pH (acidity) of solutions. Buffer Solutions. Hydrolysis

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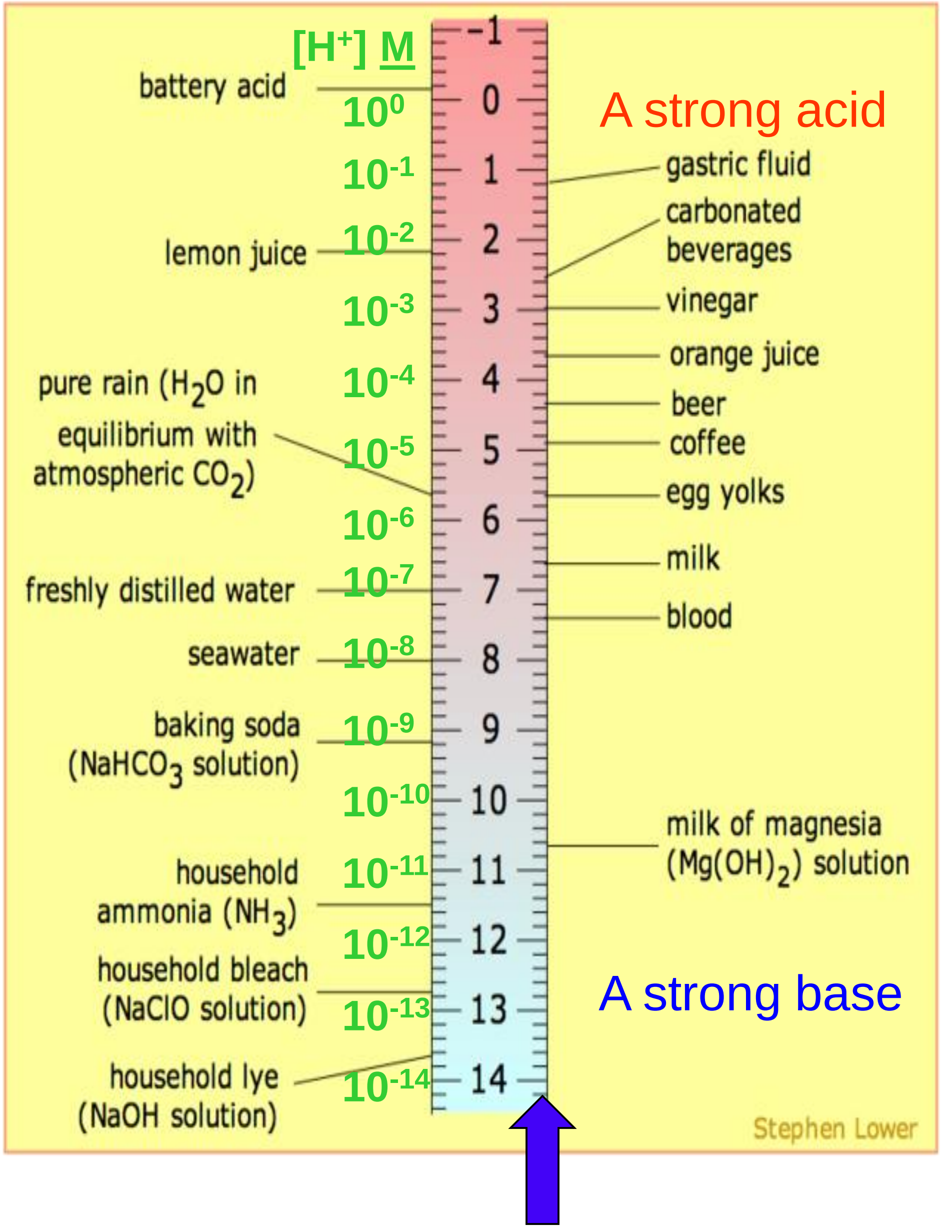
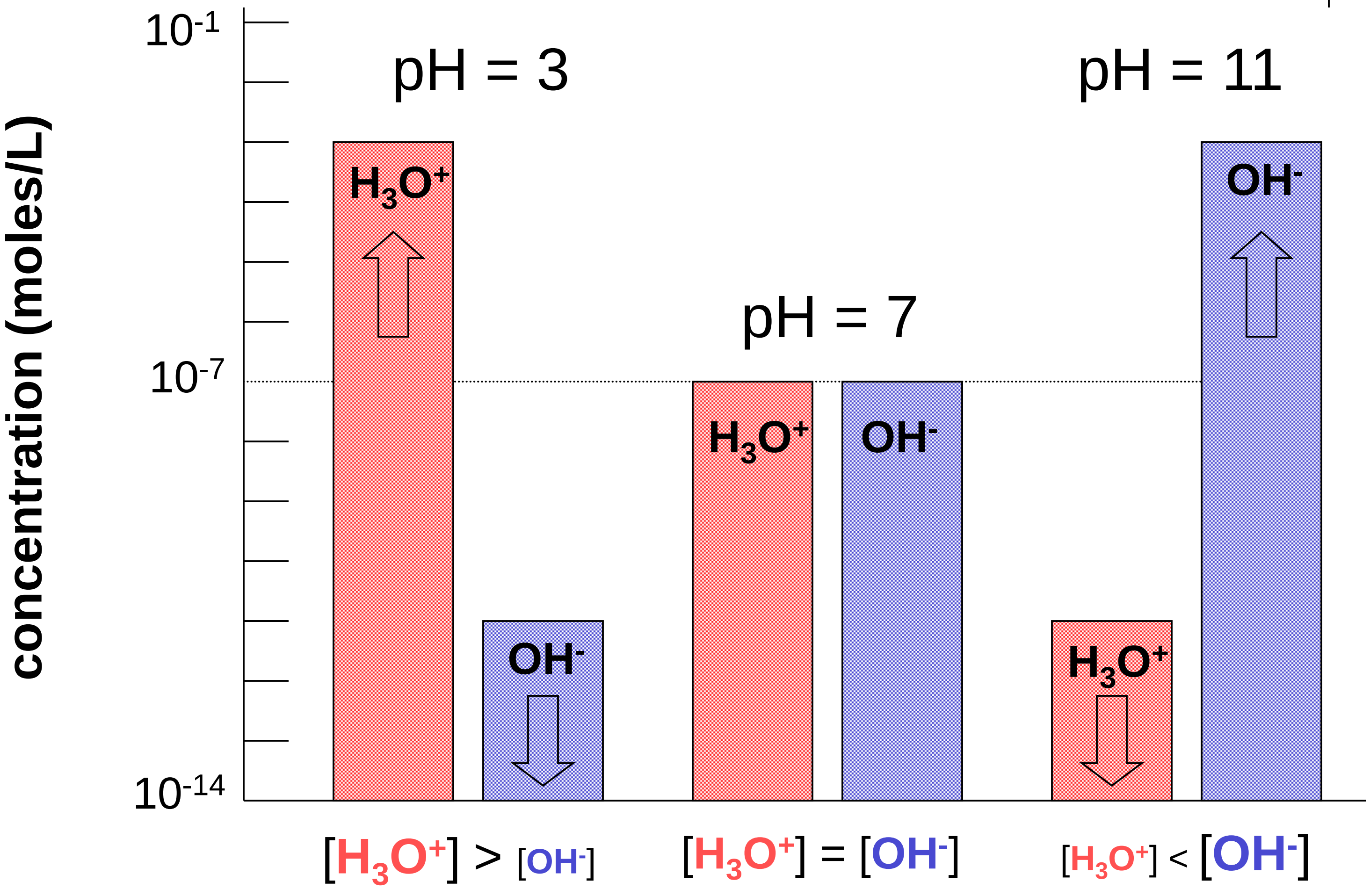
pH



- Most living cells have a *very* narrow range of tolerance for pH, i.e. $[H^+]$.
- The $[H^+]$ concentration will be important (either explicitly or implicitly) for many other topics in biology.
- $[H^+]$ is controlled in all biological organisms, and in virtually all biochemical experiments.
- Each pH unit represents a factor of 10 difference in $[H^+]$.

The pH scale goes from 0 to 14—because $[H^+][OH^-] = 10^{-14}$

Acid – Base Concentrations



acidic solution

neutral solution

basic solution

Electrolytes

(a) Nonelectrolyte **(b) Weak electrolyte** **(c) Strong electrolyte**

Strong electrolytes completely dissociated on ions

Dissociation of weak electrolytes

$$KtAn \Leftrightarrow Kt^+ + An^-.$$

$$[KtAn] = C - \alpha C = C(1 - \alpha). \qquad [Kt^+] = [An^-] = \alpha C,$$

$$K_a = \frac{[Kt^+]\,[An^-]}{[KtAn]} = \frac{\alpha C \cdot \alpha C}{C(1-\alpha)} = \frac{\alpha^2 C}{1-\alpha}.$$

$$K = \alpha^2 C,$$

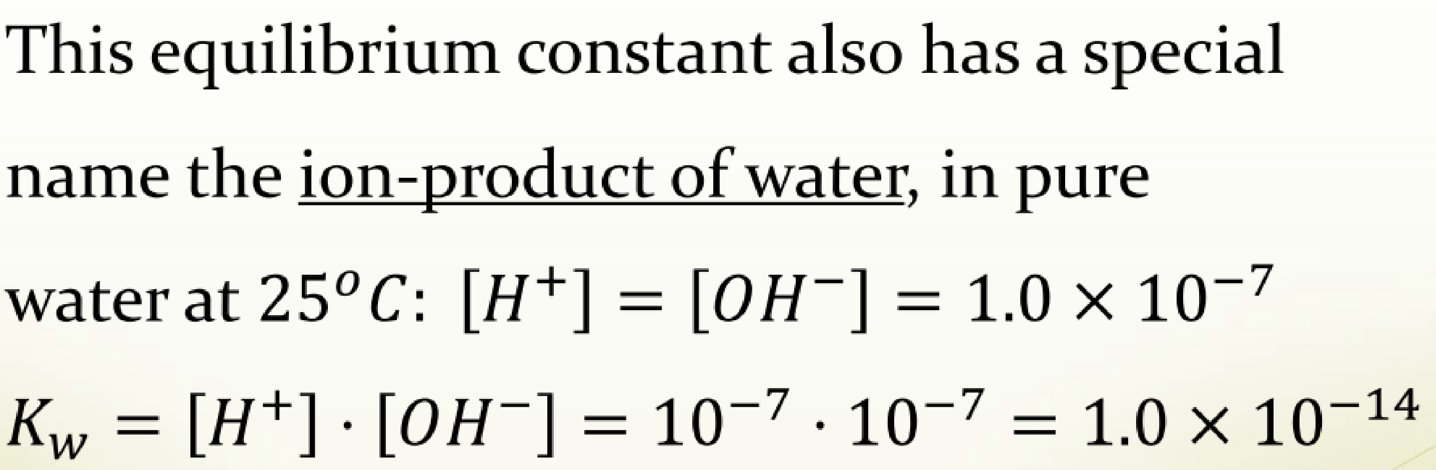
$$C_{M(H_2O)} = [H_2O] = \frac{\nu(H_2O)}{V(H_2O)} = \frac{m(H_2O)}{M_{H_2O} \cdot V_{H_2O}}.$$

