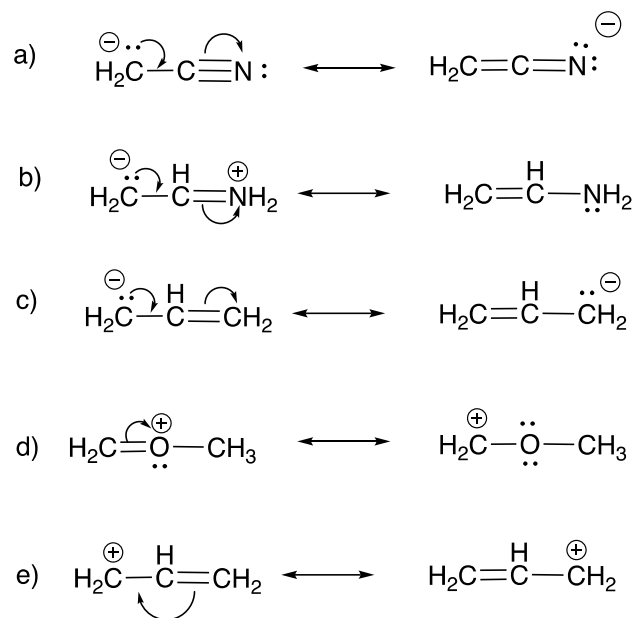
- (ii) Give another resonance structure for each of the following species. Use *curved arrow formalism* to indicate electron movement and include formal charges where appropriate.



Problem 1.12 Solution

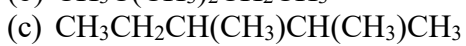


(ii)

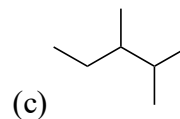
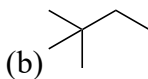
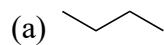


Problem 1.13

Give the structural formulas (line-angle representation) for each of the following molecules.

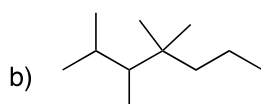
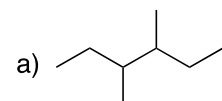


Problem 1.13 Solution

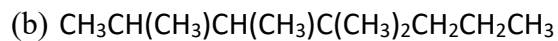
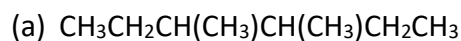


Problem 1.14

Give the condensed formulas for the following molecules.

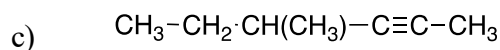
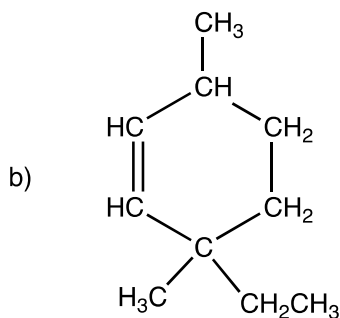
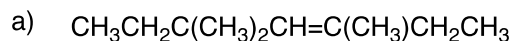


Problem 1.14 Solution

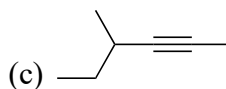
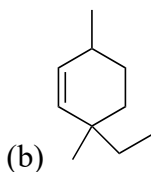
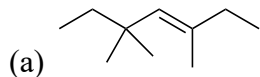


Problem 1.15

Give the structural formulas (line-angle representations) for each of the following molecules shown below.



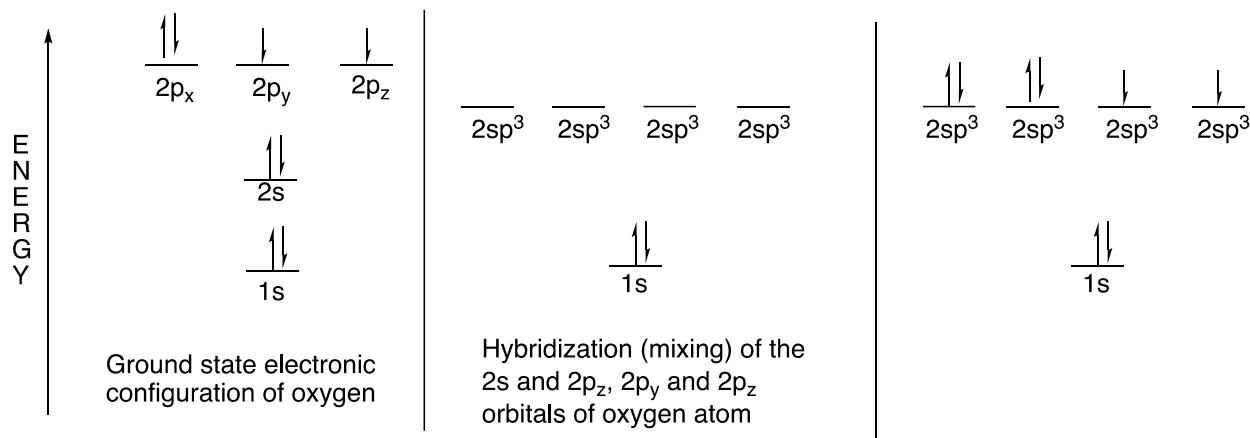
Problem 1.15 Solution



Problem 1.16

- (a) The oxygen atom of water (H_2O) also has four sp^3 hybridized orbitals, two to the hydrogen atoms and two that contain two lone pairs of nonbonding electrons. Use a figure similar to that in Figure 1.21 to represent the orbitals and electrons in water.
- (b) Explain why the observed bond angle of water is 104.5° and not 109.5° , similar to methane (CH_4) or ammonia (NH_3), which is 107° .

