

Chapter 2 – Polar Covalent Bonds; Acids and Bases

Chapter Outline

I. Polar covalent bonds (Sections 2.1 – 2.3).

A. Electronegativity (Section 2.1).

1. Although some bonds are totally ionic and some are totally covalent, most chemical bonds are polar covalent bonds.
In these bonds, electrons are attracted to one atom more than to the other atom.
2. Bond polarity is due to differences in electronegativity (EN).
 - a. Elements on the right side of the periodic table are more electronegative than elements on the left side.
 - b. Carbon has an EN of 2.5.
 - c. Elements with $EN > 2.5$ are more electronegative than carbon.
 - d. Elements with $EN < 2.5$ are less electronegative than carbon.
3. The difference in EN between two elements can be used to predict the polarity of a bond.
 - a. If $\Delta EN < 0.4$, a bond is nonpolar covalent.
 - b. If ΔEN is between 0.4 and 2.0, a bond is polar covalent.
 - c. If $\Delta EN > 2.0$, a bond is ionic.
 - d. The symbols $\delta+$ and $\delta-$ are used to indicate partial charges.
 - e. A crossed arrow is used to indicate bond polarity.

The tail of the arrow is electron-poor, and the head of the arrow is electron-rich.

4. Electrostatic potential maps are also used to show electron-rich (red) and electron-poor (blue) regions of molecules.
 5. An inductive effect is an atom's ability to polarize a bond.
- ##### B. Dipole moment (Section 2.2).
1. Dipole moment is the measure of a molecule's overall polarity.
 2. Dipole moment (μ) = $Q \times r$, where Q = charge and r = distance between charges.
Dipole moment is measured in debyes (D).
 3. Dipole moment can be used to measure charge separation.
 4. Water and ammonia have large values of D; methane and ethane have $D = 0$.

C. Formal charge (Section 2.3).

1. Formal charge (FC) indicates electron "ownership" in a molecule.
2.
$$(FC) = \left[\begin{array}{c} \text{\# of valence} \\ \text{electrons} \end{array} \right] - \left[\frac{\text{\# of bonding electrons}}{2} \right] - \left[\begin{array}{c} \text{\# nonbonding} \\ \text{electrons} \end{array} \right]$$
3. Molecules that are neutral overall but that have + and – charges on individual atoms are dipolar.

II. Resonance (Sections 2.4 – 2.6).

A. Chemical structures and resonance (Section 2.4).

1. Some molecules (acetate ion, for example) can be drawn as two (or more) different electron-dot structures.
 - a. These structures are called resonance structures.
 - b. The true structure of the molecule is intermediate between the structures.
 - c. The true structure is called a resonance hybrid.
2. Resonance structures differ only in the placement of π and nonbonding electrons.
All atoms occupy the same positions.
3. Resonance is an important concept in organic chemistry.

- B. Rules for resonance forms (Section 2.5).
1. Individual resonance forms are imaginary, not real.
 2. Resonance forms differ only in the placement of their π or nonbonding electrons.
A curved arrow is used to indicate the movement of electrons, not atoms.
 3. Different resonance forms of a molecule don't have to be equivalent.
If resonance forms are nonequivalent, the structure of the actual molecule resembles the more stable resonance form(s).
 4. Resonance forms must be valid Lewis structures and obey normal rules of valency.
 5. The resonance hybrid is more stable than any individual resonance form.
- C. A useful technique for drawing resonance forms (Section 2.6).
1. Any three-atom grouping with a multiple bond adjacent to a nonbonding p orbital has two resonance forms.
 2. One atom in the grouping has a lone electron pair, a vacant orbital or a single electron.
 3. By recognizing these three-atom pieces, resonance forms can be generated.
- III. Acids and bases (Sections 2.7 – 2.11).
- A. Brønsted-Lowry definition (Section 2.7).
1. A Brønsted-Lowry acid donates an H^+ ion; a Brønsted-Lowry base accepts H^+ .
 2. The product that results when a base gains H^+ is the conjugate acid of the base; the product that results when an acid loses H^+ is the conjugate base of the acid.
- B. Acid and base strength (Section 2.8 – 2.10).
1. A strong acid reacts almost completely with water (Section 2.8).
 2. The strength of an acid in water is indicated by K_a , the acidity constant.
 3. Strong acids have large acidity constants, and weaker acids have smaller acidity constants.
 4. The pK_a is normally used to express acid strength.
 - a. $pK_a = -\log K_a$
 - b. A strong acid has a small pK_a , and a weak acid has a large pK_a .
 - c. The conjugate base of a strong acid is a weak base, and the conjugate base of a weak acid is a strong base.
 5. Predicting acid–base reactions from pK_a (Section 2.9).
 - a. An acid with a low pK_a reacts with the conjugate base of an acid with a high pK_a
 - b. In other words, the products of an acid–base reaction are more stable than the reactants.
 6. Organic acids and organic bases (Section 2.10).
 - a. There are two main types of organic acids:
 - i. Acids that contain hydrogen bonded to oxygen.
 - ii. Acids that have hydrogen bonded to the carbon next to a $C=O$ group.
 - b. The main type of organic base contains a nitrogen atom with a lone electron pair.
- C. Lewis acids and bases (Section 2.11).
1. A Lewis acid accepts an electron pair.
 - a. A Lewis acid may have either a vacant low-energy orbital or a polar bond to hydrogen.
 - b. Examples include metal cations, halogen acids, Group 3 compounds and transition-metal compounds.
 2. A Lewis base has a pair of nonbonding electrons.
 - a. Most oxygen- and nitrogen-containing organic compounds are Lewis bases.
 - b. Many organic Lewis bases have more than one basic site.
 3. A curved arrow shows the movement of electrons from a Lewis base to a Lewis acid.

IV. Molecular models (Section 2.12).

- A. Space-filling models show the crowding within a molecule.
- B. Ball-and-stick models show bonds.

V. Noncovalent interactions in molecules (Section 2.13).

- A. Dipole-dipole interactions occur between polar molecules as a result of electrostatic interactions among dipoles.
- B. Dispersion forces result from the constantly changing electron distribution within molecules.

These forces are small but their cumulative effect may be important.

C. Hydrogen bonding.

1. Hydrogen bonds form between a hydrogen bonded to an electronegative atom and an unshared electron pair on another electronegative atom.
2. Hydrogen bonds are extremely important in living organisms.
3. Hydrophilic substances dissolve in water because they are capable of forming hydrogen bonds.
4. Hydrophobic substances don't form hydrogen bonds.

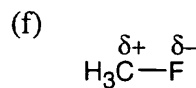
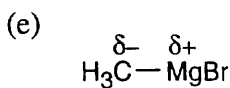
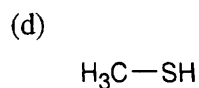
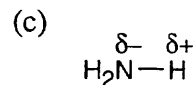
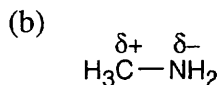
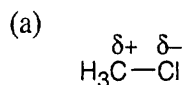
Answers to Problems

- 2.1 After solving this problem, use Figure 2.2 to check your answers. The larger the number, the more electronegative the element.

<i>More electronegative</i>	<i>Less electronegative</i>
(a) H (2.1)	Li (1.0)
(b) Br (2.8)	B (2.0)
(c) Cl (3.0)	I (2.5)
(d) C (2.5)	H (2.1)

Carbon is slightly more electronegative than hydrogen.