$$\alpha = \sqrt{K / C}\,.$$

$$[H_2O] = m / M \cdot V = 1000 / 18 \cdot 1 = 55{,}56\,;$$

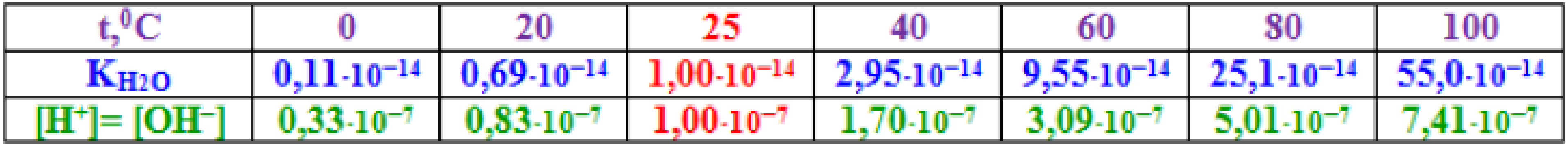
$$[H^+] = [OH^-] = \alpha \cdot [H_2O] = 1{,}8 \cdot 10^{-9} \cdot 55{,}56 = 1 \cdot 10^{-7}\,;$$

Equilibrium constant for water

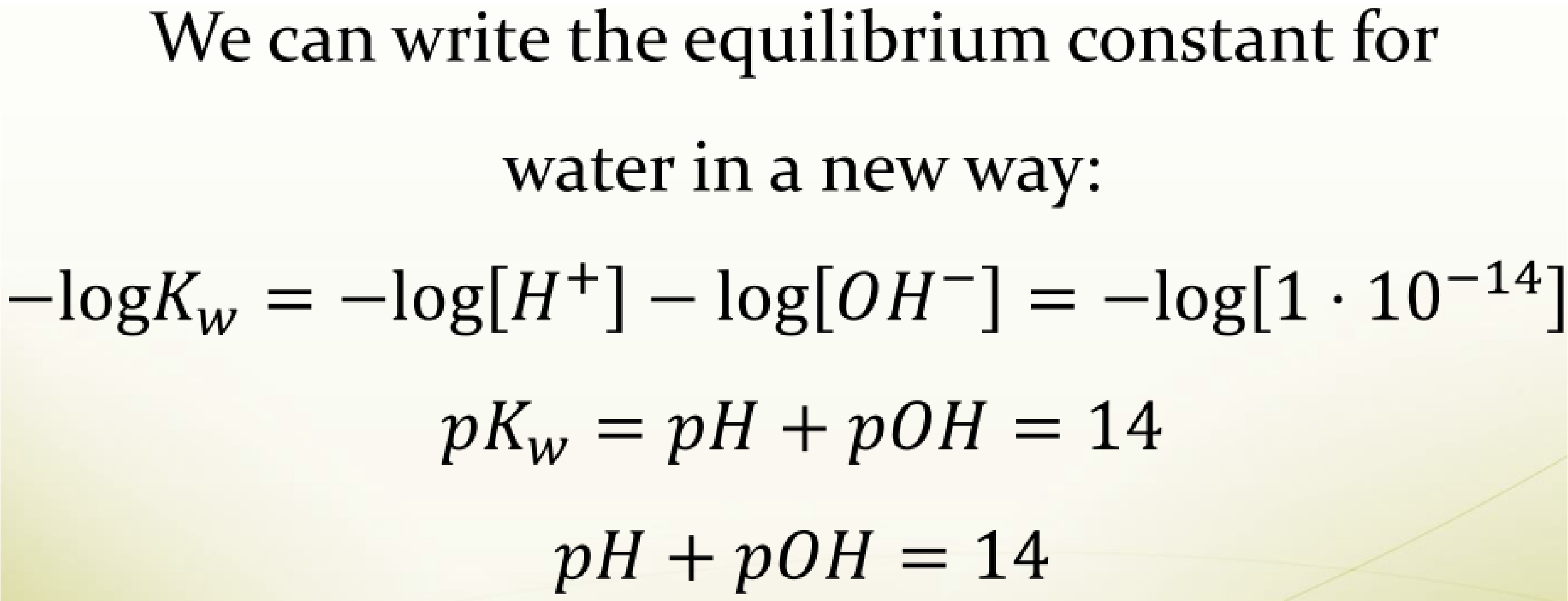


This equilibrium constant also has a special name the <u>ion-product of water</u>, in pure water at 25°C: $[H^+] = [OH^-] = 1.0 \times 10^{-7}$

$$K_w = [H^+] \cdot [OH^-] = 10^{-7} \cdot 10^{-7} = 1.0 \times 10^{-14}$$



$t, ^0C$	0	20	25	40	60	80	100
K_{H_2O}	$0{,}11 \cdot 10^{-14}$	$0{,}69 \cdot 10^{-14}$	$1{,}00 \cdot 10^{-14}$	$2{,}95 \cdot 10^{-14}$	$9{,}55 \cdot 10^{-14}$	$25{,}1 \cdot 10^{-14}$	$55{,}0 \cdot 10^{-14}$
$[H^+] = [OH^-]$	$0{,}33 \cdot 10^{-7}$	$0{,}83 \cdot 10^{-7}$	$1{,}00 \cdot 10^{-7}$	$1{,}70 \cdot 10^{-7}$	$3{,}09 \cdot 10^{-7}$	$5{,}01 \cdot 10^{-7}$	$7{,}41 \cdot 10^{-7}$



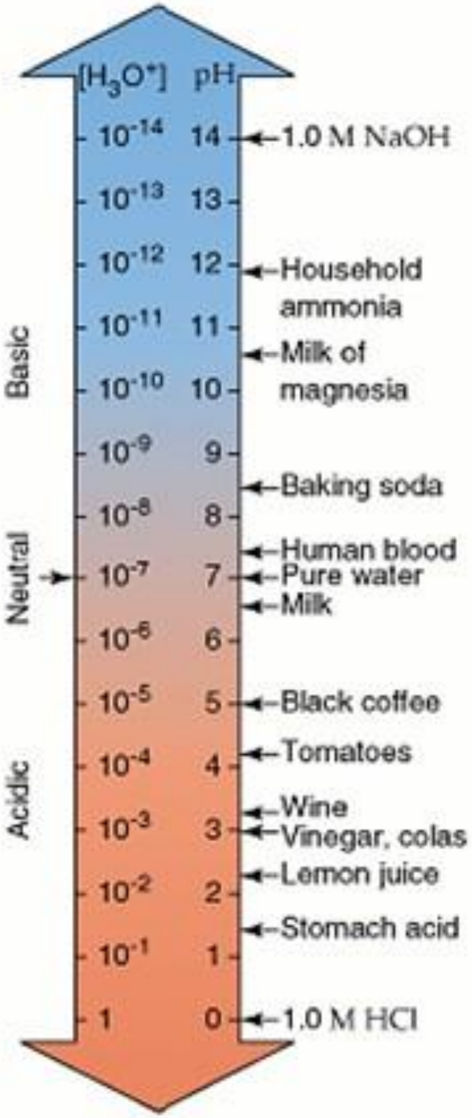
We can write the equilibrium constant for water in a new way:

$$-\log K_w = -\log[H^+] - \log[OH^-] = -\log[1 \cdot 10^{-14}]$$

$$pK_w = pH + pOH = 14$$

$$pH + pOH = 14$$

pH Scale



$$[H^+][OH^-] = 1x10^{-14}$$

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

$$[H^+] = \text{antilog}(-pH)$$

$$pOH = -\log[OH^-]$$

$$[OH^-] = \text{antilog}(-pOH)$$

Acidic solution: $pH < 7$

Neutral solution: $pH = 7$

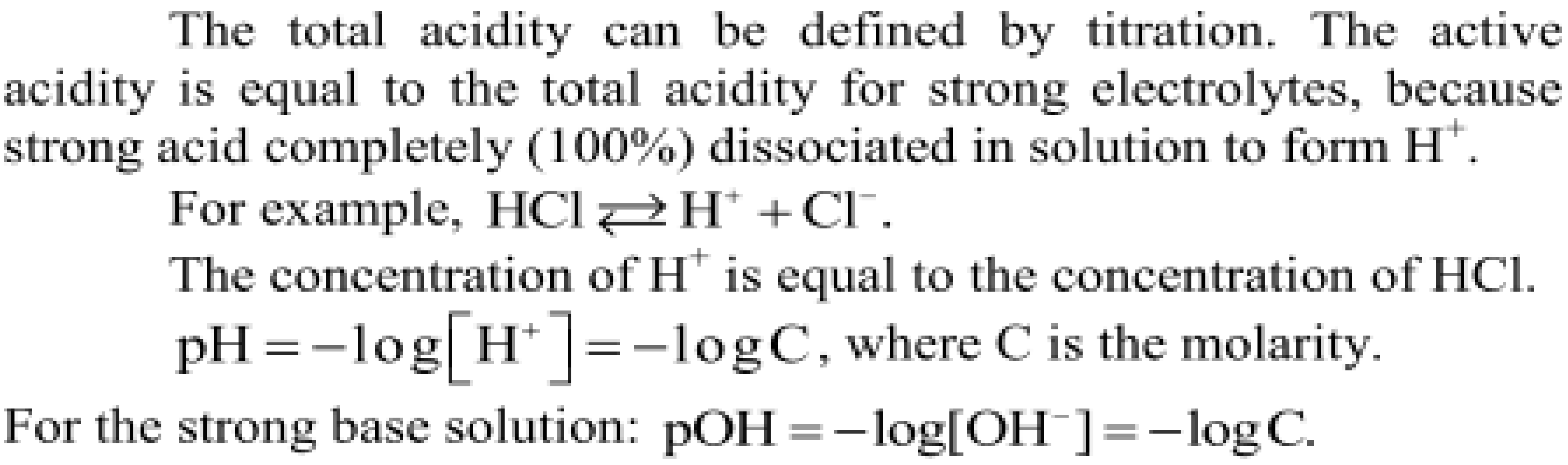
Basic solution: $pH > 7$

Ion concentration

$$[H^+] = \alpha \cdot [acid] \quad [OH^-] = \alpha \cdot [base]$$

- For strong electrolytes α=1, so concentration of [H+] and [OH-] is equal concentration of base or concentration of acid respectively.
- For weak acids and bases we use degree of dissociation α. If we know Ka or Kb we use following formula