Problem 1.16 Solution



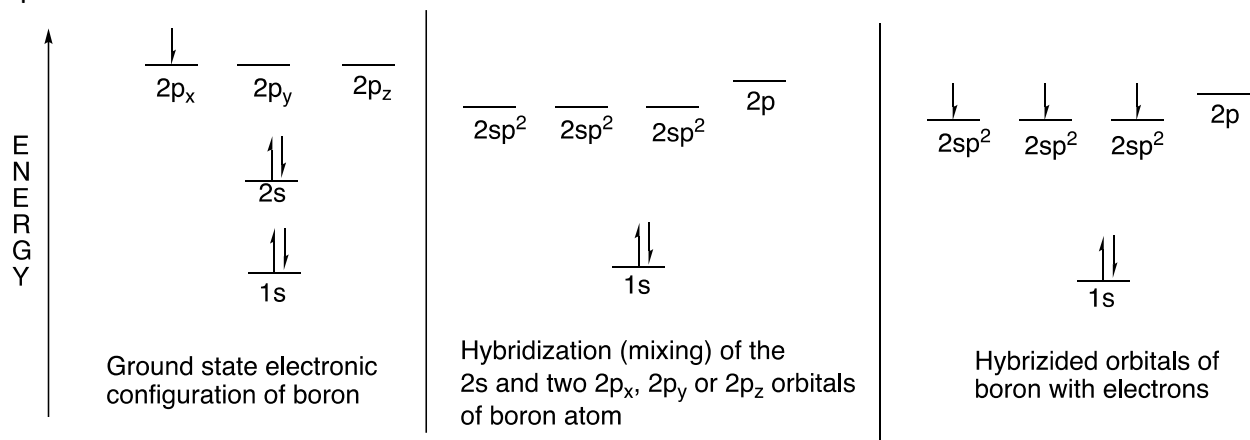
- (b) Unshared electrons occupy more space than bonding electrons. Water has two pairs of nonbonding electrons, compared to one pair of nonbonding electrons for ammonia and none for methane.

Problem 1.17

- What type of orbitals would you expect for the bonding from the central atom in BF_3 ?
- Give a diagram similar to that in Figure 1.17 for BF_3 .
- Based on the location of Al in the periodic table, predict the geometry that would exist for AlCl_3 .

Problem 1.17 Solution

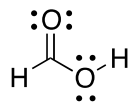
- (i) sp^2



- (iii) Since Al is in the same column as B on the periodic table, the expectation is that the geometry about AlCl_3 would be similar to that of BF_3 , trigonal planar.

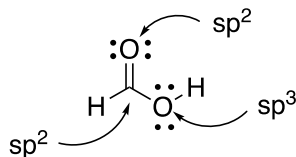
Problem 1.18

Formic acid has two oxygen atoms that use different types of orbitals for bonding. (i) For each oxygen atom, determine the type of orbitals used for bonding? (i) Determine the type of orbitals used for bonding around the carbon atom?



Formic acid

Problem 1.18 Solution



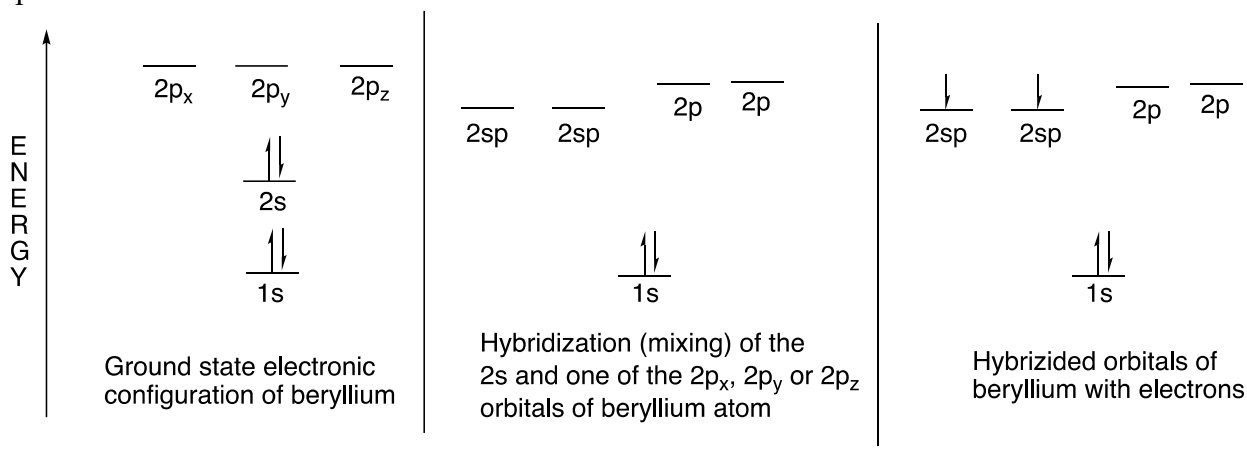
Formic acid

Problem 1.19

- What type of orbitals would you expect for the bonding in BeCl_2 ?
- Give a diagram similar to that in Figure 1.29 for this molecule.

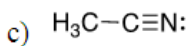
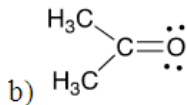
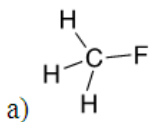
Problem 1.19 Solution

- sp



Problem 1.20

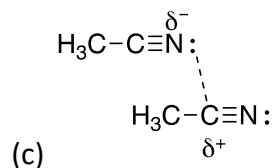
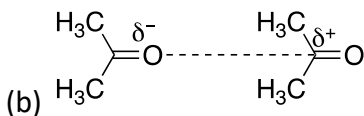
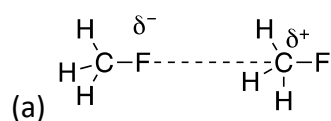
- Using an illustration as shown in Figure 1.32, show the intermolecular attractions between each of the following molecules.



- For compounds (b) and (c) above, what factors should be considered to determine which will form a stronger intermolecular attraction?

Problem 1.20 Solution

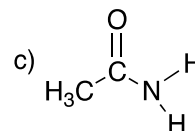
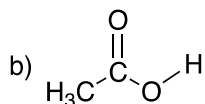
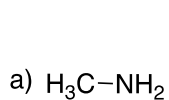
(i)



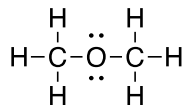
- (ii) Factors that must be considered include the electronegativity of the atoms involved, oxygen is more electronegative than nitrogen; note that nitrogen is sp hybridized and oxygen is sp^2 hybridized.

Problem 1.21

- (i) Using an illustration as shown in Figure 1.33, show the hydrogen bond intermolecular attractions between each of the following molecules.

