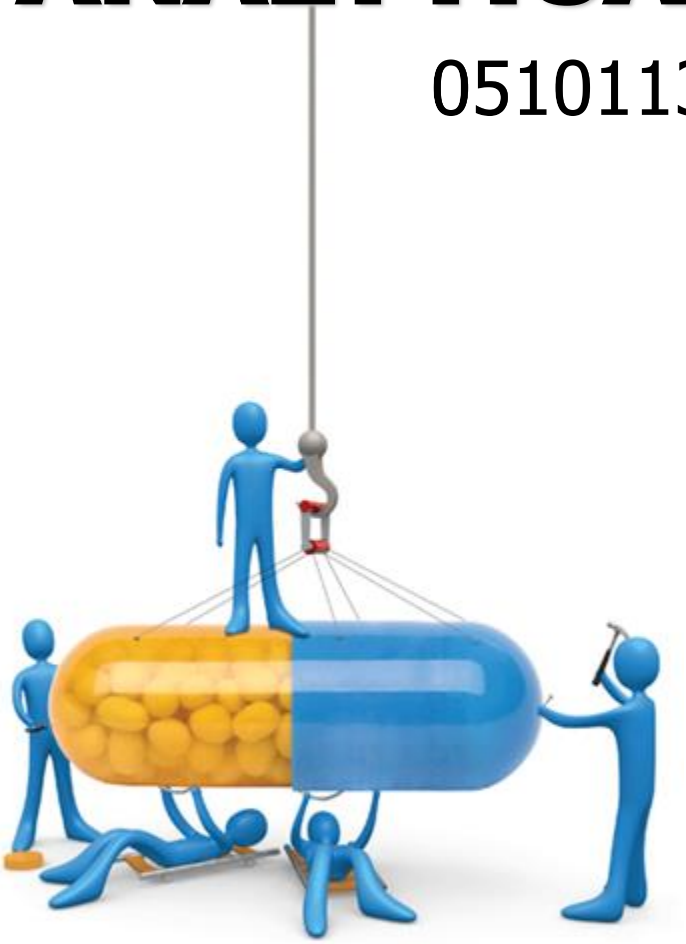
PHARMACEUTICAL ANALYTICAL CHEMISTRY

0510113/0510122/0521123

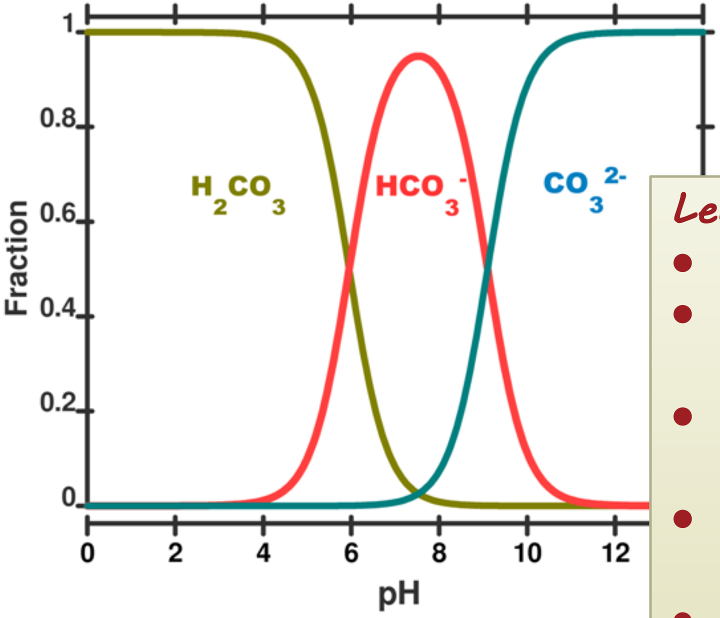
Acid–Base Equilibria

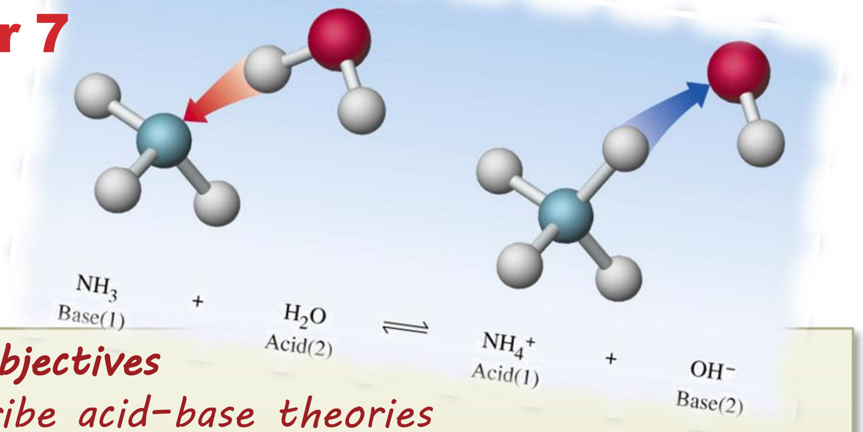


Dr. Ahmad Najjar
Philadelphia University
Faculty of Pharmacy
Department of Pharmaceutical Sciences
Second Semester, 2021/2022

