

a substance that produces  $H^+$  when dissolved in water. it is a proton donor and an electron pair acceptor.

# Acids.

Sour

acid pH = 0-7

Vinegar

aq. solutions conduct electricity.

Strong = 0-4 pH  
Weak = 3-6 pH

Stomach acid

Corrosive to metals.

changes litmus from red to blue.

Proton donors.

lemons

acid rain

splits off ions.



HCl

acid!

Soda

a solution that has an excess of  $H^+$  ions.

substances that dissolve in water to release hydroxide ions into a solution.

# Bases.

bitter.

base pH = 7-14

slippery.

Strong = 10-14 pH  
Weak = 8-10 pH

Proton receivers.

baking soda

antacid

NaOH

Takes ions.



base!

$Ca(OH)_2$  don't change the color of litmus.

a solution that has an excess of  $OH^-$  ions.

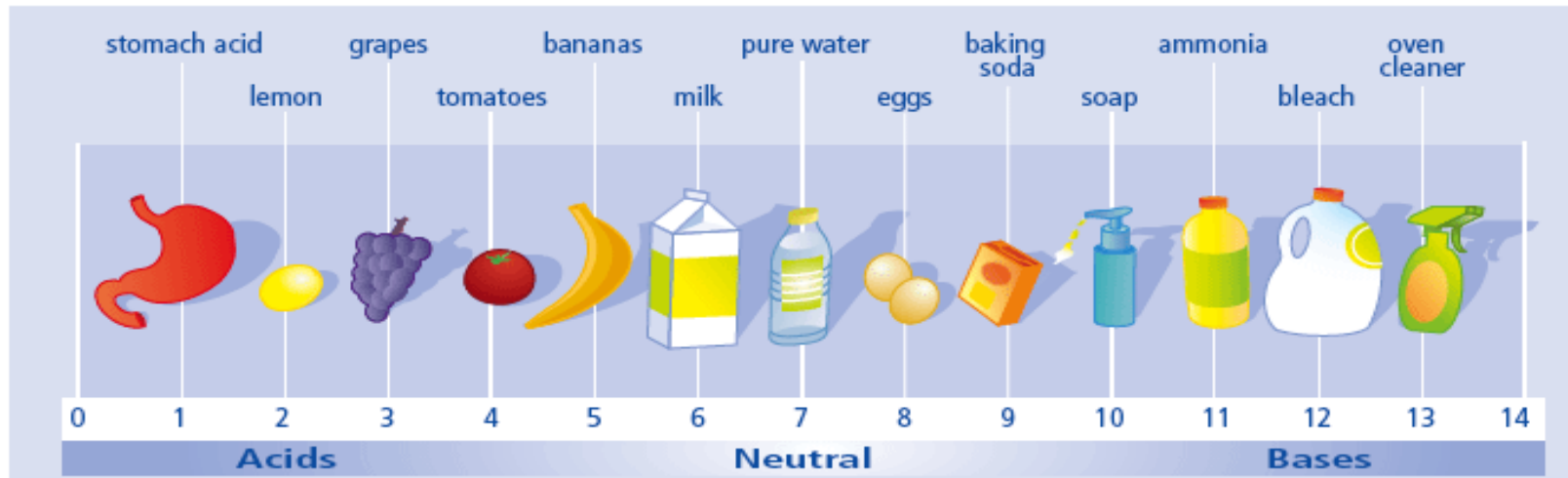
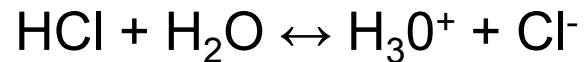
ammonia