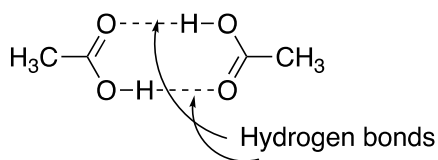
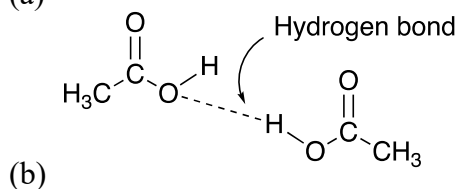
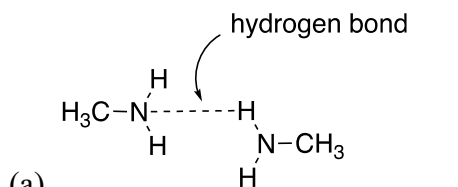


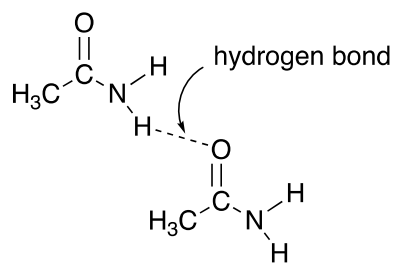
- (ii) Would you expect intermolecular hydrogen bonding to exist for the following molecule? Explain your answer.



Problem 1.21 Solution

(i)

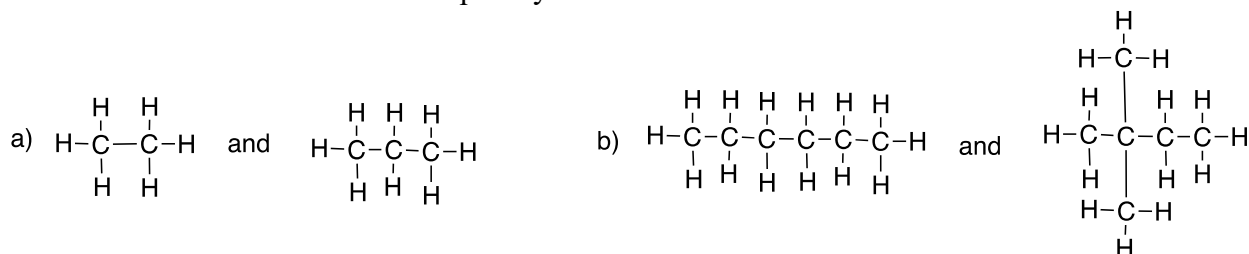




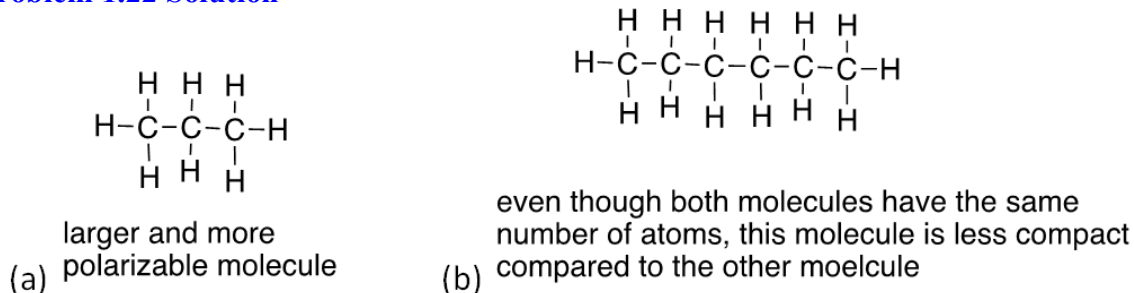
- (c)
 (ii) No, there are no hydrogen atoms bonded to an electronegative atom to create a partially positive hydrogen atom and partially negative charge on another atom.

Problem 1.22

Of the following pairs of molecules, determine which molecule would have the stronger Van der Waals intermolecular attraction. Explain your answer.



Problem 1.22 Solution



Problem 1.23

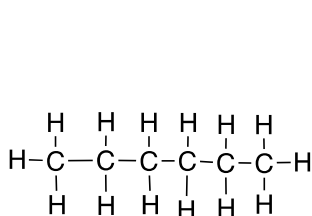
Butaneamine ($\text{CH}_3\text{CH}_2\text{CH}_2\text{NH}_2$) has approximately the same molecular weight as propanol ($\text{CH}_3\text{CH}_2\text{CH}_2\text{OH}$), yet the boiling point of butaneamine is 48°C and that of propanol is 97°C . Explain this difference of boiling points.

Problem 1.23 Solution

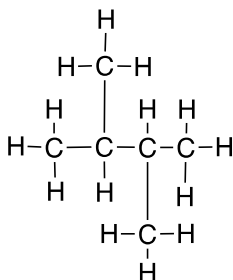
The hydrogen bonds of propanol are stronger than that of butaneamine. Since oxygen is more electronegative than nitrogen, the hydrogen bonds formed among propanol molecules are stronger than that of butaneamine. As a result, more energy is required to break the intermolecular attractions that exist with propanol resulting in a higher boiling point, compared to butaneamine.

Problem 1.24

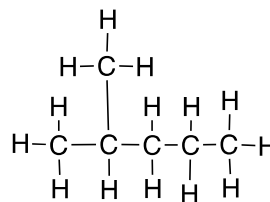
Arrange the following compounds in order of boiling points, i.e. the lowest boiling liquid first:



Hexane



2,3-Dimethylbutane



2-Methylpentane

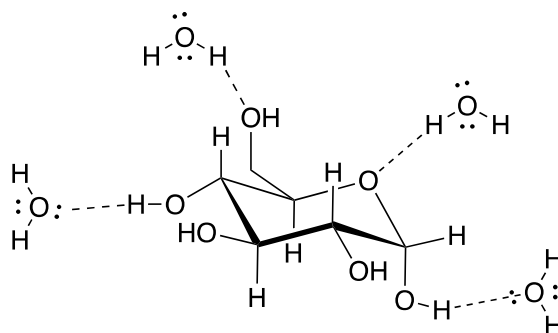
Problem 1.24 Solution

2,3-Dimethylbutane (b.p. = 58 °C) < 2-methylpentane (b.p. = 60 °C) < hexane (b.p. = 69 °C).

Problem 1.25

Draw structures to demonstrate the hydrogen bonds that are possible with water and glucose (structure shown above).