- 2.2 As in Problem 2.1, use Figure 2.2.



Carbon and sulfur have identical electronegativities.

- 2.3 **Strategy:** Use Figure 2.2 to find the electronegativities of each element. Calculate ΔEN and rank the answers in order of increasing ΔEN .

Solution:

Carbon: EN = 2.5
 Lithium: EN = 1.0

$$\Delta\text{EN} = 1.5$$

Carbon: EN = 2.5
 Potassium: EN = 0.8

$$\Delta\text{EN} = 1.7$$

Fluorine: EN = 4.0
 Carbon: EN = 2.5

$$\Delta\text{EN} = 1.5$$

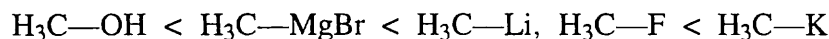
Carbon: EN = 2.5
Magnesium: EN = 1.2

$$\Delta\text{EN} = 1.3$$

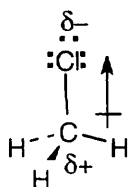
Oxygen: EN = 3.5
Carbon: EN = 2.5

$$\Delta\text{EN} = 1.0$$

The most polar bond has the largest ΔEN . Thus, in order of increasing bond polarity:

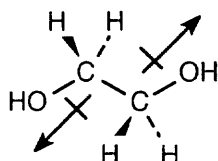


- 2.4 In an electrostatic potential map, the color red indicates regions of a molecule that are electron-rich. The map shows that chlorine is the most electronegative atom in methyl alcohol, and the direction of polarity of the C–Cl bond is:



Chloromethane

2.5



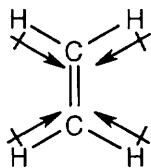
Ethylene glycol

The dipole moment of ethylene glycol is zero because the bond polarities of the two carbon-oxygen bonds cancel.

- 2.6 **Strategy:** For each bond, identify the more electronegative element, and draw an arrow that points from the less electronegative element to the more electronegative element. Estimate the sum of the individual dipole moments to arrive at the dipole moment for the entire molecule.

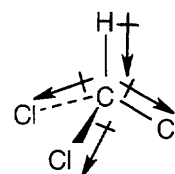
Solution:

(a)



0 dipole moment

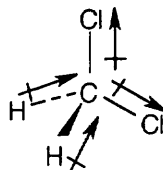
(b)



↑

net dipole moment

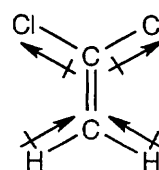
(c)



↗

net dipole moment

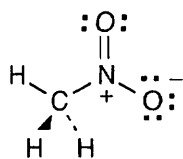
(d)



↑

net dipole moment

2.7 Use the formula in Section 2.3 to confirm the formal charges shown for nitromethane.



Nitromethane

$$\text{Formal charge (FC)} = \left[\begin{array}{c} \text{\# of valence} \\ \text{electrons} \end{array} \right] - \left[\frac{\text{\# of bonding electrons}}{2} \right] - \left[\begin{array}{c} \text{\# nonbonding} \\ \text{electrons} \end{array} \right]$$

$$\text{For nitrogen: FC} = 5 - \frac{8}{2} - 0 = +1$$

$$\text{For oxygen: FC} = 6 - \frac{2}{2} - 6 = -1$$

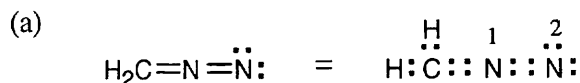
2.8 **Strategy:** To find the formal charge of an atom in a molecule, follow these two steps:

(1) Draw an electron-dot structure of the molecule.

(2) Use the formula in Section 2.3 to determine formal charge for each atom. The periodic table shows the number of valence electrons of the element, and the electron-dot structure shows the number of bonding and nonbonding electrons.

$$\text{Formal charge (FC)} = \left[\begin{array}{c} \text{\# of valence} \\ \text{electrons} \end{array} \right] - \left[\frac{\text{\# of bonding electrons}}{2} \right] - \left[\begin{array}{c} \text{\# nonbonding} \\ \text{electrons} \end{array} \right]$$

Solution:

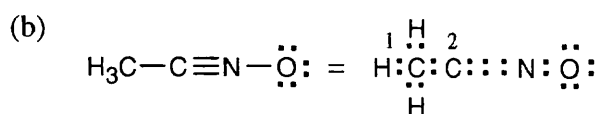


$$\text{For carbon: FC} = 4 - \frac{8}{2} - 0 = 0$$

$$\text{For nitrogen 1: FC} = 5 - \frac{8}{2} - 0 = +1$$

$$\text{For nitrogen 2: FC} = 5 - \frac{4}{2} - 4 = -1$$

Remember: *Valence electrons* are the electrons characteristic of a specific element. *Bonding electrons* are those electrons involved in bonding to other atoms. *Nonbonding electrons* are those electrons in lone pairs.

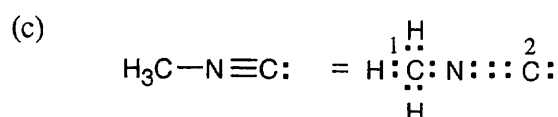


$$\text{For carbon 1: FC} = 4 - \frac{8}{2} - 0 = 0$$

$$\text{For carbon 2: FC} = 4 - \frac{8}{2} - 0 = 0$$

$$\text{For nitrogen: FC} = 5 - \frac{8}{2} - 0 = +1$$

$$\text{For oxygen: FC} = 6 - \frac{2}{2} - 6 = -1$$



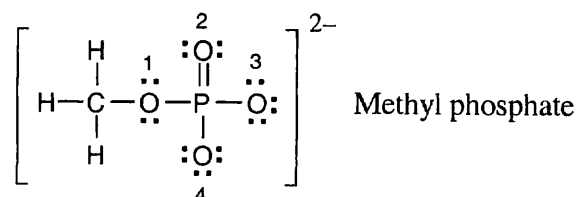
$$\text{For carbon 1: FC} = 4 - \frac{8}{2} - 0 = 0$$

$$\text{For carbon 2: FC} = 4 - \frac{6}{2} - 2 = -1$$

$$\text{For nitrogen: FC} = 5 - \frac{8}{2} - 0 = +1$$

2.9

$$\text{Formal charge (FC)} = \left[\begin{array}{c} \# \text{ of valence} \\ \text{electrons} \end{array} \right] - \left[\frac{\# \text{ of bonding electrons}}{2} \right] - \left[\begin{array}{c} \# \text{ nonbonding} \\ \text{electrons} \end{array} \right]$$



$$\text{For oxygen 1: FC} = 6 - \frac{4}{2} - 4 = 0$$

$$\text{For oxygen 2: FC} = 6 - \frac{4}{2} - 4 = 0$$

$$\text{For oxygen 3: FC} = 6 - \frac{2}{2} - 6 = -1$$

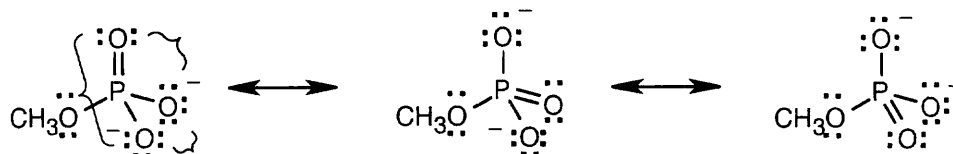
$$\text{For oxygen 4: FC} = 6 - \frac{2}{2} - 6 = -1$$

Oxygen atoms 3 and 4 each have a formal charge of -1.

