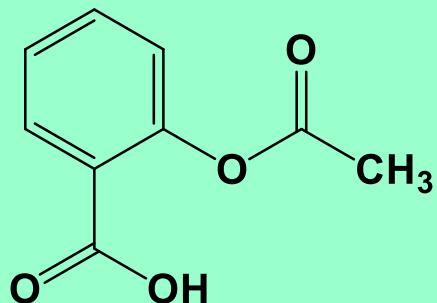
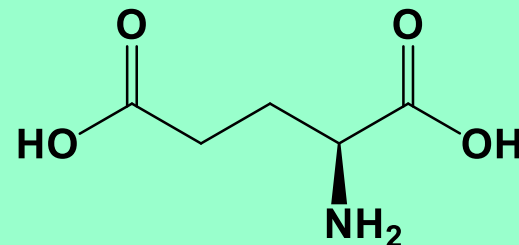


Morphine (pain killer)

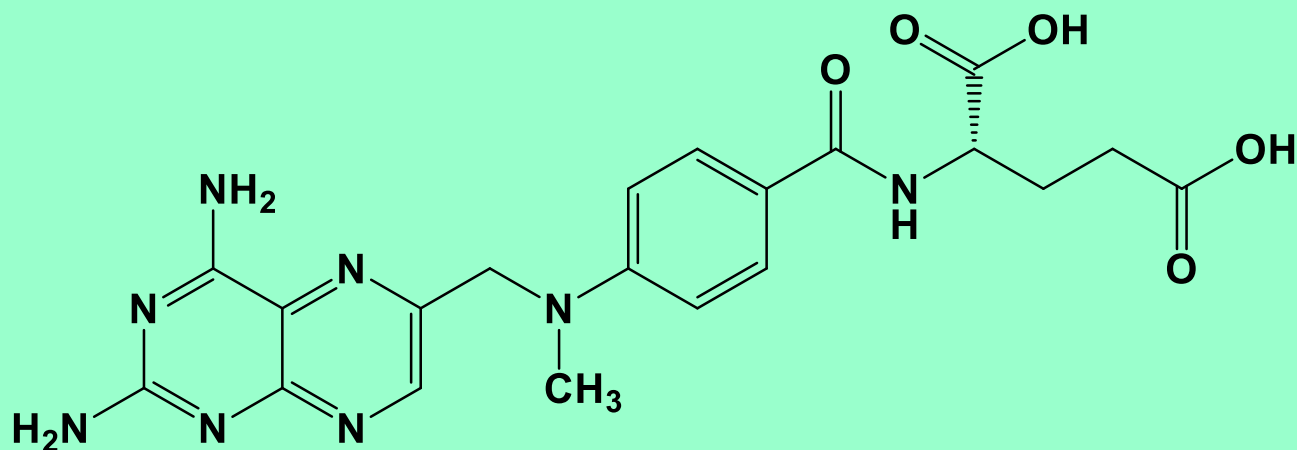


**Acetyl salicylic acid
(Aspirin)**



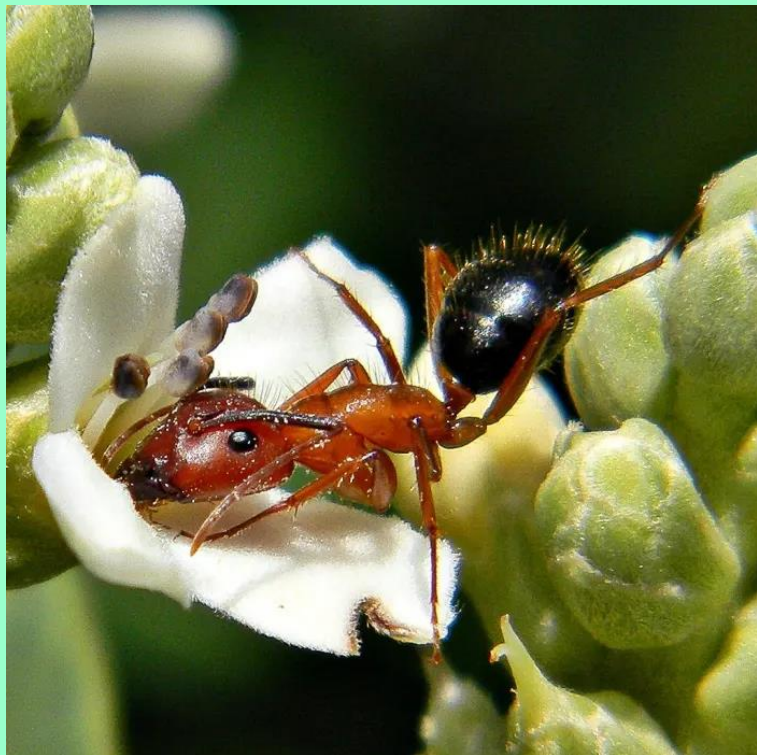
**Glutamic acid
(an amino acid)**

Organic Acids and Bases



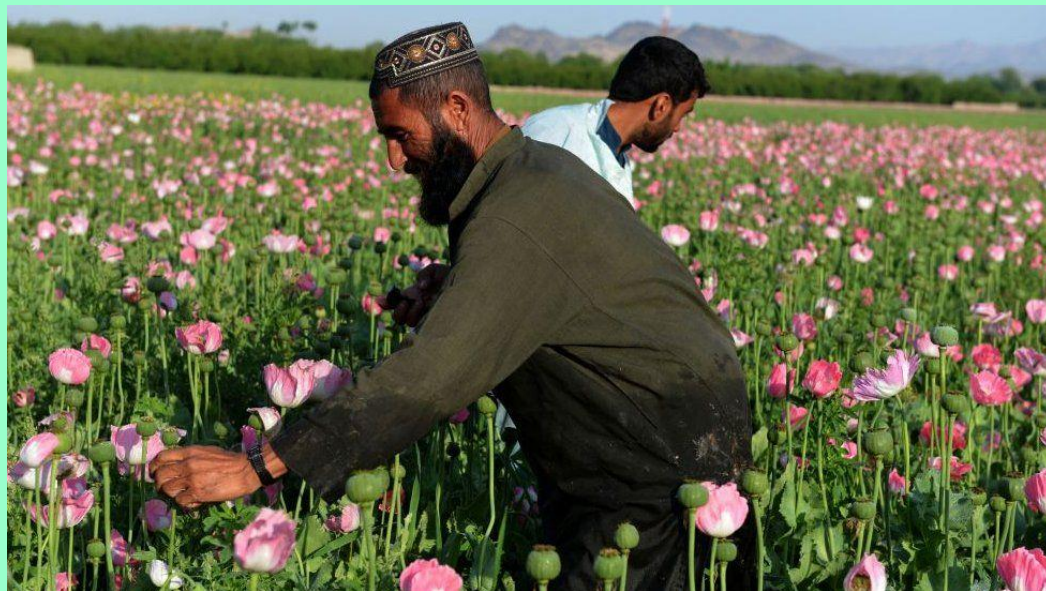
Methotrexate (MTX) for cancer treatment

Organic Acids and Bases



Carpenter ant

<https://www.flickr.com/photos/pondapple/9112104230/>



Opium poppy cultivated in Afghanistan (the world's largest source of opium, around 80% of the world's supply, UNODC).

(Source of morphine – used as a pain killer, and some other regulated drugs like codeine, heroin, etc.). One of the first “vegetable alkali”, or alkaloids to be isolated.

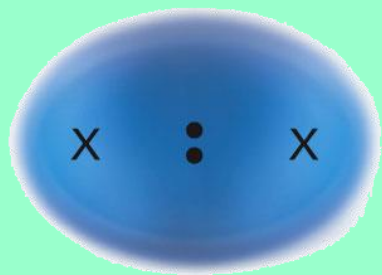
A Short Review on Previous Lecture

- In our previous lectures 1 & 2, we learned how *covalent bonds between atoms are formed* with their corresponding molecular geometry due to hybridization.
- We've learned that *hybrid orbitals* (like sp , sp^2 , sp^3) can be formed by mixing s - & p -orbitals. In some cases, d -orbitals are also involved in the like of sulphur & phosphorous atoms.
- Today, we're going to explore the consequences of these covalent bonds when bonding electrons are not shared equally between the two atoms.
- **What type of covalent bond will it form?**

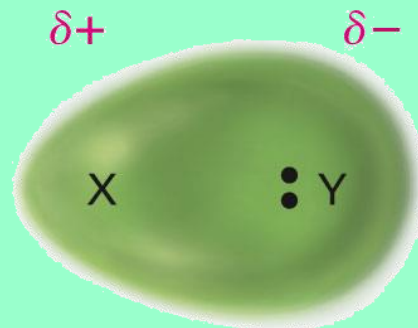
Bond Polarity

Examine the type of bonds below:

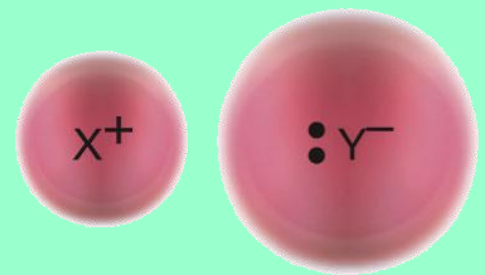
- ✿ **Covalent bonds can have ionic character also**
 - Bonding electrons attracted more strongly by one atom than by the other
 - Electron distribution between atoms is not symmetrical



Covalent bond



Polar covalent bond



Ionic bond

Bond Polarity

- Such unsymmetrical electron distribution in the bond results in neither fully ionic nor fully covalent but are somewhere between the two extremes
- Such bonds are called polar covalent bonds
- This bond polarity is due to the differences in electronegativity of the bonded atoms
- As a loose guide, bonds between atoms whose electronegativities differ by less than 0.5 are nonpolar covalent