$$\alpha = \sqrt{K / C} \, .$$



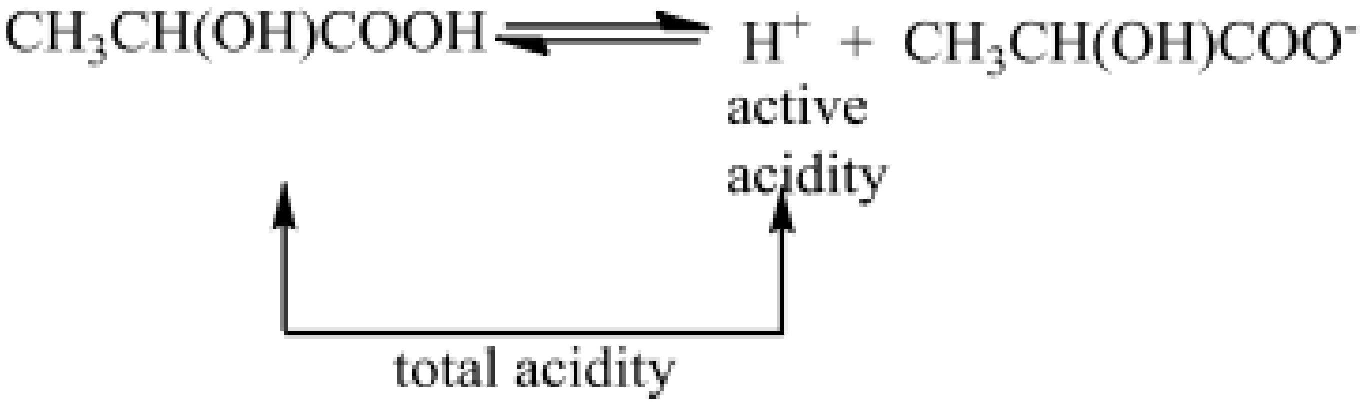
The total acidity can be defined by titration. The active acidity is equal to the total acidity for strong electrolytes, because strong acid completely (100%) dissociated in solution to form H^+.

For example, $HCl \rightleftarrows H^+ + Cl^-$.

The concentration of H^+ is equal to the concentration of HCl.

$pH = -\log\left[H^+\right] = -\log C$, where C is the molarity.

For the strong base solution: $pOH = -\log[OH^-] = -\log C$.



$$CH_3CH(OH)COOH \rightleftharpoons H^+ + CH_3CH(OH)COO^-$$

To calculate a pH of weak acid solution it is necessary to know besides the concentrations either the degree of the dissociation (α), or the dissociation constant (Ka):

$$\mathbf{pH = -\log C \cdot \alpha},\ \text{or}$$

$$\mathbf{or: pH = -\log \sqrt{Ka \times C}}.$$

For the weak base solution:

$$pOH = -\log C \times \alpha,\ \text{or}\ = -\log \sqrt{\mathbf{K_b \times C}},$$

where α is the degree of dissociation (ionization) of the base;

K_b is the base dissociation (ionization) constant;

C is the molarity.

Calculation of the pH of Weak Acid Solutions

Problem: What is the pH of a solution of 1.00 M nitrous acid, $K_a = 4.0 \times 10^{-4}$

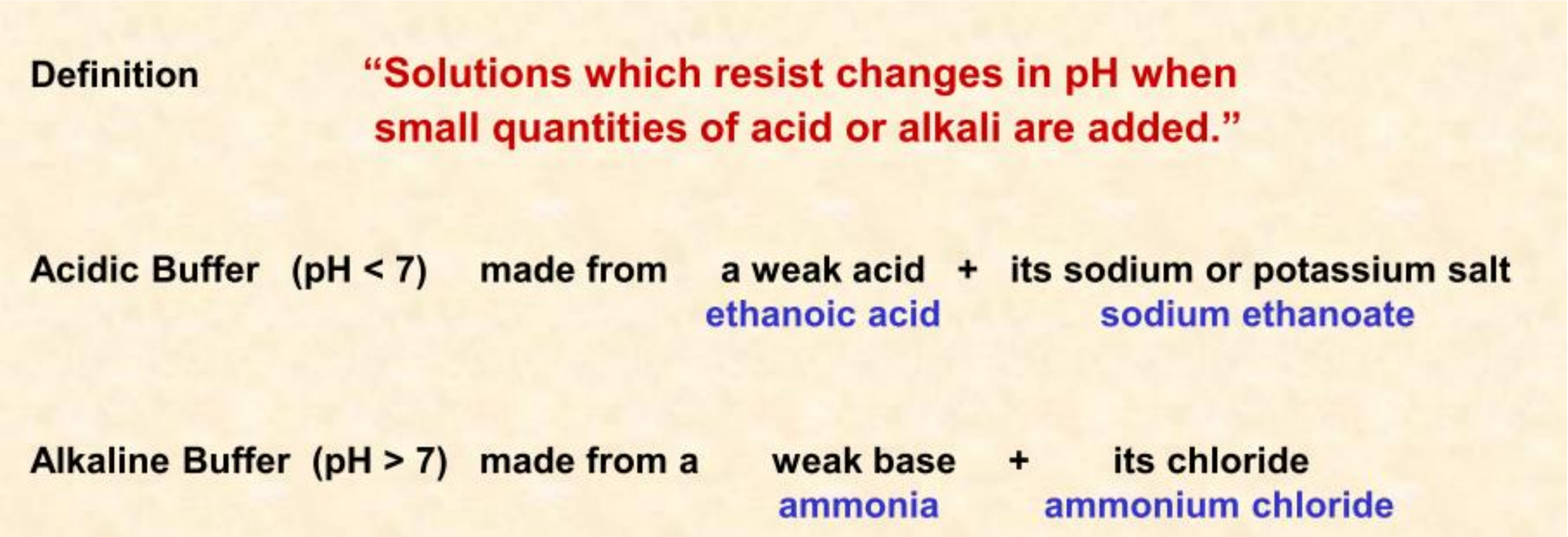
- Major species in solution: HNO_2 and H_2O
- Which species can generate H^+ ions?
  - $HNO_2(aq) = H^+(aq) + NO_2^-(aq)$ $K_a = 4.0 \times 10^{-4}$
  - $H_2O(aq) = H^+(aq) + OH^-(aq)$ $K_w = 1.0 \times 10^{-14}$
  - Ignore contribution from water, $K_a >> K_w$

$$\left[H^+\right] = \sqrt{4.0 \cdot 10^{-4} \cdot 1.0} = 2 \cdot 10^{-2}$$

pH = $-\log [H^+] = -\log 0.02 = 1.70$

Characteristics of Buffer Solutions

- Contain weak acids or weak bases and their corresponding conjugate partners (common ions).
- Resist changes in pH.
- Buffering capacity depends on concentrations of weak acid or weak base and their common ions.
- Effective pH buffering range ~ p$K_a \pm 1$



Definition "Solutions which resist changes in pH when small quantities of acid or alkali are added."

Acidic Buffer	**(pH < 7)**	**made from**	**a weak acid** ethanoic acid	**+**	**its sodium or potassium salt** sodium ethanoate
Alkaline Buffer	**(pH > 7)**	**made from a**	**weak base** ammonia	**+**	**its chloride** ammonium chloride

