

strong acid \rightarrow completely ionize.

weak \rightleftharpoons \rightarrow parti \rightarrow equi $\rightarrow K_{eq} = \frac{[\text{prod.}]}{[\text{reactan.}]}$

$$K_w = \frac{K_a K_b}{1 \times 10^{-14}}$$

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

Weak acids and bases

Strong acid and base are completely ionized ✓

$$K_a = \infty$$

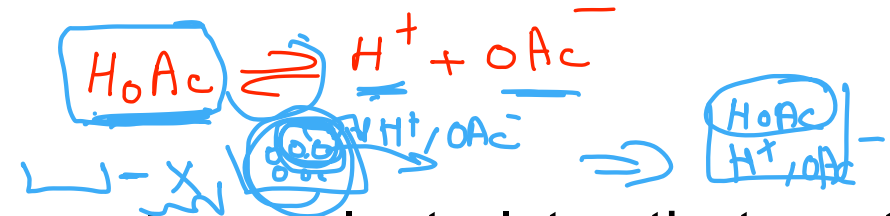
Weak acid and base are partially ionized

incomplete

Mineral (inorganic) acids e.g. HCl, HClO₄, HNO₃ are strong electrolytes that are totally ionized in H₂O

Most organic acids and bases as found in clinical applications, are weak. (e.g. trichloroacetic acid, acetylsalicylic acid (Aspirin))

*الاسوية التي تكون معظمها organic تكون weak acids
weak base.*



Equilibrium constant for dissociating species

weak

strong

Compounds are either partially (e.g. CH_3COOH) or completely (HCl) dissociated when dissolved in H_2O

$K_a = \infty$, can't use equilibrium constant.

Weak electrolytes: are partially dissociated electrolytes



equilib

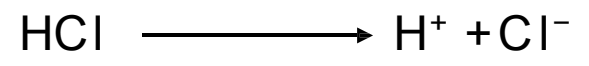
$$K_a \text{ or } K_{eq} = \frac{[\text{CH}_3\text{COO}^-][\text{H}^+]}{[\text{CH}_3\text{COOH}]}$$



$K_{eq} = \frac{[\text{A}][\text{B}]}{[\text{AB}]}$ = molar equilibrium constant

Some species dissociate stepwise, and an equilibrium constant can be written for each step ***

Strong electrolytes: completely dissociated electrolytes e.g. *there isn't equilibrium for strong acid.*



Calculations using equilibrium constant

use previous chapter → strong
 use this chapter → weak
 use equilibrium constant.

Equilibrium concentration (mol/L) = (Initial analytical concentration) – (An amount reacted)

Steps to be followed:

1. Write the chemical reaction.
2. Write the expression.
3. Let x represents the concentration of unknown small compound amount compared to others
4. In calculations, for simplicity: In a dissociation, neglect x compared to the initial concentration C if $C \geq 100K_{eq}$

	1 HOAc	\rightleftharpoons	1 H^+	$+$	1 OAc^-	✓
before equil	# of moles		0		0	
after equil.	# of moles - x		x		x	

$K_{eq} = \frac{[x][x]}{[x - \text{moles} - x]}$

* ضروري نتأكد اننا متقاربه
 - ضروري معرفه stoichiometry

* نعمل check ← اذا كان التركيز اكبر اربعاوي 100 K_{eq} نعمل x من المقام



① اسأل نفسك ← ضعيف ← تفاعل جزئي (X) تقابل.

صفا قوی کے حصے میں

Example

Calculate the pH & pOH of a 1×10^{-3} M of acetic acid? $K_a = 1.75 \times 10^{-5}$ at 25°C .



② اكتب معادلة التفاعل ضعيف وتأكد انك متوازنة

Initial [before]

change

Equilibrium [after]

	HOAc	H ⁺	OAc ⁻
I (M)	1×10^{-3}	0	0
C (M)	-x	+x	+x
E (M)	$1 \times 10^{-3} - x$	x	x

make your table. ③

نواحي متعادلا

$K_a = \frac{[\text{H}^+][\text{OAc}^-]}{[\text{HOAc}]}$

$K_a = \frac{x^2}{[1 \times 10^{-3} - x]} = 1.75 \times 10^{-5}$

[after]

write (equilibrium constant) equation ④

Note: if $[C_{HA}] > 100 K_a$, x can be neglected compared to C_{HA}

$1 \times 10^{-3} \geq 100 \times 1.75 \times 10^{-5}$

if we can check ⑤ if we can neglect x or not.



X is not neglected

The solution from a quadratic equation

$$x = 1.32 \times 10^{-4} \text{M} = [\text{H}^+]$$

$$\text{pH} = -\log[\text{H}^+] = -\log 1.32 \times 10^{-4} = 3.88$$

$$\text{pOH} = 14 - 3.88 = 10.12$$

as same as previous chapter.

in our case:

$$100K_a = 100 \times 1.75 \times 10^{-5} = 1.75 \times 10^{-3}$$

$$[1 \times 10^{-3} \geq 1.75 \times 10^{-3}] \text{ No}$$

$$1.75 \times 10^{-5} = \frac{x^2}{(1 \times 10^{-3} - x)}$$
$$x^2 = (1.75 \times 10^{-5} \times 1 \times 10^{-3}) - 1.75 \times 10^{-5} x$$

[X should not be neglected]

- ترتيب جميع الحدود على نفس الطرف لتصبح معادلة
تربيعية $ax^2 + bx + c = 0$

$$x^2 + 1.75 \times 10^{-5} x - 1.75 \times 10^{-8} = 0 \longrightarrow \text{use your calculator}$$

$$x = 1.32 \times 10^{-4} \text{M} = [\text{H}^+]$$



→ acid → K_a
 → base → K_b

$K_w = K_a K_b$

Example

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 $1.00 \times 10^{-3}\text{M}$ solution of ammonia

use equilibrium constant. ← partially weak dissociation ①

اكتب معادلة كيميائية موزونة. ②



	NH_3	NH_4^+	OH^-
I (M)	1×10^{-3} ✓	0	0
C (M)	-x	+x	+x
E (M)	$1 \times 10^{-3} - x$	x	x

