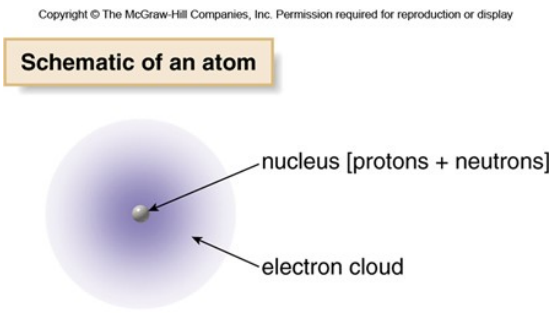
Unit *1: Structure and AB Overview*

Part 1 Notes: *Periodic Table & Bonding*

What is Organic Chem?

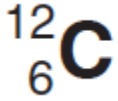


The Atom and the Periodic Table.

- Nucleus
- Electron Cloud
- Cations
- Anions

Atomic Structure

- Atomic Number:



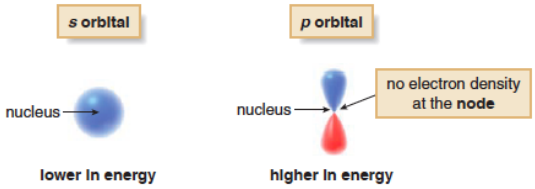
- Atomic Mass:

- In a neutral atom:

    - For Ions:
        - Cations:
        - Anions

Isotopes

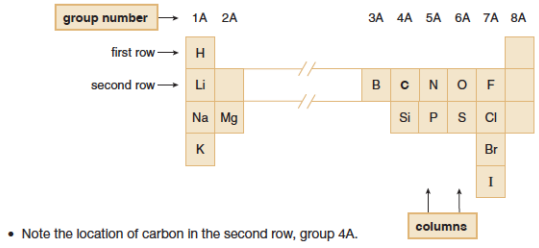
Atomic Orbitals:



Orbital Shapes (Applicable to OChem)

Periodic Table

- Elements in the same row are



- Elements in the same column

The First Row:

The Second Row:

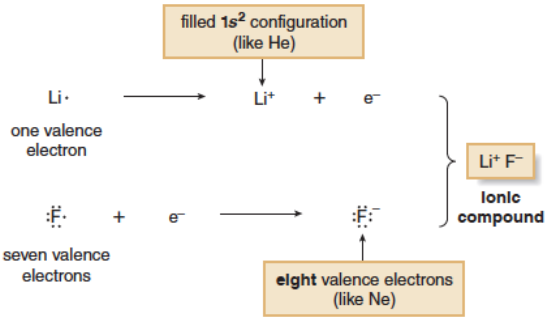
Example

1. While the most common isotope of nitrogen has a mass number of 14 (nitrogen-14), a radioactive isotope of nitrogen has a mass number of 13 (nitrogen-13). Nitrogen-13 is used in PET (positron emission tomography) scans by physicians to monitor brain activity and diagnose dementia. For each isotope, give the following information:
    a. The number of protons
    b. The number of neutrons
    c. The number of electrons in the neutral atom
    d. The group number

Bonding

- Bonding is:

- Occurs because:



Ionic Bonding

Covalent Bonding

- Results from

- Referred to as a

Example

1. Label each bond in the following compounds as ionic or covalent
    a. F_2
    b. LiBr
    c. CH_3CH_3
    d. $NaNH_2$

Valence Electrons

- Outermost electrons
- Loosely held
- Participate in chemical reactions
- Fore Second row elements…

predicted number of bonds = 8 − **number of valence electrons**

Octet Rule: 2 important exceptions

Example

1. How many covalent bonds are predicted for each atom?
    a. O

    b. Al

    c. Br

    d. Si

Part 1 Practice: *Periodic Table & Bonding*

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1. Two radioactive isotopes of iodine used for the diagnosis and treatment of thyroid disease have mass numbers of 123 and 131. For each isotope, give the following information:
    a. The number of protons

    b. The number of neutrons

    c. The number of electrons (in the neutral atom)

    d. The group number

2. Label each bond in the following compounds as ionic or covalent
    a. NaI
    b. BrCl
    c. HCl
    d. CH_3NH_2
    e. $NaOCH_3$

Part 2 Notes: *Lewis Structures, Resonance, and Shapes*

Lewis Structures

- Electron dot representation
- Three general rules:

Practice

Draw the lewis structure for:

CH_4

CH_4O

CH_3CH_3

CH_5N

CH_2CH_2

C_2H_2

HCN

$HOCH_2CO_2H$

Formal Charge

formal charge = **number of valence electrons** − **number of electrons an atom "owns"**

number of electrons owned = **number of unshared electrons** + $\frac{1}{2}$ [**number of shared electrons**]

Practice

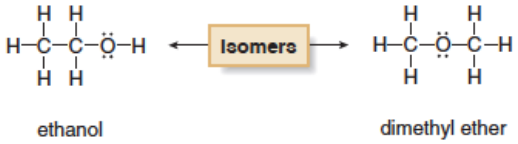
Determine the formal charge on each atom of H_3O^+

Determine the formal charge on N in NH_4^+

Determine the formal charge on N in CH_3NC

Isomers

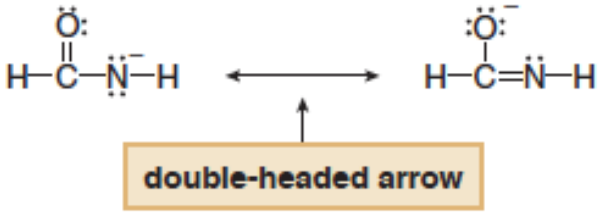
- Isomers are….
  - The atoms are connected in a different order!



same molecular formula
C_2H_6O

Resonance

- How are they different than isomers?



- Major & minor contributors

X
more bonds
fewer charges
major contributor

Y
minor contributor

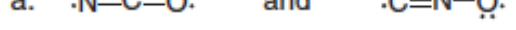
3 principles of resonance theory

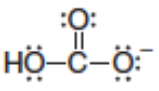
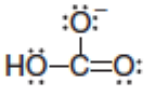
1. Resonance structures are not REAL. An individual resonance structure does not represent the structure of the molecule or ion. Only the hybrid does.

2. Resonance structures are not in equilibrium with each other. There is no movement of electrons from one form to another.
3. Resonance structures are not isomers. Two isomers different in the arrangement of atoms & electrons, whereas resonance structures differ only in the arrangement of the electrons

Practice

Classify each pair as isomers or resonance

a.  and

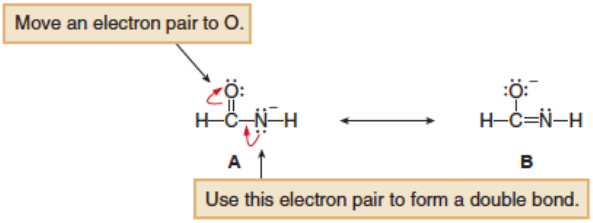
b.  and 

Drawing resonance structures

1. Two resonance structures differ in the position of multiple bonds and nonbonded electrons
2. Two resonance structures must always have the same number of unpaired electrons
3. Must be valid Lewis structures

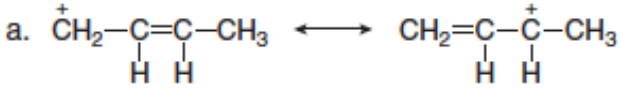
Curved Arrow Notation

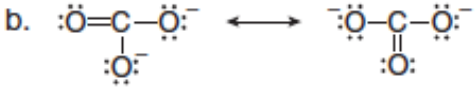
- Curved arrows show
- Head points to
- Tail starts where
- Always use a



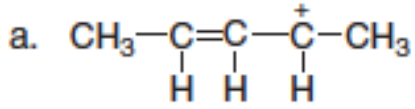
Examples

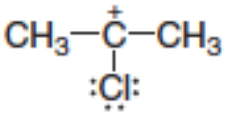
Use a curved arrow notation to show how the first resonance structure can be converted into the second.

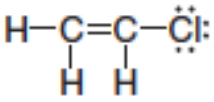
a. 

b. 

Draw a second resonance structure for each species

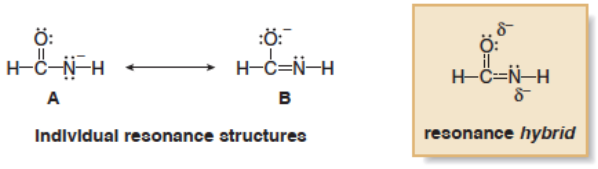
a. 

b. 

c. 

The resonance Hybrid

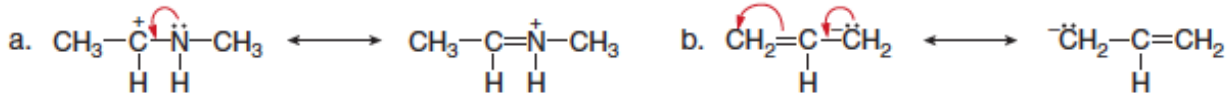
- Why is the hybrid more stable than any given resonance structure?



- How do you know which structure is better?

Examples

Label the resonance structures in each pair as major, minor, or equal contributors to the hybrid. Then draw the hybrid.



