تفريغ فيزيكال

اسم الموضوع:
Chapter 5
Ionic equilibria

إعداد الصيدلاني/ـة: حلا سامر 🌞

Ionic equilibria

بسم الله الرحمن الرحيم

Introduction

Many drugs are:

- Weak acids such as acetylsalicylic acid and ibuprofen مضاد التهاب
- Weak bases such as procaine and lidocaine مخدر موضعي
- Salts such as sodium diclofenac مضاد للالتهاب and metformin hydrochloride → لعلاج مرض السكري

أمثلة فقط

Ionization of drug is important in:

ليش عملية تأين الدواء جدًا مهمة ؟!

- Formulation: ionized drug is more soluble
- Absorption: unionized drugs easily diffuse across membrane.
- Distribution: unionized drugs have high volumes of distribution and protein binding
- Excretion: ionized drugs excreted more readily.

بستنتج انه درجة تأين الدواء يلعب دور هام في تحديد كيفية تصرف الدواء.

① الأدوية المتأينة تكون أكثر قابلية للذوبان في السوائل الجسم مما يسهل تحضير الأدوية بشكل محلول / سائل.

② من ناحية الامتصاص:
فالأدوية الغير المتأينة تعبر بسهولة الغشاء (لانه غالبا متكون من طبقات دهنية)

③ من ناحية التوزيع:
الأدوية الغير متأينة تنتشر في الجسم بشكل أكبر و(تنتقل الى أماكن أكثر)

④ من ناحية الاخراج:
فالأدوية المتأينة تخرج بشكل اسهل عن طريق الكلى لأنها أكثر قابلية للذوبان في البول فيتم ترشيحها بسهولة واخراجها.

① Arrhenius Theory

الحمض يحرر (H^+) عند ذوبانه
القاعدة تحرر (OH^-) =

- Arrhenius defined an acid as a substance that liberates hydrogen ions, H^+ and a base as a substance that supplies hydroxyl ions, OH^- on dissociation.

مفهوم ارهينيوس محدود التعريف

- However; Arrhenius definition could not explain the basic behavior of many compounds that do not contain hydroxyl ions, OH^- (e.g.NH_3) ما يقدر يفسر سلوك المركبات التي لا تحتوي على (OH) مثل

$$\underset{\text{قاعدة}}{NH_3} + \underset{\text{حمض}}{HCl} \rightarrow NH_4^+ + Cl^-$$

- Therefore; the Brönsted–Lowry theory is more useful than the Arrhenius theory for the representation of ionization in both aqueous and non-aqueous systems.

لهذا السبب ← برونستد لوري > ارهينيوس

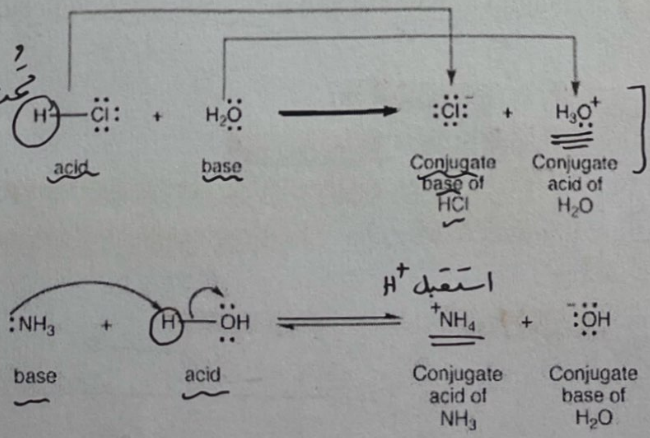
حتى نعبر عن التأين من
aqueous soln.
non aqueous soln.

② Brönsted–Lowry Theory

الحمض يمنح/يتبرع (H^+)
القاعدة: تستقبل (H^+)

- According to the Brönsted–Lowry theory, an acid is a substance that is capable (القدرة) of donating a proton, and a base is a substance that is capable of accepting a proton from an acid.

تمنح للقاعدة



استقبل #

استقبل H^+

Brönsted–Lowry Theory

قوة الحمض / القاعدة
تعتمد على أو تقاس
بمدى قدرة المادة
على التبرع بـ H^+
أو استقبالها.

