

8: Acids and Bases

Date

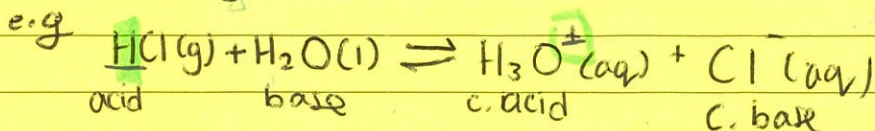
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Definitions of Acids and Bases

Bronsted Lowry acid - Proton Donor $[H^+]$

Bronsted Lowry base - Proton Acceptor $[H^+]$



The difference between a conjugate base/acid and acid/base is the presence of a H^+ ion (electrophile)

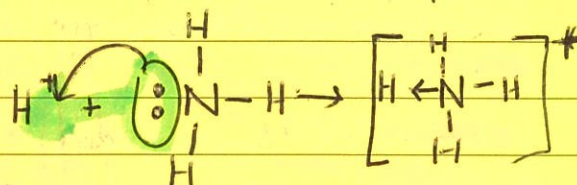
Lewis Acid: A lone electron pair acceptor

Lewis Base: A lone electron pair donor

(nucleophile)

Lewis theory

- o includes BL acids and alkali but includes those that has a full valence shell, the H^+ can be replaced with an atom with an empty change orbital
- o Always involves formation of dative bond,
- o See the NH_3 reaction



An amphiprotic ^{gain or lose protons} substance is one that can act as a Lowry/Bronsted Acid or Base.

E.g. H_2O - Oxygen has 2 lone pairs [can function as a Lewis base] and has 2 hydrogen atoms attached that can be lost as H^+ ions, hence H_2O would act as a proton donor.

Reactions

- 1) Acid + Carbonate \rightarrow Salt (carbonate) + Water + carbon dioxide
- 2) Acid + ammonia base \rightarrow ammonia gas + salt
- 3) Acid + Base \rightarrow salt + Water [but e.g. of exception $\rightarrow NH_3(aq) + HCl(aq) \rightarrow NH_4Cl(aq)$]
- 4) Acid + Metal ^(reactive) \rightarrow salt + hydrogen gas
- 5) Acid + metal oxide \rightarrow salt + water $[-\Delta H]$
- 6) Acid + hydrogencarbonate \rightarrow salt + water + CO_2

fully dissociates
weak conjugate base
strong conjugate base

Strong vs Weak Acid/base (partially dissociates)

Date No.

- Determined by the extent to which they ionise in solution
- We can determine this by ~~determining~~ ^{observing} the equilibrium of the reaction, if it lies to the reactants - the acid/base has not fully ionised → ergo it is a weak acid/base
- There are certain acids and bases whose reactions go to completion

Strong Acids [weak c. base]

- HCl
- HNO₃
- H₂SO₄ [only first ionisation]

Strong bases [weak c. acid]

- LiOH
- NaOH
- KOH and Ba(OH)₂

◦ weak conjugate acid/bases are v. stable, hence equil favours r.h.s

Weak Acids [strong c. base]

- H₃PO₄
- CH₃COOH or any organics
- H₂CO₃

Weak Bases [strong c. acid]

- NH₃
- ethylamine or any amines

Experimental Methods [based on principle that greater strength = higher concentration]

1) Electrical Conductivity

- higher concentration of mobile ions begets a larger number of charged particles available to carry a charge throughout a solution [for strong base/acids]
- Use electrodes + ammeter → ∴ greater electrical conductivity
- if asked, say test conductivity and that one is a better conductor than the other

2) Rate of Reaction

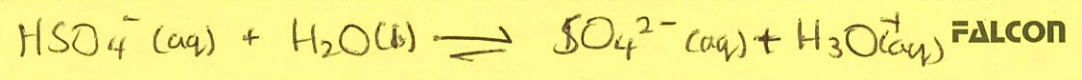
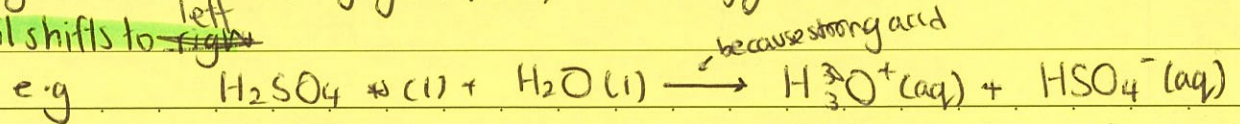
- ↑ [H⁺] / [OH⁻], more collision per unit time per unit volume, increasing the rate → this matches that of a strong acid/base (as ↑ [H⁺]/[H⁻])
- use a reaction with a strong metal or acid (al/zn), say the reaction is more vigorous

3) pH

- $-log_{10} [H^+]$: conc of H⁺ ions in mol/dm³
- The lower this number, the more acidic something is
- direct measurement of [H⁺] → defines strength

X-Protic Acids

- X → number of protons that can be lost
- ionises in X steps and % ionisation decreases with each step because negative charge builds on the conjugate base, more energy needed to remove a H⁺ from it, equil shifts to ~~right~~ ^{left}



pH calculations

pH (potential hydrogen) = $-\log_{10} [H^+]$ and $[H^+] = 10^{-pH}$

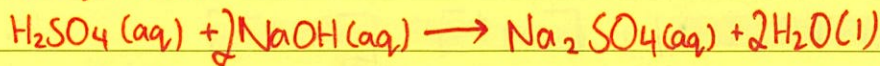
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Example Question

- ① 10ml of H_2SO_4 [0.1 mol/dm^3] was added to 5ml of 0.1 mol/dm^3 NaOH. What is the pH of the final solution

① Write the equation



② Determine the limiting reagent

$$\text{Moles of } H_2SO_4 \text{ available} = \frac{10}{1000} \cdot 0.1 = 0.001$$

$$\text{Moles of NaOH available} = \frac{5}{1000} \cdot 0.1 = 0.0005 \rightarrow \therefore \text{this is the limiting reagent as } 0.00025 \text{ moles of } H_2SO_4 \text{ need react with it}$$