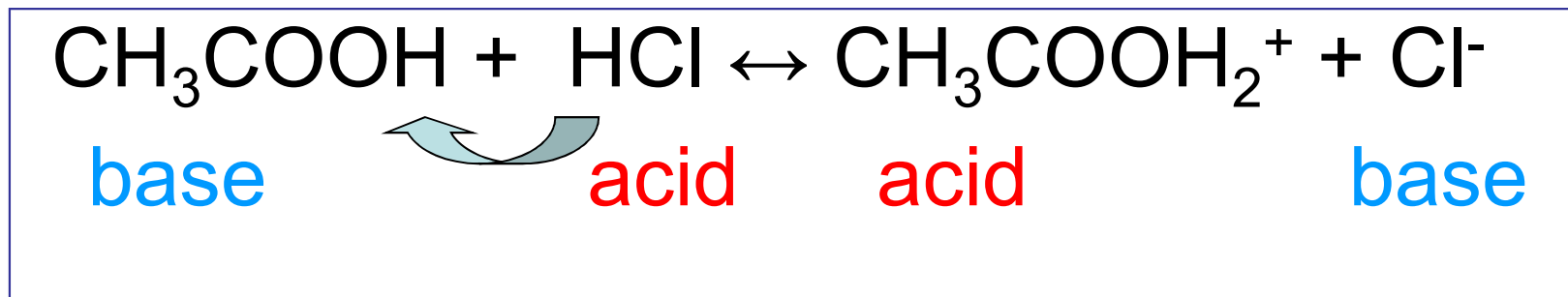
-Divina'09-

# Acids and bases

- Definition of Brønsted and Lowry is: **Acid** is a substance that gives the proton, ie the **proton-donor**, and the **base** receives a proton, or **proton-acceptor**.
- So base is formed by seceding protons from the acid:

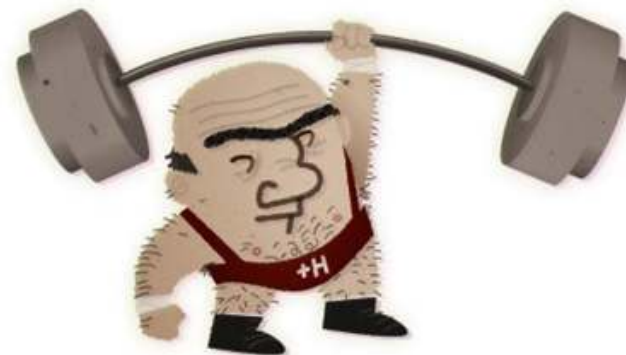
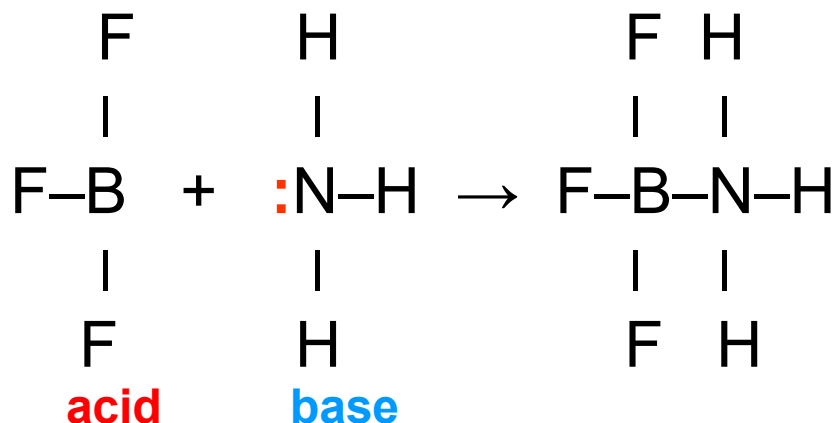


- Each acid has its own conjugated base and vice versa.
- What is an acid stronger, its conjugate base is weaker and vice versa.
- The concept of acids and bases is relative, for example:



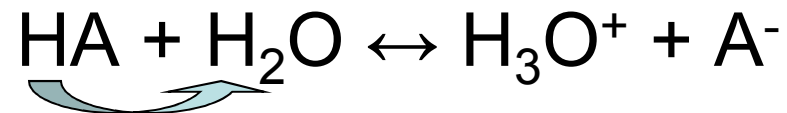
# Strong and weak acid and bases

- **strong acid and base** (strong electrolyte) completely dissociated
- **weak acid and base** (weak electrolyte) partially dissociated
- **Lewis:** The base is a substance which has a free electron pair. Acid is, due to the lack of electrons, bound to the free electron pair of the base.
- Acid does not have to have a proton

