بسم الله الرحمن الرحيم

قال سبحانه: وَمَنْ يَتَّقِ اللَّهَ يَجْعَلْ لَهُ مَخْرَجًا * وَيَرْزُقْهُ مِنْ حَيْثُ لاَ يَحْتَسِبُ وَمَنْ يَتَوَكَّلْ عَلَى اللَّهِ فَهُوَ حَسْبُهُ إِنَّ اللَّهَ بَالِغُ أَمْرِهِ قَدْ جَعَلَ اللَّهُ لِكُلِّ شَيْءٍ قَدْرًا {الطلاق:2-3}.

pH and buffers

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Kw

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

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

- Kw is called the ion product for water ز

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

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}

What is pH?

$$\frac{10^{-4}}{10^{-3}} = 10^{(-1)}$$

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

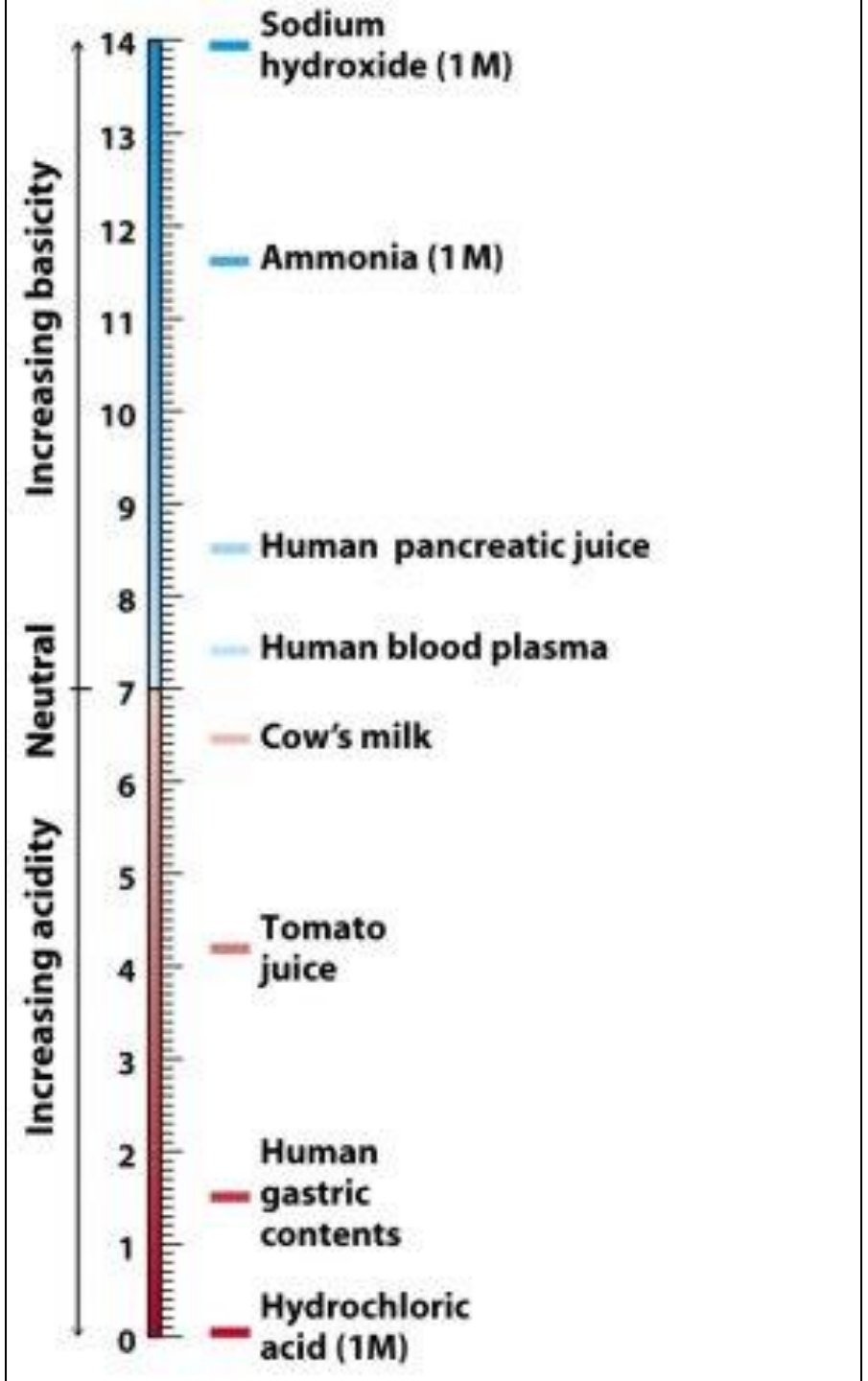
$pH_1 = 3$, $pH_2 = 4$

$\frac{[H^+]_1}{[H^+]_2} = \frac{10^{-3}}{10^{-4}} = 10$ times more acidic

$pH_1 = 3$ $pH_2 = 3.5$, $\frac{[H^+]_1}{[H^+]_2}$??

$\Rightarrow \frac{10^{-3}}{10^{-3.5}} = 10^{.5} \simeq 3$

PH = -Log[H+] ‖ [H+] = 10^{-PH}



Example 1:

Find the K_a of a 0.04 M weak acid HA whose $[H^+]$ is 1×10^{-4}?

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

$$K_a = [A^-][H^+] / [HA] = [H^+]^2 / [HA] = 10^{-4} \times 10^{-4} / 0.04 = 2.5 \times 10^{-7}$$

Example 2:

What is the $[H^+]$ of a 0.05 M $Ba(OH)_2$?

$$Ba(OH)_2 \longrightarrow Ba + 2OH^-$$

$$[OH^-] = 2 \times 0.05 = 0.10\ M = 1 \times 10^{-1}$$

$$[H^+] = 1 \times 10^{-13}$$

Example 3:

The $[H^+]$ of a 0.03 M weak base solution is 1×10^{-10} M. Calculate pKb?

$$B + H_2O \rightleftharpoons BH^+ + OH^-$$

$[OH^-] = 10^{-4}$

$K_b = (10^{-4} \times 10^{-4}) / 0.03 = 3.33 \times 10^{-7}$ M

$pK_b = -\log K_b = 6.48$

حل أسئلة من الكتاب

Exercises

- What is the pH of
  - 0.01 M HCl?
  - 0.01 N H_2SO_4?
  - 0.01 N NaOH?
  - 1 x 10^{-11} M HCl? (this is a tricky one)
  - 0.1 M of acetic acid (CH_3COOH)? Remember Ka = 1.76×10^{-5}

normality → تركيز مكافئ جاهز H^+

$pH = -\log 10^{-2} \rightarrow 2$

$2 \times [H_2SO_4] = 0.01 \rightarrow [H^+]$

$[H_2SO_4] = 0.005$

molarity → تركيز الحمض أو أحماض

0.01 M of H_2SO_4

$[H^+] = 0.02$

$pH = -\log 2 \times 10^{-2}$

$= 2 - \log 2$

$[OH^-] = 10^{-2} \rightarrow [H^+] = 10^{-12} \rightarrow pH = 12$

Solution: $Ka = \frac{[H^+][CH_3COO^-]}{[CH_3COOH]}$

$1.76 \times 10^{-5} = \frac{x^2}{0.1}$

$x^2 = 1.76 \times 10^{-6}$

$[H^+] = x = \sqrt{\quad}$

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

$[OH^-]$ or

← حتى يعطيل N بكون تركيز $[H^+]$ الحمض أو القاعدة

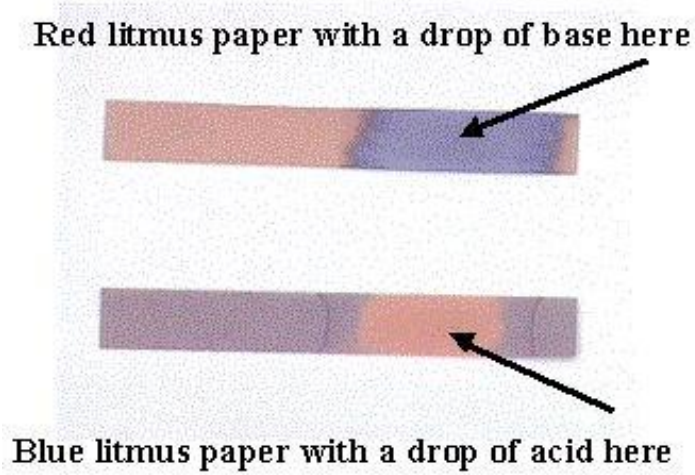
← " " M " "