Bond Length

- Bond length _____ across a row because
- Bond leant _____ down a column because

Table 1.2 Average Bond Lengths

Bond	Length (pm)	Bond	Length (pm)	Bond	Length (pm)
H–H	74	H–F	92	C–F	133
C–H	109	H–Cl	127	C–Cl	177
N–H	101	H–Br	141	C–Br	194
O–H	96	H–I	161	C–I	213

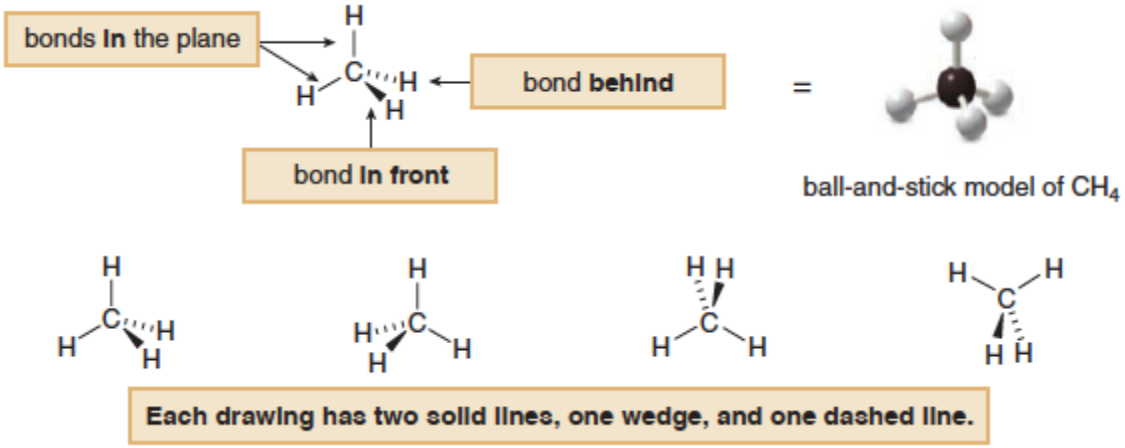
VSEPR For Organic

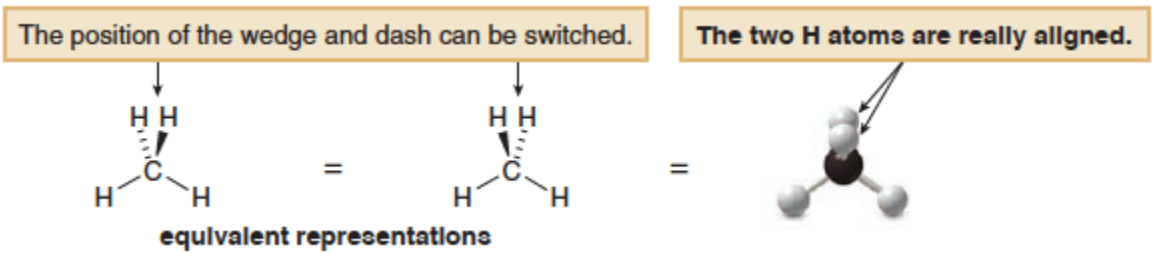
- VSEPR: the most stable arrangement of electrons keeps them…

Number of groups	Geometry	Bond angle
• two groups	linear	180°
• three groups	trigonal planar	120°
• four groups	tetrahedral	109.5°

3D Structures

- Solid line for a bond IN the plane
- Wedge for a bond coming OUT of the plane
- Dashed line for a bond going INTO the plane





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EXAMPLES

Draw two different 3-D representation for CH_2Cl_2 using wedges and dashes

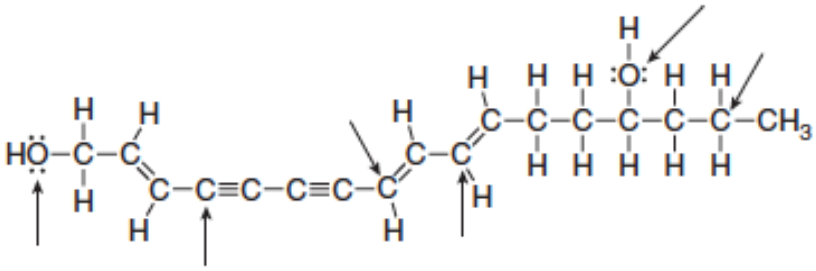
Determine the geometry around all second row elements in each compound

a. $CH_3-\overset{\overset{:O:}{\|}}{C}-CH_3$ b. $CH_3-\ddot{O}-CH_3$ c. $\ddot{N}H_2^-$ d. $CH_3-C\equiv N:$

Predict the indicated bond angles in each compound

a. $CH_3-C\equiv C-Cl$ b. $CH_2=C-Cl$ c. $CH_3-C(H)(H)-Cl$

Predict the geometry around the indicated atoms in enanthotoxin



enanthotoxin

Part 2 Practice: *Lewis Structures, Resonance, and Shapes*

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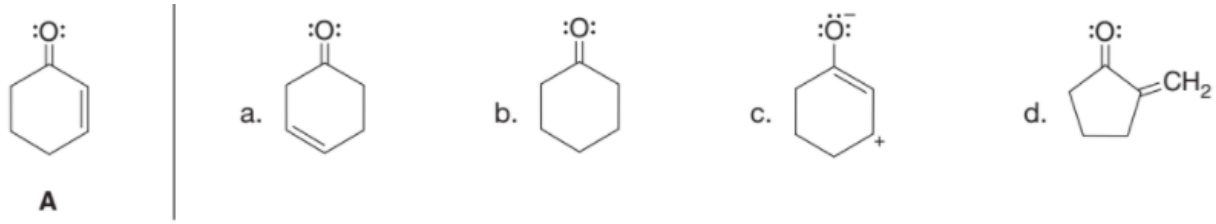
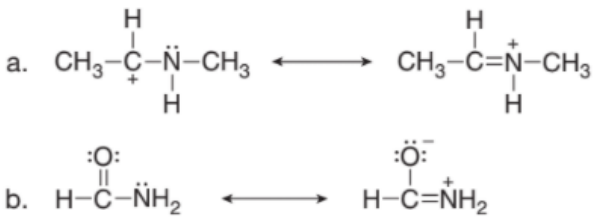
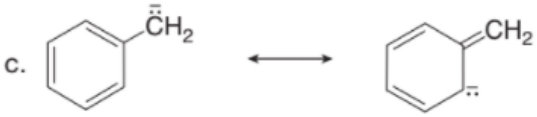
1. Draw all possible isomers for C_3H_7Cl (two isomers)

2. With Reference to species A, label each compound as an isomer, a resonance structure, or neither.

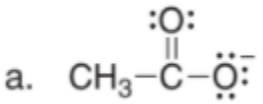


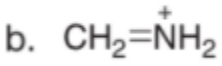
3. Add curved arrows to show how the first resonance structure can be converted into the second.





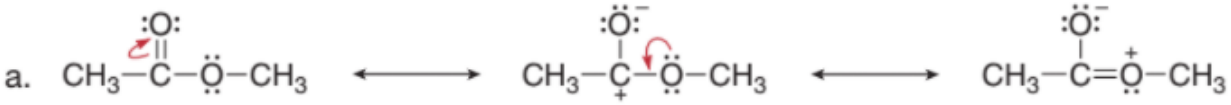
4. Draw a second resonance structure for each ion.