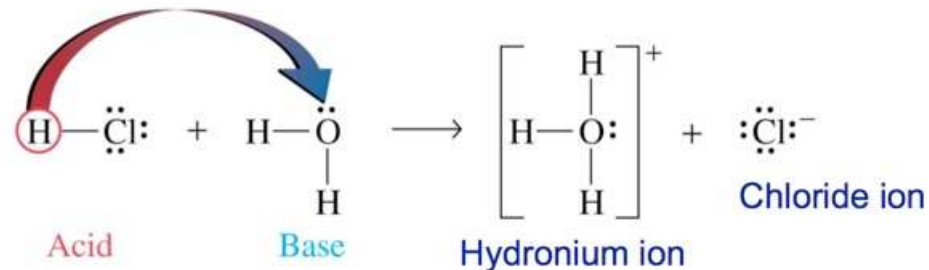
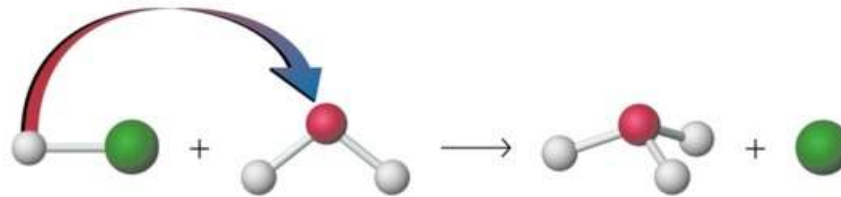


# Dissociation in water

- Because the weak acids and bases are partially dissociated in water, the dynamic equilibrium is established:



HA....weak acid

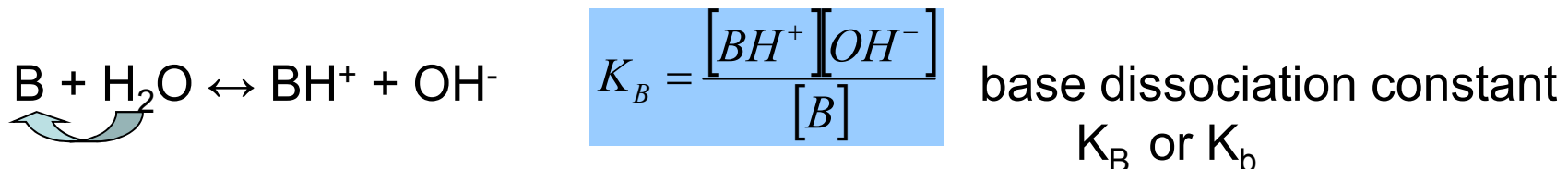
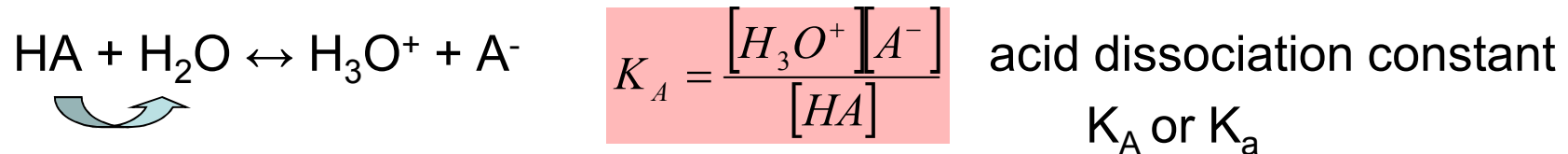


# The equilibrium constant

- For each equilibrium reaction the expression for the equilibrium constant applies:

$$K_c = \frac{[H_3O^+][A^-]}{[HA][H_2O]}$$
 equilibrium constant  $K_c$

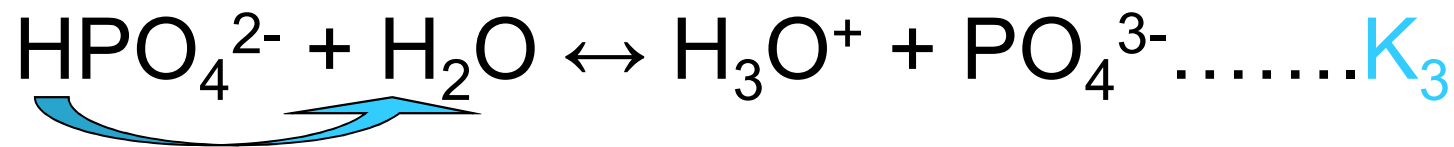
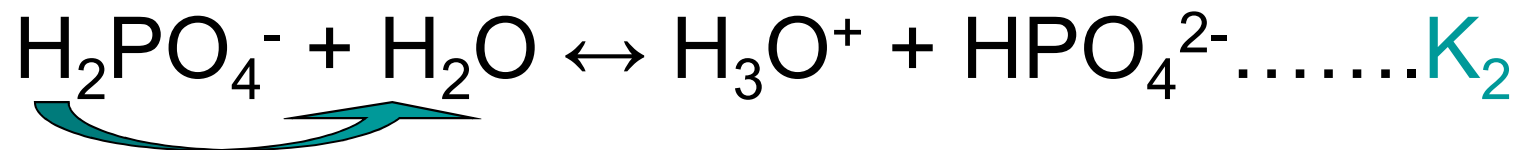
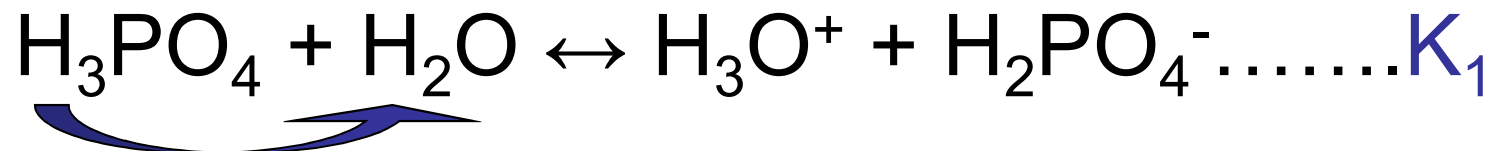
$[H_2O]$ ...constant



$$K_a(\text{CH}_3\text{COOH}) = 1,74 \cdot 10^{-5} \text{ mol dm}^{-3}$$
$$K_b(\text{NH}_3) = 1,74 \cdot 10^{-5} \text{ mol dm}^{-3}$$

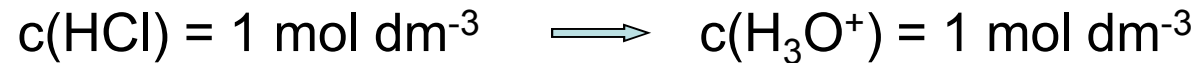
# Polyprotic acids

- dissociate in several stages:

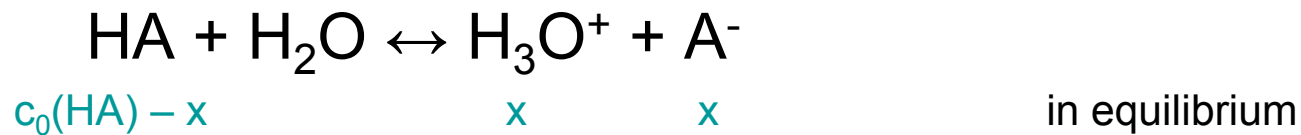


# H<sub>3</sub>O<sup>+</sup> i OH<sup>-</sup>

- Concentration of H<sub>3</sub>O<sup>+</sup> in strong acids:



- In weak acids:



- dissociated part of the acid is very small and is neglected

$$K_A = \frac{[\text{H}_3\text{O}^+][\text{A}^-]}{[\text{HA}]} = \frac{x^2}{c_0(A)} \implies [\text{H}_3\text{O}^+] = \sqrt{K_A \cdot c_0(A)}$$

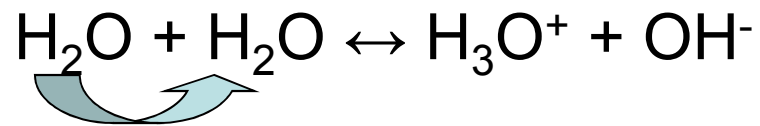
- analogously in weak bases:



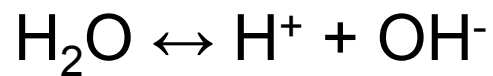


# Ionic product of water

- Water is very poor dissociated :



or



$$K_W = [\text{H}^+][\text{OH}^-] = 10^{-14} \text{ mol}^2 \text{ dm}^{-6}$$

ionic product of water



# pH

- in each aqueous solution ionic product of water applies
- Danish chemist **Sørensen** (1909.) Introduced the concept of pH :

$$\text{pH} = -\log[\text{H}^+]$$

$$\text{pOH} = -\log[\text{OH}^-]$$

$$[\text{H}^+] = [\text{OH}^-] = 10^{-7} \text{ mol dm}^{-3}$$

$$[\text{H}^+] > 10^{-7} \text{ mol dm}^{-3}$$

$$[\text{H}^+] < 10^{-7} \text{ mol dm}^{-3}$$

$$\text{pH} + \text{pOH} = 14$$

pH = pOH = 7  
neutral solution

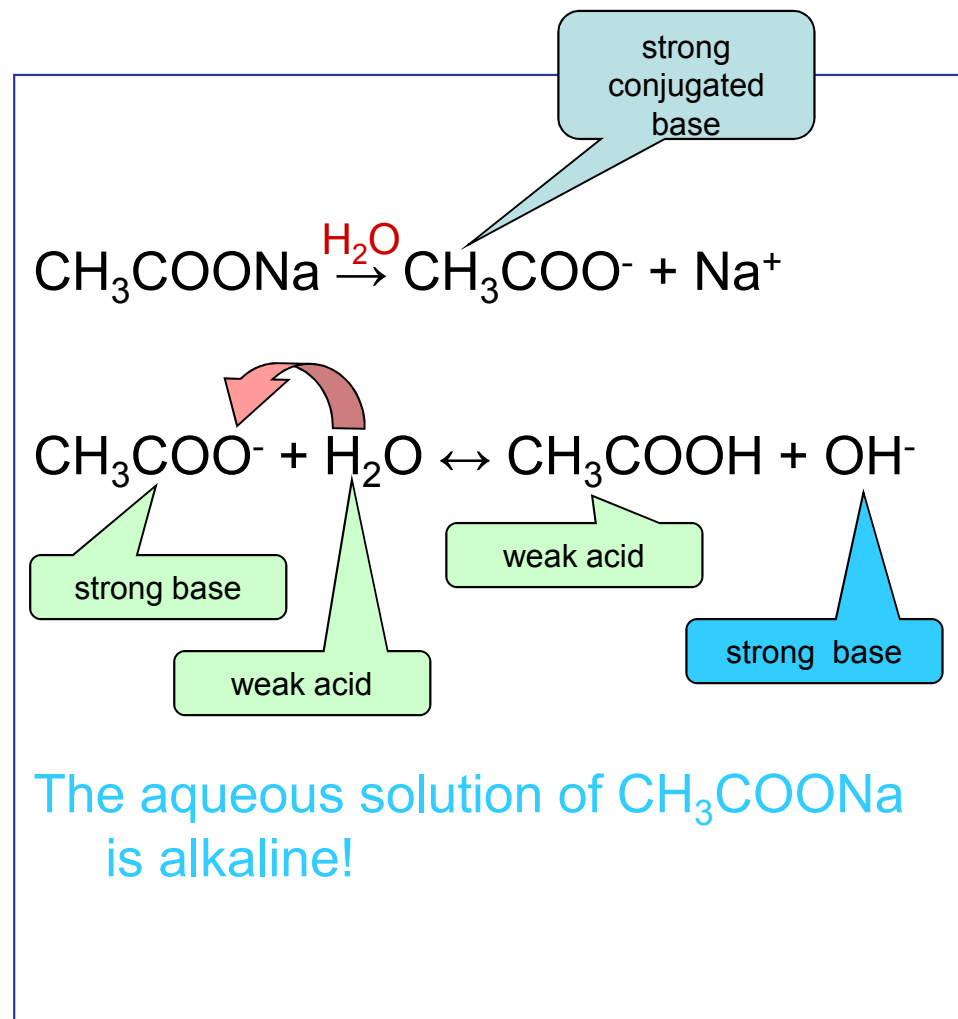
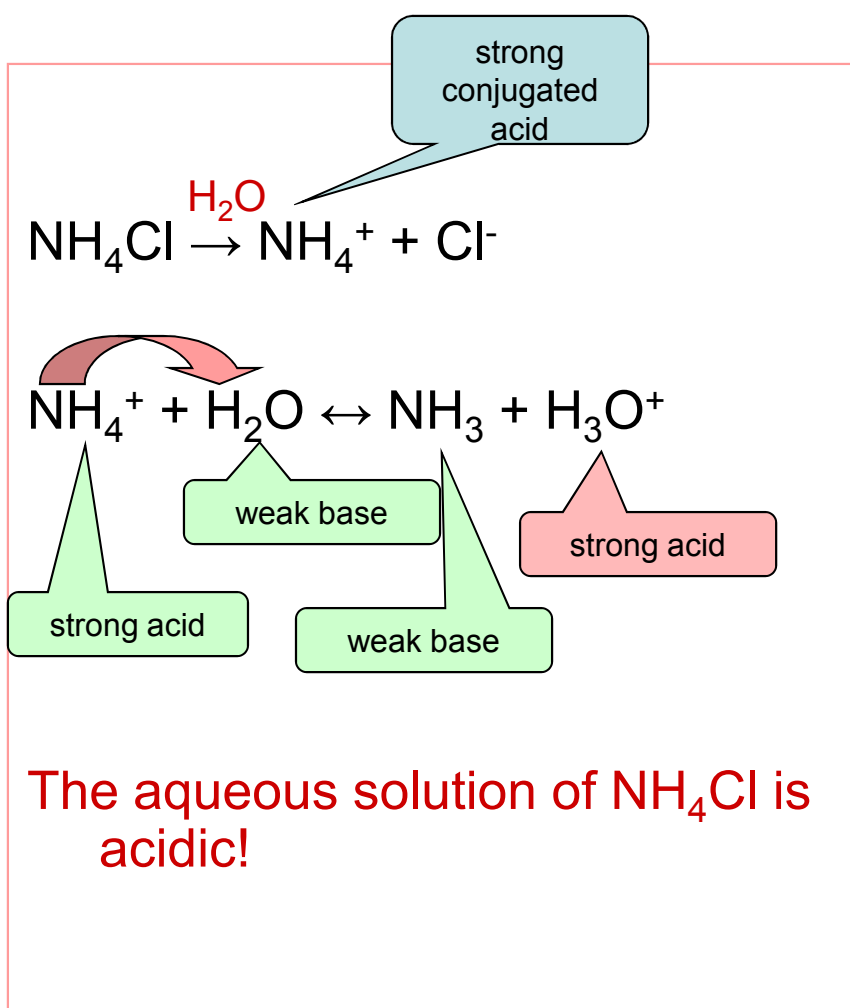
pH < 7  
acid solution