Acid–Base Equilibria

Chapter 7





Learning Objectives

- *To describe acid-base theories*
- *To explain the acid-base equilibria in water solutions*
- *To predict the weak acids and bases in chemical equilibrium*
- *To write the ionization reaction for different weak acids and bases*
- *To specify the acidity or basicity of salts*
- *To prepare Buffers with certain pH*
- *To draw logarithmic concentration diagrams*

Acid–base Theories

- Arrhenius theory: donate H^+ or OH^-
- Theory of solvent systems: solvent cations and anions
- Brønsted-Lowry theory: taking and giving protons (conjugate pairs)
- Lewis theory: taking and giving electrons

In general: Acid is a substance that increases the concentration of H_3O^+ (hydronium ion). Conversely, a base decreases the concentration of H_3O^+ in aqueous solution.

$$\underset{\substack{\text{Hydrochloric acid}\\\text{(acid: proton donor)}}}{HCl(g)} + \underset{\substack{\text{Ammonia}\\\text{(base: proton acceptor)}}}{NH_3(g)} \rightleftharpoons \underset{\substack{\text{Ammonium chloride}\\\text{(salt)}}}{NH_4^+Cl^-(s)}$$

A more general definition of acids and bases given by Brønsted and Lowry is that an acid is a proton donor and a base is a proton acceptor.

Acid–base Equilibria In Water

- **Hydrochloric acid is a strong acid, and in water, its ionization is complete:**

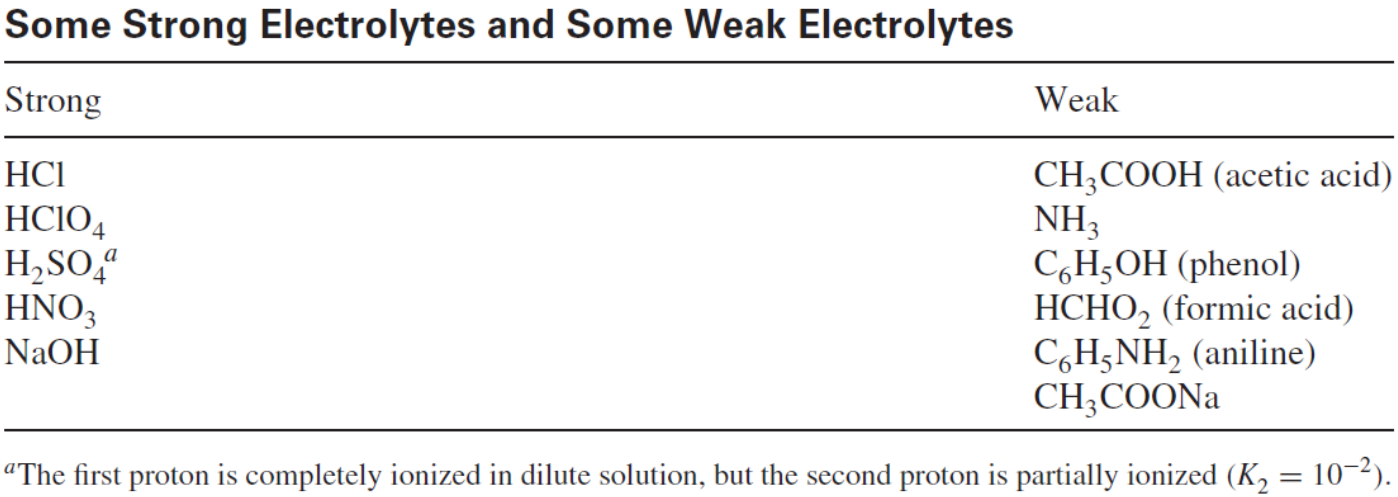
$$HCl + H_2O \rightarrow H_3O^+ + Cl^-$$