Problem 1.25 Solution



Some of the possible hydrogen bonds from water to Glucose

Problem 1.26

Write the ground state electronic configuration for each of the following atoms.

- (a) Carbon (b) Oxygen (c) Nitrogen (d) Fluorine (e) Sodium
(f) Lithium (g) Neon (h) Beryllium (i) Helium (j) Boron

Problem 1.26 Solution

- (a) Carbon (${}^6\text{C}$): $1s^2 2s^2 2p_x^1 2p_y^1 2p_z^0 3s^0$
- (b) Oxygen (${}^8\text{O}$): $1s^2 2s^2 2p_x^2 2p_y^1 2p_z^1 3s^0$
- (c) Nitrogen (${}^7\text{N}$): $1s^2 2s^2 2p_x^1 2p_y^1 2p_z^1 3s^0$
- (d) Fluorine (${}^9\text{F}$): $1s^2 2s^2 2p_x^2 2p_y^2 2p_z^1 3s^0$
- (e) Sodium (${}^{11}\text{Na}$): $1s^2 2s^2 2p_x^2 2p_y^2 2p_z^2 3s^1$
- (f) Lithium (${}^3\text{Li}$): $1s^2 2s^1 2p_x^0 2p_y^0 2p_z^0 3s^0$
- (g) Neon (${}^{10}\text{Ne}$): $1s^2 2s^2 2p_x^2 2p_y^2 2p_z^2 3s^0$
- (h) Beryllium (${}^4\text{Be}$): $1s^2 2s^2 2p_x^0 2p_y^0 2p_z^0 3s^0$
- (i) Helium (${}^2\text{He}$): $1s^2 2s^0 2p_x^0 2p_y^0 2p_z^0 3s^0$
- (j) Boron (${}^5\text{B}$): $1s^2 2s^2 2p_x^1 2p_y^0 2p_z^0 3s^0$

Problem 1.27

How many valence electrons are there in each of the following atoms?

- (a) Carbon (b) Oxygen (c) Nitrogen (d) Hydrogen (e) Fluorine
(f) Boron (g) Lithium (h) Sodium (i) Beryllium

Problem 1.27 Solution

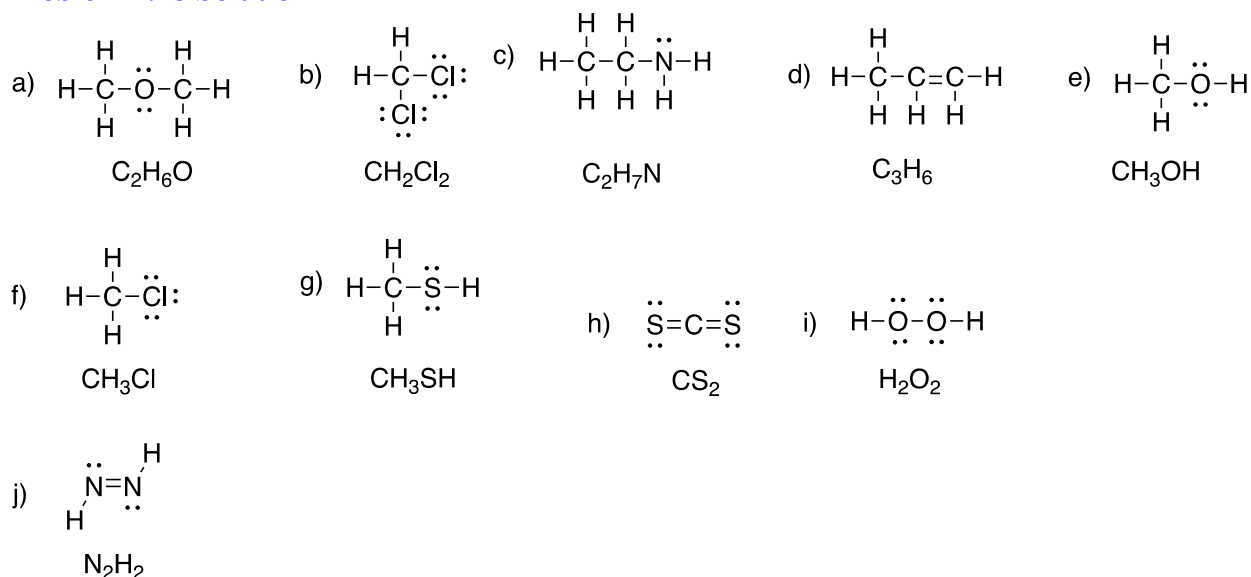
- | | |
|---|-----------------------------|
| (a) Carbon (${}^6\text{C}$): $1s^2 \mathbf{2s^2 2p_x^1 2p_y^1} 2p_z^0 3s^0$ | Four (4) valence electrons |
| (b) Oxygen (${}^8\text{O}$): $1s^2 \mathbf{2s^2 2p_x^2 2p_y^1 2p_z^1} 3s^0$ | Six (6) valence electrons |
| (c) Hydrogen (${}^1\text{H}$): $\mathbf{1s^1} 2s^0 2p_x^0 2p_y^0 2p_z^0 3s^0$ | One (1) valence electron |
| (d) Nitrogen (${}^7\text{N}$): $1s^2 \mathbf{2s^2 2p_x^1 2p_y^1 2p_z^1} 3s^0$ | Five (5) valence electrons |
| (e) Fluorine (${}^9\text{F}$): $1s^2 \mathbf{2s^2 2p_x^2 2p_y^2 2p_z^1} 3s^0$ | Seven (7) valence electrons |
| (f) Boron (${}^5\text{B}$): $1s^2 \mathbf{2s^2 2p_x^1} 2p_y^0 2p_z^0 3s^0$ | Three (3) valence electrons |
| (g) Lithium (${}^3\text{Li}$): $1s^2 \mathbf{2s^1} 2p_x^0 2p_y^0 2p_z^0 3s^0$ | One (1) valence electron |
| (h) Sodium (${}^{11}\text{Na}$): $1s^2 2s^2 2p_x^2 2p_y^2 2p_z^2 \mathbf{3s^1}$ | One (1) valence electron |
| (i) Beryllium (${}^4\text{Be}$): $1s^2 \mathbf{2s^2} 2p_x^0 2p_y^0 2p_z^0 3s^0$ | Two (2) valence electrons |

Problem 1.28

Give Lewis dot structures for the following molecules.