pH > 7  
alkaline solution



- Analogous to pH there has been introduced:  
 $\log K_w = \text{p}K_w = 14$   
 $-\log K_a = \text{p}K_a$   
 $-\log K_b = \text{p}K_b$

# Hydrolysis

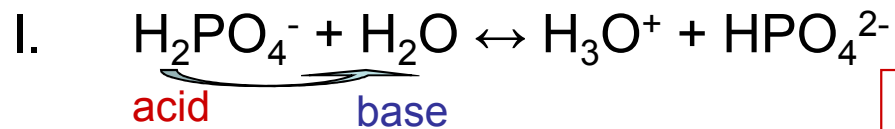
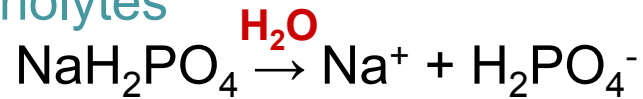


# Hydrolysis

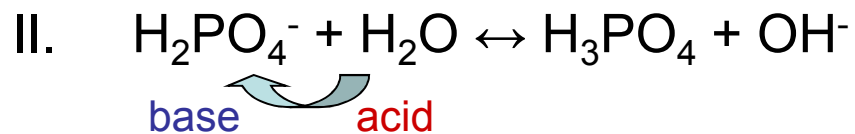
- The chemical reaction of salt ions and water; the equilibrium reaction
- In these solutions  $K_W$ ,  $K_A$  and  $K_B$  must be fulfilled
- Water is ampholyte and behaves as a weak base or a weak acid