5. For (a) and (b) above, draw the resonance hybrid.

6. Rank the resonance structures in order of increasing contribution to the resonance hybrid.



7. Predict all bond angles in these compounds

a. CH_3Cl

b. $CH_2{=}NCH_3$

8. Predict the geometry around each indicated atom

a. $CH_3CH_2\underset{\uparrow}{C}H_2CH_3$

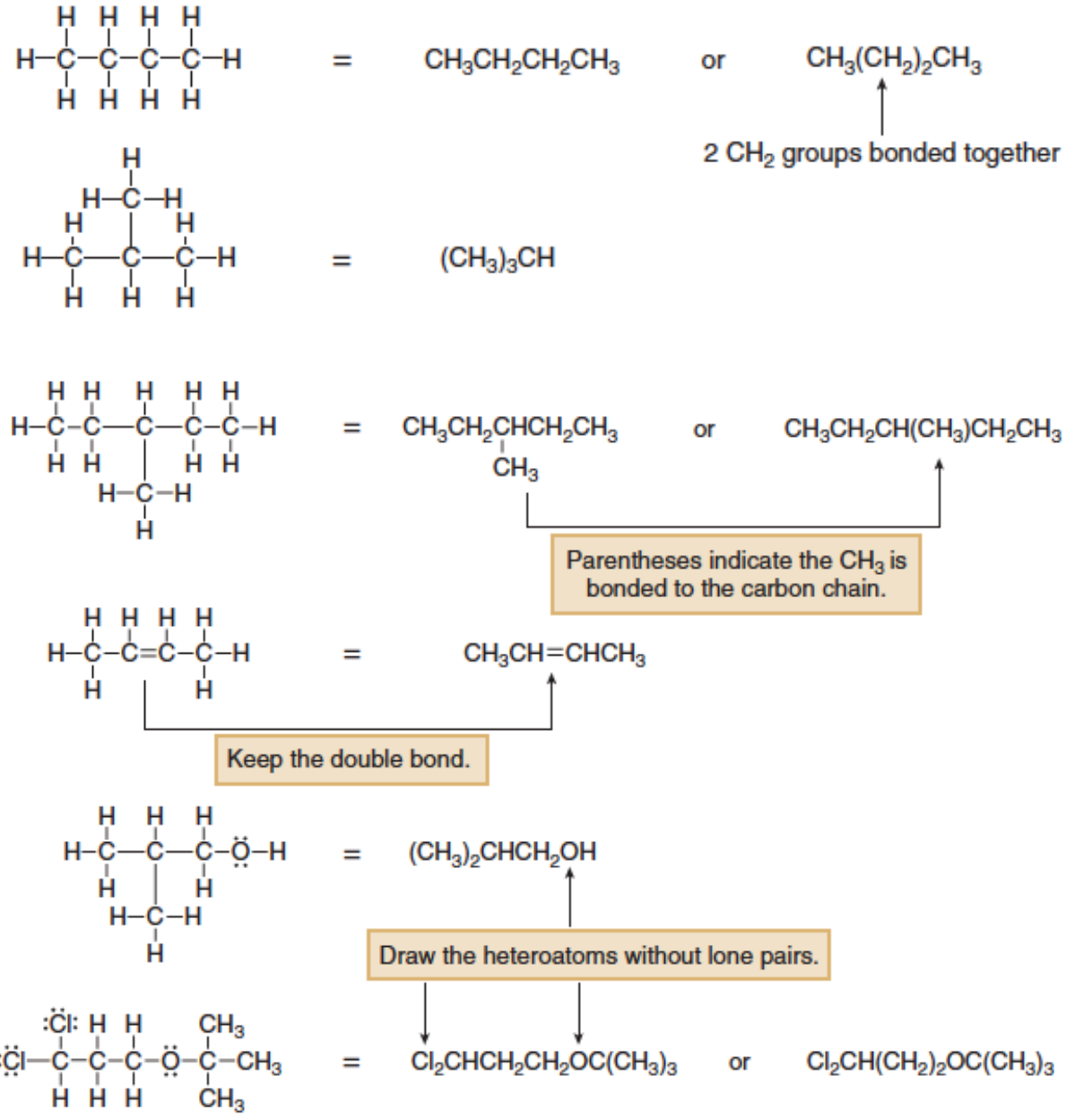
b. $(CH_3)_2\underset{\uparrow}{N}^-$

9. Draw the structure of halothane, $CF_3CHClBr$, in three dimensions, using solid lines, wedges, and dashes to illustrate the position of atoms. Halothane is a safe and widely used anesthetic.

Part 3 Notes: *Drawing Organic Structures*

Condensed Structures

- All of the atoms are drawn in, but the two-electron bond lines are generally omitted
- Atoms are usually drawn next to the atoms to which they are bonded
- Parentheses are used around similar groups bonded to the same atom
- Lone pairs are omitted



Examples

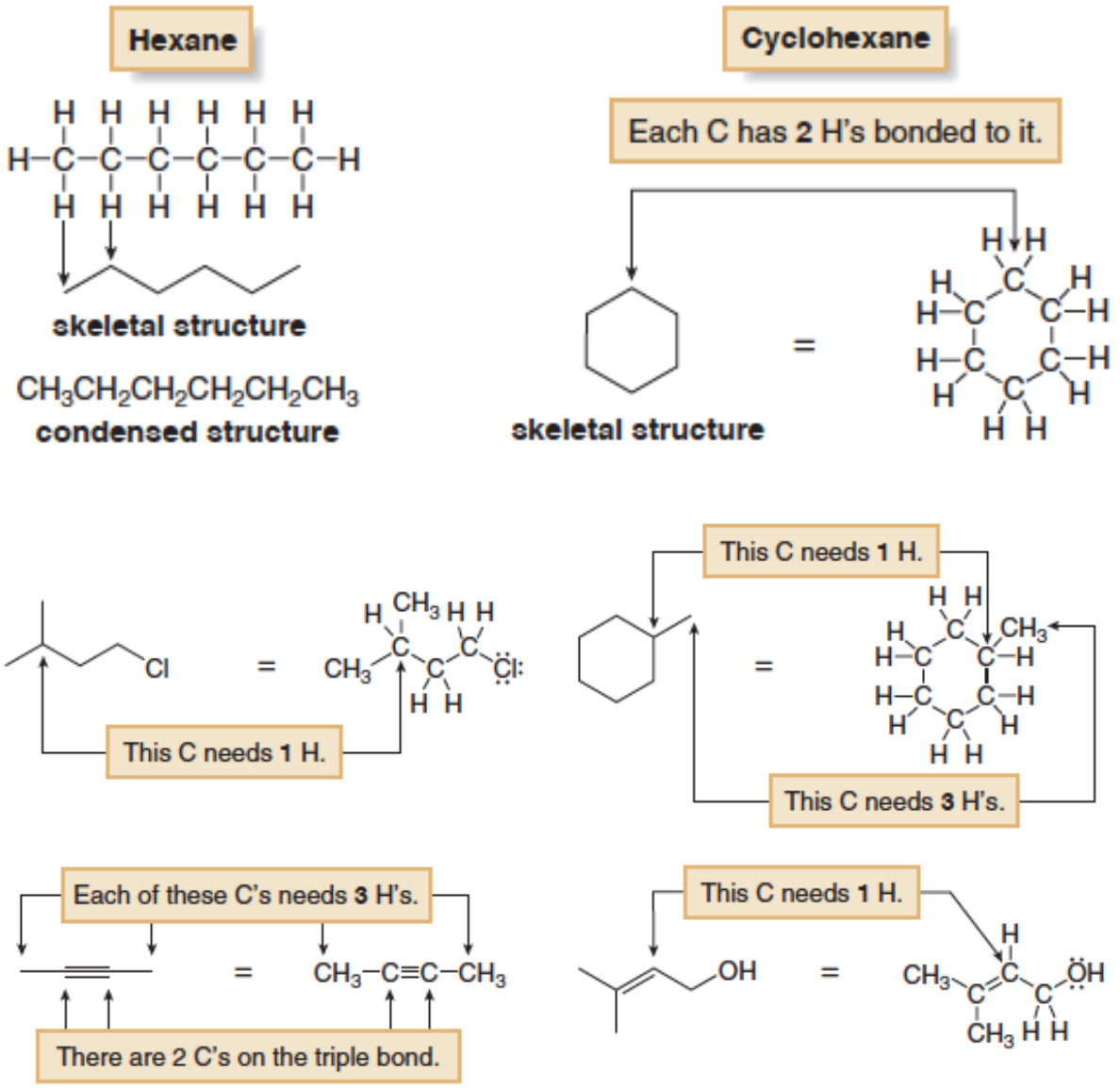
Convert each condensed formula to a Lewis Structure

a. $(CH_3)_2CHOCH_2CH_2CH_2OH$ b. $CH_3(CH_2)_2CO_2C(CH_3)_3$

During periods of strenuous exercise, the buildup of lactic acid [$CH_3CH(OH)CO_2H$] causes the aching feeling in sore muscles. Convert this condensed structure to a Lewis structure of lactic acid.

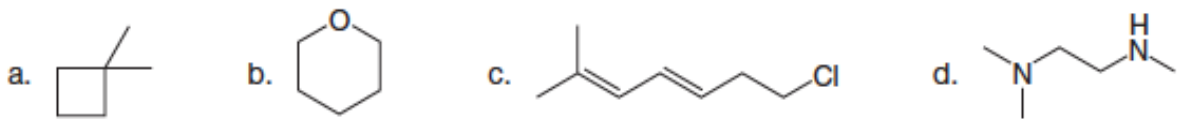
Skeletal structures

- Assume there is a carbon atom at the junction of any two lines or at the end of any line
- Assume there are enough hydrogens around each carbon to make it tetravalent
- Draw in all heteroatoms and the hydrogens directly bonded to them.



