

Bio-Chem

Lecture 1
-intro

★ pH, Buffer and Bonds
14/oct/2024

Logarithms: $b^x = a$
 $\hookrightarrow \log_b a = x$

ex 10 goes with a 10
 $\hookrightarrow \log_{10} 10^3 = 3$

$\log 1 = \text{zero}$

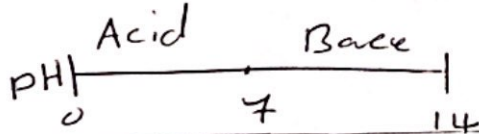
Acid \rightarrow releases a Proton (H^+)

Base \rightarrow accepts a Proton (H^+)

pH \rightarrow concentration of H^+

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

using pH is easier than using $[H^+]$ alone



$[H^+] \uparrow \rightarrow \downarrow pH \rightarrow \uparrow \text{Acidity}$

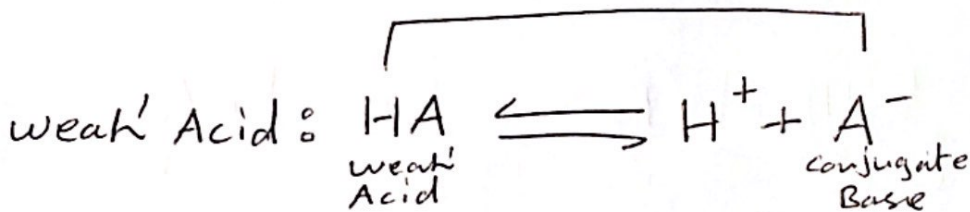
$[H^+] \uparrow = \uparrow \text{Acidity}$

* The relationship between any value and its p(value) is inverse

Dissociation constants (K_a)

تساوي ثابت التفكك

The ~~excess~~ tendency of any acid (HA) to lose a proton and form its conjugate base

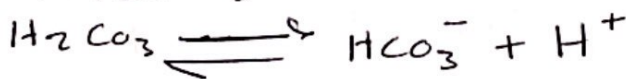


Positive ion: cation
 Negative ion: anion

The stronger the Acid, the greater its tendency to lose its protons

Strong Acids like HCl dissociate completely in solutions $\rightarrow \text{HCl} \dots$

Weak Acids do not completely dissociate



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

$\uparrow \text{Acidity} = \uparrow [\text{H}^+] = \downarrow \text{pH} = \uparrow K_a$

Hence

ex: on HCl
 $\downarrow \text{HCl} \rightarrow \uparrow \text{H}^+ + \uparrow \text{Cl}^-$
 $\sim \text{zero} \quad \sim 100\% \dots$
 $K_a = \frac{100\%}{\text{zero}} = \infty$

(2)

pK_a is a measure of the acidity of a substance

$$pK_a = -\log K_a$$

made the relation inverse

↳ Just like pH it is easier to work with and remember

So: $\uparrow \text{Acidity} = \uparrow [H^+] = \uparrow K_a = \downarrow pH = \downarrow pK_a$

the stronger the tendency of an acid to dissociate the higher is the K_a and thus the lower the pK_a

at equilibrium $pK_a = pH$ so when 50% of the acid is dissociated, the $[HA] = [A^-]$

$$\begin{aligned} pK_a &= -\log K_a \\ &= -\log \frac{[A^-][H^+]}{[HA]} \\ &= -\log \frac{[A^-][H^+]}{[HA]} \\ &= -\log [H^+] \end{aligned}$$

