3

BIOCHEMISTRY
faculty of medicine - JU2017

ketone
starch
isomers
lipid
protein
amine
carbohydrates
carbo

◉ Sheet   ○ Slides

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Last time we talked about the Ionization of water and then started talking about k_w (ion product for water) which is a constant and it equals the concentration of protons (hydronium ions) times the concentration of hydroxyl ions.

$$K_{eq} \, (55.5 \text{ M}) = [H^{\oplus}] \, [OH^{\ominus}]$$

$$K_w = [H^{\oplus}] \, [OH^{\ominus}] = 1.0 \times 10^{-14} \, M^2$$

In this lecture we will continue talking about pH.

Remember that:

pH	$[H^{\oplus}]$ (M)	$[OH^{\ominus}]$ (M)
0	1	10^{-14}
1	10^{-1}	10^{-13}
2	10^{-2}	10^{-12}
3	10^{-3}	10^{-11}
4	10^{-4}	10^{-10}
5	10^{-5}	10^{-9}
6	10^{-6}	10^{-8}
7	10^{-7}	10^{-7}
8	10^{-8}	10^{-6}
9	10^{-9}	10^{-5}
10	10^{-10}	10^{-4}
11	10^{-11}	10^{-3}
12	10^{-12}	10^{-2}
13	10^{-13}	10^{-1}

TABLE 2.3   Relation of $[H^{\oplus}]$ and $[OH^{\ominus}]$ to pH

- If we have 10^{-14} concentration of hydroxyl ions in a solution, then the concentration of protons in that solution MUST be 1.
- If we have 10^{-7} concentration of hydroxyl ions in a solution, then the concentration of protons in that solution MUST be 10^{-7} and so on.

Why? Because when we multiply them together they must be equal to 1.0×10^{-14} according to the equation.

We said that concentration of protons is an indication for the acidity of the solution

So, if we have a high concentration of protons that means the solution is highly acidic.

We mentioned before that scientists are lazy, it is a struggle to say long numbers like 5.65×10^{-6} *every time*, so they took the minus logarithm to the base of 10 for the concentration of protons. This gives you the **pH** which is a small nice and simple number. It is easier to say and is a better indication for the concentration of protons.

So, pH is equal to minus logarithm to the base of 10 for the concentration of protons.

$$pH = \log_{10}(1/[H^+]) = -\log_{10}[H^+]$$

We should also know that the higher the concentration of protons, the lower the pH.

So, pH $\propto$ 1/[H⁺]   acidity $\propto$ [H⁺]   Ka $\propto$ acidity   pKa $\propto$ 1/acidity   pH $\propto$ 1/acidity

As above, the lower the pH, the higher the acidity. Therefore, any solution that has:

pH<7 is acidic   pH>7 is basic

The pH scale is a logarithmic scale. One pH unit difference implies a 10-fold difference in [H+].
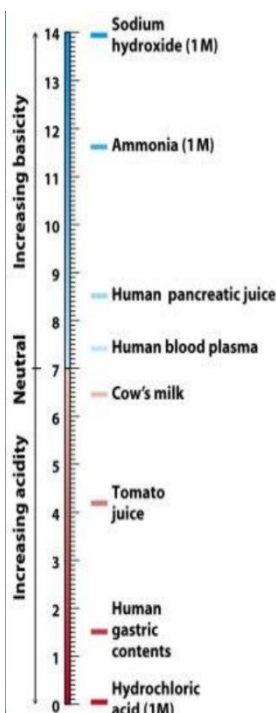
Example: lemon juice concentration of [H⁺] at pH 2.0 is higher 100 times than orange juice at pH 4.0.

- the picture on the right shows some natural products with their respective pH values. Human blood plasma is 7.4, and milk is about 6.5 which is a bit acidic. However, notice that Hydrochloric acid (1M) and human gastric contents must be very acidic in order to perform their function.

(I asked the doctor, you don't have to memorize this table ☺).

The following 3 examples are related to Ka (review sheet 2).

The pH of Various Common Fluids	
Fluid	**pH**
Household lye	13.6
Bleach	12.6
Household ammonia	11.4
Milk of magnesia	10.3
Baking soda	8.4
Seawater	8.0
Pancreatic fluid	7.8–8.0
Blood plasma	7.4
Intracellular fluids	
Liver	6.9
Muscle	6.1
Saliva	6.6
Urine	5–8
Boric acid	5.0
Beer	4.5
Orange juice	4.3
Grapefruit juice	3.2
Vinegar	2.9
Soft drinks	2.8
Lemon juice	2.3
Gastric juice	1.2–3.0
Battery acid	0.35

Example (1): Find the Ka of a 0.04 M weak acid HA whose [H⁺] is 1 x 10⁻⁴?

$$ HA \rightleftharpoons H^+ + A^- $$
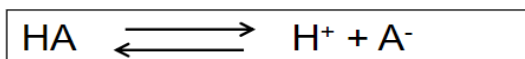
From the question: [HA]=0.04 M   [H⁺] is 1 x 10⁻⁴M   Ka=??