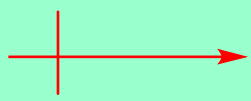
Bond Polarity

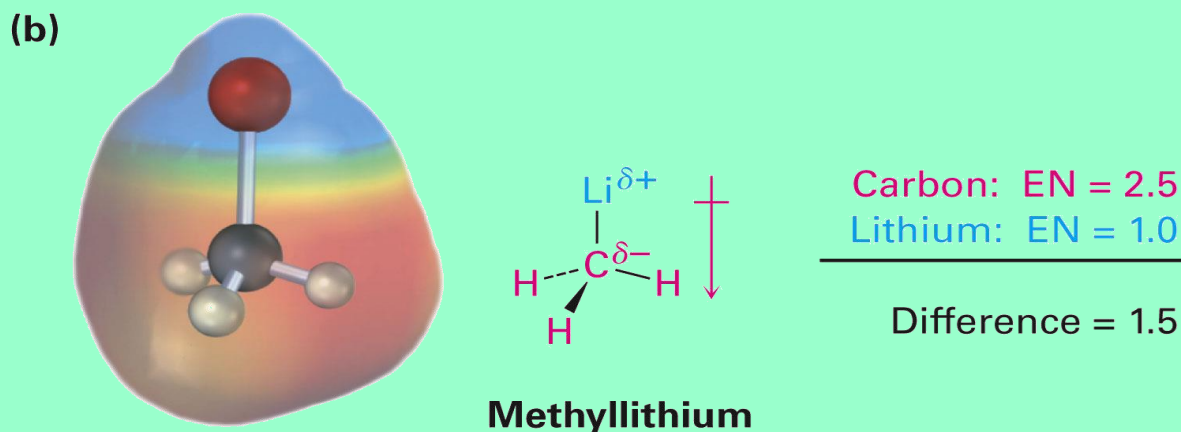
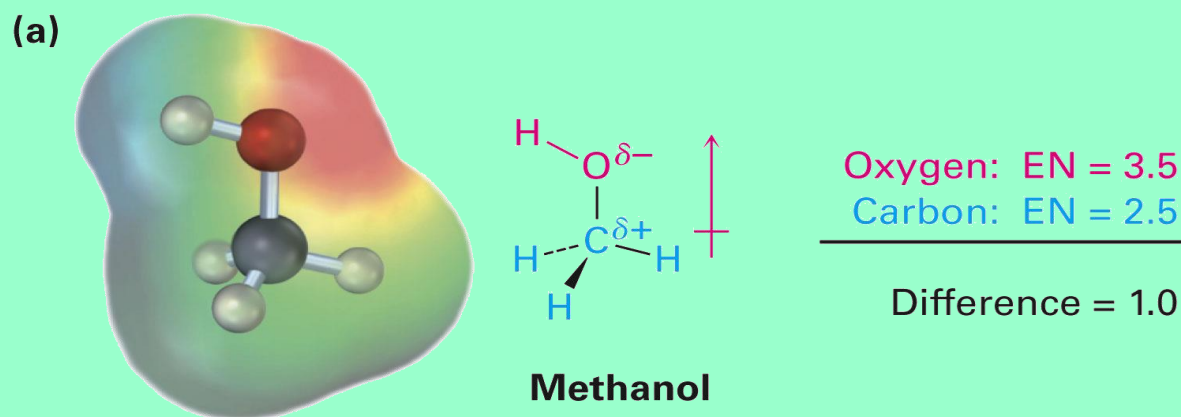
- Bonds between atoms whose electronegativities differ by 0.5 – 2 are polar covalent
- And bonds between atoms whose electronegativities differ by more than 2 are *largely ionic*
- The value of electronegativities shown in the chart on the next slides are based on an arbitrary scale (in measurements, the absolute value can't be obtained)
- Fluorine is the most electronegative (EN = 4.0)

Bond Polarity

- Cesium the least (EN = 0.7)
- Red blocks are the most electronegative, yellow are medium and dark greens are the least electronegative

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

- We use the lower case Greek letter delta (δ) to indicate partial charges
- Note also the representation of crossed arrow  to indicate the direction of bond polarity



Consequences of Bond Polarity

- When we talk about the atom's ability to polarize a bond, we often use the term *inductive effect* which is simply the shifting of electrons in σ bond in response to the electronegativity of nearby atoms
- Since electronegativities affect bond polarity, this will also affect the substance's acidity or basicity.
- Now let's take a look at the concept of acidity and basicity as applied to organic chemistry

Acids and Bases: The Brønsted–Lowry Definition

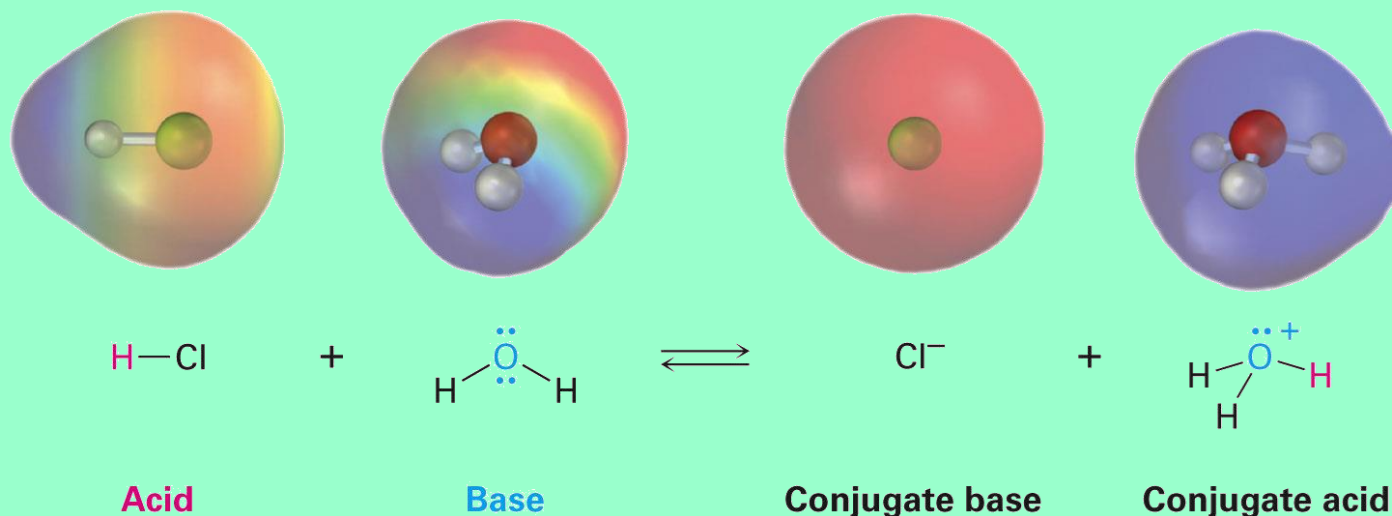
- The terms “acid” and “base” can have different meanings in different contexts
- For that reason, we specify the usage with more complete terminology
- The idea that acids are solutions containing a lot of “ H^+ ” and bases are solutions containing a lot of “ OH^- ” is not very useful in organic chemistry
- Instead, Brønsted–Lowry theory defines acids and bases by their role in reactions *that transfer protons (H^+) between donors and acceptors*

Brønsted Acids and Bases

- “Brønsted-Lowry” is usually shortened to “Brønsted”
- A Brønsted acid is a substance that donates a hydrogen ion (H^+)
- A Brønsted base is a substance that accepts the H^+
- “proton” is a synonym for H^+ (*loss of an electron from H leaving the bare nucleus*) - a proton

The Reaction of HCl with H₂O

- When HCl gas dissolves in water, a Brønsted acid-base reaction occurs
- HCl donates a proton to a water molecule, yielding hydronium ion (H₃O⁺) and Cl⁻
- This acid-base reaction is reversible, so we can write them with double, forward-and-backward arrows.



H_2O acting as an Acid/Base

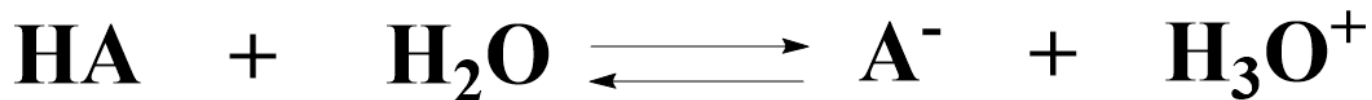
- Notice that water can act either as an acid or a base, depending on the circumstances
- In its reaction with HCl , water is a base that accepts a proton to give the hydronium ion, H_3O^+
- In its reaction with ammonia (NH_3), water act as an acid that donates a proton to give ammonium ion (NH_4^+) and hydroxide ion, HO^-

Quantitative Measures of Acid Strength

- In general, acids differ in their ability to donate a proton, H^+
- Organic acids are not that strong as mineral acids such as HCl that react completely with water
- Organic acids such as acetic acid (CH_3CO_2H), react only slightly with water
- The exact strength of a certain acid (we denote acid as HA) in water solution is described using the **acidity constant (K_a)** for the acidity dissociation equilibrium (K_{eq})

Quantitative Measures of Acid Strength

- Below is our equilibrium expression, we usually ignored the concentration of the solvent
- The brackets around the substance refer to the concentration of the enclosed species in moles per liter.



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

Quantitative Measures of Acid Strength

- Stronger acids have their equilibria toward the right and thus have larger K_a , whereas weaker acids have their equilibria toward the left and have smaller K_a
- Because of this, the range of K_a for different acids is enormous, running from 10^{15} for the strongest acids to about 10^{-60} for the weakest one
- To aid our understanding and have a rough feeling of which acids are “strong” and which are “weak”, it is practical to express this value in terms of pK_a (the negative logarithm of the K_a)

$$pK_a = -\log K_a$$

Quantitative Measures of Acid Strength

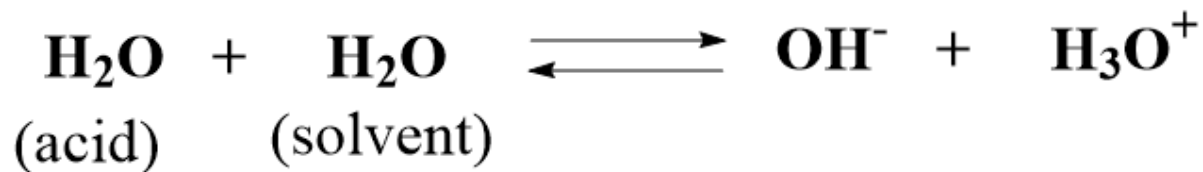
- A *stronger* acid (larger K_a) has a smaller pK_a
- A *weaker* acid (smaller K_a) has a larger pK_a

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
	H_2O	Water	15.74	HO^-	Hydroxide ion	
	HCN	Hydrocyanic acid	9.31	CN^-	Cyanide ion	
	H_2PO_4^-	Dihydrogen phosphate ion	7.21	HPO_4^{2-}	Hydrogen phosphate ion	
	$\text{CH}_3\text{CO}_2\text{H}$	Acetic acid	4.76	CH_3CO_2^-	Acetate ion	
	H_3PO_4	Phosphoric acid	2.16	H_2PO_4^-	Dihydrogen phosphate ion	
	HNO_3	Nitric acid	-1.3	NO_3^-	Nitrate ion	
Stronger acid	HCl	Hydrochloric acid	-7.0	Cl^-	Chloride ion	Weaker base

Quantitative Measures of Acid Strength

- Notice that the pK_a value of water is 15.74 because water is both the acid and the solvent



$$K_a = \frac{[\text{H}_3\text{O}^+][\text{A}^-]}{[\text{HA}]} = \frac{[\text{H}_3\text{O}^+][\text{OH}^-]}{[\text{H}_2\text{O}]} = \frac{[1.0 \times 10^{-7}][1.0 \times 10^{-7}]}{[55.4]} = 1.8 \times 10^{-16}$$

$$pK_a = -\log \text{ of } 1.8 \times 10^{-16}$$

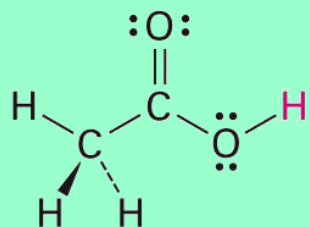
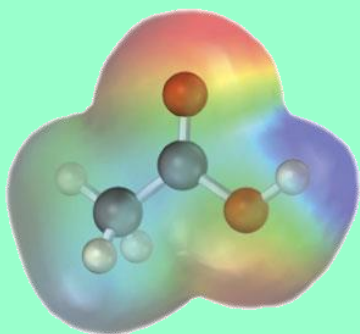
$$pK_a = 15.74$$

- The numerator is the ion-product constant for water, $K_w = [\text{H}_3\text{O}^+][\text{OH}^-] = 1.00 \times 10^{-14}$ and the denominator is the molar concentration of pure water, $[\text{H}_2\text{O}] = 55.4 \text{ M at } 25^\circ\text{C}$