- **Acetic acid is a weak acid, which ionizes only partially in water (a few percent):**

$$HOAc + H_2O \rightleftharpoons H_3O^+ + OAc^-$$

We can write an equilibrium constant for this reaction:

$$K_a^\circ = \frac{a_{H_3O^+} \cdot a_{OAc^-}}{a_{HOAc} \cdot a_{H_2O}}$$



Some Strong Electrolytes and Some Weak Electrolytes

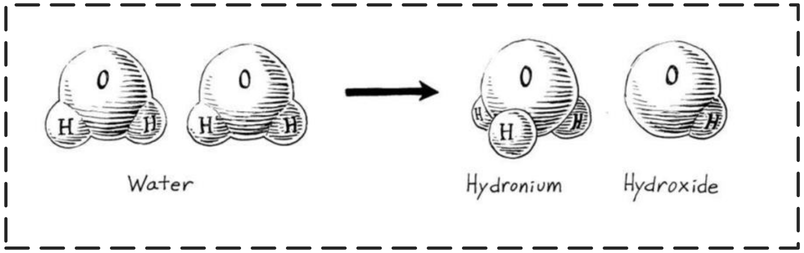
Strong	Weak
HCl	CH_3COOH (acetic acid)
$HClO_4$	NH_3
H_2SO_4[a]	C_6H_5OH (phenol)
HNO_3	$HCHO_2$ (formic acid)
NaOH	$C_6H_5NH_2$ (aniline)
	CH_3COONa

[a] The first proton is completely ionized in dilute solution, but the second proton is partially ionized ($K_2 = 10^{-2}$).

Acid–base Equilibria In Water

$$2H_2O \rightleftharpoons H_3O^+ + OH^-$$

$$K_w^\circ = \frac{a_{H_3O^+} \cdot a_{OH^-}}{a_{H_2O^2}} = a_{H_3O^+} \cdot a_{OH^-}$$



$$K_w = [H^+][OH^-]$$

K_w is the **molar equilibrium constants** and is exactly 1.00×10^{-14} at 24°C

Autoprotolysis is the self-ionization of a solvent to give a characteristic cation and anion, e.g.,

$$2CH_3OH \rightleftharpoons CH_3OH_2^+ + CH_3O^-$$

A 1.0×10^{-3} M solution of hydrochloric acid is prepared. What is the hydroxide ion concentration?

Solution $[OH^-] = 1.0 \times 10^{-11}$ M

The pH Scale

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

In general, **p**Anything = **−log** Anything, and it is a mathematical function to express the very low quantities.

- Calculate the pOH and the pH of a 5.0×10^{-2} *M* solution of NaOH at 25°C.

Solution $pOH = 1.30$ $pH = 12.70$

- The pH of a solution is 9.67. Calculate the hydrogen ion concentration in the solution.

Solution $[H^+] = 2.1 \times 10^{-10}\ M$

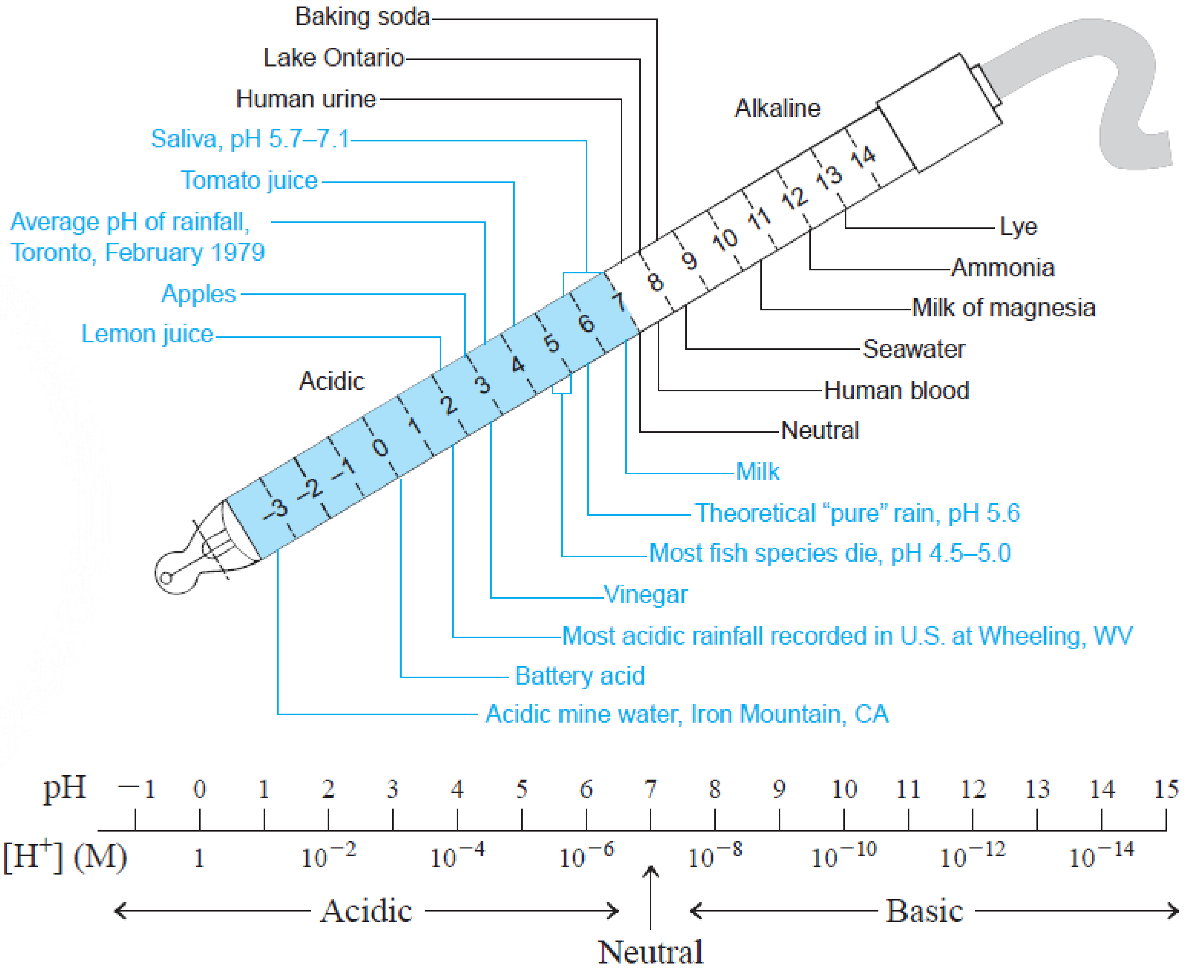
- Calculate the pH and pOH of a 1.0×10^{-7} *M* solution of HCl.