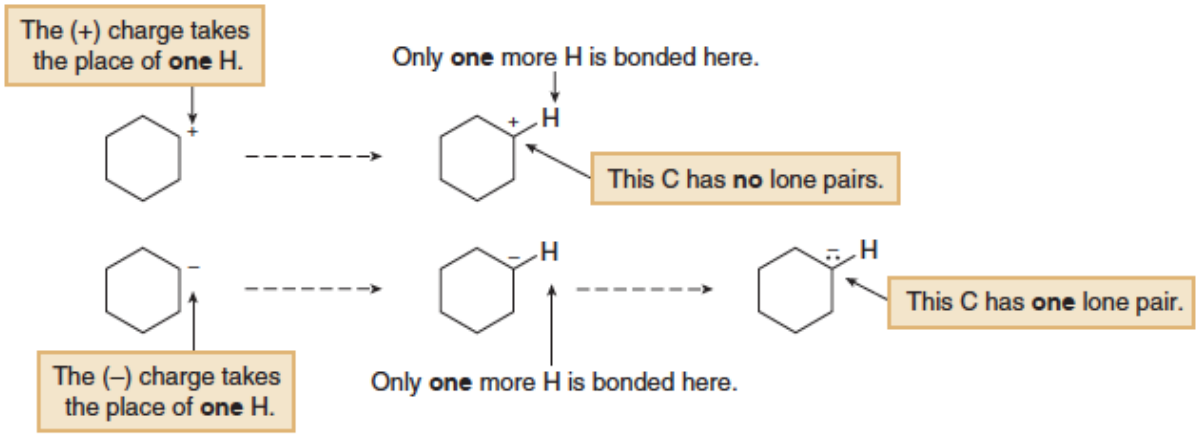
Examples

Convert each skeletal structure to a Lewis structure

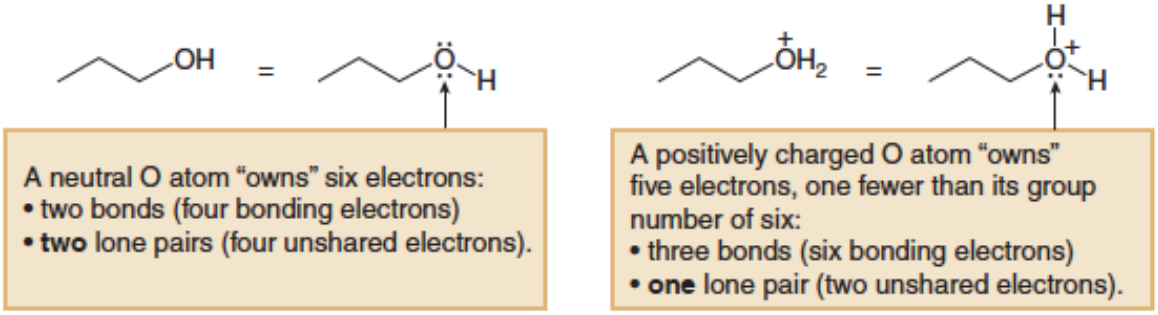


Skeletal structures with charged carbon atoms

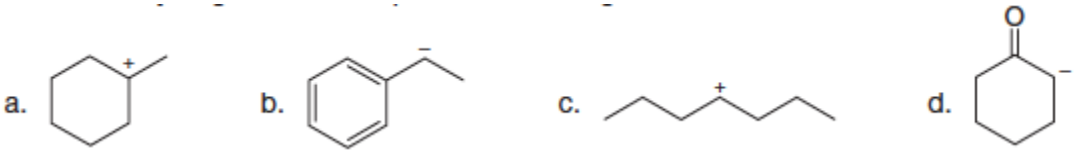
- A charge on a carbon takes the place of a hydrogen
- The charge determines the number of lone pairs. Negatively charged carbon atoms have on pair, and positively charged carbon atoms have none.



Lone pairs on Heteroatoms!

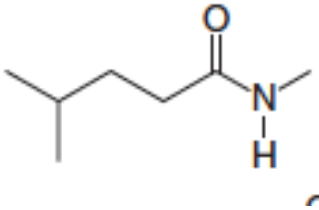


Examples
Draw in all hydrogens and lone pairs on the charged carbons in each ion



Draw a skeletal structure for molecule (a) and a condensed structure for molecule ©

a. $(CH_3)_2C{=}CH(CH_2)_4CH_3$



Part 3 Practice: *Drawing Organic Structures*

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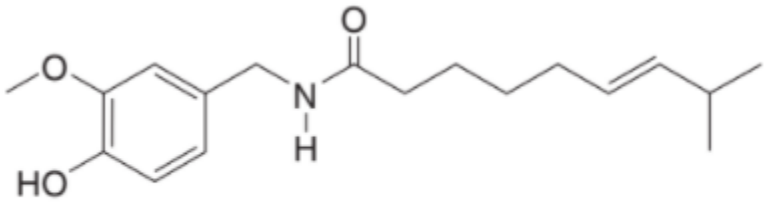
1. How many carbon atoms are in capsaicin? How many hydrogens are present around each carbon atom in capsaicin?



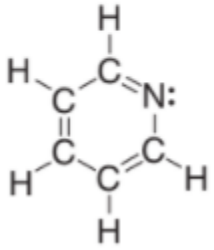
capsaicin
(spicy component of hot peppers)

2. Convert each molecule into a skeletal structure.

   a. $(CH_3)_2CHCH_2CH_2CH(CH_3)_2$

   b. $CH_3CH(Cl)CH(OH)CH_3$

   c. $(CH_3)_3C(CH_2)_5CH_3$

   d. 

3. Convert the following condensed formulas into Lewis structures

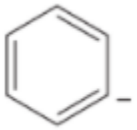
   a. CH_3CH_2COOH

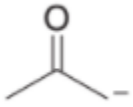
   b. $CH_3CONHCH_3$

   c. CH_3COCH_2Br

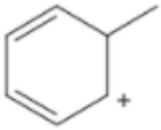
   d. $(CH_3)_3COH$

4. Draw in all H atoms and nonbonded electron pairs in each ion.

   a. 

   b. 

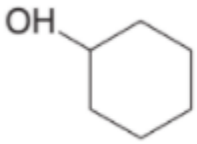
   c. 

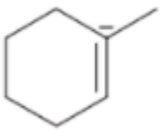
   d. 

5. Each of the following condensed or skeletal structures is an incorrect representation of a molecule or ion. Explain what is wrong in each structure.

   a. $CH_3CH=CH=CHCH_3$

   b. $(CH_3)_3CHCH_2CH_2CH_3$

   c. 

   d. 

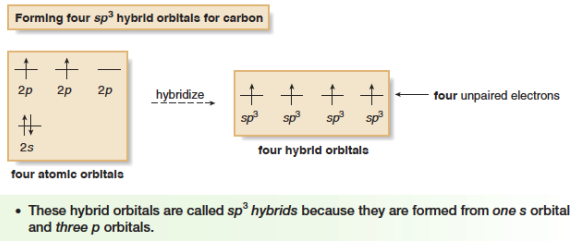
Part 4 Notes: *Hybridization - Bond Length and Strength*

Hybridization in Hydrogen

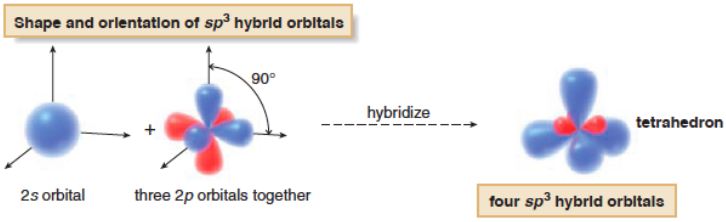
- Sigma Bonds:

Hybridization in Methane

- Carbon ground state makes only two bonds!
  - No octet rule
  - Must hybridize



- These hybrid orbitals are called sp^3 *hybrids* because they are formed from *one* s orbital and *three* p orbitals.

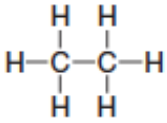


number of groups around an atom	number of orbitals used	type of hybrid orbital
2	2	two *sp* hybrid orbitals
3	3	three sp^2 hybrid orbitals
4	4	four sp^3 hybrid orbitals

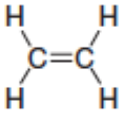
What orbitals are used to form each of the following?

a. CH_3BeH b. $(CH_3)_3B$ c. CH_3OCH_3

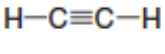
Ethane, Ethylene, Acetylene



ethane



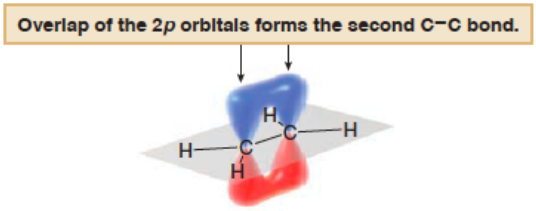
ethylene



acetylene

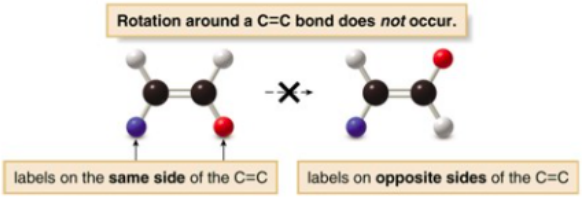
Sigma and Pi Bonds

- Sigma bond formed by
- Pi bond formed by



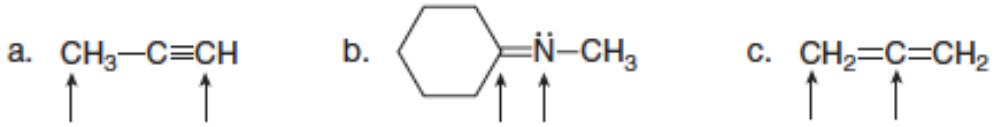
Bond Rotation

- Sigma bonds allow for free rotation
- Pi bonds DO NOT

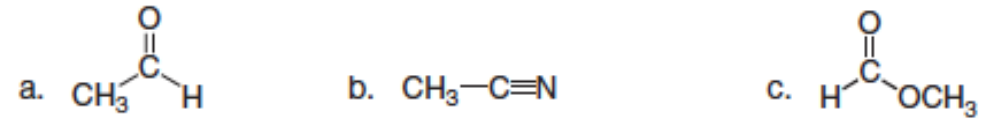


Example

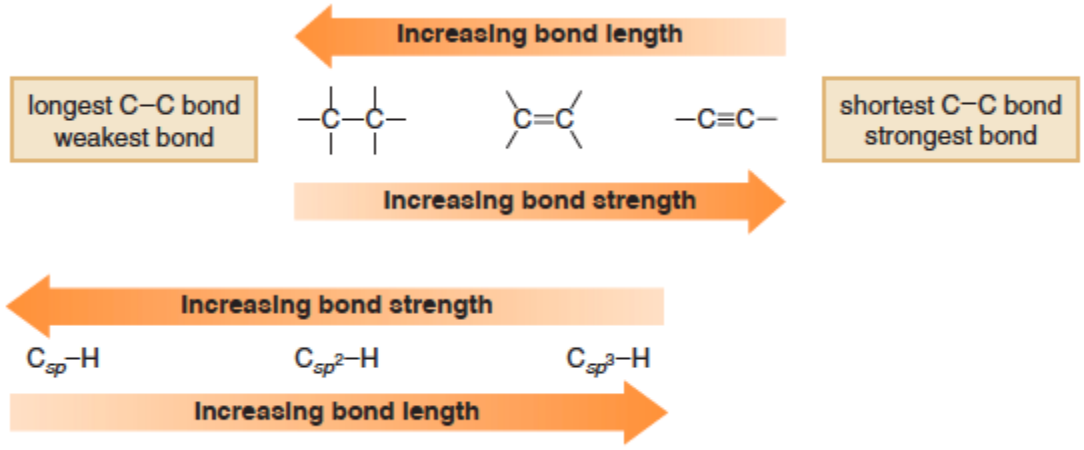
Determine the hybridization around the indicated atoms in the following molecules