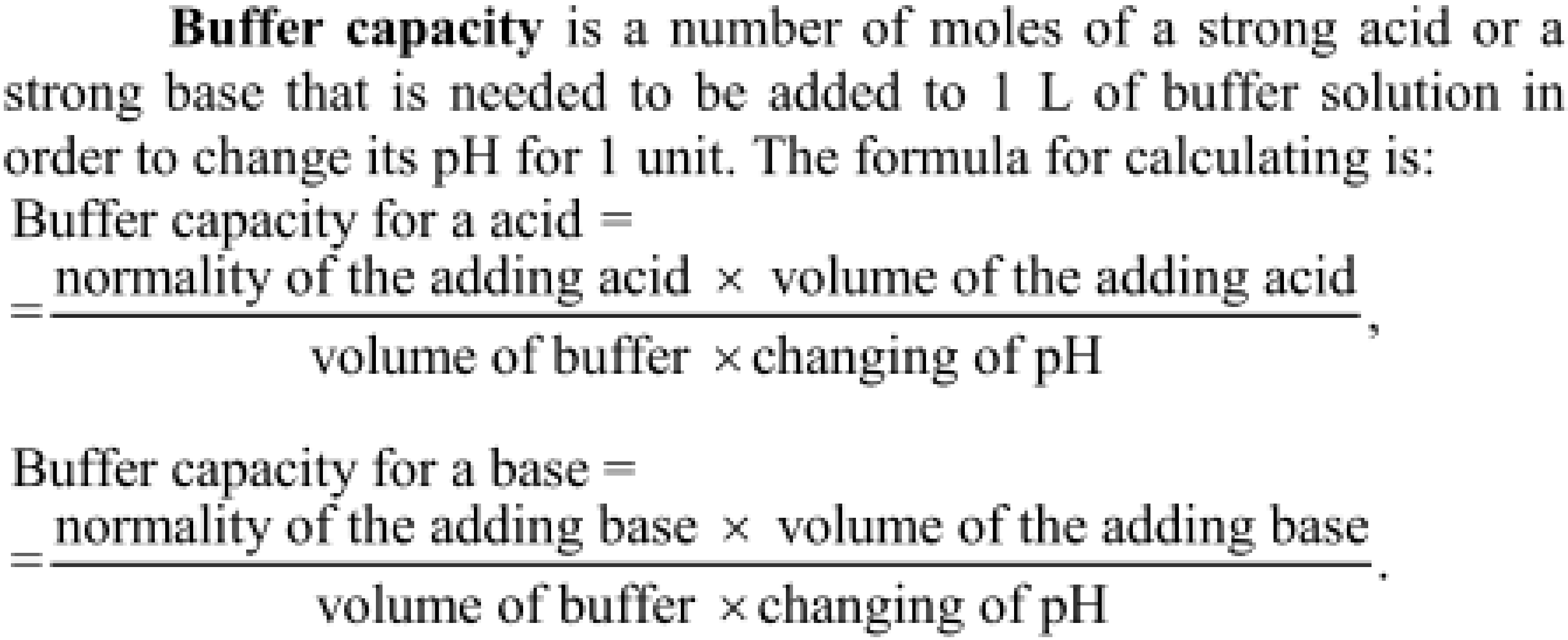
Characteristics of Buffer Solutions

1. Buffers contain relatively large amounts of the weak acids (HA) and their conjugate base (A^-), (or weak bases and their conjugate acids)
2. Buffer pH is determined by the pK_a of the acid HA and the molar ratio of the conjugate base to acid: $[A^-]/[HA]$.
3. Buffer pH changes very little because the ratio $[A^-]/[HA]$ changes very little when a small amount of strong acid or strong base is added.
4. $[H_3O^+]$ in buffer solutions remains more or less constant:

   Most of H^+ from strong acid is absorbed by the conjugate base A^-; most of OH^- added from strong base reacts with acid HA in the buffer to yield A^- and H_2O.

Buffering Capacity

- How much H_3O^+ or OH^- the buffer can absorb without significantly changing its pH.
- Depends on the concentrations of HA and A^-.
- High [HA] and [A^-] lead to large buffering capacity.
- Optimal buffering occurs when [HA] = [A^-];
- Ratio [A^-] / [HA] ~ 1 strong resist to change when either H_3O^+ or OH^- is added.



Buffer capacity is a number of moles of a strong acid or a strong base that is needed to be added to 1 L of buffer solution in order to change its pH for 1 unit. The formula for calculating is:

Buffer capacity for a acid =

$$= \frac{\text{normality of the adding acid} \times \text{volume of the adding acid}}{\text{volume of buffer} \times \text{changing of pH}},$$

Buffer capacity for a base =

$$= \frac{\text{normality of the adding base} \times \text{volume of the adding base}}{\text{volume of buffer} \times \text{changing of pH}}.$$

Some Common Buffers

Buffers	pK_a	pH Range
$HCOOH - HCOONa$	3.74	2.74 – 4.74
$CH_3COOH - CH_3COONa$	4.74	3.74 – 5.74
$KH_2PO_4 - K_2HPO_4$	7.21	6.20 – 8.20
$CO_2/H_2O - NaHCO_3$	6.37	5.40 – 7.40
$NH_4Cl - NH_3$	9.25	8.25 – 10.25

Buffer solutions

- Definition: a solution that resists change in pH
  - Typically a mixture of the acid and base form of a chemical
  - Can be adjusted to a particular pH value

$$H_2O + CO_2 \leftrightarrow H_2CO_3 \leftrightarrow H^+ + HCO_3^-$$

Blood: pH = 7.35-7.45

$$[H+] = K_a \text{ x } [acid]/[base]$$

Then take the log of both sides

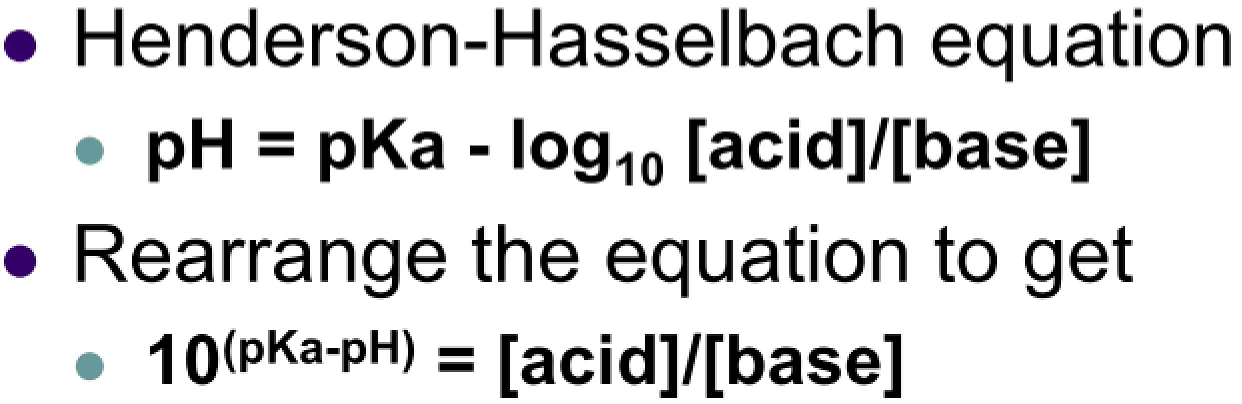
- $\log_{10}[H+] = \log_{10}K_a + \log_{10} [acid]/[base]$

Henderson-Hasselbach equation

- **$pH = pKa - \log_{10} [acid]/[base]$**

Weak acids, their conjugate bases, and buffers…

- Weak acids have only a modest tendency to shed their protons (definition of an acid).
- When they do, the corresponding negatively charged anion becomes a willing proton acceptor, and is called the *conjugate base.*
- The properties of a *buffer* rely on a balance between a weak acid and its conjugate base.



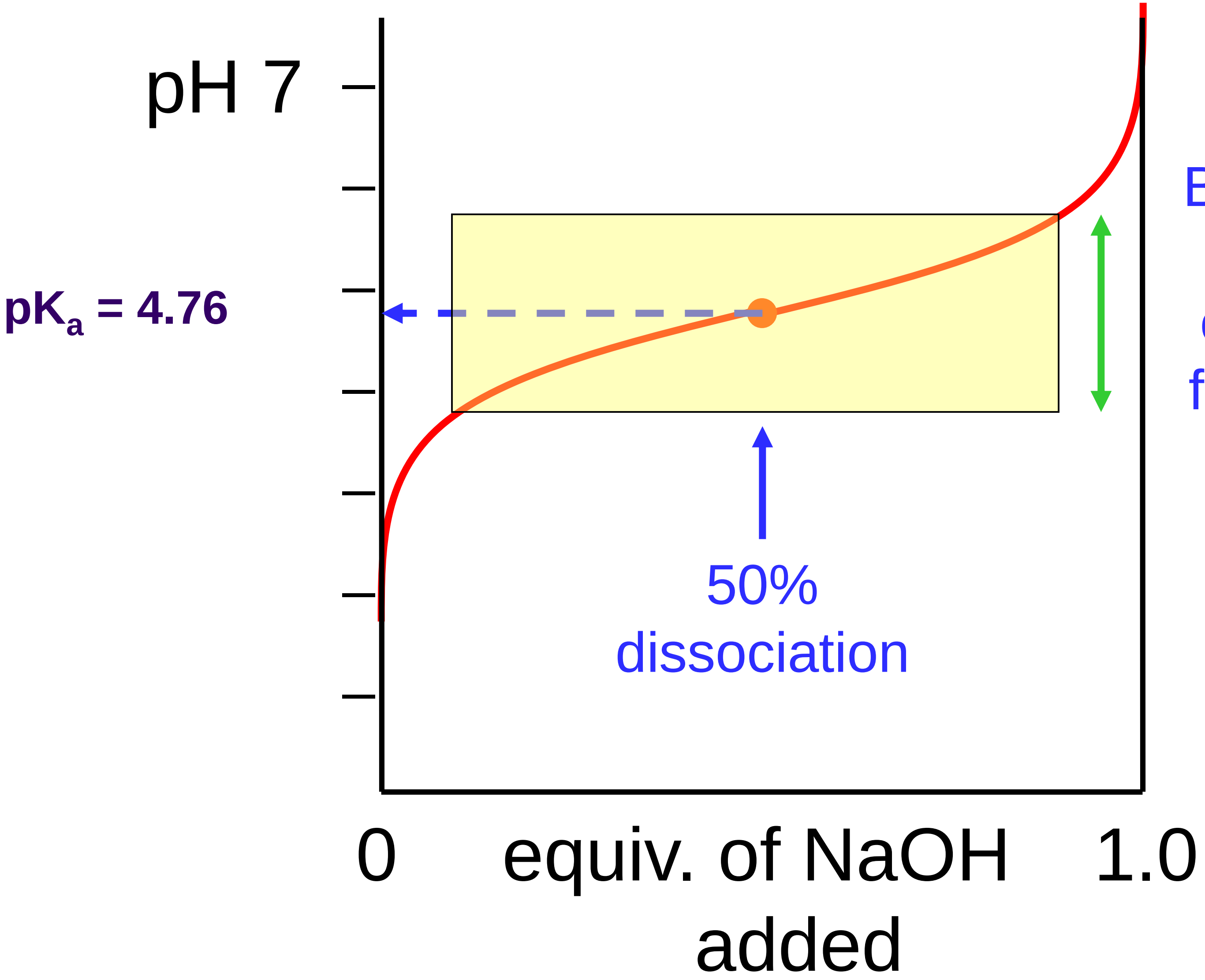
- Henderson-Hasselbach equation
  - **$pH = pKa - \log_{10} [acid]/[base]$**
- Rearrange the equation to get
  - **$10^{(pKa-pH)} = [acid]/[base]$**

Chemistry of buffers

- What happens when the concentration of the acid and base are equal?
  - Example: Prepare a buffer with 0.10M acetic acid and 0.10M acetate
    - $pH = pKa - \log_{10} [acid]/[base]$
    - $pH = pKa - \log_{10} [0.10]/[0.10]$
    - $pH=pKa$
    - Thus, the pH where equal concentrations of acid and base are present is defined as the pKa