الكمية الاساسية فقيرت من مقدارها الكمية التي تفاعلت جزئياً

$K_b = \frac{[x][x]}{1 \times 10^{-3} - x}$ ، تصرون قيمة OH^-

$x = 1.32 \times 10^{-4}\text{M} = [\text{OH}^-]$

$\text{pOH} = -\log 1.32 \times 10^{-4} = 3.88$

$\text{pH} = 14.00 - 3.88 = 10.12$

* الارقام نفس المتكاتبين و نفس طريقة

ولكن $x = [\text{OH}^-]$ عند pOH نحو pH .



Salts of weak acids and weak bases

strong electrolyte

$K_a \rightarrow H^+$ (Acid)
 $K_b \rightarrow OH^-$ (base)
 $K_w = K_a * K_b$
 1×10^{-14}

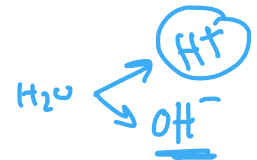
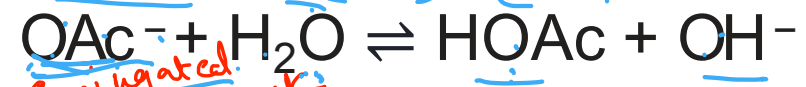
HOAc \rightarrow H^+

The salt of a weak acid, for example, NaOAc, is a strong electrolyte that completely ionizes



Na^+
 K^+ (salt of)

② \rightarrow behaves as base.



$K_b = \frac{[HOAc][OH^-]}{[OAc^-]} = \frac{[OH^-]^2}{[OAc^-]} = \frac{K_w}{K_a}$
1:1

$K_w = K_a * K_b$
 $K_b = \frac{K_w}{K_a}$
 $\therefore K_w = 1 \times 10^{-14}$
قيمة ثابتة

For any salt of weak acid

$$[OH^-] = \sqrt{K_b[A^-]}$$

$[OH^-]^2 = \frac{K_w[OAc^-]}{K_a}$

take square roots.

$[OH^-] = \sqrt{\frac{K_w[OAc^-]}{K_a}} = \sqrt{K_b[OAc^-]}$

لنستخرج العلاقة

* عندما نلقي نظرة عامة على سلوك ملح الحمض الضعيف نستنتج انه في معادله كيميائية انتج ايون OH^- كما تفصل القاعدة.

لنتوصل الى ان ملح الحمض الضعيف يمتلك سلوك القاعد بعد التذبذبة
 salt of weak acid After dissociate it's behaves as a base.



Example

وجود اسم ايون موجب غير H^+ دلاله على انه ملح لعنصر



Calculate the pH of 0.1 M sodium acetate ($CH_3COONa = NaOAc$)? $K_a = 1.75 \times 10^{-5}$



ionized partially in water

تصرف القاعص
يكون انه يهترق
للحمض الضعيف
لا تتعد

salt of weak acid

بعبارة ثانية لربما

- Salt of weak acid means
 - behaves as base
 - complete ionized.

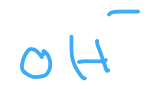
	OAc^-	$HOAc$	OH^-
I (M)	0.1	0	0
C (M)	-x	+x	+x
E (M)	<u>0.1 - x</u>	<u>x</u>	<u>x</u>

لانه سيكيز الملح
بالتكافؤ
ionized

K_b المرتبطة بـ $[OH^-]$
 $K_w = K_a K_b$ لايجاد K_b عن طريقها

تذكر ان الملح لحمض ضعيف يسلك سلوك قاعص يعني فنحتاج الى K_b وليس K_a لكن بما ان المعنى K_a نبدأ الى قانون





$0.1 - x$

x

x

$x \equiv [\text{OH}^-]$

base

K_b

Find K_b ①

check ②

next

→ K_a
→ K_b



$$K_b = \frac{[\text{X}][\text{X}]}{[\text{reactant}]} = \frac{[x]^2}{[0.1-x]}$$

* $K_w = K_a K_b$
 $\frac{1 \times 10^{-14}}{1.75 \times 10^{-5}} = \frac{1.75 \times 10^{-5}}{1.75 \times 10^{-5}} \times K_b$

$$\Rightarrow K_b = 5.7 \times 10^{-10}$$

$$x = \sqrt{5.7 \times 10^{-10} \times 0.1} = [\text{OH}^-]$$

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

$$\boxed{C \gg 100 K_b}$$

$0.1 \gg 100 \times 5.7 \times 10^{-10}$

Yes, then we can neglect x

Solution

$$pH = pK_a + pOH$$

نقطة التكافؤ

$$K_b = K_w/K_a = 1 \times 10^{-14} / 1.75 \times 10^{-5} = 5.7 \times 10^{-10} \quad K_b = 5.7 \times 10^{-10} = [\text{HOAC}][\text{OH}^-] / [\text{OAC}^-]$$

At Equilibrium $[\text{HOAC}] = [\text{OH}^-] = x$

$$K_b = \frac{[x][x]}{[0.1-x]} \quad , \quad [x] = [\text{OH}^-]$$

$[\text{OAC}^-] = 0.1$ (x is neglected because $0.1 > 100$ times than K_b)