- The relative strengths of acids and bases are measured by the tendencies of these substances to give up and take on protons:
  - HCl is a strong acid in water because it gives up its proton readily — يمنح H^+ بسهولة
  - CH_3COOH is a weak acid because it gives up its proton only to a small extent. — يمنح H^+ بشكل جزئي (لا يذوب بشكل كامل)

  و بشكل كامل عندما يذوب في الماء.

  حمض ضعيف / قليلة التأين

- The strength of an acid or a base varies with the solvent:
  - HCl is a weak acid in glacial acetic acid
  - CH_3COOH is a strong acid in liquid ammonia.
- i.e. the strength of an acid depends not only (1) on its ability to give up a proton but also (2) on the ability of the solvent to accept the proton from the acid.

عاملين
2 factor

قوة الحمض/القاعدة
لا تعتمد فقط على
قدرة المادة بالتبرع او
استقبال (H^+)
بل أيضاً تعتمد على (المذيب solvent)
يعني من solvent لـ solvent تختلف قوة الحمض/القاعدة

① قوة الحمض / القاعدة
كمنح / استقبال لل H^+
② قدرة الـ solvent
على استخراج H
من الحمض

Brönsted–Lowry Theory

- Solvents can be classified as protophilic, protogenic, amphiprotic, and aprotic.

➢ **Protophilic** or basic solvent is one that is capable of accepting protons from the solute (e.g. liquid ammonia **NH_3**). → لها ميول قوية لاستقبال H^+ (basic solvent)

➢ **Protogenic** solvent is a proton donating compound (e.g. acetic acid **CH_3COOH**) → لها ميول قوية لمنح H^+ (acidic solvent)

➢ **Amphiprotic** solvents act as both proton acceptors and proton donor (e.g. water **H_2O**). → مرة حمض ومرة قاعدة

➢ **Aprotic** solvents neither accept nor donate protons (e.g. methane **CH_4**).

لا تمنح ولا تستقبل

لا حمض ولا قاعدة
يعني ما بتشارك
باللعبة أبداً

③ Lewis Electronic Theory

According to the Lewis theory:

- An acid is a molecule or an ion that accepts an electron pair to form a covalent bond.
- A base is a substance that provides the pair of unshared electrons to coordinate with an acid.

الحمض (جزيء / ايون) يقبل زوج من (e^-) اما القاعدة تُقدم زوج من (e^-)

- Certain compounds such as BF_3 are considered acids even when they are not proton donors (do not contain hydrogen).

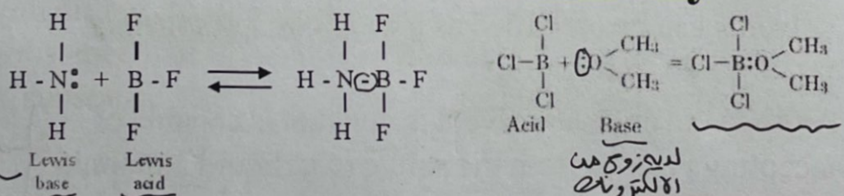
ما عنده H^+ حتى يتبرع فيه، لكنه يعتبر حمض لانه يستقبل ازواج الالكترونات بدلاً من التبرع.

- Other compounds such as ethers and NH_3 are considered bases even when they do not accept proton.

بنعتبرهم Base حتى لو ما عندهم القدرة يستقبلوا H^+ بس بتبرعوا بالالكترونات غير المرتبطة ويكونوا روابط تساهمية مع الاحماض.

Lewis Electronic Theory

لديه زوج من الالكترونات (unshared)



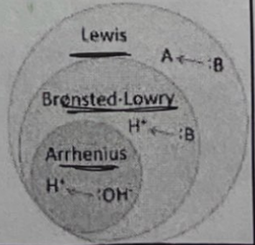
لديه زوج من الالكترونات

- The Lewis systems is probably too broad for convenient application to ordinary acid-base reactions.
- These reactions can be described as a form of electron sharing rather than as acid-base reactions.