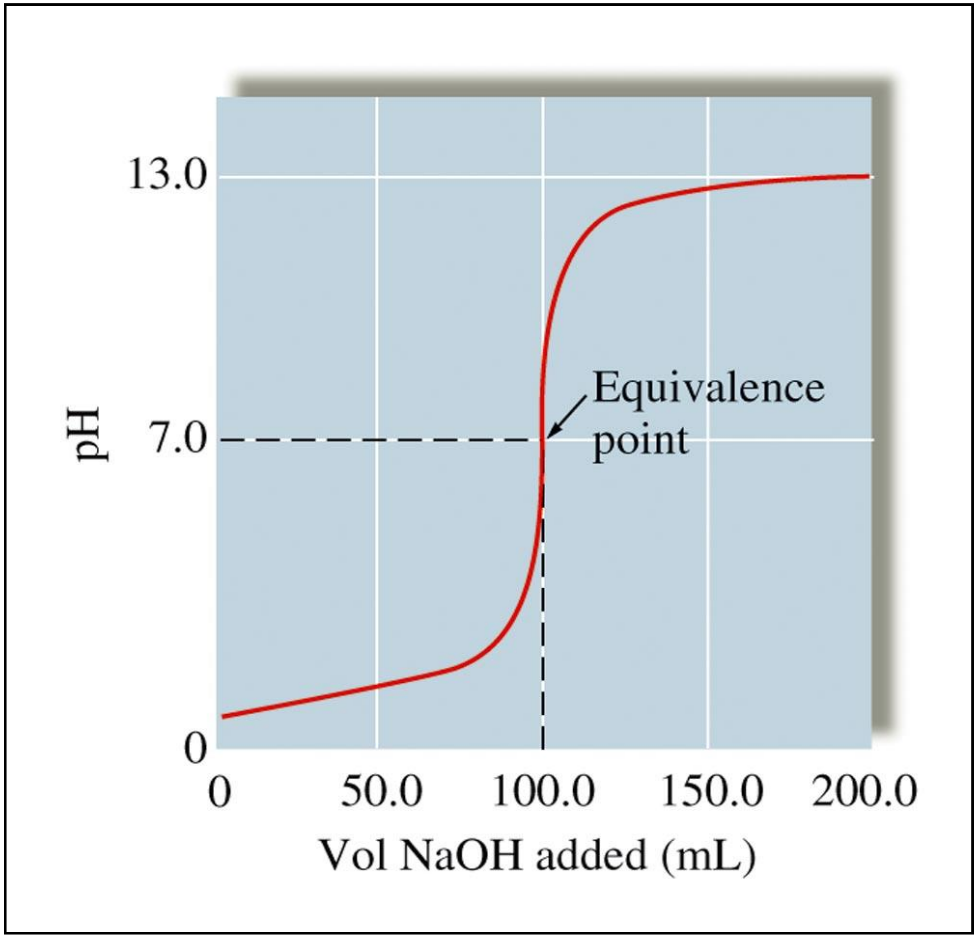
- A buffer works most effectively at pH values that are $\pm$ 1 pH unit from the pKa (the buffer range)

Titration of acetic acid with sodium hydroxide

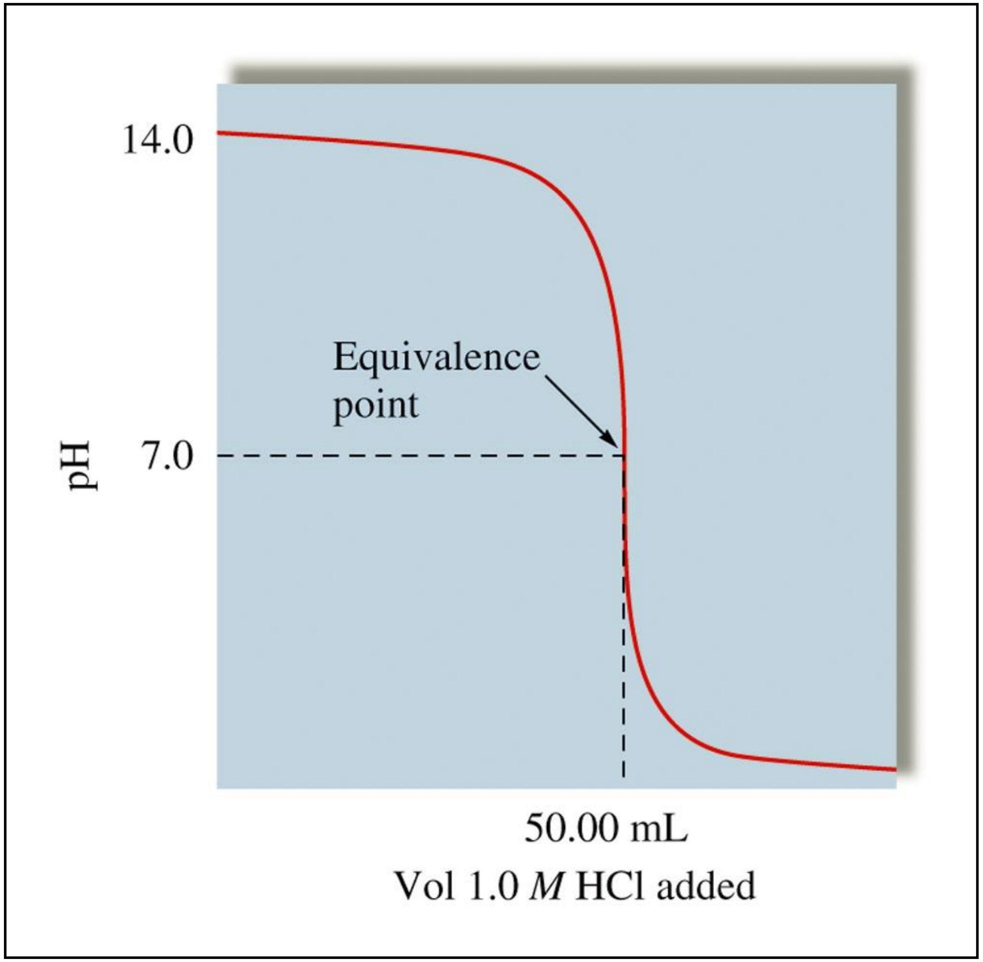


Buffering range: only small pH changes result from addition of base or acid

The pH Curve for the Titration of 50.0 mL of 0.200 *M* HNO_3 with 0.100 *M* NaOH



The pH Curve for the Titration of 100.0 mL of 0.50 *M* NaOH with 1.0 *M* HCl



Ways to measure pH

- Indicator dyes and test strips
  - Less precise
  - Each indicator is only good for a small pH range (1-2 pH units)
  - But may be good for field usage, or measuring small volumes, or dealing with noxious samples.

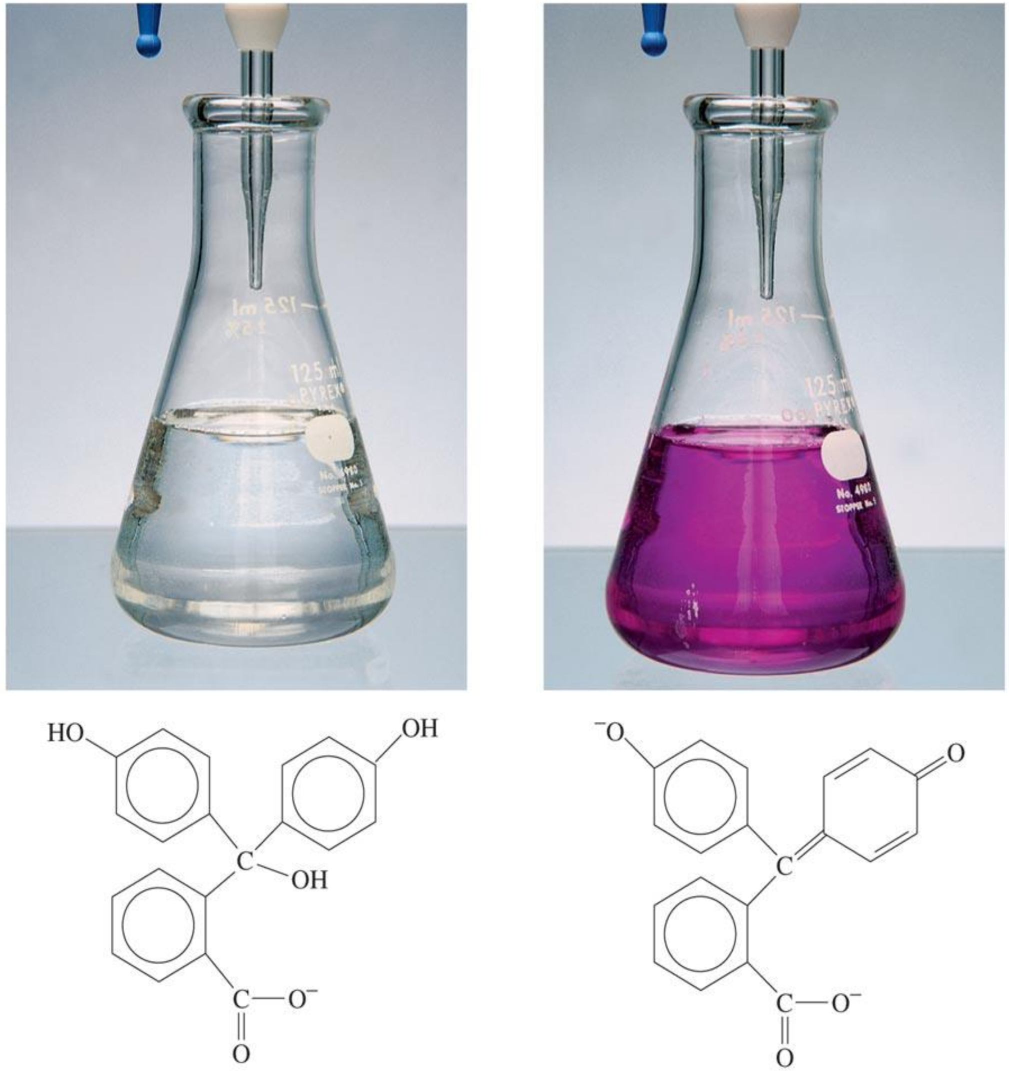
Acid-Base Indicators

- An indicator is a substance added to acid or base solution to marks the *end point* of a titration by the change of its color.

  For example, phenolphthalein changes from *colorless to pink* at the *end point* when an acid is titrated with a base.

