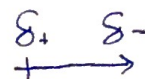


Electron Sharing: Polarity, Resonance, Acids and Bases

Spencer
Brucks
2017

Today, we will cover several fundamental concepts in organic chemistry that underlie chemical reactivity.

Polar covalent bonds: Arise between atoms of different electronegativity with e^- polarized towards high electronegativity



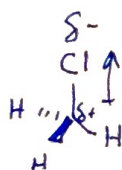
Resonance: A way of describing e^- where bonding cannot be written with a single Lewis structure

Acids: Bronsted-Lowry proton (H^+) donor Lewis e^- pair acceptor

Bases: H^+ acceptor e^- pair donor



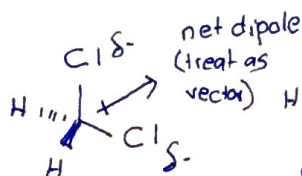
non-polar



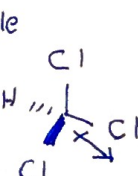
* Not a formal charge!
only "slightly negative"
and "slightly positive"

dipole moment $\rightarrow \mu = 1.9 \text{ D}$

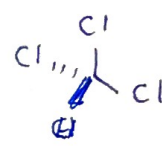
\uparrow
Debye



$\mu = 1.6 \text{ D}$



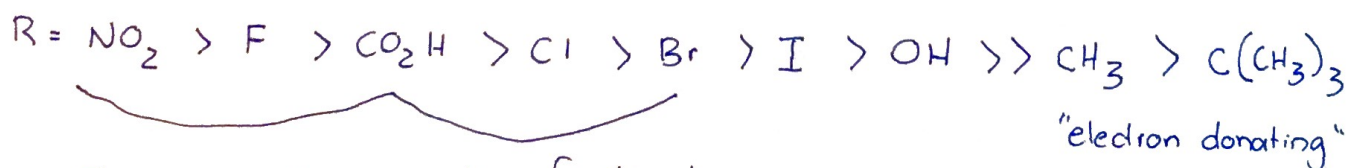
$\mu = 1.08 \text{ D}$



non-polar

All dipoles cancel.
They pull the same
magnitude, away from each other

Most organic functional groups are more electronegative than Carbon



The most electronegative functional groups are referred to as "electron withdrawing groups"

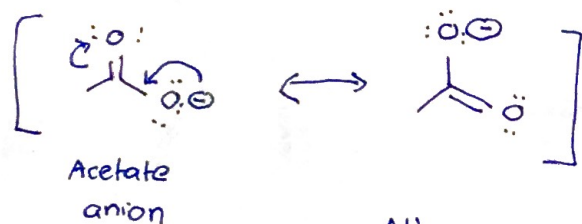
Resonance: A way of describing delocalized electrons where

Add to board bonding cannot be written with a single Lewis structure.

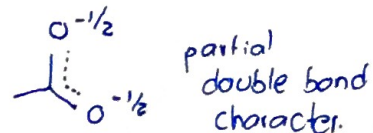
We use resonance structures to accurately represent electron density and orbital interactions in a single molecule.

- RULES:
- Electrons can move (lone pair and π)
 - Atoms cannot move (can't break σ bonds!)
 - This is a resonance arrow \longleftrightarrow v.s. \rightleftharpoons means equilibrium
 - ★ Resonance is NOT interconverting species; all the same molecule!

ex:



we can draw a resonance hybrid:

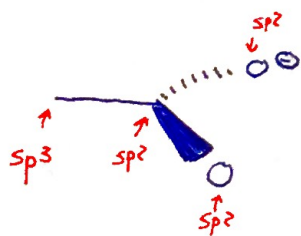


All resonance structures must obey valency rules.

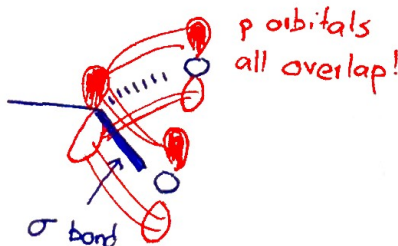


Texas Carbon!