أكثر من

مفهوم لويس

Arrhenius < Bronsted < Lewis



Ionization of Weak Electrolytes

Weak Acids

The ionization of an uncharged weak acid, HA, in water:

$$\underset{\text{Acid 1}}{HA} + \underset{\text{Base 1}}{H_2O} \leftrightarrow \underset{\text{Acid 2}}{H_3O^+} + \underset{\text{Base 2}}{A^-}$$

conjugate acid-base pair

قاعدة وحمضها المرافق

او حمض و قاعدته المرافقة

مثل HA , A⁻
حمض القاعدة ما المرافقة

The *acidity constant* K_a is expressed as:

$$K_a = \frac{[H_3O^+][A^-]}{[HA]}$$

Note: An acid and a base in equilibrium is termed a conjugate acid-base pair. E.g. A^- is the conjugate base of the weak acid HA

For a charged acid, BH^+, the reaction is written as:

$$\underset{\text{Acid 1}}{BH^+} + \underset{\text{Base 1}}{H_2O} \leftrightarrow \underset{\text{Acid 2}}{H_3O^+} + \underset{\text{Base 2}}{B}$$

بدون شحنة

The acidity constant K_a is expressed as :

$$K_a = \frac{[H_3O^+][B]}{[BH^+]}$$

عند التأين تصبح

uncharged acid → القاعدة -1 (A^-)
الحمض +1 (H_3O^+)

charged acid → الحمض +1 (H_3O^+)
القاعدة (الشحنة -1) (B)
مرضا 2 - 1 = 1

1 - 1 = 0 ماعليها شحنة تصبح

يعني
$BH^+ \rightarrow B$
القاعدة شو شحنتها؟
0 = 1 - (1+)
من BH^+
بالتالي عشان هيك ماعليها شحنة.

Ionization of Weak Electrolytes

Weak Bases

The ionization of an uncharged weak base, B, in water can be written as:

الحمض المرافق القاعدة المرافقة

$$\underset{\text{Base 1}}{B} + \underset{\text{Acid 1}}{H_2O} \leftrightarrow \underset{\text{Base 2}}{OH^-} + \underset{\text{Acid 2}}{BH^+}$$

uncharged base
القاعدة (OH) ← -1
الحمض (BH^+) ← +1

The *basicity constant* K_b is expressed as :

$$K_b = \frac{[OH^-][BH^+]}{[B]}$$

النواتج
المتفاعلات

For anionic base, A^-, the reaction is written as:

$$\underset{\text{Base 1}}{A^-} + \underset{\text{Acid 1}}{H_2O} \leftrightarrow \underset{\text{Base 2}}{OH^-} + \underset{\text{Acid 2}}{HA}$$

Base لما تكون مشحونة

The basicity constant K_b is expressed as :

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

القاعدة (OH) ← -1
الحمض (HA) ←
[الشحنة +1]
الشحنة الاصلية (-1)
-1 + 1 = 0

عشان هيك ماعليها شحنة

Ionization of salts

ناتجة عن علاقة التعادل بين الحمض والقاعدة

- Salts are the non-water product of an acid base neutralization. (H_2O مش فيها)
- Drug salts are often used due to their complete ionization, and thus better aqueous solubility than weak acids and bases.

قدرة الاملاح على التأين الكامل يعزز ذوبانها في السوائل لذلك تستخدم في الادوية

- Depending on the strength of the acid and base that form the salt, there are four possible types:

تصنف الاملاح الى ٤ اصناف حسب قوة الحمض والقاعدة .

1. Salt of strong acid and a strong base (e.g. NaCl)

- This salt dissociates to give ions that practically do not consume or release protons.
- E.g. NaCl → $Na^+ + Cl^-$

  Salt (Practically neither Acid or Base)

لما يتفكك ما بنتج عندي حمض او قاعدة

Ionization of salts

2. Salt of weak acid and strong base (e.g. NaOAc)

This salt dissociates into ions; from which one acts as a base and consumes a proton to give its conjugate weak acid:

E.g. $\underset{\text{Salt}}{NaOAc} \rightarrow Na^+ + \underset{\text{Base1}}{\boxed{OAc^-}}$

$\underset{\text{Base1}}{OAc^-} + \underset{\text{Acid1}}{H_2O} \leftrightarrow \underset{\text{Acid2}}{HOAc} + \underset{\text{Base2}}{OH^-}$