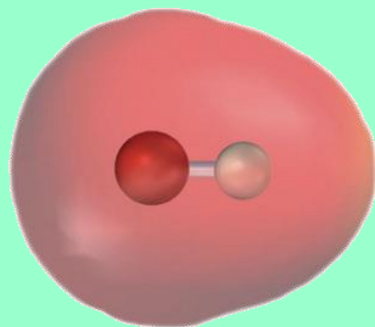
Predicting Acid–Base Reactions from pK_a Values

- With the pK_a values given or calculated, it is now easy for us to predict whether a given acid-base reaction will take place
- **How?** Always remember that proton (H^+) will always transfer from the *stronger acid* to the *stronger base*
- The stronger base always holds the proton more tightly
- Now let us take a look at the reaction of acetic acid and hydroxide ion on the next slide, **will the reaction be significant?**

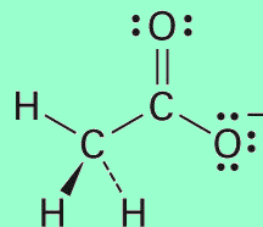
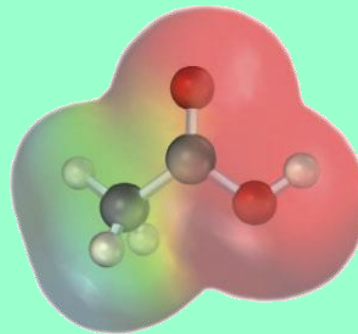
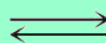
Predicting Acid–Base Reactions from pK_a Values



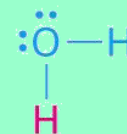
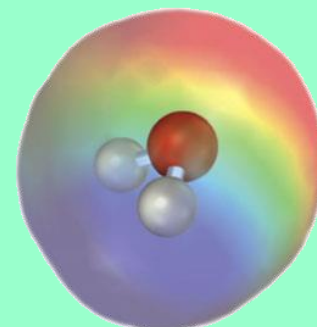
Acetic acid
($pK_a = 4.76$)



Hydroxide ion



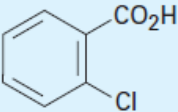
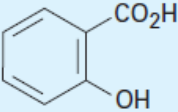
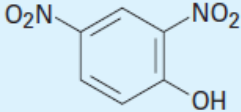
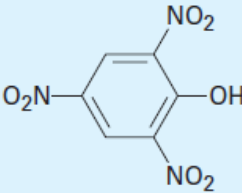
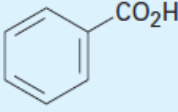
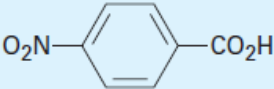
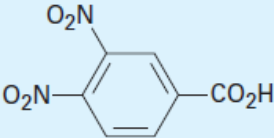
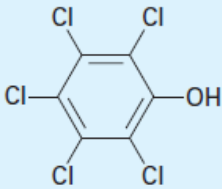
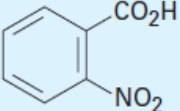
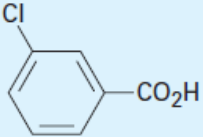
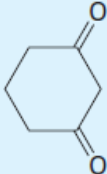
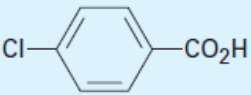
Acetate ion

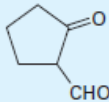
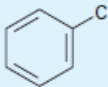
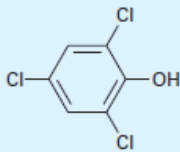
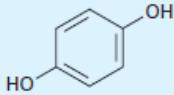
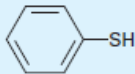
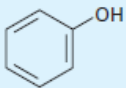
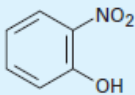
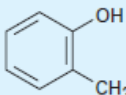
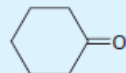
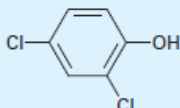
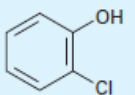
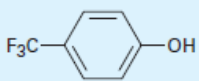
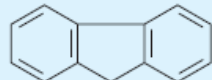
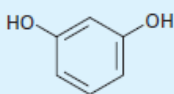
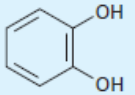

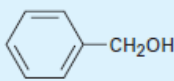
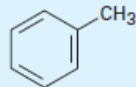



Water
($pK_a = 15.74$)

Predicting Acid–Base Reactions from pK_a Values

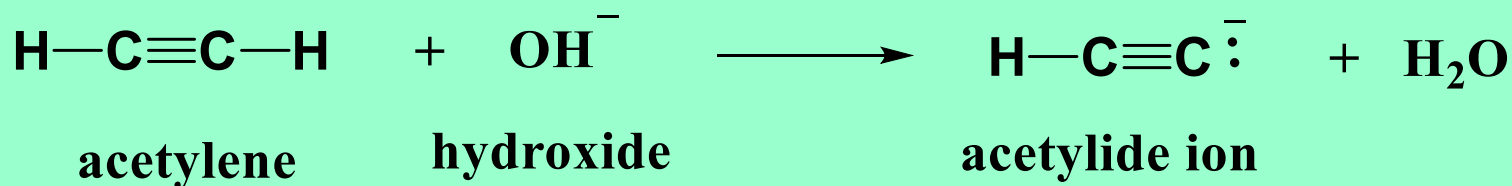
- Some very important concepts to remember in predicting acid-base reaction:
- To predict acid-base reactivity is to remember that the products of an acid-base reaction must be **more stable** than the reactants
- In other words, **the product acid must be weaker and less reactive than the starting acid**, and the product base must be weaker and less reactive than the starting base.

Compound	pK _a	Compound	pK _a	Compound	pK _a
CH ₃ SO ₃ H	-1.8		3.0	CH ₂ BrCH ₂ CO ₂ H	4.0
CH(NO ₂) ₃	0.1		3.0		4.1
	0.3	CH ₂ ICO ₂ H	3.2		4.2
CCl ₃ CO ₂ H	0.5	CHOCO ₂ H	3.2	H ₂ C=CHCO ₂ H	4.2
CF ₃ CO ₂ H	0.5		3.4	HO ₂ CCH ₂ CH ₂ CO ₂ H	4.2; 5.7
CBr ₃ CO ₂ H	0.7		3.5	HO ₂ CCH ₂ CH ₂ CH ₂ CO ₂ H	4.3; 5.4
HO ₂ CC≡CCO ₂ H	1.2; 2.5	HSCH ₂ CO ₂ H	3.5; 10.2		4.5
HO ₂ CCO ₂ H	1.2; 3.7	CH ₂ (NO ₂) ₂	3.6	H ₂ C=C(CH ₃)CO ₂ H	4.7
CHCl ₂ CO ₂ H	1.3	CH ₃ OCH ₂ CO ₂ H	3.6	CH ₃ CO ₂ H	4.8
CH ₂ (NO ₂)CO ₂ H	1.3	CH ₃ COCH ₂ CO ₂ H	3.6	CH ₃ CH ₂ CO ₂ H	4.8
HC≡CCO ₂ H	1.9	HOCH ₂ CO ₂ H	3.7	(CH ₃) ₃ CCO ₂ H	5.0
(Z) HO ₂ CCH=CHCO ₂ H	1.9; 6.3	HCO ₂ H	3.7	CH ₃ COCH ₂ NO ₂	5.1
	2.4		3.8		5.3
CH ₃ COCO ₂ H	2.4		4.0	O ₂ NCH ₂ CO ₂ CH ₃	5.8
NCCH ₂ CO ₂ H	2.5				
CH ₃ C≡CCO ₂ H	2.6				
CH ₂ FCO ₂ H	2.7				
CH ₂ ClCO ₂ H	2.8				
HO ₂ CCH ₂ CO ₂ H	2.8; 5.6				
CH ₂ BrCO ₂ H	2.9				

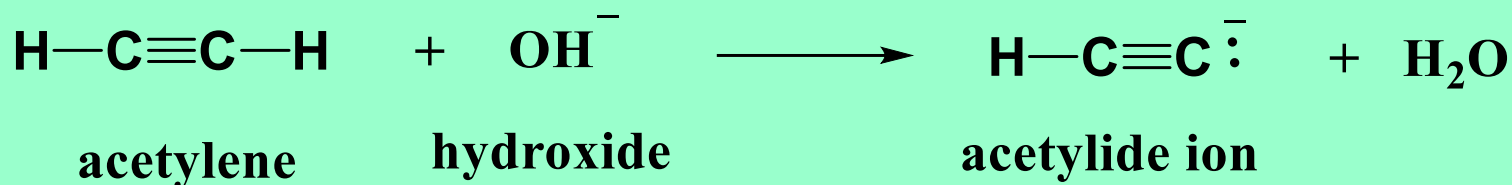
Compound	pK _a	Compound	pK _a	Compound	pK _a
	5.8		9.4	$\text{H}_2\text{C}=\text{CHCH}_2\text{OH}$	15.5
	6.2		9.9; 11.5	$\text{CH}_3\text{CH}_2\text{OH}$	16.0
	6.6		9.9	$\text{CH}_3\text{CH}_2\text{CH}_2\text{OH}$	16.1
HCO_3H	7.1	$\text{CH}_3\text{COCH}_2\text{SOCH}_3$	10.0	$\text{CH}_3\text{COCH}_2\text{Br}$	16.1
	7.2		10.3		16.7
$(\text{CH}_3)_2\text{CHNO}_2$	7.7	CH_3NO_2	10.3	CH_3CHO	17
	7.8	CH_3SH	10.3	$(\text{CH}_3)_2\text{CHCHO}$	17
$\text{CH}_3\text{CO}_3\text{H}$	8.2	$\text{CH}_3\text{COCH}_2\text{CO}_2\text{CH}_3$	10.6	$(\text{CH}_3)_2\text{CHOH}$	17.1
	8.5	CH_3COCHO	11.0	$(\text{CH}_3)_3\text{COH}$	18.0
$\text{CH}_3\text{CH}_2\text{NO}_2$	8.5	$\text{CH}_2(\text{CN})_2$	11.2	CH_3COCH_3	19.3
	8.7	$\text{CCl}_3\text{CH}_2\text{OH}$	12.2		23
$\text{CH}_3\text{COCH}_2\text{COCH}_3$	9.0	Glucose	12.3	$\text{CH}_3\text{CO}_2\text{CH}_2\text{CH}_3$	25
	9.3; 11.1	$(\text{CH}_3)_2\text{C}=\text{NOH}$	12.4	$\text{HC}\equiv\text{CH}$	25
	9.3; 12.6	$\text{CH}_2(\text{CO}_2\text{CH}_3)_2$	12.9	CH_3CN	25
		$\text{CHCl}_2\text{CH}_2\text{OH}$	12.9	$\text{CH}_3\text{SO}_2\text{CH}_3$	28
		$\text{CH}_2(\text{OH})_2$	13.3	$(\text{C}_6\text{H}_5)_3\text{CH}$	32
		$\text{HOCH}_2\text{CH}(\text{OH})\text{CH}_2\text{OH}$	14.1	$(\text{C}_6\text{H}_5)_2\text{CH}_2$	34
		$\text{CH}_2\text{ClCH}_2\text{OH}$	14.3	CH_3SOCH_3	35
			15.0	NH_3	36
			15.4	$\text{CH}_3\text{CH}_2\text{NH}_2$	36
		CH_3OH	15.5	$(\text{CH}_3\text{CH}_2)_2\text{NH}$	40
					41
					43
				$\text{H}_2\text{C}=\text{CH}_2$	44
				CH_4	~60

An acidity list covering more than 5000 organic compounds has been published: E.P. Serjeant and B. Dempsey (eds.), "Ionization Constants of Organic Acids in Aqueous Solution," IUPAC Chemical Data Series No. 23, Pergamon Press, Oxford, 1979.

