

ORGANIC  
CHEMISTRY | John McMurry  
NINTH EDITION



*John E. McMurry*

[www.cengage.com/chemistry/mcmurry](http://www.cengage.com/chemistry/mcmurry)

# Chapter 2

## Polar Covalent Bonds; Acids and Bases

# Learning Objectives



(2.1)

- Polar covalent bonds: Electronegativity

(2.2)

- Polar covalent bonds: Dipole moments

(2.3)

- Formal charges

(2.4)

- Resonance

(2.5)

- Rules for resonance forms

# Learning Objectives



(2.6)

- Drawing resonance forms

(2.7)

- Acids and bases: The Brønsted–Lowry definition

(2.8)

- Acid and base strength

(2.9)

- Predicting acid–base reactions from  $pK_a$  values

(2.10)

- Organic acids and organic bases

# Learning Objectives



(2.11)

- Acids and bases: The Lewis definition

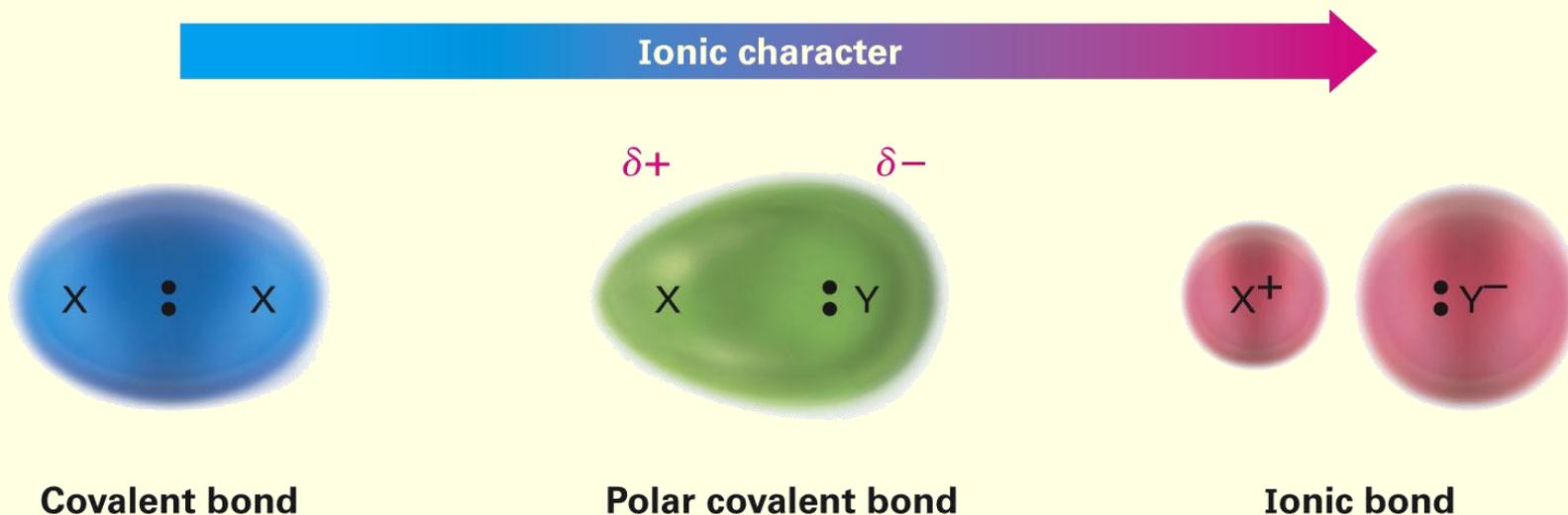
(2.12)

- Noncovalent interactions between molecules

# Polar Covalent Bonds: Electronegativity



- Covalent bonds can have ionic character
- **Polar covalent bonds:** Bonding electrons are attracted more strongly by one atom than by the other
  - Electron distribution between atoms is not symmetrical



# Electronegativity



- Intrinsic ability of an atom to attract the shared electrons in a covalent bond
- Differences in EN produce bond polarity
- F is most electronegative (EN = 4.0), Cs is least (EN = 0.7)
- Metals on left side of periodic table attract electrons weakly
- Halogens and other reactive nonmetals on right side of periodic table attract electrons strongly
- EN of C = 2.5

# Figure 2.2 - Electronegativity Values and Trends



H 2.1																	He
Li 1.0	Be 1.6											B 2.0	C 2.5	N 3.0	O 3.5	F 4.0	Ne
Na 0.9	Mg 1.2											Al 1.5	Si 1.8	P 2.1	S 2.5	Cl 3.0	Ar
K 0.8	Ca 1.0	Sc 1.3	Ti 1.5	V 1.6	Cr 1.6	Mn 1.5	Fe 1.8	Co 1.9	Ni 1.9	Cu 1.9	Zn 1.6	Ga 1.6	Ge 1.8	As 2.0	Se 2.4	Br 2.8	Kr
Rb 0.8	Sr 1.0	Y 1.2	Zr 1.4	Nb 1.6	Mo 1.8	Tc 1.9	Ru 2.2	Rh 2.2	Pd 2.2	Ag 1.9	Cd 1.7	In 1.7	Sn 1.8	Sb 1.9	Te 2.1	I 2.5	Xe
Cs 0.7	Ba 0.9	La 1.0	Hf 1.3	Ta 1.5	W 1.7	Re 1.9	Os 2.2	Ir 2.2	Pt 2.2	Au 2.4	Hg 1.9	Tl 1.8	Pb 1.9	Bi 1.9	Po 2.0	At 2.1	Rn

# Bond Polarity and Inductive Effect

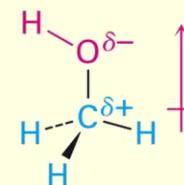
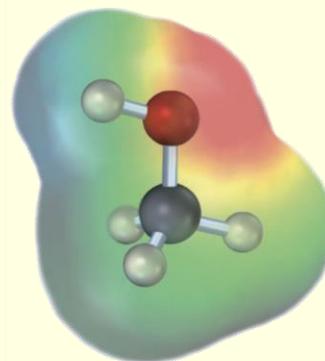


- Difference in EN of atoms  $< 2$  in polar covalent bonds
- Difference in EN  $> 2$  in ionic bonds
  - C–H bonds are relatively nonpolar
  - C–O, C–X bonds are polar
- Bonding electrons toward electronegative atom
  - C acquires partial positive charge,  $\delta+$
  - Electronegative atom acquires partial negative charge,  $\delta-$
- **Inductive effect:** Shifting of electrons in a  $\sigma$  bond in response to EN of nearby atoms



# Electrostatic Potential Maps

- Show calculated charge distributions
- Colors indicate electron-rich (red) and electron-poor (blue) regions
- Arrows indicate direction of bond polarity

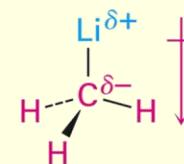
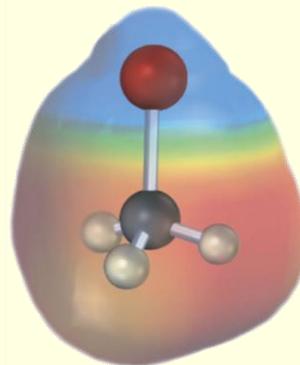


Methanol

Oxygen: EN = 3.5  
Carbon: EN = 2.5

---

Difference = 1.0



Methyllithium

Carbon: EN = 2.5  
Lithium: EN = 1.0

---

Difference = 1.5

# Worked Example



- Which element in each of the following pairs is more electronegative?
  - (a) Li or H
  - (b) Cl or I
- Solution:
  - Using Figure 2.2
  - (a) Li (1.0) is less electronegative when compared to H (2.1)
  - (b) Cl (3.0) is more electronegative when compared to I (2.5)

# Polar Covalent Bonds: Dipole Moments

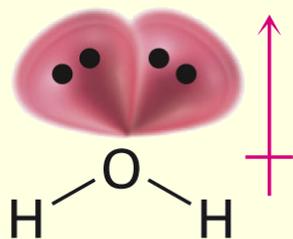


- Molecules are often polar from vector summation of individual bond polarities and lone-pair contributions
- Strongly polar substances are soluble in polar solvents like water
  - Nonpolar substances are insoluble in water
- **Dipole moment** ( $\mu$ ): Net molecular polarity, due to difference in summed charges
  - $\mu$  - Magnitude of charge  $Q$  at end of molecular dipole times distance  $r$  between charges

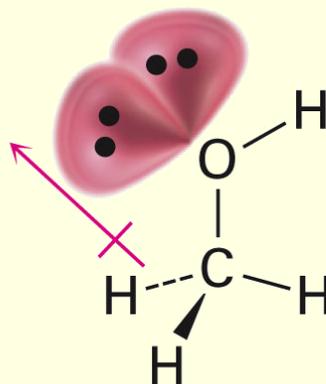
# Polar Covalent Bonds: Dipole Moments



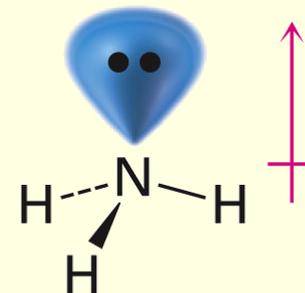
- $\mu = Q \times r$ , in debyes (D),  
1 D =  $3.336 \times 10^{-30}$  coulomb meter
- Length of an average covalent bond, the dipole moment would be  $1.60 \times 10^{-29}$  C·m, or 4.80 D