The calculation in this example is more academic than practical!!

Solution $pH = 6.79$ $pOH = 7.21$

$$[H^+] = 10^{-pH}$$

The pH Scale



Blood pH is 7.4 at 37°C. The pH of blood must be measured at body temperature to accurately reflect the status of blood buffers.

Weak Acids and Bases, K_a and K_b

The basicity constant K_b for ammonia is 1.75×10^{-5} at 25°C. (It is only coincidental that this is equal to K_a for acetic acid.) Calculate the pH and pOH for a 1.00×10^{-3} M solution of ammonia.

Solution $\text{pOH} = 3.88$ $\text{pH} = 10.12$

Important Notes

❑Salts of weak acids are bases (conjugate bases) and have $K_b = K_w / K_a$

❑ Also salt of weak base is a conjugate acid and has $K_a = K_w / K_b$

❑ For a diprotic acid–base pair like ethylenediamine, $NH_2C_2H_4NH_2$,

$$K_{b1} = K_w/K_{a2} \text{ and } K_{b2} = K_w/K_{a1}$$

Weak Acids and Bases, K_a and K_b

Calculate the pH of a 0.10 *M* solution of sodium acetate.

Solution

Write the equilibria

$$\text{NaOAc} \rightarrow \text{Na}^+ + \text{OAc}^- \text{(ionization)}$$

$$\text{OAc}^- + \text{H}_2\text{O} \rightleftharpoons \text{HOAc} + \text{OH}^- \text{(hydrolysis)}$$

Write the equilibrium constant

$$\frac{[\text{HOAc}][\text{OH}^-]}{[\text{OAc}^-]} = K_b = \frac{K_w}{K_a} = \frac{1.0 \times 10^{-14}}{1.75 \times 10^{-5}} = 5.7 \times 10^{-10}$$

Let x represent the concentration of HOAc and OH^- at equilibrium. Then, at equilibrium,

$$[\text{HOAc}] = [\text{OH}^-] = x$$

$$[\text{OAc}^-] = C_{\text{OAc}^-} - x = 0.10 - x$$

Since $C_{\text{OAc}^-} \gg K_b$, neglect x compared to C_{OAc^-}. Then,

$$\frac{(x)(x)}{0.10} = 5.7 \times 10^{-10}$$

$$x = \sqrt{5.7 \times 10^{-10} \times 0.10} = 7.6 \times 10^{-6}\ M$$

$$[\text{OH}^-] = 7.6 \times 10^{-6}\ M$$

$$[\text{H}^+] = \frac{1.0 \times 10^{-14}}{7.6 \times 10^{-6}} = 1.3 \times 10^{-9}\ M$$

$$\text{pH} = -\log 1.3 \times 10^{-9} = 9 - 0.11 = 8.89$$

Fractions of dissociating species at a given pH: α values — how much of each species?

$$H_3PO_4 \rightleftharpoons H^+ + H_2PO_4^- \qquad K_{a1} = 1.1 \times 10^{-2} = \frac{[H^+][H_2PO_4^-]}{[H_3PO_4]}$$

$$H_2PO_4^- \rightleftharpoons H^+ + HPO_4^{2-} \qquad K_{a2} = 7.5 \times 10^{-8} = \frac{[H^+][HPO_4^{2-}]}{[H_2PO_4^-]}$$

$$HPO_4^{2-} \rightleftharpoons H^+ + PO_4^{3-} \qquad K_{a3} = 4.8 \times 10^{-13} = \frac{[H^+][PO_4^{3-}]}{[HPO_4^{2-}]}$$