- 2.10 Strategy:** Look for three-atom groupings that contain a multiple bond next to an atom with a p orbital. Exchange the positions of the bond and the electrons in the p orbital to draw the resonance form of each grouping.

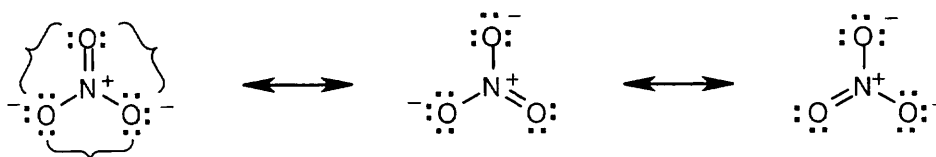
Solution:

(a) Methyl phosphate anion has 3 three-atom groupings and thus has 3 resonance forms.

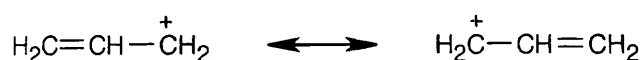


Recall from Chapter 1 that phosphorus, a third-row element, can form more than four covalent bonds

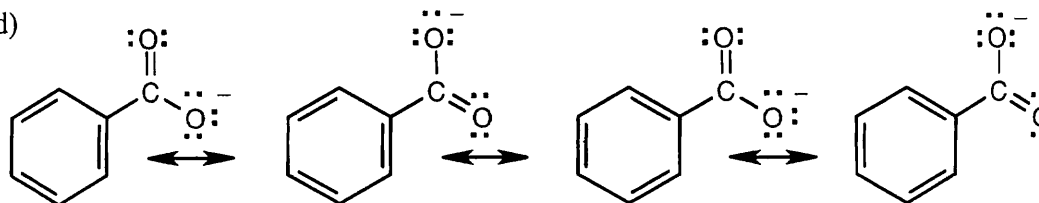
(b)



(c)

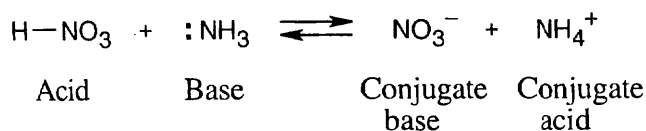


(d)



- 2.11 Strategy:** When an acid loses a proton, the product is the conjugate base of the acid. When a base gains a proton, the product is the conjugate acid of the base.

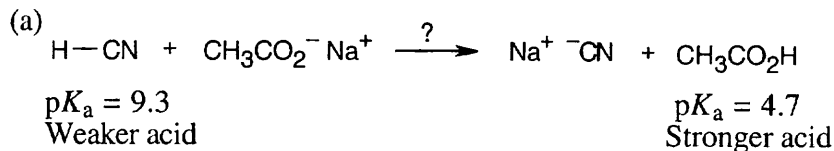
Solution:



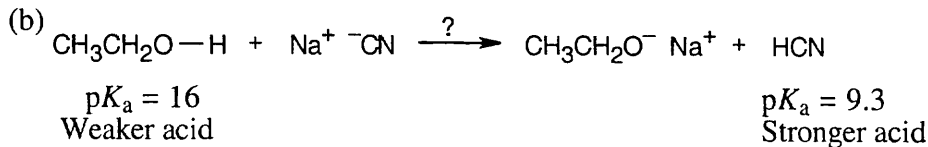
- 2.12** Recall from Section 2.8 that a stronger acid has a smaller $\text{p}K_a$ and a weaker acid has a larger $\text{p}K_a$. Accordingly, phenylalanine ($\text{p}K_a = 1.83$) is a stronger acid than tryptophan ($\text{p}K_a = 2.83$).
- 2.13** $\text{HO}-\text{H}$ is a stronger acid than $\text{H}_2\text{N}-\text{H}$. Since H_2N^- is a stronger base than HO^- , the conjugate acid of H_2N^- ($\text{H}_2\text{N}-\text{H}$) is a weaker acid than the conjugate acid of HO^- ($\text{HO}-\text{H}$).

- 2.14 Strategy:** Use Table 2.3 to find the strength of each acid. A reaction takes place as written if the stronger acid is the reactant.

Solution:

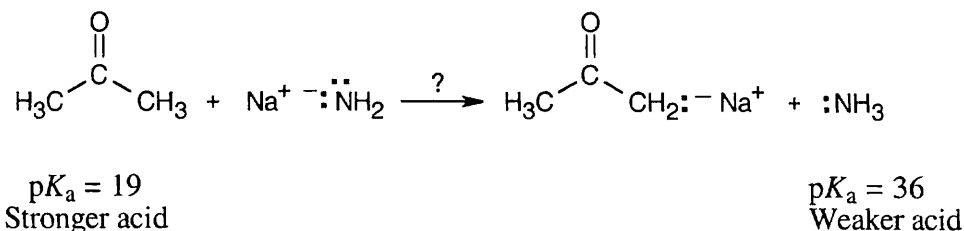


Remember that the lower the $\text{p}K_a$, the stronger the acid. Thus $\text{CH}_3\text{CO}_2\text{H}$, not HCN , is the stronger acid, and the above reaction will not take place in the direction written.



Using the same reasoning as in part (a), we can see that the above reaction will not occur.

2.15

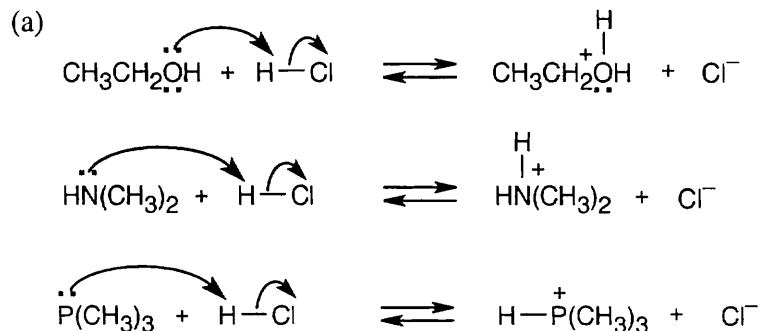


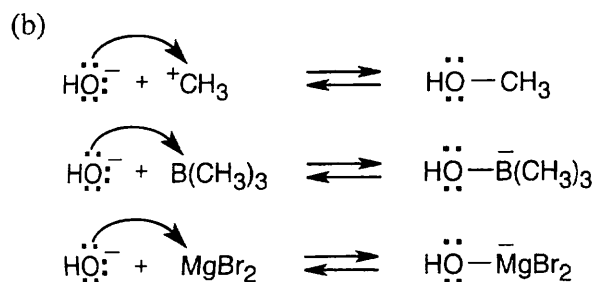
The above reaction will take place as written.

- 2.16** Enter -9.31 into a calculator and use the INV LOG function to arrive at the answer $K_a = 4.9 \times 10^{-10}$.