Since the concentration of H⁺ = the concentration of the conjugate base (A⁻)

Ka= [Products] / [Reactants] = [A⁻] [H⁺]/ [HA] = [H⁺]²/ [HA] = 10⁻⁴ x10⁻⁴ / 0.04 = 2.5 x 10⁻⁷

Note: since we have studied the pH now, we can make the question a little bit harder by requiring an additional step:

Find the Ka of a 0.04 M weak acid HA whose pH is 4.0?

In that case we just use [H+] =10⁻ᵖᴴ equation to get the concentration of the proton then the rest of the answer will be the same as above.

Example (2): What is the [H⁺] of a 0.05 M Ba(OH)₂?

$$ Ba(OH)_2 \rightleftharpoons Ba + 2OH^- $$
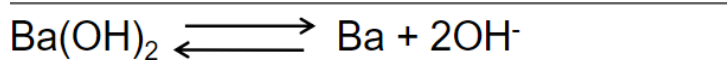
Barium Hydroxide doesn't give a proton; it gives two hydroxyl ions so if we calculate the concentration of the hydroxyl ions, we can use kw = [H⁺] [OH⁻] equation to calculate the concentration of the proton.

[OH-] = 2x 0.05 = 0.10 M = 1 x 10⁻¹

[H+] = 1x 10⁻¹³

$$pH + pOH = 14$$
$$pH = - \log [H^+]$$
$$pOH = - \log [OH^-]$$

Remember

Example (3): The [H$^+$] of a 0.03 M weak base solution is 1 x 10^{-10}M. Calculate pKb?

Note: careful when you read the question to decide each number for what

$$B + H_2O \rightleftharpoons BH^+ + OH^-$$
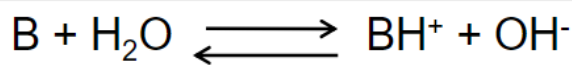
You should already know from previous lecture that Kw=(Ka x Kb)= 10^{-14}

[OH-] = 10^{-4}

Kb= (10^{-4} x 10^{-4}) / 0.03 = 3.33 x 10^{-7} M

 pKb= -log Kb= 6.48.

Measuring acidity in aqueous solutions- pH scale

pH = - log [H$_3$O$^+$]

[H$_3$O$^+$] = 10^{-pH}

Example 1: [H$_3$O$^+$] is 10^{-12} M what is the pH

pH = - log [10^{-12}] = 12

Example 2: Orange juice has a pH of 4 what is the concentration of [H$^+$]

[H$_3$O$^+$] = 10^{-4} M

Enzyme activity and pH

■ **FIGURE 2.12 pH versus enzymatic activity.** Pepsin, trypsin, and lysozyme all have steep pH optimum curves. Pepsin has maximum activity under very acidic conditions, as would be expected for a digestive enzyme that is found in the stomach. Lysozyme has its maximum activity near pH 5, while trypsin is most active near pH 6. found in lysosomes and tears
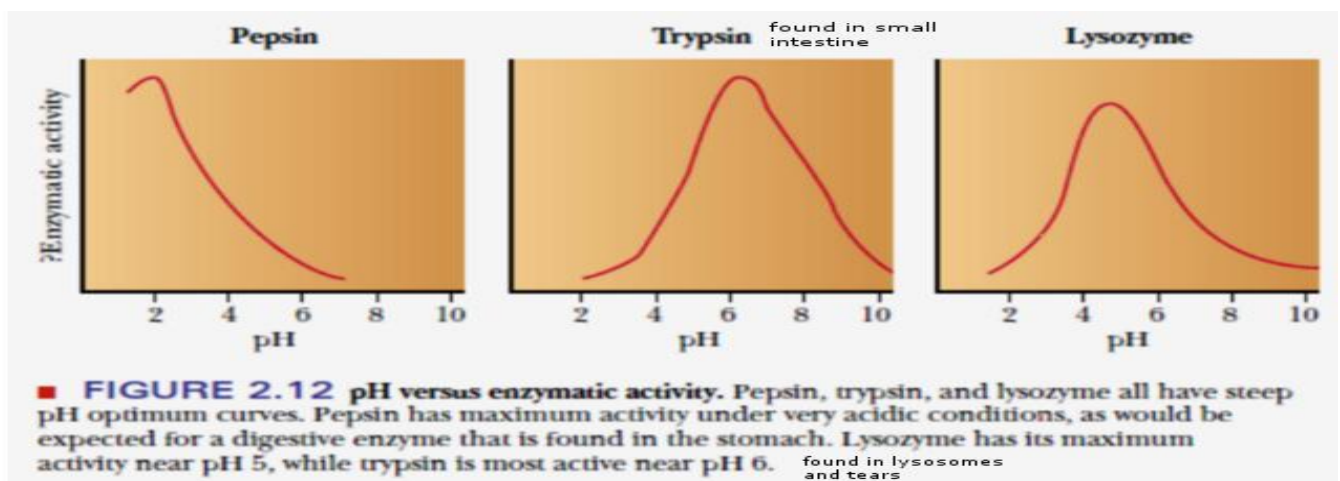
a) 0.01 M HCl?