Classify each bond in the following molecules as sigma or pi

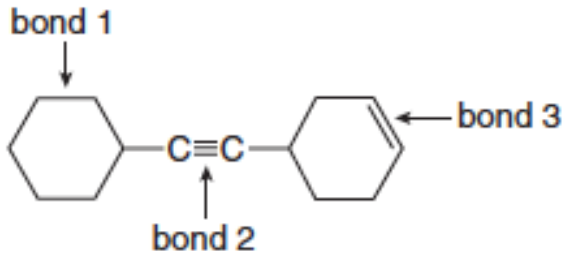


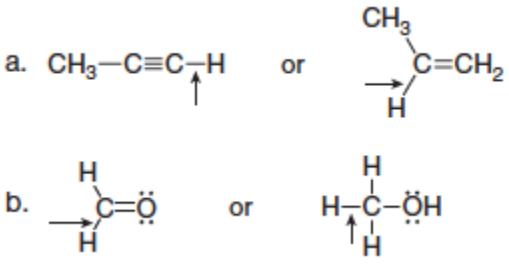
Bond Length and Bond Strength



Example

Rank the indicated bonds in order of increasing bond strength, and increasing bond length



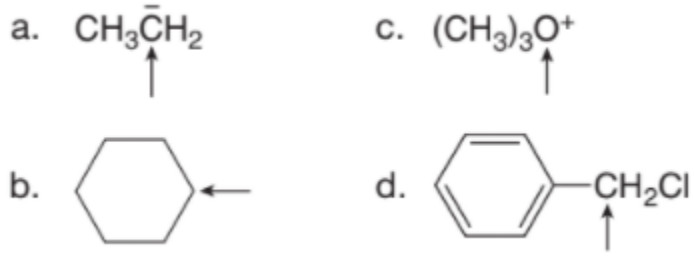


Which of the indicated bonds in each pair is shorter?

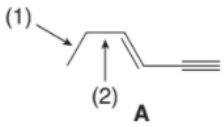
Part 4 Practice: *Hybridization - Bond Length and Strength*

---

1. Predict the hybridization and geometry around each indicated atom



2. Answer the following questions about compound A

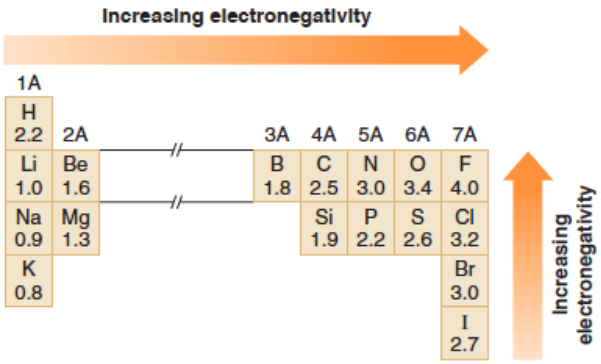


a. Label the shortest C–C single bond.
b. Label the longest C–C single bond.
c. Considering all the bonds, label the shortest C–C bond.
d. Label the weakest C–C bond.
e. Label the strongest C–H bond.
f. Explain why bond (1) and bond (2) are different in length, even though they are both C–C single bonds.

Part 5 Notes: *Polarity*

Electronegativity

- Measure of
- Increases as
- Decreases as

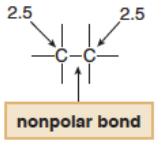


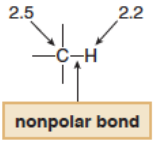
Rank the following in order of increasing electronegativity

a. Se, O, S b. P, Na, Cl c. Cl, S, F d. O, P, N

Nonpolar bonds

- Made by

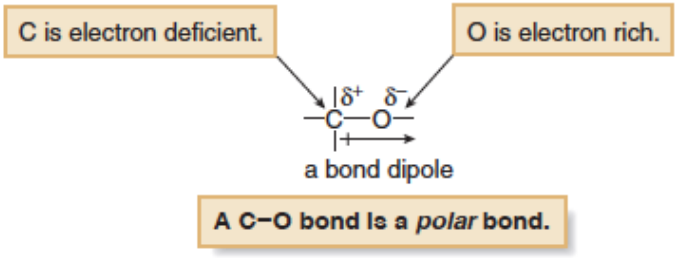




The small electronegativity difference between C and H is ignored.

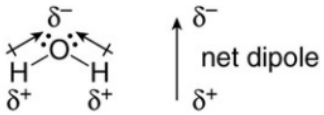
Polar bonds

- Bonding between atoms that cause



Polarity of molecules

- Vector of the dipole: decide if individual dipoles cancel or reinforce each other

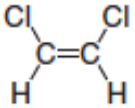
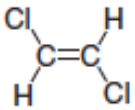


The net dipole bisects the H–O–H bond angle.
The two individual dipoles reinforce.

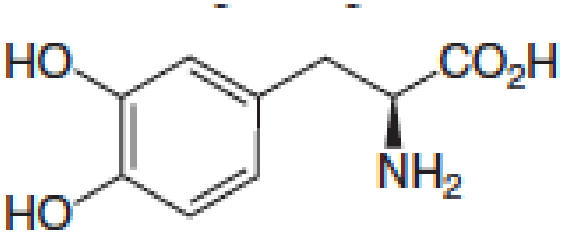
H_2O is a polar molecule.

Example

Indicate which of the following molecules is polar. Show the direction

a. CH_3Br b. CH_2Br_2 c. CF_4 d.  e. 

Big Example Time!
Here is L-Dopa (a drug used to treat Parkinson's).



L-dopa

Do the following

1. Convert to a Lewis Structure
2. Determine the hybridization around all atoms
3. Label 2 polar and 2 nonpolar bonds
4. Compare bond length and strength
5. Draw another resonance structure

Part 5 Practice: *Polarity*

---

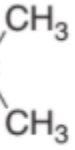
1. Answer each question with a brief explanation and an example to illustrate your answer.
   a. Can a compound be nonpolar if it contains one polar bond?

   b. Can a compound be nonpolar if it contains two or more polar bonds?

   c. Can a compound be polar if it contains no polar bonds?

2. Label the polar bonds in each molecule. Indicate the direction of the net dipole (if there is one).

a. $CHBr_3$
b. $CH_3CH_2OCH_2CH_3$
c. CBr_4

d. 

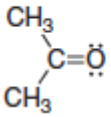
Part 6 Notes: *Bronsted-Lowry Acids and Bases*

Brønsted–Lowry acids [H – A]

Inorganic	Organic
HCl H_2SO_4 HSO_4^- H_2O H_3O^+	CH_3CO_2H acetic acid $HO_2CCH_2-C(OH)(COOH)-CH_2CO_2H$ citric acid

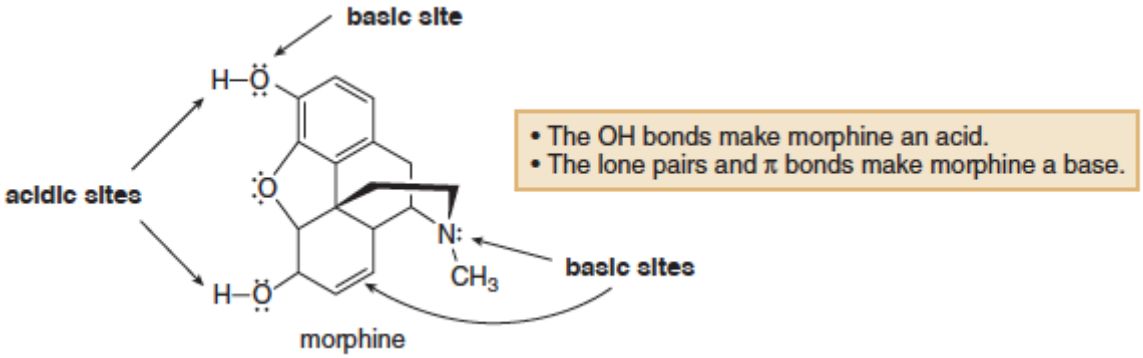
- All Brønsted–Lowry acids contain a proton.
- The net charge may be zero, (+), or (–).

Brønsted–Lowry bases [B:]

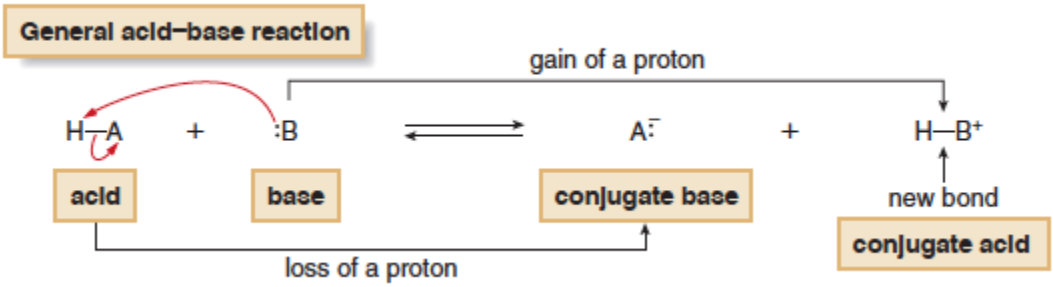
Inorganic		Organic	
		$CH_3\ddot{N}H_2$ methylamine	$CH_3\ddot{O}:^-$ methoxide
$H_2\ddot{O}:$	$:NH_3$		
$:\ddot{O}H^-$	$:\ddot{N}H_2^-$	 acetone	$CH_2=CH_2$ ethylene

- All Brønsted–Lowry bases contain a lone pair of electrons or a π bond.
- The net charge may be zero or (–).