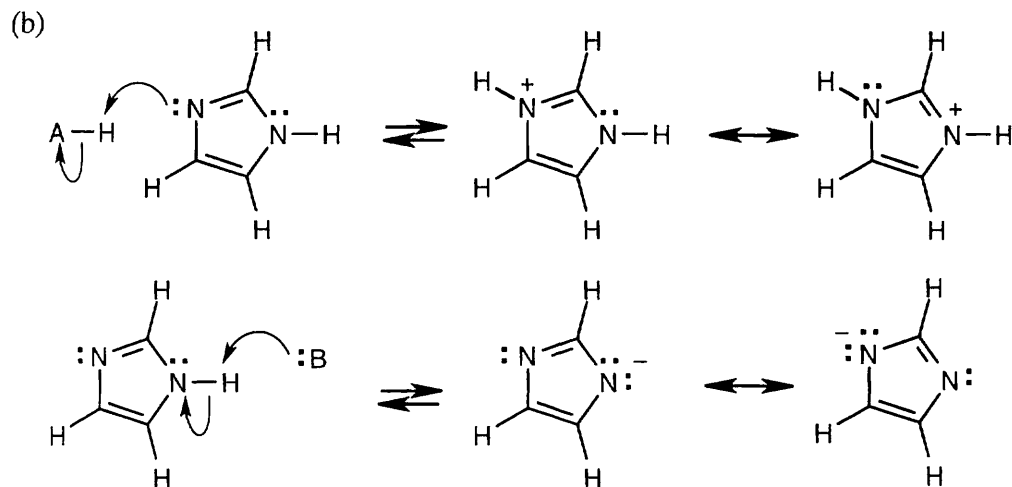
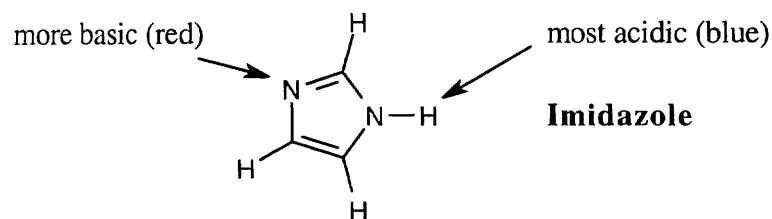
- 2.17 Strategy:** Locate the electron pair(s) of the Lewis base and draw a curved arrow from the electron pair to the Lewis acid. The electron pair moves from the atom at the tail of the arrow (Lewis base) to the atom at the point of the arrow (Lewis acid).

Solution: (Note: electron dots have been omitted from Cl^- to reduce clutter.)

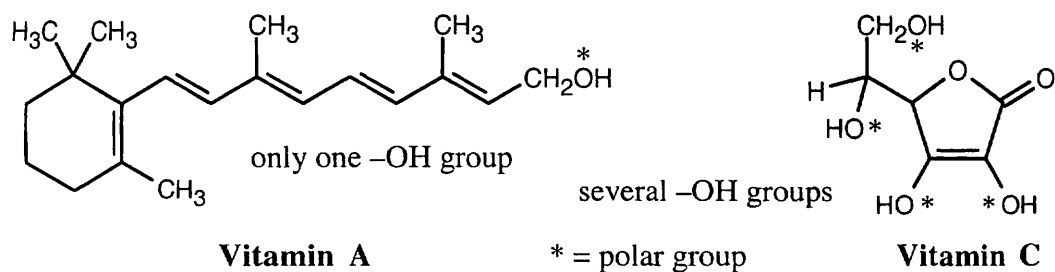




2.18 (a) The nitrogen on the left is more electron-rich and more basic. The indicated hydrogen is most electron-poor (bluest) and is most acidic.



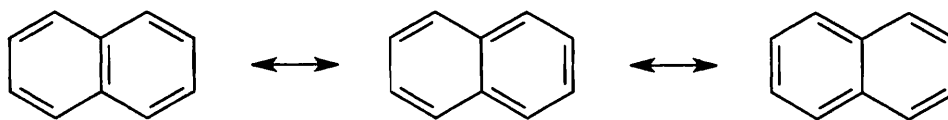
2.19



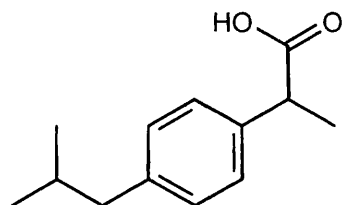
Vitamin C is water-soluble (hydrophilic) because it has several polar -OH groups that can form hydrogen bonds with water. Vitamin A is fat-soluble (hydrophobic) because most of its atoms can't form hydrogen bonds with water.

Visualizing Chemistry

2.20 Naphthalene has three resonance forms.



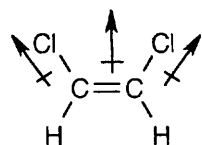
2.21



Ibuprofen

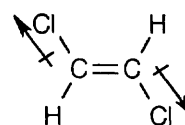
2.22 Electrostatic potential maps show that the electron-rich regions of the cis isomer lie on the same side of the double bond, leading to a net dipole moment. Because the electron-rich regions of the trans isomer are symmetrical about the double bond, the individual bond dipole moments cancel, and the isomer has no overall dipole moment.

net dipole moment



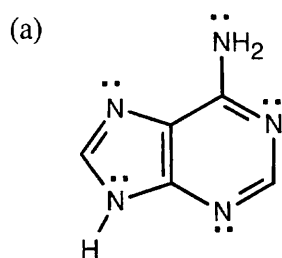
cis-1,2-Dichloroethylene

zero dipole moment

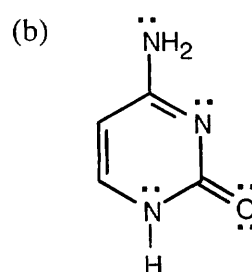


trans-1,2-Dichloroethylene

2.23



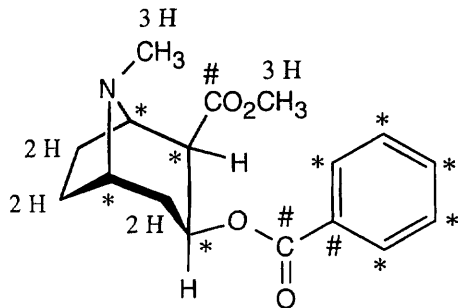
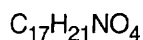
Adenine



Cytosine

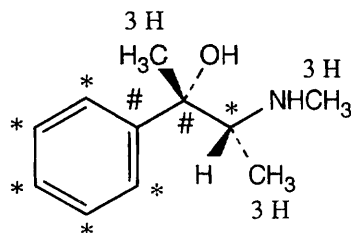
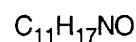
Additional Problems

2.24

**Cocaine**

= 0 hydrogens

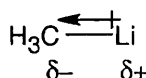
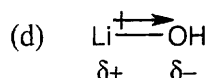
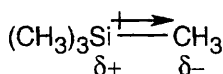
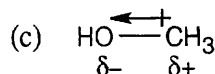
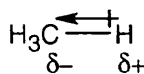
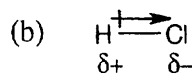
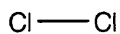
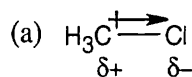
* = 1 hydrogen

**Ephedrine**

2.25 Use Figure 2.2 if you need help. The most electronegative element is starred.

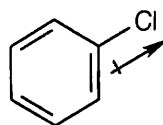
- (a) CH_2^*FCl (b) $FCH_2CH_2CH_2Br$ (c) $HOCH_2CH_2NH_2$ (d) $CH_3^*OCH_2Li$

2.26

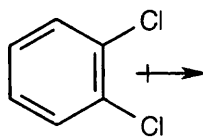
*More polar**Less polar*

2.27

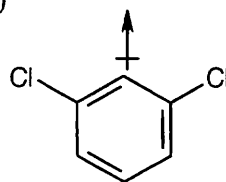
(a)



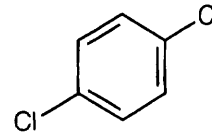
(b)



(c)



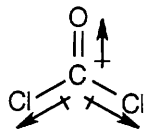
(d) no dipole moment



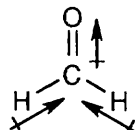
2.28 (a) In Section 2.2, we found that $\mu = Q \times r$. For a proton and an electron separated by 100 pm, $\mu = 4.8$ D. If the two charges are separated by 136 pm, $\mu = 6.53$ D.