- The *end point* of a titration should correspond to the *equivalence points* of the acid-base reaction.

The Acid and Base Forms of the Indicator Phenolphthalein



(Colorless acid form, HIn)

(Pink base form, In$^-$)

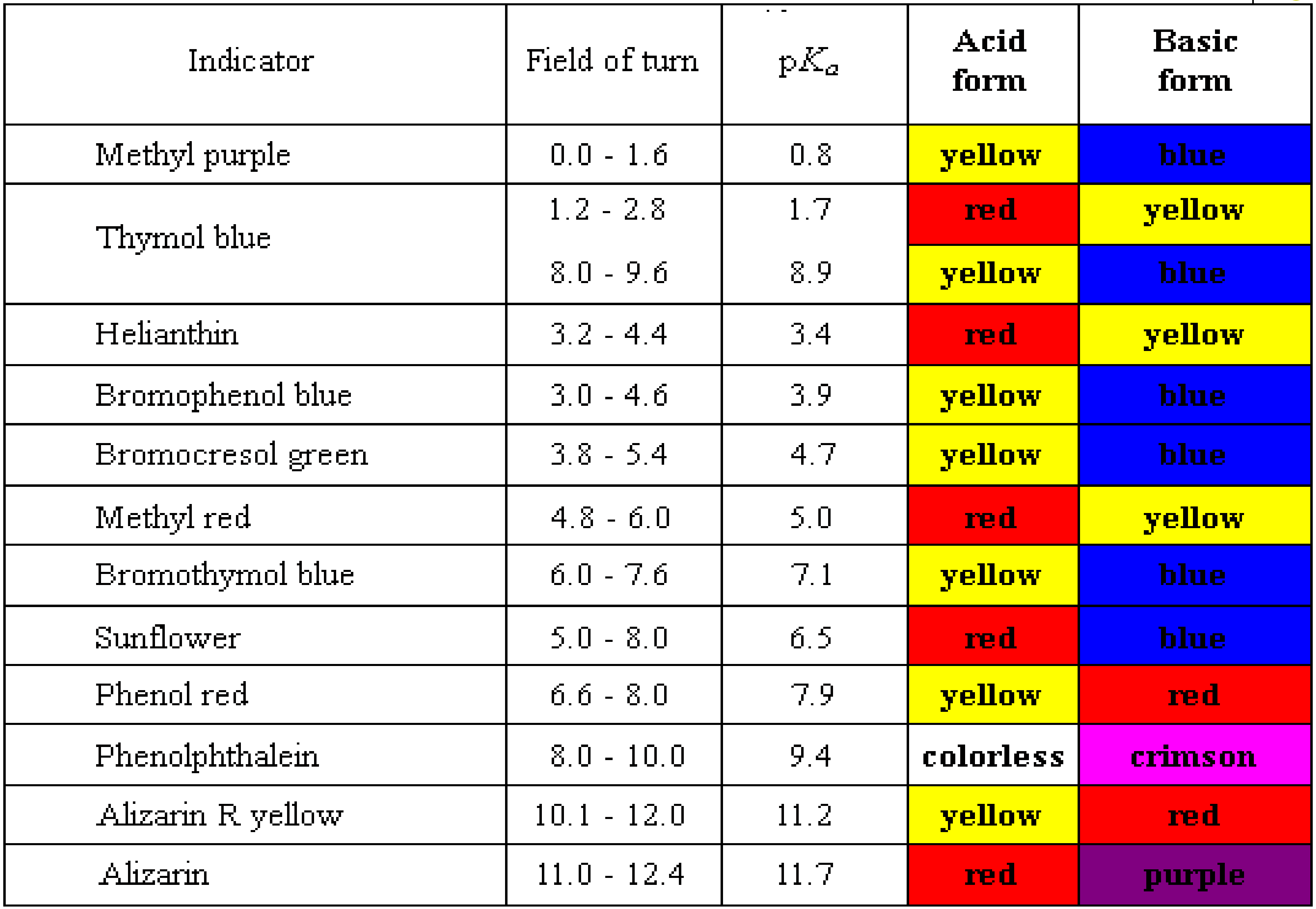
The Methyl Orange Indicator is Yellow in Basic Solution and Red in Acidic Solution

Choosing Indicators

1. The pH range for color changes should occur within the sharp vertical rise (or drop) in the pH-titration curves.
2. An indicator changes color at pH = pK_a ± 1, where pK_a is that of the indicator used.

	Indicators:	Acid Color	Base Color	pH Range	Type of Titration
1.	Methyl orange	Orange	Yellow	3.2 – 4.5	strong acid-strong base strong acid-weak base
2.	Bromocresol green	Yellow	Blue	3.8 – 5.4	strong acid-strong base strong acid-weak base
3.	Methyl red	Red	Yellow	4.5 – 6.0	strong acid-strong base strong acid-weak base
4.	Bromothymol blue	Yellow	Blue	6.0 – 7.6	strong acid-strong base
5.	Phenol Red	Orange	Red	6.8 – 8.2	strong acid-strong base weak acid-strong base

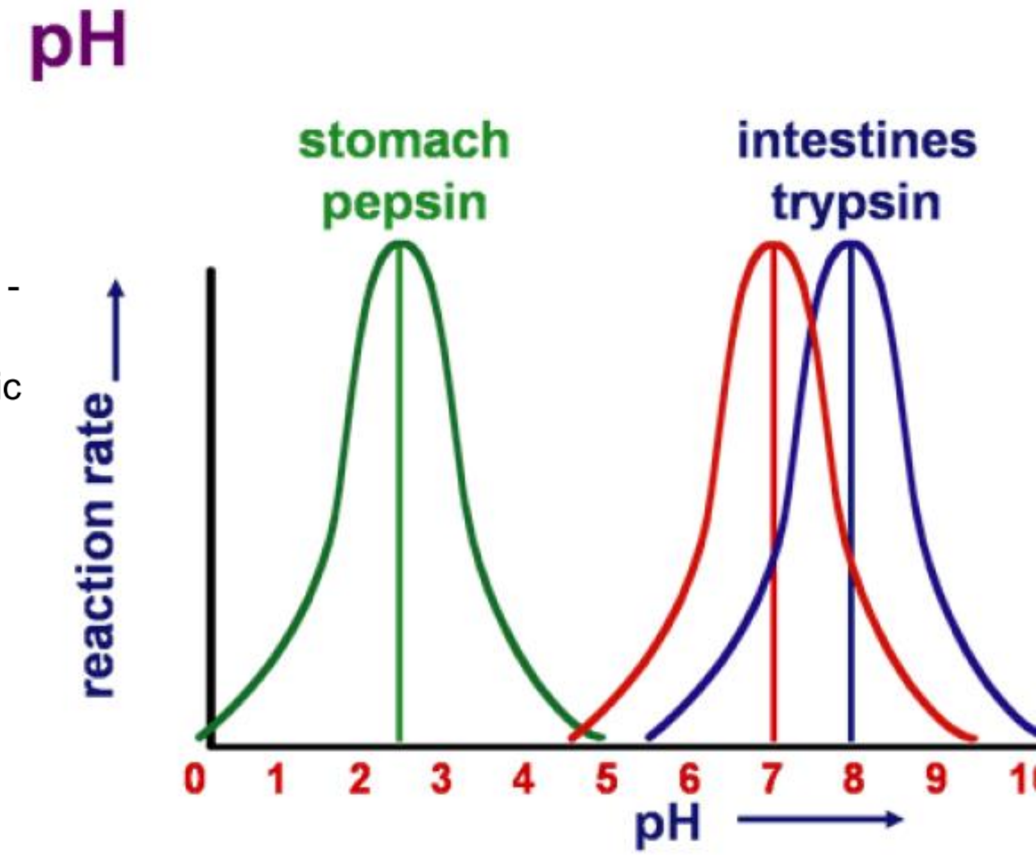
pH Ranges for Indicators



Indicator	Field of turn	pK_a	**Acid form**	**Basic form**
Methyl purple	0.0 - 1.6	0.8	**yellow**	**blue**
Thymol blue	1.2 - 2.8	1.7	**red**	**yellow**
	8.0 - 9.6	8.9	**yellow**	**blue**
Helianthin	3.2 - 4.4	3.4	**red**	**yellow**
Bromophenol blue	3.0 - 4.6	3.9	**yellow**	**blue**
Bromocresol green	3.8 - 5.4	4.7	**yellow**	**blue**
Methyl red	4.8 - 6.0	5.0	**red**	**yellow**
Bromothymol blue	6.0 - 7.6	7.1	**yellow**	**blue**
Sunflower	5.0 - 8.0	6.5	**red**	**blue**
Phenol red	6.6 - 8.0	7.9	**yellow**	**red**
Phenolphthalein	8.0 - 10.0	9.4	**colorless**	**crimson**
Alizarin R yellow	10.1 - 12.0	11.2	**yellow**	**red**
Alizarin	11.0 - 12.4	11.7	**red**	**purple**