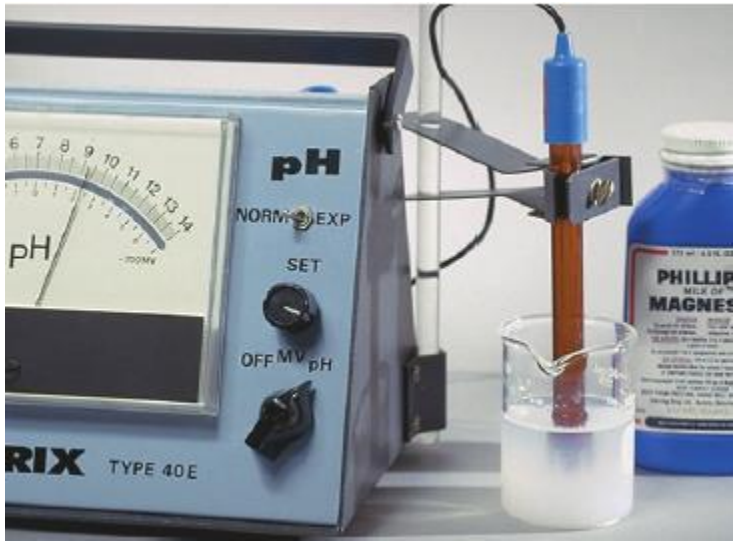
مستحيل pH=12 ∴ 0.1 HCl

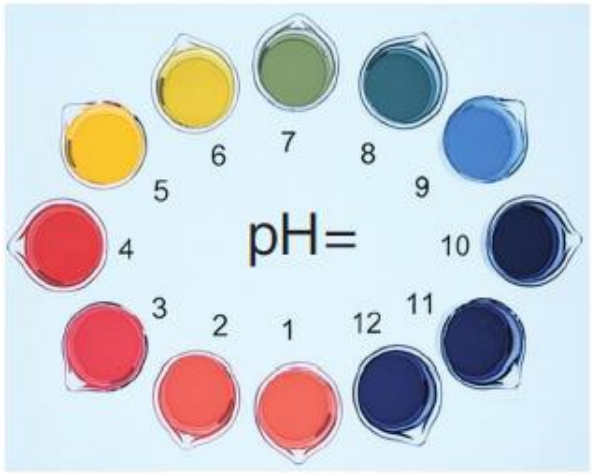
$[H^+] = [H^+]_{HCl} + [H^+]_{H_2O} = 10^{-11} + 10^{-7} \simeq 10^{-7} \rightarrow pH = 7$

Determination of pH

- Acid-base indicator
  - Litmus paper (least accurate)
  - Universal indicator
- An electronic pH meter (most accurate)







(a)

(b)

Henderson-Hasselbalch equation

$$K_a = \frac{H^+ \times A^-}{HA} \rightarrow H^+ = \frac{K_a \times HA}{A^-}$$

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

deprotonated

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

Protonated

$$5 = 5 + ??$$

$$?! = 0 \rightarrow \frac{A^-}{HA} = 1 \rightarrow$$ %50 protonated, %50 deprotonated

$$7 = 5 + x$$

$$x = 2 \rightarrow \frac{A^-}{HA} = 100 \rightarrow$$ for 100 [A⁻] → we have 1 [HA]

pH	=	acidity of a buffer solution
pKa	=	negative logarithm of Ka
Ka	=	acid disassociation constant
[HA]	=	concentration of an acid
[A-]	=	concentration of conjugate base

pKa is the pH where 50% of an acid is dissociated into its conjugate base.

pKa = pH

pKa is the pH where the concentration of the acid is equal to that of the conjugate base.

A comparison of the change in pH (water vs. acetic acid)

Base
0.01 mol
V = 1L
$CH_3COOH \rightarrow 0.1$
$CH_3COO^- \rightarrow 0.1$
↳ pKa = pH

- 0.010 mol of base are added to 1.0 L of pure water and to 1.0 L of an acetate buffer composed of 0.10 M acetic acid and 0.10 M acetate ion buffer, the pH of the water varies between 12 and 2, while the pH of the buffer varies only between 4.85 and 4.68.

↳ pH = pKa
[HA] = [A⁻]

$H_2O + H_2O \rightleftharpoons H_3O^+ + OH^- \rightarrow pH = 7$

$[OH^-] = 10^{-7} \not\rightarrow 10^{-2} \simeq 10^{-2}$

$\rightarrow pH = 12$

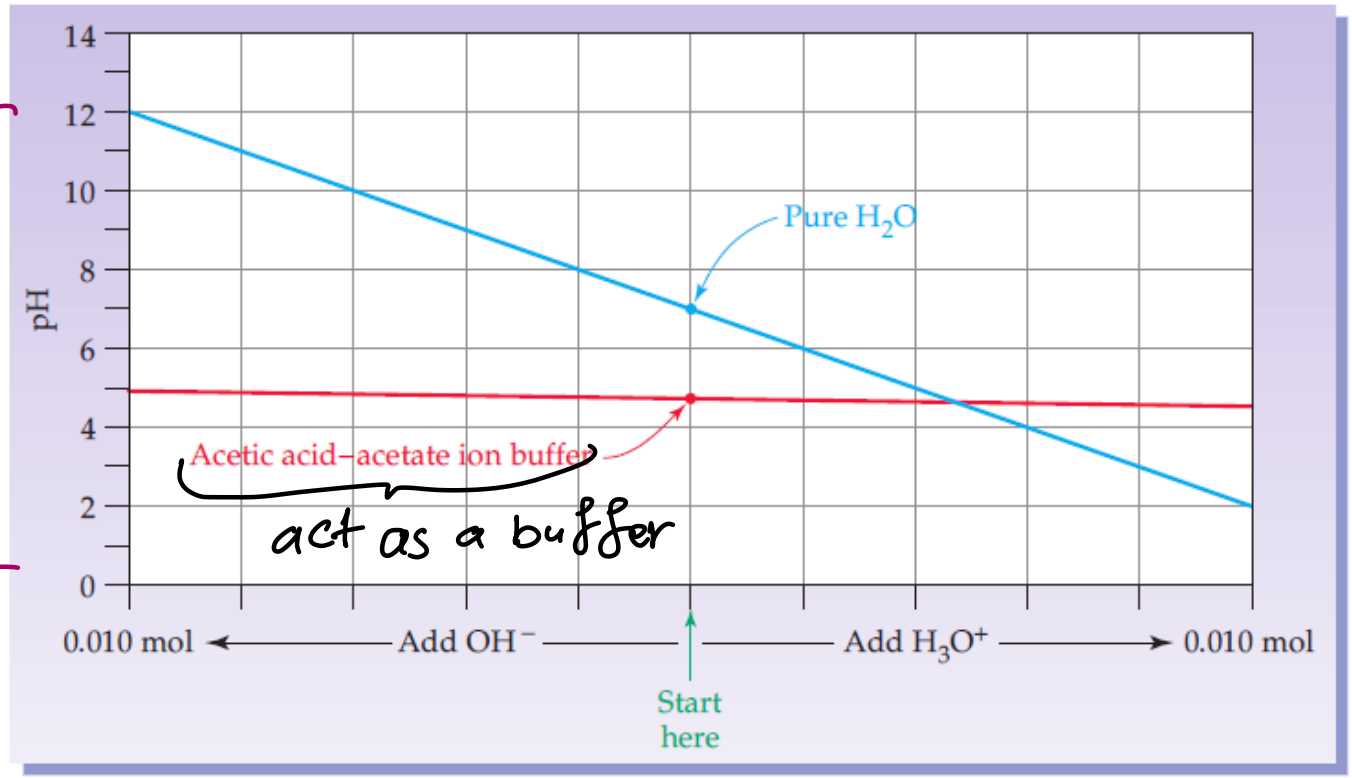
$CH_3CO_2H(aq) + H_2O(l) \rightleftharpoons H_3O^+(aq) + CH_3CO_2^-(aq)$

Weak acid
Acetic acid

Conjugate base
($NaCH_3CO_2$)
Salt of the weak acid

pH depending in 3 elements according to Henderson equation :-

1 [H+] → pKa 2 [HA] 3 [A⁻]



What is a buffer?

- Buffers are solutions that resist changes in pH by changing reaction equilibrium.
- They are usually composed of mixtures of a weak acid and an equal concentration of its conjugate base (*salt*).

