$HA + B \rightarrow A^{\ominus} + \overset{\oplus}{B}H$

↑ Proton acceptor → lone pair, π-bond, σ-bond

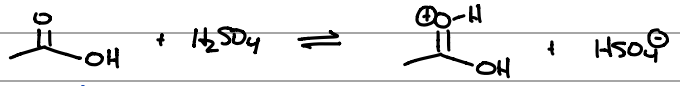
Brønsted Acid = a hydron donor

↳ H^+, D^+, T^+

$^1H^+$ $^2H^+$ $^3H^+$

Stable A^- → Strong acid → Low pKa

Compounds may act as an acid or a base depending on environment.

Ex:  + H_2SO_4 ⇌ + $HSO_4^{\ominus}$

typically an acid — but here it acts as the base

Aqueous Solutions

$HA + H_2O \rightleftharpoons A^{\ominus} + \underbrace{H_3O^{\oplus}}$

lyonium ion (for H_2O → hydronium ion)

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

* For $K_a > 1$ $(pK_a < 0)$ HA is a stronger acid than H_3O^+

A^- is a weaker base than H_2O

A pKa unit corresponds to 10x change in acid strength

	H_2O	vs	^tBuOH
pKa	16		17

(10x stronger acid)

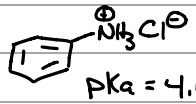
pKa = fundamental property of the acid in a particular environment

pH = acidity of the solution; related to pKa of acid and concentration of acid

$pH = -\log [H_3O^+]$

Unlike pKa, you can experimentally adjust pH.

Henderson-Hasselbach Eq: $\boxed{pH = pK_a + \log \frac{[A^-]}{[HA]}}$

Ex: 

pKa = 4.6 in H_2O $\xrightarrow[pH\ 3]{\text{Adjust to}}$ $3.0 = 4.6 + \log \frac{[A^-]}{[HA]}$ $\Rightarrow 0.025 = \frac{[A^-]}{[HA]}$

40:1

[HA] : [A⁻]

Adjust to pH 7

$251 = \frac{[A^-]}{[HA]}$

1:251

[HA] : [A⁻]

Rules: ① When pH = pKa [HA] + [A⁻] exist 1:1
② When pH > pKa acid exists mostly in A⁻ form
③ When pH < pKa acid exists mostly in HA form

The Leveling Effect

Consider: $HCl + H_2O \rightleftharpoons H_3O^+ + Cl^-$

pKa -6 pKa -1.74

With a strong acid, the equilibrium exists predominately as products

essentially the real acid in the soln (pKa -1.74)

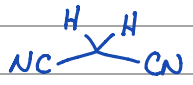
* The lowest pKa you can measure in a given solvent is the protonated solvent (H_3O^+ for H_2O).

* In H_2O, pKa's can be determined from -1.74 → 15.7

Non-Aqueous Systems

Solvation is an important factor for acidity

pKa of organic acids are often measured in DMSO → doesn't ion pair well $DMSO\text{-}H^{\oplus}$ $CB^{\ominus}$

Ex: HCN vs 

9.4 11 pKa in H_2O

12.9 11.1 pKa in DMSO

* pKa's are usually larger in org. solvents

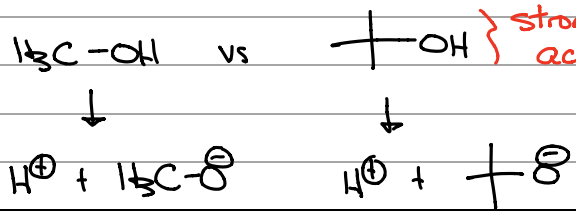
DMSO is less effective at stabilizing CN^-

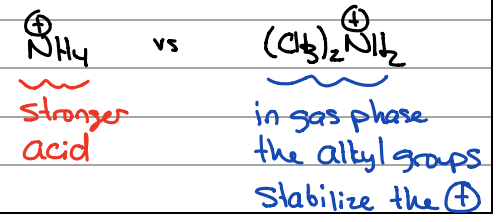
little charge since ⊖ is less localized & spread across the molecule