➤ Water has $pK_a = 15.74$, and acetylene has $pK_a = 25$. Which is the stronger acid? Will hydroxide ions react to a significant extent with acetylene?



In comparing two acids, the one with the lower pK_a is the stronger acid. Thus, water is a stronger acid than acetylene and gives up H^+ more easily.



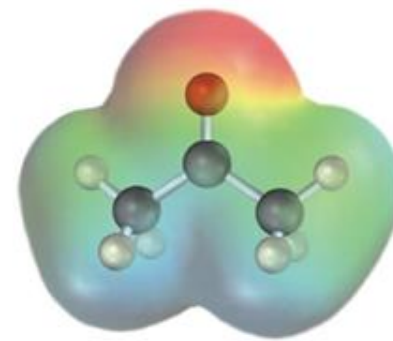
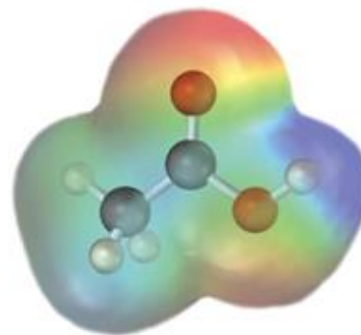
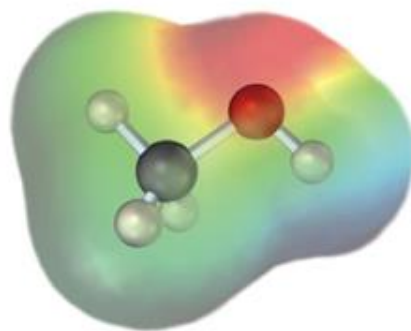
Analysis: Because water is stronger acid and gives up H^+ more easily than acetylene does, the HO^- ion must have less affinity for H^+ than the $\text{HC}\equiv\text{C}^-$ ion has.

In other words, the anion of acetylene is a stronger base than the hydroxide ion, and the reaction will not proceed significantly.

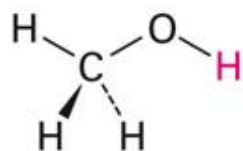
What are Organic Acids and Bases

- The reaction patterns of organic compounds often are acid-base combinations
- The transfer of a proton from strong Brønsted acid to a Brønsted base, for example, is a very fast process and will always occur along with other reactions
- Organic acids are characterized by the presence of a positively polarized hydrogen atom (blue in electrostatic potential maps) and are of two main kinds:

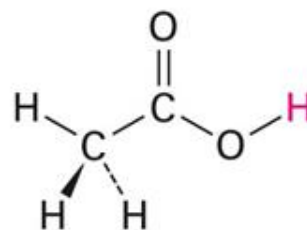
What are Organic Acids and Bases



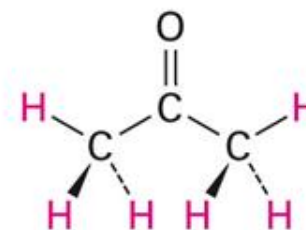
Some organic
acids



Methanol
($\text{pK}_a = 15.54$)



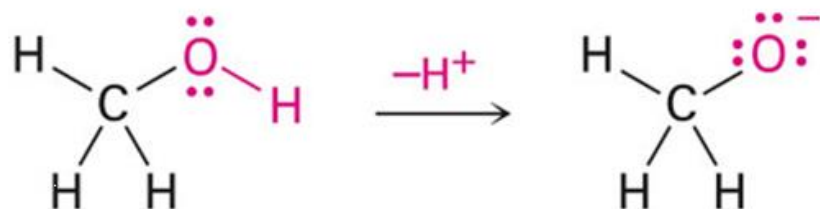
Acetic acid
($\text{pK}_a = 4.76$)



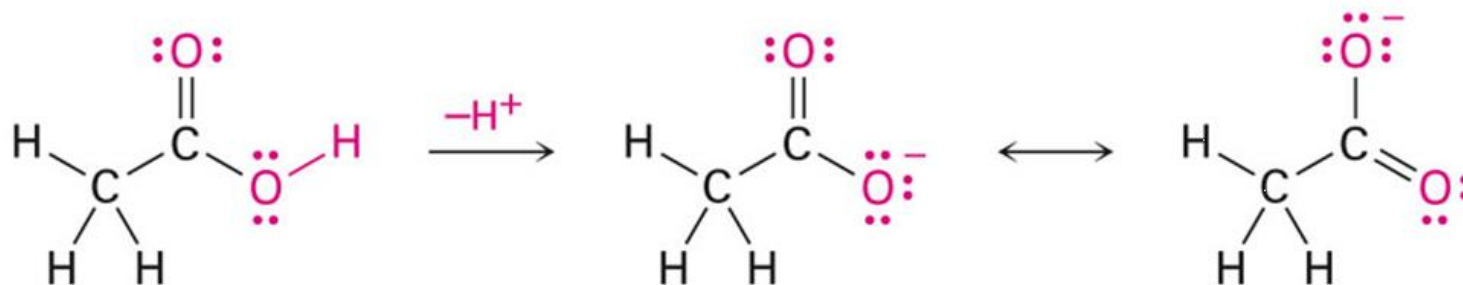
Acetone
($\text{pK}_a = 19.3$)

Organic Acids

- Those that lose a proton (H^+) from O–H, such as methanol and acetic acid



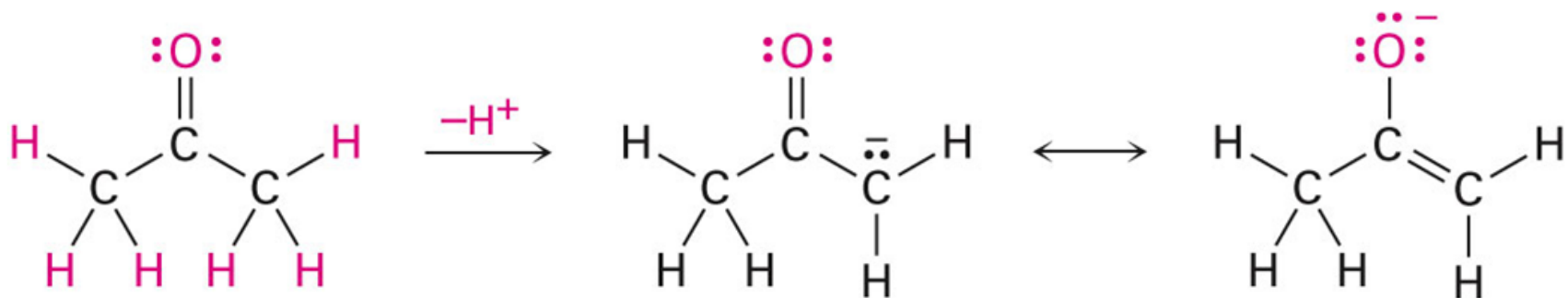
Anion is stabilized by having negative charge on a highly electronegative atom.



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

Organic Acids

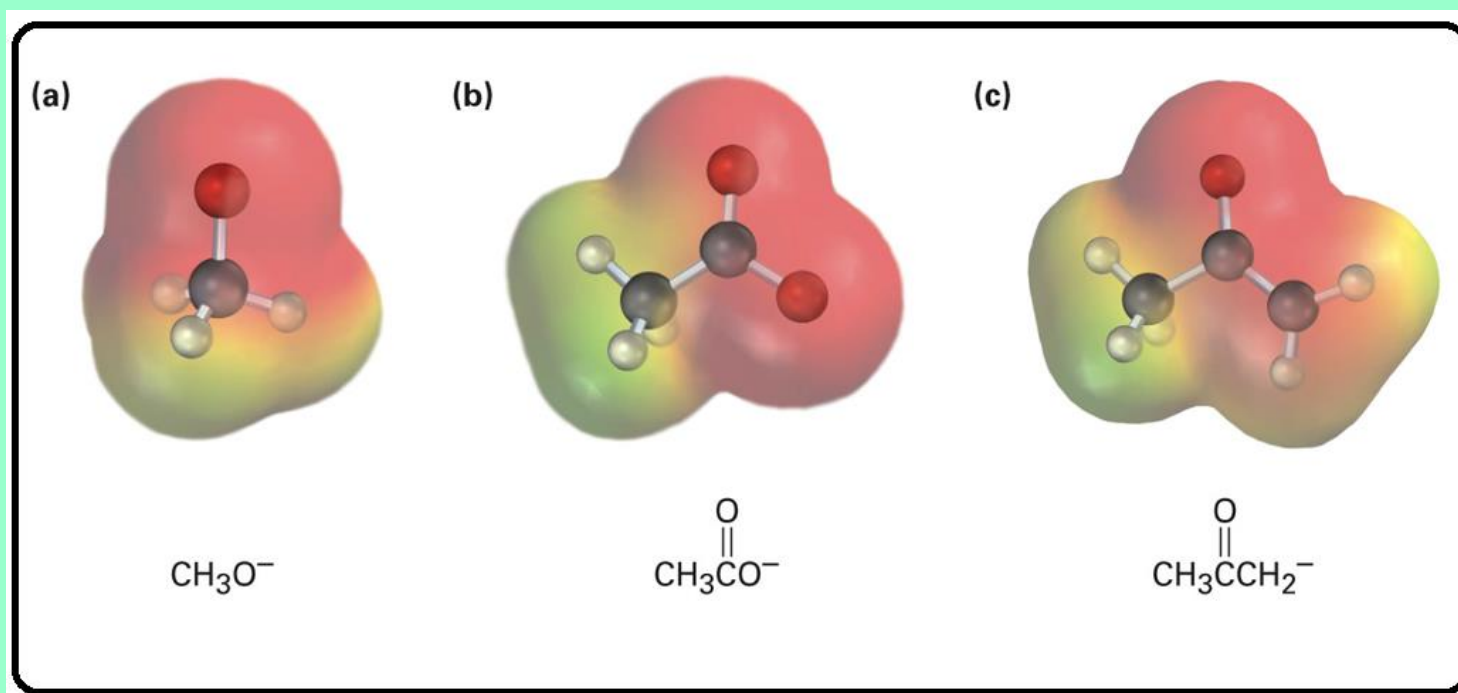
- Those that lose a proton (H^+) from C–H, usually from a carbon atom next to a C=O double bond ($\text{O}=\text{C}-\text{C}-\text{H}$)



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

Organic Acids

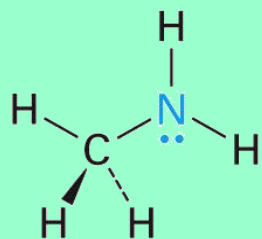
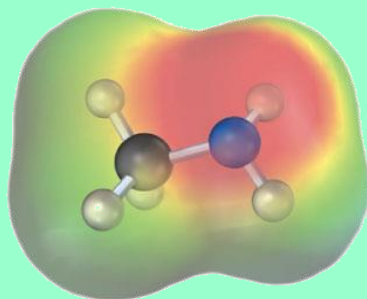
- Below are the electrostatic potential maps of the conjugate bases from methanol, acetic acid, and acetone
- All three shows a substantial amount of negative charge (red) on oxygen