(b) Since the observed dipole moment is 1.08 D, the H-Cl bond has $(1.08 \text{ D} / 6.53 \text{ D}) \times 100\% = 16.5\%$ ionic character.

- 2.29 In phosgene, the individual bond polarities tend to cancel, but in formaldehyde, the bond polarities add to each other. Thus, phosgene has a smaller dipole moment than formaldehyde.

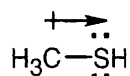


Phosgene

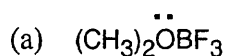


Formaldehyde

- 2.30 The magnitude of a dipole moment depends on both charge and distance between atoms. Fluorine is more electronegative than chlorine, but a C–F bond is shorter than a C–Cl bond. Thus, the dipole moment of CH₃F is smaller than that of CH₃Cl.
- 2.31 The observed dipole moment is due to the lone pair electrons on sulfur.

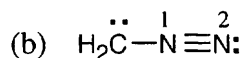


- 2.32 To save space, molecules are shown as line-bond structures with lone pairs, rather than as electron-dot structures.



$$\text{Oxygen: FC} = 6 - \frac{6}{2} - 2 = +1$$

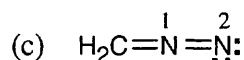
$$\text{Boron: FC} = 3 - \frac{8}{2} - 0 = -1$$



$$\text{Carbon: FC} = 4 - \frac{6}{2} - 2 = -1$$

$$\text{Nitrogen 1: FC} = 5 - \frac{8}{2} - 0 = +1$$

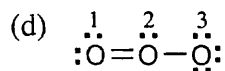
$$\text{Nitrogen 2: FC} = 5 - \frac{6}{2} - 2 = 0$$



$$\text{Carbon: FC} = 4 - \frac{8}{2} - 0 = 0$$

$$\text{Nitrogen 1: FC} = 5 - \frac{8}{2} - 0 = +1$$

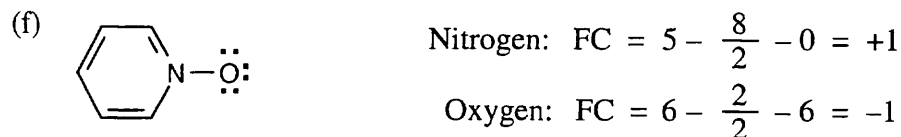
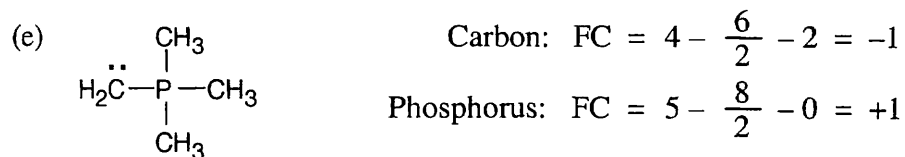
$$\text{Nitrogen 2: FC} = 5 - \frac{4}{2} - 4 = -1$$



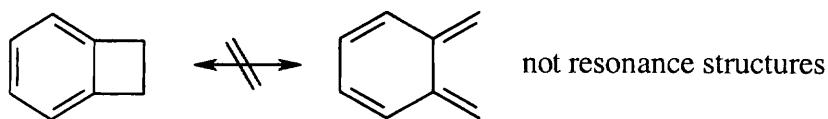
$$\text{Oxygen 1: FC} = 6 - \frac{4}{2} - 4 = 0$$

$$\text{Oxygen 2: FC} = 6 - \frac{6}{2} - 2 = +1$$

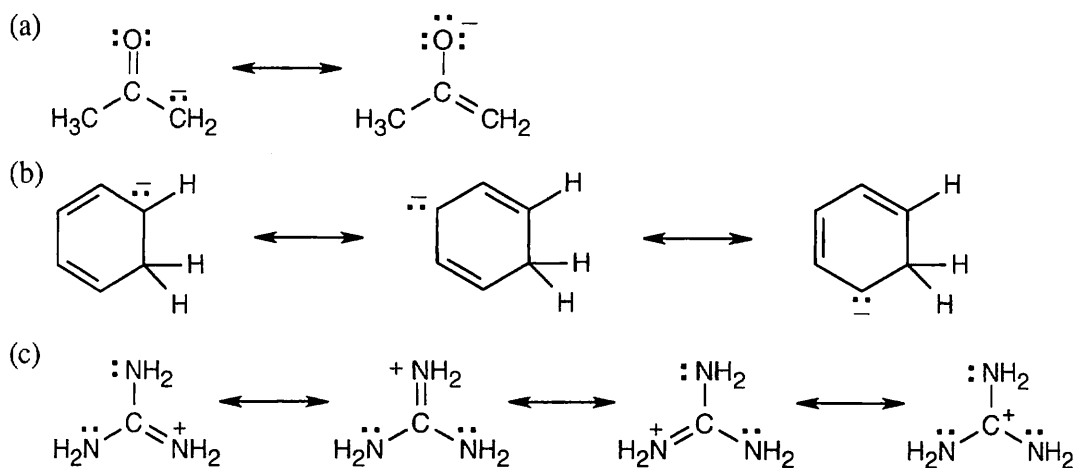
$$\text{Oxygen 3: FC} = 6 - \frac{2}{2} - 6 = -1$$



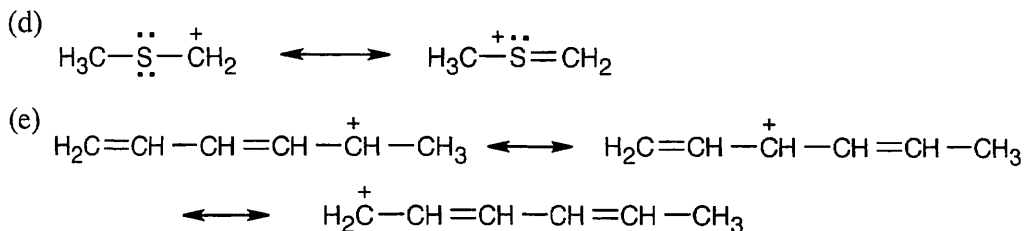
2.33 Resonance forms do not differ in the position of nuclei. The two structures in (a) are not resonance forms because the positions of the carbon and hydrogen atoms outside the ring are different in the two forms.



