

Acids and Bases

Overview of Acids and Bases

Indicators: Acids and Bases

Acids: Donates its proton to base (proton donor)

Base: Donates its lone pair to acid (LBP)



Remember: (1) only H atoms, (2) only OH is a base

What about organic indicators and organic bases/acids?

Organic Acids and Bases

Acids are proton donors, Bases are proton acceptors.



1) Look at the structure of acid/base molecules

to find the atom that has the lone pair/negative charge of the base

and check for any other lone pairs of the acid.

2) Look for the lone pair of the hydrogen and move the electrons

to the other atom

Remember: correct arrows show movement of electrons

3) Acids: Electron and Proton Donations are from Proton Properties

What if there is no proton transfer?

Organic Acids and Bases

Acids: acid is an electron-proton acceptor

Acids: base is an electron-proton donor



So, for the 3 acids/bases, acids and bases are described in terms of electron

3) None of the acid-base theories completely fit our needs

As a general practice, consider acids as (1) and bases as (2)

Acid Strength

What does it mean "strong acid"?

Informally speaking, for acids the acid dissociates into a proton, the stronger it is



The dissociation of the acid is a reversible reaction that is described by equilibrium constant



Notice that: $K_a = K_{eq} [\text{H}_2\text{O}]$

Acid Strength and pK_a

The K_a values of weak acids (1) and weak bases (2) are shown below



The acidities of a stronger and weaker in the 1st step

To quantify the acidities, pK_a is used, which is the pK_a of the given acid

$$pK_a = -\log K_a$$

$$pK_{a1} = -\log K_{a1} = 10 \quad pK_{a2} = -\log K_{a2} = 15$$

So, the stronger the acid, the lower the pK_a value (It's the opposite the K_a)

Below are pK_a values of some common functional groups



What makes the Acid Strong?

The strength of the acid is dictated by the stability of the conjugate base (the lower the acid dissociates)

The stronger the acid, the more stable its conjugate base



Stability is more about acid dissociation, does provide their strength and

There are 3 factors to be considered when assessing the stability of the conjugate base

1) The first and the last factor is the **charge** to which the hydrogen is connected. Usually

to **anionic** than **Phenol/Alcohols** and **Thiophenols**

2) Order the most stable anionic atom that charge ends up electronegative is the best



Stays in more electronegative, that is why the alcohol is more acidic

$$pK_{a1} = 10 \quad pK_{a2} = 15$$

3) When comparing elements in different rows, the atomic size is factor

Larger atoms have higher polarizability and results for negative charge better in the charge is spread over



There are more acidic atom charge negative is more electronegative than sulfur

The sulfur atom is larger and stabilizes the negative charge better

3) When comparing for same atom, check which conjugate base is **Resonance stabilized**



1) The first factor for determining the acidity is called **Inductive**

Inductive or inductive effect is the transmission of charge through atoms

Electronegative atoms pull the negative charge and helps to stabilize the conjugate base



The more electronegative, the more electron withdrawing, the stronger the acid



The closer the electronegative atom to the hydrogen, the stronger the acid



2) The second factor is the **Resonance** of the atom connected to the hydrogen

The explains the great difference of pK_a values of alcohols, aldehydes, and ketones



3) which is more electronegative than O , so the more O character, the stronger the acid

Therefore, for **acids**, **aldehydes** > **ketones** > **alcohols**

The primary acids: **Alcohols**, **Aldehydes**, **Ketones**, defined the decreasing the acidity

of a compound is shown as **acidity**

Remember: also consider the conjugate base



The Position of Equilibrium for an Acid-Base Reaction

Is the following reaction reaction possible?



Acid-base reactions are reversible, so another way of asking the question is

Is the equilibrium of the reaction shifted to the right (products) or to the left (reactants)?

To answer this question, you need to remember that the equilibrium of acids and bases reaction is shifted towards the weaker acid and base

So, you need to compare the pK_a of the acid left with the conjugate acid right with

The equilibrium will be shifted to the higher pK_a reaction with only



Is the acid the pK_a of the conjugate acid is greater higher the acid pK_a , so the

equilibrium is shifted to the right. Therefore the right pointing arrow is shown larger

Choosing a Base to Deprotonate a Given Compound

Suppose you need to deprotonate the following compound: HCl

(Deprotonate means remove the acidic proton)

First, find its pK_a value: $\text{HCl} \rightarrow \text{H}^+ + \text{Cl}^-$

Find the chemical equation of the conjugate reaction with a conjugate base (2)



The pK_a of the conjugate acid must be higher than 7, so weaker acid is formed

Find a conjugate base with pK_a value > 7. Acids are weak when $pK_a > 10$



Choosing an Acid to Protonate a Given Compound

Suppose you need a conjugate to protonate compound **alcohol** (1) that has a pK_a value of 15 (very reactive base and needs to be quenched before being disposed)

To do that, write the chemical equation of alcohol with a weak acid



The pK_a of the acid (the conjugate acid of alcohol) is 7. Therefore, the pK_a of **acid** and must be smaller than 15, stronger acid reacts, weaker acid forms

Therefore a strong acid is a proton donor (acid) for protonating the alcohol because its pK_a is 7 and is more stronger of an acid than bases



Hammett-Hansch Equation

This equation shows whether the compound with a given pK_a is an acid or base (deprotonated form is an organic solution depending on the pH of the solution)

$$pK_a = pH + \log \frac{[A]}{[A^-]}$$

Let's graph and implications of the Hammett-Hansch equation

1) If the $pH < pK_a$ 2) If the $pH = pK_a$ 3) If the $pH > pK_a$

$$\log \frac{[A]}{[A^-]} = 1 \quad \log \frac{[A]}{[A^-]} = 0 \quad \log \frac{[A]}{[A^-]} = -1$$

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T. R. Dickson



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