Total analytical concentration of phosphoric acid, C_{H3PO4}

$$C_{H_3PO_4} = [PO_4^{3-}] + [HPO_4^{2-}] + [H_2PO_4^-] + [H_3PO_4]$$

$$\alpha_0 = \frac{[H_3PO_4]}{C_{H_3PO_4}} \qquad \alpha_1 = \frac{[H_2PO_4^-]}{C_{H_3PO_4}} \qquad \alpha_2 = \frac{[HPO_4^{2-}]}{C_{H_3PO_4}} \qquad \alpha_3 = \frac{[PO_4^{3-}]}{C_{H_3PO_4}}$$

$$\alpha_0 + \alpha_1 + \alpha_2 + \alpha_3 = 1$$

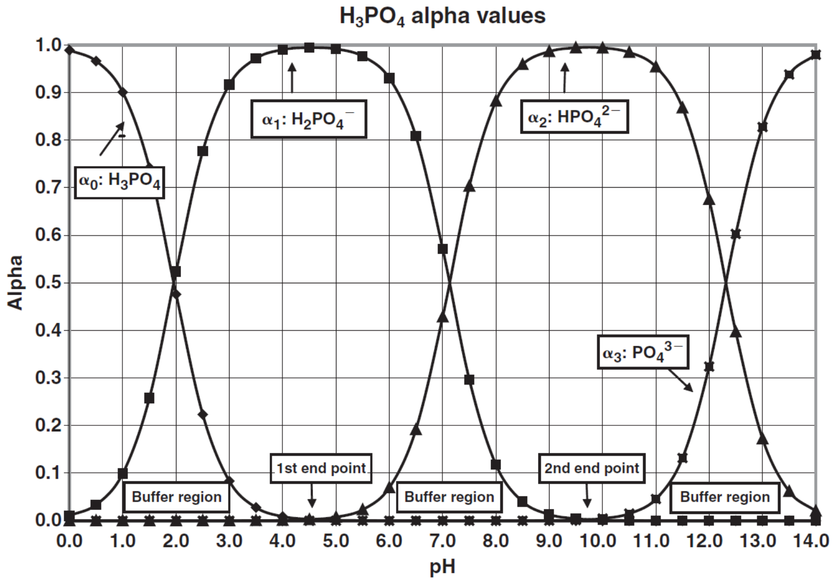
Fractions of dissociating species at a given pH: α values — how much of each species?

$$\alpha_0 = \frac{[\mathrm{H}^+]^3}{[\mathrm{H}^+]^3 + K_{a1}[\mathrm{H}^+]^2 + K_{a1}K_{a2}[\mathrm{H}^+] + K_{a1}K_{a2}K_{a3}}$$

$$\alpha_1 = \frac{K_{a1}[\mathrm{H}^+]^2}{[\mathrm{H}^+]^3 + K_{a1}[\mathrm{H}^+]^2 + K_{a1}K_{a2}[\mathrm{H}^+] + K_{a1}K_{a2}K_{a3}}$$

$$\alpha_2 = \frac{K_{a1}K_{a2}[\mathrm{H}^+]}{[\mathrm{H}^+]^3 + K_{a1}[\mathrm{H}^+]^2 + K_{a1}K_{a2}[\mathrm{H}^+] + K_{a1}K_{a2}K_{a3}}$$

$$\alpha_3 = \frac{K_{a1}K_{a2}K_{a3}}{[\mathrm{H}^+]^3 + K_{a1}[\mathrm{H}^+]^2 + K_{a1}K_{a2}[\mathrm{H}^+] + K_{a1}K_{a2}K_{a3}}$$



Calculate the equilibrium concentration of the different species in a 0.10 *M* phosphoric acid solution at pH 3.00 ($[\mathrm{H}^+] = 1.0 \times 10^{-3}$ *M*).

Buffers

Buffer is a solution that resists changes in pH when a small amount of acid or base is added or when the solution is diluted.

Buffer solution is usually made of (weak acid with its conjugated base) or (weak base with its conjugated acid).

Examples:

CH_3COO^- and CH_3COOH

HCN and CN^-

C_6H_5COOH and C_6H_5COONa

NH_3 and NH_4Cl

$(C_2H_5)_3N$ and $(C_2H_5)_3NH^+$

Buffers, Henderson-Hasselbalch equation

$$\mathrm{pH} = \mathrm{p}K_a + \log \frac{[\mathrm{A}^-]}{[\mathrm{HA}]}$$

$$\mathrm{pH} = \mathrm{p}K_a + \log \frac{[\text{conjugate base}]}{[\text{acid}]}$$

$$\mathrm{pH} = \mathrm{p}K_a + \log \frac{[\text{proton acceptor}]}{[\text{proton donor}]}$$

Calculate the pH of a buffer prepared by adding 10 mL of 0.10 *M* acetic acid to 20 mL of 0.10 *M* sodium acetate.