③ Find the moles of the excess solution left

$$\frac{0.0005}{2} = 0.00025 \text{ [moles of } H_2SO_4 \text{ needed]}$$

$$0.001 - 0.00025 = 0.00075 \text{ moles}$$

④ Depending on the protic nature of the solution, find the moles of H^+

$$\text{As this is } H_2SO_4, 0.00075 \times 2 = 0.0015 \text{ moles}$$

⑤ \div this by total volume of solution [✓IP]

$$\frac{0.0015 \text{ moles}}{0.015 \text{ dm}^3} = 0.1 \text{ mol/dm}^3$$

⑥ Apply pH equation

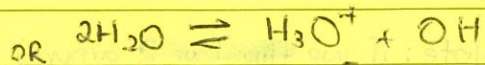
$$-\log_{10} [0.1] = 1$$

$-\log_{10} [OH^-]$

$-\log_{10} (K_w)$

$pOH, K_w, pK_w, K_b, pK_b, K_a, pK_a$ that's a lot to know!

$$pOH = -\log_{10} [OH^-]$$



$$K_w = K_c [55 \text{ mol/dm}^3] \rightarrow H_2O \rightleftharpoons H^+ + OH^-$$

$$K_w = 1 \times 10^{-14} \text{ mol}^2 \text{ dm}^{-6} \text{ at } 298 \text{ K}$$

$$\therefore K_c = \frac{[H^+][OH^-]}{[H_2O]^2}$$

$K_w \uparrow, T \uparrow$ because as $\uparrow T, \uparrow K_c$

As $T \uparrow, \uparrow [H^+]$ and $[OH^-]$, hence

$[H_2O]$ is constant, at 55

\downarrow pH and \downarrow pOH

$\therefore K_w$ [ionic product of water]

$$-\left[\log_{10} [H^+] + \log_{10} [OH^-] \right] = 14 \text{ at } 298 \text{ K}$$

OR

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

$$pH + pOH = pK_w$$

$$pK_w = -\log_{10} [K_w]$$

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

A note on pH and neutrality.

Note the equation $2\text{H}_2\text{O} \rightleftharpoons \text{H}_3\text{O}^+ + \text{OH}^-$

The $[\text{OH}^-]_{\text{m}} : [\text{H}^+]_{\text{m}}$ is 1:1.

Even though pH would change with T, the $\text{OH}^- : \text{H}^+$ stays at 1:1, this makes a solution neutral.

FOR NEUTRALITY $\rightarrow [\text{H}^+] = [\text{OH}^-]$

FOR ACIDITY $\rightarrow [\text{H}^+] > [\text{OH}^-]$

FOR BASIC NATURE $\rightarrow [\text{H}^+] < [\text{OH}^-]$

\rightarrow [SEE FINAL PAGE]

K_a (acid dissociation constant) \rightarrow Expression of the strength of weak acid / base

\rightarrow used for weak acids

For the equation $\text{HA} \rightleftharpoons \text{A}^- + \text{H}^+$ $\text{H}_2\text{O} + \text{HA} \rightleftharpoons \text{H}_3\text{O}^+ + \text{A}^-$

$$K_c = \frac{[\text{H}_3\text{O}^+][\text{A}^-]}{[\text{HA}][\text{H}_2\text{O}]}$$

but, we assume $[\text{H}_2\text{O}]$ is constant so at 55M

$$K_c [\text{H}_2\text{O}] = \frac{[\text{H}_3\text{O}^+][\text{A}^-]}{[\text{HA}]} = K_a$$

$$\text{p}K_a = -\log_{10}(K_a)$$

$$\text{p}K_b = -\log_{10}(K_b)$$

another assumption we make is that $[\text{HA}]$ does not change, as we know that this is a weak acid

K_b is identical (-ish)

For 2) $\text{B} + \text{H}_2\text{O} \rightleftharpoons \text{BH}^+ + \text{OH}^-$

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

the same story unfolds.

Note: if we flip 1 or 2 around

1' $\rightarrow \text{A}^- + \text{H}_3\text{O}^+ \rightleftharpoons \text{H}_2\text{O} + \text{HA} \rightarrow$ base (base is A^-)

2' $\rightarrow \text{BH}^+ + \text{OH}^- \rightleftharpoons \text{B} + \text{H}_2\text{O} \rightarrow$ acid (acid is OH^-)

so, their K_a 's/ K_b 's become K_b/K_a , with the reciprocal of the old ∇ one.

$$\text{E.g. } 1' \rightarrow K_b = \frac{[\text{HA}]}{[\text{H}_3\text{O}^+][\text{A}^-]}$$

monoprotonic

Example: The pH of a $0.020 \text{ mol dm}^{-3}$ solution of weak acid is 3.9. Find K_a .

Step 1: Write equation for $K_a \rightarrow \frac{[\text{H}^+][\text{A}^-]}{[\text{HA}]} = \frac{[\text{H}^+]^2}{[0.02]} = K_a$

$$K_a = (10^{-3.9})^2$$

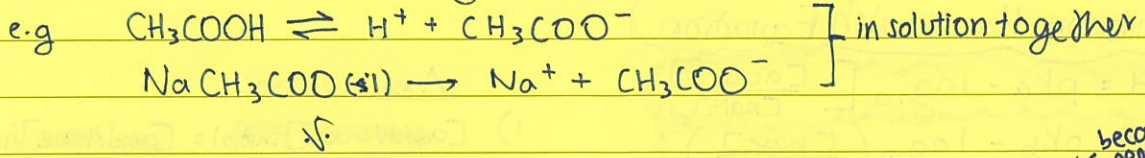
Buffer solutions

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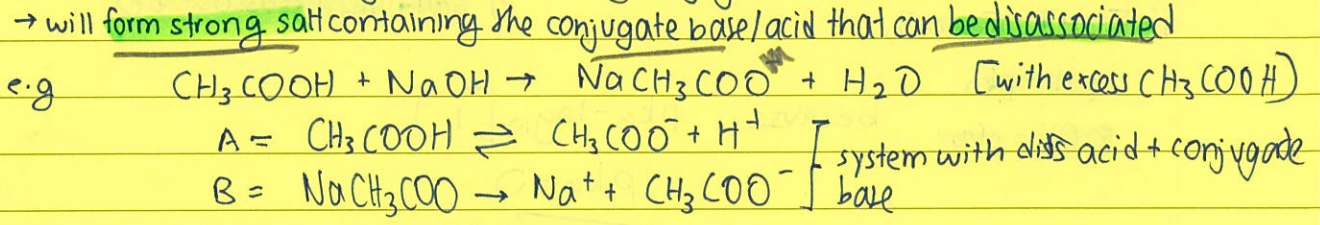
- Definition: a solution that resists changes in pH when small amounts of acids or alkali are added
- Consists of a weak acid / base with its strong conjugate base