Since it is a strong acid, $[H^+] = [HCl] = 0.01$ M

$$HCl \rightarrow H^+ + Cl^-$$

$$Ka = \frac{[H^+]\ [Cl^-]}{[HCl]}$$

$$[H^+] = [acid]$$

pH = -log (1.0×10^{-2}) = 2
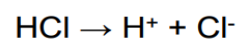
b) 0.01 N H₂SO₄?

Recall that for an acid solution, n is the number of H+ provided by a **formula unit** of acid. Similarly, for a base solution n is the number of OH- a base can donate.

N= n x M (where n is an integer)

But you don't need to find the M because it's a diprotic acid and it will give you 2 protons (hence a definition for N)

[H+] = 0.01 M

pH = -log (0.01) = 2.0

Note: if the question gave you 0.01 M instead of N, you will need to multiply by 2.

c) 0.01 N NaOH?

N= n x M

0.01 = 1 x M

$[OH^-]$ = 0.01 $\longrightarrow$ kw = $[H^+]$ $[OH^-]$ | $[H^+]$ = 1.0 x 10^{-12} | pH = -log (1.0 x 10^{-12}) = 12

OR

$\longrightarrow$ pOH = -log (0.01) = 2 | 14 = pH + pOH | 14 = pH + 2 | pH = 12

d) 1 x 10^{-11} M HCl? (this is a tricky one)

We have two sources of protons in that solution.

First source is Hydrochloric acid which gives concentration of 1 x 10^{-11} M of protons.

Second source is water which gives concentration of 1 x 10^{-7} M of protons.

So, the number of protons for this solution is equal to

10^{-7} (from the water) + 10^{-11} (from the hydrochloric acid) = 1.0001 x 10^{-7}.

10^{-11} is too small relative to the 10^{-7}, too small that makes it negligible, the pH is 6.999956573 which is very close to 7.

Note: in previous exercises we neglected proton concentration from water while here we did the opposite, why?

e) 0.1 M of acetic acid (CH3COOH)? Remember Ka

Notice that acetic acid (Vinegar) is a weak acid which means not all the 0.1 M dissociates to products (i.e. proton and conjugate base of the acid [acetate]).

$$CH_3COOH + H_2O \longleftrightarrow CH_3COO^- + H_3O^+$$
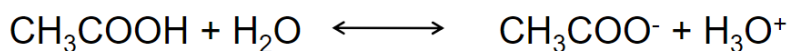
$$Ka = \frac{[H+][CH_3COO^-]}{[CH_3COOH]}$$

All we need is the concentration of the protons, we know:

1- the equilibrium constant of acetic acid, Ka = [Products] / [Reactants]
2- 2-the concentration of the proton is equal to the concentration of the conjugate base
3- we have Ka and concentration of the acetic acid and from the equation we can calculate x.

$Ka = 1.74 \times 10^{-5}$

$1.74 \times 10^{-5} = x^2/0.1$

$x^2 = 1.74 \times 10^{-6}$

$x = 1.32 \times 10^{-3} M = [H^+]$

$pH = -\log(1.32 \times 10^{-3}) = 2.88$

Note: we neglected the amount of acid the will change to the H^+ and conjugate base to simplify the equation (it is too small compared to 0.1M).

11:38

Additional questions from Dr.Diala slides

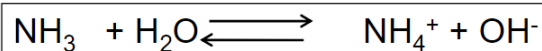**Q1)** What is the pH of a solution, if the hydrogen ion concentration is 3.2×10^{-4} mol/L?