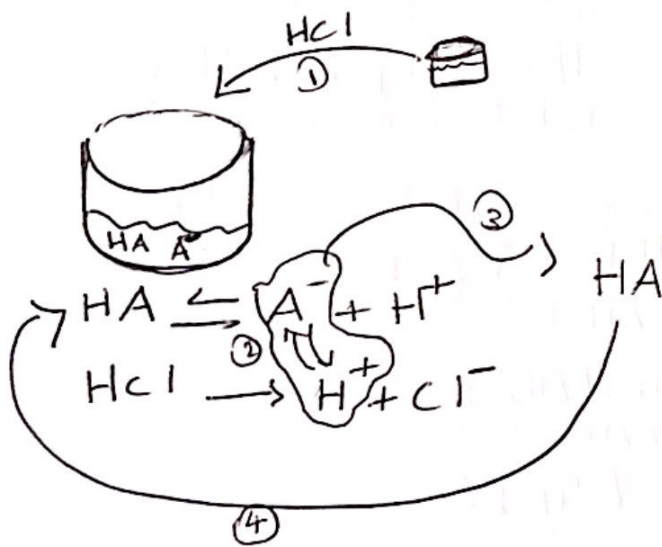
$$pK_a = pH$$

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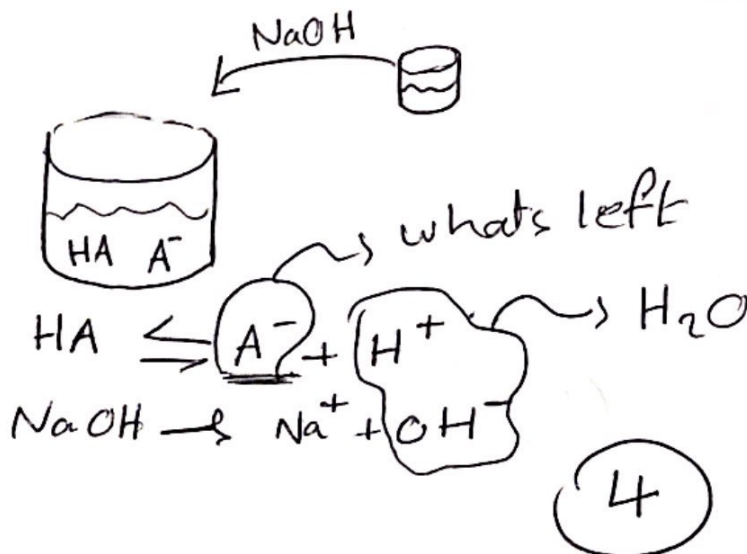
Buffers : a solution that resists pH changes when bases or acids are added to the solution

* Buffer solutions consist of a weak acid (undissociated acid) and its conjugate base

* it works because added acids (H^+) are neutralized by the conjugate base (A^-) which converts to the acid (HA)



* Added bases are neutralized by the acid (HA) which is converted to the conjugate base to give protons to the base coming in



- * Two Factors determine the effectiveness of a buffer
 - 1- its pKa relative to the pH of the solution
 - 2- its concentration

* Henderson-Hasselbalch Equation:

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

In Log's when
numerator > denominator
| Log = + |
numerator < denominator
| Log = - |

* an adjusted equation to describe the relationship between the acid and its conjugate base with pH and pKa

* the most effective buffers are when $pH = pK_a$ means it has equal concentration of the acid and its conjugate base \rightarrow 50% dissociated

* at $pH = pK_a \pm 1$ the Buffer capacity falls ~~to~~ 33% from the maximum value, therefore the buffer is effective one point up or down the pH pKa value

* important conclusions

$pH > pK_a$: more conjugate base than acid
 $pH < pK_a$: more acid than conjugate base
 $pH = pK_a$: Equal concentration of acid and conj base

(5)

why :

when $pH > pK_a$ is basic

$$bc \quad pH = pK_a + \text{Log} \frac{[A^-]}{[HA]}$$

$$ex: 5 = 4 + \text{Log} \frac{[A^-]}{[HA]} \rightsquigarrow$$

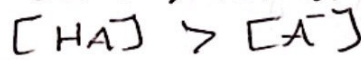
to get the
equation to be
equal
the Log must
be positive
so the numer > deno
 $[A^-] > [HA]$



Same thing but inverse

for when $pH < pK_a$

deno > numer



for the Log to be negative


$$4 = 5 - \text{Log} \frac{[A^-]}{[HA]}$$

(6)

Chemical Bonds

* Refers to attractive forces that hold atoms together in compounds

* They happen when an atom:

- Loses  electrons (give)
- accept electrons
- Share electrons with another atom

* A stable compound occurs when the total energy of the combination has lower energy than the separated atoms

$$\begin{array}{r} x + y = z \\ 7 \quad 20 \quad 15 \end{array}$$

$$27 > 15$$

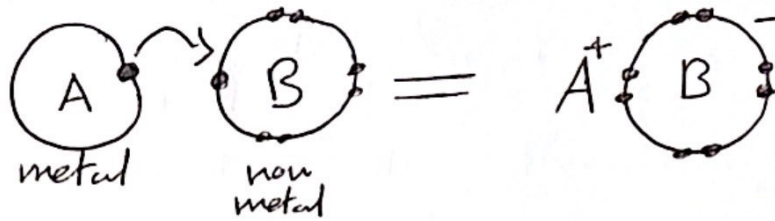
stable

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Bonds :