**Water**  
( $\mu = 1.85$  D)



**Methanol**  
( $\mu = 1.70$  D)



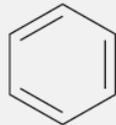
**Ammonia**  
( $\mu = 1.47$  D)

# Dipole Moments in Water and Ammonia



- Large dipole moments
  - EN of O and N > H
  - Both O and N have lone-pair electrons oriented away from all nuclei

TABLE 2-1 Dipole Moments of Some Compounds

Compound	Dipole moment (D)	Compound	Dipole moment (D)
NaCl	9.00	NH <sub>3</sub>	1.47
CH <sub>2</sub> O	2.33	CH <sub>3</sub> NH <sub>2</sub>	1.31
CH <sub>3</sub> Cl	1.87	CO <sub>2</sub>	0
H <sub>2</sub> O	1.85	CH <sub>4</sub>	0
CH <sub>3</sub> OH	1.70	CH <sub>3</sub> CH <sub>3</sub>	0
CH <sub>3</sub> CO <sub>2</sub> H	1.70		0
CH <sub>3</sub> SH	1.52	<b>Benzene</b>	

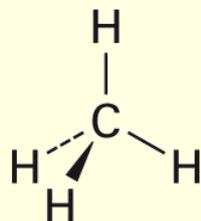
# Absence of Dipole Moments



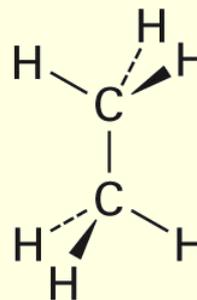
- In symmetrical molecules, the dipole moments of each bond have one in the opposite direction
  - The effects of the local dipoles cancel each other



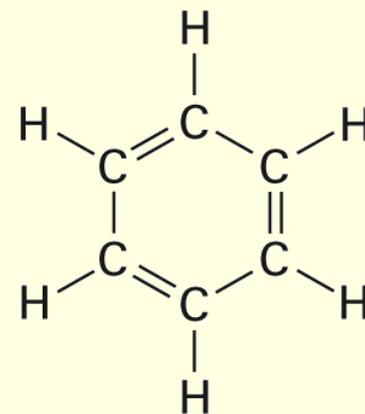
**Carbon dioxide**  
( $\mu = 0$ )



**Methane**  
( $\mu = 0$ )



**Ethane**  
( $\mu = 0$ )

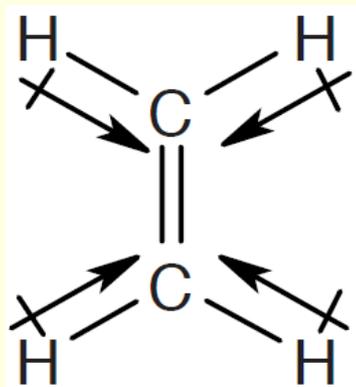


**Benzene**  
( $\mu = 0$ )

# Worked Example



- Draw three-dimensional drawing of  $\text{H}_2\text{C}=\text{CH}_2$  molecules
  - Predict whether it has dipole moment
- Solution:
  - Drawing an arrow that points from the least electronegative element to the most electronegative element

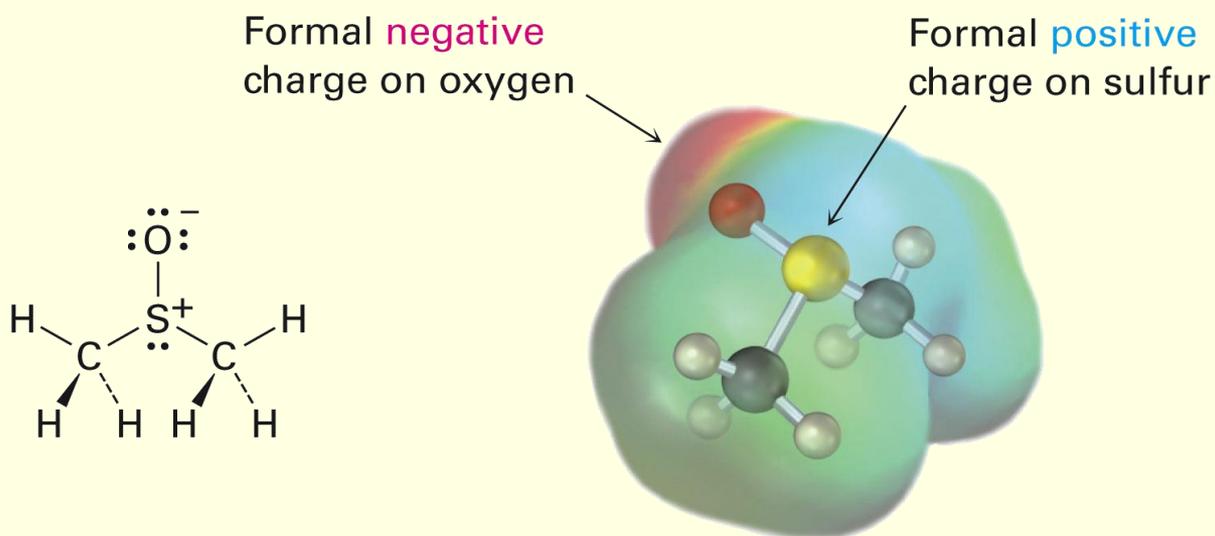


Has zero dipole moment

# Formal Charges



- At times it is necessary to have structures with formal charges on individual atoms
- Bonding of the atom in the molecule is compared to valence electron structure

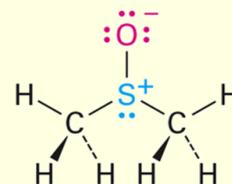


**Dimethyl sulfoxide**

# Formal Charge for Dimethyl Sulfoxide



- Atomic sulfur has 6 valence electrons
- Dimethyl sulfoxide sulfur has only 5
- It has lost an electron and has positive charge



For sulfur:

Sulfur valence electrons	= 6
Sulfur bonding electrons	= 6
Sulfur nonbonding electrons	= 2

$$\text{Formal charge} = 6 - 6/2 - 2 = +1$$

For oxygen:

Oxygen valence electrons	= 6
Oxygen bonding electrons	= 2
Oxygen nonbonding electrons	= 6

$$\text{Formal charge} = 6 - 2/2 - 6 = -1$$

$$\text{Formal charge} = \left( \begin{array}{c} \text{Number of} \\ \text{valence electrons} \\ \text{in free atom} \end{array} \right) - \left( \begin{array}{c} \text{Number of} \\ \text{valence electrons} \\ \text{in bonded atom} \end{array} \right)$$

$$= \left( \begin{array}{c} \text{Number of} \\ \text{valence electrons} \\ \text{in free atom} \end{array} \right) - \left( \frac{\text{Number of} \\ \text{bonding electrons}}{2} \right) - \left( \begin{array}{c} \text{Number of} \\ \text{nonbonding} \\ \text{electrons} \end{array} \right)$$

# Table 2.2 - A Summary of Common Formal Charges

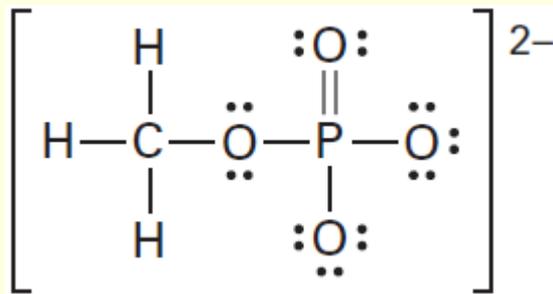


Atom	C			N		O		S		P
Structure	$\begin{array}{c} \cdot \\   \\ \text{---C---} \\   \end{array}$	$\begin{array}{c} + \\   \\ \text{---C---} \\   \end{array}$	$\begin{array}{c} \cdot\cdot \\   \\ \text{---C---} \\   \end{array}$	$\begin{array}{c}   \\ \text{---N}^+ \\   \end{array}$	$\begin{array}{c} \cdot\cdot \\   \\ \text{---N} \\   \\ \cdot\cdot \end{array}$	$\begin{array}{c} \cdot\cdot \\   \\ \text{---O}^+ \\   \end{array}$	$\begin{array}{c} \cdot\cdot \\   \\ \text{---O}^- \\   \\ \cdot\cdot \end{array}$	$\begin{array}{c} \cdot\cdot \\   \\ \text{---S}^+ \\   \end{array}$	$\begin{array}{c} \cdot\cdot \\   \\ \text{---S}^- \\   \\ \cdot\cdot \end{array}$	$\begin{array}{c}   \\ \text{---P}^+ \\   \end{array}$
Valence electrons	4	4	4	5	5	6	6	6	6	5
Number of bonds	3	3	3	4	2	3	1	3	1	4
Number of nonbonding electrons	1	0	2	0	4	2	6	2	6	0
Formal charge	0	+1	-1	+1	-1	+1	-1	+1	-1	+1