Some molecules can be both!



Reactions of Bronsted Lowry acids and bases (note curved arrow notation)



Examples
Draw the conjugate acid of each base: NH_3, Cl^-, $(CH_3)_2C{=}O$

Draw the conjugate base of each acid: HBr, HSO_4^-, CH_3OH

Label the acid/base, and conjugate acid/base in the following. Use curved arrow notation to show the movement of electron pairs.

Key Ideas:

ELECTRON RICH species react with **ELECTRON-DEFICIENT** ones

Given two starting materials, how do you know which is the acid and which is the base?

1. Common acids and bases introduced in gen chem will often be used in teh same way in organic reactions: HCl & H_2SO_4 are strong acids, and ^-OH is a strong base
2. When only one starting material contains a hydrogen, it must be the acid!
3. If only one starting material has a lone pair or a pi bond, it must be the base!
4. A starting material with a net positive charge is usually the acid. A starting material with a negative charge is usually the base.

Part 6 Practice: *Bronsted-Lowry Acids and Bases*

---

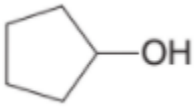
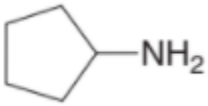
1. What is the conjugate acid of each base?

   a. H_2O b. $^-NH_2$ c. HCO_3^-

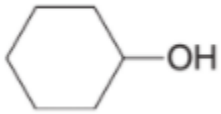
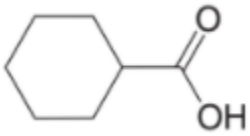
2. What is the conjugate base of each acid?

   a. HCN b. HCO_3^- c. $(CH_3)_2\overset{+}{N}H_2$

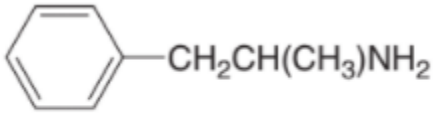
3. Draw the products formed from the acid-base reaction of H_2SO_4 with each compound

   a.  b. 

4. Draw the products formed from the acid-base reaction of KOH with each compound.

   a.  b. 

5. Amphetamine is a powerful stimulant of the central nervous system. Draw the products formed from the acid-base reaction of amphetamine with each reagent. (a) HCl; (b) NaH



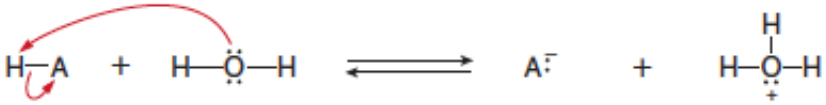
amphetamine

Part 7 Notes: *Acid Strength*

Acid Strength and pK_a

- The readily a compound donates a proton, the stronger the acid it is!

Dissolving an acid in water



Equilibrium constant

$$K_{eq} = \frac{[\text{products}]}{[\text{starting materials}]} = \frac{[H_3O^+][A:^-]}{[H{-}A][H_2O]}$$

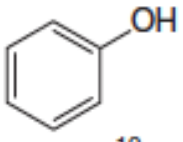
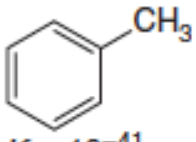
$$\textbf{Acidity constant} = K_a = [H_2O]K_{eq} = \frac{[H_3O^+][A:^-]}{[H{-}A]}$$

- Generally more convenient when describing acid strength to use pK_a (less numbers to write)

Definition: $pK_a = -\log K_a$

Which compound in each pair is the stronger acid?

a. $CH_3CH_2CH_3$ ($pK_a = 50$) or CH_3CH_2OH ($pK_a = 16$)

b.  ($K_a = 10^{-10}$) or  ($K_a = 10^{-41}$)

Acid strength vs Base Strength

- An _____ relationship exists
- A strong acid
- A strong base

Table 2.1 Selected pK_a Values

Acid	pK_a	Conjugate base
H–Cl	–7	Cl^-
CH_3COO-H	4.8	CH_3COO^-
HO–H	15.7	HO^-
CH_3CH_2O-H	16	$CH_3CH_2O^-$
HC≡CH	25	$HC\equiv C^-$
H–H	35	H^-
H_2N-H	38	H_2N^-
$CH_2=CH_2$	44	$CH_2=\bar{C}H$
CH_3-H	50	CH_3^-

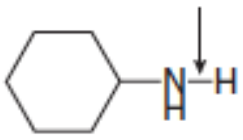
Increasing acidity of the acid ↑

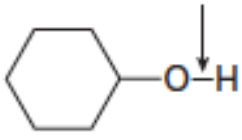
Increasing basicity of the conjugate base ↓

Examples

1. Consider to acids: HCO_2H (formic acid, pK_a= 3.8) and pivalic acid [$(CH_3)_3CCO_2H$, pK_a = 5.0].
   a. Which acid has the larger K_a?
   b. Which acid is the stronger acid?
   c. Which acid forms the stronger conjugate base?
   d. When each acid is dissolved in water, for which acid does the equilibrium lie further to the right?

2. Estimate the pK_a of each of the indicated bonds?

a. 

b. 

c. $BrCH_2COO-H$

Predicting outcome

- Equilibrium will always favor

Draw the products and predict the direction of equilibrium

a. $CH_2{=}CH_2 + H^-$ ⇌

b. $CH_4 + {}^-OH$ ⇌

Part 7 Practice: *Acid Strength*

NONE. Wrapped into Part 8 Practice

Part 8 Notes: *Factors that Determine A/B Strength*

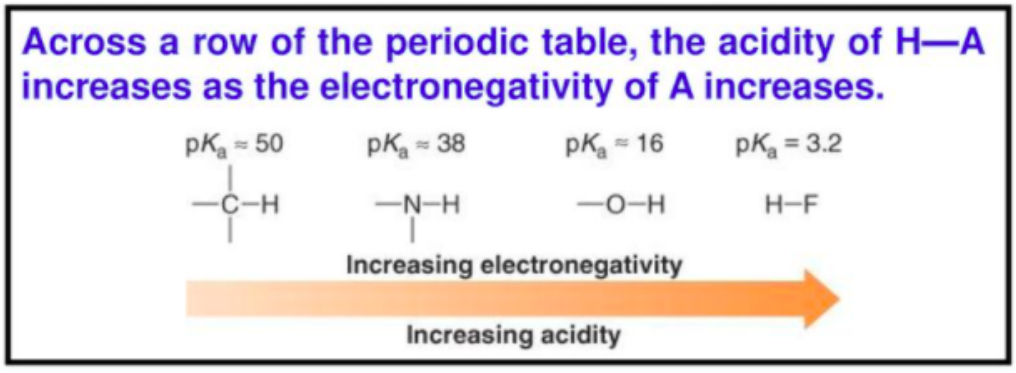
Anything that stabilizes the conjugate base makes the starting acid H-A more acidic.
4 main factors

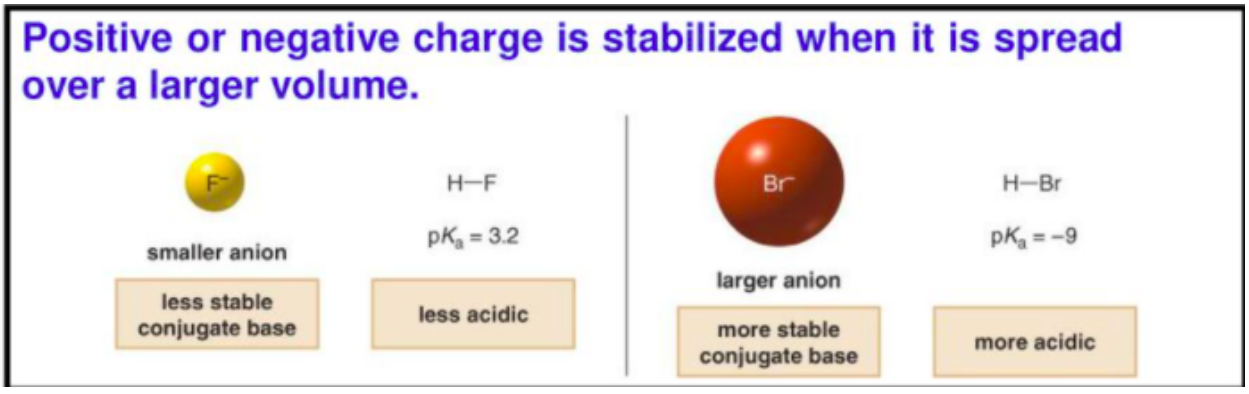
1. Element effects
2. Inductive effects
3. Resonance effects
4. Hybridization effects