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

$= -\log(3.2 \times 10^{-4})$

$= -\log(3.2) - \log(10^{-4})$

$= -0.5 + 4.0 = 3.5$

Q2) The Kb for ammonia is 1.8×10^{-5} M. What is the pH of 1×10^{-2} M of ammonia?

$$NH_3 + H_2O \rightleftharpoons NH_4^+ + OH^-$$

$Kb = [NH_4^+][OH^-]/[NH_3]$

$1.8 \times 10^{-5} = [OH^-]^2/0.01$

$[OH^-] = 4.24 \times 10^{-4}$ M

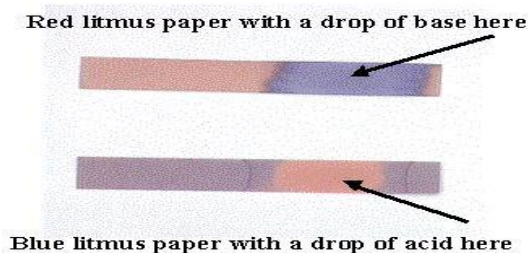$pOH = -\log 4.24 \times 10^{-4} = 3.37$

$pH = 14 - 3.37 = 10.63$

Determination of pH

In laboratories, we can calculate pH in several ways, such as:

1) Acid-base indicator

a) Litmus paper (least accurate)

Red litmus paper with a drop of base here

Blue litmus paper with a drop of acid here

b) Universal indicator: (consists of natural products, it gives a certain pH value) this is a more accurate way than litmus paper since it doesn't just display whether substance is acidic or basic (red or blue). It gives a wide range of colors which can indicate whether the pH is for example 5.0 or 5.5 as accurate as possible.
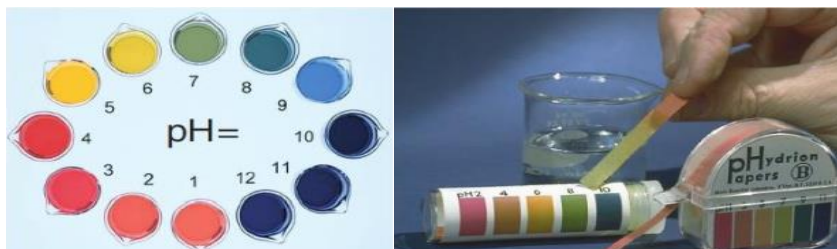
2) An electronic pH meter (most accurate): an electrical way of measuring concentration of protons.

Henderson-Hasselbalch Equation

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

Dr.Mamoun only mentioned the equation and said to make sure to memorize it, while Dr.Diala derived it in her slides as following:

1-The dissociation of a weak acid is as follows:

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

2-The acid dissociation constant is as follows:

$$K_a = \frac{[H^+][A^-]}{[HA]}$$

3- Rearranging this expression in terms of the parameter of interest [H+] gives the following:

$$[H^+] = \frac{K_a[HA]}{[A^-]}$$

4- Take the log of both sides:

$$\log[H^+] = \log K_a + \log \frac{[HA]}{[A^-]}$$

5- Change the signs , remember $pK_a = -\log K_a$:

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

or

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

Ka equals to concentration of the products over concentration of the reactants. The products include H+, so we can clearly see that there is a relationship between Ka and concentration of protons.

Henderson-Hasselbalch Equation it shows the relationship between the pH and the pKa (relationship between concentration of protons with equilibrium constant).

This equation fits weak acids and weak bases (we are talking about weak acids here). It **DOESN'T** fit strong acids because they dissociate **completely**.

In an acidic solution (weak acid), we have 3 components:

a) the acid.   b) the protons.   c) the conjugate base.

So, we take in consideration **all these 3 components**. Therefore, the pH of the solution doesn't only depend on how much protons we have in the solution, but also depends on amounts of conjugate base and acid.

→What does this equation mean? How can we define/explain it?

pH in a weak acid solution is equal to its pKa when the concentration of the conjugate base is equal to the concentration of the acid. How? When they are equal, we get the logarithm of 1 which is zero and therefore pH is equal to pKa.

Summary: we can define pH in many ways:

pH is equal to the pKa of the weak acid when the concentration of the conjugate base is equal to the concentration of the acid.

 OR

pH is equal to the pKa of the weak acid when half the weak acid has dissociated.

OR

pH is equal to the pKa of the weak acid when 50% of it is in the conjugate base form and 50% is in the acidic form.

What happens when (dr.Diala slides)…..

pH = pKa

Substance protonation and deprotonation are in equilibrium**.**

pH < pKa

this means the logarithm is negative, concentration of the acid is more than the concentration of the conjugate base which means most of the acid is protonated (didn't give its protons).

H+ on, substance is protonated.

pH > pKa

this means the logarithm is positive, concentration of conjugate base is more than the concentration of the acid which means the acid is deprotonated (gave most of its protons)

H+ off, substance is deprotonated.