Solution $\mathrm{pH} = 5.06$

Buffers, Henderson-Hasselbalch equation

Calculate the pH of a solution prepared by adding 25 mL of 0.10 M sodium hydroxide to 30 mL of 0.20 M acetic acid.

The solution formed after mixing is containing weak acid and strong base, these will react forming water and salt. Note that the salt formed is containing the conjugate base of the weak acid.

$$HA + NaOH \rightleftharpoons NaA + H_2O$$

We have three possibilities,

1. The amount of weak acid=strong base, so the final solution is containing the salt only.
2. The amount of weak acid < strong base, so the final solution containing the formed salt and excess of the strong base.
3. The amount of weak acid > strong base, so the final solution containing the formed salt and excess of the weak acid.

- If possibility #1 is happened, then the acidity of the solution depends on the concentration of formed salt (which is not neutral!).
- If possibility #2 is happened, then the acidity of the solution depends on the concentration of excess strong base, while the contribution of the formed salt is negligible.
- If possibility #3 is happened, then the acidity of the solution depends on the concentration of excess weak acid and concentration of the formed salt (**We have BUFFER!**).

Buffers, Henderson-Hasselbalch equation

- Calculate the volume of concentrated ammonia and the weight of ammonium chloride you would have to take to prepare 100 mL of a buffer at pH 10.00 if the final concentration of salt is to be 0.200 *M*.

- The pH of blood is 7.40. What is the ratio of $[HPO_4^{2-}]/[H_2PO_4^{-}]$ in the blood (assume 25°C)?

Buffers, Henderson-Hasselbalch equation

- What weights of NaH_2PO_4 and Na_2HPO_4 would be required to prepare 1 L of a buffer solution of pH 7.45 that has an ionic strength of 0.100?

$pK_{a2} = 7.12$

- How many grams ammonium chloride and how many milliliters 3.0 *M* sodium hydroxide should be added to 200 mL water and diluted to 500 mL to prepare a buffer of pH 9.50 with a salt concentration of 0.10 *M*?

Buffers, Henderson-Hasselbalch equation

Calculating the pH of a buffer when strong acid or base is added

$$HA \rightleftharpoons H^+ + A^-$$

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

when strong acid is added to the buffer:

⇩ $A^- + H^+ \rightarrow HA$ ⇧

$$pH = pK_a + \log \frac{(\text{moles } A^- - \text{moles } H^+ \text{added})}{(\text{moles } HA + \text{moles } H^+ \text{added})}$$

when strong base is added to the buffer:

⇩ $HA + OH^- \rightarrow H_2O + A^-$ ⇧

$$pH = pKa + \log \frac{(\text{moles } A^- + \text{moles } OH^- \text{added})}{(\text{moles } HA - \text{moles } OH^- \text{added})}$$

Buffers, Henderson-Hasselbalch equation

Example:

Find the pH of a solution of 0.02964M HA (Ka =8.4x10^{-9}) plus 0.1026M NaA?

Answer: pH=8.614

Example:

For 1L of the above solution, calculate the pH after addition of 12mL of 1.0M HCl.

HCl is an acid, it will react with the basic content (A$^-$) of the buffer solution and forming more of HA, so it will affect the acid/base content of the solution, NaA will decrease and HA will increase. Amount of change is equal to the moles of HCl (0.012Lx1.0M)

$$\mathrm{pH} = \mathrm{pK_a} + \mathrm{Log}\frac{\mathrm{mol.A^-}}{\mathrm{mol.HA}} = 8.075 + \mathrm{Log}\frac{(1x0.1026) - (0.012x1.0)}{(1x0.02964) + (0.012x1.0)}$$

$$\Rightarrow \mathrm{pH} = 8.075 + \mathrm{Log}\frac{0.0906}{0.04164} = 8.075 + \mathrm{Log}(2.1757) = 8.075 + 0.3376 = 8.413$$

Buffers, Henderson-Hasselbalch equation

Added base will react with the acid content of the buffer, this will decrease the concentration (or moles) of acid and increase the concentration (or moles) of base content of the buffer.

Example:

For 1L of the previous buffer solution, calculate the pH after addition of 12mL of 1.0M NaOH.

NaOH will affect the acid/base content of the solution, HA will decrease and NaA will increase. Amount of change is equal to the moles of NaOH (0.012Lx1.0M)

$$\mathrm{pH} = \mathrm{p}K_a + \mathrm{Log}\frac{\mathrm{mol.A^-}}{\mathrm{mol.HA}} = 8.075 + \mathrm{Log}\frac{(1\mathrm{x}0.1026) + (0.012\mathrm{x}1.0)}{(1\mathrm{x}0.02964) - (0.012\mathrm{x}1.0)}$$

$$\Rightarrow \mathrm{pH} = 8.075 + \mathrm{Log}\frac{0.1146}{0.01764} = 8.075 + \mathrm{Log}(6.4966) = 8.075 + 0.8127 = 8.888$$

Acid-Base Titration

Titration is used to determine the concentration of analyte volumetrically, depending on the stoichiometric ratio between analyte and titrant.