2 conditions: -

1

strong not buffers

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	$NaHCO_3$

• Mechanism of Buffer Function

• Adding Acid :

protonation ←

$$HA \rightleftharpoons A^- + \boxed{H^+ \uparrow \text{(increase)}}$$

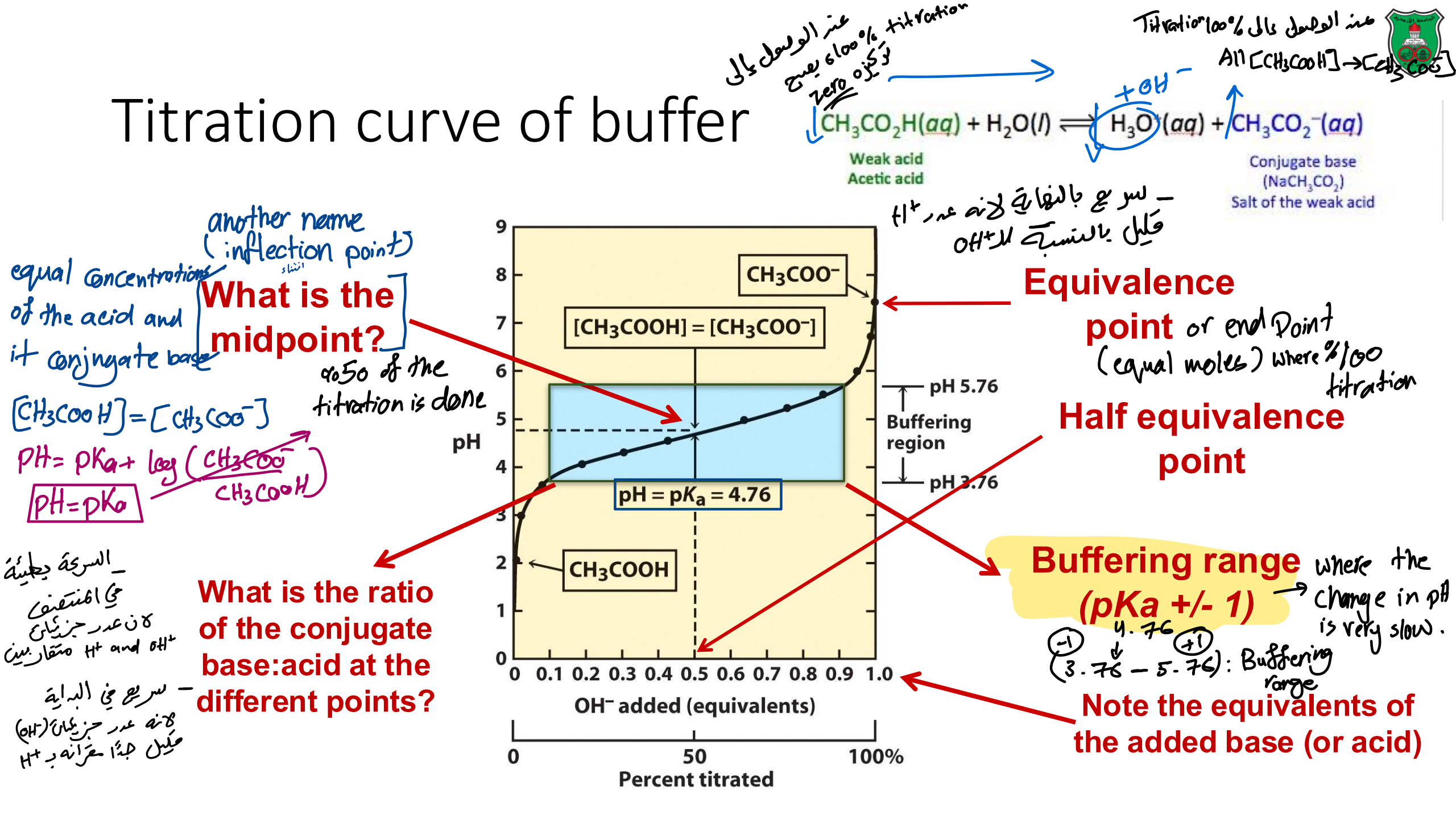
• Adding Base :—

$$HA \rightleftharpoons A^- + \boxed{H^+ \downarrow \text{(decrease)}}$$

deprotonation →

کمی (-)

Titration curve of buffer



another name (inflection point)

equal concentrations of the acid and it conjugate base

What is the midpoint?

$[CH_3COOH] = [CH_3COO^-]$

$pH = pK_a + \log \frac{[CH_3COO^-]}{[CH_3COOH]}$

$pH = pK_a$

50% of the titration is done

What is the ratio of the conjugate base:acid at the different points?

Equivalence point or end Point (equal moles) where 100% titration

Half equivalence point

Buffering range (pKa +/- 1) → where the change in pH is very slow.

4.76 (−1) (+1) (3.76 − 5.76): Buffering range

Note the equivalents of the added base (or acid)

*) Mid Point [inflection Point]

→ %50 of the Titration is done.

→ equal concentration of the acid and it conjugate base $[CH_3COOH] = [CH_3COO^-]$

↳ apply to Henderson equation :—

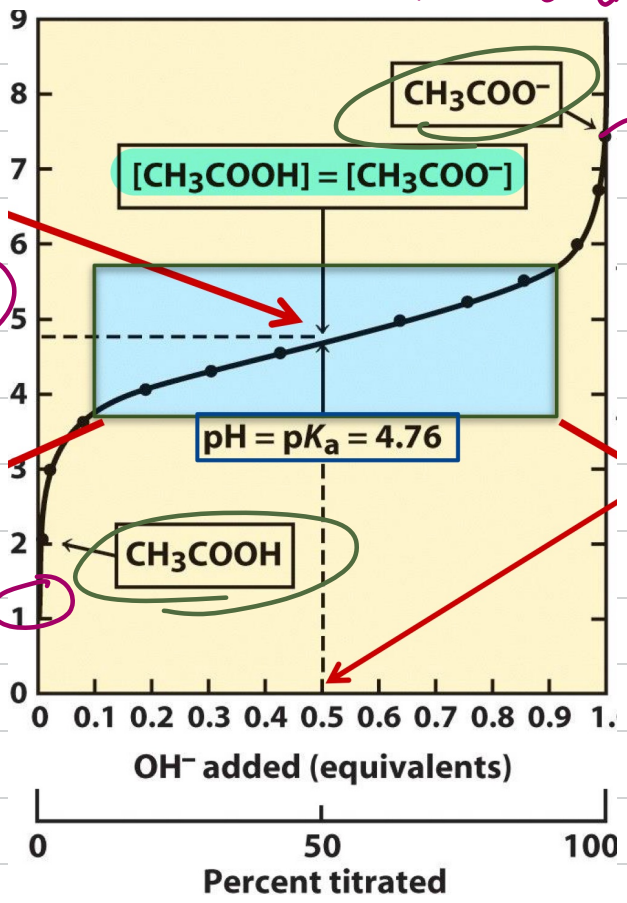
$$pH = pKa + \log \frac{[CH_3COO^-]}{[CH_3COOH]}$$

$$\boxed{pH = pKa}$$

Mid Point (inflection Point)

There is $[CH_3COOH]$ and little $[CH_3COO^-]$ from the dissociate of the acid.

Equivalence Point or end Point



* Equivalence Point or End Point :—

done → %100

→ equal moles where titration is ↑, $[CH_3COOH] = 0$ → finish the whole acid

* Buffering Range :—

$(pKa - 1, pKa + 1) \rightarrow (4.76 - 1, 4.76 + 1) \rightarrow (3.76, 5.76)$

Q) What is the predominant form in 1 ?!

CH_3COOH

Q) What is the predominant form in 2 ?!

equal amounts of $[CH_3COOH]$ and $[CH_3COO^-]$

Q) " " " " " " 3 ?!

CH_3COO^-

Q) " " " 4 ?!

CH_3COOH

Q) " " " 5 ?!

CH_3COO^-

Q) What is the ratio of the conjugate base to the acid at 1

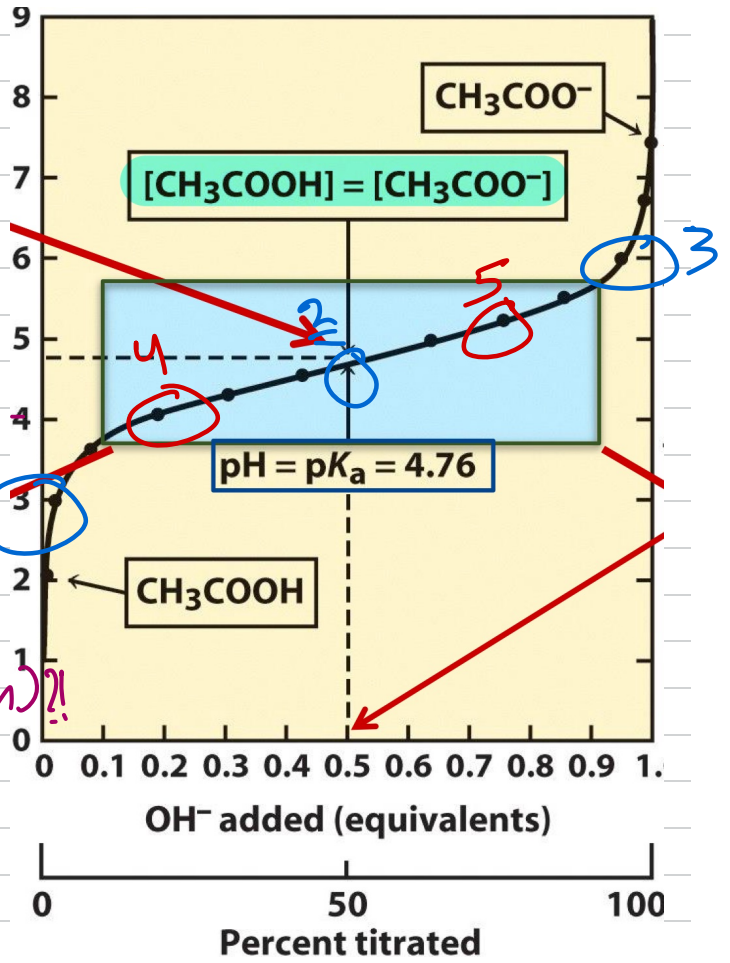
3.8 pH (lower end of buffering region) ?!