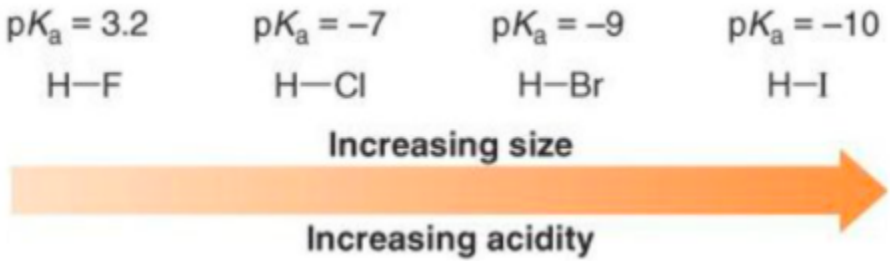
Steps to determine

1. Always draw the conjugate bases
2. Determine which conjugate base is more stable
3. The *more stable* conjugate base came from the stronger acid

Element Effects







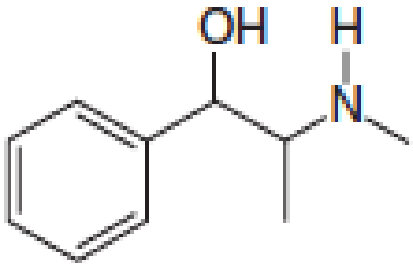
Examples

Decide which compound in each pair is the stronger acid.

a. H_2O or HF b. H_2S or H_2O

a. $CH_3CH_2CH_2NH_2$ or $(CH_3)_3N$ b. $CH_3CH_2OCH_3$ or $CH_3CH_2CH_2OH$

Which hydrogen is most acidic?



pseudoephedrine

Inductive Effects

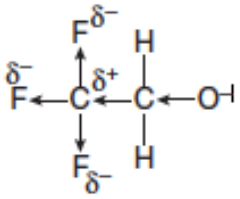
- Inductive effect is the _____ caused by

_____

CH_3CH_2O—H	CF_3CH_2O—H
ethanol	2,2,2-trifluoroethanol
$pK_a = 16$	$pK_a = 12.4$ ← **stronger acid**

$CH_3CH_2O^-$

No additional electronegative atoms stabilize the conjugate base.



CF_3 withdraws electron density, stabilizing the conjugate base.

- More electronegative atoms will

- The acidity of H-A

Examples

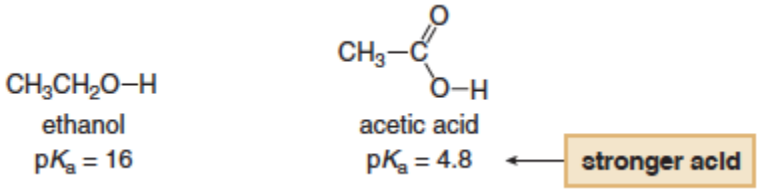
Which compound in each pair is the stronger acid

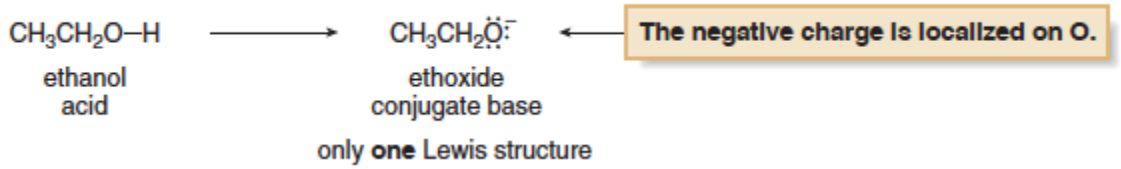
a. $ClCH_2COOH$ or FCH_2COOH

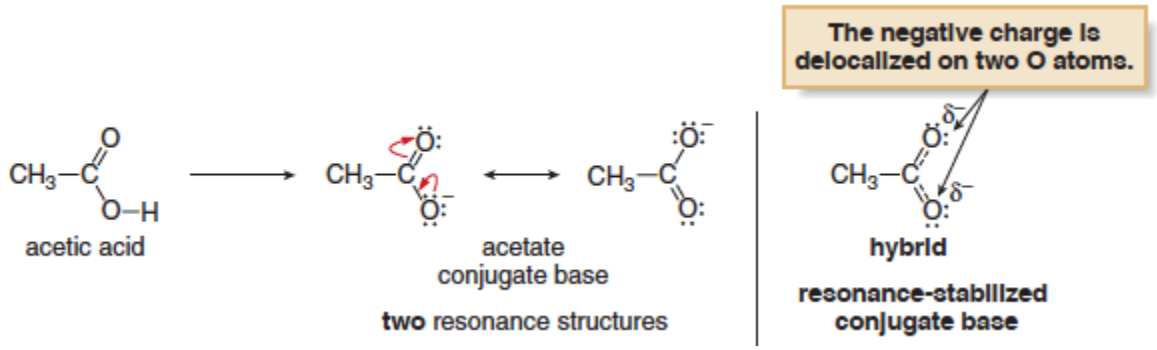
b. Cl_2CHCH_2OH or $Cl_2CHCH_2CH_2OH$

c. CH_3COOH or O_2NCH_2COOH

Resonance effects







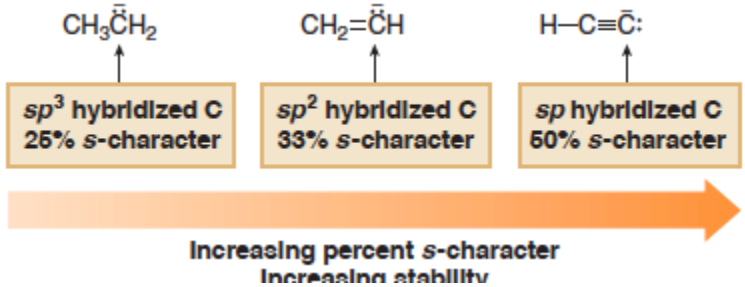
Resonance delocalization!

- The acidity of H-A increases when

Example

The C-H bond in acetone, $(CH_3)_2C{=}O$, has a pK_a of 19.2. Draw two resonance structures for its conjugate base. Then, explain why acetone is much more acidic than propane, $CH_3CH_2CH_3$ (pK_a = 50).

Hybridization



The higher percent s-character:

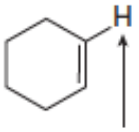
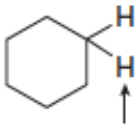
Because the:

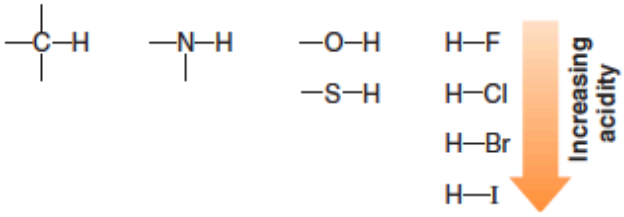
Example

For each pair answer/do the following:

1. Which indicated H is more acidic?
2. Draw the conjugate base of each acid.
3. Which conjugate base is stronger?

a. CH_3CH_2—C≡C—H or $CH_3CH_2CH_2CH_2$—H

b.  or 

Factor	Example
1. **Element effects:** The acidity of H – A increases both left-to-right across a row and down a column of the periodic table.	Increasing acidity → 
2. **Inductive effects:** The acidity of H – A increases with the presence of electron-withdrawing groups in A.	CH_3CH_2O—H CF_3CH_2O—H **more acidic**
3. **Resonance effects:** The acidity of H – A increases when the conjugate base $A:^-$ is resonance stabilized.	CH_3CH_2O—H CH_3COO—H **more acidic**
4. **Hybridization effects:** The acidity of H – A increases as the percent *s*-character of $A:^-$ increases.	CH_3CH_3 CH_2=CH_2 H—C≡C—H → Increasing acidity

Example

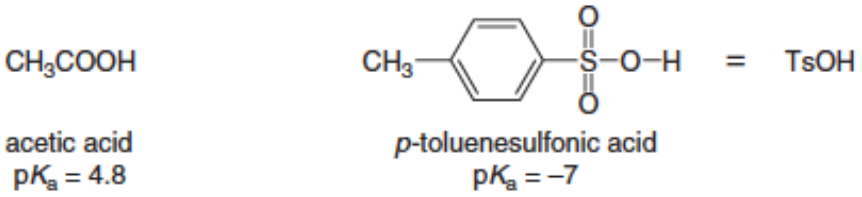
Rank the following compounds in order of increasing acidity of their most acidic hydrogen atom

$ClCH_2CH_2OH$ (A) CH_3CH_2OH (B) $CH_3CH_2NH_2$ (C)

Common acids and bases

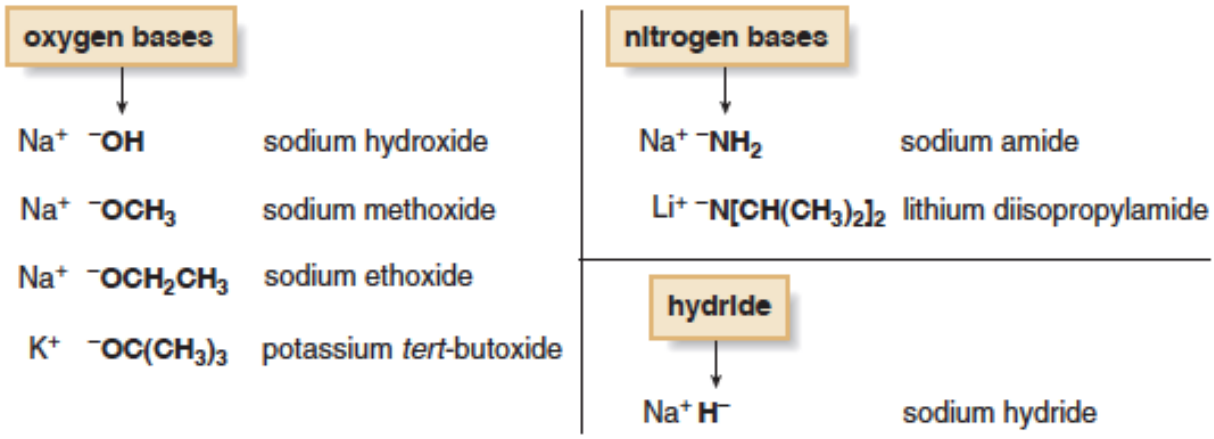
- Know all the ones from last year!