نستعمل إذاً يجب أن يكون
تفعلها أم لا

$$K_b = 5.7 \times 10^{-10} = x^2 / 0.1$$

$$5.7 \times 10^{-10} = \frac{x^2}{0.1}$$

$$C \gg K_b * 100 \quad ??$$

$$x^2 = 5.7 \times 10^{-11}$$

$$\sqrt{x^2} = \sqrt{0.1 * 5.7 \times 10^{-10}}$$

$$0.1 \gg 5.7 \times 10^{-10} \quad \text{Yes}$$

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

→ Find pOH

$$[\text{H}^+] = 1 \times 10^{-14} / 7.6 \times 10^{-6} = 1.3 \times 10^{-9} \text{ M} \quad \text{pH} = -\log 1.3 \times 10^{-9} = 9 - 0.11 = 8.89$$

$$[\text{H}^+]$$

مباشرة على الآلة حاسبة

then, ignore x



Salts of weak bases \rightarrow behaves as acid \rightarrow completely ionized

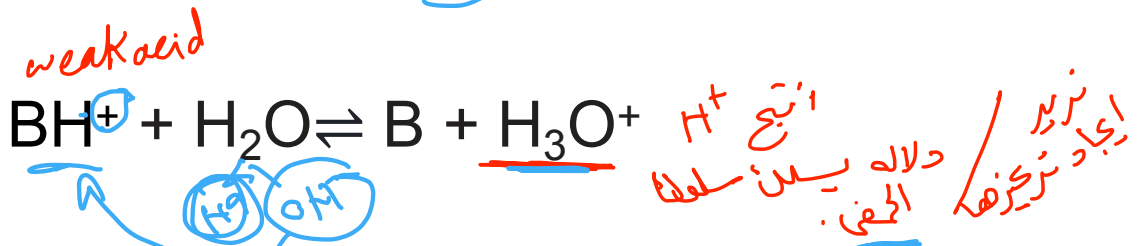
تفكك القاعدية
تفكك المبدأ

الأيون سالب
 OH^-

Cl^-



The salt of a weak base completely ionizes in water



$$K_a = \frac{[B][H_3O^+]}{[BH^+]} = \frac{[H^+]^2}{[BH^+]} = \frac{K_w}{K_b}$$

لأنه
طرح قاعدية ضعيفة
يعطينا بالمواد كات
العايون. لذلك نتخذ هذا
Kb.

$$[H^+]^2 = \frac{K_w[BH^+]}{K_b}$$

$$[H^+] = \sqrt{\frac{K_w[BH^+]}{K_b}} = \sqrt{K_a[BH^+]}$$

General equation

For any salt of weak bases

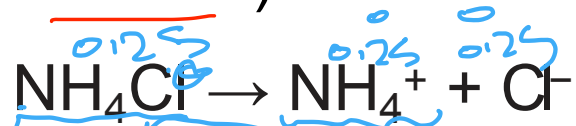
$$[H^+] = \sqrt{K_a[BH^+]}$$


Example

salt

ايون سالبا
نستنج انه ملح قاعده

Calculate the pH of a 0.25 M solution of ammonium chloride NH_4Cl ? (K_b for $\text{NH}_3 = 1.75 \times 10^{-5}$)



totally ionized.

→ RXN 1:1

$[\text{NH}_4^+] = 0.25\text{M}$
 $[\text{Cl}^-] = 0.25\text{M}$

↓
 K_a



Partially dissociate with water.

	NH_4^+	H^+	NH_4OH
I (M)	0.25	0	0
C (M)	-x	+x	+x
E (M)	0.25 - x	x	x

* act as acid → find K_a ✓ → $K_a = \frac{K_w}{K_b} = \frac{1 \times 10^{-14}}{1.75 \times 10^{-5}} = 5.7 \times 10^{-10}$

$K_a = \frac{[x^2]}{0.25 - x}$

$C \geq 100 K_a$
 $0.25 \geq 100 * 5.7 \times 10^{-10}$

Yes, then neglect x.

$\sqrt{0.25 * K_a} = x$
 $x = [\text{H}^+] = 1.2 \times 10^{-5}$
Find pH



Solution

$$K_a = K_w / K_b = [\text{NH}_4\text{OH}][\text{H}^+] / [\text{NH}_4^+]$$

$$K_a = 1 \times 10^{-14} / 1.75 \times 10^{-5} = 5.7 \times 10^{-10}$$

At equilibrium $[\text{NH}_4\text{OH}] = [\text{H}^+] = x$

$[\text{NH}_4^+] = 0.25 - x$ (x is neglected because $[\text{NH}_4^+] > 100 K_a$)

$$K_a = 5.7 \times 10^{-10} = x^2 / 0.25$$

$$x = [\text{H}^+] = \sqrt{K_a \times [\text{NH}_4^+]} = \sqrt{5.7 \times 10^{-10} \times 0.25} = 1.2 \times 10^{-5}$$

$$\text{pH} = -\log 1.2 \times 10^{-5} = 4.92$$

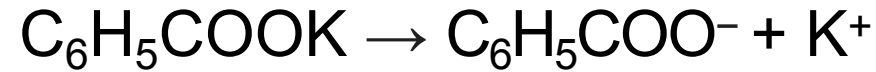
ففسى
الحل السابق.



Example

[please try to solve it]

Calculate the pH of a 0.025M solution of potassium benzoate at 25 C if (Ka for benzoic acid = 6.25×10^{-5})



	$\text{C}_6\text{H}_5\text{COO}^-$	$\text{C}_6\text{H}_5\text{COOH}$	OH^-
I (M)	0.025	0	0
C (M)	- x	+ x	+ x
E (M)	0.025 - x	x	x

هل نغفل x؟

$0.025 \geq 100 * K_b \rightarrow$ Yes, neglect it.

$$K_b = \frac{x^2}{0.025}$$

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

Find pOH

Find pH

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



Salt of weak acid



يعني ملح قوي يعني يلوك سلوك قاعدة .