# Worked Example



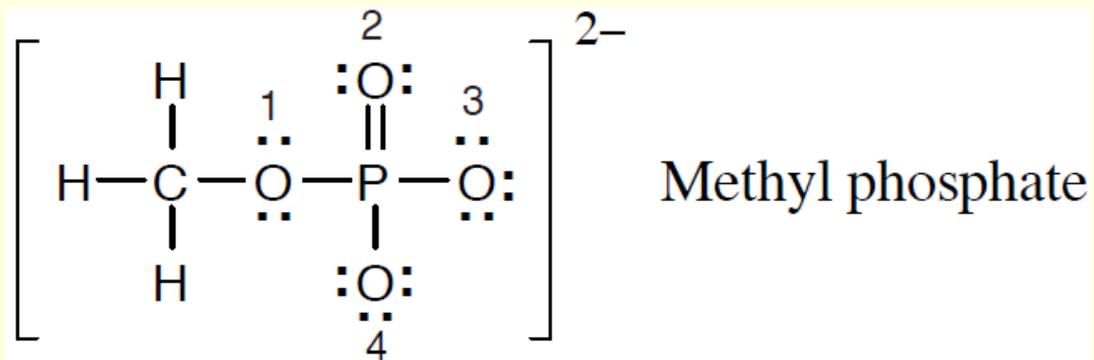
- Calculate formal charges on the four O atoms in the methyl phosphate dianion



- Solution:

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

# Worked Example



- For oxygen 1:  $\text{FC} = 6 - \frac{4}{2} - 4 = 0$
- For oxygen 2:  $\text{FC} = 6 - \frac{4}{2} - 4 = 0$
- For oxygen 3:  $\text{FC} = 6 - \frac{2}{2} - 6 = -1$

# Worked Example



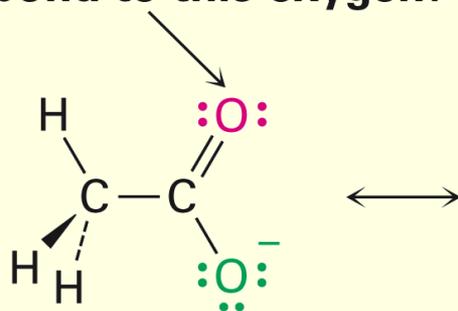
- For oxygen 4:  $FC = 6 - \frac{2}{2} - 6 = -1$
- Formal charge of oxygen atoms 1 and 2 is 0
- Formal charge of oxygen atoms 3 and 4 is -1



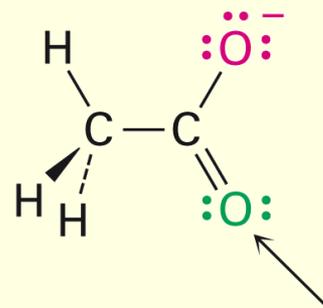
# Resonance

- Some molecules have structures that cannot be shown with a single representation
- Represented by structures that contribute to the final structure but differ in the position of the  $\pi$  bond or lone pair
  - Such structures are delocalized and are represented by **resonance forms**

Double bond to this oxygen?



Acetate ion

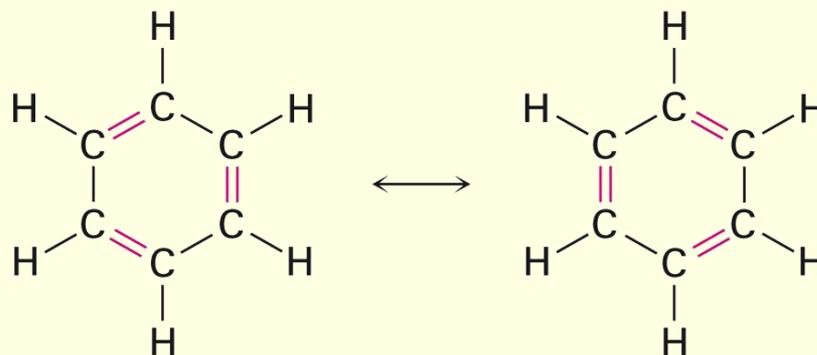
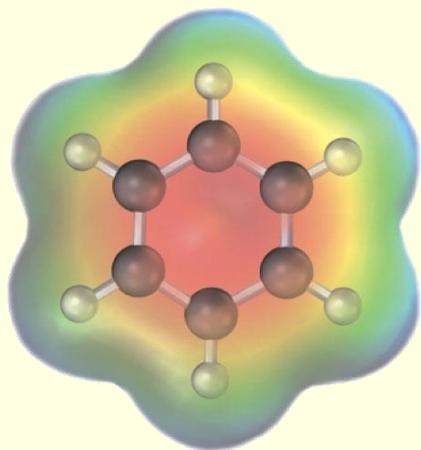


Or to this oxygen?



# Resonance Hybrids

- Structure with resonance forms, that does not alternate between the forms
- Example - Benzene ( $C_6H_6$ ) has two resonance forms with alternating double and single bonds
  - Is a hybrid of the two individual forms
  - All six carbon-carbon bonds are equivalent

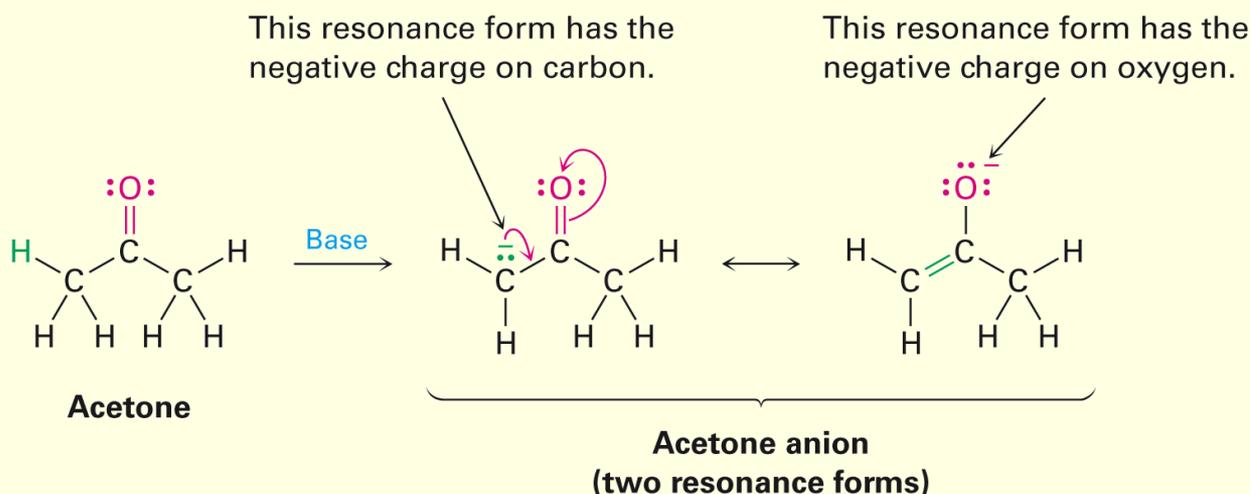


**Benzene (two resonance forms)**



# Rules for Resonance Forms

- Individual resonance forms are imaginary
  - Real structure is a hybrid of different forms
- Resonance forms differ only in the placement of their  $\pi$  or nonbonding electrons
  - Curved arrow indicates movement of electrons, not of the atoms





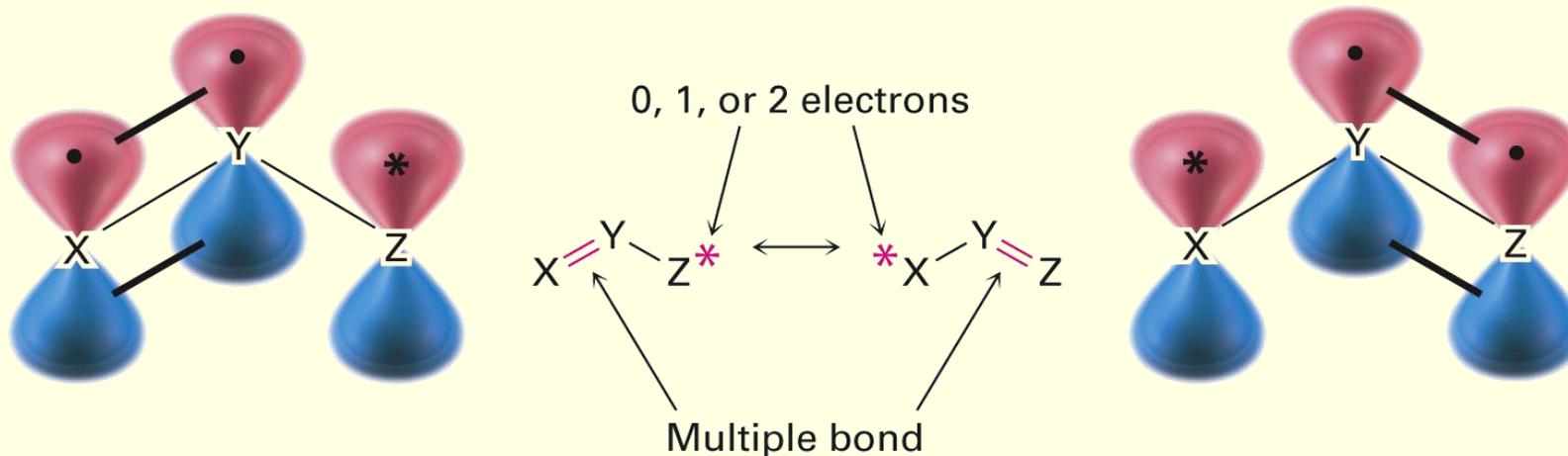
# Rules for Resonance Forms

- Different resonance forms of a substance do not have to be equivalent
  - When two resonance forms are nonequivalent, the actual structure of the resonance hybrid resembles the more stable form
- Resonance forms obey normal rules of valency
- Resonance hybrid is more stable than any individual resonance form
  - Resonance leads to stability



# Drawing Resonance Forms

- Any three-atom grouping with a  $p$  orbital on each atom has two resonance forms