لما يتفكك هاد النوع من الاملاح بنتج عندي قاعدة تتفاعل مع الماء لتعطي الحمض المرافق

3. Salt of weak base and strong acid (e.g. NH_4Cl)

This salt dissociates into ions; from which one acts as a n acid and releases a proton to give its conjugate weak base:

E.g. $\underset{\text{Salt}}{NH_4Cl} \rightarrow \underset{\text{Acid1}}{\boxed{NH4^+}} + \boxed{Cl^-}$

$\underset{\text{Acid1}}{NH4^+} + \underset{\text{Base1}}{H_2O} \leftrightarrow \underset{\text{Base2}}{NH_3} + \underset{\text{Acid2}}{H3O^+}$

لما يتفكك هاد النوع من الاملاح بنتج عندي حمض وتتفاعل مع الماء لتعطيني القاعدة المرافقة

طب سؤال ؟؟

ليه هي NH_4^+ الي بتتفاعل مع H_2O مش Cl^- ؟؟

لانه NH_4^+ مشتقة من قاعدة ضعيفة NH_3

اما Cl^- مشتقة من حمض قوي HCl فهو مستقر

يعني مارح تتفاعل وغير نشط

ايضا في ملح NaOAc لما يتفكك Na^+ مارح تتفاعل مع الماء لانها مشتقة من قاعدة قوية فما بتتفاعل مع

Ionization of salts

4. Salt of weak acid and weak base (e.g. NH_4OAc)

This salt dissociates into ions from which one acts as an acid and the other as a base.

E.g.	NH_4OAc	$\rightarrow$			NH_4^+	+	OAc^-	
	Salt				Acid 1		Base 1	
	NH_4^+	+	H_2O	$\leftrightarrow$	NH_3	+	H_3O^+	
	Acid 1		Base 1		Base 2		Acid 2	
	AcO^-	+	H_2O	$\leftrightarrow$	$AcOH$	+	OH^-	
	Base 1		Acid 1		Acid 2		Base 2	

كما يتفكك هون
رح يطيني
حمض وقاعدة
قاعدة مرافقة — حمض مرافق

Ionization of Polyprotic Electrolytes

Polyprotic acids can lose more than one H^+ ion. (H^+ فقدان أكثر من)

E.g. *Diprotic acids*, such as H_2SO_4 and H_2CO_3 which release $\boxed{\text{2 protons}}$, and *Triprotic acids*, such as H_3PO_4 which releases $\boxed{\text{3 protons}}$

Consider the ionization of the weak diprotic acid, H_2CO_3 that dissociates in two steps:

$$H_2CO_3 + H_2O \leftrightarrow H_3O^+ + HCO_3^-$$

$$HCO_3^- + H_2O \leftrightarrow H_3O^+ + CO_3^{2-}$$

Two acidity constants is used to describe the two equilibrium:

$$\boxed{K_{a1}} = \frac{[H_3O^+][HCO_3^-]}{[H_2CO_3]}, \quad \boxed{K_{a2}} = \frac{[H_3O^+][CO_3^{2-}]}{[HCO_3^-]}$$

K_{a1} is larger than K_{a2} because the polyprotic acid lose its first proton more easily than the second (and third) proton.

$K_{a1} > K_{a2}$

فقدان H^+ الاول دائ
يكون اسهل من
فقدان H^+ الثاني

Scanned with CamScanner

Ionization of Polyprotic Electrolytes

Polyprotic bases can accept more than one H^+ ion.

E.g. *Diprotic bases*, such as CO_3^{2-} which accepts 2 protons, and *Triprotic bases*, such as PO_4^{3-} accepts 3 protons.

Consider the ionization of the weak diprotic base, CO_3^{2} that consumes protons in two steps:

$CO_3^{2-} + H_2O \leftrightarrow OH^- + HCO_3^{1-}$ $-2+1=-1$

$HCO_3^{1-} + H_2O \leftrightarrow OH^- + H_2CO_3$ $-1+1=0$

Two basicity constants is used to describe the two equilibrium:

$$K_{b1} = \frac{[OH^-][HCO_3^-]}{[CO_3^{2-}]}, \qquad K_{b2} = \frac{[OH^-][H_2CO_3]}{[HCO_3^-]}$$

K_{b1} is larger than K_{b2} because the polyprotic base consumes its first proton more easily than the second (and third) proton.