اذا كان ملحون قاعدة اذا كان كاتيون الـ K^+

$$K_w = K_a K_b$$

$$10^{-14} = 6.25 \times 10^{-5} \times K_b$$

Find K_b

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

Solution

$$K_b = [\text{C}_6\text{H}_5\text{COOH}][\text{OH}^-] / [\text{C}_6\text{H}_5\text{COO}^-]$$

$$K_b = 1 \times 10^{-14} / 6.25 \times 10^{-5} = 1.6 \times 10^{-10}$$

$$K_b = 1.6 \times 10^{-10} = x^2 / [(0.025 - x)]$$

$$x^2 = 4 \times 10^{-12}$$

$$x = \sqrt{4 \times 10^{-12}} = 2.0 \times 10^{-6} = [\text{OH}^-]$$

$$\text{pOH} = -\log 2.0 \times 10^{-6} = 5.7$$

$$\text{pH} = 14 - 5.7 = \underline{8.3}$$

check your
answer



Buffers

- A buffer is a solution that resists change in pH when a small amount of acid or base is added, or when the solution is diluted

عبارة عن محاليل تقاوم التغيير في درجة الحموضة [pH] عند إضافة الحموض أو القواعد أو عند تخفيف المحاليل.

- A buffer solution consists of a mixture of weak acid & its conjugated base or a weak base & its conjugated acid

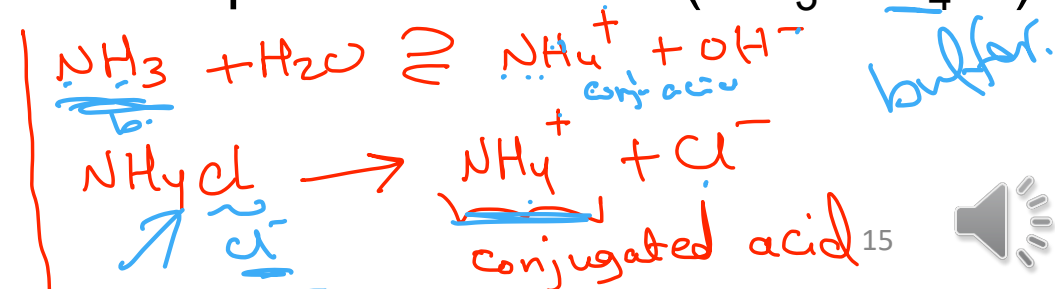
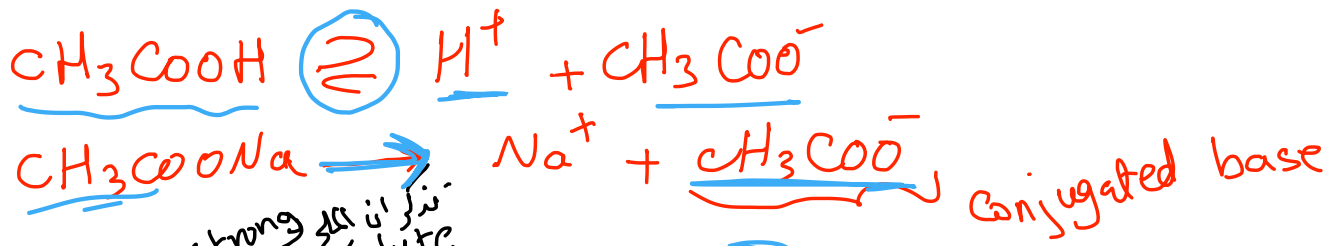
[weak acid + salt of its salt of weak acid]

[weak base + salt of weak base]

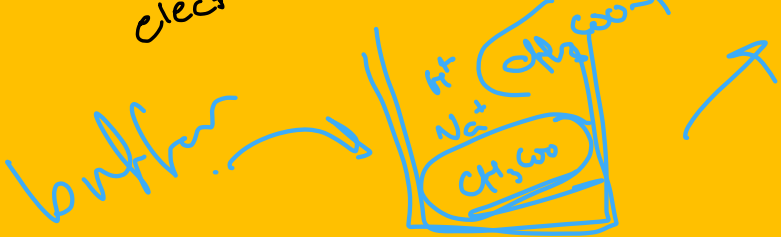
- Weak acid & its salt in solution is buffer. For example a mixture of (CH₃COOH/CH₃COONa)

+ OAc⁻
Salt of weak aci.
NaOAc → Na⁺ + OAc⁻
Conj

- Weak base & its salt in one solution is a buffer. For example a mixture of (NH₃/NH₄Cl)



Buffers



HOAc + OAc⁻ (acetic acid- acetate buffer)



K_a = $\frac{[H^+][OAc^-]}{[HOAc]}$ \rightarrow acid \rightarrow H⁺ \rightarrow use K_a

$K_a = \frac{[H^+][OAc^-]}{[HOAc]}$ *conjugate*

[H⁺] = K_a [HOAc] / [OAc⁻]

-log[H⁺] = -log (K_a · [HOAc] / [OAc⁻])

-log[H⁺] = -logK_a - log([HOAc] / [OAc⁻])

pH = pK_a - log([HOAc] / [OAc⁻])

*احد ترتيب ليكون H⁺ موضع العاينون
خذ واحد. لجميع الاطراف*

H⁺ \rightarrow pH

*العكس إشارة
بسط و مقام*

Note: (Change the sign – to +) the above equation becomes as follows:

pH = pK_a + log([OAc⁻] / [HOAc])

In general: pH = pK_a + log([salt] / [acid])

buffer *base/acid*
تركيز حمض *تركيز قاعده*

*يفضل حفظها
تسمى*

Henderson-Hasselbalch equation



Henderson-Hasselbalch equation

It is used to calculate the pH of a weak acid & its salt in solution \equiv pH of buffer solutions

$$pH = pKa + \log \left(\frac{[A^-]}{[HA]} \right)$$

buffer solution

Example: Acetate buffer HOAC/NaOAC

$$pH = pKa + \log \left(\frac{[OAC^-]}{[HOAC]} \right)$$

$$pH = pKa + \log \left(\frac{[base]}{[acid]} \right)$$

$$pH = pKa + \log \left(\frac{[proton acceptor]}{[proton donor]} \right)$$

مستقبل البروتون

مادة البروتون

General

Example: NH_3/NH_4Cl

$$pH = pKa + \log \left(\frac{[NH_3]}{[NH_4^+]}$$

$$= (pKw - pKb) + \log \left(\frac{[NH_3]}{[NH_4^+]}\right)$$

انتهى
على المقام
بنيان
acid

* buffer
↓
use henderson h
 $pH = pKa + \log \frac{base}{acid}$

هذه قاعدة سيكون قريباً
لذلك نجد
 $Ka = \frac{Kw}{Kb}$

عند اقتراب
الطرفين
المسماة تكون الى 1

$$pKa = pKw - pKb$$



خصائص اللوسيتيمات

Example



New case.