- (a) $\text{C}_2\text{H}_6\text{O}$ (b) CH_2Cl_2 (c) $\text{C}_2\text{H}_7\text{N}$ (d) C_3H_6 (e) CH_3OH
(f) CH_3Cl (g) CH_3SH (h) CS_2 (i) H_2O_2 (j) N_2H_2

Problem 1.28 Solution



Problem 1.29

Determine which of the following molecules is(are) polar or nonpolar?

- (a) O_2 (b) Cl_2 (c) N_2 (d) H_2O (e) CO_2
(f) H_2O_2 (g) HCl (h) NH_3 (i) H_2S (j) CO
(k) CH_2Cl_2 (l) BF_3 (m) CF_4

Problem 1.29 Solution

(a) Nonpolar; (b) nonpolar; (c) nonpolar; (d) polar; (e) nonpolar; (f) polar; (g) polar; (h) polar; (i) polar; (j) polar; (k) polar; (l) nonpolar; (m) nonpolar

Problem 1.30

List the following atoms in terms of increasing electronegativity, i.e. the least electronegative first: Li; C; O; N; Na

Problem 1.30 Solution

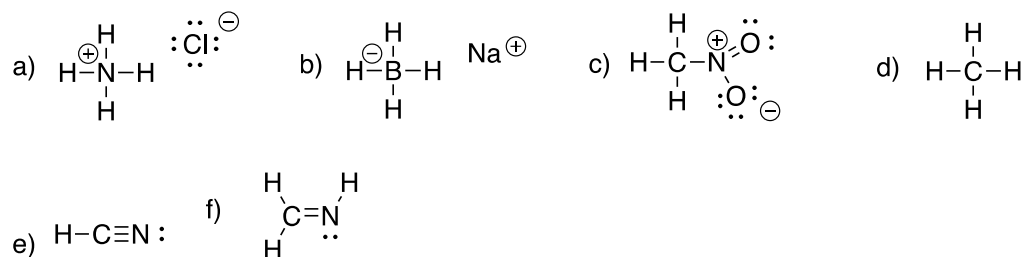
$\text{Li} < \text{Na} < \text{C} < \text{N} < \text{O}$

Problem 1.31

Draw the Lewis dot structures for the following compounds, showing appropriate formal charges.

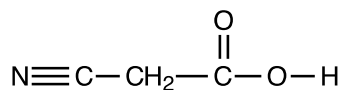
- (a) NH_4NCl (b) NaBH_4 (c) CH_3NO_2 (d) CH_4 (e) HCN (f) CH_3N

Problem 1.31 Solution

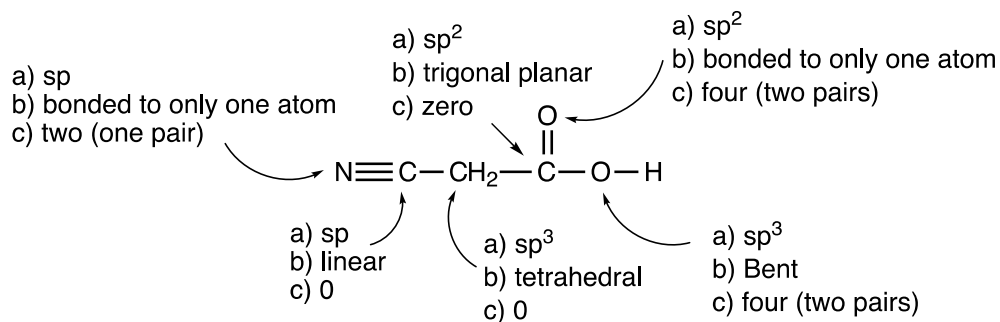


Problem 1.32

For the molecule shown below, indicate: (a) the type of hybridized orbitals for each atom (except hydrogen); (b) the geometry around each atom; and (c) the number of unshared pair(s) of electrons on each atom.

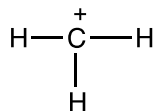


Problem 1.32 Solution



Problem 1.33

What is the geometry about the central atom of the following?



- (a) BH_3 (b) Methyl carbocation

Problem 1.33 Solution

- (a) Trigonal planar (b) Trigonal planar

Problem 1.34

Which of the following compounds would you expect to have a covalent bond(s)?

NaCl , KCl , CCl_4 , CaO

Problem 1.34 Solution

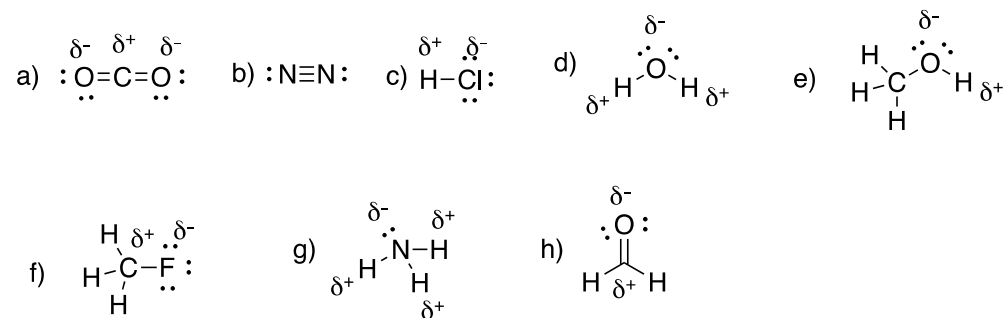
CCl_4 . The difference in electronegativity between carbon and chlorine is not great enough to make an ionic bond as is the case with the other molecules.

Problem 1.35

Draw Lewis dot structures of each of the following molecules and use δ^+ and δ^- where possible to show the polarity of each covalent bond in the following molecules.

- (a) CO_2 (b) N_2 (c) HCl (d) H_2O (e) CH_3OH
(f) CH_3F (g) NH_3 (h) CH_2O

Problem 1.35 Solution

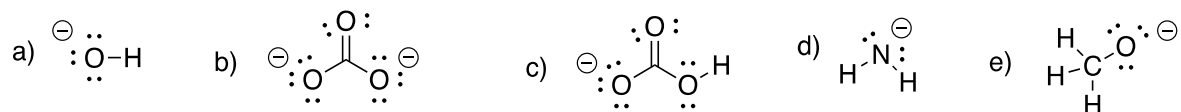


Problem 1.36

Write Lewis dot structures for each of the following ions.