# Drawing Resonance Forms

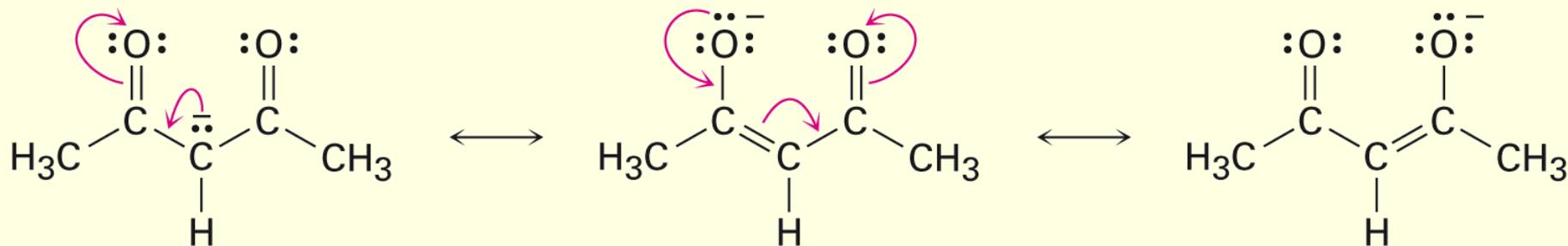


- Resonance forms differ by an exchange in position of the multiple bonds and the asterisk
  - From one end of the three-atom grouping to the other
- Recognizing three-atom groupings within larger structures help generate resonance forms, symmetrically

# 2,4-Pentanedione Anion



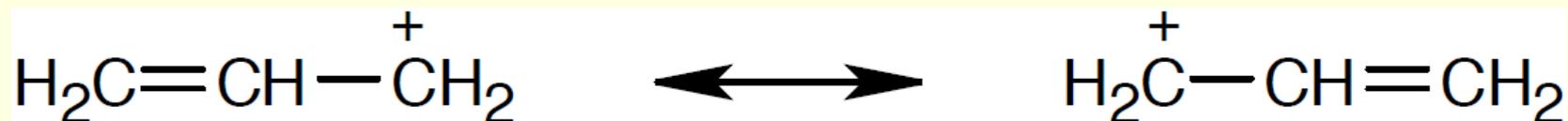
- Has a lone pair of electrons
- Has a formal negative charge on the central carbon atom, next to a C=O bond on the left and on the right
- Has three resonance structures



# Worked Example



- Draw the indicated number of resonance forms for:
  - The allyl cation,  $\text{H}_2\text{C}=\text{CH}-\text{CH}_2^+$  (2)
- Solution:
  - Locating three-atom groupings that contain a multiple bond next to an atom with a  $p$  orbital
  - Exchanging the positions of the bond and the electrons in the  $p$  orbital, we have:



# Acids and Bases: The Brønsted-Lowry Definition

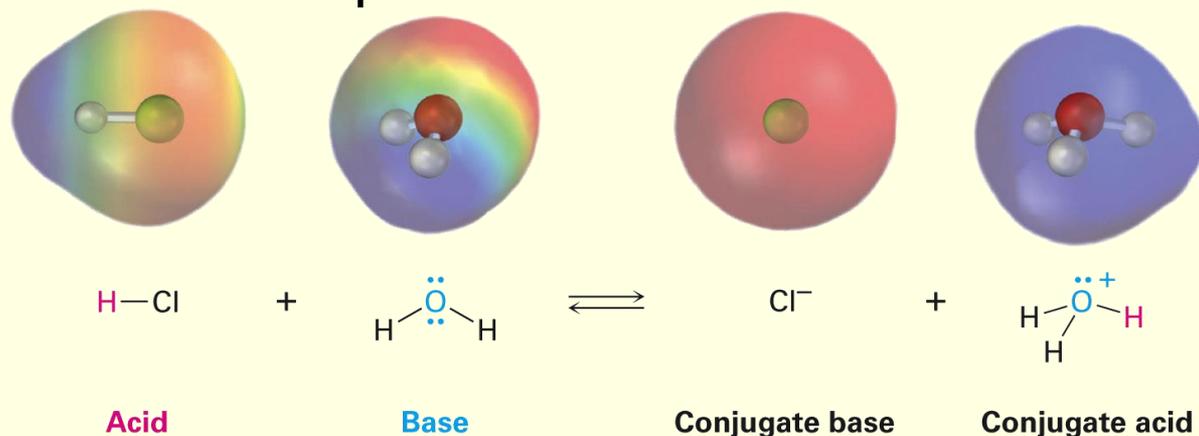


- Idea that acids are solutions containing a lot of “H<sup>+</sup>” and bases are solutions containing a lot of “OH<sup>-</sup>” is not very useful in organic chemistry
- Brønsted-Lowry theory defines acids and bases by their role in reactions that transfer protons (H<sup>+</sup>) between donors and acceptors

# Acids and Bases: The Brønsted-Lowry Definition



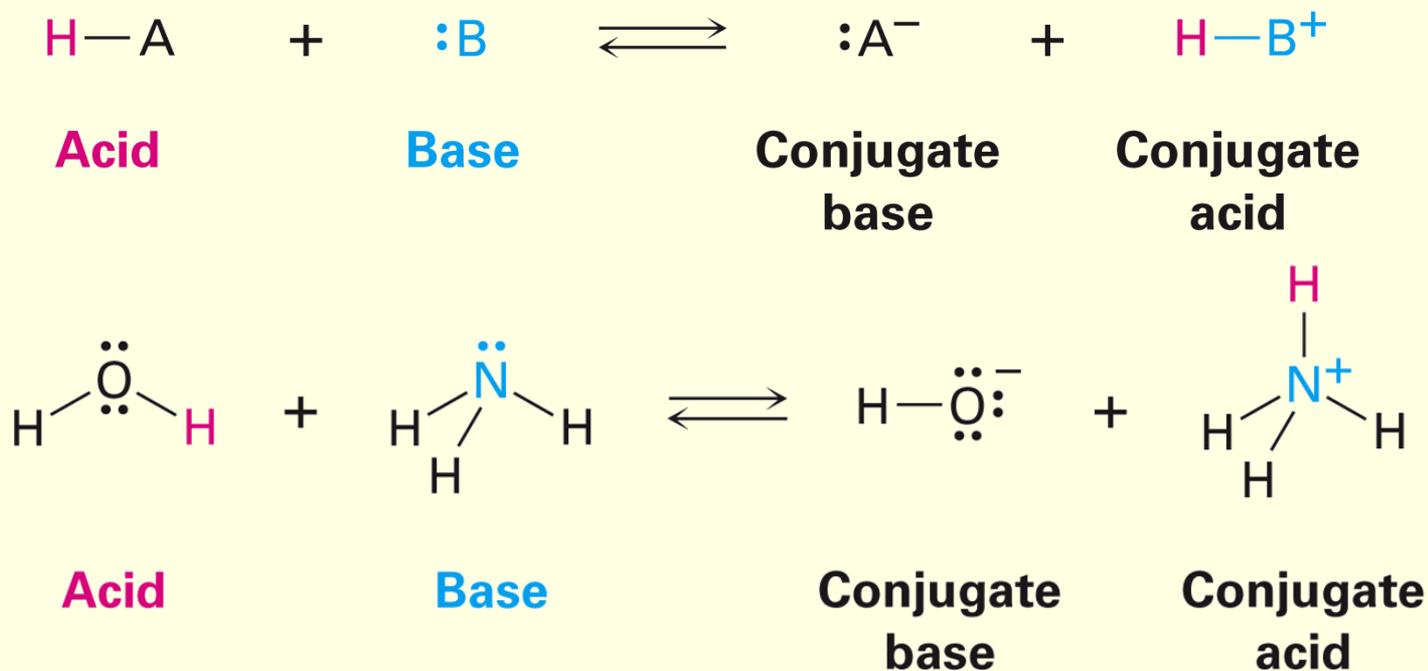
- **Brønsted-Lowry acid:** Substance that donates a hydrogen ion,  $H^+$
- **Brønsted-Lowry base:** Substance that accepts a hydrogen ion,  $H^+$ 
  - Proton is a synonym for  $H^+$ 
    - Loss of valence electron from H leaves only the nucleus—a proton



# Acids and Bases: The Brønsted-Lowry Definition



- **Conjugate base:** Product that results from deprotonation of a Brønsted-Lowry acid
- **Conjugate acid:** Product that results from protonation of a Brønsted-Lowry base



# Acid Base Strength



- **Acidity constant ( $K_a$ ):** Measure of acid strength
  - For the reaction of an acid (HA) with water to form hydronium ion
- Conjugate base ( $A^-$ ) is a measure related to the strength of the acid
- Brackets [ ] indicate concentration in moles per liter

$$K_a = \frac{[H_3O^+][A^-]}{[HA]}$$

# Acid and Base Strength



- Acid strengths are normally expressed using  $pK_a$  values
  - $pK_a$ : Negative common logarithm of the  $K_a$

