

# Acids and Bases

## Overview of Acids and Bases

### Acid-base Acids and Bases

Acids: Donates its proton in these processes (H<sup>+</sup>)  
 Base: Donates its lone pair to form a bond (H<sup>+</sup>)



Remember: It's only in water, it only HCl is a free

What does aqueous hydrochloric acid and sodium hydroxide solution?

### Acid-base Acids and Bases

Acids are proton donors (the molecule is losing its proton)  
 Bases are proton acceptors (the molecule is taking the lone pair)



### General Acids and Bases

It shows the proton donor, the lone pair acceptor, the conjugate acid, and the conjugate base.  
 It shows the lone pair of the hydroxide and where the electron is in the other base.

Remember: **acid accepts lone pair** (proton donor)

Both Acids and Bases are defined in terms of Proton Properties. What if there is no proton involved?

### Acids and Bases

Acids act as proton donors.  
 Bases act as proton acceptors.



Acid-base Acids and Bases are defined in terms of electron.

Acid: A general proton donor (acid) and base (base)

## Acid Strength

### What does it mean "strong acid"?

Generally speaking, the more 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.



## Acid Strength and pKa

The Ka value of acid dissociation (Ka) and acid-base (Kb) are shown below.



It's related to a stronger acid because in the K<sub>a</sub> equation.  
 To simplify the equation,  $\text{pH} = -\log[\text{H}^+]$  is the pKa of the proton acid.

$\text{pH} = -\log[\text{H}^+]$   
 $\text{pKa} = -\log[\text{Ka}]$   
 $\text{pH} = \text{pKa} + \log\left(\frac{[\text{A}^-]}{[\text{HA}]}\right)$

So, the stronger the acid, the lower the pKa value. (It's the opposite for Kb)

Below are pKa values of some common functional groups.



## What makes the Acid Strong?

The strength of the acid is defined by the stability of the conjugate base. The more stable the conjugate base, the stronger the acid.



Water is a weak acid and hydroxide ions are a strong base. Water is in the middle of its ionization state, meaning the stability of the conjugate base.

- 1) The first and the best factor is the **charge** to which the hydrogens is connected. Strongly **electron withdrawing** and **electronegative**.
- 2) When the same **charge** atom, then the charge is **electron withdrawing** in the base.



When comparing elements in different rows, the more electronegative, the more stable the conjugate base, the stronger the acid.



Electron withdrawing groups (EWG) are more electronegative than carbon. The carbon atom is larger and stabilizes the negative charge better.

When comparing for same atom, think which conjugate base is **more electronegative**.



1) The best 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 help to stabilize the conjugate base.



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



The closer the electronegative atom to the hydrogens, the stronger the acid.

## Resonance Acids

2) The best factor is the **charge** of the atom connected to the hydrogens. The opposite the proton difference of pKa values of alcohols, aldehydes, and ketones.



3) When the same electronegative atom, or the same element, the stronger the acid. Therefore, for **acids**, **alcohols** < **aldehydes** < **ketones**.

The protonic acids, **Alcohols**, **Aldehydes**, **Ketones**, defined the increasing the acidity of a compound is shown in **pKa**.

## Remember! the stability of the conjugate base



## The Position of Equilibrium for an Acid-Base Reaction

1) In the following reactions, choose the stronger acid.



Acid-base reactions are reversible, so another way of asking the question is: In the equilibrium of the reaction, which is the right product or is the left reactant? 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 pKa of the acid with acid and to compare with acid with acid. The equilibrium will be shifted to the higher pKa reaction with acid.



In the case the pKa of the conjugate acid is **greater** higher than the acid pKa, or the equilibrium is shifted to the right. Therefore, the right product side is shown higher.

## Choosing a Base to Deprotonate a Given Compound

Suppose you need to deprotonate the following compound. —OH  
 (Deprotonate means remove the proton acidic proton)

First, find its pKa value. —OH → pKa = 17

Then, the chemical equation of the conjugate reaction with a conjugate base (B<sup>-</sup>).



The pKa of the conjugate acid must be higher than 17, so consider acid is formed.

Find a conjugate base with pKa value with pKa > 17. Answer are used when pKa > 17.



## Choosing an Acid to Protonate a Given Compound

Suppose you need to protonate a given compound. —OH (Alcohol that has a hydroxyl group, which is a very reactive base and needs to be protonated before being displaced)

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



The pKa of the conjugate acid of alcohol is 17. Therefore, the pKa of acid must be lower than 17 — stronger acid means, weaker acid base.

Therefore, a strong acid is a proton donor (acid) for protonating the alcohol because its pKa is 17 and it is a weaker strength of an acid than base.



## Remember: Hammett's Equation

The equation shows whether the compound with a given pKa is in the acidic or basic (deprotonated) form in an aqueous solution depending on the pH of that solution.

$\text{pH} = \text{pKa} + \log\left(\frac{[\text{A}^-]}{[\text{HA}]}\right)$

Two graphs and implications of the Hammett-Henderson equation.

1) If the pH < pKa	2) If the pH = pKa	3) If the pH > pKa
$\text{pH} = 7$ $\text{pKa} = 10$ $\Delta \log\left(\frac{[\text{A}^-]}{[\text{HA}]}\right) = -3$	$\text{pH} = 10$ $\text{pKa} = 10$ $\Delta \log\left(\frac{[\text{A}^-]}{[\text{HA}]}\right) = 0$	$\text{pH} = 13$ $\text{pKa} = 10$ $\Delta \log\left(\frac{[\text{A}^-]}{[\text{HA}]}\right) = 3$
Impregnated with mostly in its acidic form	Equal parts	Impregnated with mostly in its basic form



The pKa of typical carboxylic acids is 4-5. Therefore, it is usually in the acidic form because its pKa is around 4-5, which is less than the pKa of the acid (pKa = 7). So, they will be in the acidic form. Here.

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**SB Merriam**

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