- (a) OH^- (b) CO_3^{2-} (c) HCO_3^- (d) NH_2^- (e) CH_3O^-

Problem 1.36 Solution

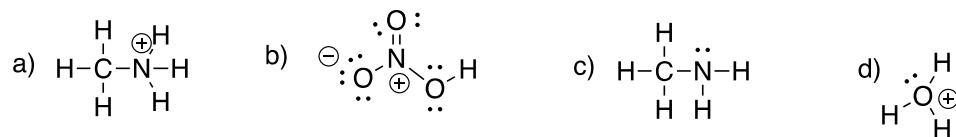


Problem 1.37

Give Lewis dot structures of the following and determine the formal charge on the nitrogen and the oxygen atoms.

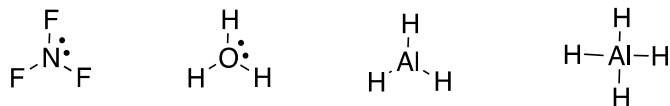


Problem 1.37 Solution

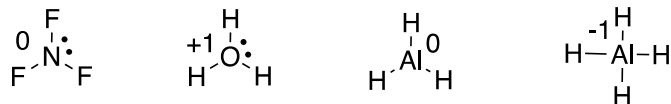


Problem 1.38

What is the formal charge on the central atom of each of the species shown below?

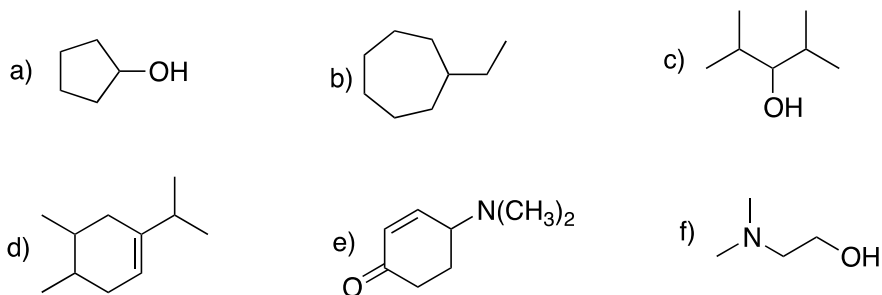


Problem 1.38 Solution

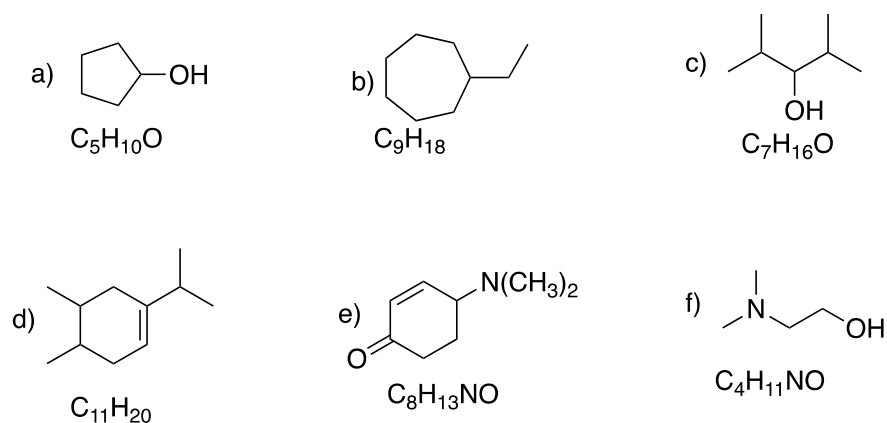


Problem 1.39

What are the molecular formulas for the following compounds?

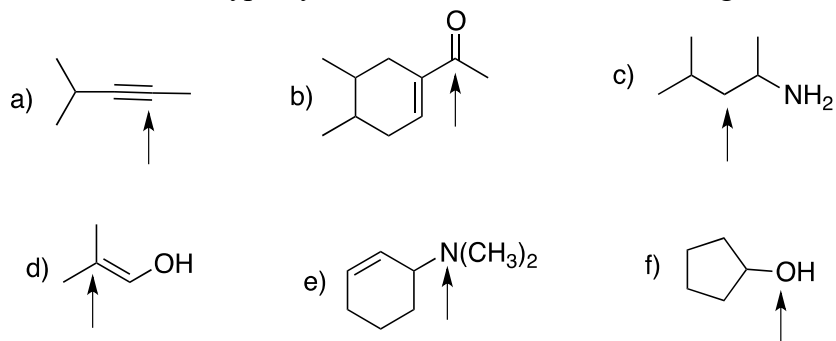


Problem 1.39 Solution

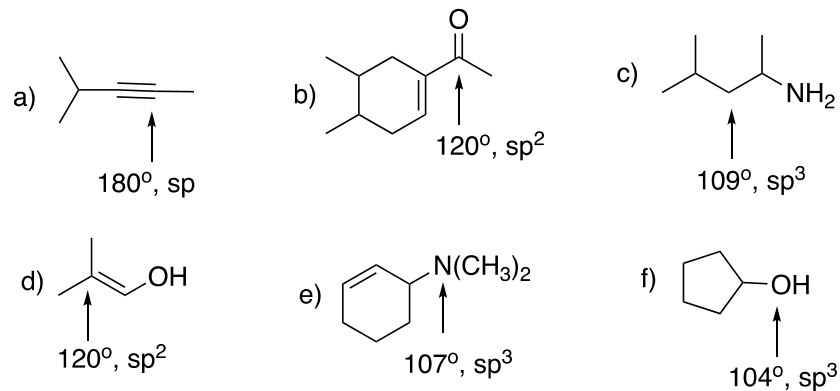


Problem 1.40

For the following molecules, give the approximate bond angles about each of the atoms indicated by an arrow and determine the type hybridized orbitals used in bonding to the other atoms.