$pH = pKa + \log \frac{A^-}{HA}$

$3.8 = 4.8 + \log x$

$-1 = \log x \rightarrow x = 0.1$

Q) Same question but in the upper end of buffering region ?!



$pH = pKa + \log x$

$5.8 = 4.8 \log x$

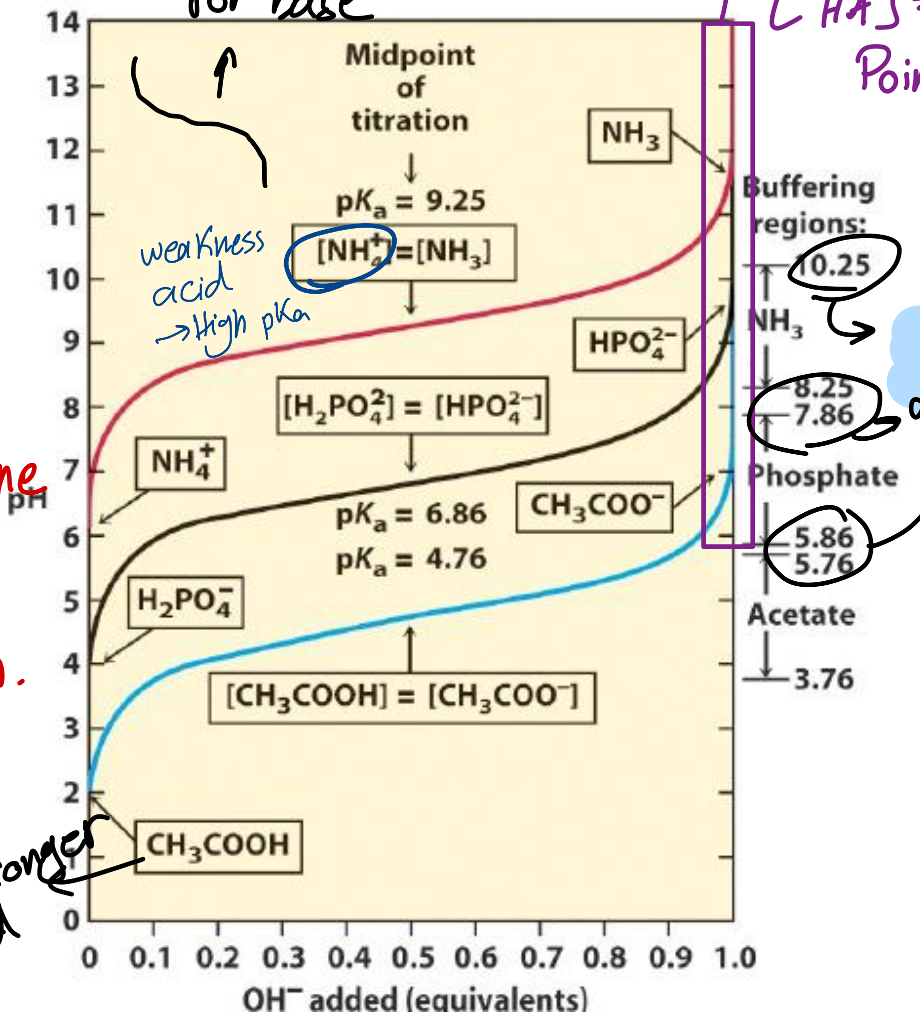
$1 = \log x$

$x = 10$

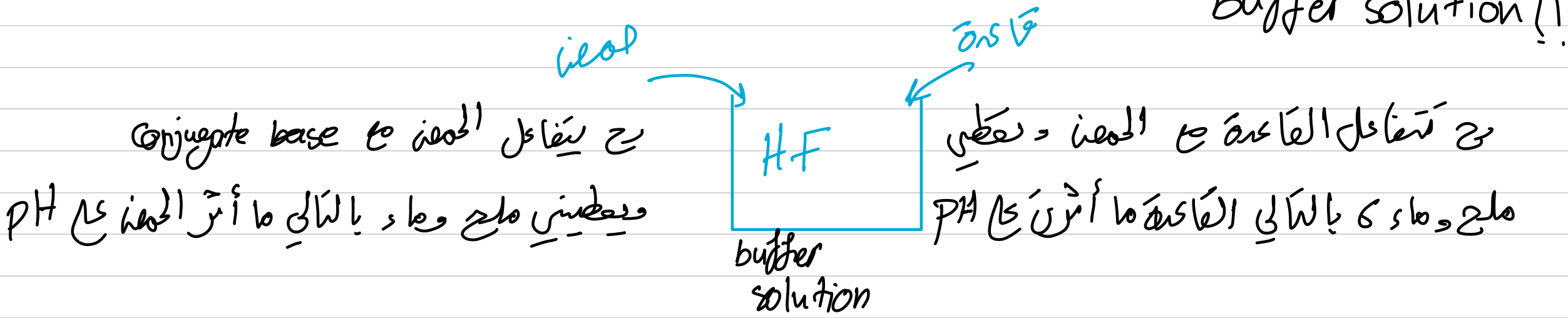
How do we make/choose a buffer?

- A buffer is made by combining weak acid/base and its salt.
- The buffering capacity of a buffer to function depends on:
  - Buffer concentration
  - Buffering range
    - pKa of the buffer
    - The desired pH

Note: increasing the concentration does not change the buffering range but it increases buffering capacity or strength.



How the PH doesn't change when I add a strong base or acid to a buffer solution?!



Conjugate base ح يتفاعل الحمض ح يتفاعل القاعدة ح الحمض ، يعطي
وبعطيني ملح وماء ، بالتالي ما أثر الحمض على PH ملح وماء ، بالتالي القاعدة ما أثرت على PH

* How I can prepare a buffer?!
get equal concentration of HA and A^-,
the question is: what would be the PH?!

the PH = pKa, because equal concentration

$$pH = pK_a + \log\frac{1}{1}$$

Ka equation ← (acidic) solution في السؤال طلب التعرف -

Henderson equation ← buffer " " " " " -

Exercise

- A solution of 0.1 M acetic acid and 0.2 M acetate ion. The pKa of acetic acid is 4.8. Hence, the pH of the solution is given by

$$\mathrm{pH} = 4.8 + \log(0.2/0.1) = 4.8 + \log 2.0 = 4.8 + 0.3 = 5.1$$

- Similarly, the pKa of an acid can be

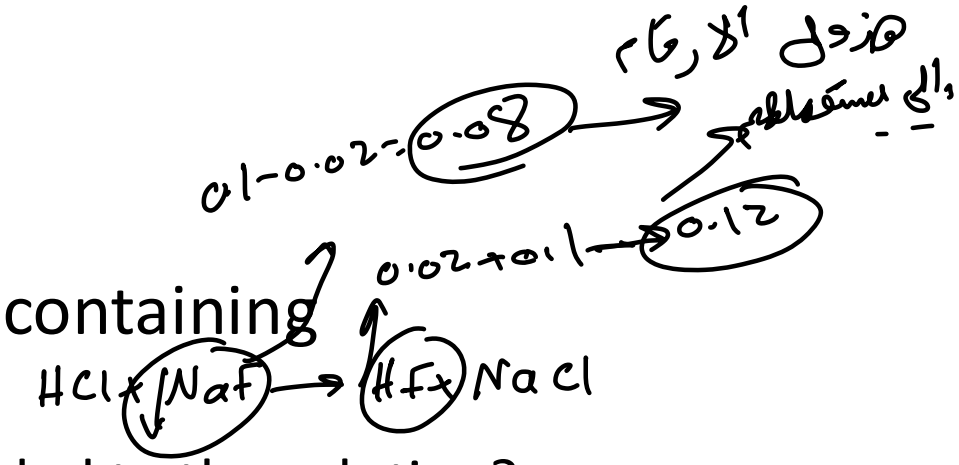
Since you are smart, you can estimate it.
No calculator is needed.

Base 10 Logarithms

$\mathrm{Log}_{10}(0.01)=-2$	since $10^{-2}=1/100$
$\mathrm{Log}_{10}(0.1)=-1$	since $10^{-1}=1/10$
$\mathrm{Log}_{10}(1)=0$	since $10^{0}=1$
$\mathrm{Log}_{10}(10)=1$	since $10^{1}=10$
$\mathrm{Log}_{10}(100)=2$	since $10^{2}=100$
$\mathrm{Log}_{10}(1000)=3$	since $10^{3}=1000$

Exercise

1. Predict then calculate the pH of a buffer containing
   - 0.1M HF and 0.12M NaF? ($K_a = 3.5 \times 10^{-4}$)
   - 0.1M HF and 0.1M NaF, when 0.02M HCl is added to the solution?
2. What is the pH of a lactate buffer that contain 75% lactic acid and 25% lactate? (pKa = 3.86)
3. What is the concentration of 5 ml of acetic acid that can be titrated completely by 44.5 ml of 0.1 N of NaOH? Also, calculate the normality of acetic acid.
   - The number of equivalents of OH^- required for complete neutralization is equal to the number of equivalents of hydrogen ion present as H^+ and HA.