2 Ways of formation

1) weak acid/base + salt containing conjugate acid/base



2) weak acid/base + a strong acid/base holding conjugate base/acid in limited quantities



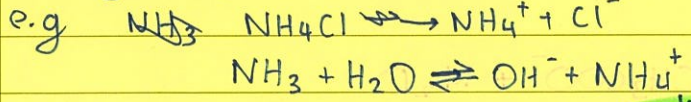
because we need acid leftover for buffer

Resisting Changes [Acidic Buffers]

Consider the system above

- addition of H^+ ions → equilibrium of A shifts to the left, using up the added H^+ to produce additional CH_3COOH , hence $[H^+]$ does not change (for calc) reacts with CH_3COO^- to prod CH_3COOH
- addition of OH^- → reacts with CH_3COOH to produce H_2O and CH_3COO^- , no change in $[H^+]$, by consuming added OH^-

Resisting Change [Basic Buffers]



- addition of H^+ → react with OH^- to form H_2O , ensure $[H^+]$ stays constant reacts with NH_3 to form NH_4^+ , $[H^+]$ doesn't change
- addition of OH^- → shifts equil to left in order to consume added OH^- , $[H^+]$ stays constant
reacts with the NH_4^+ to produce NH_3 and H_2O → $[H^+]$ constant

Why do we add more of the conjugate?

- we need high concs as the conjugate base/acid is used in maintaining pH,
- weak base/acid only weakly, not enough of conjugate
- Hence ionic salt holding it added so it fully dissociates, allowing for sufficient quantities

Factors Affecting Buffers

- Dilution: ↑ dilution, ↓ buffering capacity: amount of acid/base that can be absorbed without change in pH, this decreases as the molar conc of buffering components ↓
- Temp: affects K_a and K_b , hence affects pH. VTI to have constant temp in blood transfusions - due to blood buffers

Why don't we use strong bases?



No equilibrium to be shifted when H^+ added, hence $[\text{H}^+] \uparrow$, $\downarrow \text{pH}$

Addition of OH^- consumes H^+ , hence $[\text{H}^+] \downarrow$, $\uparrow \text{pH}$

Henderson-Hasselbalch Equation

$$\text{pH} = \text{pK}_a - \log_{10} \left[\frac{[\text{acid}]}{[\text{salt}]} \right]$$

$$\text{pOH} = \text{pK}_b - \log_{10} \left(\frac{[\text{base}]}{[\text{salt}]} \right)$$

if $[\text{salt}] = [\text{acid}]$,

$$\text{pH} = \text{pK}_a$$

$$\text{pOH} = \text{pK}_b$$

because $\text{pK}_a - \log_{10} [1]$
 $= \text{pK}_a - 0$

Example

A buffer has 0.1 mol of CH_3COOH and 0.13 mol of NaCH_3COO in 1 dm^3 of solution.

(a) pH of buffer?

(b) pH when 0.02 mol of KOH added?

(c) pH when 0.02 mol of HNO_3 added?

$$\text{pH} = \text{pK}_a - \log_{10} \left(\frac{0.10}{0.13} \right)$$

$$\text{pH} = 4.76 - \log_{10} \left(\frac{0.10}{0.13} \right)$$

$$\text{pH} = 4.87$$

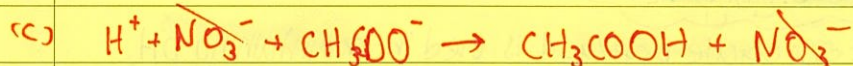


I	0.02	0.10	0.13
---	------	------	------

E	0	0.08	0.13
---	---	------	------

$$\therefore \text{pH} = 4.76 - \log_{10} \left(\frac{0.08}{0.13} \right)$$
$$= 5.03$$

Note: consideration of c. base and acid



I	0.02	0.02	0.13	0.10
---	------	------	------	------

E	0	0	0.11	0.12
---	---	---	------	------

$$\text{pH} = 4.76 - \log_{10} \left(\frac{0.12}{0.11} \right)$$

$$= 4.72$$

Assumptions

1) $[\text{acid/base}]_{\text{initial}} = [\text{acid/base}]_{\text{final}}$ - assumes v. weak dissociation

2) salt fully dissociates

Pre-requisites

1) If acid/alkali used to form salt, must be 1:1 reagent, otherwise no weak acid or base left \rightarrow cannot result change in pH

2) conjugate needed \rightarrow to react with $\text{OH}^- / \text{H}^+ \rightarrow \therefore$ we add salt