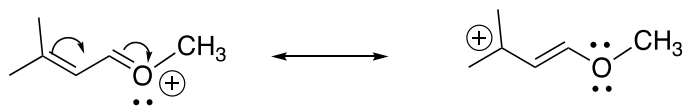


Problem 1.40 Solution

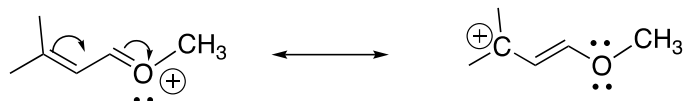


Problem 1.41

Below are the resonance structures of an ion. Determine which is the major resonance contributor and explain your answer.



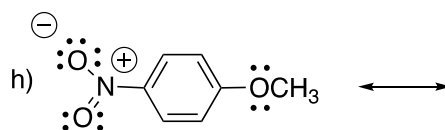
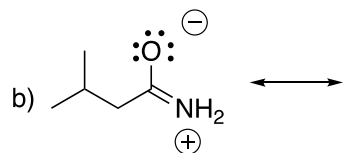
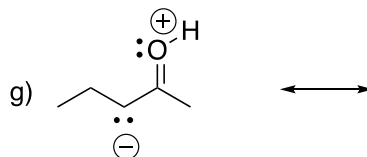
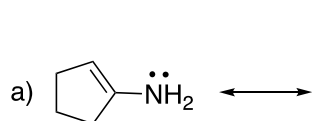
Problem 1.41 Solution



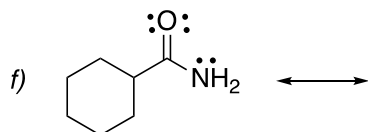
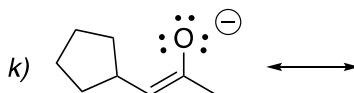
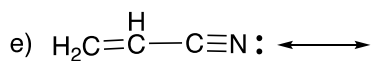
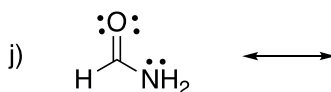
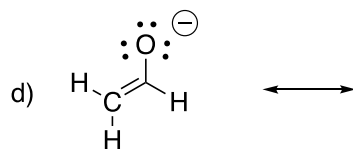
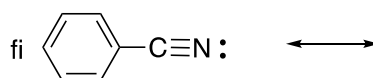
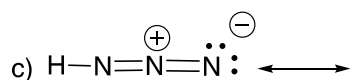
Major resonance contributor.
The positive charge is on the least electronegative atom, carbon, compared to oxygen.

Problem 1.42

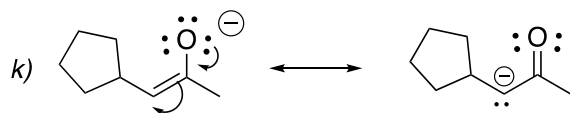
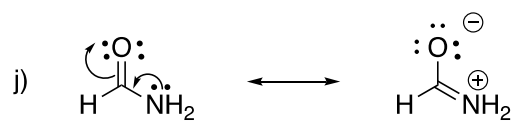
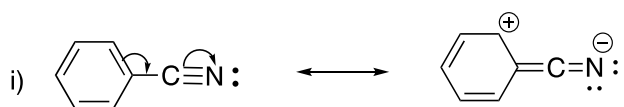
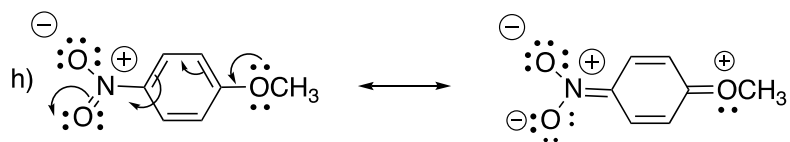
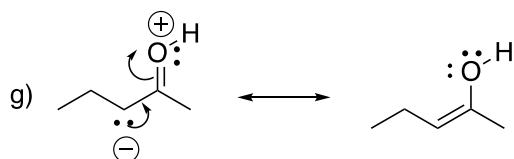
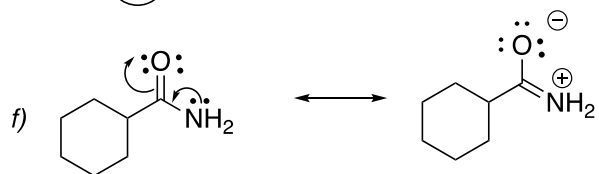
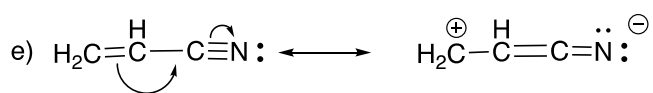
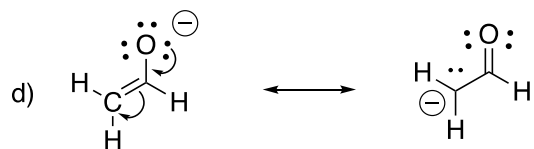
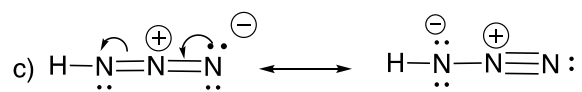
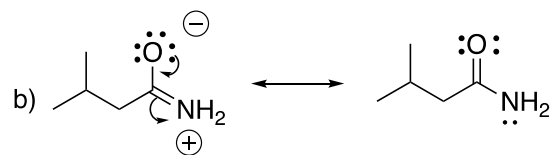
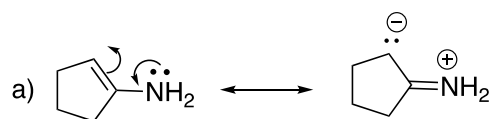
For the molecules shown below, draw another resonance structure. Use curved arrows to indicate electron movement and include formal charges where appropriate.