$K_{b2} < K_{b1}$

لانه اكتساب البروتون الاول
يكون اسهل من البروتون الثاني.

Ionization of Water

Water ionizes slightly (تشكل ضعيف) to yield hydrogen (H^+) and hydroxyl (OH^-) ions by reacting with another molecule of water (*autoprotolytic* reaction):

$H_2O + H_2O \leftrightarrow OH^- + \boxed{H_3O^+}$

The equilibrium constant is expressed as:

$$K = \frac{[OH^-][H_3O^+]}{[H_2O]^2} \Rightarrow K[H_2O]^2 = [OH^-][H_3O^+]$$

[H2O]² is considered as a constant and is combined with K to give a new constant, K_w, known as the *autoprotolysis constant*, or the *ion product* of water:

$K_w = [OH^-][H_3O^+]$

$1 \times 10^{-14} = [OH^-] \times [H_3O^+]$

Scanned with CamScanner

Ionization of Water

[H₃O⁺]=[OH⁻] ←

In *pure* water: $[OH^-] = [H_3O^+] = 1 \times 10^{-7}$ M at 25°C.

$K_w = [OH^-][H_3O^+] = (1 \times 10^{-7}) \times (1 \times 10^{-7}) = 1 \times 10^{-14}$ M

When an acid is added to pure water, the increase in hydrogen ions is offset by a decrease in the hydroxyl ions so that K_w remains constant at about 1×10^{-14} M at 25°C.

لما بضيف الحمض للماء النقي تزداد H^+ وتقل OH^- يعني الزيادة يقابلها نقصان (نفس المقدار) فبالتالي K_w ثابت. 1×10^{-14} الزيادة = النقصان حتى يبقى K_w ثابت

A simple relationship exists between K_a of a weak acid (**HB**) and K_b of its conjugate base (**B−**), and between K_a of **BH⁺** and K_b of **B** when the solvent is amphiprotic (e.g. water).

$$\underbrace{K_a}_{\text{weak acid}}\underbrace{K_b}_{\text{conjugate Base}} = K_w$$

في محلول مائي

1M of strong acid → $[H_3O^+] = 1$ عالي جداً

″ ″ ″ base → $[H_3O^+] = 1 \times 10^{-14}$ منخفض جداً

pH

- $[H_3O]^+$ varies from 1 (in a 1 M solution of a strong acid) to 1×10^{-14} (in a 1 M solution of a strong base).
- Sorensen suggested a simplified method of expressing (طريقة) $[H_3O]^+$ via the term pH.
- pH is defined as the negative logarithm of $[H_3O]^+$

$$pH = -\log [H_3O]^+$$ → طريقة حتى نعبر عن $[H_3O^+]$

- The pH of a solution is a numeric scale from 0 to 14, which expresses the degree of acidity (7-0) and alkalinity (7 - 14).

acidity ← (7 − 0)
alKalinity ← (14 − 7)

- The value 7 at which $[H_3O]^+ = [OH]^-$ at room temperature is referred to as the neutral point

مراجعة

Mathematical revision

Exponential Laws	Logarithm Laws
$x^a \cdot x^b = x^{a+b}$	$\log(ab) = \log(a) + \log(b)$
$\dfrac{x^a}{x^b} = x^{a-b}$	$\log\left(\dfrac{a}{b}\right) = \log(a) - \log(b)$
$(x^a)^b = x^{ab}$	$\log(a^b) = b \cdot \log(a)$
$x^{-a} = \dfrac{1}{x^a}$	$\log_x\left(\dfrac{1}{x^a}\right) = -a$
$x^0 = 1$	$\log_x 1 = 0$

pK and pOH

- The term "p" is also used to express the negative logarithm of each of $[OH^-]$, K_a, K_b, K_w as pOH, pK_a, pK_b, and pK_w

$$\boxed{pK_w} = pH + pOH$$
$$pK_w = pK_a + pK_b$$

- pk_a and pK_b values provide a means of comparing the strengths of weak acid and weak bases:
- Lower pk_a values correspond to stronger acids
- Lower pK_b values correspond to stronger Bases

للتعبير عن قوة الحمض/القاعدة.