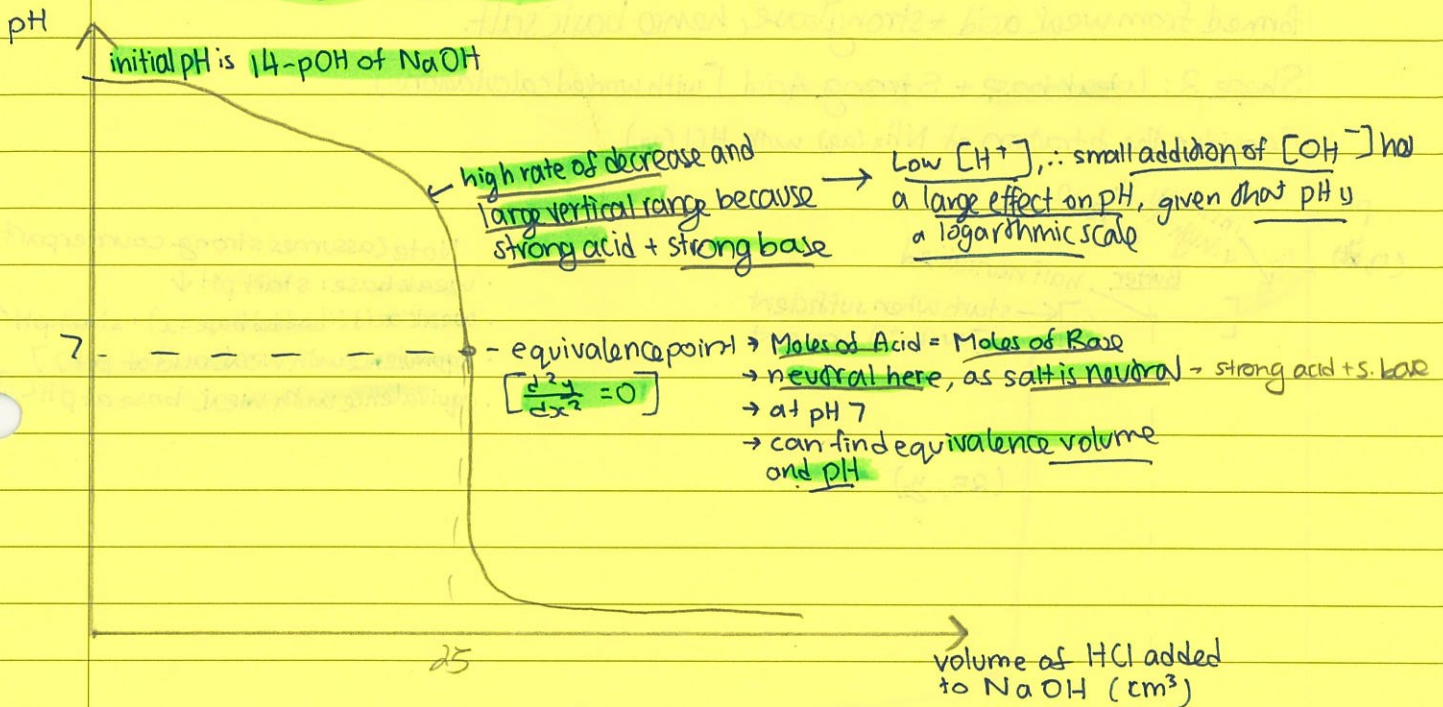
pH curves

Date

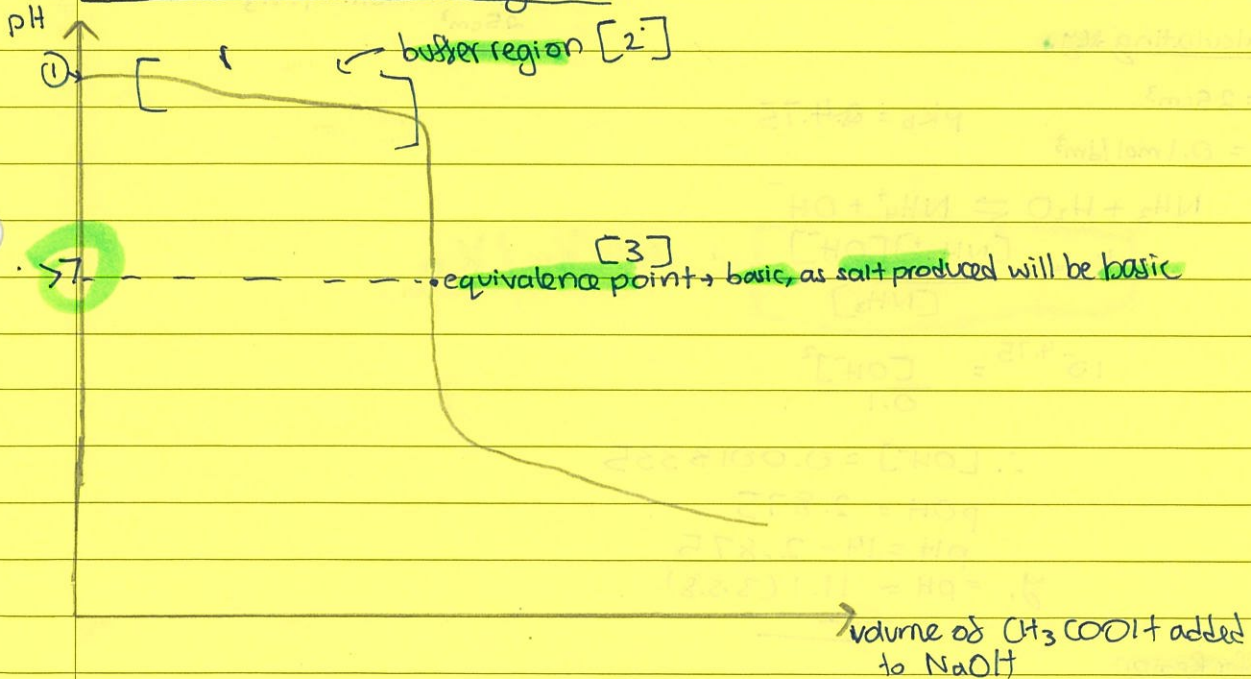
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- with pH on y and volume of acid/base added on the x axis

Shape 1 (strong Acid + Strong Base)



Shape 2 (weak acid + strong base)