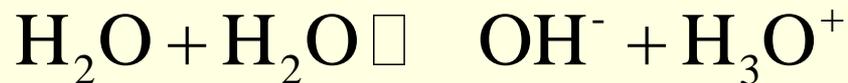
$$pK_a = -\log K_a$$

- Stronger acids have smaller  $pK_a$
- Weaker acids have larger  $pK_a$

# Acid and Base Strength



- Water is both an acid and base solvent



$$K_a = \frac{[\text{H}_3\text{O}^+][\text{OH}^-]}{[\text{HA}]} = \frac{[\text{H}_3\text{O}^+][\text{OH}^-]}{[\text{H}_2\text{O}]}$$

$$K_a = \frac{[1.0 \times 10^{-7}][1.0 \times 10^{-7}]}{55.4} = 1.8 \times 10^{-15}$$

$$\text{p}K_a = 15.74$$

- Ion product of water,  $K_w = [\text{H}_3\text{O}^+][\text{OH}^-]$
- Molar concentration of pure water,  $[\text{H}_2\text{O}] = 55.4 \text{ M}$  at  $25^\circ \text{ C}$

# Table 2.3 - Relative Strengths of Some Common Acids and Their Conjugate Bases



	Acid	Name	$pK_a$	Conjugate base	Name	
Weaker acid	$\text{CH}_3\text{CH}_2\text{OH}$	Ethanol	16.00	$\text{CH}_3\text{CH}_2\text{O}^-$	Ethoxide ion	Stronger base
	$\text{H}_2\text{O}$	Water	15.74	$\text{HO}^-$	Hydroxide ion	
	$\text{HCN}$	Hydrocyanic acid	9.31	$\text{CN}^-$	Cyanide ion	
	$\text{H}_2\text{PO}_4^-$	Dihydrogen phosphate ion	7.21	$\text{HPO}_4^{2-}$	Hydrogen phosphate ion	
	$\text{CH}_3\text{CO}_2\text{H}$	Acetic acid	4.76	$\text{CH}_3\text{CO}_2^-$	Acetate ion	
	$\text{H}_3\text{PO}_4$	Phosphoric acid	2.16	$\text{H}_2\text{PO}_4^-$	Dihydrogen phosphate ion	
	$\text{HNO}_3$	Nitric acid	-1.3	$\text{NO}_3^-$	Nitrate ion	
Stronger acid	$\text{HCl}$	Hydrochloric acid	-7.0	$\text{Cl}^-$	Chloride ion	Weaker base

# Worked Example



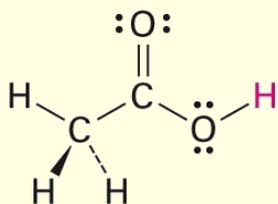
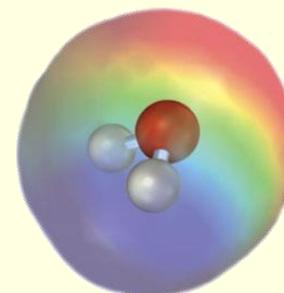
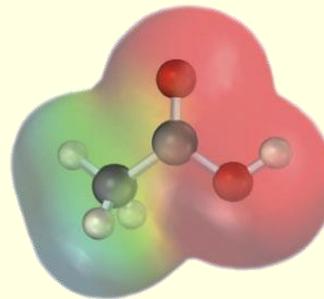
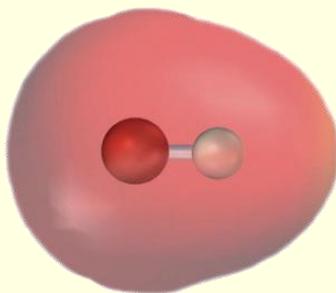
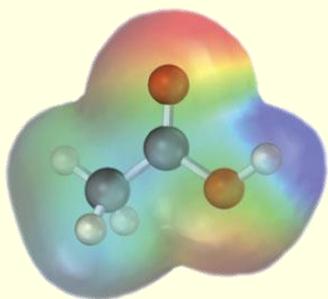
- The amino acid phenylalanine has  $pK_a = 1.83$ , and tryptophan has  $pK_a = 2.83$ 
  - Which is the stronger acid?
- Solution:
  - Stronger acid has a smaller  $pK_a$  and a weaker acid has a larger  $pK_a$
  - Accordingly, phenylalanine ( $pK_a = 1.83$ ) is a stronger acid than tryptophan ( $pK_a = 2.83$ )

# Predicting Acid-Base Reactions from $pK_a$ Values



- $pK_a$  values are related as logarithms to equilibrium constants
- Useful for predicting whether a given acid-base reaction will take place
- Difference in two  $pK_a$  values is the log of the ratio of equilibrium constants, and can be used to calculate the extent of transfer

# Predicting Acid-Base Reactions from $pK_a$ Values

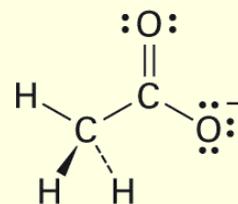


**Acetic acid**  
( $pK_a = 4.76$ )

+

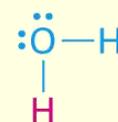


**Hydroxide ion**



**Acetate ion**

+

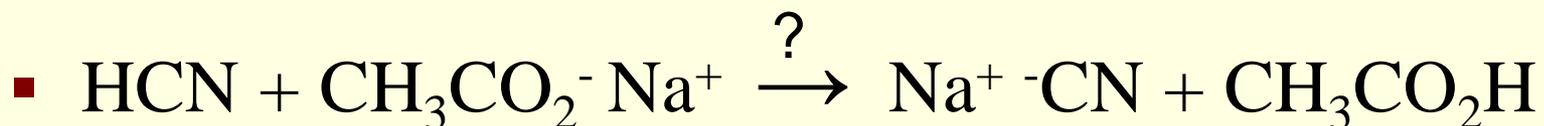


**Water**  
( $pK_a = 15.74$ )

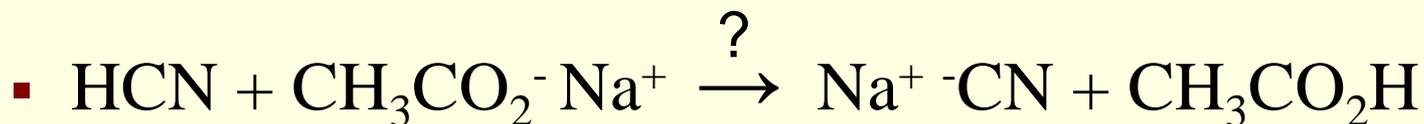
# Worked Example



- Will the following reaction take place to a significant extent as written, according to the data in Table 2.3?



- Solution:



$\text{pK}_a = 9.3$   
Weaker acid

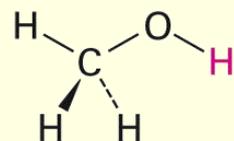
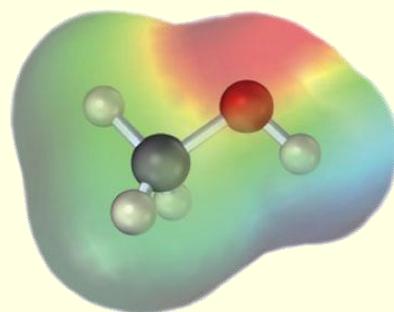
$\text{pK}_a = 4.7$   
Stronger acid

- Since  $\text{CH}_3\text{CO}_2\text{H}$  is stronger than  $\text{HCN}$  the reaction will not take place to a significant extent the direction written

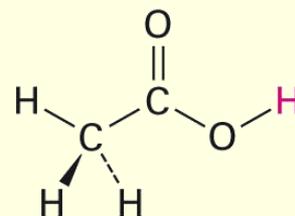
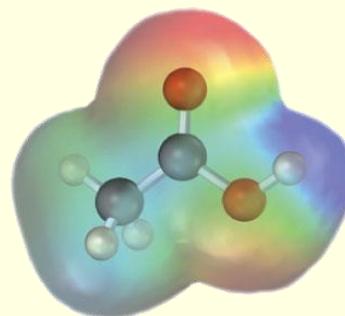
# Organic Acids



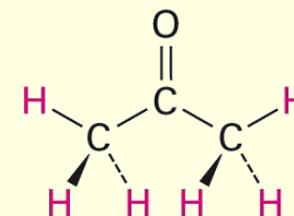
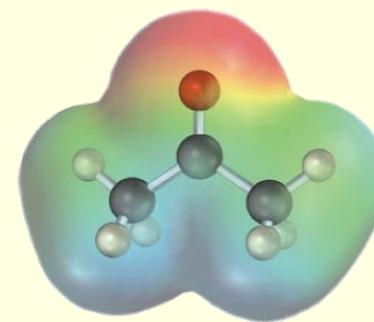
- Characterized by the presence of positively polarized hydrogen atom



**Methanol**  
( $pK_a = 15.54$ )



**Acetic acid**  
( $pK_a = 4.76$ )



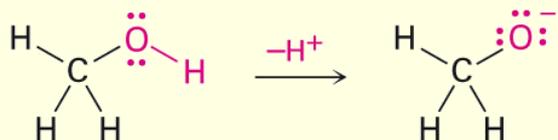
**Acetone**  
( $pK_a = 19.3$ )

Some organic acids

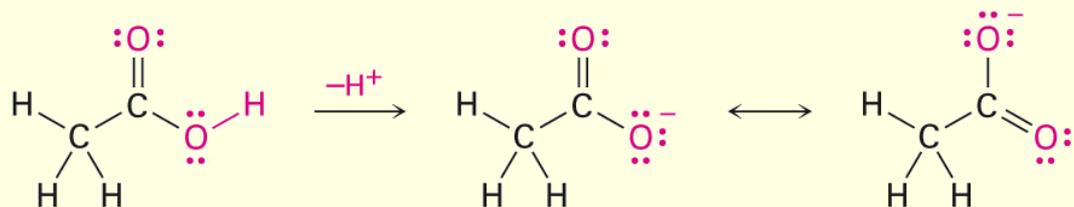
# Organic Acids



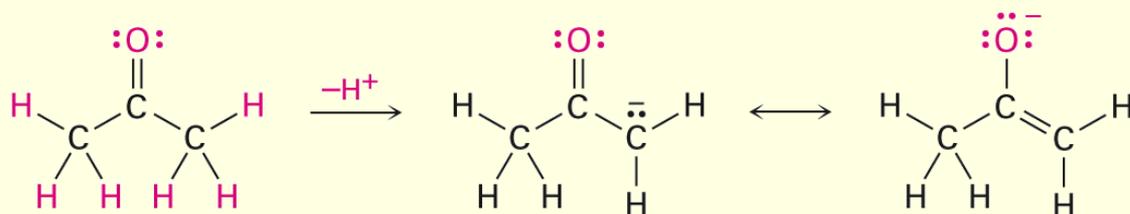
- Two main kinds, those that contain:
  - Hydrogen atom bonded to a electronegative oxygen atom(O–H)
  - A hydrogen atom bonded to a carbon atom next to a C=O bond(O=C–C–H)



Anion is stabilized by having negative charge on a highly electronegative atom.