Calculate the pH of a buffer solution prepared by adding 10mL of 0.1M HOAC to 20 mL of 0.1M sodium acetate ($K_a=1.75 \times 10^{-5}$)

$$pH = -\log K_a + \log\left(\frac{[OAC^-]}{[HOAC]}\right)$$

Final volume = 10+20=30mL

$$M \text{ of HOAC} = (0.1 \times 10/30) = 0.033 \text{ M}$$

$$M \text{ of OAC}^- = (0.1 \times 20/30) = 0.067 \text{ M}$$

$$pH = -\log (1.75 \times 10^{-5}) + \log ([0.067]/[0.033])$$

$$pH = 4.76 + \log 2.0 \rightarrow pH = 5.06$$

For the log ([salt]/[acid]), volume gets cancelled and the number of mole ratio can be used

هنا السؤال اننا نستطيع من خلال Henderson ان نضع عدد المولات او التركيز

henderson ← buffer ← acid + its salt

$pH = -\log K_a + \log \left[\frac{\text{Base}}{\text{acid}} \right]$

$pH = -\log (1.75 \times 10^{-5}) + \log \left[\frac{V M_B / V_{total}}{V M_A / V_{total}} \right]$

$pH = -\log K_a + \log \frac{\# \text{ of moles Base}}{\# \text{ of moles Acid}}$

اصبح لينا mixing نفس الحالة في chapter 1 نبدأ لينا عدد المولات ليحبر عنها بطريقة اصح

$$pH = 4.76 + \log \frac{0.1 \times 0.02}{0.1 \times 0.01} \rightarrow pH = 5.06$$

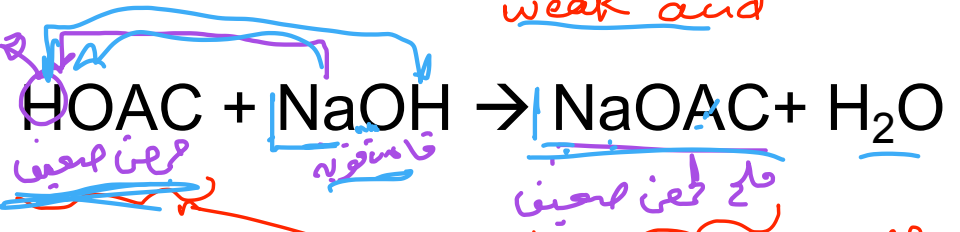


Example

mix or add → use number of moles
NaOH

Calculate the pH of a solution prepared by adding 25 mL of 0.1 M sodium hydroxide to 30 mL of 0.2 M acetic acid ($K_a = 1.75 \times 10^{-5}$)

NaOH
[strong base]



buffer
→ Henderson

this is buffer solution.

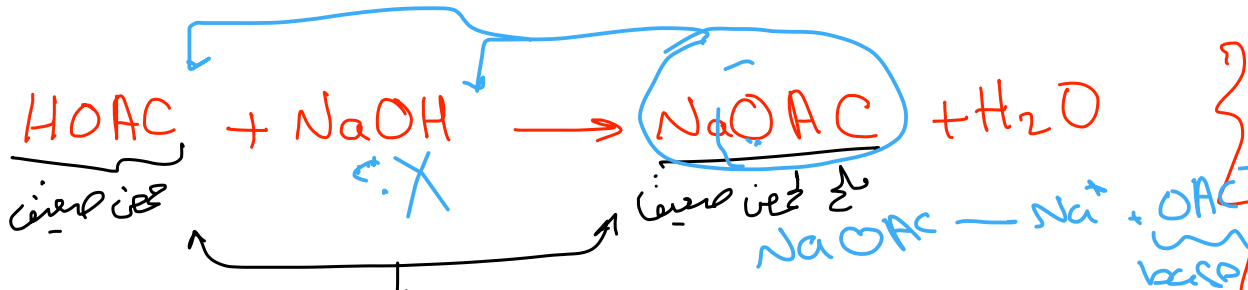
* نلاحظ من السؤال انه يوجد Case جديدة
وهي عبارة عن strong base مع weak acid
لكن ماذا لو لم نسرع بإلقاء الحكم

Solution:

mmol HOAC = $0.2 \times 30 = 6.0$
 mmol NaOH = $0.1 \times 25 = 2.5$ mmol of OAC⁻
 $6 - 2.5 = 3.5$ mmol HOAC excess
 $pH = pK_a + \log[\text{salt}]/[\text{acid}] =$
 $pK_a + \log[\text{OAC}^-]/[\text{HOAC}]$
 $pH = -\log 1.75 \times 10^{-5} + \log (2.5/3.5)$
 $pH = 4.76 + \log 0.71 = 4.61$

* الفاتحة
strong base + weak acid → salt of weak acid
means it buffer solution





القاعدة القاعده ان كمية المتفاعلات تحد بالاعتماد على المتفاعلات الاقل التي تنتهي اول لتفاعل

Limiting reactants

Case: Buffer
 ↓
 Henderson

mix - use number of moles.

$$\text{pH} = \text{pKa} + \log \frac{\text{base}}{\text{acid}}$$

$$\text{pH} = \text{pKa} + \log \frac{\# \text{ of moles base}}{\# \text{ of moles acid}}$$

$$\text{pH} = \text{pKa} + \log \frac{\# \text{ of moles OAC}^-}{\# \text{ of moles HOAC}}$$

انتبه اي قاعده
 كت في buffer
 القاعدة المتبقية من الملح
 = salt of weak acid.