**pH= pKa -1**

Base/acid = 0.1

pH= pKa + 1

Base/acid = 10

Table 2.7	
pH Values and Base/Acid Ratios for Buffers	
If the pH equals	**The ratio of base form/acid form equals**
$pK_a - 3$	1/1000
$pK_a - 2$	1/100
$pK_a - 1$	1/10
pK_a	1/1
$pK_a + 1$	10/1
$pK_a + 2$	100/1
$pK_a + 3$	1000/1

Le Châtelier's principle

When there is disturbance in equilibrium, the direction of arrows will change until we reach equilibrium again. For example, if we add A to the solution, equilibrium is disturbed, we no longer have the equilibrium constant. So, in order to decrease the concentration of A, the reaction will shift towards the right. A will decrease, 2B will decrease too, C and D will increase. So, we will reach Keq again **BUT** <u>with different values for A, B, C and D.</u>

When more reactants, A and/or B is added, the equilibrium shifts to reduce A and B by producing more C and D.

$$A + 2B \rightleftharpoons C + D$$
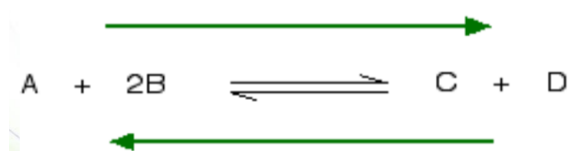
When more products (C and/or D is added), the equilibrium shifts to reduce C and D by producing more A and B.

19:05

A comparison of the change in pH (Water Vs. Acetic acid)

Let's see how the pH changes when we add a base to both solutions,

1- 0.01 mol of a strong base is added to 1 L of pure water=> pH changes drastically from 7 to 12 (100000 folds change!)
2- 0.01 mol of a strong base is added to 1L acetic acid/acetate ion buffer which has $[CH_3COOH]$ = 0.1M and $[CH_3COO^-]$ =0.1M => pH may change from 4.7 to 4.85.
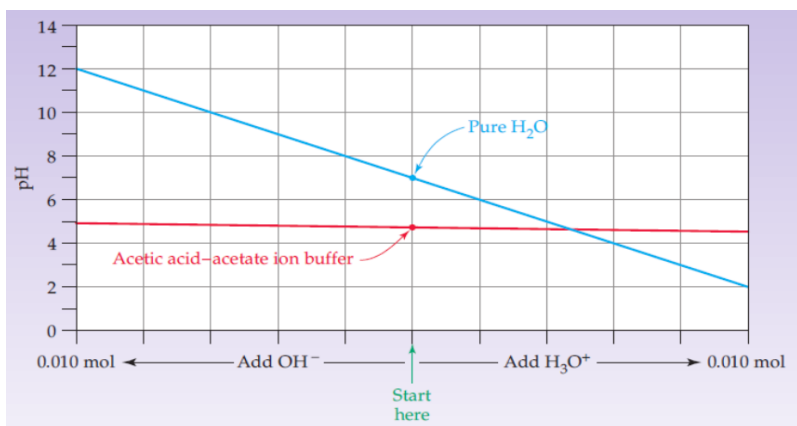
This picture shows what happens when we add .01 M of $[H^+]/[OH^-]$ to both pure water and acetic acid-acetate buffer, notice that the pH varies between:

2-12 in pure water

4.85-4.68 in the buffer

When we add HCl, why does the pH change in water drastically while in acetic acid-acetate ion buffer its change is minimal? In water: pH was 7, it drops drastically (for example) to 4. But acetic acid is different, it has an equilibrium constant which must be preserved. Acetic acid is in equilibrium with its products; the conjugate base and the proton. If we increase concentration of protons, the reaction will shift toward the formation of acetic acid, therefore the concentration of both protons and acetate ion (conjugate base) will decrease and the concentration of acetic acid will increase. This all happened in order to preserve the equilibrium constant.

So, if we add hydrochloric acid to a solution with acetic acid, equilibrium will change in a way in which the extra protons that were added will be removed(absorbed). Acetate will react with the proton and therefore we will have more formation of acetic acid in order to reach the equilibrium constant, so the number of protons that we added will decrease. So, the drop in pH **WON'T** be from 7 to 4 like the previous example, it will be from 7 to 6.8. Why? Because the number of protons that we added have been absorbed/removed immediately. This property is present in weak acids, why not in strong acids?

Summary: if we add hydrochloric acid to a solution of <u>water only,</u> the drop in pH will be high but if we add it to a solution of a weak acid (acetic acid) the pH (Y axis) change will be minimal and this is how a **buffer** works.

Buffer: a solution of a weak acid or it can be a solution of a weak base (as long as it doesn't dissociate completely),or you can define it as a combination of a weak acid and its conjugate base. Its function is to resist changes in pH when concentration of protons is changed (whether you add or remove protons in a solution).

The word buffer is used in many concepts of our life like Buffer zone: a zone that experiences no movement, no reactions. Similarly, the buffer resists any change in the pH whether by adding or removing protons from a solution.