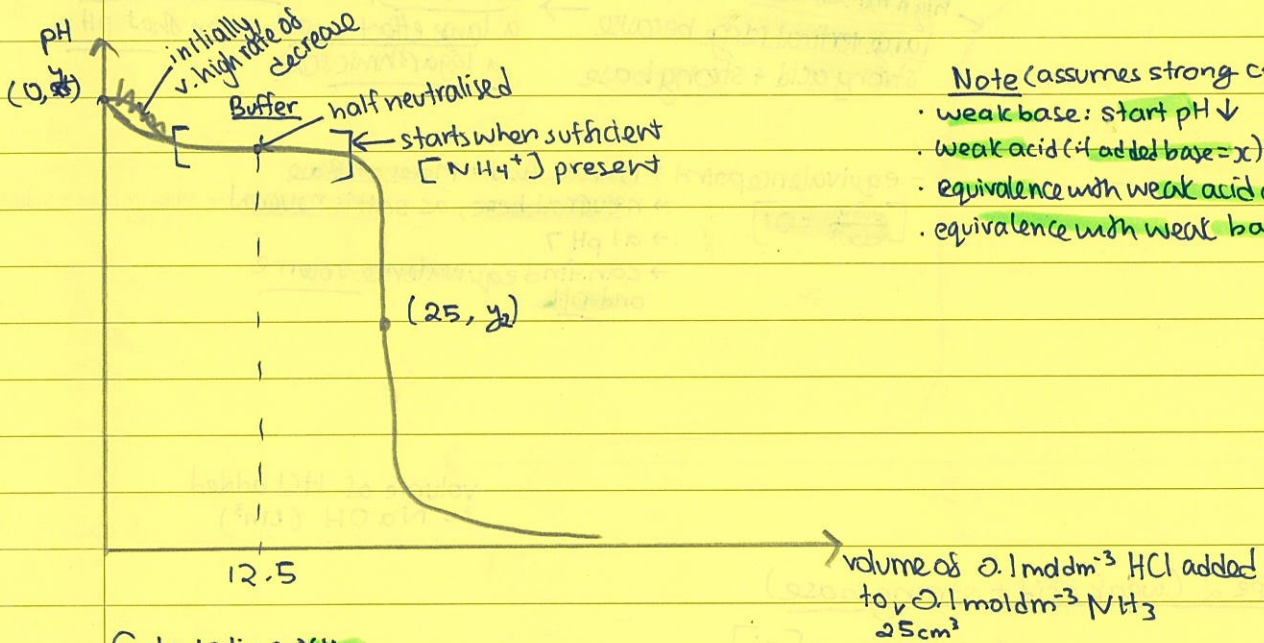
① Determined by strength of acid or base depending on x axis. Use either pK_a , pK_b or full dissociation to determine

② Buffer Region: Consider the salt of $CH_3COOH + NaOH$, $NaCH_3COO$. It dissociates into Na^+ and CH_3COO^- (strong conjugate base). CH_3COO^- has a lone pair and a negative charge. Once sufficient CH_3COO^- is present, it forms a dative covalent bond with H^+ in solution, $\downarrow [H^+]$. If more H^+ is added, it simply reacts with the added H^+ to produce CH_3COOH , consuming added H^+ and resisting

③ - Equivalence Point: Still when moles of acid = moles of base. However, the salt produced here would be basic as $\text{NaCH}_3\text{COO} \rightarrow \text{Na}^+(\text{aq}) + \text{CH}_3\text{COO}^-(\text{aq})$. The CH_3COO^- forms a dative covalent bond with a surrounding H^+ , $\downarrow [\text{H}^+] \therefore \uparrow \text{pH}$ at equivalence. Say that salt was formed from weak acid + strong base, hence basic salt.

Shape 3: Weak base + Strong Acid [with worked calculations]

Consider the titration of $\text{NH}_3(\text{aq})$ with $\text{HCl}(\text{aq})$.



Calculating y_1

$V = 25 \text{ cm}^3$

$pK_b = 4.75$

$C = 0.1 \text{ mol/dm}^3$



$K_b = \frac{[\text{NH}_4^+][\text{OH}^-]}{[\text{NH}_3]}$ ← use K_b / K_a

$10^{-4.75} = \frac{[\text{OH}^-]^2}{0.1}$

$\therefore [\text{OH}^-] = 0.0013335$

$p\text{OH} = 2.875$

$p\text{H} = 14 - 2.875$

$y_1 = p\text{H} = 11.1 (3.s.s)$

Buffer Region

Ionic salt produced by $\text{NH}_3 + \text{HCl}$ (NH_4Cl), dissociates into NH_4^+ and Cl^- .

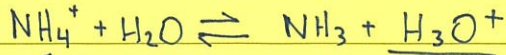
\therefore we have a system in place, once enough NH_4^+ is produced, that can be a buffer

Thus, adding more HCl results in limited pH change as added H^+ is reacted with NH_3 to produce NH_4^+ , consuming added NH_4^+ . This can be proven by the Henderson Hasselbach

Equation

- γ_2
- Realistically, γ_2 should be < 7 . This is because NH_4Cl is acidic in solution

Proof



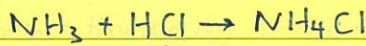
cation hydrolysis

added H_3O^+ , $\therefore \uparrow [\text{H}^+]$ and $\downarrow \text{pH}$

Half-Neutralisation

- for $\text{pH} = \text{pK}_a$ or $\text{pOH} = \text{pK}_b$, $[\text{base/acid}] = [\text{concentration of salt}]$. This can happen when half of the solution being titrated is neutralised, because

Calculation



Ini: 0.0025 0.00125 0

Cha: -0.00125 +0.00125 +0.00125

End: 0.00125 0 0.00125

$\text{pH at this point} = \text{pK}_a - \text{pK}_b$

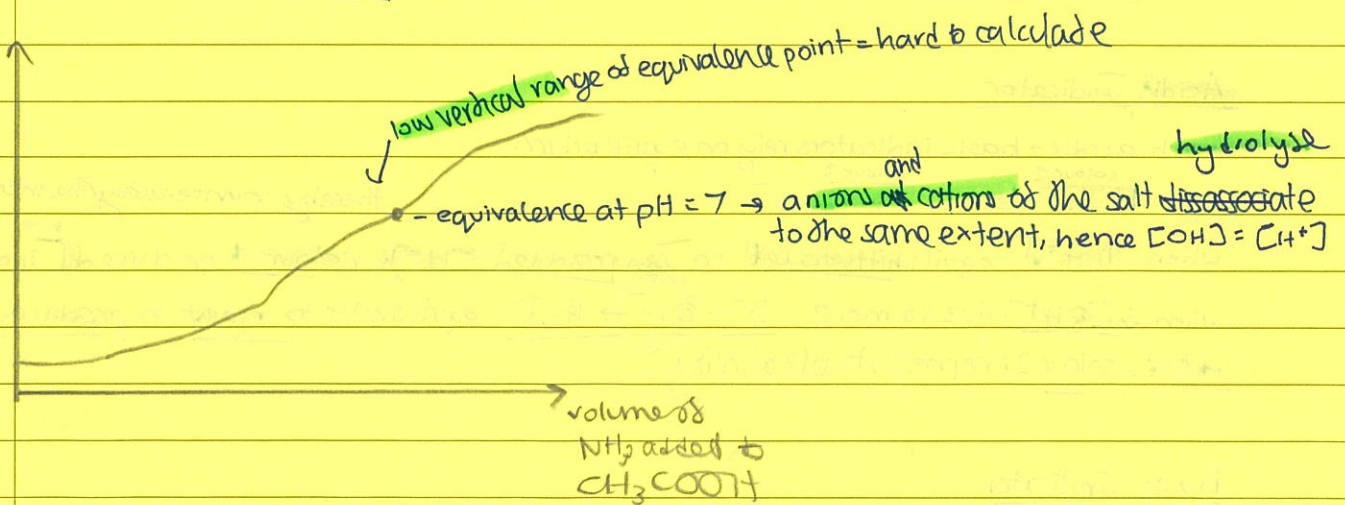
$$\text{pOH} = \text{pK}_b - \log_{10} \left[\frac{[\text{NH}_3]}{[\text{NH}_4\text{Cl}]} \right]$$