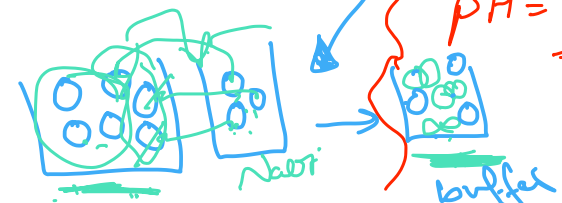
of moles acid = 6 mmol
 # of moles NaOH = 25 x 0.1 = 2.5 mmol
 الاقل ينتهي بسرعة

1 mole NaOH = 1 mole NaOAC = 2.5 mmol
 عدد مولات الحزن هي المولات التي بقيت باطون
 لتكون buffer اذا هي المولات التي لم تتفاعل.

of moles reacted HOAC = # of moles NaOH = 2.5
 # of moles remains = 6 - 2.5 = 3.5 mmol
 Limiting reactant.
 Ac.

$$\text{pH} = -\log [1.75 \times 10^{-5}] + \log \frac{2.5}{3.5} = 4.76 + \log(0.71) = 4$$

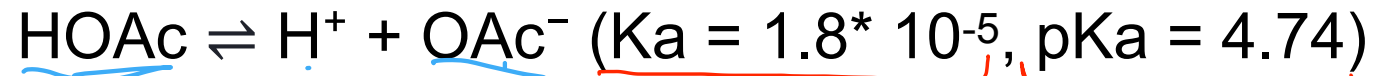
لما عرفت تركيز OAC⁻ من NaOH
 HOAC من HOAC



Buffer capacity

من خصائص buffer
كم يمكن ان اصيف H^+ على OH^- buffer
كث لي قام التغيير pH
] 10^5 كانت buffer Capacity كبيرة يكون افضل ويكون strong buffer.

A higher buffer capacity means that we can add more H^+ or OH^- before the pH changes



base/acid

takes $-\log$

Assume that $[A^-]/[HA] = 1.84$

نفس طور الـ A^- والـ HA

$$pH = pK_a + \log \left(\frac{[OAc^-]}{[HOAc]} \right) = 4.74 + \log 1.84 = 5.00$$

Henderson

$$pH = 5.00$$



Buffer capacity

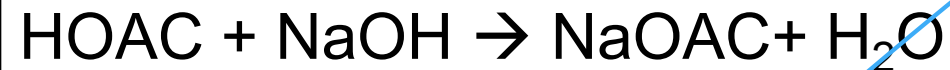
1] قبل إضافة NaOH
2] بعد إضافة NaOH

Buffer 1 **Has a higher buffer capacity**

$$[A^-] = 0.90 \text{ M} + 0.04 \text{ M} = 0.94 \text{ M}$$

$$[HA] = 0.49 \text{ M} - 0.04 \text{ M} = 0.45 \text{ M}$$

Add 0.04 mol NaOH to 1L buffer



$$\text{pH} = \text{pKa} + \log \left(\frac{[\text{OAC}^-]}{[\text{HOAC}]} \right) = 4.74 + \log [0.94/0.45]$$

$$\text{pH} = 5.06$$

Buffer 2

$$[A^-] = 0.090 \text{ M} + 0.04 \text{ M} = 0.13 \text{ M}$$

$$[HA] = 0.049 \text{ M} - 0.04 \text{ M} = 0.009 \text{ M}$$

X

$$\text{pH} = \text{pKa} + \log \left(\frac{[\text{OAC}^-]}{[\text{HOAC}]} \right) = 4.74 + \log [0.13/0.009]$$

$$\text{pH} = 5.90$$



Buffer 1

تركيز عيبي

قائمة

$$[A^-] = 0.9 \text{ M} \quad [\text{ملح الحمض ضعيف}]$$

$$[HA] = 0.49 \quad [\text{الحمض}]$$



} use Henderson.

$$pH = pK_a + \log \frac{[\text{base}]}{[\text{acid}]}$$

$$= 4.74 + \log \frac{0.9}{0.49}$$

1.84

$$pH = 5$$

صحة نفس ايجاد
التي Capacity

Buffer 2

تركيز طبق

$$[A^-] = 0.09 \text{ M}$$

$$[HA] = 0.049 \text{ M}$$

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

$$= 4.74 + \log \frac{[0.09]}{[0.049]}$$

1.84

$$= 4.74 + \log 1.84$$

$$pH = 5$$

two buffers have the same value of pH.

Buffer capacity can be determine after adding H^+ or OH^-

If we add 0.04 mol NaOH to 1L of buffer

عبارة عن قابلية صوف تتفاعل مع الهيدروجين .

Buffer 1

$$[A^-] = 0.9 + 0.04 = 0.94$$

سيزداد عدد مولات القاعدة لاننا اضفنا قاعدته .

$$[HA] = 0.49 - 0.04 = 0.45$$

حين عدد مولات الهيدروجين بعد مولات القاعدة المضافة .