The components of a buffer are the weak acid and its conjugate base, or the weak base and its conjugate acid.

Buffers are solutions that resist changes in pH by changing reaction equilibrium. They are composed of mixtures of a weak acid and a roughly equal concentration of its conjugate base.

Acetic acid, its conjugate base is the salt of acetate (Sodium acetate)

When we say acet*ic* acid, we mean the molecule that has the proton(it can donate it).

Acid	Conjugate base
CH_3COOH	CH_3COONa ($NaCH_3COO$)
H_3PO_4	NaH_2PO_4
$H_2PO_4^-$ (or NaH_2PO_4)	Na_2HPO_4
H_2CO_3	$NaHCO_3$

when we say acet*ate*, we talk about the conjugate base, the molecule with no proton, which is $CH3COO^-$ or salt of acetate (Sodium acetate or Potassium acetate)

Phosphoric acid, its conjugate base is mono sodium dihydrogen phosphate, the word phosphate indicates it's a conjugate base NaH_2PO_4.

Mono sodium dihydrogen phosphate, that is the conjugate base we just talked about, but it has a proton, is it an acid too? **Yes**, because it can donate a proton therefore it can act as an acid as well. What is its conjugate base? Disodium monohydrogen phosphate Na_2HPO_4.

If a compound can act as an acid and a base, how do we distinguish between the acid and the conjugate base in this scenario?

The acid is the molecule that has the extra proton, the conjugate base is the molecule that has one less proton.

So, if we have $H_2PO_4^-$ and HPO_4^{2-}, what is the acid and what's the base?

Note: We add sodium to remove the charges. Since its triprotic, it has the ability to donate one more proton.

We can't say conjugate base of phosphoric acid is disodium monohydrogen because it is less by 2 protons which is not acceptable. The difference in protons between them MUST be one.

If **Disodium monohydrogen phosphate** is an acid as well what is its conjugate base?

It's Trisodium phosphate which is not an acid anymore because it doesn't have a proton.

Carbonic acid: its conjugate base is hydrogen carbonate.

Another example is lactic acid, its conjugate base is lactate which is also an acid, it can act as a buffer.

30:00

- The conjugate base can exist alone (charged/ionized) but it exists mostly as a salt especially when you buy it from different companies, you buy it as a salt (the cation of the salt won't affect the reaction pathway since its inert).
- The weak acid and its conjugate base can resist changes in pH by manipulating/changing the equilibrium. So, it maintains the concentration of protons but at the same time other concentrations are reduced/increased.

(When you add A, you will change B, C and D at the same time.)

▪ Where do the absorbed protons go? The conjugate base reacts with them. So, we maintain the concentration of protons in the solution, but we change the concentration of acid and conjugate base. Don't forget that pH depends on all 3 of these components, so the pH will change, but this change is minimal.

Titration

Titration is an experiment in which measured amounts of a base are added to a measured amount of an acid while following up the changes in pH using a pH meter.
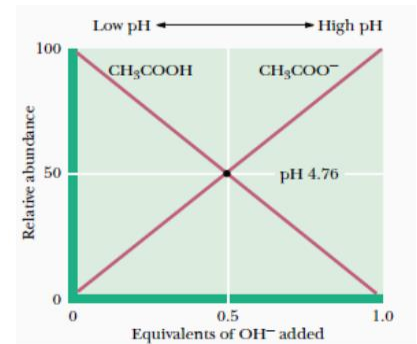

Note: **This is not what happens in reality, the following picture is more precise depiction of what happens in weak acids:**