$H_3O^+ \rightarrow pH$
$OH^- \rightarrow pOH$
$K_a \rightarrow pK_a$
$K_b \rightarrow pK_b$
$K_w \rightarrow pK_w$

$\downarrow pK_a \rightarrow \uparrow K_a \rightarrow$ stronger acid

$\downarrow pK_b \rightarrow \uparrow K_b \rightarrow$ stronger bases

Calculation of pH

Strong Acids and Bases

- A strong acid, HA, ionizes completely to H_3O^+ and A^-. Therefore $[H_3O^+]$ =[HA] and pH is calculated as:

$$\mathbf{pH = -log\ [HA]} = -\log [H_3O^+]$$

- While a strong base, B ionizes completely to BH+ and OH-. Therefore [OH-] = [B] and pOH is calculated as:

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

Since pH = pKw – pOH

Then: pH = pKw – (- log [B]) or:

$$\mathbf{\boxed{pH} = \boxed{pKw} + log\ [B]}$$

في حالة strong

base acid

ركزوا منيح

Calculation of pH

Weak Acids and Bases

A weak acid, HA, ionizes partially H_3O^+ and A^-:

$$K_a = \frac{[H_3O^+][A^-]}{[HA]} \quad \text{Since } [A^-] = [H_3O^+]$$

بنفترض

$$\text{Then} \quad K_a = \frac{[H_3O^+]^2}{[HA]}$$

$$\sqrt{[H_3O^+]^2} = \sqrt{K_a\ [HA]} \quad \Rightarrow \quad [H_3O^+] = (K_a\ [HA])^{1/2}$$

→ pH = p(Ka [HA

$$pH = \frac{1}{2}(pK_a - \log[HA])$$

Calculation of pH

Weak Acids and Bases

Example

Calculate the pH of a 50 mg mL^{-1} solution of ascorbic acid (MW 176.1, pK_a 4.17)

حمض ضعيف

Convert concentration to M (or mol/L)

$C = 50\ mg/mL = 50\ g/L$

$$\frac{50 \times 10^{-3}\ g}{1 \times 10^{-3}\ L}$$

$\boxed{50\ g/L}$

$M = 50 / 176.1 = \boxed{0.284\ mol/L}$ — $\frac{mol}{L}$ M

$$pH = \frac{1}{2}(pKa - \log[HA]) \rightarrow$$

استخدمنا هاد القانون لانه حمض ضعيف

معطى

$$pH = \frac{1}{2}(4.17 - \log 0.284)$$

$$pH = \frac{1}{2}(4.17 + 0.574) = 2.36$$

Weak base (B)

slide 24

$$K_b = \frac{[OH^-]^2}{[B]}$$

$$\sqrt{[OH^-]^2} = \sqrt{K_b [B]}$$

$$[OH^-] = (K_b[B])^{\frac{1}{2}}$$

بدي اخد P للطرفين

$$P(OH) = P\left((K_b[B])^{\frac{1}{2}}\right)$$

$$P(OH) = -\log(K_b[B])^{\frac{1}{2}}$$

$$P(OH) = -\frac{1}{2}\log\underline{(K_b[B])}$$ فك

مش مطلوب الاشتقاق بس للفهم

$$P(OH) = -\frac{1}{2}(\log K_b + \log[B])$$

$$P(OH) = +\frac{1}{2}PK_b - \frac{1}{2}\log[B]$$

بدي هسا PH ليش !؟ لانه بالامتحان بطلبني PH

+ بكون في المعطيات PKa كيف بدي احل وقتها ؟!

$$PK_w - PH = \frac{1}{2}(PK_w - PK_a) - \frac{1}{2}\log[B]$$

$$PH = \left(PK_w - \frac{1}{2}PK_w\right) + \frac{1}{2}PK_a + \frac{1}{2}\log[B]$$

$\frac{1}{2}PK_w$

$$PH = \frac{1}{2}(PK_w + PK_a + \log[B])$$

Calculation of pH

Weak Acids and Bases

Example

Calculate the pH of a saturated solution of codeine monohydrate (MW 317.4, pKa 8.2, solubility at room temperature is 1 g in 120 mL