(delocalize electrons from the OCH₃ group)



Problem 1.42 Solution

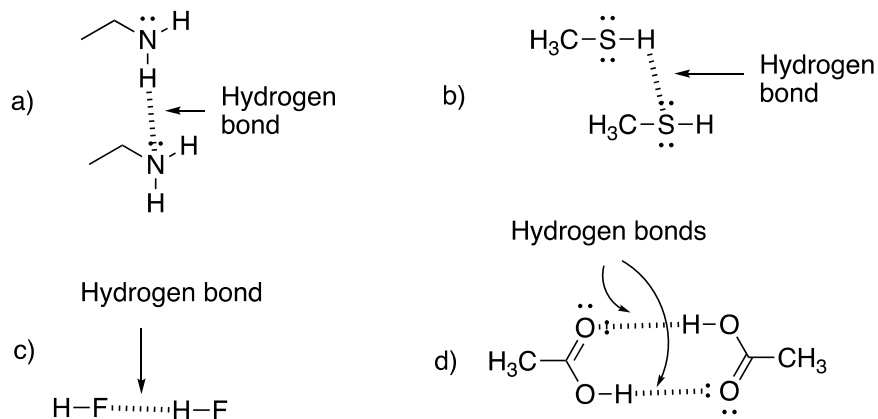


Problem 1.43

Using dotted lines to indicate intermolecular interactions, indicate the hydrogen bond formed for each of the following molecules.

- (a) $\text{CH}_3\text{CH}_2\text{NH}_2$ (b) CH_3SH (c) HF (d) $\text{CH}_3\text{CO}_2\text{H}$

Problem 1.43 Solution



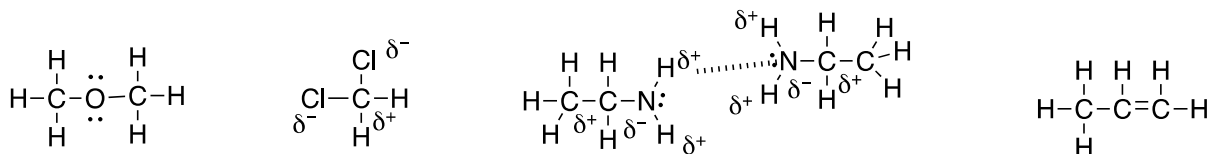
Problem 1.44

Consider the molecular formulas shown below, then answer the following questions.



- (a) For each molecule, draw Lewis dot structure(s). [Note: there may be more than one structure for each molecule.]
- (b) Using the partial charge notation (δ^+ or δ^-), indicate the polarity of each covalent bond for each isomer that you have drawn.
- (c) Is the molecule that you have drawn capable of forming hydrogen bonds to water?
- (d) Use dotted lines ($\cdots$) to indicate the hydrogen bonded structure that would exist between two molecules that you have drawn in question b.

Problem 1.44 Solution



Capable of forming H-bonds to water

Problem 1.45

Even though the molecular weight of methanamine (CH_3NH_2) is approximately the same as that of methanol (CH_3OH), the boiling point of methanol is greater than that of methanamine (65.0°C vs. -6.3°C). Briefly explain this observation.

Problem 1.45 Solution

The hydrogen bonds formed with methanol are stronger than those formed with methanamine.

Problem 1.46

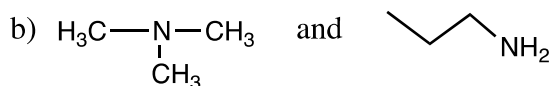
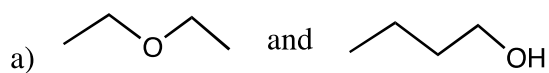
The molecular weight of CH_3SH (methanethiol) is greater than that of methanol (CH_3OH), yet the boiling point of methanethiol is 6°C , compared to the boiling point of methanol which is 65.0°C . Briefly explain this observation.

Problem 1.46 Solution

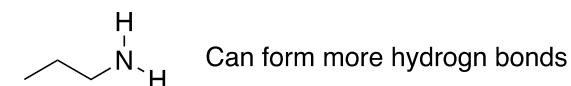
The hydrogen bonds formed with methanol are stronger than those formed with either methanamine or methanethiol.

Problem 1.47

Which of the following pairs of molecules has a higher boiling point? Explain your answer.

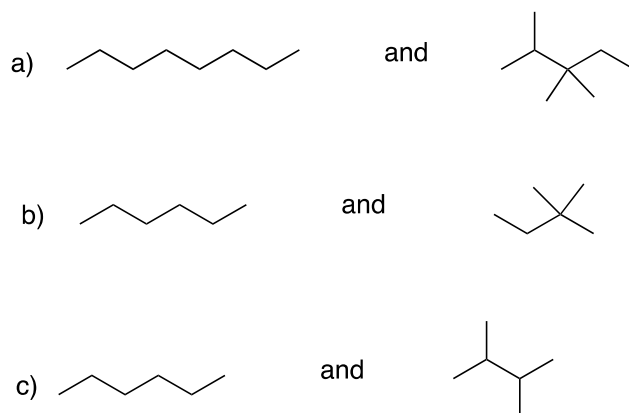


Problem 1.47 Solution

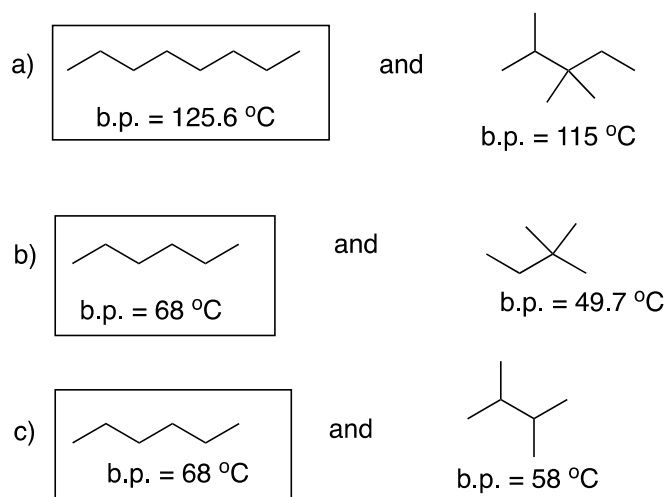


Problem 1.48

For each of the following pairs of compounds, which has a higher boiling point? Explain your answer.



Problem 1.48 Solution



Problem 1.49

Which of the following compounds have dipole moments greater than zero?

- (a) CH₂Cl₂ (b) BF₃ (c) CF₄ (d) H₂O (e) NH₃

Problem 1.49 Solution

- (a) CH₂Cl₂ ($\mu > 0$); (b) BF₃ ($\mu = 0$); (c) CF₄ ($\mu = 0$); (d) H₂O ($\mu > 0$); (e) NH₃ ($\mu > 0$)