# Ampholytes

- Polyprotic acid anions behave as acids and bases, ie as ampholytes



$$K_A = 1,2 \cdot 10^{-7} \text{ mol dm}^{-3}$$

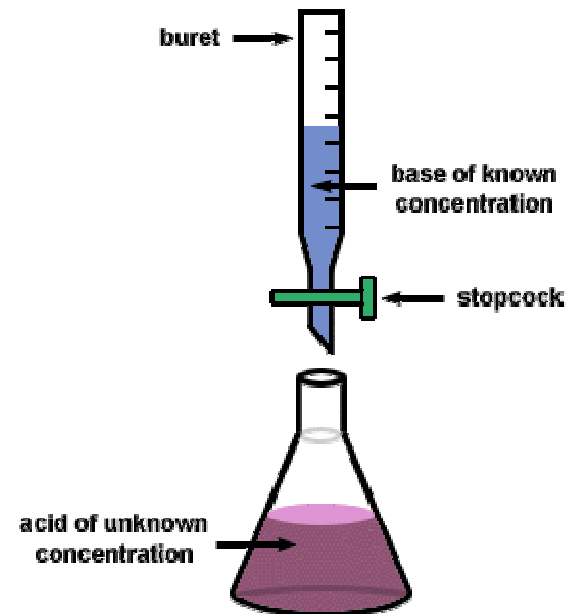


$$K_B = \frac{K_W}{K_{\text{H}_3\text{PO}_4}} = 9,1 \cdot 10^{-13} \frac{\text{mol}}{\text{dm}^3}$$

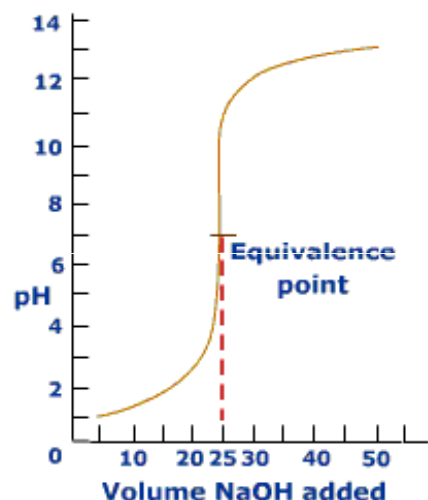
- Since the  $K_A \gg K_B$  the first reaction dominates and the pH is, without great errors, calculated from these reactions.
- If  $K_A$  and  $K_B$  are equal, both reactions must be considered in calculating the pH.

# Titration

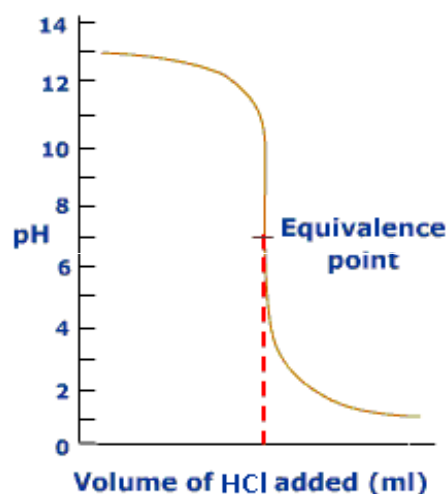
- Practical performing of reaction: acid (base) with base (acid), or neutralization
- For the end point of the titration the salt hydrolysis which occurs is significant
- At this point the exact amount of acid (base) to turn all base (acid) into salt, is added.



# Titration of strong acids and strong bases



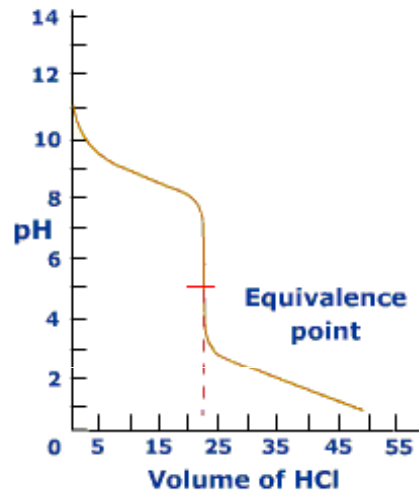
Titration curve of strong acid (HCl) with a strong base (NaOH)