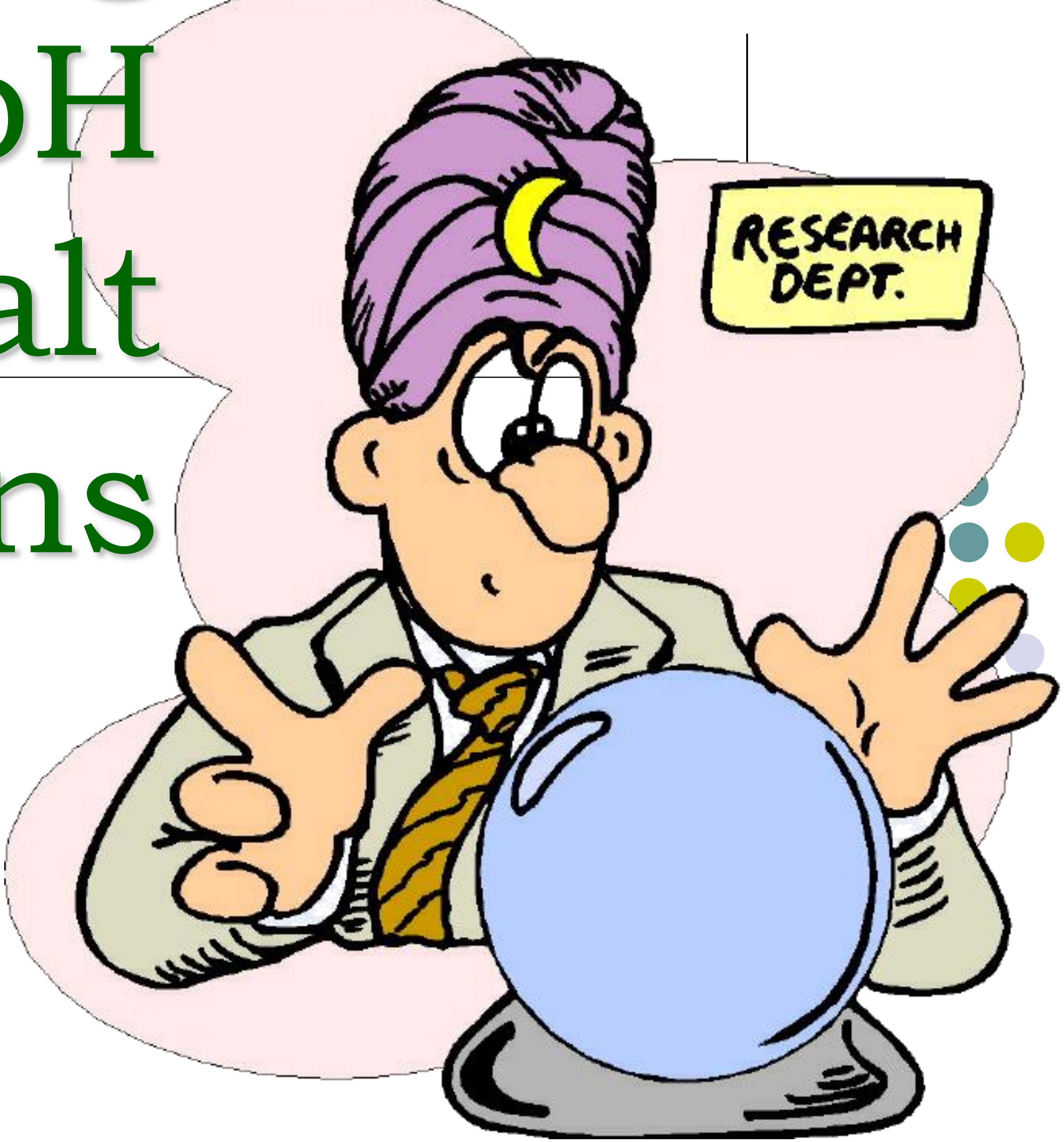
Why is pH important in biology?

- pH affects solubility of many substances.
- pH affects structure and function of most proteins - including enzymes.



- pH affects solubility of many substances.
- pH affects structure and function of most proteins - including enzymes.
- Many cells and organisms (esp. plants and aquatic animals) can only survive in a specific pH environment.

Predicting the pH of salt solutions



Hydrolysis

This reaction, in which water reacts with an ionic species in solution, is called <u>hydrolysis</u>.

Salts of strong acids and strong bases (for example, NaCl) do not hydrolyze: PH=7

$$NaCl + H_2O = no\ reaction$$

The ions of salts can have an influence on the pH of a solution.

Ions that come from a strong acid or base do not influence the pH.

WHY?

Since strong acids and bases are 100% ionized in water, the ions are unable to reform the molecular acid or the base in water.

$HCl + H_2O \rightarrow H_3O^+ + Cl^-$ **"one way arrow"**

$NaOH + H_2O \rightarrow Na^+ + OH^- + H_2O$ **"one way arrow"**

NaCl is the salt that comes from a strong acid and a strong base.

What would the pH of a sodium chloride solution (at 25^{o}C)?

pH = 7

What gives rise to this pH?

Auto hydrolysis of water.

$$H_2O + H_2O \leftrightarrow H_3O^+ + OH^-$$

Salts that contain ions that come from a weak acid or base.

weak acid: HNO_2

A salt containing the anion of the weak acid and the cation from a strong base.

KNO_2

Add water: $KNO_2(s) + H_2O \rightarrow K^+ + NO_2^- + H_2O$

Hydrolysis: $NO_2^- + H_2O \leftrightarrow HNO_2 + OH^-$

A basic solution.

Calculate the pH of a 0.10 M KNO_2 solution. $K_a(HNO_2) = 4.5\ 10^{-4}$.

CHEMISTRY: $KNO_2(s) \rightarrow K^+ + NO_2^-$

More Chemistry: $NO_2^- + H_2O \leftrightarrow HNO_2 + OH^-$

Equilibrium:

I	0.10	N/A	0	0
C	-X	N/A	+X	+X
E	0.10-X	N/A	+X	+X

$$K_b = \frac{[HNO_2][OH^-]}{[NO_2^-]}$$

$K_aK_b = 1.0 \times 10^{-14}$

$$2.22 \times 10^{-11} = \frac{[x][x]}{[0.10 - x]}$$

Try dropping

$pOH = -\log[OH^-] = 5.83$

$x = 1.49 \times 10^{-6} = [OH^-]$

$pH = 14 - 5.83 = 8.17$

Salts that contain ions that come from a weak acid or base.

Weak Base: $(CH_3)_3N$ **trimethylamine**

A salt containing the cation of the weak base and the anion from a strong acid.

$(CH_3)_3NHCl$ trimethylammonium chloride

Add water: $(CH_3)_3NHCl\ (s) + H_2O \rightarrow (CH_3)_3NH^+ + Cl^- + H_2O$

Hydrolysis: $(CH_3)_3NH^+ + H_2O \leftrightarrow (CH_3)_3NHOH + H^+$

An acidic solution.

Calculate the pH of a 0.10 $(CH_3)_3NHCl$ solution.

$K_b((CH_3)_3NHCl) = 7.4\ 10^{-5}$.

CHEMISTRY: $(CH_3)_3NHCl(s) + H_2O \rightarrow (CH_3)_3NH^+ + Cl^- + H_2O$

More Chemistry: $(CH_3)_3NH^+ + H_2O \leftrightarrow (CH_3)_3NHOH^- + H^+$

Equilibrium:

I	0.10	N/A	0	0
C	-x	N/A	+x	+x
E	0.10-x	N/A	+x	+x

$$K_a = \frac{[(CH)_3NHOH^-][H^+]}{[(CH)_3NH^+]}$$

$$K_aK_b = 1.0 \times 10^{-14}$$

$$1.35 \times 10^{-10} = \frac{[x][x]}{[0.10 - x]}$$

Try dropping

$$pH = -\log[H^+] = 5.43$$

$$x = 3.68 \times 10^{-6} = [H^+]$$

What if both ions of a salt come from weak acid and a weak base?

Then the K_a and K_b of the acid or base from which the ions come from must be compared.

$NH_4CN(aq)$ $K_a(NH_4^+) = 5.6 \times 10^{-10}$ $K_b(CN^-) = 2.04 \times 10^{-5}$

note: K_a was calculated from $K_b(NH_3)$ and K_b from K_a HCN

Since $K_b(CN^-)$ is greater than $K_a(NH_4^+)$, the solution is basic.

Salts of weak acids and strong bases (for example CH_3COONa) hydrolyze, producing a basic solution: pH>7.

$$CH_3COONa + H_2O \rightleftarrows CH_3COOH + NaOH$$

$$CH_3COO^- + Na^+ + HOH \rightleftarrows CH_3COOH + Na^+ + OH^-$$

$$Na^+ + HOH = no\ reaction$$

Water reacts with a CH_3COO^- to produce OH^-.

A solution of CH_3COONa has a pH>7.

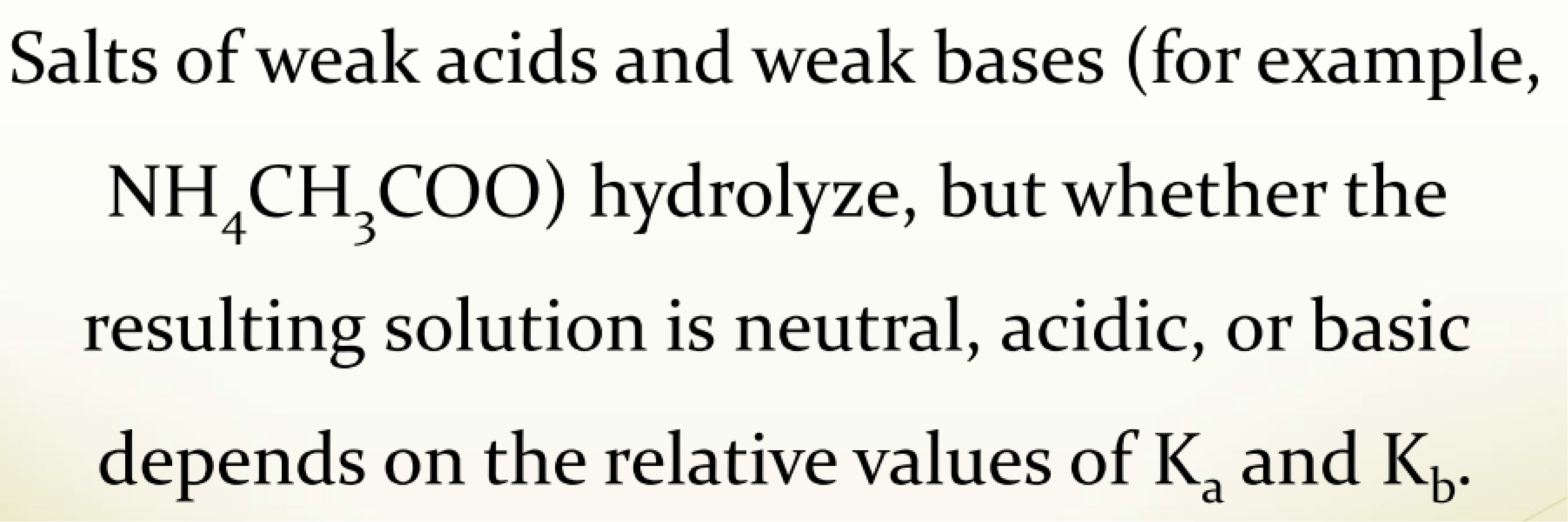
Salts of strong acids and weak bases (for example, NH_4Cl) hydrolyze, producing a acidic solution: pH<7.

$$NH_4Cl \rightleftarrows NH_4^+ + Cl^-$$

$$Cl^- + H_2O \rightleftarrows no\ reaction$$

$$NH_4^+ + H_2O \rightleftarrows H^+ + NH_4OH$$

$$NH_4Cl + H_2O \rightleftarrows NH_4OH + HCl$$



Salts of weak acids and weak bases (for example, NH_4CH_3COO) hydrolyze, but whether the resulting solution is neutral, acidic, or basic depends on the relative values of K_a and K_b.