Orbital picture:
(sideview)

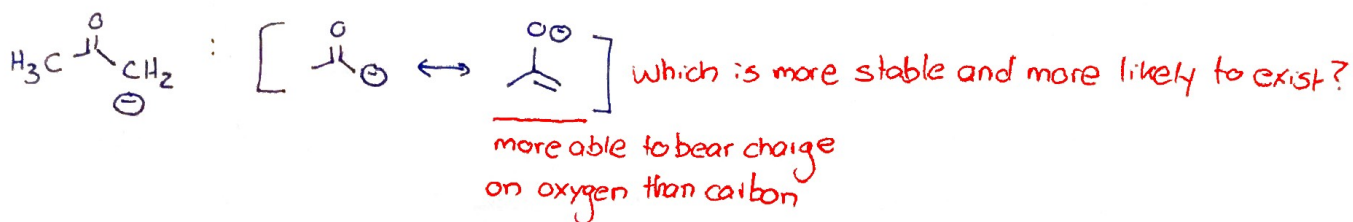


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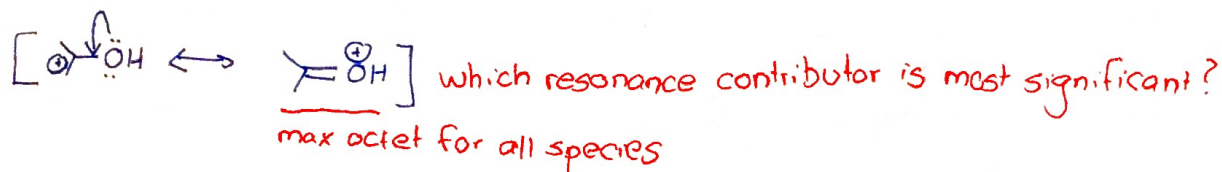


we have 4 e^- 's shared in 3p orbitals.

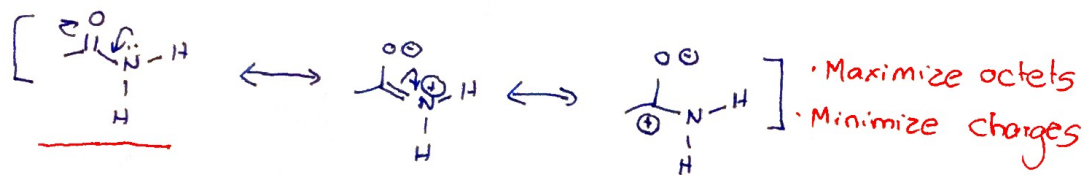
More Resonance Structures



Positively charged:



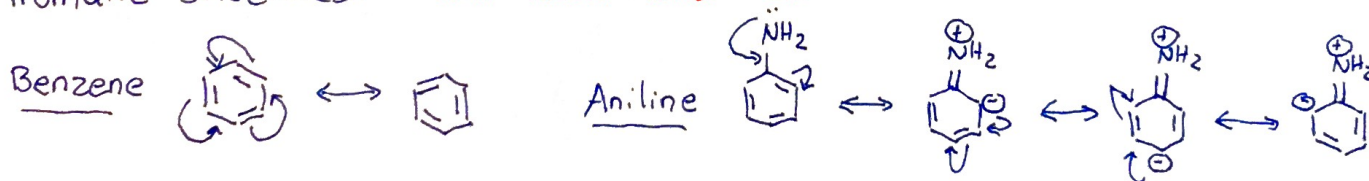
Neutral:



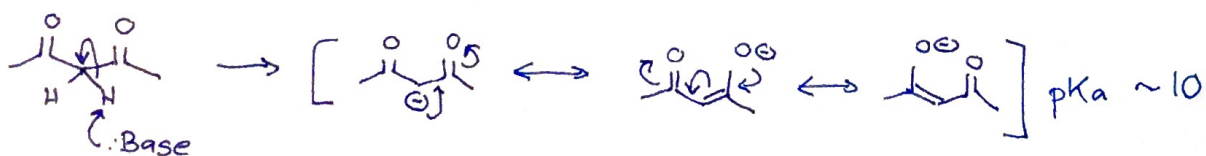
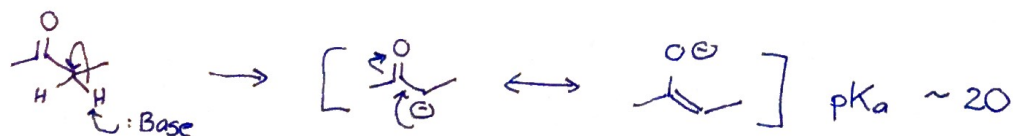
Radical:



Aromatic structures: will learn more about these later!