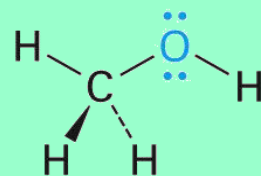
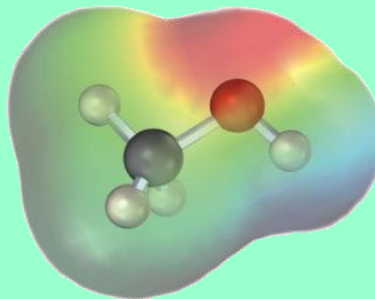
Organic Bases

- Organic bases are characterized by the presence of an atom (reddish in electrostatic potential maps) with a lone pair of electrons that can bond to a proton (H^+)
- Nitrogen-containing compounds derived from ammonia are the most common organic bases
- Oxygen-containing compounds can react as bases depending on the circumstances just like water
- Methanol, and acetone act as an acid when they donate a proton but as bases when their oxygen atom accepts a proton (H^+)

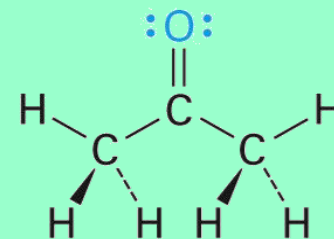
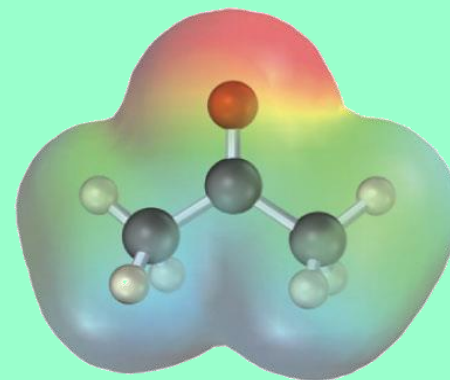
Organic Bases



Methylamine



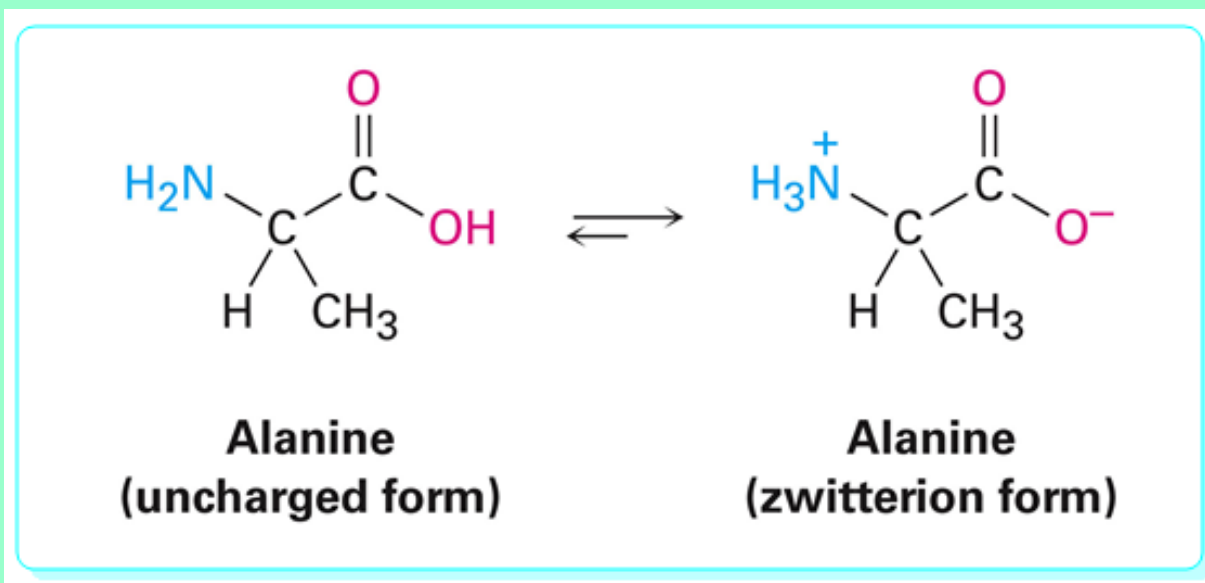
Methanol



Acetone

Some organic
bases

Organic Bases

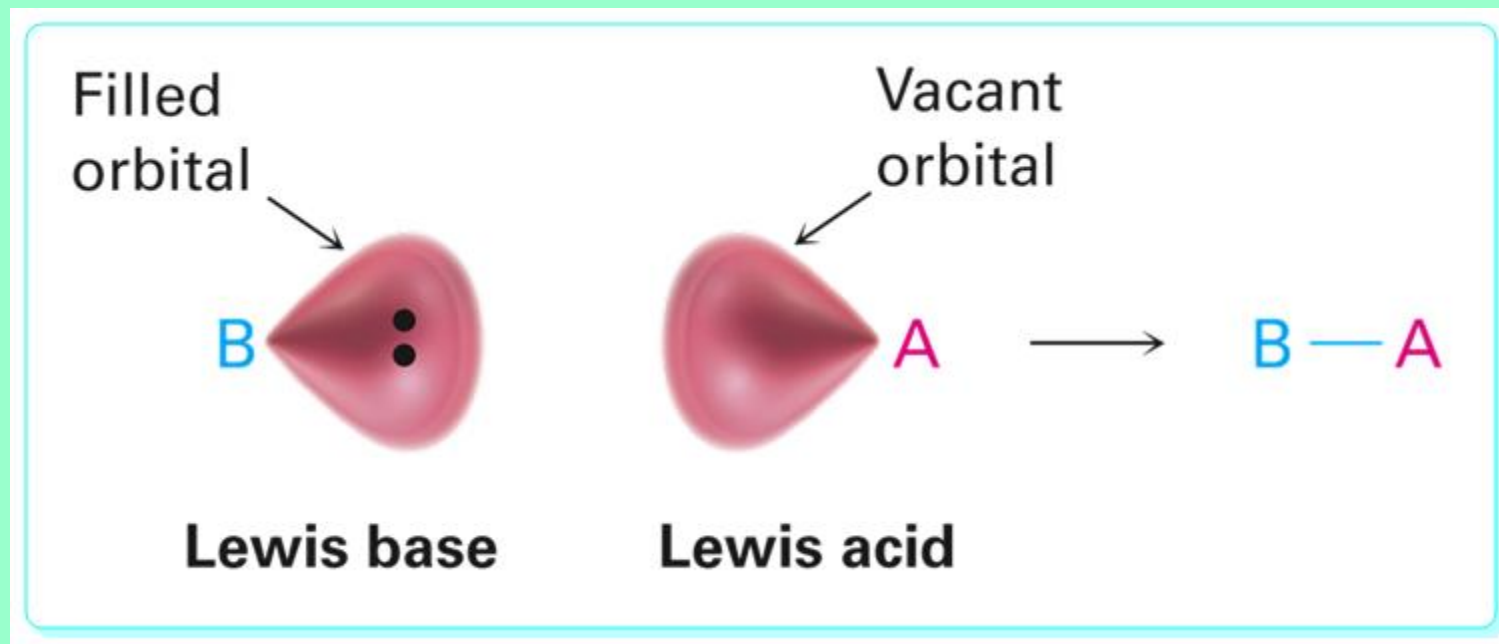


- Alanine, an amino acid and one of the building blocks of protein exist primarily in a doubly charged form called *zwitterion* rather than in the uncharged form
- The zwitterion form arises because alanine have both acidic and basic sites within the same molecule and therefore undergo an internal acid-base reaction

Acids and Bases: The Lewis Definition

- Lewis definition of acids and bases is much more broader and encompassing since it is not limited to substances that donate or accept protons (H^+)
- A **Lewis acid** is a substance that accepts an electron pair, while a **Lewis base** is a substance that donates an electron pair
- This donated electron pair is shared between the acid and the base in a covalent bond
- Now let us explore the mechanism of the Lewis acid –base reaction

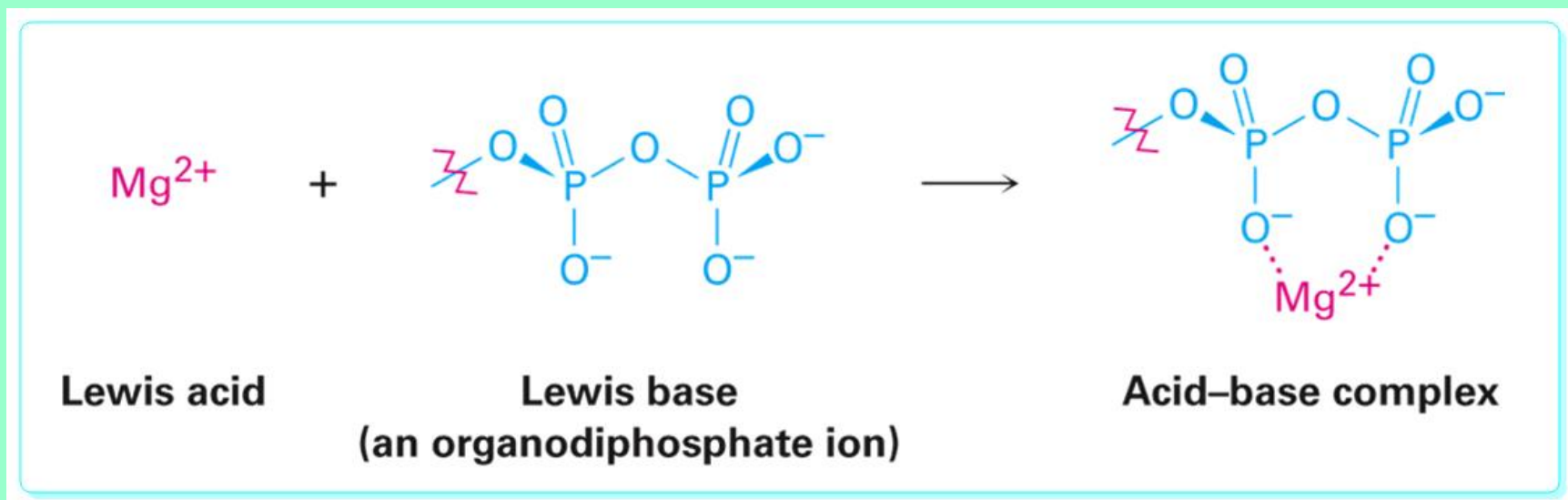
Acids and Bases: The Lewis Definition



- Since Lewis acid is able to accept an electron pair, means that it must have either a **vacant, low-energy orbital or a polar bond to hydrogen** so that it can donate a proton (H^+ , having an empty $1s$ orbital)