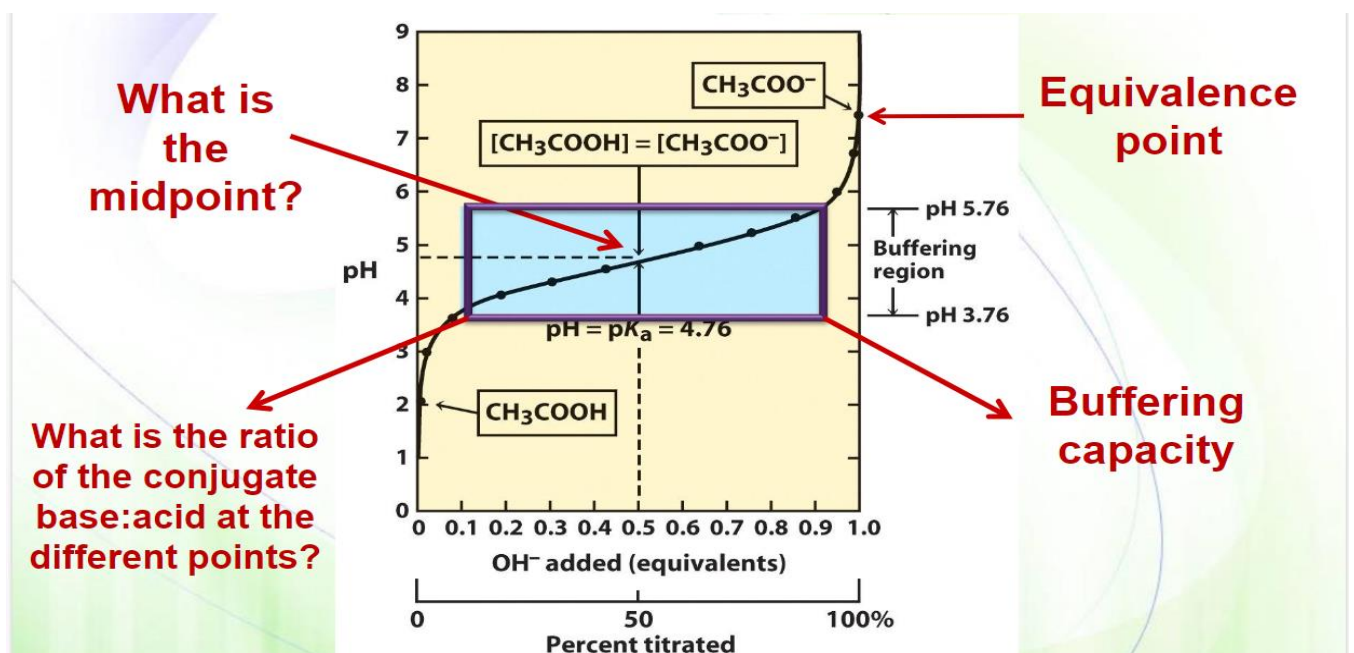$$CH_3COOH + OH^- \rightleftharpoons CH_3COO^- + H_2O$$
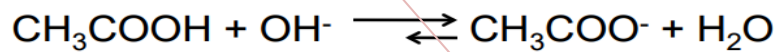
In the region small pH changes upon addition of acid or base, the acid/base ratio varies within a narrow range (10:1 at one end and 1:10 at the other end).

The point in the titration at which the acid is exactly neutralized is called the **equivalence point.**
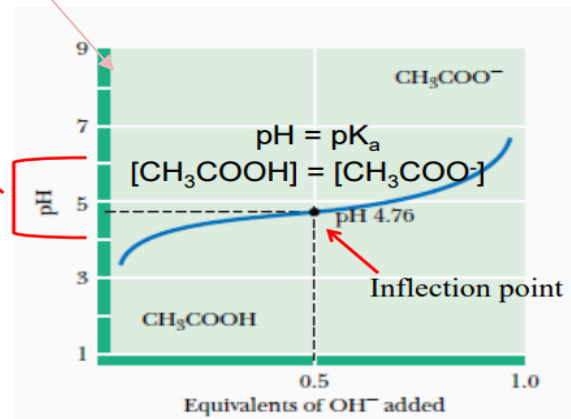
Titration curve of buffer

Titration curve of the buffer (acetate buffer)

Note: try to study each of the following points on the curve (لا تشلفق)

→**Y axis** represents pH from 0 to 9.

→**On x axis** we have equivalents which is a better measure of acids and bases, but what does that mean? if we use molarity in x axis we will need to make sure that the number of protons donated by an acid is the same that is absorbed by the base. In other words, you can't titrate 1 mole of H_2SO_4 (diprotic) by 1 mole of NaOH.

→ By using equivalents, we can neglect the amounts of protons that can be donated/accepted. We only consider equivalents just like how we say 1 mole of sodium equals 1 mole of potassium.

→When pH is low, the concentration of hydroxyl is approximately 0. CH3COOH concentration is high because the concentration of protons is high, that's why it is found in large amounts (you can make sense of that using Le Châtelier's principle or using Henderson-Hasselbalch Equation).

- When we add hydroxyl ions, it will react with protons to produce water. Equilibrium is disturbed. So, the acetic acid dissociates to compensate for the number of lost protons, therefore resisting the change in pH.
- When we add OH^- ions, the pH will change rapidly at first, until it reaches a point where the change is slow, why? Because there are high amounts of H^+ for the OH^- to react with at first, then the acid starts to dissociate to make up for the H^+ loss.
- The change will stay slow until the curve reaches another point, there, when we add hydroxyl ions pH increases rapidly again. Why isn't there any more resistance? The concentration of the weak acid can't keep up with the high amount of concentration of hydroxyl ions that are being added! So, we no longer have CH_3COOH (very small amount that can't resist the change, so the pH rises rapidly). →The region where we have resistance in the change in pH is called **buffering capacity.**
- Why doesn't the resistance happen from the start? Because we have a very small amount of hydroxyl ions. (We don't have any conjugate base to do any buffering dr.Diala said that it changed rapidly at first because there are high amounts of H^+, SO THE OH^- can react with them easily).
- We have a point in the middle: the midpoint which represents the pKa. In other words, this is the point where the pH is equal to the pKa, the point when the concentration of the conjugate base is equal to the concentration of the acid.
- When we add more OH^-, we consume more concentration of protons, so pH increases. However, there are other reasons for the increase of pH (which are the concentration of the acid and the conjugate base).
- The ratio between the acid and the conjugate base in the midpoint is 1:1.