اعطاني pKa يعني مضطر استخدم القانون الثالث

Convert concentration to mol/L

C = 1 g / 120 mL = 8.33 g/L

M = 8.33/317.4 = 0.0263 mol/L

$$pH = \frac{1}{2}(pK_w + pKa + \log[B])$$

$$pH = \frac{1}{2}(14 + 8.2 + \log 0.0263)$$

$$pH = \frac{1}{2}(14 + 8.2 - 1.58) = 10.31$$

Calculation of pH

Salts of Strong Acid and Strong Base

- The salt of strong acid and strong base (e.g. NaCl) dissociates in water into Na^+ and Cl^-.

$$NaCl \rightarrow Na^+ + Cl^-$$

ما رح يتفاعلوا مع H_2O

- Neither Na^+ nor Cl^- ions are capable of consuming or releasing protons from /to water (they are neither acids nor bases)

Na^+, Cl^- ما رح يأثروا على pH للمحلول water

فرح ضل $\boxed{7 = pH}$

- Therefore these ions have no effect on pH. The pH of the solution remains the same as that of pure water, 7.

PH = 7

رح ضل 7

لانه مالهم تأثير

Calculation of pH

Salts of Weak Acid and Strong Base

The salt of weak acid and strong base (e.g. Sodium acetate, AcONa (designated as S)) dissociates into Na^+ and AcO^-.

$AcONa \rightarrow AcO^- + Na^+$

AcO^- acts as a base and consumes one proton to form AcOH

$AcO^- + H_2O \leftrightarrow AcOH + OH^-$

The pH is calculated in the same way as in weak base:

$$pH = \frac{1}{2}(pK_w + pKa + \log[AcO^-])$$

Since $[AcO^-] = [S]$

Then $$pH = \frac{1}{2}(pK_w + pKa + \log[S])$$

Note: the pH is always > 7

لأنه تأثير القاعدة
على المحلول

Calculation of pH

Salts of Weak Base and Strong Acid

The salt of weak base and strong acid (e.g. ammonium chloride, NH_4Cl (designated as S)) dissociates into NH_4^+ and Cl^-.

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

NH_4^+ acts as an acid and releases one proton to form NH_3

$NH_4^+ + H_2O \leftrightarrow NH_3 + H_3O^+$

The pH is calculated in the same way as in weak acid:

$$pH = \frac{1}{2}(pK_a - \log[NH_4^+])$$

Since $[NH_4^+] = [S]$

Then $$pH = \frac{1}{2}(pKa - log[S])$$

Note: the pH is always < 7

لأنه تأثير الحمض
على المحلول

Calculation of pH

Salts of Weak Acid and Weak Base

The salt of weak acid and weak base (e.g. ammonium acetate, $AcONH_4$) dissociates into NH_4^+ and AcO^-.

$AcONH_4 \rightarrow AcO^- + NH_4^+$

NH_4^+ acts as an acid and releases one proton to form NH_3, while AcO^- acts as a base and consumes one proton to form AcOH

$NH_4^+ + H_2O \leftrightarrow NH_3 + H_3O^+$

$AcO^- + H_2O \leftrightarrow AcOH + OH^-$

The pH can be calculated by:

قانون جديد

$$pH = \frac{1}{2}(pK_w + pKa - pK_b)$$

Note: the pH does not depend on the concentration of the salt, but rather depends on the strength of the weak acid and weak base

Common Ion

Calculation of pH

Weak Acids and their Salts

When a weak acid and a salt of that acid exist in solution (e.g., acetic acid and sodium acetate), both compound dissociate to give the conjugate base of that acid (in this case OAc^-).

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

$NaOAc \rightarrow Na^+ + OAc^-$

الايون المشترك

OAc^- in this case is called *a common ion.*

Most OAc^- will come from the salt NaOAc, therefore;

$[OAc^-] = [NaOAc]$ (designated as [S])

[تركيز الملح] = [تركيز القاعدة]

= [تركيز الحمض]

The pH is calculated by the following equation:

$$pH = pKa + \log\frac{[S]}{[HA]}$$

تركيز الملح / تركيز الحمض

The solution above is considered a *buffer* solution, and the equation above is named *Henderson-Hasselbalch* equation for buffers of weak acids.