Q1) pH ?!

A) $pH = -\log(3.5*10^{-4}) + \log 1.2 = 3.4$

or

$3.5*10^{-4} = \frac{[H^+]*0.12}{0.1} \rightarrow pH = 3.4$

* إذا كان عندي حمض وضفت حمض قوي بزيد تركيز الحمض القوي الضعيف وبنقص من القاعدة المرافقة نفس تركيز الحمض القوي

في حال ضفت قاعدة M=0.02

$K_a = \frac{[H^+][0.1+0.02]}{[0.1-0.02]}$

B) $HF \quad F^- \quad HCl \rightarrow H^+ \quad Cl^-$

$0.1 \quad 0.1 \quad 0.02 \quad 0.02 \quad 0.02$

قاعدة - حمض

$K_a = \frac{[H^+][0.1-0.02]}{[0.1+0.02]}$

حمض حمض

$[H^+] = 3.5*10^{-4} * 0.12/0.08 \Longrightarrow pH = 3.28$

Q2) HA A^- $pKa = 3.86$

%75 %25

$$pH = 3.86 + \log\left(\frac{1}{3}\right) = 3.38$$

Q3) HA NaoH

M ?!

0.1 N

5 mL 44.5 mL

$$MV = MV$$

$$M = \frac{44.5 \times 0.1}{5} = 0.89\ M$$

$CH_3COOH + H_2O \rightleftharpoons H_3O^+ + CH_3COO^-$

يزداد بمقدار تركيز الحمض القوي

when add strong acid

يقل بمقدار تركيز الحمض القوي

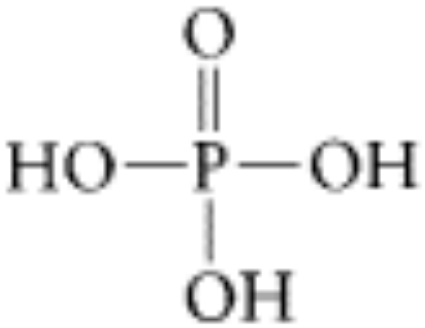
← هذه التغيرات في التراكيز وانخفاض لا تزال تحدث من أجل الحفاظ على constant Ka

$CH_3COOH + H_2O \rightleftharpoons H_3O^+ + CH_3COO^-$

يقل بمقدار تركيز القاعدة القوية

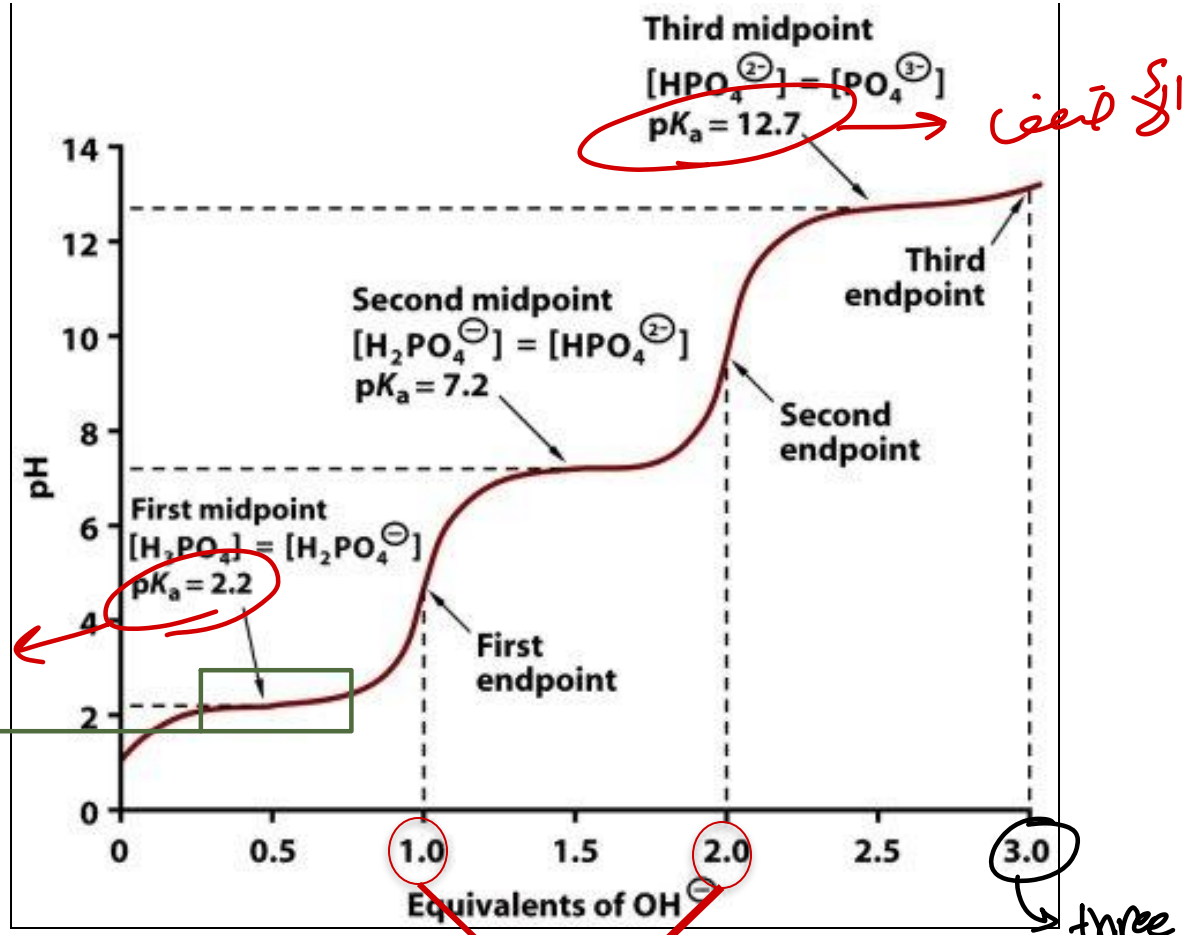
when add strong base

يزداد بمقدار القاعدة القوية



$$H_3PO_4 \underset{}{\overset{pK_{a(1)} = 2.14}{\rightleftharpoons}} H_2PO_4^- + H^+ \overset{pK_{a(2)} = 7.20}{\rightleftharpoons} HPO_4^{2-} + H^+ \overset{pK_{a(3)} = 12.4}{\rightleftharpoons} PO_4^{3-} + H^+$$

Titration curve of phosphate buffer



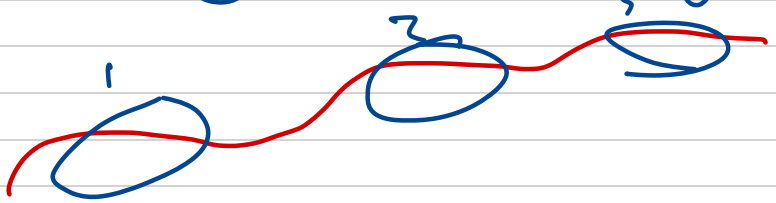
buffering range [1.2 – 3.2]

* below pH = 1.2, There is $H_2PO_4^-$ but in little amounts and there is no resistance for pH, the resistance starts at (1.2–3.2) and (6.2–8.2) and (11.7–13.7)

Note values

three eq. is used = 3 H^+

Q) How many buffering capacities did you see?!