Steps in determining pH

1. Write the ions that form: e.g. NH_4CN

   $\rightarrow NH_4^+ + CN^-$

2. Determine the reaction ions have with water:

$NH_4^+ + OH^- \leftrightarrow NH_3 + H_2O$, $NH_3 + H_2O \leftrightarrow NH_4^+ + OH^-$

$CN^- + H^+ \rightarrow HCN$, $HCN \rightarrow CN^- + H^+$

3. Look up the Ka of the conjugate acid and the Kb of the conjugate base:

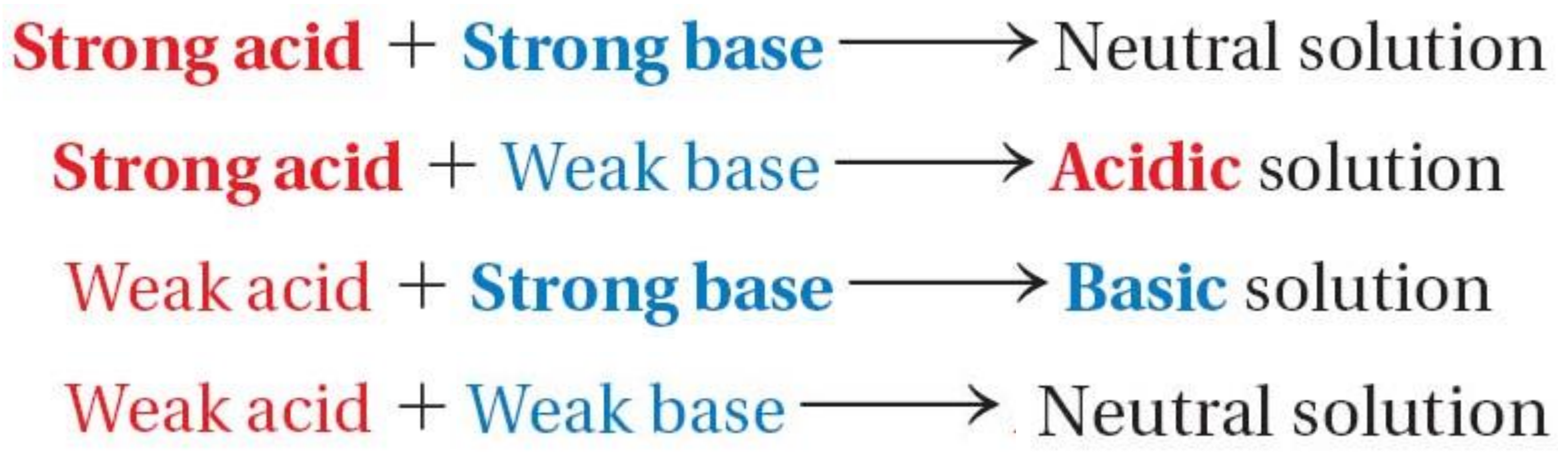
$$Kb = \frac{[NH_4^+][OH^-]}{[NH_3]} = 1.8 \times 10^{-5} \qquad Ka = \frac{[CN^-][H^+]}{[HCN]} = 6.2 \times 10^{-10}$$

4. Determine if more H^+ or OH^- is removed:

   More H^+ is removed, therefore BASIC

Salt Hydrolysis

To determine whether a salt solution is acidic or basic, remember the following rules:



Strong acid + **Strong base** ⟶ Neutral solution

Strong acid + Weak base ⟶ **Acidic** solution

Weak acid + **Strong base** ⟶ **Basic** solution

Weak acid + Weak base ⟶ Neutral solution

Making Buffer Solution (example)

A phosphate buffer with pH = 7.40 is prepared using KH_2PO_4 and K_2HPO_4.

(a) What is the molar ratio of $[HPO_4^{2-}]$ to $[H_2PO_4^-]$ in the buffered solution?

(b) If $[H_2PO_4^-] = 0.20\ M$, what is $[HPO_4^{2-}]$?

(c) How many grams of KH_2PO_4 and K_2HPO_4, respectively, are needed to make 500. mL of this solution? ($H_2PO_4^-$ has $K_a = 6.2 \times 10^{-8}$)

Solutions to Buffer

(a) Use Henderson-Hasselbalch equation:

- $\mathrm{pH} = \mathrm{p}K_a + \log([\mathrm{HPO_4}^{2-}]/[\mathrm{H_2PO_4}^-])$
- $7.40 = 7.21 + \log([\mathrm{HPO_4}^{2-}]/[\mathrm{H_2PO_4}^-])$
- $\log([\mathrm{HPO_4}^{2-}]/[\mathrm{H_2PO_4}^-]) = 7.40 – 7.21 = 0.19$
- $[\mathrm{HPO_4}^{2-}]/[\mathrm{H_2PO_4}^-] = 10^{0.19} = 1.55$

(b) If $[\mathrm{H_2PO_4}^-] = 0.20\ M$,

- $[\mathrm{HPO_4}^{2-}] = 1.55 \text{ x } 0.20\ M = 0.31\ M$

Solutions to Buffer

(c) Moles of KH_2PO_4 needed =

500. mL x (1 L/1000 mL) x 0.20 mol/L = 0.10 mole

- Moles of K_2HPO_4 needed =

  500. mL x (1 L/1000 mL) x 0.31 mol/L = 0.155 mole

- Grams of KH_2PO_4 needed =

  0.10 mol x (136.086 g/mol) = 14 g

- Grams of K_2HPO_4 needed =

  0.155 mol x (174.178 g/mol) = 27 g

Titration Problem

Strong Acid-Weak Base Titration:

- A 20.0 mL aliquot of 0.100 M NH_3 is titrated with 0.100 M HCl. (a) What is the pH of the solution before titration? (b) What is the pH of the solution after 10.0 mL of HCl has been added? (c) What is the pH of the solution at equivalent point (after 20.0 mL of 0.100 M HCl is added)?

  (K_b of $NH_3 = 1.8 \times 10^{-5}$)

Solution to Titration Problem

(a) Solving initial concentration of OH^- by approximation method:

$$[OH^-] = \sqrt{[NH_3] \text{ x } K_b} =$$

$$= \sqrt{(0.100 \text{ x } 1.8 \text{ x } 10^{-5})} = 1.3 \text{ x } 10^{-3} \text{ M}$$

$$[H_3O^+] = K_w/[OH^-] = (1.0 \text{ x } 10^{-14})/(1.3 \text{ x } 10^{-3})$$

$$= 7.5 \text{ x } 10^{-12} \text{ M}$$

$$\text{pH} = -\log(7.5 \text{ x } 10^{-12} \text{ M}) = 11.13$$

Solution to Titration Problem

(b) Concentration after 10.0 mL of HCl is added:

Reaction:	$NH_{3(aq)}$ +	$H_3O^+_{(aq)} \rightarrow$	$NH_4^+_{(aq)}$ +	H_2O
[**I**]$_{\text{before rxn}}$:	0.067 M	0.033 M	0.000	
[**C**]$_{\text{from rxn}}$:	-0.033 M	-0.033 M	+0.033 M	
[**E**]$_{\text{after rxn}}$:	0.034 M	0.000	0.033 M	

- $pH = pKa + \log([NH_3]_f/[NH_4^+]_f)$
- $= -\log(5.6 \times 10^{-10}) + \log(0.034/0.033)$
- $= 9.25 + (0.0) =$ **9.25**

Solution to Titration Problem

(c) Calculating pH at equivalent point:

Reaction:	$NH_{3(aq)}$ +	$H_3O^+(aq) \rightarrow$	$NH_4^+(aq)$ +	H_2O
$[\mathbf{I}]_{\text{before rxn}}$:	0.050 M	0.050 M	0.000	
$[\mathbf{C}]_{\text{from rxn}}$:	-0.050 M	-0.050 M	+0.050 M	
$[\mathbf{E}]_{\text{after rxn}}$:	0.000 M	0.000	0.050 M	

- At equivalent point, $[NH_4^+] = 0.050$ M
- K_a for $NH_4^+ = K_w/K_b = (1.0 \times 10^{-14})/(1.8 \times 10^{-5})$

  $= 5.6 \times 10^{-10}$

Solution to Titration Problem: example #3

(c) Calculating pH at equivalent point (continue):

$$K_a = x^2/(0.050 - x) = 5.6 \text{ x } 10^{-10}$$

- $x = [H_3O^+]$,
- Using approximation method,

$$[H_3O^+] = \sqrt{[NH_4^+] \text{ x } K_a}$$

$$= \sqrt{(0.050 \text{ x } (5.6 \text{ x } 10^{-10})} = 5.3 \text{ x } 10^{-6} \text{ M}$$

- pH = -log(5.3 x 10^{-6} = **5.28**

Solution to Titration Problem: example #3

(c) Calculating pH at equivalent point (continue):

$$K_a = x^2/(0.050 - x) = 5.6 \text{ x } 10^{-10}$$

- $x = [H_3O^+]$,
- Using approximation method,

$$[H_3O^+] = \sqrt{[NH_4^+] \text{ x } K_a}$$

$$= \sqrt{(0.050 \text{ x } (5.6 \text{ x } 10^{-10})} = 5.3 \text{ x } 10^{-6} \text{ M}$$

- pH = -log(5.3 x 10^{-6} = **5.28**

pH of Buffer Solution

- What is the pH of a buffer solution that is 0.45 *M* acetic acid ($HC_2H_3O_2$) and 0.85 M sodium acetate ($NaC_2H_3O_2$)? The K_a for acetic acid is 1.8×10^{-5}.

- Solution:
- $pH = pK_a + \log([C_2H_3O_2-]/[HC_2H_3O_2])$
- $pH = -\log(1.8 \times 10^{-5}) + \log(0.85/0.45)$
- $pH = 4.74 + 0.28 =$ **5.02**

Thank you for your attention!

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