- Plus these two especially



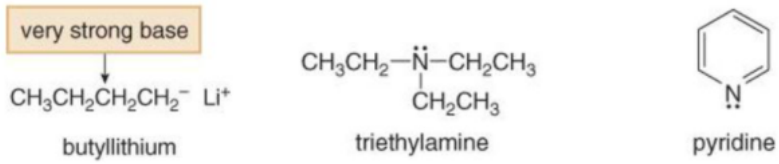
Common Bases

- Three basic traits
  - Negatively charged oxygen bases: ^{-}OH (hydroxide) and organic derivatives
  - Negatively charged nitrogen bases: $^{-}NH_2$ (amide) and organic derivatives
  - Hydride (H^{-})



Important base ideas:

- Strong bases have weak (or negligible) conjugate acids with high pK_a values, usually > 12
- Strong bases have a net negative charge, but not all negatively charged species are strong bases. For example, none of the halides are strong bases.
- Two other weaker organic bases are triethylamine and pyridine



Part 8 Practice: F*actors that Determine A/B Strength*

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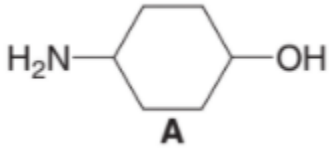
1. Rank the following compounds in order of increasing acidity
   a. NH_3, H_2O, HF
   b. HBr, HCl, HF
   c. H_2O, H_3O^+, HO^-

2. Rank the following ions in order of increasing basicity

   a. $CH_3\bar{C}H_2$, CH_3O^-, $CH_3\bar{N}H$

   b. CH_3^-, HO^-, Br^-

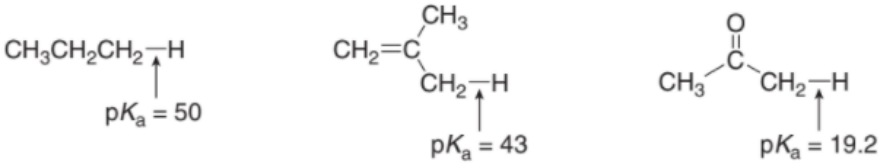
3. What is the conjugate acid of A? What is the conjugate base of A?



A

4. Explain why the pK_a of propiolic acid ($HC{\equiv}CCO_2H$, pK_a = 1.8) is significantly lower than the pK_a of propanoic acid ($CH_3CH_2CO_2H$, pK_a = 4.9).

5. The pK_a of three different C-H bonds is given below.

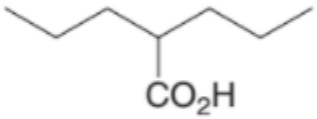


   a. For each compound, draw the conjugate base, including all possible resonance structures.

   b. Explain the observed trend in pK_a.

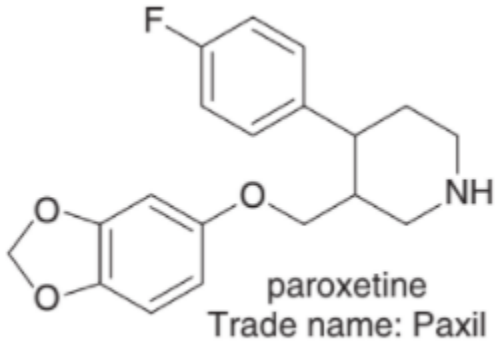
6. Label the most acidic hydrogen in each drug. Explain your choice.

a.



valproic acid
(used to treat epilepsy)

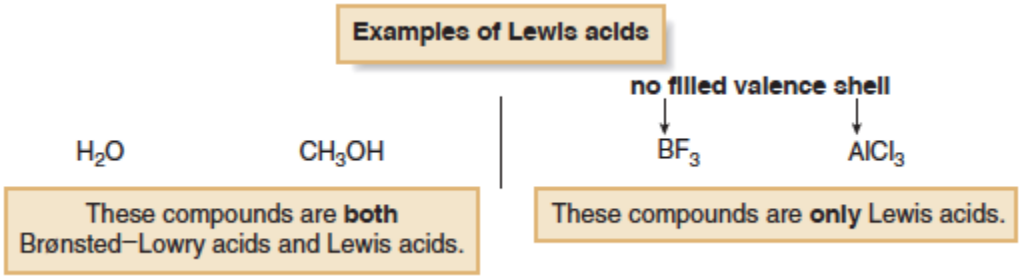
b.



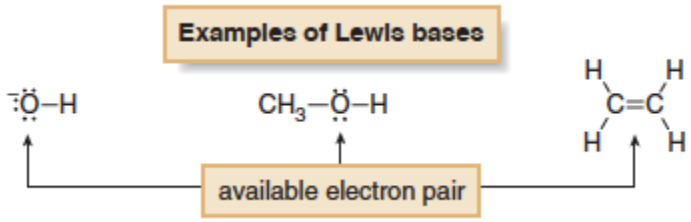
paroxetine
Trade name: Paxil
(used to treat depression)

Part 9 Notes: *Lewis Acids and Bases*

- Lewis Acid is defined as:



- Lewis Base is defined as:



Examples

- Which compounds are Lewis Bases?

a. NH_3 b. $CH_3CH_2CH_3$ c. H^- d. $H-C\equiv C-H$

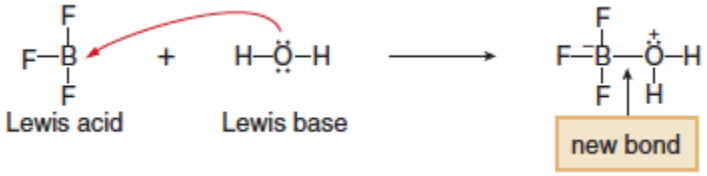
- Which compounds are Lewis Acids?

a. BBr_3 b. CH_3CH_2OH c. $(CH_3)_3C^+$ d. Br^-

Lewis Acid/Base Reactions

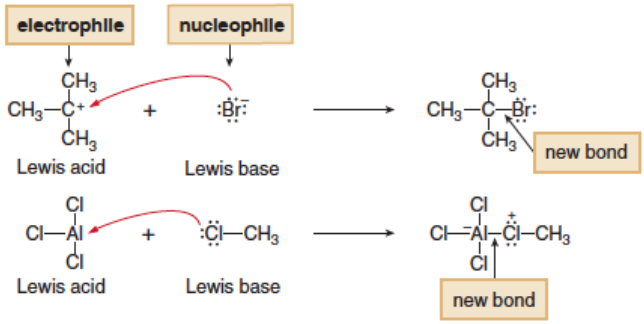
- Any reaction where _____ is a Lewis acid-base reaction

- SUPER COMMON in OChem! Electron _____ react with electron _____.



Electrophile:

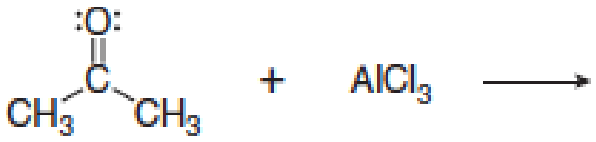
Nucleophile:



Example

Draw the products of each reaction, and label the nucleophile and electrophile.

a. $CH_3CH_2-\ddot{O}-CH_2CH_3$ + BBr_3 ⟶

b. 

Steps to remember in Lewis Acid-Base Reactions

1. Identify the Lewis acid and base FIRST
2. Draw a curved arrow from the electron pair of the base to the electron deficient atom of the acid
3. Count electron pairs and break a bond when needed to keep the correct number of valence electrons

Example.

- Label the Lewis Acid and Base. Use curved arrow notation to show the movement of electron pairs.

Part 9 Practice: *Lewis Acids and Bases*

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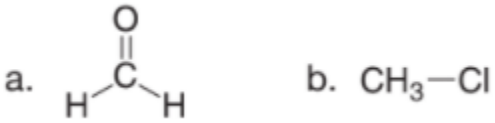
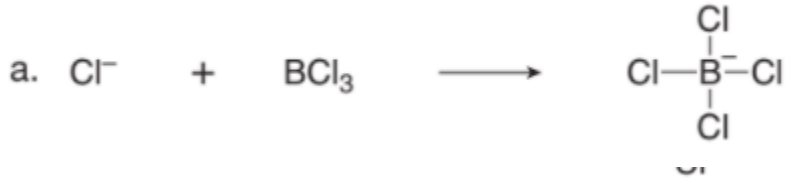
1. Classify each compound as a Lewis base, a Bronsted-Lowry base, both or neither.



2. Classify each species as a Lewis acid, a Bronsted-Lowry acid, both or neither.

a. H_3O^+ b. Cl_3C^+

3. Label the Lewis acid and Lewis base in each reaction. Use curved arrows to show the movement of electron pairs.



4. Draw the products of each Lewis acid-base reaction. Label the electrophile and nucleophile.

a. $CH_3CH_2OH + BF_3 \longrightarrow$

b. $CH_3SCH_3 + AlCl_3 \longrightarrow$