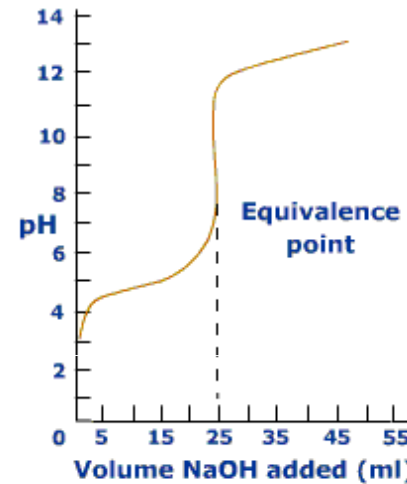
Titration curve of strong base (NaOH) with strong acid (HCl)

- Example: **strong acid + strong base** (formed salt doesn't hydrolyze)
- **pH changes rapidly, in titration end point pH=7**
- Immediately prior to the end point, concentration of base is  $10^{-3} \text{ mol/dm}^3 \rightarrow \text{pH}=11$
- Immediately after the end point, concentration of added acid is  $10^{-3} \text{ mol/dm}^3 \rightarrow \text{pH}=3$

# Titration of weak acid (base) with strong base (acid)



The pH titration curve of weak base( $\text{NH}_4\text{OH}$ ) and strong acid (HCl)



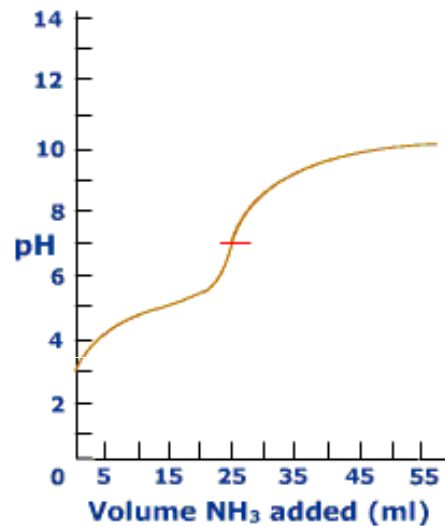
The pH titration curve of weak acid ( $\text{CH}_3\text{COOH}$ ) and strong base (NaOH)

- In both cases, the pH does not change rapidly, must be fulfilled two equilibriums:





# Titration of weak acid and weak base



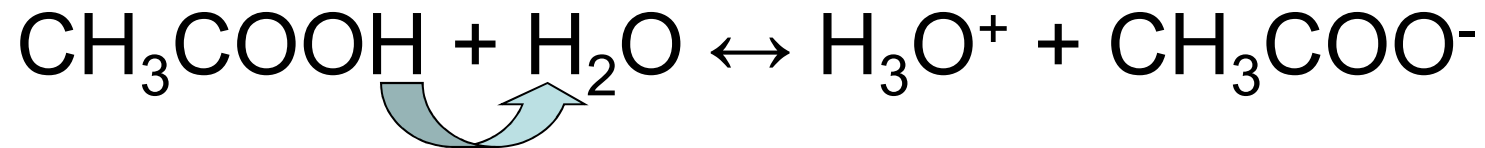
Titration curve of weak base (NH<sub>4</sub>OH) and weak acid (CH<sub>3</sub>COOH)

- very slow change in pH
- must be fulfilled three equilibriums:  $K_w$ ,  $K_a$  i  $K_b$

# Buffers

- the solution consisting of a weak acid and its salt or a weak base and its salt
- minimal change in pH with the addition of a strong base or strong acid
- important in living systems because most of the metabolic processes take place at the constant pH
- example:  $\text{CH}_3\text{COOH}$  and  $\text{CH}_3\text{COONa}$ ; aqueous solution of  $\text{NH}_3$  and  $\text{NH}_4\text{Cl}$

# Buffers



$$K_{acid} = \frac{[\text{H}_3\text{O}^+][\text{CH}_3\text{COO}^-]}{[\text{CH}_3\text{COOH}]}$$

$$K_{acid} = \frac{[\text{H}_3\text{O}^+] \cdot c_{salt}}{c_{acid}}$$