called a **polar covalent bond.** In a polar covalent bond the more electronegative atom assumes a partial negative charge and the less electronegative atom assumes a partial positive charge. The polarity of a bond may be indicated by the symbol $\overset{+}{\longrightarrow}$. The head of the arrow points in the direction of the more electronegative atom. The tail, marked with a plus sign, is located at the less electronegative atom. More frequently, the partial charges are denoted by the Greek letter symbols $\delta+$ and $\delta-$ (pronounced *delta plus* and *delta minus*).

$$\overset{+}{\longrightarrow}\quad\text{or}\quad \overset{\delta+}{\text{C}}—\overset{\delta-}{\text{F}}$$

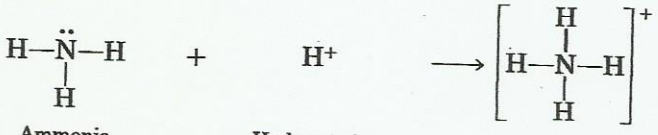
As we proceed through our study of organic chemistry, we will find that polar bonds exert special effects on the physical and chemical properties of organic molecules. In general, most reactions involve changes in polar covalent bonds (C—O, C—Cl, etc.) while nonpolar covalent bonds (C—C, C—H, etc.) remain unaltered.

Problem 1.2 Show the partial charges by placing the $\delta+$ and $\delta-$ symbols on the appropriate atoms in the following polar covalent bonds.

(a) H—Cl **(b)** O—H **(c)** C—Cl
(d) N—H **(e)** C—O **(f)** C—N

C Coordinate Covalent Bonding

In the covalent bonding discussed so far each of the two atoms contributed one electron to the electron pair shared between them. There are molecules in which one atom supplies *both* electrons to another atom in the formation of a covalent bond. A covalent bond thus formed is called a **coordinate covalent bond.** For example, when ammonia, $:NH_3$, reacts with a proton, H^+, to form an ammonium ion, NH_4^+, the nitrogen atom in ammonia supplies both electrons to the new bond.



Ammonia (Lewis base—furnishes electron pair) | Hydrogen ion (Lewis acid—accepts electron pair) | Ammonium ion

The species that furnishes the electron pair to form a coordinate covalent bond is called a **Lewis base.** The species that accepts the electron pair to complete its valence shell is called a **Lewis acid.** In subsequent chapters we shall have numerous occasions to refer to the concept of Lewis acids and Lewis bases to explain how chemical reactions occur.

Problem 1.3 Each of the following interactions involves the formation of a coordinate covalent bond. Indicate (1) the structure of the product formed and (2) which species acts as a Lewis acid and which acts as a Lewis base.

Exercises

Valence Electrons and Electron-Dot Structures [Sec. 1.3]

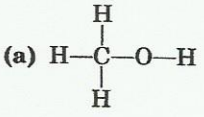
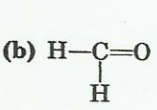
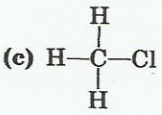
1.1 Write an electron-dot structure for each compound.

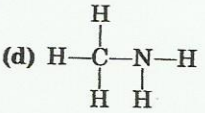
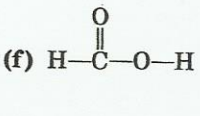
(a) H_2O **(b)** NH_3 **(c)** CH_4 **(d)** KBr
(e) $CaCl_2$ **(f)** H_2CO **(g)** CO_2 **(h)** HCN

Ionic, Covalent, and Polar Covalent Bonding [Sec. 1.4A, B]

1.2 List which compound(s) in Exercise 1.1 contain(s) **(a)** only ionic bonds, **(b)** only nonpolar covalent bonds, **(c)** only polar covalent bonds, and **(d)** both nonpolar and polar covalent bonds.