Anion is stabilized both by having negative charge on a highly electronegative atom and by resonance.



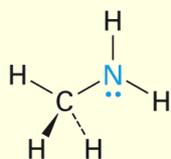
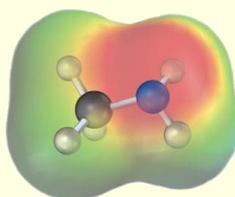
Anion is stabilized both by resonance and by having negative charge on a highly electronegative atom.

# Organic Bases

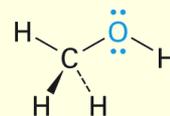
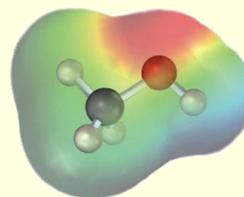


- Have an atom with a lone pair of electrons that can bond to  $H^+$
- Nitrogen-containing compounds derived from ammonia are the most common organic bases
- Oxygen-containing compounds can react as bases with a strong acid or as acids with strong bases

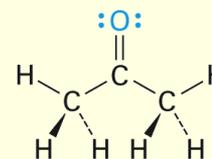
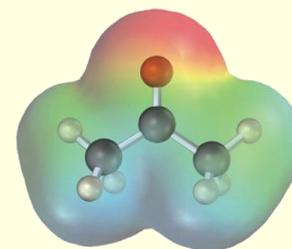
Some organic bases



**Methylamine**



**Methanol**

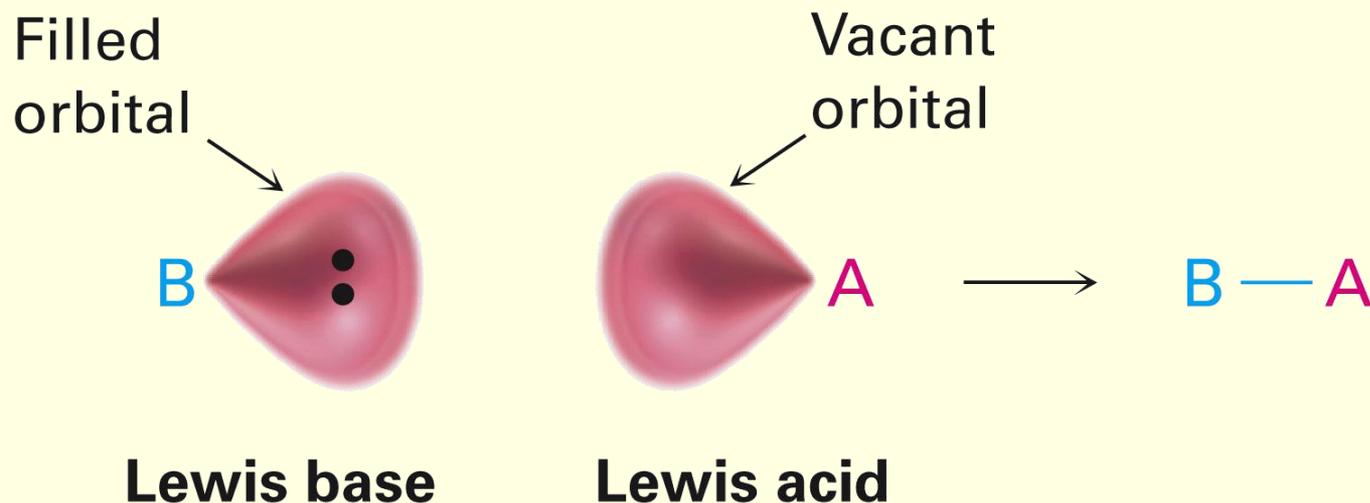


**Acetone**

# Acids and Bases: The Lewis Definition



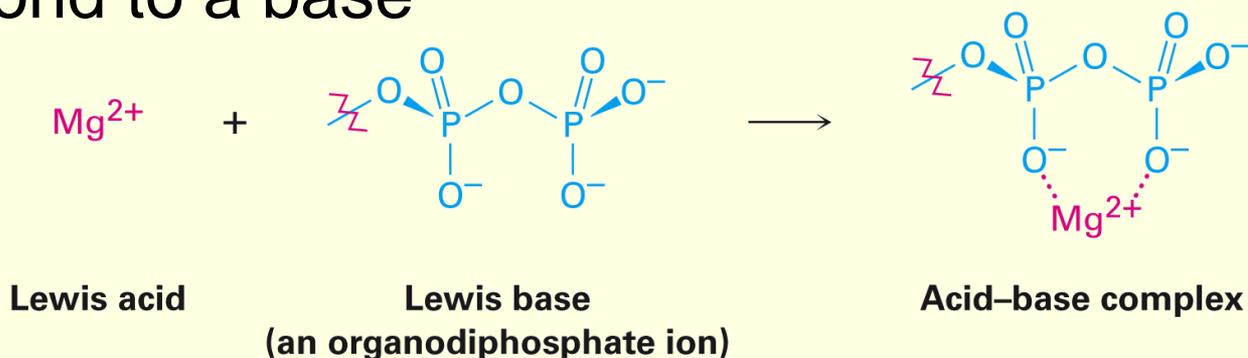
- **Lewis acid:** Electron pair acceptors
- **Lewis bases:** Electron pair donors
- Brønsted acids are not Lewis acids because they cannot accept an electron pair directly



# Lewis Acids and the Curved Arrow Formalism



- Lewis definition of acidity includes metal cations, such as  $\text{Mg}^{2+}$ 
  - They accept a pair of electrons when they form a bond to a base



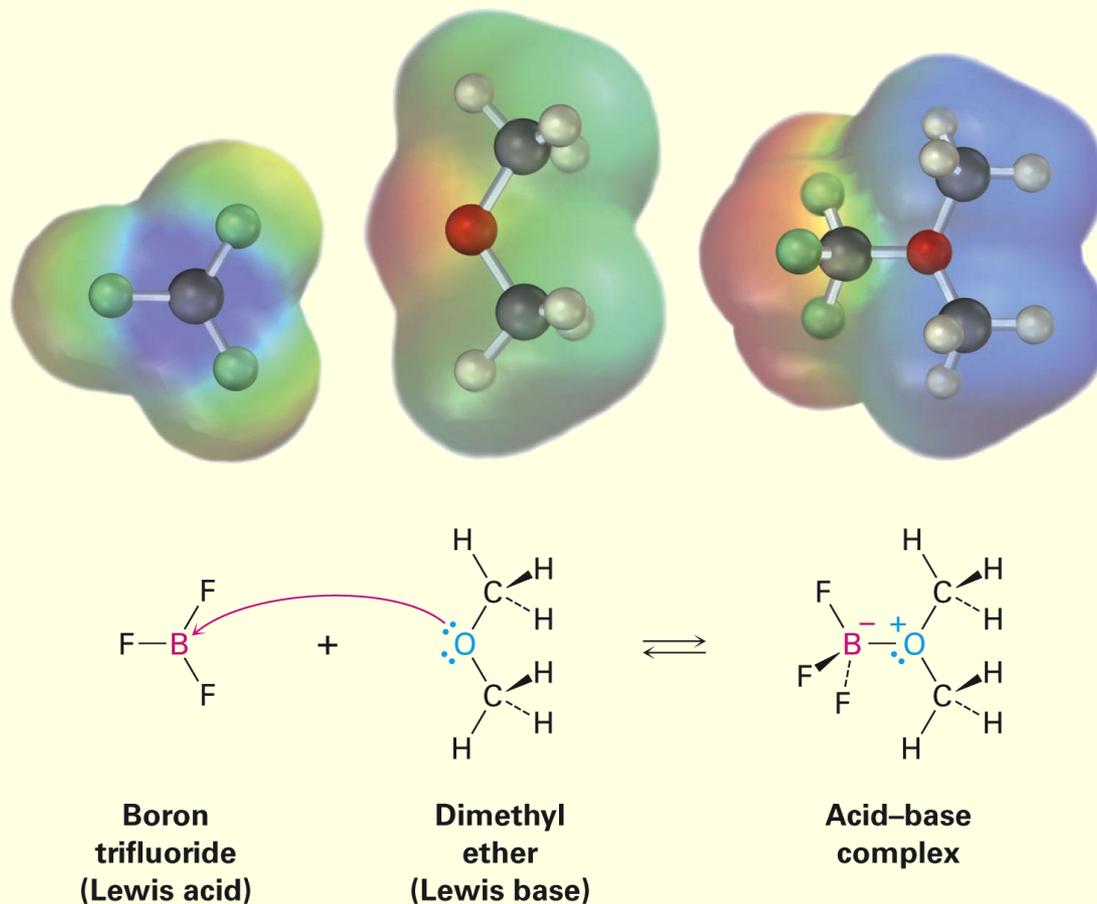
- Group 3A elements, such as  $\text{BF}_3$  and  $\text{AlCl}_3$ , are Lewis acids
  - Have unfilled valence orbitals and can accept electron pairs from Lewis bases