Notice that the buffering capacity **is constant in all buffers**.

its range is **pKa -1 to pKa+1**. If the pka was 4.8, the range is 3.8 to 5.8, the zone where it can act as buffer and resist change in pH.

If someone has a solution and wants its pH to be 4.5

we recommend him to use acetate buffer even though its pH is 4.8 not 4.5 because buffering capacity is 3.8 to 5.8 which includes that pH value, so it can still act as a buffer at 4.5.

BUT if we want the pH in my solution to be 3.5

we can't use acetate buffer because that pH value is not within its buffering capacity.

It is also recommended to use a buffer which has pKa close to the desired pH:

Can we use acetate buffer to keep a pH value of 4 unchanged? Yes, because it is within the buffering capacity of acetate, **BUT** we are very close to the edge of the range, any acid that is added to the solution might lead to a decrease in pH value because we are close to the end of the range of the buffering capacity.

45:00

Additional questions for better understanding
- **Question (1): What is the midpoint?**

Midpoint is the point where pH = pKa because the concentration of the acid and concentration of its conjugate base are equal because 50% is neutralized and 50% is not.

Note: notice first part of the curve is concave downwards then it becomes concave upwards that's why the midpoint is also called inflection point.

- **Question (2): Why at the start and end points the change was quick, where as in the middle it was slow?**

Near start point: we have high concentration of acid and we added few amounts of the base. So, the base is facing many molecules of the acid, so pH goes up rapidly.

(we don't have any conjugate base to do any buffering)

in the middle: by time, acid molecules number decreases but in relative amounts to base molecules. Ratio becomes almost 1:1 which is why in the middle we have the plateau.

Near end point: base concentration is high, there isn't enough acid left to give protons which can react with OH^-. OH^- will accumulate and make the solution basic.

(there is no more acidic form that can donate the proton to make any resistance in pH)

- **Question (3): What is the ratio of conjugate base: acid at different points?**

Midpoint: 1:1

Near start point of the buffer zone: 1:10

Near end point of the buffer zone: 10:1

(will be explained in next lecture)

- **Note:** Titration curve that we explained previously was for a monoprotic acid, polyprotic acid titration curve will also be explained in next lecture.
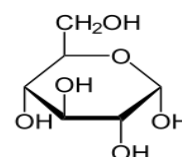
Check your understanding:
These questions may include ideas from previous sheets

1- **Which of the following is a true measure of acid's strength?**
   a- pKa   b- pH   c- conjugate base concentration   d- a+b

2- **How many water molecules can hydrogen bond directly to the molecules of glucose?**
   a- 11   b- 17   c- 15   d- 21

3- **Aspirin is an acid with pka of 3.5; its structure includes a carboxyl group. To be absorbed into blood stream, it must pass through the membrane lining the stomach and the small intestine, electrically neutral molecules can pass through a membrane more easily than can charged molecules. Would you expect more aspirin to be absorbed in the stomach (pH=1) or in small intestine (pH=6)? why?**
   a- Stomach, because carboxyl group would be deprotonated
   b- Stomach, because carboxyl group would be protonated
   c- Small intestine, because carboxyl group would be deprotonated
   d- Small intestine, because carboxyl group would be protonated

4- **How many folds (by what factor?) does the [H$^+$] increase when you change the pH from 7 to 6.5?**
   a- It doesn't increase it decreases   b- 3 folds   c- 5 folds   d- 0.5 folds

5- **Which of the following is a good definition for buffer capacity?**
   a- The amount of an acid or a base that can be added to a buffer before experiencing a sharp change in pH.
   b- The point at the titration curve at which the added acid or base equals the amount of buffer originally present.
   c- An experiment in which acid or base is added stepwise to a solution of a compound and the pH is measured as a function of the added substance
   d- The point where the ratio of acid to conjugate base equals 1.

Can you tell me the concept of each definition? ↑

6- **What is the [CH$_3$COO$^-$]/[CH$_3$COOH] ratio in an acetate buffer at pH 5.00 (kb of acetate = 5.568x10^{-10}).**
   a- 5.56x10^{-5} b- 2.1   c- 0.57   d- 1.7