1.3 Show the partial charges by placing $\delta+$ and $\delta-$ symbols on the atoms involved in a polar covalent bond.

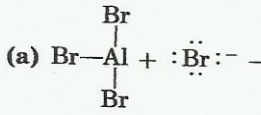
(a)  **(b)**  **(c)** 

(d)  **(e)** **(f)** 

1.4 Arrange the hydrogen halides (HI, HBr, HCl, HF) in order of polarity, from the most polar to the least polar.

Coordinate Covalent Bonding; Lewis Acid and Lewis Base [Sec. 1.4C]

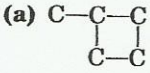
1.5 For each of the following Lewis acid-base interactions (1) indicate the structure of the product formed and (2) identify the Lewis acid and the Lewis base.

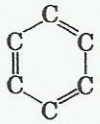
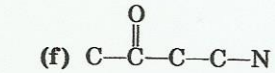
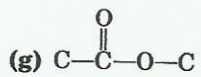
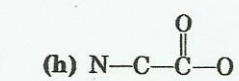
(a)  **(b)** $CH_3\ddot{N}H_2 + H^+ \longrightarrow$

(c) $H^+ + :\ddot{O}H^- \longrightarrow$ **(d)** $C_2H_5\ddot{O}H + H^+ \longrightarrow$

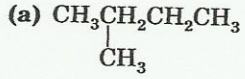
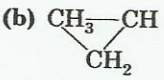
Covalence Number and Structural Formula [Secs. 1.5, 1.6]

1.6 Given the skeletal structure, and assuming that only hydrogen atoms are missing, draw the correct structural formula for each of the following.

(a)  **(b)** C=C—C **(c)** C≡C—C **(d)** Cl—C—C=C

(e)  **(f)**  **(g)**  **(h)** 

1.7 Check the following structures to see whether or not they represent possible compounds within the rules of covalence. State either "possible" or "impossible" for each.

(a)  **(b)** 

(c) $CH_3CH_2—O—CH_2—O—CH_3$

(d) $CH_3\underset{\displaystyle OH}{\underset{|}{C}}HCH_2\overset{\displaystyle O}{\overset{\|}{C}}—OH$

(e) $CH_3CH_2\underset{\displaystyle Cl}{\underset{|}{C}}H_2Cl$

(f) $CH_3CH_2=CH_2$

(g) $CH_3—NH_2—CH_3$

(h) $CH_3C\equiv CCH_2CH_3$

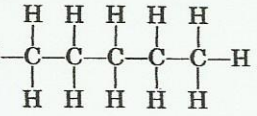
1.8 Draw a correct structural formula for each "impossible" representation in Exercise 1.7 by either adding or removing hydrogen atoms.

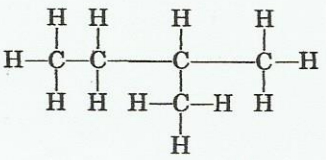
1.9 Draw an expanded structural formula corresponding to each molecular formula.

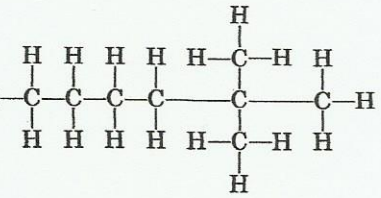
(a) $CHCl_3$ (b) CH_2F_2 (c) CH_4O (d) CH_5N (e) C_2H_2
(f) C_2H_4 (g) C_2H_6 (h) C_2H_5Cl (i) CH_3N (j) H_3NO

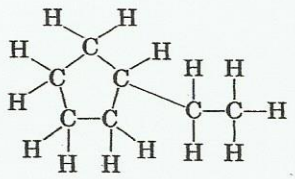
Condensed and Expanded Structural Formulas [Sec. 1.7]

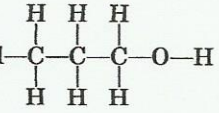
1.10 Convert each of the following expanded structural formulas into (1) a partially condensed and (2) a fully condensed formula.

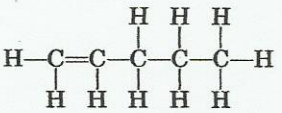
(a)


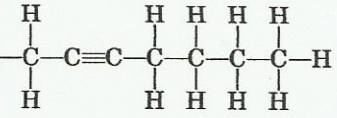
(b)


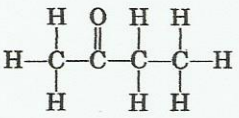
(c)


(d)


(e)


(f)


(g)


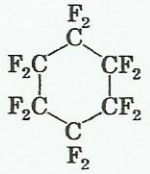
(h)


1.11 Draw the fully expanded structures corresponding to the partially condensed formulas (a–d) and fully condensed formulas (e–h).