Resonance structures help us understand and predict reactivity



Reminder: Lower pK_a is easier to remove H^+

★ The second reaction is easier to achieve because the anion is more easily delocalized = more resonance structures

Introduction to Organic Reactions: Acids and Bases

Some definitions first:

Brønsted-Lowry Acid: proton (H^+) donor

Brønsted-Lowry Base: proton acceptor

Lewis Acid: electron pair acceptor

Lewis Base: electron pair donor

↓
Sour
taste

↓
bitter
taste

↓
Add to
board

Strength of Acids

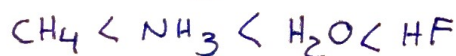
Reminder: Lower pK_A = stronger acid

① Acid strength \uparrow as size \uparrow

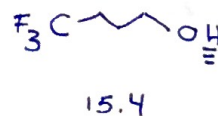
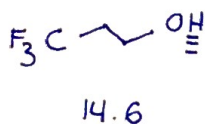
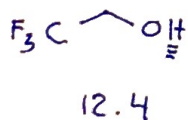
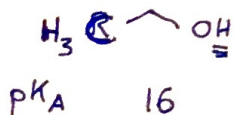


② If atoms are similar in size the strongest acid will have its hydrogen attached to the most electronegative atom.

\therefore Acid strength \uparrow with electronegativity \uparrow



③ Inductive effect: Nearby electronegative halogens increase strength of acid



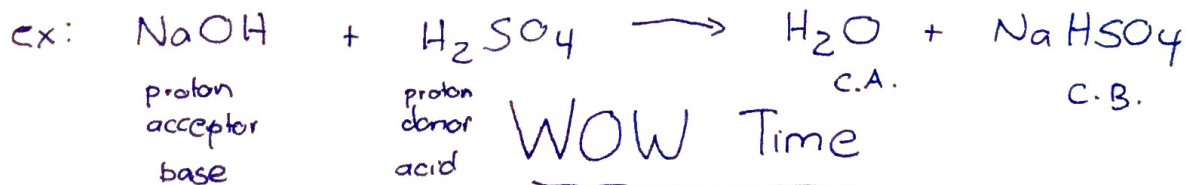
Effect falls off through bond interactions

= weakest

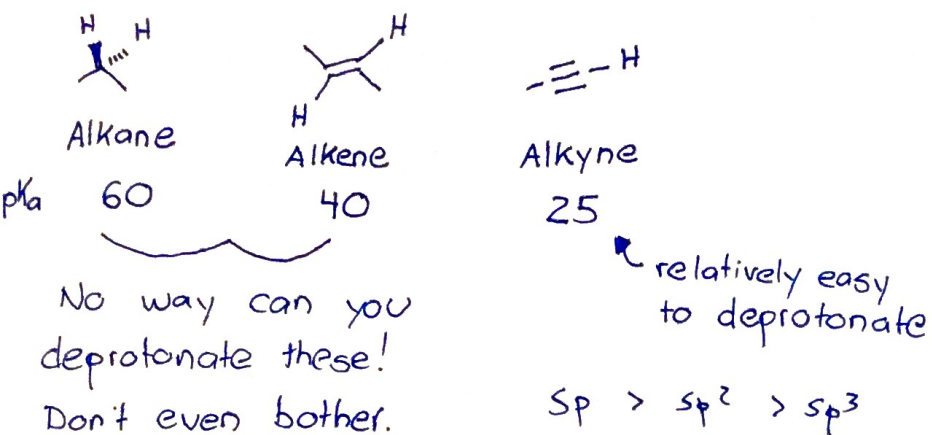
★ The strongest acids have the most stable conjugate bases ★

Reactions must "run downhill" such that stronger acid/base \rightarrow weaker acid/base

We have gathered information about structure in order to understand reactivity.



Last thing:



Why?

$sp > sp^2 > sp^3$
 most electronegative least electronegative

Alkyne protons are the most acidic because alkynes are sp hybridized with sp hybrid orbitals. These orbitals have the most s character (50% s and 50% p).

Therefore they are the closest to the positively charged nucleus - this best stabilizes the conjugate base.

$\text{--}\equiv\ominus$ most stable conjugate base!

Terms to Know

Nucleophile: "nucleus - loving"
electron rich

Reminder: nucleus is positively charged

Electrophile: "electron - loving"
electron - deficient