$$= \text{pK}_b$$

$$= 4.75$$

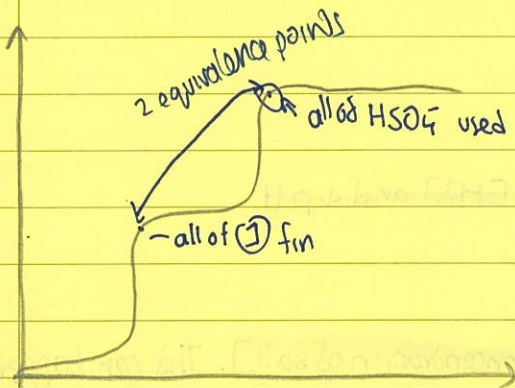
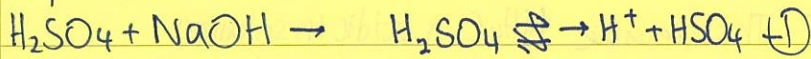
$$\text{pH} = 9.25 \text{ at } x = 12.5$$

Weak Acid + Weak Base (Curve 4)



Polyprotic Acids

Date No.



→ NOT ASSESSED!

Strong vs Weak [Acids]

- | | |
|--------------------------------|-------------------------------------|
| ◦ lower equivalence | ◦ higher equivalence |
| ◦ lower start | ◦ higher start |
| ◦ high vertical range of equiv | ◦ low vertical range of equivalence |

Salts [consider conjugates]

- s.a + s.b → neutral salt + water
- w.a + s.b → basic salt + water
- s.a + w.b → acidic salt + water
- w.a + w.b → neutral salt + water

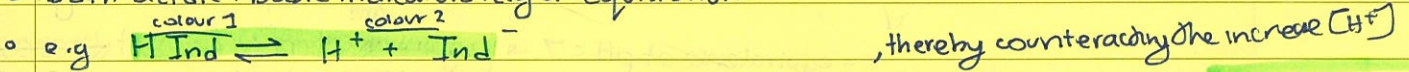
cations and anions hydrolyse to same extent

Indicators

- Definition: An indicator is a weak acid or base that changes colour when it becomes dissociated
- point at which it changes colour = end point
- Equivalence point must match end point, ∴ endpoint must be within vertical range of the equiv
- Especially considered with weak acids/bases as their vertical range is smaller
- usually weak acids or bases

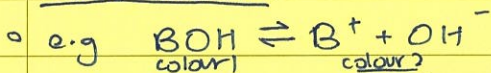
Acidic Indicator

◦ both acidic + basic indicators rely on equilibrium



when $\uparrow [H^+]$, equil shifts to left to use up added $[H^+]$, colour 1 produced [indicates acidity]
 when $\downarrow [H^+]$ due to more $H^+ + OH^- \rightarrow H_2O$, equil shifts to right to produce more H^+ , ∴ colour 2: represents alkalinity

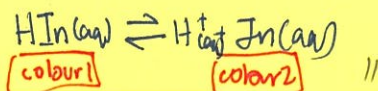
Basic Indicator



when $\uparrow [H^+]$, more OH^- reacts with it, equil shifts to right, colour 2 expressed, thus the acidic colour. When $[OH^-] \uparrow$, equil shifts to left, colour 1 expressed, acidic colour

NOTE

Always write indicators like this



K_{ind} and pK_{ind}

$$K_{\text{ind}} = \frac{[\text{In}^-][\text{H}^+]}{[\text{HIn}]} \rightarrow \text{halfway through colour change } [\text{In}^-] \text{ or } [\text{H}^+] = [\text{HIn}]$$

↳ ∴ mix of colour 1 and 2

$$\therefore K_{\text{ind}} = [\text{H}^+]$$

and pK_{ind} would = pH. Only at the halfway point.

However, this does not always happen, so we do ± 1 , to produce a colour change interval. e.g for methyl orange \rightarrow pK_{ind} = 3.7 but $3.1 < \text{pH} < 4.4$

Must select appropriate range that will be the same range of the equivalence point.
↳ phrase: "indicator end point within range of pH at equivalence."

Why is there a change in colour?

The energy gap of undissociated and dissociated will vary, hence as different degrees of π bond conjugation in each states. Hence different wavelengths of light absorbed and then transmitted.

Table of how different combos affect range of equivalence point

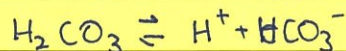
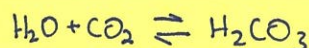
S.A + S.B	3 - 11, 7 in middle
^W S.A + ^S S.B	7 - 11, 9 in middle
^S S.B + ^W S.A	3 - 7, 5 in middle
W.A + W.B	No sig change of pH at equiv., but usually 7

± 2 for S.A or S.B

Acid Deposition

Why is rainwater acidic?