Three

Q) ① what respresnt?!

it is midpoint or pKa for the first H^+

→ the buffering rang (1.2 – 3.2)

Q) In ①, what is the form of the phosphate buffer?! H_3PO_4 and $H_2PO_4^-$

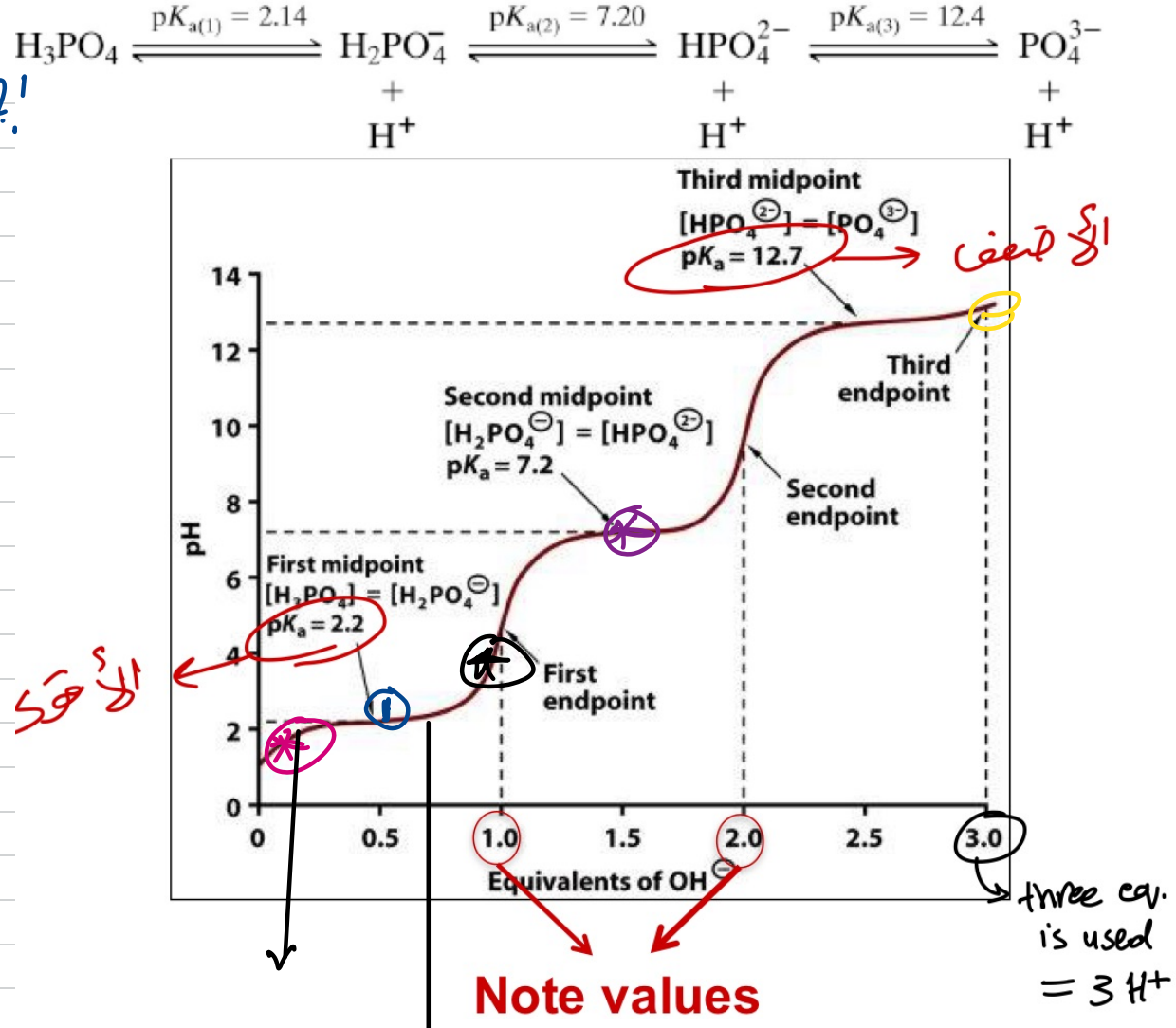
Q) What is the predominant form at:—

Ⓐ: H_3PO_4

Ⓑ: $H_2PO_4^-$

★: ($H_2PO_4^- + HPO_4^{-2}$) equal concentrations

Ⓐ: PO_4^{-3}



The ratio capacity in the: — $H_3PO_4 : H_2PO_4^-$

first of the buffering range?! 10 : 1

end of the buffering range?! 1 : 10

- I need 0.5 Equivalent to reach ?!

first pKa or (midpoint or inflectionpoint) for (H_3PO_4 – $H_2PO_4^-$) buffer.

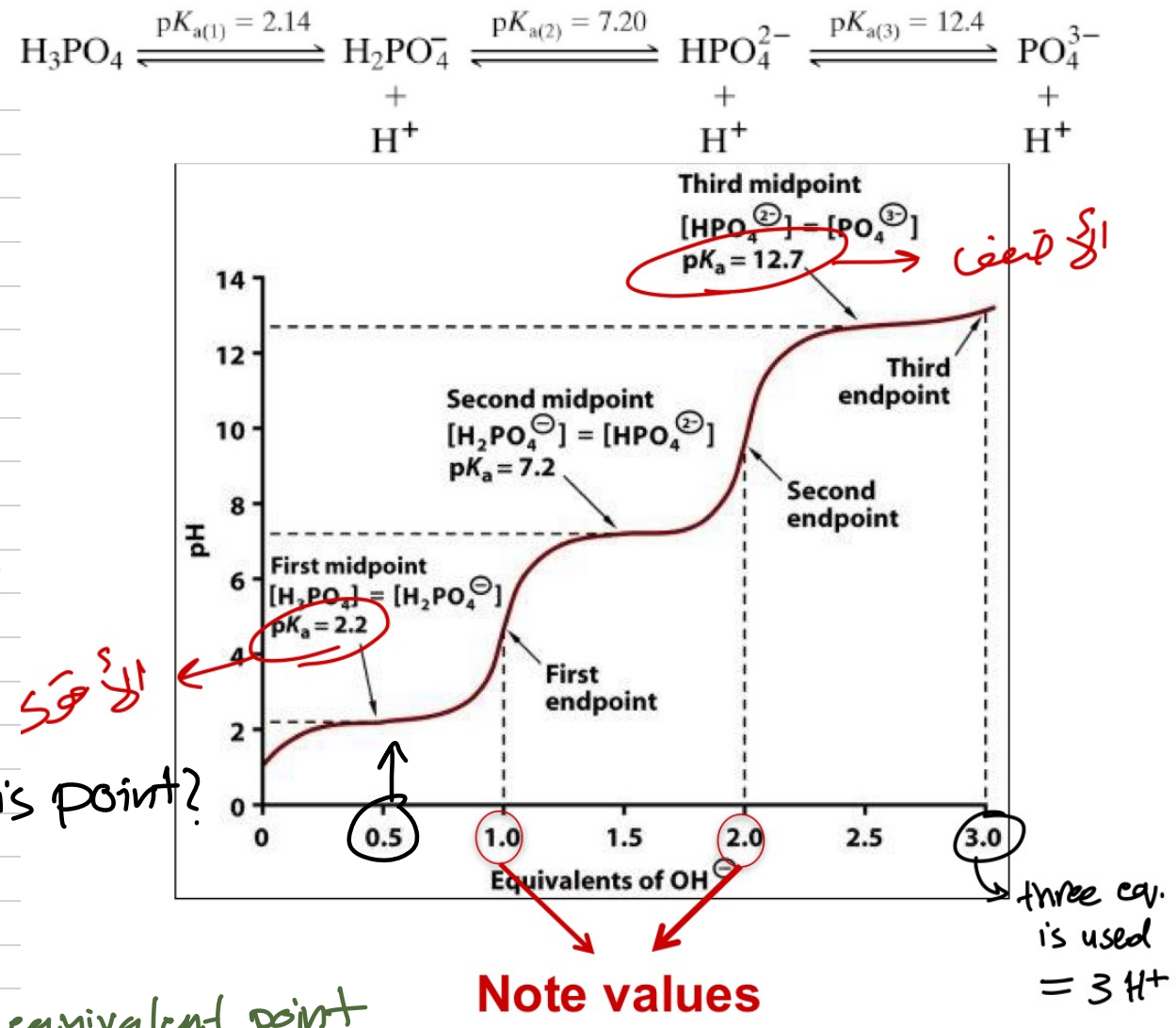
- I need 1 Equivalent to reach ?!

Equivalent point or end point (equal moles where 100 % Titration.

and what is the predominant form in this point?

100 % $H_2PO_4^-$

- I need 2 equivalent (of the base) to reach the 2nd equivalent point



Excercises

Henderson معادلة

$H_2PO_4^-$

1. What is the pKa of a dihydrogen phosphate buffer when a pH of 7.2 is obtained when 100 ml of 0.1 M NaH_2PO_4 is mixed with 100 ml of 0.3 M Na_2HPO_4?

$7.2 = pKa + \log\left(\frac{0.3}{0.1}\right) \rightarrow pKa = 6.72$

2. A solution was prepared by dissolving 0.02 moles of acetic acid (pKa = 4.8) in water to give 1 liter of solution.
   - What is the pH? $Ka = 10^{-pKa} \rightarrow Ka = \frac{x^2}{HA} \rightarrow pH = -\log[x]$
   - To this solution, 0.008 moles NaOH were added. What is the new pH? (ignore changes in volume).