Acids and Bases: The Lewis Definition

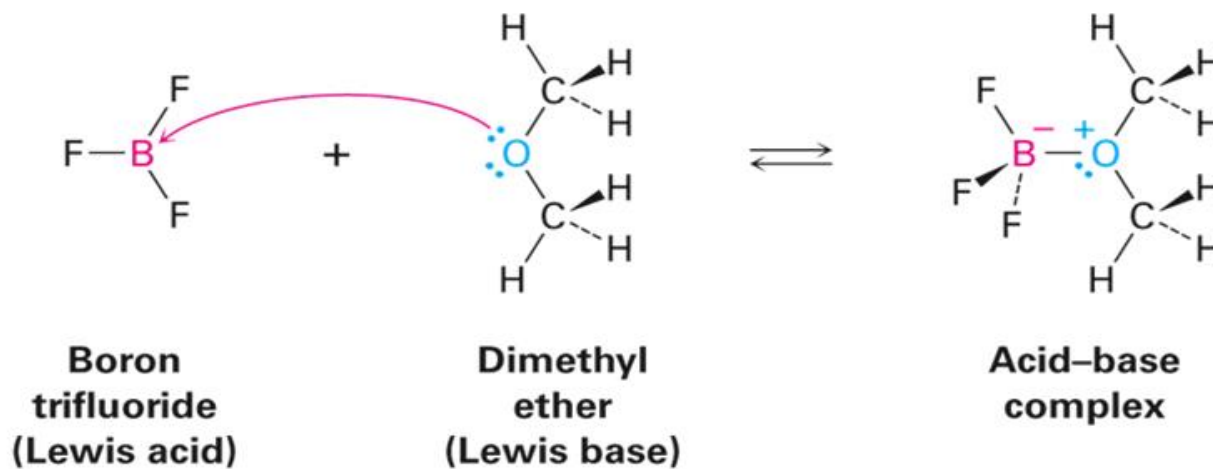
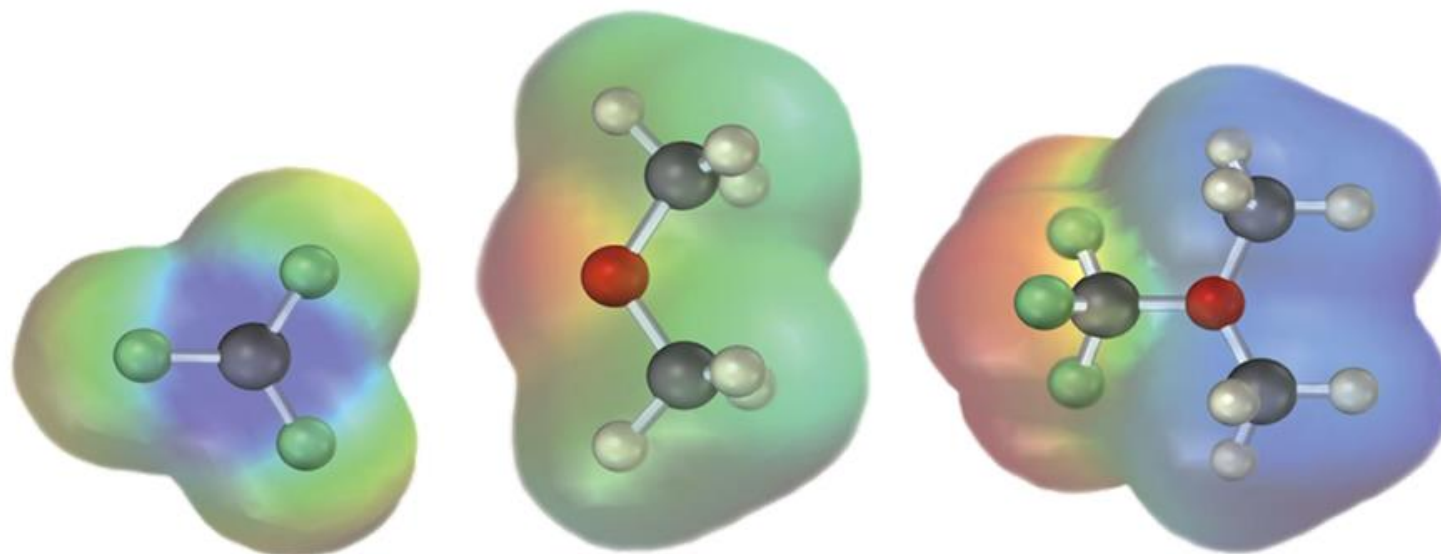
- The Lewis acid definition may now include metal cations, such as Mg^{2+} since they accept a pair of electrons when they form a bond to a base



Acids and Bases: The Lewis Definition

- **Group 3A elements, such as BF_3 and AlCl_3 , are Lewis acids because they have unfilled valence orbitals and can accept electron pairs from Lewis bases**
- **Transition-metal compounds, such as TiCl_4 , FeCl_3 , ZnCl_2 , and SnCl_4 , are Lewis acids**
- **Organic compounds that undergo addition reactions with Lewis bases (discussed later) are called electrophiles and therefore Lewis Acids**

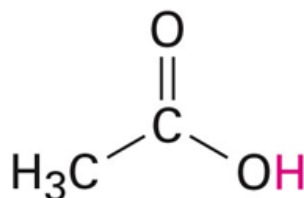
Acids and Bases: The Lewis Definition



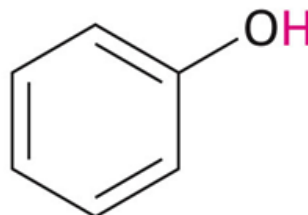
Acids and Bases: The Lewis Definition

Some
Lewis
acids

Some neutral proton donors:



A carboxylic acid



A phenol



An alcohol

Some cations:

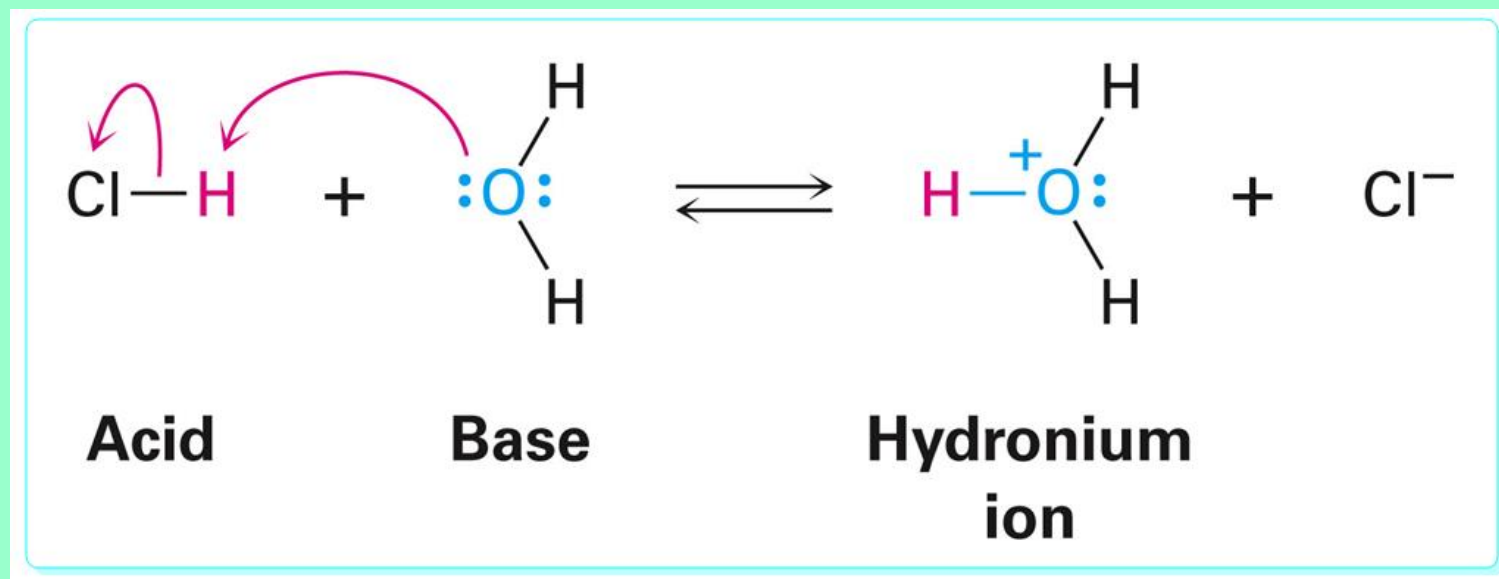


Some metal compounds:



Lewis Bases

- Lewis definition of a base – (a compound with a pair of nonbonding electrons that it can use to bond to a Lewis acid – similar to the Brønsted-Lowry definition)
- Water, can be a good example in this case:

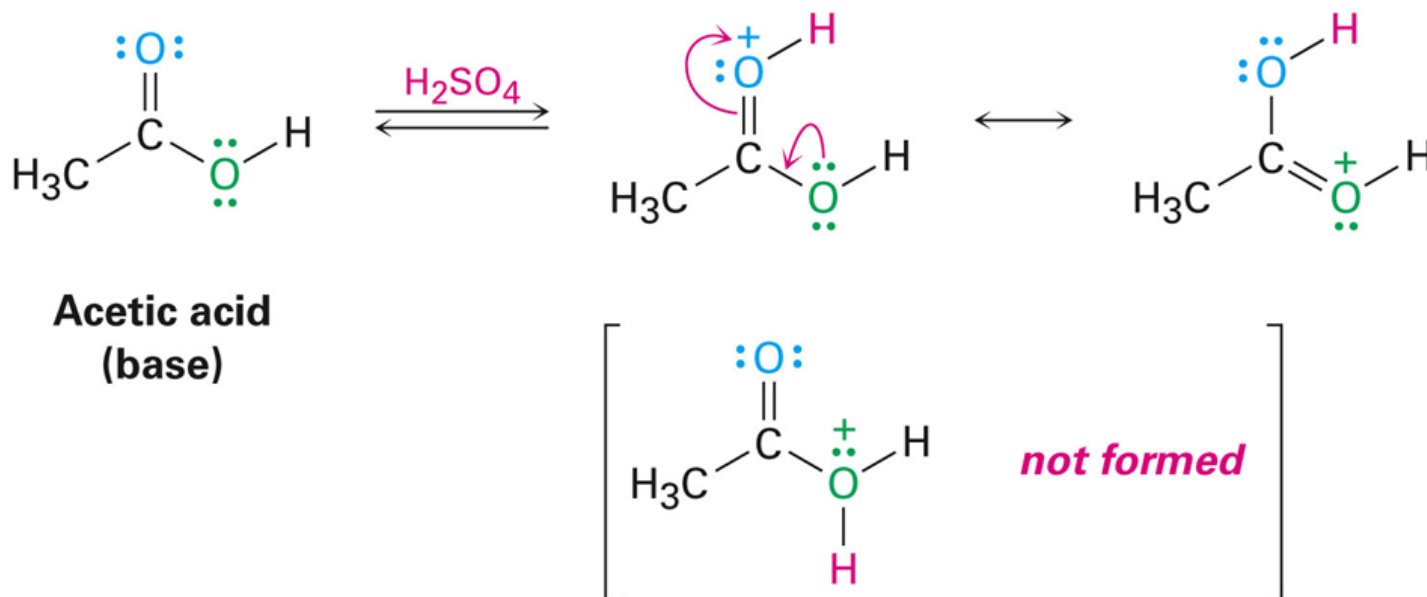


Lewis Bases

- Lewis bases can accept protons as well as Lewis acids, therefore the definition encompasses that for Brønsted bases
- Most oxygen- and nitrogen-containing organic compounds are Lewis bases because they have lone pairs of electrons
- Some compounds can act as both acids and bases, depending on the reaction and circumstances, just in the case of alcohol and carboxylic acid which can act also as bases when their oxygen atom accepts a proton (H^+)