Buffer 2.

$$[A^-] = 0.09 + 0.04 = 0.13$$

$$[HA] = 0.049 - 0.04 = 0.009$$

* قبل اضافة القاعدة كانت لـ $\frac{[base]}{[acid]}$ نسبة لكن الان اختلفت

$$pH = pK_a + \log \frac{base}{acid}$$

$$= 4.75 + \log \frac{0.94}{0.45}$$

has a higher buffer capacity

$$pH = 5.06$$

$$pH = 4.75 + \log \frac{0.13}{0.009}$$

$$pH = 5.9$$

قبل اضافة كانت $pH = 5$ وكلاهما بعد اضافة القاعدة / الاخر زاد بعد اضافة 0.9 يعني بان الهيدروجين الاول قاوم التغير / زاد لـ 0.06 اما

Buffering mechanism

- For a mixture of a weak acid and its salt $\text{pH} = \text{pK}_a + \log \left(\frac{[\text{A}^-]}{[\text{HA}]}\right)$ the pH is governed by $\log \left(\frac{[\text{A}^-]}{[\text{HA}]}\right)$
weak base or fits salt
- Upon dilution the ratio stays constant so the pH does not change
لا يتغير pH بتخفيف المحلول
- Upon addition of a strong acid or strong base, the change will be small compared to unbuffered solutions
 $\Delta \text{pH}_{\text{unbuffered}} > \Delta \text{pH}_{\text{buffered}}$ [resist the Δ] *والا يتغير باضافة*
- The buffering capacity increases with the conc. Of $[\text{A}^-]$ and $[\text{HA}]$
كما زاد تركيز buffer capacity تزيد
- The buffering capacity is maximum at $\text{pH} = \text{pK}_a$
- The buffering capacity is satisfactory over a pH range of $\text{pK}_a \pm 1$

$\text{pH} = \text{pK}_a + \log \frac{\text{base}}{\text{acid}}$
 $\log [1] = 0$
base = acid



Example

buffer
weak acid
salt of weak acid

pH before and
M after addi.

A buffer solution 0.2 M HOAc and NaOAc. Calculate the change in pH upon addition of 1 mL 0.1 M HCl to 10 mL of this solution ($K_a = 1.75 \times 10^{-5}$)

$\Delta \text{pH} = \text{pH}_{\text{after}} - \text{pH}_{\text{before}}$
start in this

At the beginning $[\text{HOAc}] = [\text{NaOAc}]$
 $\text{pH} = \text{pK}_a + \log \left(\frac{[\text{OAc}^-]}{[\text{HOAc}]} \right) \rightarrow \text{pH} = \text{pK}_a = 4.76$

When adding HCl it will react with OAc^-

$0.2 \times 10 = 2 \text{ mmol HOAc and NaOAc}$

$0.1 \times 1 = 0.1 \text{ mmol HCl that will react with OAc}^- \text{ to produce more HOAc}$

After the addition of HCl: $2 + 0.1 = 2.1 \text{ mmol HOAc}; 2 - 0.1 = 1.9 \text{ mmol OAc}^-$

Total volume 11 mL



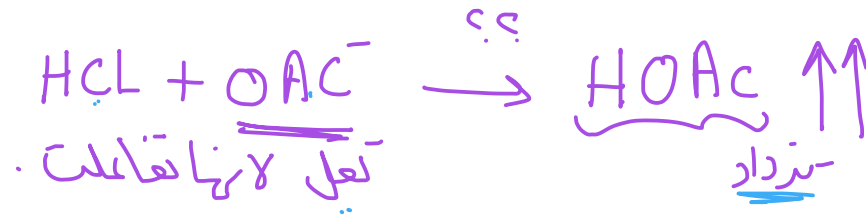
* before adding Acid.

$$pH = pK_a + \log \frac{[base]}{[acid]}$$

$$= -\log 1.75 \times 10^{-5} + \log \left[\frac{0.2}{0.2} \right]$$

$pH = 4.76$

* After adding Acid → Acid will increase.
base will decrease.



* 1 mL / 0.1 M HCl → # of moles HCl = 1 * 0.1 = 0.1 mmol

10 mL / 0.2 M

→ # of moles HOAc / NaOAc = 10 * 0.2 = 2 mmol
acid base

After the addition

→ acid increases. (HOAc)

→ 2 + 0.1 = 2.1 mmol

base will decrease (OAc⁻)

→ 2 - 0.1 = 1.9 mmol

+ $pH_{after} = pK_a + \log \frac{base}{acid}$

= 4.76 + log [1.9 / 2.1]

= 4.71

ΔpH = 4.71 - 4.76

= -0.05

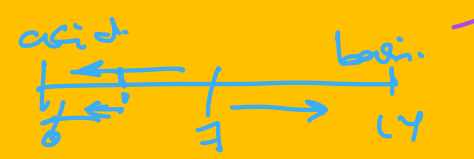
الأس الهيدروجيني سيبتنا

أن PH قلنا

الأس الهيدروجيني acid كاننا

Solution...Cont.

0.1 0.1



$$\text{pH} = 4.76 + \log \left(\frac{[1.9]}{[2.1]} \right) = 4.71$$

Change in pH is -0.05

Done. 17



Example

weat NH_3 (base) mass buffer solution NH_4Cl

Calculate the volume of conc. ammonia and the weight (g) of ammonium chloride you need to add to prepare 100 mL buffer at pH 10 if the final conc. of salt is 0.2 M. $[\text{NH}_3] = 14.8\text{M}$, $K_b = 1.75 \times 10^{-5}$, (NH_4Cl 53.5 g/mol).