- Usually pH 5.65
- dissolved CO₂, which forms carbonic acid
- ∴ acidic rain = rain with pH < 5.60



* Acidic Deposition: Process by which acidic particles, gases and precipitates leave the atmosphere

Wet Acid Deposition

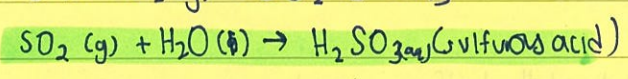
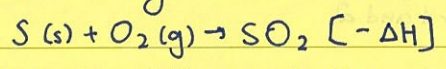
- fall to ground as aqueous precipitates (referring to the acidic material), e.g snow, rain, hail

Dry Acid Deposition

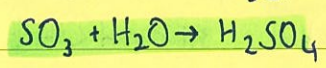
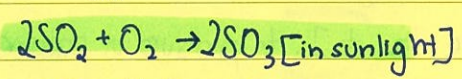
- Acidifying particles, gases fall to ground as smoke/dust and later dissolve to form acids.

Sulfur Oxides

SO₂ from burning fossil fuels [coal + heavy oil] and smelting [iron from ore]

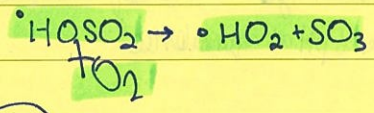
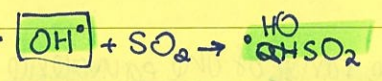


OR



} dry acid deposition

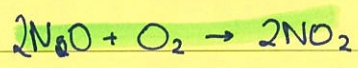
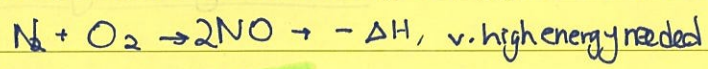
hydroxyl radical in atmosphere [one unpaired e⁻]



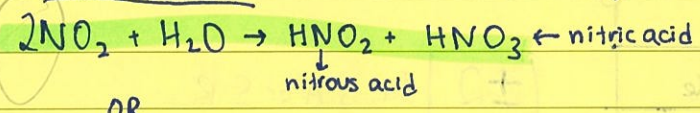
} Photooxidation

Nitrogen Oxides

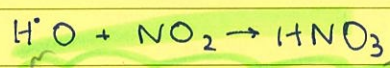
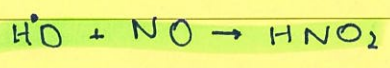
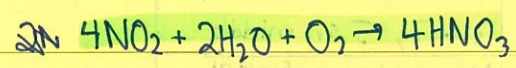
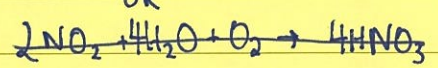
NO for instance is formed from the heat of internal combustion engines or from lightning strikes



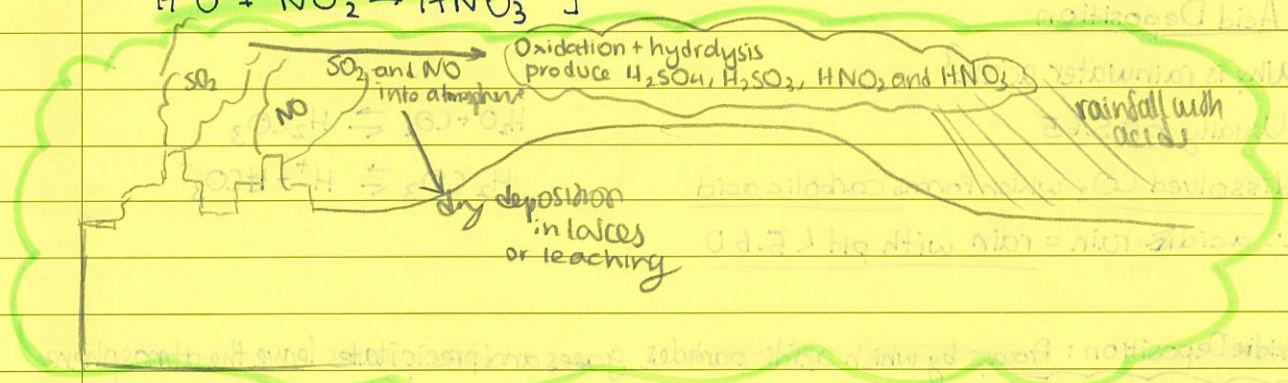
Acidification



OR



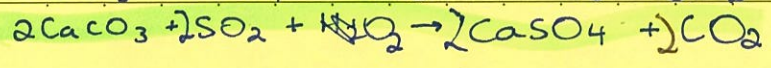
} Photooxidative reaction



Effects

1) Materials

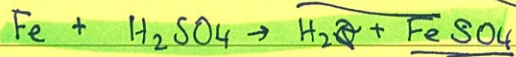
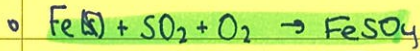
- marble and limestone made of CaCO₃
- react with H₂SO₄, H₂SO₃ and SO₂ (dry deposition) to form CaSO₄



- Also reacts with nitric acid $\rightarrow 2\text{HNO}_3 + \text{CaCO}_3 \rightarrow \text{Ca}(\text{NO}_3)_2 + \text{H}_2\text{O} + \text{CO}_2$
- CaSO_4 and $\text{Ca}(\text{NO}_3)_2$ very soluble, hence washes off the limestone easily
- CaSO_4 and $\text{Ca}(\text{NO}_3)_2$ bigger than CaSO_3 in Mr, hence creation causes expansion and cracking

Metals

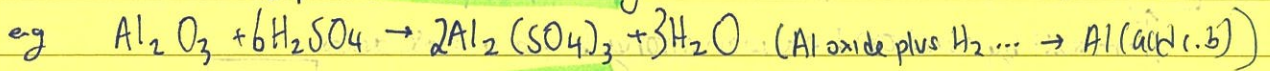
- dry deposition and wet react with metals, e.g. Iron



enables ionic conductivity

increase in rate of electrochemical corrosion reactions

- Acid rain can remove protective oxide coatings

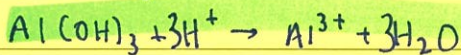


\therefore causes damage to bridges, roads + tracks.

Lakes + Water Bodies

- Causes leaching of important

- Reacts with $\text{Al}(\text{OH})_3$ in rock to form soluble, toxic Al^{3+}



- Al^{3+} interferes with gill operation \rightarrow cannot take in O_2

- Cause eutrophication due to nitrates in acid rain \rightarrow algal blooms result in increased BOD demand, death of ecosystem.