$0.02 - 0.008 = 0.012$

HA
0.02 mol

$pKa = 4.8$

$$Ka = \frac{[H^+][A^- + 0.008]}{[HA - 0.008]}$$

Regarding Excersise 1 c-

100 mL of 0.1 M NaH_2PO_4 mixed with the 100ml of 0.3 M of Na_2HPO_4

* the new volume is 200 mL, so the concentrations will change

for Na_2HPO_4 :-

$$n = M \cdot V = 0.3 * 0.1 = \boxed{0.03 \text{ mole}}$$

The new M is:-

$$M = \frac{0.03}{0.2} = \boxed{0.15\ M}$$

* # of moles doesn't change

for NaH_2PO_4

$$n = M \cdot V = 0.1 * 0.1 = 0.01 \text{ moles}$$

The new M :- $\frac{0.01}{0.2} = 0.05$

$$pH = pKa + \log \frac{Na_2HPO_4}{NaH_2PO_4}$$

$$7.2 = pKa + \log \frac{0.15}{0.05}$$

$$pKa = 6.72$$

لإجابة نفسها لأن Na_2HPO_4 and $NaHPO_4$ من نفس الحجم، لكن إذا كانوا أخذ حجم مختلفة يسوي نفس هاي الخطوة

Assume: 200 mL of 0.1M of NaH_2PO_4 mixed with 100 ml of 0.3 of Na_2HPO_4.

$7.2 = pKa + \log\left(\frac{0.3}{0.1}\right) \rightarrow pKa = 6.72$ ✗ wrong

The correct answer is: —

The new volume is 200ml + 100ml = 300ml = 0.3L

① NaH_2PO_4

$n = M \cdot V = 0.1 * 0.2 =$ 0.02 moles

$M_{new} = \frac{n}{V_{new}} = \frac{0.02}{0.3} =$ 0.067

for Na_2HPO_4

$n = M \cdot V = 0.1 * 0.3 = 0.03$

$M_{new} = \frac{n}{V_{new}} = \frac{0.03}{0.3} =$ 0.1

$pH = pKa + \log \frac{Na_2HPO_4}{NaH_2PO_4}$

$7.2 = pKa + \log \frac{0.1}{0.067}$

$pKa = 7$ different answer ✓

Biological buffers in human body

it is open system (acts as 1st line to defence) / while chemical buffers are closed system.

- Carbonic acid-bicarbonate system (blood) [Extracellular]
- Dihydrogen phosphate-monohydrogen phosphate system (intracellular) because the PKa ≅ 7
  - ATP, glucose-6-phosphate, bisphsphoglycerate (RBC) [Intracellular]
- Proteins (why?) both (intra+extra)cellular, because they are everywhere
  - Hemoglobin in blood
  - Other proteins in blood and cells

] act as buffer system

Bicarbonate buffer

$$CO_2(d) + H_2O \underset{}{\overset{K_h}{\rightleftharpoons}} H_2CO_3$$

dissolved

$$H_2CO_3 \underset{}{\overset{K_a}{\rightleftharpoons}} H^+ + HCO_3^-$$

→ ما بحمل
يسمح & pKa

Blood (instantaneously)

على الفور

*the acid

$$\mathbf{CO_2 + H_20 \rightleftharpoons H_2CO_3 \rightleftharpoons H^+ + HCO_3^-}$$

not stable (dissociate) in 2 ways

if the blood more acidity
بنزيد الزفير للتخلص من
CO_2 لأنه مصدر الحموضة
والعكس صحيح

Lungs + RBCs (2nd line deffence) **(within minutes)**

افراز

3rd line deffence

Excretion via kidneys (hours to days)

Titration curve of bicarbonate buffer

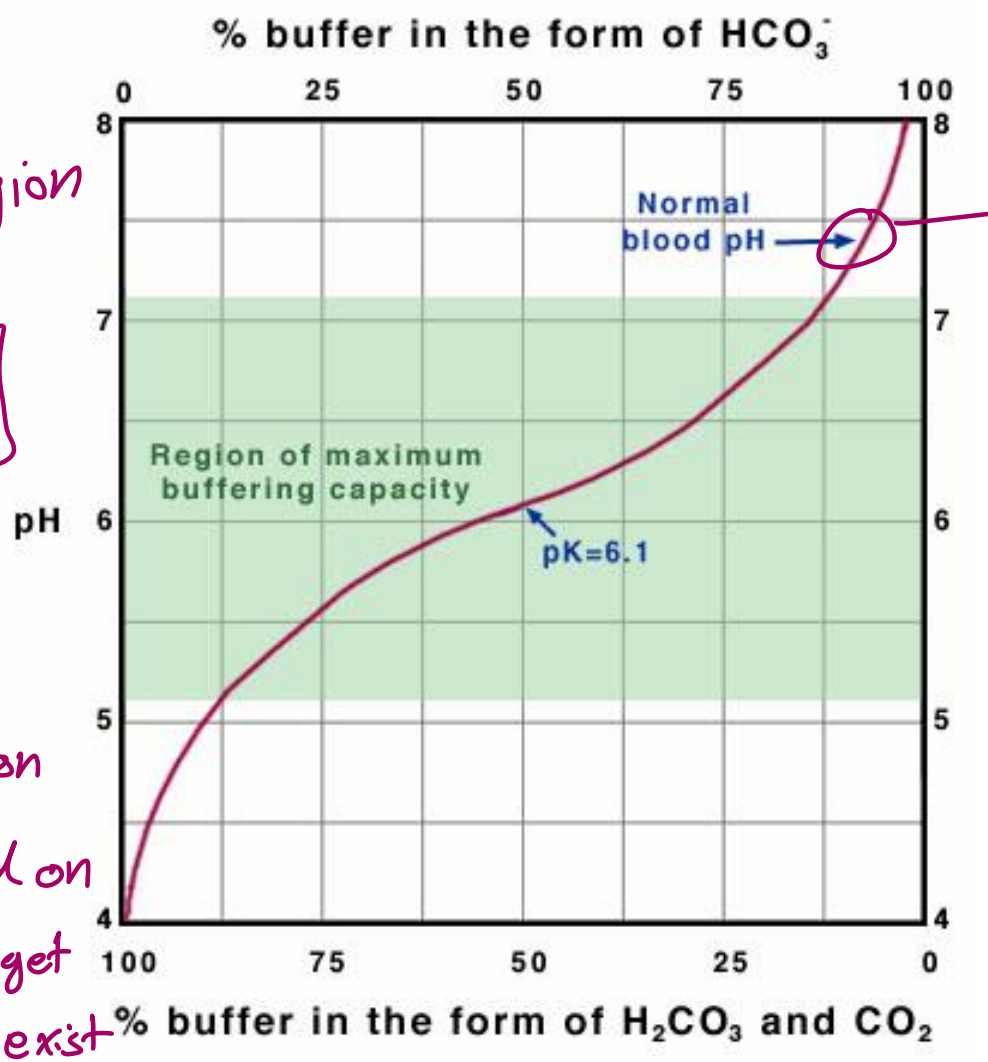
Note pKa

[5.1 – 7.1] buffering region

$7.4 = 6.1 + \log\left[\frac{HCO_3^-}{CO_2}\right]$

$\frac{HCO_3^-}{CO_2} \cong 20$

* I need high concentration of HCO_3^- and I can't depend on the dissociation of H_2CO_3 to get HCO_3^- because it is slow and not exist in a large amounts



→ out of buffering rang

Why is this buffer effective?

- Even though the normal blood pH of 7.4 is outside the optimal buffering range of the bicarbonate buffer, which is 6.1, this buffer pair is important due to two properties:
  1) bicarbonate is present in a relatively high concentration in the ECF (24mmol/L).
  2) the components of the buffer system are effectively under physiological control: the CO_2 by the lungs, and the bicarbonate by the kidneys.
  3) It is an open system (not a closed system like in laboratory).
     - An open system is a system that continuously interacts with its environment.

Why pH out of buffering rang?!

they are regulated

Acidosis and alkalosis

- Both pathological conditions can be either metabolic or respiratory.

Acidemya in blood — in all body

- Acidosis (pH< 7.35)
  - Metabolic: production of ketone bodies (starvation and diabetes)
  - Respiratory: pulmonary (asthma; emphysema)

- Alkalosis (pH > 7.45)
  - Metabolic: excessive administration of salts
  - Respiratory: hyperventilation (anxiety)

Pathological conditions → metabolic → kidney → not well control H_2CO_3

Pathological conditions → respiratory → lungs → not well control CO_2

Respiratory conditions

↑CO_2 → ↑H^+ → ↑Acidity

problem in CO_2

Respiratory Acidosis

↑ H^+ + ↑HCO_3^- ↔ ↑H_2CO_3 ↔ ↑CO_2 + H_2O

- when you breath quickly

[hyperventilation] → [↓CO_2]

So you can't control your breathing

عشان طبيعي يعمي عليك (Pass out)