1- Ionic Bonds

* Happens between a metal and a non-metal



* electrons are completely transferred from one atom to another

* Produces ~~two~~ oppositely charged ions that attract and are attracted together by electrostatic forces

* Ionic compounds dissolve easily in water and other Polar Solvents (Likes dissolves Likes)

Polar compounds dissolve Polar Compounds
- water is Polar
non Polar dissolves non Polar

* Ionic compounds easily conduct electricity in solutions

* Ionic compounds tend to form crystalline solids with high melting temperatures

2- Covalent Bonds :

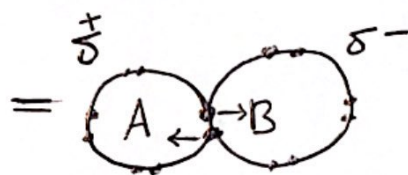
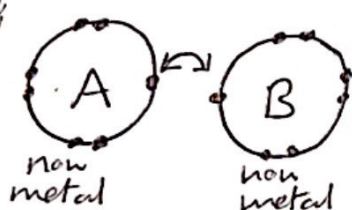
* Results from sharing one or more electron pairs between two atoms

* happens between a non-metal and a non-metal

* occurs because the atoms in the compound have a similar tendency for electrons (generally gain e^- s)

* The elements involved will share electrons in an effort to fill their valence shell (become stable)

exg



↳ this is a polar bond

Covalent Bonds have two types :

Polar C.B : happens when electrons are unequally shared between atoms (between atoms with different electronegativity)

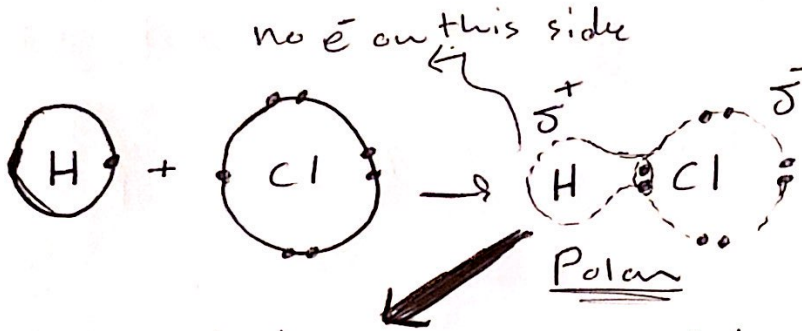
is the ability of an atom to attract electrons to itself in a covalent bond

~~all~~

Non Polar C.B : when e^- s are equally shared (between atoms with the same electronegativity)

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ex:

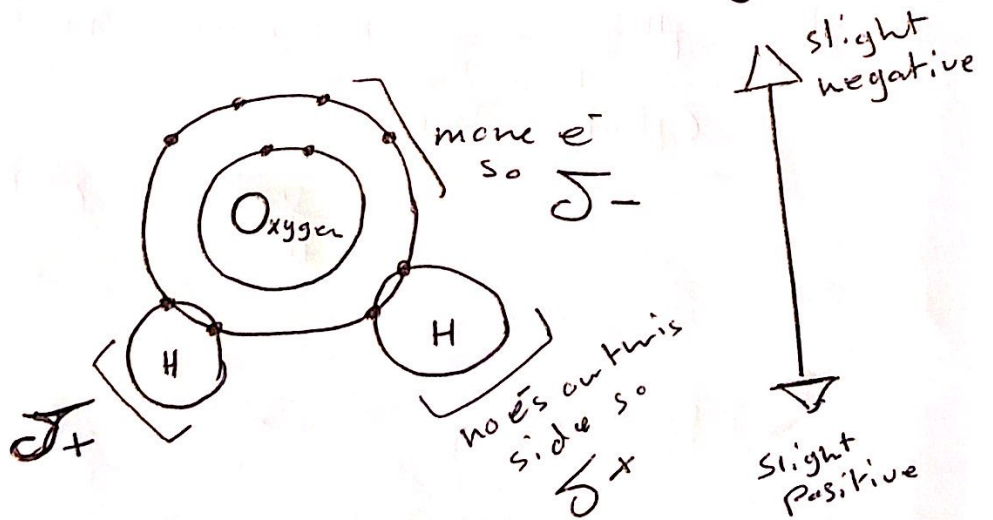


electrons are not completely transferred
 they are closer to the atom with
 more electronegativity
 so one atom is just Partially (+) charged
 and the other is Partially (-)