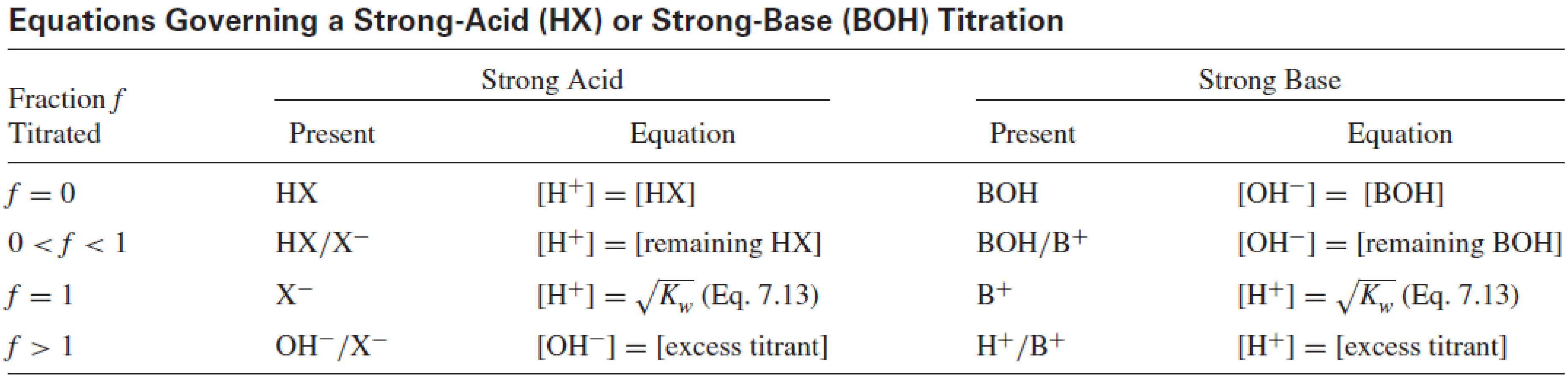
Examples:

$1HCl + 1NaOH \rightarrow NaCl + H_2O$ (ratio 1:1)

$1H_2SO_4 + 2NaOH \rightarrow Na_2SO_4 + 2H_2O$ (ratio 1:2)

$2CH_3COOH + 1Ca(OH)_2 \rightarrow (CH_3COO)_2Ca + 2H_2O$ (ratio 2:1)

$1H_2SO_4 + 1Ca(OH)_2 \rightarrow CaSO_4 + 2H_2O$ (ratio 1:1)



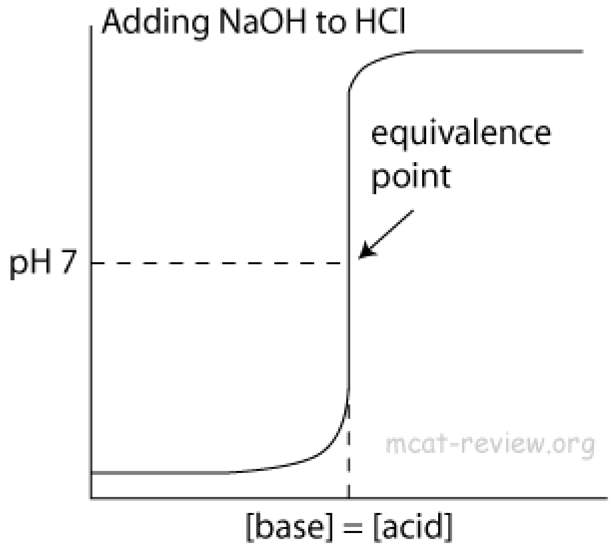
Equations Governing a Strong-Acid (HX) or Strong-Base (BOH) Titration

Fraction f Titrated	Strong Acid		Strong Base	
	Present	Equation	Present	Equation
$f = 0$	HX	$[H^+] = [HX]$	BOH	$[OH^-] = [BOH]$
$0 < f < 1$	HX/X^-	$[H^+] = [\text{remaining HX}]$	BOH/B^+	$[OH^-] = [\text{remaining BOH}]$
$f = 1$	X^-	$[H^+] = \sqrt{K_w}$ (Eq. 7.13)	B^+	$[H^+] = \sqrt{K_w}$ (Eq. 7.13)
$f > 1$	OH^-/X^-	$[OH^-] = [\text{excess titrant}]$	H^+/B^+	$[H^+] = [\text{excess titrant}]$

pH curve

During the acid-base titration, the pH will change. If the analyte is acid, pH will be low and start increasing through titrant addition (base).