2.34 The pairs of structures in parts (b), (c), and (d) represent resonance forms.

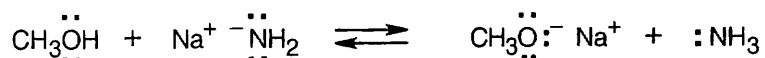
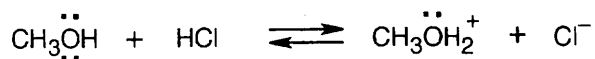


The last resonance structure is a minor contributor because its carbon lacks a complete electron octet.

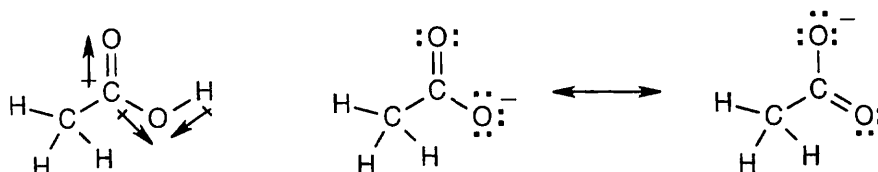


2.35 The two structures are not resonance forms because the positions of the carbon atoms are different in the two forms.

2.36



2.37



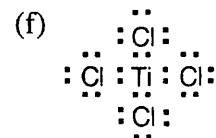
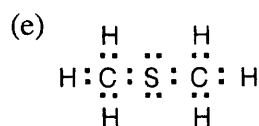
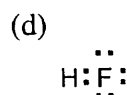
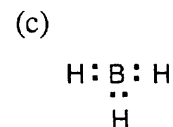
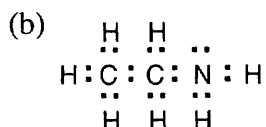
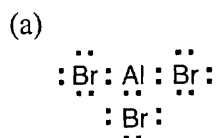
The O-H hydrogen of acetic acid is more acidic than the C-H hydrogens. The -OH oxygen is electronegative, and, consequently, the -O-H bond is more strongly polarized than the -C-H bonds. In addition, the acetate anion is stabilized by resonance.

2.38 The Lewis acids shown below can accept an electron pair either because they have a vacant orbital or because they can donate H^+ . The Lewis bases have nonbonding electron pairs.

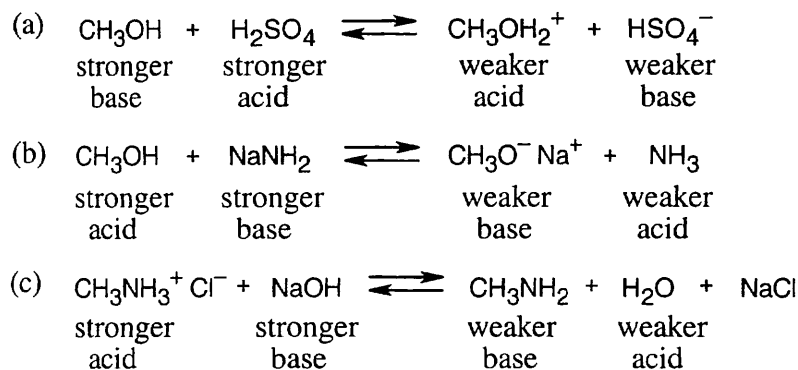
Lewis acids: AlBr_3 , BH_3 , HF , TiCl_4

Lewis bases: $\text{CH}_3\text{CH}_2\ddot{\text{N}}\text{H}_2$, $\text{H}_3\text{C}-\ddot{\text{S}}-\text{CH}_3$

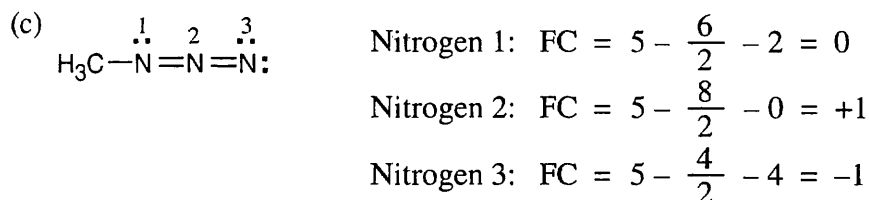
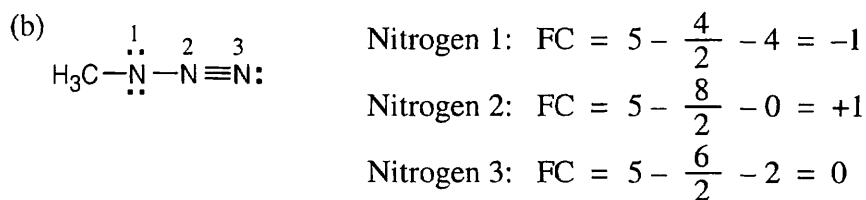
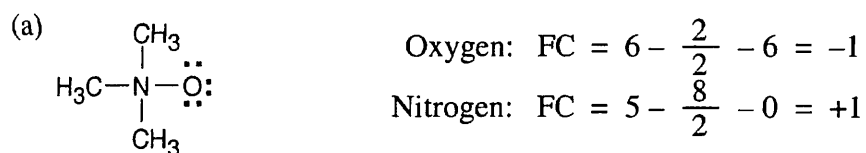
2.39



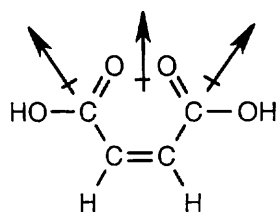
2.40



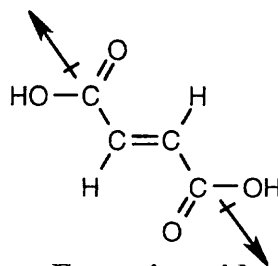
2.41 As in Problem 2.32, molecules are shown as line-bond structures with lone-pair electrons indicated. Only calculations for atoms with non-zero formal charge are shown.



2.42 In maleic acid, the individual dipole moments add to produce a net dipole moment for the whole molecule. The individual dipole moments in fumaric acid cancel, resulting in a zero dipole moment.

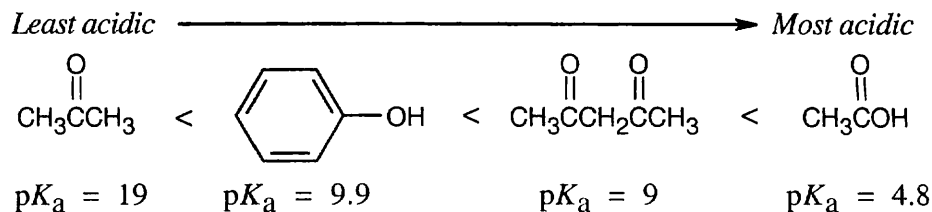


Maleic acid



Fumaric acid

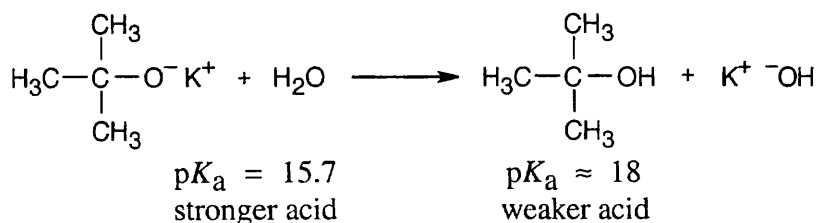
2.43 The substances with the largest values of pK_a are the least acidic.



2.44 To react completely (> 99.9%) with NaOH, an acid must have a pK_a at least 3 units smaller than the pK_a of H_2O . Thus, all substances in the previous problem except acetone react completely with NaOH.