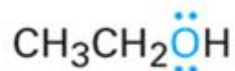
Lewis Bases

- In the case of acetic acid, protonation can happen either on the *doubly bonded* oxygen atom or on the *singly bonded* oxygen atom, and the more stable of the two protonation products is formed



Lewis Bases

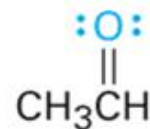
Some
Lewis
bases



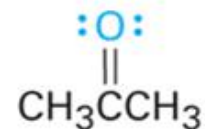
An alcohol



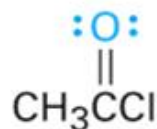
An ether



An aldehyde



A ketone



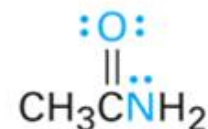
An acid chloride



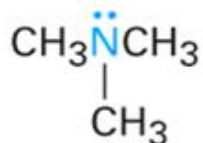
A carboxylic
acid



An ester



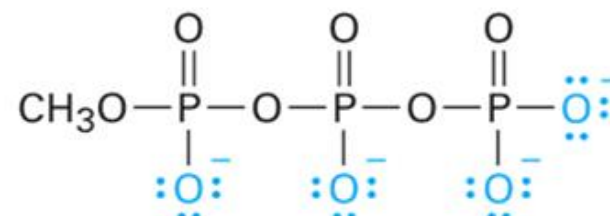
An amide



An amine



A sulfide



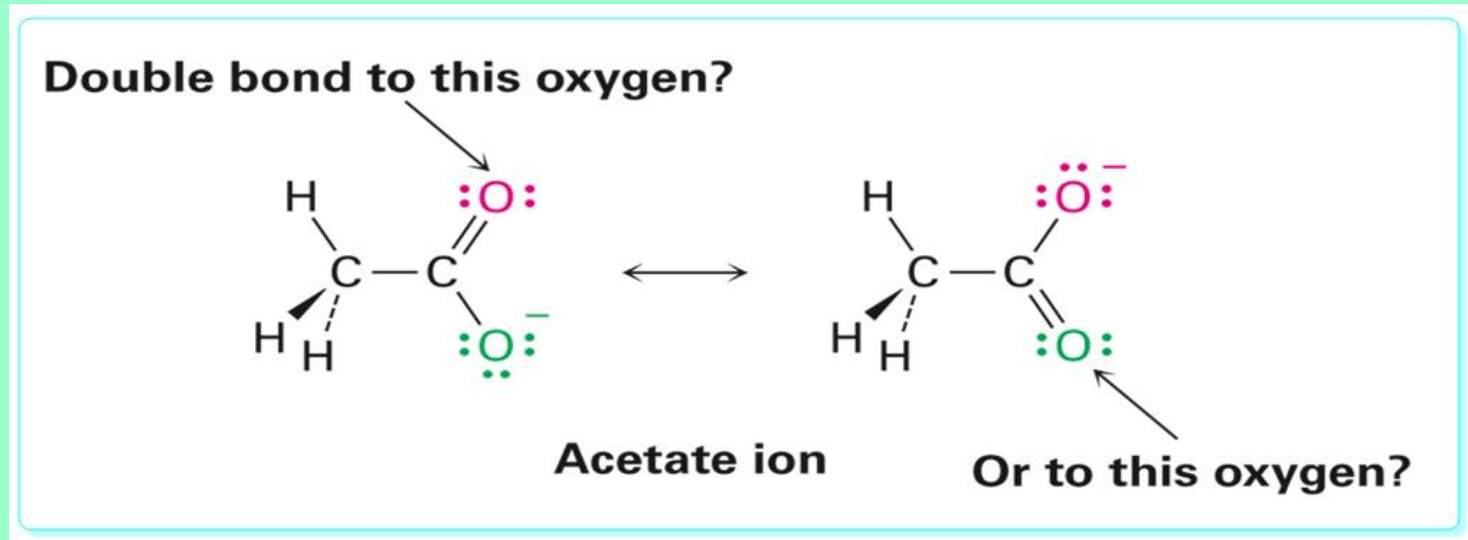
An organotriphosphate ion

Resonance

- There are some molecules that have structures that cannot be shown with a single representation
- In these cases we draw structures that contribute to the final structure but which differ in the position of the π bond(s) or lone pair(s)
- Such a structure is delocalized and is represented by resonance forms

Resonance

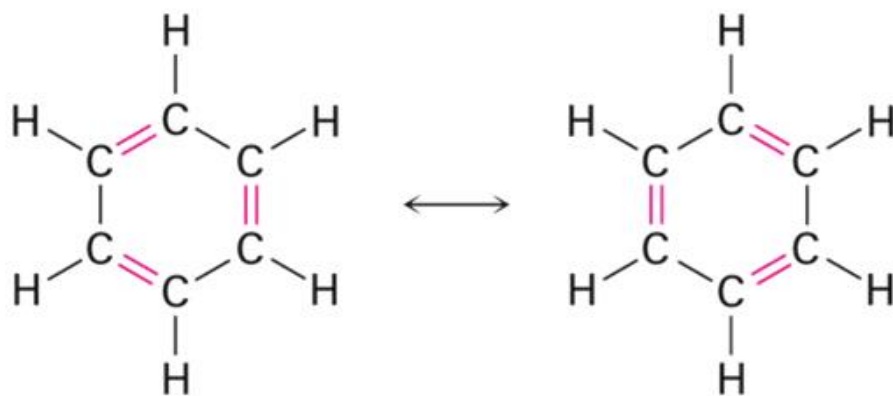
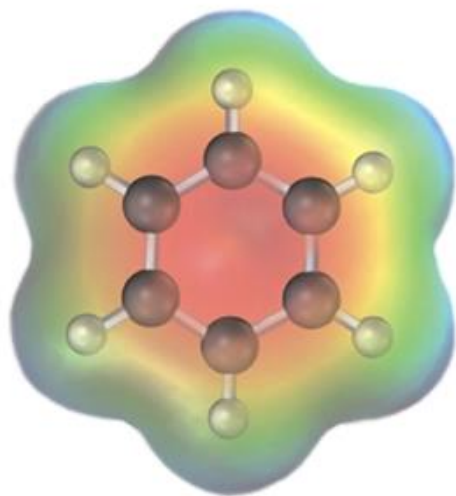
- The resonance forms are connected by a double-headed arrow



- A structure with resonance forms **does not** alternate between the forms
- Instead, it is a *hybrid* of the two resonance forms, so the structure is called a resonance hybrid

Resonance Hybrids

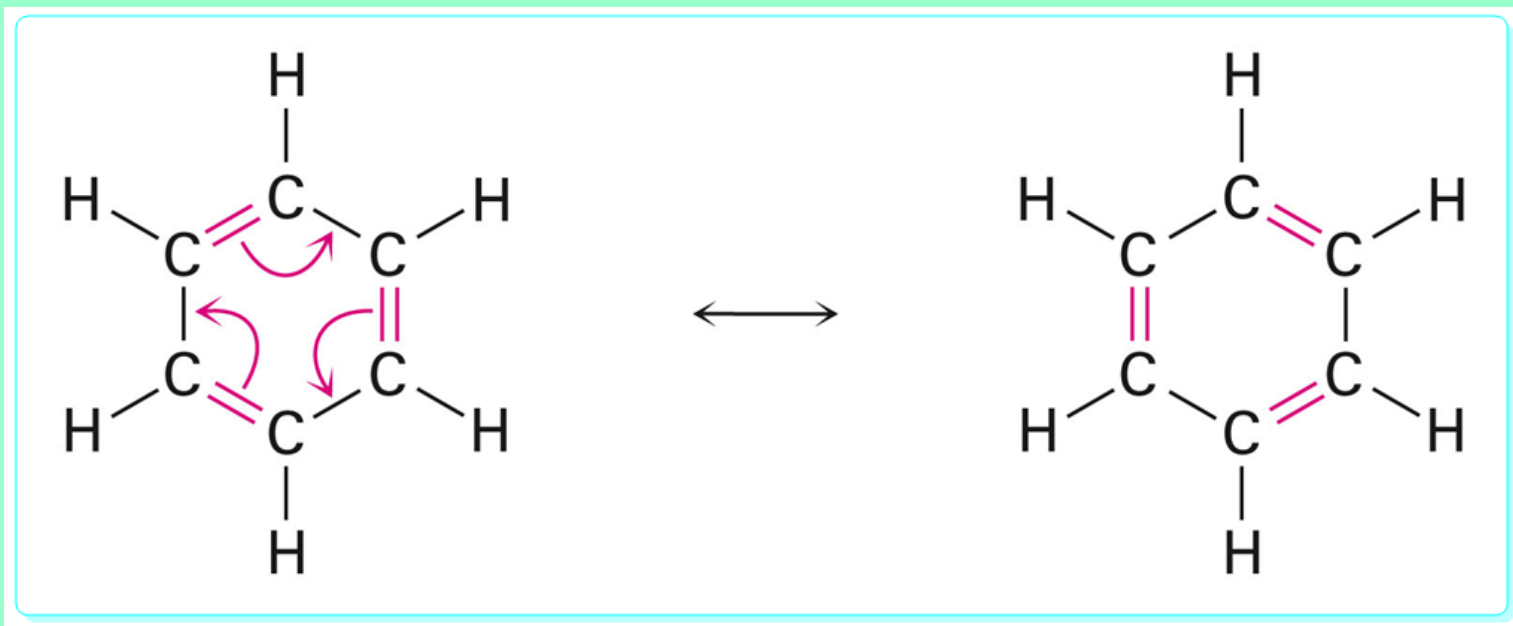
- For example, benzene (C_6H_6) has two resonance forms with alternating double and single bonds
- In the resonance hybrid, the actual structure, all its C-C bonds equivalent, midway between double and single



Benzene (two resonance forms)