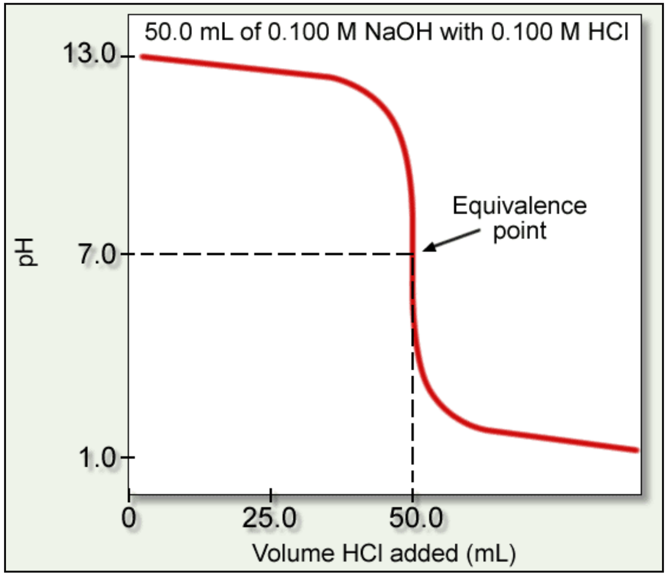
Example: (titration 50 mL of 0.10M HCl with 0.10M NaOH)



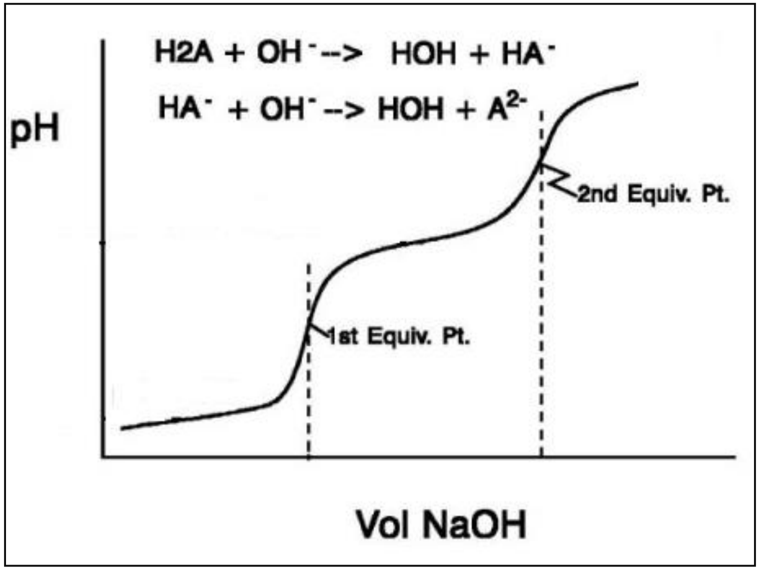
Volume of NaOH added	pH	
0	1.00	Acid excess
10	1.18	
20	1.37	
30	1.60	
40	1.95	
50	7.00 equiv. point	$[H][OH]=K_w$
60	11.96	Base excess
70	12.22	
80	12.36	

pH curve

Titration strong base with strong acid



Titration diprotic acid with strong base



pH curve

Question

50 mL of 0.10M HCl was titrated with 0.24M NaOH. Calculate the pH after addition of 0, 5, 20, 25 and 30 mL of NaOH.

Question

17.4mL of 0.1102M HCl was needed to titrate 25.0mL of NaOH with unknown concentration. Calculate the concentration of NaOH.

References

- Analytical Chemistry, 7th ed., by G. D. Christian, P. K. Dasgupta, and K. A. Schug, Wiley & Sons Publisher, 2014.
- Fundamentals of Analytical Chemistry, 9th ed., by D. A. Skoog, D. M. West, F. J. Holler, and S. R. Crouch, Cengage Learning Publisher, 2013.
- Problem Solving in Analytical Chemistry, by T. P. Hadjiioannou, G. D. Christian, C. E. Efstathiou, and D. Nikolelis, Pergamon Publisher, 1988.
- Calculations in Analytical Chemistry, by Q. Fernando and M. D. Ryan, Harcourt Brace Jovanovich Publisher, 1982.
- Solving Problems in Analytical Chemistry, by S. Brewer, Wiley Publisher, 1980
- Concepts and Calculations in Analytical Chemistry. A Spreadsheet Approach, by H. Freiser, CRC Press, 1992.
- https://www.khanacademy.org/science/chemistry/acid-and-bases-topic
- YouTube https://www.youtube.com/watch?v=GfMveIZXq9A&list=PL9fwy3NUQKwbICh15S1cAUCJm1SxVhq19 أ.د. منذر عبداللطيف
- YouTube https://www.youtube.com/channel/UCh-WhopyMvQaMv0VF_h5O-Q/videos
  أ.د. ابراهيم زامل الزامل

Thank You!