Gas Phase Acidity

- Gives insight into intrinsic anion stability
- No solvent effects

Often see a reversal of expected trends in the gas phase





As organic chemists, we are primarily concerned w/ solution phase acidity

Predicting Relative Acidity

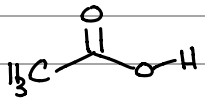
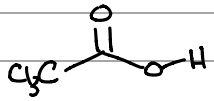
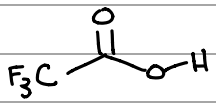
Choices: Assess relative stability of HA
↳ more stable HA = less acidic — Best for $HA^{\oplus}$

Assess relative stability of $A^{\ominus}$
↳ more stable $A^{\ominus}$ comes from the stronger acid — Best for neutral HA

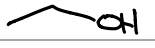
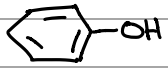
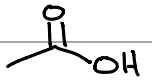
Electronegativity

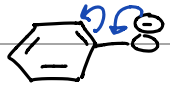
	pKa
H_3C-H	50
H_2N-H	38
$HO-H$	16
$F-H$	3.2

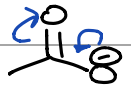
Inductive Effect

	pKa
	5
	0.65
	-0.2

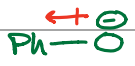
Resonance

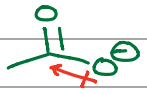
	pKa
	16
	10
	5


Benefits from delocalization of ⊖

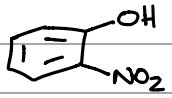

Benefits more when ⊖ can delocalize onto e⁻ neg atom

Also benefit from inductive effect

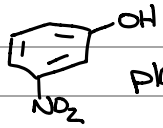




The power of inductive effect:


pKa 7.22

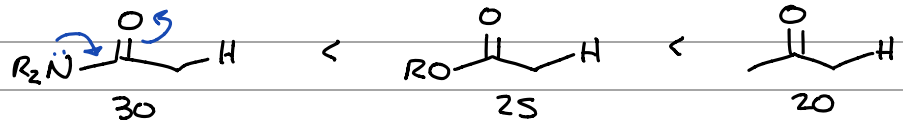
Resonance onto $-NO_2$ + Inductive w/ $-NO_2$


pKa 8.36

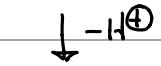
Inductive effect w/ $-NO_2$ only

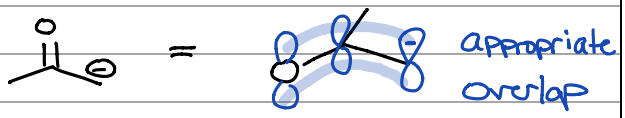
* Here resonance only lowers pKa by ~1 unit over the inductive effect.

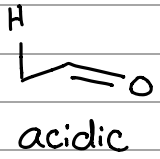
Enolates

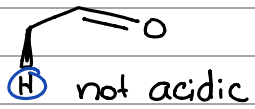


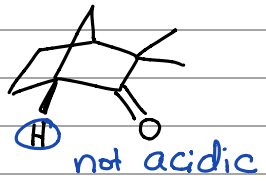
Resonance donation lowers α-H acidity





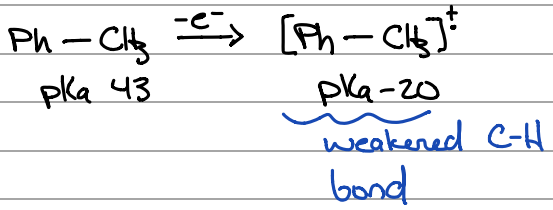




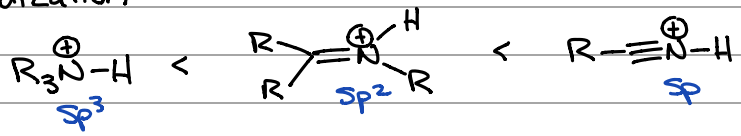


Bond Strength (i.e. atom size)

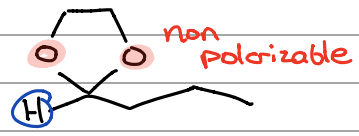
H_2O | decreasing bond
H_2O | Strength (b/c of atom size)
H_2Se | increasing acidity
H_2Te |