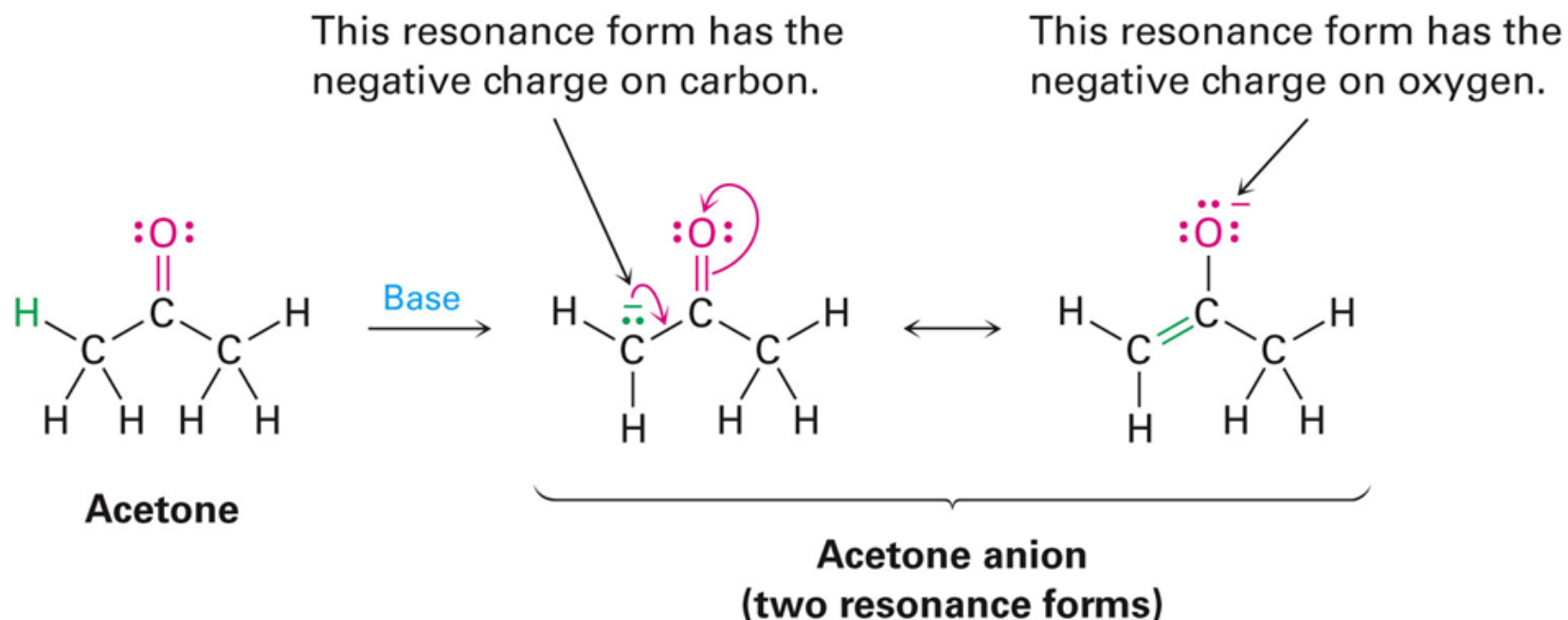
Rules for Resonance Forms

1. Individual resonance forms are imaginary - the real structure is a hybrid (only by knowing the contributors you can visualize the actual structure)
2. Resonance forms differ only in the placement of their π or nonbonding electrons



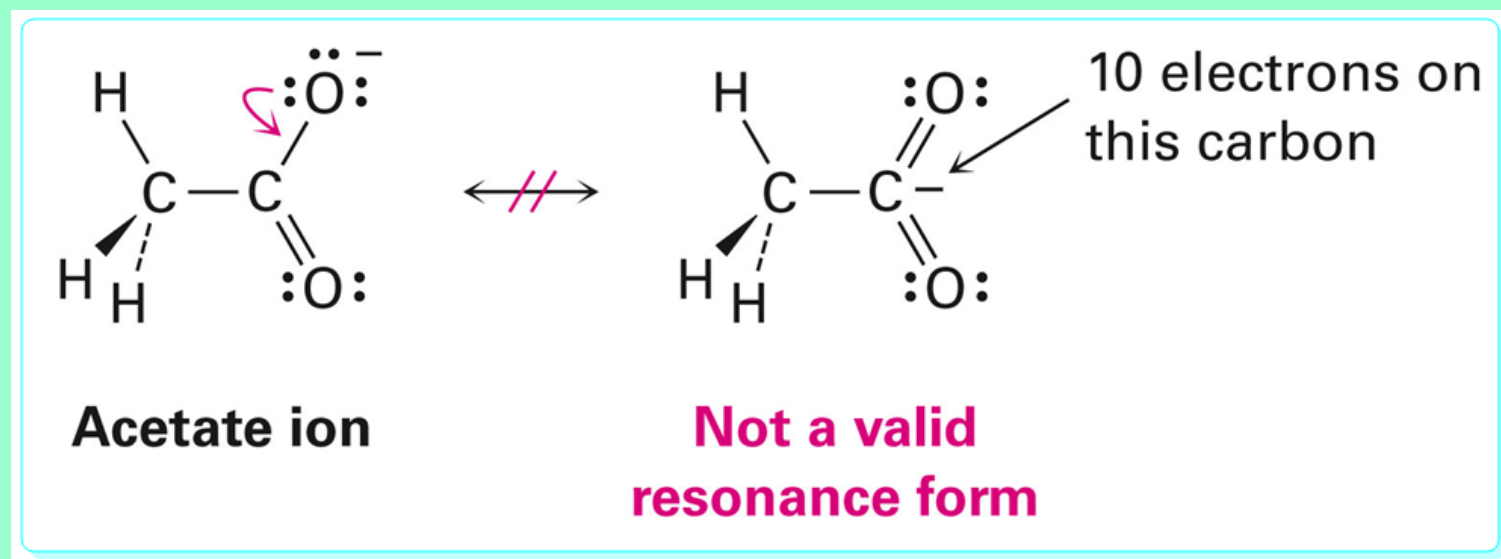
Rules for Resonance Forms

3. Different resonance forms of a substance don't have to be equivalent



Rules for Resonance Forms

- 4) Resonance forms must be valid Lewis structures and obey normal rules of valency: the octet rule applies



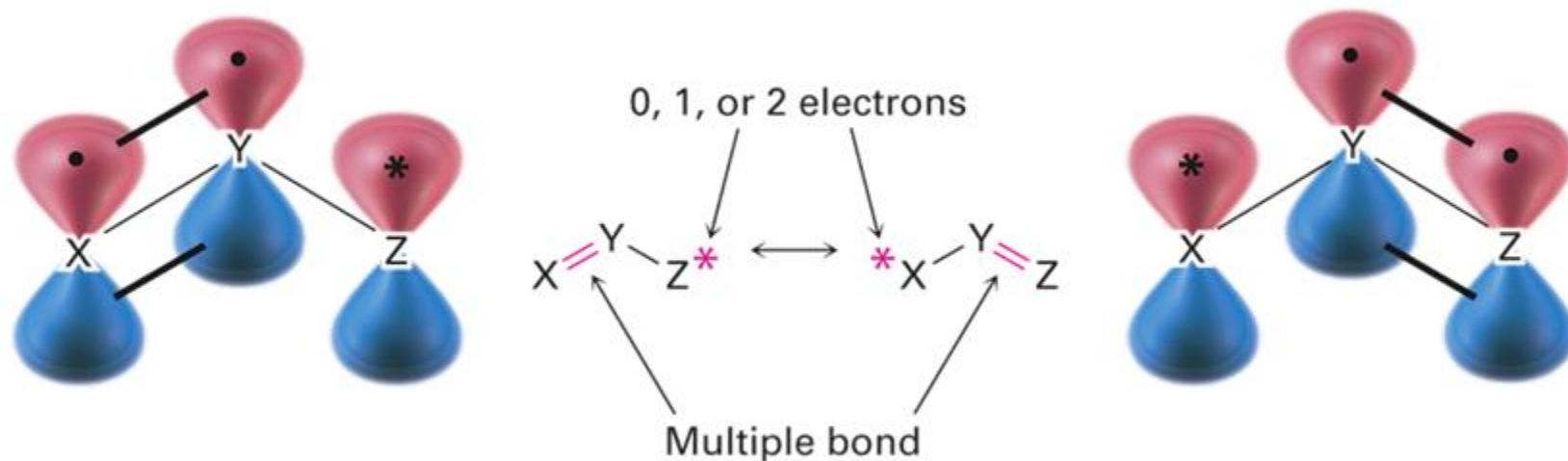
- 5) The resonance hybrid is more stable than any individual resonance form

Drawing Resonance Forms

- **Generally, any three-atom grouping with multiple bonds or with a p orbital on each atom has two resonance forms**
- **Sometimes resonance forms involve different atom types as well as locations**
- **We can imagine that electrons move in pairs to convert from one resonance form to another**

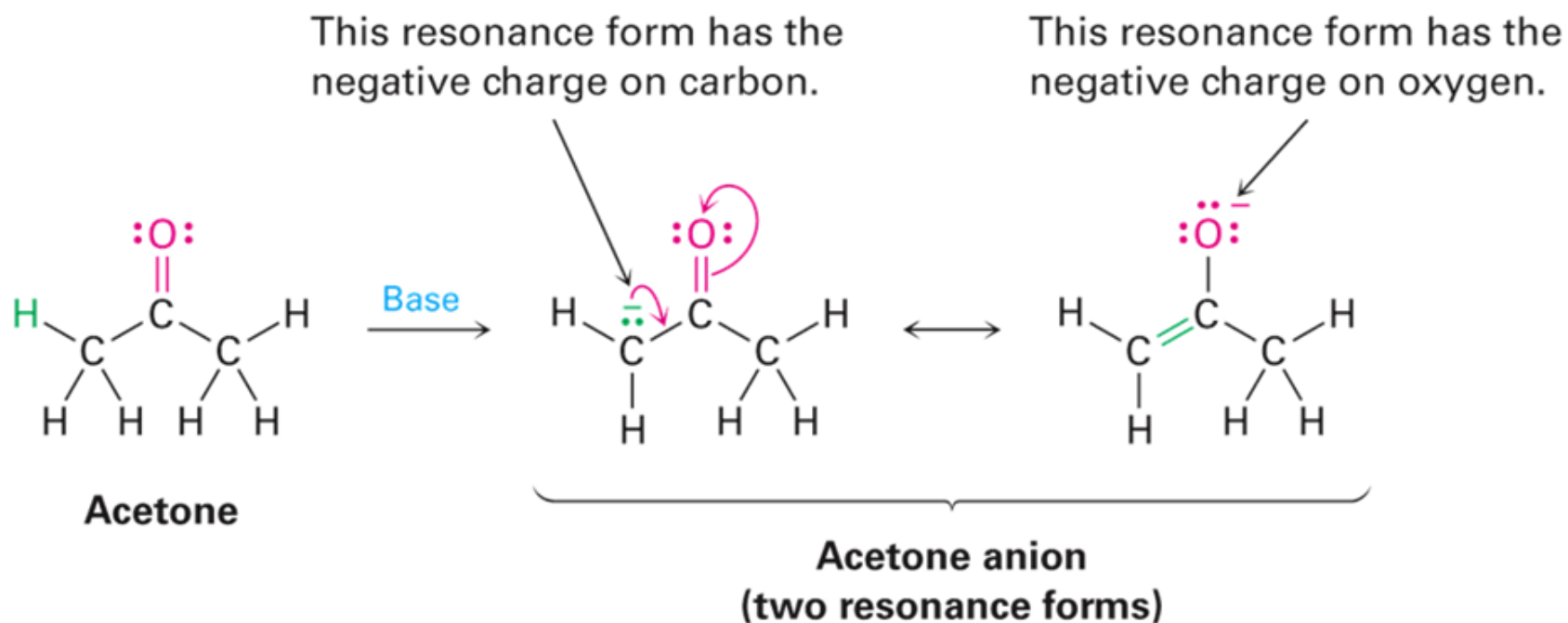
Drawing Resonance Forms

- In the illustration below, the asterisk (*) might mean that the *p* orbital on atom Z is vacant, or that it contains a single electron, or that it contains a lone pair of electrons.



Drawing Resonance Forms

- Here in the case of acetone, the “enolate” derived is a good illustration, with delocalization between carbon and oxygen



In the case of 2,4-Pentanedione

- The anion derived from 2,4-pentanedione
 - Lone pair of electrons and a formal negative charge on the central carbon atom, next to a C=O bond on the left and on the right
 - Three resonance structures result