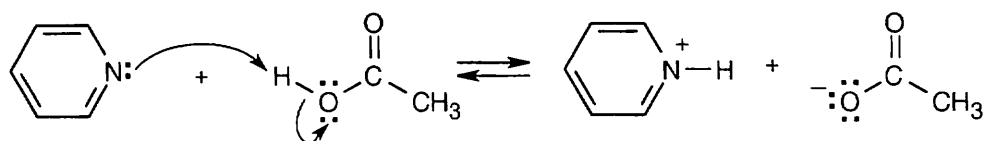
2.45 The stronger the acid (smaller pK_a), the weaker its conjugate base. Since NH_4^+ is a stronger acid than CH_3NH_3^+ , CH_3NH_2 is a stronger base than NH_3 .

2.46



The reaction takes place as written because water is a stronger acid than *tert*-butyl alcohol. Thus, a solution of potassium *tert*-butoxide in water can't be prepared.

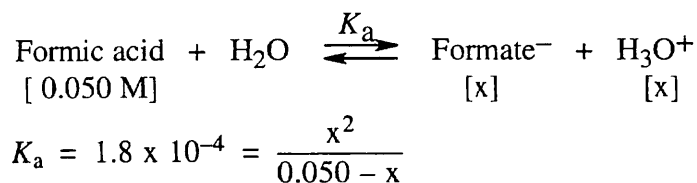
2.47



2.48 (a) Acetone: $K_a = 5 \times 10^{-20}$ (b) Formic acid: $K_a = 1.8 \times 10^{-4}$

2.49 (a) Nitromethane: $pK_a = 10.30$ (b) Acrylic acid: $pK_a = 4.25$

2.50

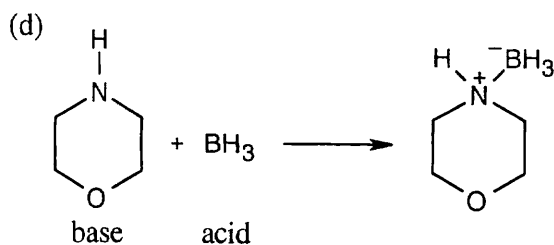
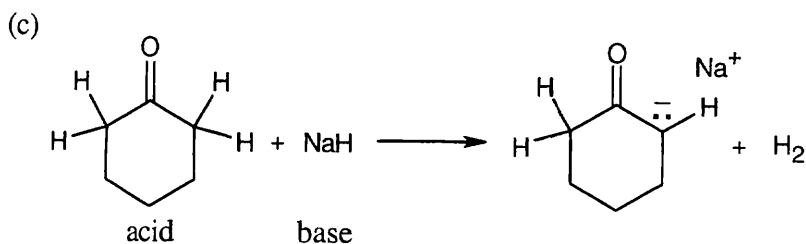
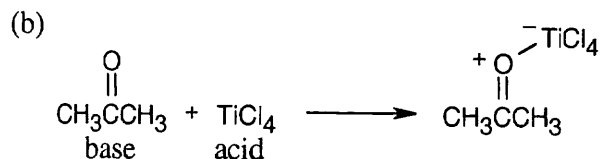
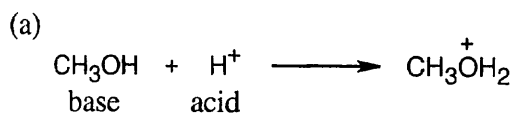


If you let $0.050 - x = 0.050$, then $x = 3.0 \times 10^{-3}$ and $\text{pH} = 2.52$. If you calculate x exactly, then $x = 2.9 \times 10^{-3}$ and $\text{pH} = 2.54$.

2.51 Only acetic acid will react with sodium bicarbonate. Acetic acid is the only substance in Problem 2.43 that is a stronger acid than carbonic acid.

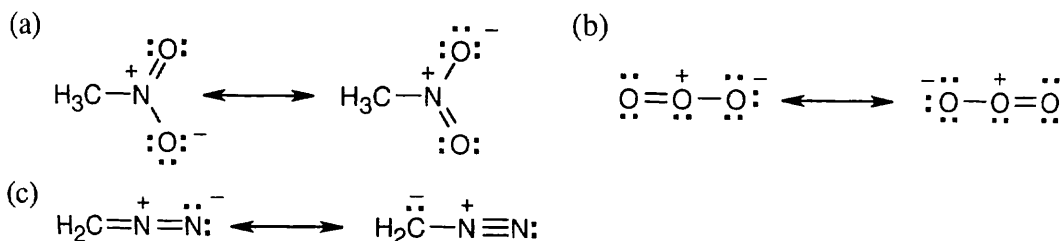
2.52 Sodium bicarbonate reacts with acetic acid to produce carbonic acid, which breaks down to form CO_2 . Thus, bubbles of CO_2 indicate the presence of an acid stronger than carbonic acid, in this case acetic acid. Phenol does not react with sodium bicarbonate.

2.53 Reactions (a) and (c) are reactions between Brønsted–Lowry acids and bases; the stronger acid and stronger base are identified. Reactions (b) and (d) occur between Lewis acids and bases.

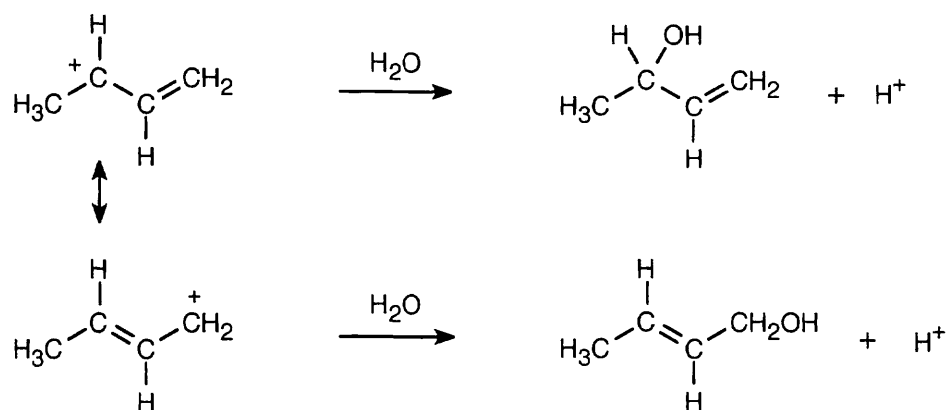


2.54 Pairs (a) and (d) represent resonance structures; pairs (b) and (c) do not. For two structures to be resonance forms, all atoms must be in the same positions in all resonance forms.

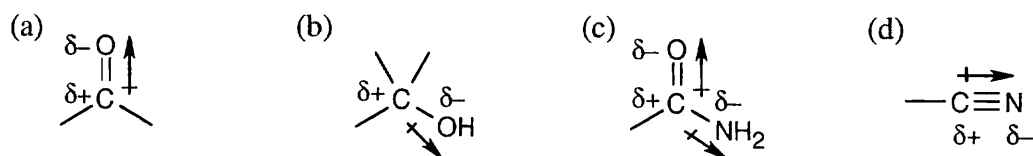
2.55



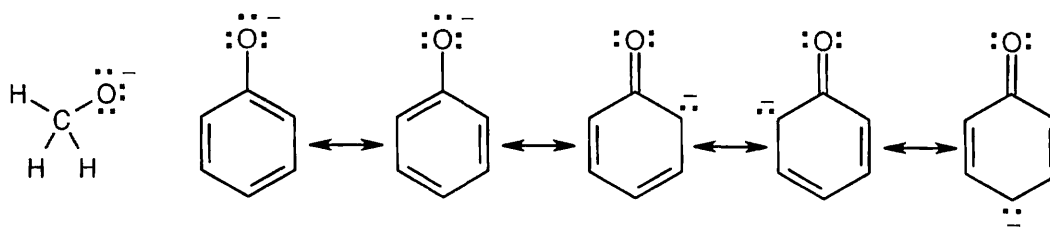
2.56 The cation pictured can be represented by two resonance forms. Reaction with water can occur at either positively charged carbon, resulting in two products.