(a) $CH_3—CH_2—CH_2—CH_3$

(b) $CH_3—CHF—\underset{\displaystyle CH_3}{\underset{|}{C}}H—CH=CH_2$

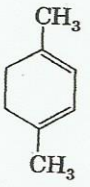
(c) $HO—\overset{\displaystyle O}{\overset{\|}{C}}—\underset{\displaystyle OH}{\underset{|}{C}}H—\underset{\displaystyle OH}{\underset{|}{C}}H—\overset{\displaystyle O}{\overset{\|}{C}}—OH$

(d)


(e) $(CH_3)_2CH(CH_2)_4CH_3$

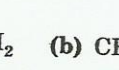
(f) $CH_3CH=CHCHO$

(g) $CH_3C{\equiv}CCH_2COOH$

(h) 

Bond Length, Bond Strength, and Bond Angle [Secs. 1.9, 1.10]

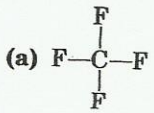
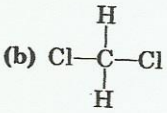
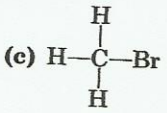
1.12 For each structure, predict which of the two bonds shown has (1) the greater bond dissociation energy and (2) the longer bond length.

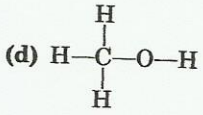
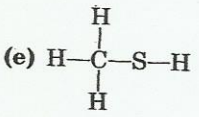
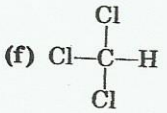
(a) $CH_3{-}CH{=}CH_2$ (b) $CH_3{-}C{\equiv}CH$ (c)  (d) $H_2N{-}CH_2C{\equiv}N$

1.13 (a) What is the size of the H—C—H bond angle in methane, CH_4?
(b) What would you expect the size of the F—C—F bond angle in carbon tetrafluoride, CF_4, to be?

Hybridization and Shape of Molecules [Sec. 1.10]

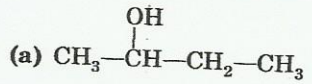
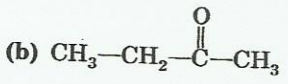
1.14 Indicate (1) the type of hybridized orbital utilized by carbon in each of the following structures and (2) the shape of each molecule.

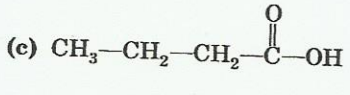
(a)  (b)  (c) 

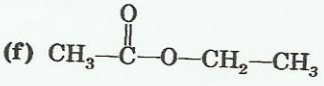
(d)  (e)  (f) 

Functional Groups and Classification of Compounds [Sec. 1.11]

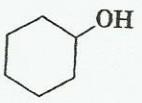
1.15 Name the class to which each of the following compounds belongs.

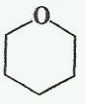
(a)  (b) 

(c)  (d) $CH_3CH{=}CH_2$

(e) $CH_3{-}C{\equiv}CH$ (f) 

1.16 Group together those compounds that you expect to behave chemically in a similar manner.

(a) CH_3OH (b) CH_3CH_2Cl (c) $CH_3OCH_2CH_3$ (d) 

(e) CH_2F_2 (f) $CH_3CH{=}CH_2$ (g) CH_3COOH (h) 

(i) C_2H_4 (j) $CHBr_3$ (k) $HOCH_2CH_2OH$ (l) 

Covalence Number, Structural Formula, and Functional Group

[Secs. 1.5, 1.6, 1.11]

1.17 (a) One alcohol and one ether correspond to C_2H_6O. Draw their structures.
(b) One aldehyde and one ketone correspond to C_3H_6O. Draw their structures.
(c) One carboxylic acid and one ester correspond to $C_2H_4O_2$. Draw their structures.