① $V = 100\text{ mL}$ Final Con Salt = 0.2 M
 $\text{pH} = 10$
 * of moles $\text{NH}_4\text{Cl} = 0.2 \times 100 = 20\text{ mmol}$
 weight $\text{NH}_4\text{Cl} = \text{* moles} \times \text{Mwt}$
 $= 20 \times 10^{-3} \times 53.5 \frac{\text{g}}{\text{mol}}$
 $= 1.07\text{ g}$

$\text{pH} = \text{pKa} + \log \frac{\text{base}}{\text{acid}}$
 $K_a = \frac{K_w}{K_b} = \frac{1 \times 10^{-14}}{1.75 \times 10^{-5}} = 0.571 \times 10^{-9}$

$10 = 9.24 + \log \frac{[\text{NH}_3]}{[\text{NH}_4^+]}$
 $-9.24 \quad -9.24$
 $0.76 = \log \frac{[\text{NH}_3]}{0.2}$
 $\neq 10$
 $5.76 = \frac{[\text{NH}_3]}{0.2} \Rightarrow [\text{NH}_3] = 1.15\text{ M} = 1.15 \frac{\text{mol}}{\text{L}} = 1.15 \frac{\text{mmol}}{\text{mL}}$

الحجم النهائي هو 100 mL اذا كانا يظب
 NH₃ ؟؟؟

$\text{NH}_3 \Rightarrow 14.8\text{ M} \rightarrow V_1 M_1 = V_2 M_2$
 $V_1 \times 14.8 = 100 \times 1.15$



$$pK_a = 9.243$$

$$V_{NH_3} = 7.77 \text{ mL}$$

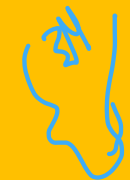
Example

Calculate the volume of conc. ammonia and the weight (g) of ammonium chloride you need to add to prepare 100 mL buffer at pH 10 if the final conc. of salt is 0.2 M. $[NH_3] = 14.8M$, $K_b = 1.75 \times 10^{-5}$, $(NH_4Cl \text{ } 53.5 \text{ g/mol})$.

Done



Example



Calculate the weight (g) NH_4Cl (53.5 g/mol) and the volume (mL) of 3M NaOH needed to prepare 500 mL buffer at pH 9.5 and salt conc. 0.1 M ($K_b = 1.75 \times 10^{-5}$)



$$\text{pH} = \text{pKa} + \log \frac{[\text{base}]}{[\text{acid}]}$$

$$9.5 = 9.24 + \log \frac{[\text{NH}_3]}{[\text{NH}_4^+]}$$

نفس مولات ما بين

$$9.5 = 9.24 + \log \frac{[\text{NH}_3]}{0.1}$$

$$[\text{NH}_3] = 0.18 \text{ M}$$

$$\# \text{ of mmol } \text{NH}_3 = 0.18 * 500 = 90 \text{ mmol}$$

$$\text{NH}_4\text{Cl} = 0.1 \rightarrow 0.1 * 500 = 50 \text{ mmol}$$

buffer solution

$$\text{Total mmol } \text{NH}_4\text{Cl} = 90 + 50 = 140 \text{ mmol}$$

$$\text{weight} = \# \text{ of mol} * \text{mwt} = 140 * 10^{-4} * 53.5$$

وزن مطلوب =

$$[\text{NaOH}] = \frac{90 \times 10^{-3}}{500 \times 10^{-3}} = 0.18 \text{ M}$$

نفس تسمى

$$M_1 V_1 = M_2 V_2$$

$$3 * V_1 = 0.18 * 500$$

$$V_1 = 30 \text{ mL}$$

هنا قمنا بالجمع وكانه التجميع عليه

كم كان لدينا من NH_4Cl

حتى تفاعل جزء منها مع NaOH واطل 90 mmol من NH_3

ونبقى بالحلول 50 mmol

$$90 + 50 = 140$$



Example

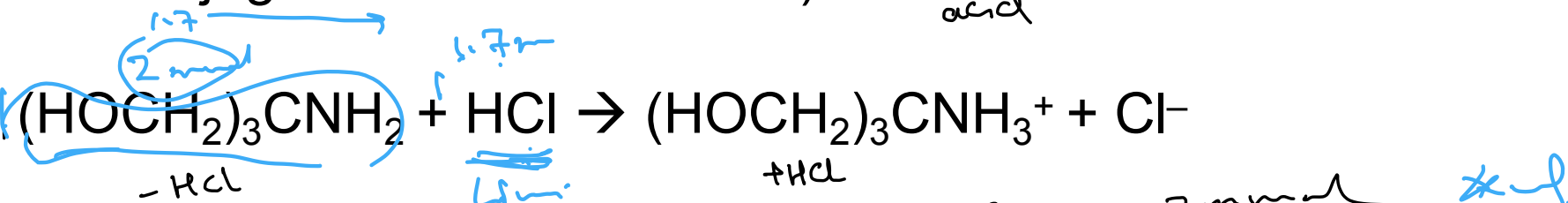
Calculate the weight (g) NH_4Cl (53.5 g/mol) and the volume (mL) of 3M NaOH needed to prepare 500 mL buffer at pH 9.5 and salt conc. 0.1 M ($K_b = 1.75 \times 10^{-5}$)

Done

Example

weak
base.

Calculate the pH of a buffer solution prepared by dissolving 242.2 mg of $((\text{HOCH}_2)_3\text{CNH}_2$ = 121.1 g/mol) in 10 mL of 0.17 M HCl and diluting to 100 mL with H_2O (pKa = 8.08 for the conjugate acid of the amine) strong acid



$$\# \text{ of moles amine} = \frac{\text{mass}}{\text{mwt}} = \frac{242.2}{121.1} = \underline{2 \text{ mmol}}$$

$$\# \text{ of mol HCl} = 0.17 \times 10 = \underline{1.7 \text{ mmol}}$$

$$\# \text{ of mmol amine after react} = 2 - 1.7 = \underline{0.3 \text{ mmol}}$$

UU

$$\text{pH} = \underbrace{\text{pKa}}_{8.08} + \log \frac{0.3}{1.7} = \underline{7.27}$$



Example

Calculate the pH of a buffer solution prepared by dissolving 242.2 mg of $(\text{HOCH}_2)_3\text{CNH}_2$ ($M = 121.1 \text{ g/mol}$) in 10 mL of 0.17 M HCl and diluting to 100 mL with H_2O ($\text{pK}_a = 8.08$ for the conjugate acid of the amine)



Done