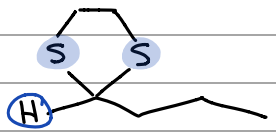


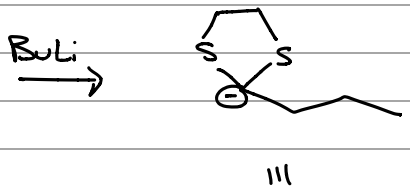
Hybridization

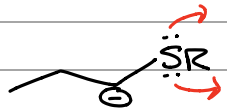


Polarizable Neighboring Atoms



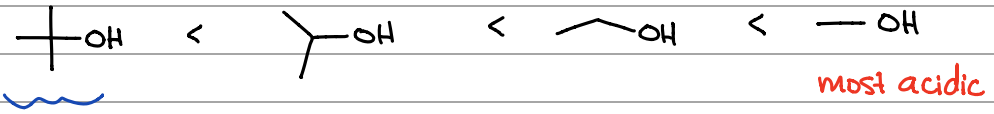






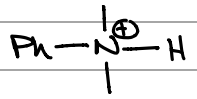
Polarizability stabilizes anions
lone pairs move out of the way

Solvation Effects



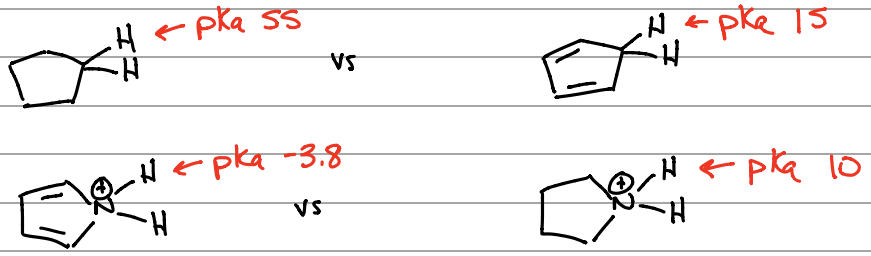
w/ the bulkier R-group it is more difficult for the solvent to make close contact to solvate and stabilize ⊖

Polar Solvents Stabilize Charged Species



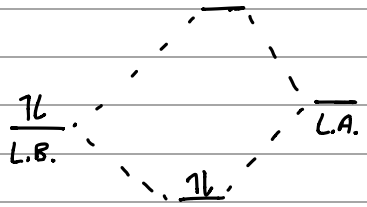
pKa = 5.2 in H_2O } Stabilizes the ⊕
pKa = 2.5 in DMSO