# Lewis Acids and the Curved Arrow Formalism



- Transition-metal compounds, such as  $\text{TiCl}_4$ ,  $\text{FeCl}_3$ ,  $\text{ZnCl}_2$ , and  $\text{SnCl}_4$ , are Lewis acids
- Curved arrow means that a pair of electrons move from the atom at the tail of the arrow to the atom at the head of the arrow

# Figure 2.5 - The Reaction of Boron Trifluoride with Dimethyl Ether



# Lewis Bases

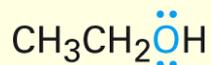


- Compound with a pair of nonbonding electrons that it can use to bond to a Lewis acid
- Can accept protons as well as Lewis acids
  - Definition encompasses that for Brønsted bases
- Oxygen- and nitrogen-containing organic compounds are Lewis bases; they have lone pairs of electrons
- Some compounds can act as both acids and bases



# Lewis Bases

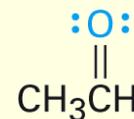
Some  
Lewis  
bases



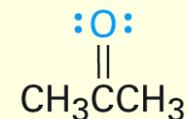
An alcohol



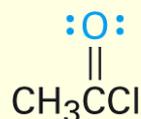
An ether



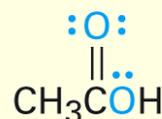
An aldehyde



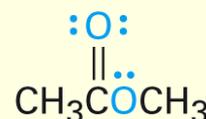
A ketone



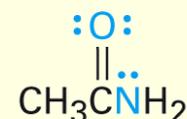
An acid chloride



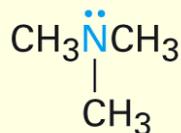
A carboxylic  
acid



An ester



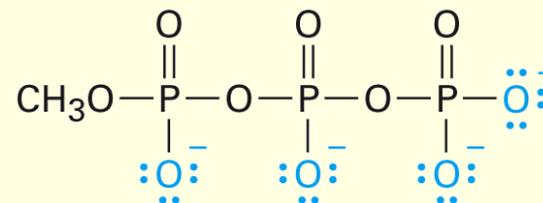
An amide



An amine



A sulfide

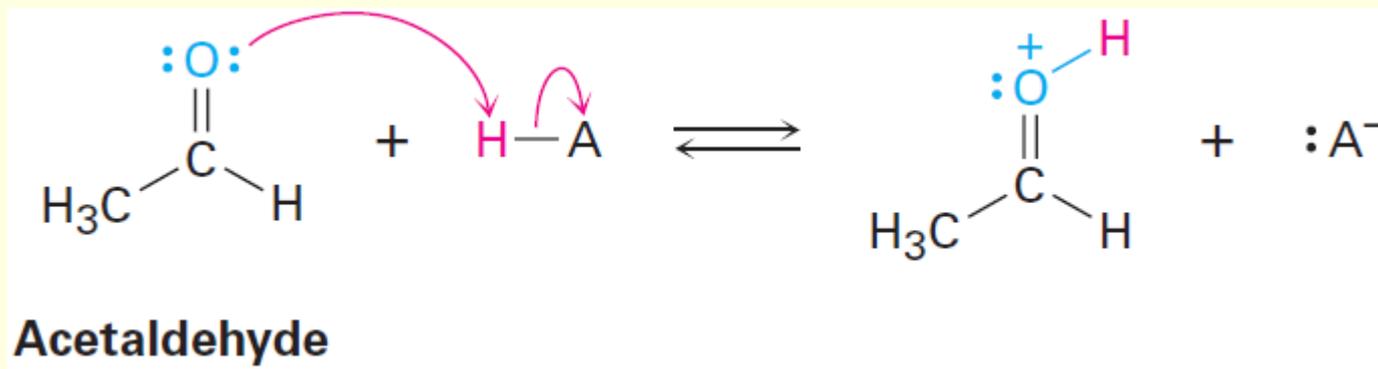


An organotriphosphate ion

# Worked Example



- Using curved arrows, show how acetaldehyde,  $\text{CH}_3\text{CHO}$ , can act as a Lewis base
- Solution:
  - A Lewis base donates an electron pair to a Lewis acid
  - Using a curved arrow to show the movement of a pair toward the H atom of the acid



# Noncovalent Interactions Between Molecules

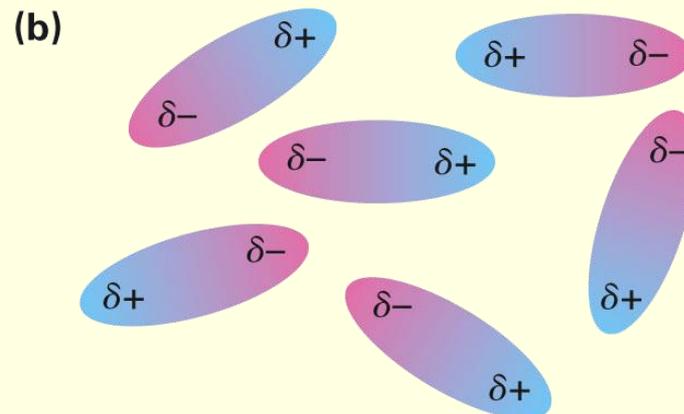
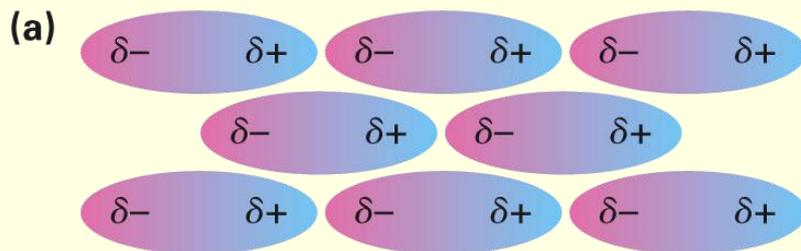


- **Noncovalent interactions:** One of a variety of nonbonding interactions between molecules
  - Dipole–dipole forces
  - Dispersion forces
  - Hydrogen bonds



# Dipole–Dipole Forces

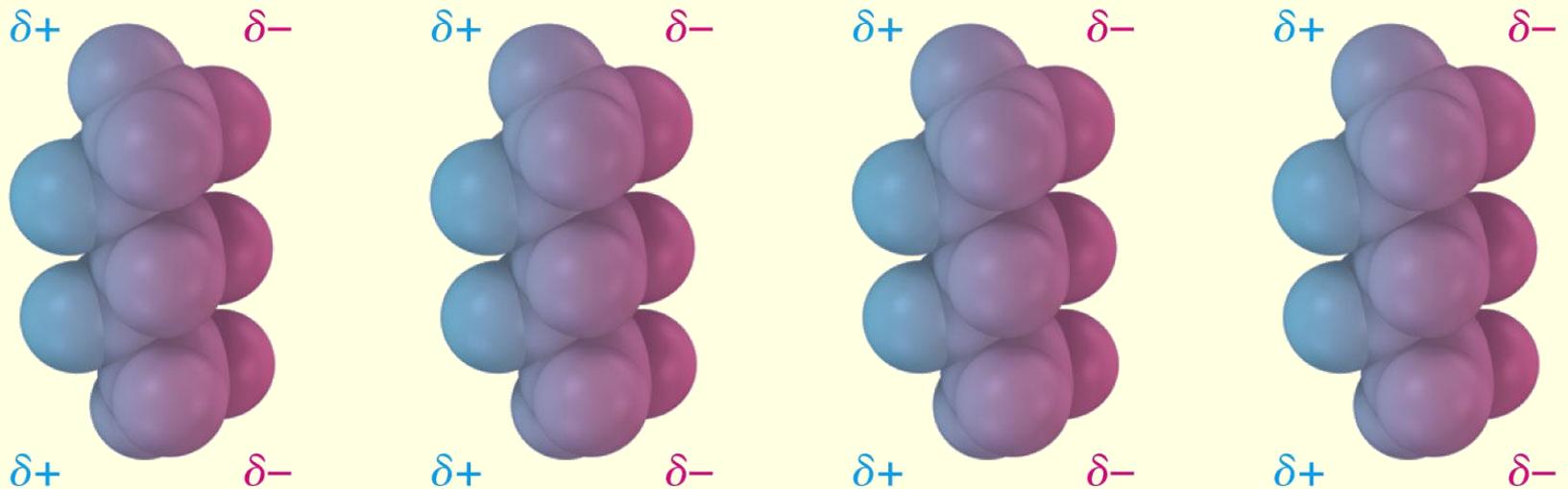
- Occur between polar molecules as a result of electrostatic interactions among dipoles
- Depending on orientation of the molecules, the forces can be either attractive or repulsive