2.57



2.58



When phenol loses a proton, the resulting anion is stabilized by resonance.

Review Unit 1: Bonds and Bond Polarity

Major Topics Covered (with vocabulary:)

Atomic Structure:

atomic number mass number wave equation orbital shell node electron configuration

Chemical Bonding Theory:

covalent bond Lewis structure lone-pair electrons line-bond structure valence-bond theory
sigma (σ) bond bond strength bond length molecular orbital theory bonding MO
antibonding MO

Hybridization:

sp^3 hybrid orbital bond angle sp^2 hybrid orbital pi (π) bond sp hybrid orbital

Polar covalent bonds:

polar covalent bond electronegativity (EN) electrostatic potential maps inductive effect
dipole moment formal charge dipolar molecule

Resonance:

resonance form resonance hybrid

Acids and Bases:

Brønsted-Lowry acid Brønsted-Lowry base conjugate acid conjugate base acidity constant
 K_a pK_a organic acid organic base Lewis acid Lewis base

Chemical Structures:

condensed structure skeletal structure space-filling models ball-and-stick models

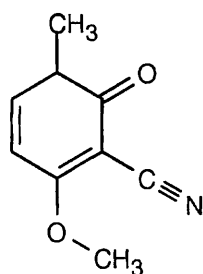
Types of Problems:

After studying these chapters you should be able to:

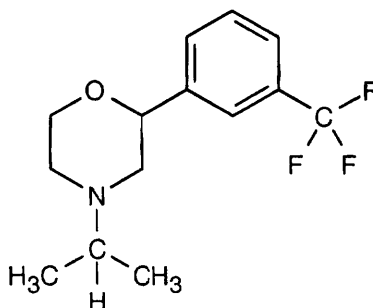
- Predict the ground state electronic configuration of atoms.
- Draw Lewis electron-dot structures of simple compounds.
- Predict and describe the hybridization of bonds in simple compounds.
- Predict bond angles and shapes of molecules.
- Predict the direction of polarity of a chemical bond, and predict the dipole moment of a simple compound.
- Calculate formal charge for atoms in a molecule.
- Draw resonance forms of molecules.
- Predict the relative acid/base strengths of Brønsted acids and bases.
- Predict the direction of Brønsted acid/base reactions.
- Calculate: pK_a from K_a , and vice versa.
pH of a solution of a weak acid.
- Identify Lewis acids and bases.
- Draw chemical structures from molecular formulas, and vice versa.

Points to Remember:

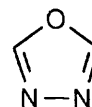
- * In order for carbon, with valence shell electron configuration of $2s^2 2p^2$, to form four sp^3 hybrid orbitals, it is necessary that one electron be promoted from the $2s$ subshell to the $2p$ subshell. Although this promotion requires energy, the resulting hybrid orbitals are able to form stronger bonds, and compounds containing these bonds are more stable.
- * Assigning formal charge to atoms in a molecule is helpful in showing where the electrons in a bond are located. Even if a bond is polar covalent, in some molecules the electrons "belong" more to one of the atoms than the other. This "ownership" is useful for predicting the outcomes of chemical reactions, as we will see in later chapters.
- * Resonance structures are representations of the distribution of π and nonbonding electrons in a molecule. Electrons don't move around in the molecule, and the molecule doesn't change back and forth, from structure to structure. Rather, resonance structures are an attempt to show, by conventional line-bond drawings, the electron distribution of a molecule that can't be represented by any one structure.
- * As in general chemistry, acid-base reactions are of fundamental importance in organic chemistry. Organic acids and bases, as well as inorganic acids and bases, occur frequently in reactions, and large numbers of reactions are catalyzed by Brønsted acids and bases and Lewis acids and bases.

Self-Test:

A
Ricine
(a toxic component
of castor beans)



B
Oxaflozane
(an antidepressant)



C
1,3,4-Oxadiazole

For **A** (ricinine) and **B** (oxaflozane): Add all missing electron lone pairs. Identify the hybridization of all carbons. Indicate the direction of bond polarity for all bonds with $\Delta \text{EN} \geq 0.5$. In each compound, which bond is the most polar? Convert **A** and **B** to molecular formulas.

Draw a resonance structure for **B**. Which atom (or atoms) of **B** can act as a Lewis base?

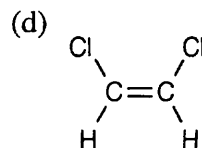
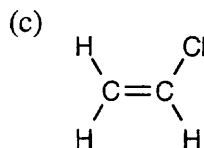
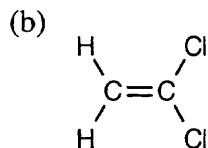
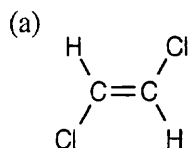
Add missing electron lone pairs to **C**. Is it possible to draw resonance forms for **C**? If so, draw at least one resonance form, and describe it.

Multiple Choice:

- Which element has $4s^2 4p^2$ as its valence shell electronic configuration?
(a) Ca (b) C (c) Al (d) Ge
- Which compound (or group of atoms) has an oxygen with a +1 formal charge?
(a) NO_3^- (b) O_3 (c) acetone anion (d) acetate anion

The following questions involve these acids: (i) HW ($\text{p}K_a = 2$); (ii) HX ($\text{p}K_a = 6$); (iii) HY ($\text{p}K_a = 10$); (iv) HZ ($\text{p}K_a = 20$).

- Which of the above acids react almost completely with water to form hydroxide ion?
(a) none of them (b) all of them (c) HY and HZ (d) HZ
- The conjugate bases of which of the above acids react almost completely with water to form hydroxide ion?
(a) none of them (b) all of them (c) HZ (d) HY and HZ
- If you want to convert HX to X^- , which bases can you use?
(a) W^- (b) Y^- (c) Z^- (d) Y^- or Z^-
- If you add equimolar amounts of HW, X^- and HY to a solution, what are the principal species in the resulting solution?
(a) HW, HX, HY (b) W^- , HX, HY (c) HW, X^- , HY (d) HW, HX, Y^-
- What is the approximate pH difference between a solution of 1 M HX and a solution of 1 M HY?
(a) 2 (b) 3 (c) 4 (d) 6
- If you wanted to write the structure of a molecule that shows carbon and hydrogen atoms as groups, without indicating many of the carbon-hydrogen bonds, you would draw a:
(a) molecular formula (b) Kekulé structure (c) skeletal structure (d) condensed structure
- Which of the following molecules has zero net dipole moment?



- In which of the following bonds is carbon the more electronegative element?
(a) C—Br (b) C—I (c) C—P (d) C—S