Just like 2 ppl
 Pulling a rope with a ball in
 the mid and one is stronger so the ball
 is closer to the stronger

* the bonding e^- 's will spend a greater time
 around the atom that has the stronger affinity
 for electrons

ex: water



* The electronegativities of selected elements :

F	4.0
O	3.5
N	3.0
Cl	3.0
C	2.5
H	2.1
P	2.1
Mg	1.2
Ca	1.0
Na	0.9
K	0.8

in the bond :
if the electronegativity Difference is :

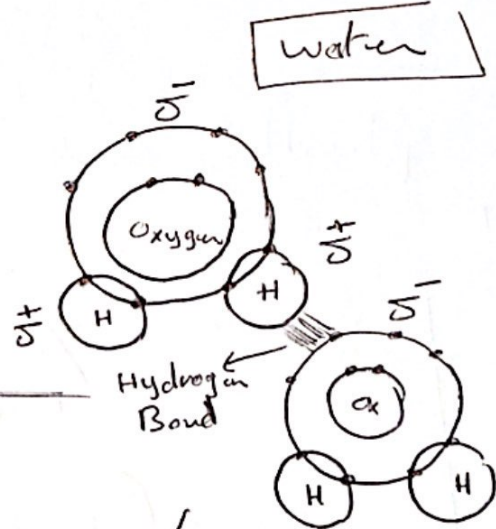
- 1 | $\Delta EN < 0.5$: it is non-Polar bond
- 2 | $0.5 \leq \Delta EN < 2.0$: Polar bond
- 3 | $\Delta EN > 2.0$: Ionic Bond

* Hydrogen Bond :

- it is the attractive interaction of a Hydrogen atom that is covalently Bonded to an electro Neg atom with another electro neg atoms like - Nitrogen, oxygen, fluorine

the Partial (+) region of the (H) is attracted to the Partial (-) region of (O)

(H) must be covalently Bonded to another electro neg atom to create a bond



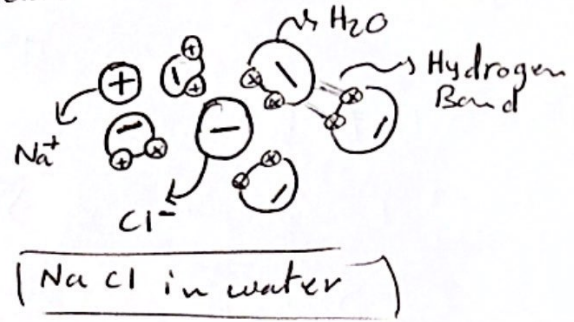
strength of Bonds

- ↑
- Covalent
 - Ionic
 - Hydrogen
 - Van der Waals

Why and how salts dissolve in water?

* Both water and salt compounds are Polar

* When the salt and water are mixed, the salt dissolves because the covalent bonds of the water are stronger than Ionic Bonds thus pulling NaCl apart



* The positive side of water molecule is attracted to the negative charge of Cl⁻ and the (-) side to the (+) Na⁺

* This results in a homogeneous solution

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Hydrophobic interactions:

* Water is Polar so it only interacts with Polars (Likes dissolve likes)

* Non Polar groups do not form hydrogen bonds with water so they are insoluble in water

* Hydrophobic substances are "excluded" from aqueous solutions which drives these molecules to cluster together

* No affinity between non polar substances except van der Waals forces that promote the weak bonding of non polar substances

* Van der Waals forces :

* weak attractive forces between electrically neutral atoms or molecules

* much weaker than Ionic or covalent bonds

* magnitude of the forces depends on the distance between neighbouring molecules

* these forces may develop cuz ~~the~~ the rapid shifting of electrons within molecules causes some parts to become momentarily charged, either (+) or (-)

* that's why weak transient forces of attraction can develop between particles that are actually neutral