# Dispersion Forces



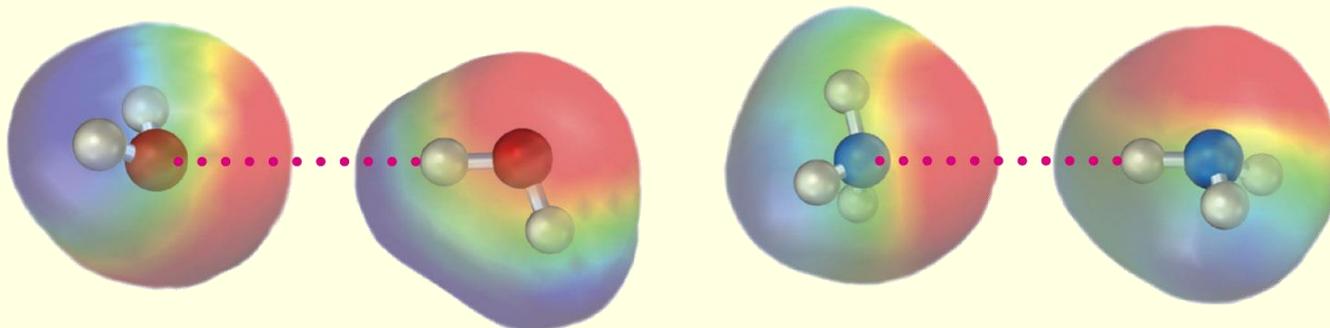
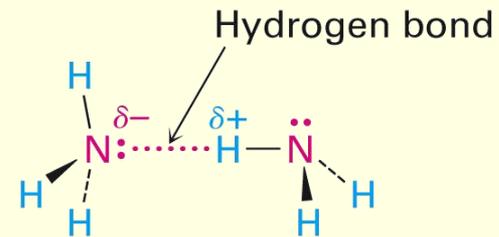
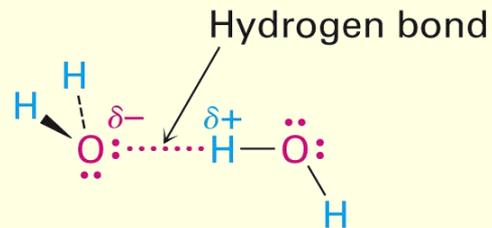
- Occur between all neighboring molecules
  - Arise due to constant change in electron distribution within molecules



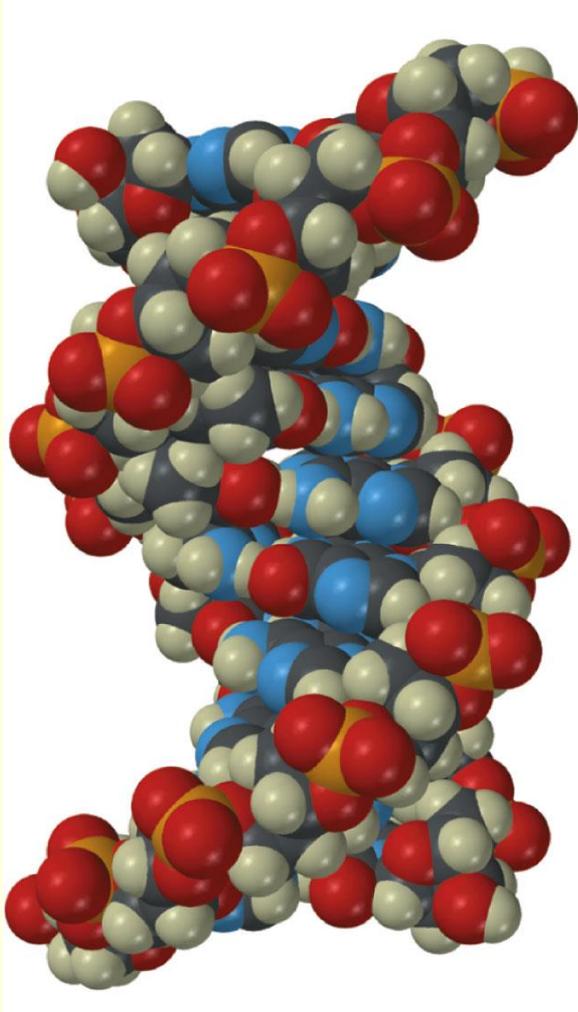
# Hydrogen Bond Forces



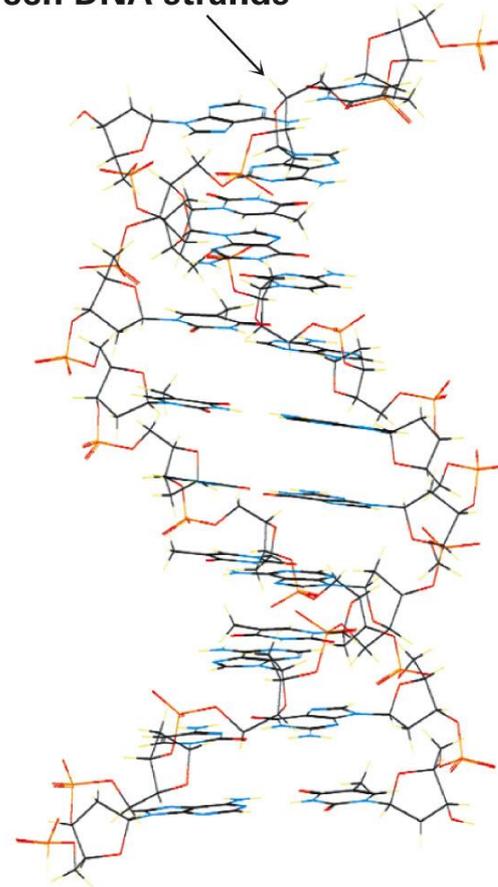
- Forces are the result of attractive interaction between a hydrogen bonded to an electronegative O or N atom and an unshared electron pair on another O or N atom



# A Deoxyribonucleic Acid Segment



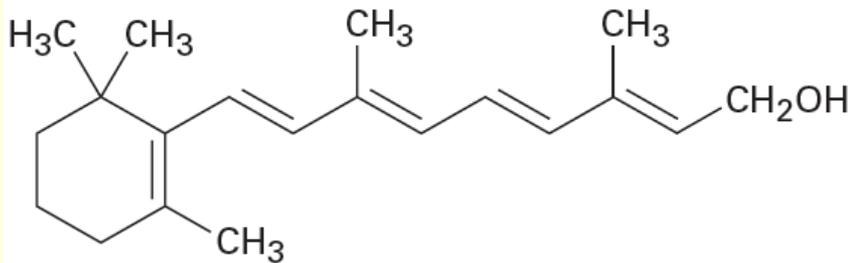
Hydrogen bonds  
between DNA strands



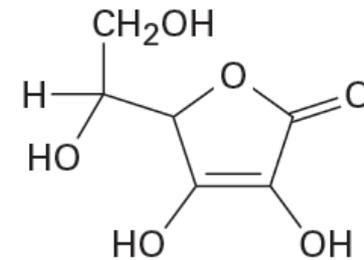
# Worked Example



- Of the two vitamins A and C, one is hydrophilic and water-soluble while the other is hydrophobic and fat-soluble
  - Which is which?



**Vitamin A**  
(retinol)

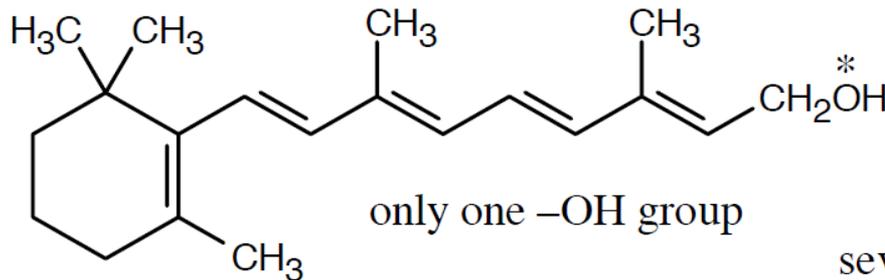


**Vitamin C**  
(ascorbic acid)

# Worked Example



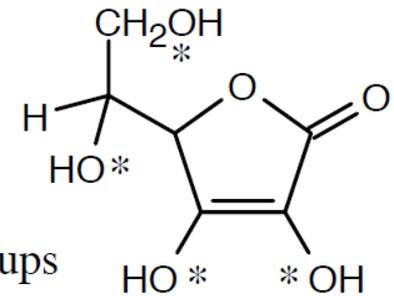
- Solution:



Vitamin A

several  $-OH$  groups

\* = polar group



Vitamin C

- Vitamin C has several polar  $-OH$  groups that can form hydrogen bonds with water
  - It is water soluble(hydrophilic)
- Most of Vitamin A's atoms can't form hydrogen bonds with water
  - It is fat-soluble(hydrophobic)

# Summary



- Organic molecules often have polar covalent bonds as a result of unsymmetrical electron sharing caused by differences in the electronegativity of atoms
- Polarity of a molecule is measured by its dipole moment,  $\mu$
- (+) and (–) indicate formal charges on atoms in molecules to keep track of valence electrons around an atom

# Summary



- Some substances must be shown as a resonance hybrid of two or more resonance forms that differ by the location of electrons
- A Brønsted(–Lowry) acid donates a proton
- A Brønsted(–Lowry) base accepts a proton
- Strength of Brønsted acid is related to the negative logarithm of the acidity constant,  $\text{p}K_a$
- Weaker acids have higher values of  $\text{p}K_a$

# Summary



- Lewis acid has an empty orbital that can accept an electron pair
- Lewis base can donate an unshared electron pair
- Noncovalent interactions have several types – Dipole–dipole, dispersion, and hydrogen bond forces