عند اضافة حمض ضعيف الى الملح الذي يكون من نفس الحمض

يزداد الايون المشترك وتقل H_3O^+ وتزداد PH يعني (تقل الحموضة) وتزداد القاعدية

(*) لذلك هاي المحاليل بنستخدمها (Buffer soln.)

حتى تنظم PH المحلول (لانه بقدر الحكم في تأيين الحمض)

Calculation of pH

Weak Acids and their Salts

Example

What is the pH of a solution containing acetic acid 0.3 M and sodium acetate 0.05 M? (Ka for acetic acid = 1.75×10^{-5})

$$pK_a = -\log K_a$$

$$pK_a = -\log 1.75 \times 10^{-5} = 4.76$$

$$pH = pKa + \log \frac{[S]}{[HA]}$$

قانون ⊛ common Ion

$$pH = 4.76 + \log \frac{0.05}{0.3} = 3.98$$

حمضي

تركيز الملح / تركيز الحمض

⊛ هسا في شغلة بدنا نفهمها
كا نيجي نعوض في المعادلة
المفروض نعوض log [الايون المشترك] / [الحمض]

اغلب الايون المشترك بكون جاي
من الملح ف تركيز الايون المشترك
من الحمض نفسه كمهمل
فرح نعوض [الايون المشترك] من الملح
والي هو نفسه
تركيز الملح .

معلومة :
الملح يتأين بشكل كامل
في العادة لذلك
[الملح] = [القاعدة] = [الحمض]
$NaCl \rightarrow Na^+ + Cl^-$
$[NaCl] = [Cl^-] =$ [مشترك]

Calculation of pH

Weak Bases and their Salts

When a weak base and a salt of that base exist in solution (e.g., NH_3 and NH_4Cl), both compound dissociate to give the conjugate acid of that base (in this case NH_4^+).

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

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

NH_4^+ in this case is called *a common ion*.

Most NH_4^+ will come from the salt NH_4Cl, therefore;

$[NH_4^+] = [NH_4Cl]$ (designated as [S])

لذلك تركيز NH_4^+ هو نفسه تركيز الملح

The pH is calculated by the following equation:

$$pH = pKa + \log \frac{[B]}{[S]}$$

القاعدة — الايون المشترك

The solution above is considered a *buffer* solution, and the equation above is named *Henderson-Hasselbalch* equation for buffers of weak bases.

Scanned with CamScanner

Calculation of pH

Weak Bases and their Salts

Example

What is the pH of a solution containing ephedrine 0.1 M and ephedrine HCl 0.01 M? (K_b for ephedrine = 2.3×10^{-5})

$$pK_b = -\log K_a$$
$$pK_b = -\log 2.3 \times 10^{-5} = 4.64$$
$$pK_a = pKw - pKb$$
$$pK_a = 14 - 4.64 = 9.36$$
$$pH = pKa + \log\frac{[B]}{[S]}$$
$$pH = 9.36 + \log\frac{0.1}{0.01} = 10.36$$

$PKw = PKa + PKb$

$14 = PKa + 4.638$

$\boxed{9.36 = PKa}$

$PH = 9.36 + \log \frac{\text{تركيز القاعدة}}{\text{تركيز الملح (الأيون المشترك)}}$

$PH = 9.36 + \log \frac{0.1}{0.01}$

قاعدي $\boxed{PH = 10.36}$

اخر موضوع

Calculation of pH

Diprotic Acids and Bases

For weak diprotic acid, the $[H_3O^+]$ mostly comes from the first step of dissociation. Therefore; The second step is ignored during calculation of pH:

$$pH = \frac{1}{2}(pKa - \log[HA])$$

pH is calculated the same way as with monoprotic weak acid

بنحسب pH زي ما كنا نحسب Weak acid

For weak diprotic base, the $[OH^-]$ mostly comes from the first step of reaction. Therefore; The second step is ignored during calculation of pH:

$$pH = \frac{1}{2}(pK_w + pKa + \log[B])$$

pH is calculated the same way as with monoprotic weak base

بنحسب pH زي ما كنا نحسب Weak base

ادعولي ♡