Aromaticity

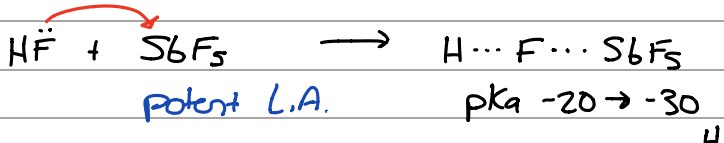


Lewis Acids and Bases

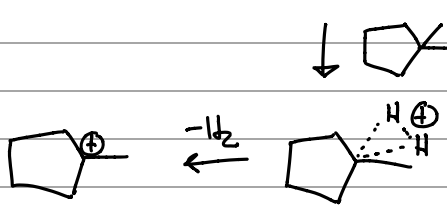
↳ e⁻ pair acceptor
↳ e⁻ pair donor

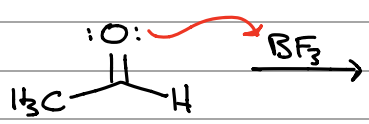


Lewis Acid = anything w/ an empty orbital



Olah
JOC 2001, 66, 5943

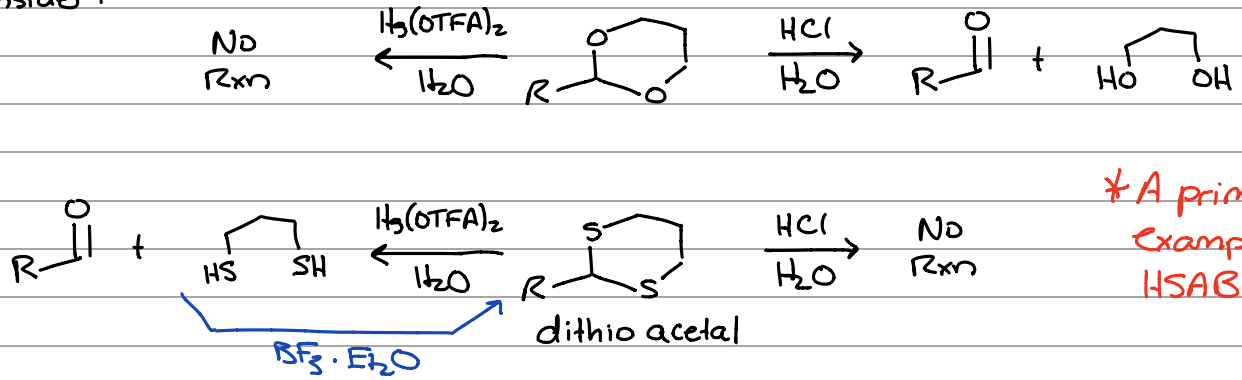




A highly electrophilic C=O

HSAB - Principle of Hard and Soft Acids and Bases

Consider:



- Hard Lewis Acids interact best with hard Lewis bases
- Soft Lewis Acids interact best with Soft Lewis bases

Related to orbital energy

$\frac{\text{H.L.}}{\text{L.B.}}$ ↓ Lower E = Harder

$\frac{\text{L.A}}{}$ ↑ Higher E = Harder

Salem-Klopman Eqn: $\Delta E = E_{\text{Steric}} + E_{\text{electrostatics}} + E_{\text{orbital overlap}}$

"energy change for chemical rxn"

Hard

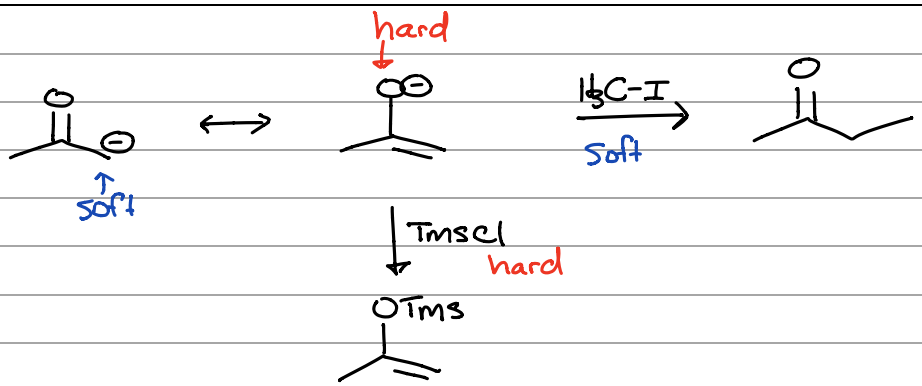
- non-polarizable
- Primarily electrostatic attractions
- Acids: H^+, Li^+, BF_3, $AlCl_3$
- Bases: H_2O, RO^-, ROR, NH_3

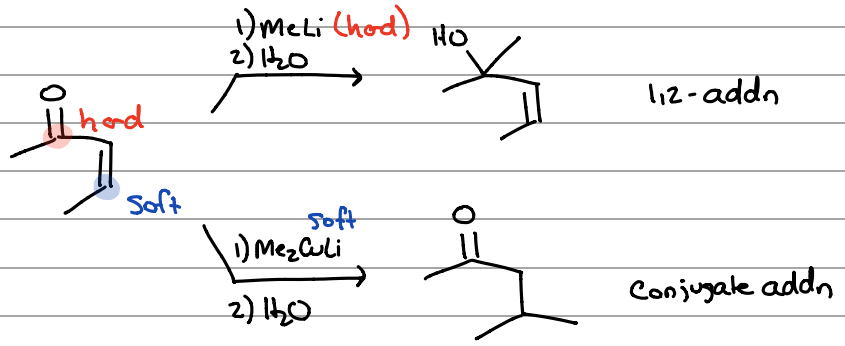
Soft

- polarizable
- Primarily orbital mixing
- Bases: RS^-, R_2S, I^-, PPh_3, CN^-, R^-
- Acids: I_2, Cu^+, Hg^+

Ex: $Ph_3P:$ + H_3C-I → $Ph_3\overset{\oplus}{P}-CH_3\ I^{\ominus}$

soft base — soft acid


hard
soft
H3C-I
soft
TmsCl
hard
OTms


1) MeLi (hard)
2) H2O
HO
1,2-addn
hard
soft
soft
1) Me2CuLi
2) H2O
Conjugate addn