Respiratory Alkalosis

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

↓pH ← ↑CO_2 بشكل طبيعي

لكي يتنفس الجسم

حتى يستقر pH ثم تستيقظ

Solution: breath in a bag

↑CO_2 → ↓pH

Metabolic conditions

problem in H+

Acidosis: - (↑)
1) Metabolic: H^+ ↑
2) Respiratory: CO_2 ↑

Alkalosis: - (↓)
1) Metabolic: ↓ H^+
2) Respiratory: ↓ CO_2

Metabolic Acidosis

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

→

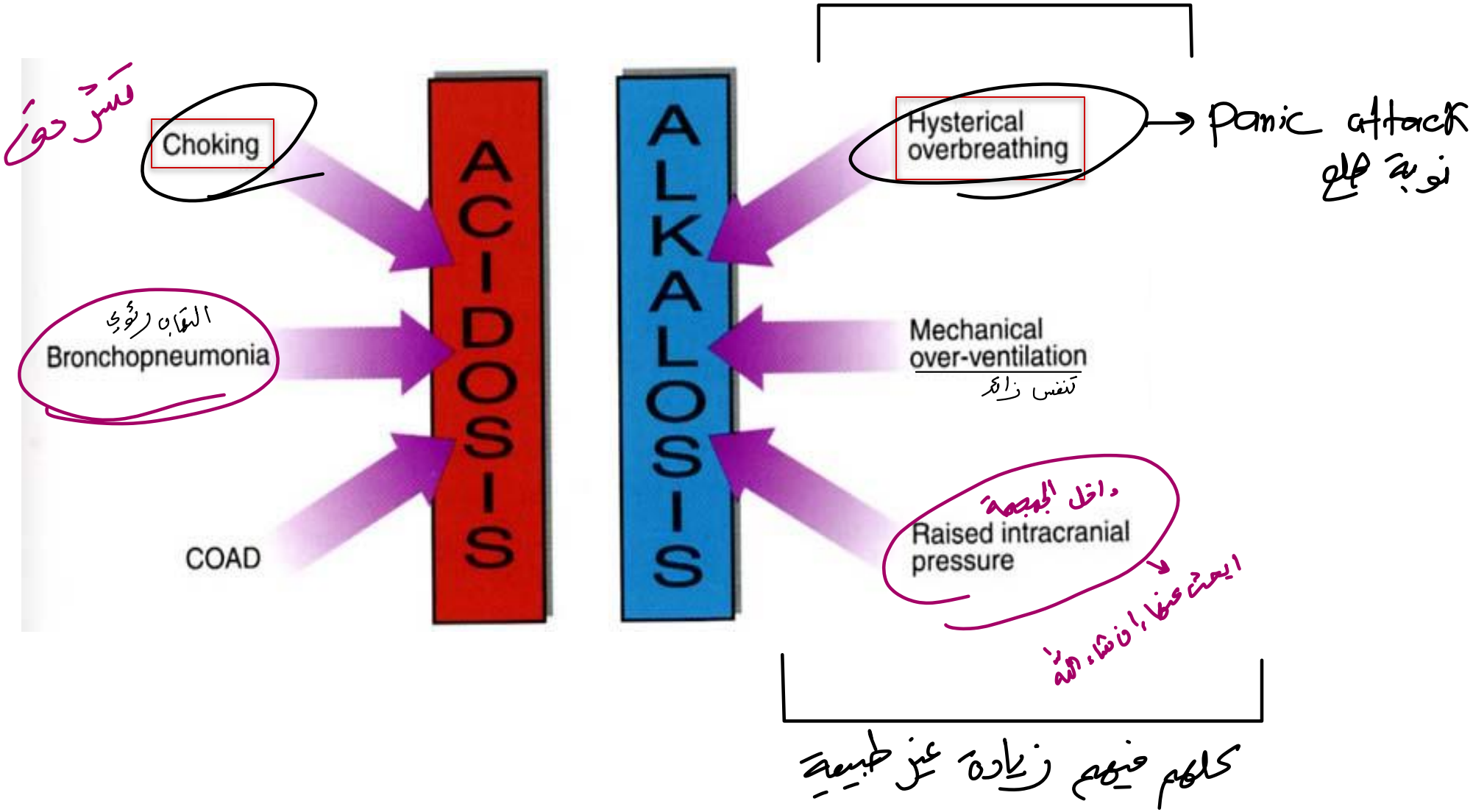
Metabolic Alkalosis

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

←

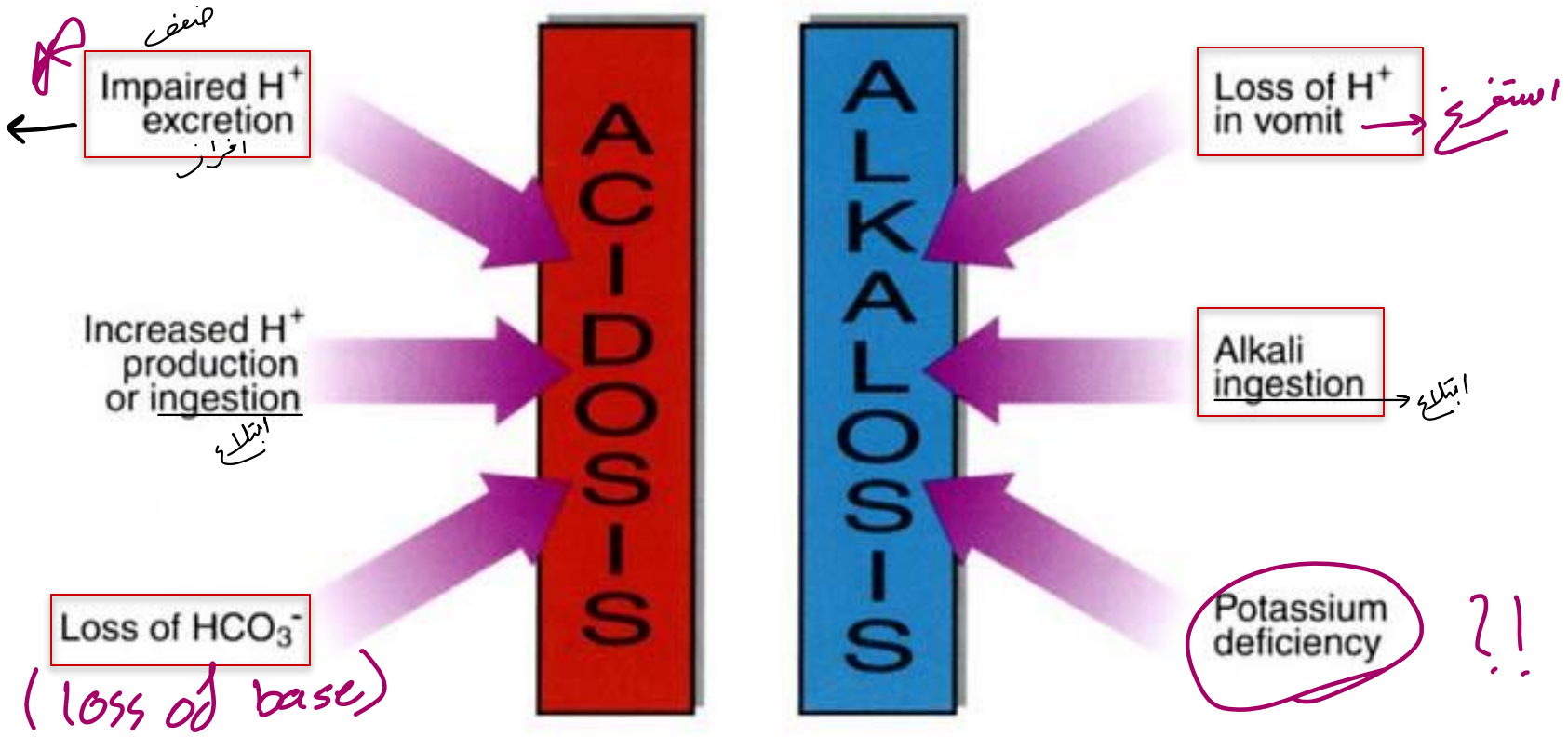
- على lungs بتقدر تتحكم فيه Kidneys بتشتغلوا
- على Kidneys بتشتغلوا lungs تتولى الأمور

Causes of respiratory acid-base disorders



Causes of metabolic acid-base disorders

Kidneys not functioning [metabolic]



تعويض

Compensation

- Compensation: A change in HCO_3^- or pCO_2 as a result of the primary event in order to return the pH to normal levels.
- If the underlying problem is metabolic, hyperventilation or hypoventilation alters pCO_2; it is called respiratory compensation.
- If the problem is respiratory, renal mechanisms drives metabolic compensation via changing [HCO_3^-].
- Complete compensation if brought back within normal limits [7.35–7.45]
- Partial compensation if the pH is still outside norms.

* When you see pCO_2 + CO_2 outside their normal ranges → Complete compensation

- but the CO_2 and pCO_2 in the blood are outside normal ranges but the ratio is still at normal one

Blood (instantaneously)

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

Lungs (within minutes)

Excretion via kidneys (hours to days)

* لذلك احفظ المعادلة منيح

Q) What is the pathological condition will happen if:-

↓ HCO_3^- ?!

lungs

$$CO_2 + H_2O \rightleftharpoons H_2CO_3 \rightleftharpoons \uparrow H^+ + \downarrow HCO_3^-$$

kidneys

→

↓ pH, so: Acidosis, metabolic (the problem in kidneys)