Plants

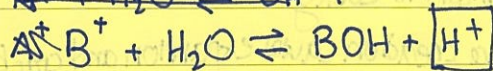
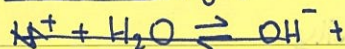
- Al^{3+} in the same reaction, damages roots, cannot take in nutrients \rightarrow stunted growth and yellowing + loss of leaves, thinning of tree tops
- Acid Rain causes leaching of nutrients like Ca^{2+} and Mg^{2+} \rightarrow needed for chlorophyll, cannot photosynthesise
- dry deposition block stomata - no gas exchange

Humans

- Fine sulfate and nitrate particles irritate mucous membranes \rightarrow ↑ risk of asthma, bronchitis and emphysema.
- Heavy metal ions like Al^{3+} in water could cause Alzheimers.

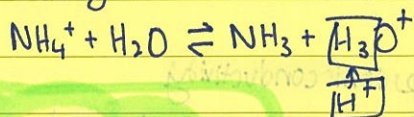
Salts [overall pH determined by the comparative hydrolysis of cations and anions]

Cation Hydrolysis



comes about from a salt created by a weak base but strong acid

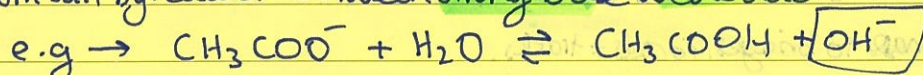
e.g



Anion Hydrolysis



from salt by reaction between strong base weak acid



If cations and anions hydrolyse to same extent, pH = 7.

* Note: Group 1 + 2 cations cannot hydrolyse to form H^+ as their charge densities are too low

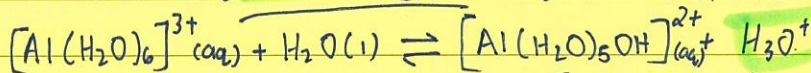
Metal Cation Hydrolysis

Consider $[Al(H_2O)_6]^{3+}$

Al^{3+} is a metal cation with a high charge density

This high charge density allows it to polarise the O-H bond in water, weakening it in the process

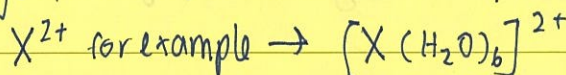
Therefore, it allows for a H^+ ion to leave a H_2O and react with another H_2O to form the hydroxonium ion



This keeps happening [addition of OH^- and loss of H^+] until the complex ion loses its charge, thus it can no longer dissolve $[Al(H_2O)_3(OH)_3] (s)$ - Forms a precipitate

\therefore Metal Cations form acidic solution by $\uparrow [H^+]$

NOTE: If you have XA , where X is a transition metal ion and A^- is an anion, you must use the hydrated complex ion of X and explain from there.

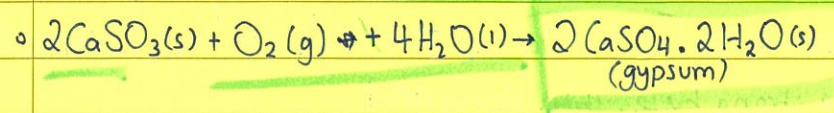
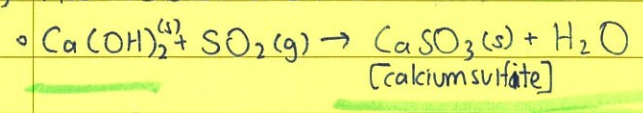


brackets: charge distributed across molecule

and NO_x

Post Combustion Methods of reducing SO₂ emissions

1) Flue Gas Desulfurisation



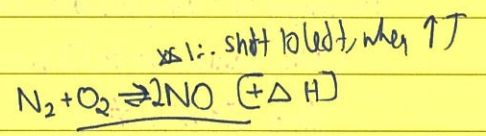
gypsum or hydrated calcium sulfate can be used to make plaster and cement

2) Catalytic Converter

- honey-comb interior to increase S.A for contact with gases
 - Uses porous Al₂O₃ as coating with platinum and other precious metals.
 - Catalyses the reduction of NO_xs into N₂ and O₂
- e.g. $2\text{CO}(\text{g}) + 2\text{NO}(\text{g}) \rightarrow 2\text{CO}_2(\text{g}) + \text{N}_2(\text{g})$

3) Lower Temperature Combustion

- formation of NO reduced at lower temperatures
- can be accomplished by recirculating exhaust gases into the engine, ↓ NO in emission



Pre Combustion

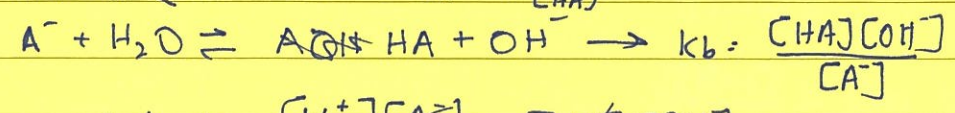
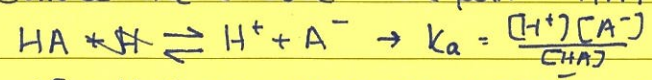
1) Hydrodesulfurisation = process by which sulfur is catalytically removed from refined petroleum, so that it can react with by reacting it with H₂ to form H₂S!
 v. toxic, so it is trapped and broken down into elemental sulfur to manufacture H₂SO₄.

- 2) Alternatives - lower demand for fossil fuels
- more efficient energy transfer systems
 - Switching to more renewable energy

3) Liming of lakes - neutralises acid and increases concentration of Ca⁺ and CO₃²⁻, good for coral

K_a and K_b [Again]

Consider the c. base - acid pair HA / A⁻



$$\therefore K_a \cdot K_b = \frac{[\text{H}^+][\text{A}^-]}{[\text{HA}]} \cdot \frac{[\text{HA}][\text{OH}^-]}{[\text{A}^-]}$$

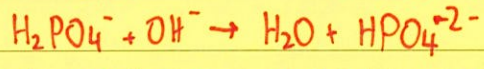
$$K_a \cdot K_b = [\text{H}^+][\text{OH}^-] = K_w \quad \therefore \boxed{pK_a + pK_b = pK_w} \text{ for c. acid / base pair}$$

This is ~~quant~~ quantitative proof that stronger acids ~~disassociate~~ have weak c. bases because, if pKa is low, pKb must be high [∴ con base is weak]

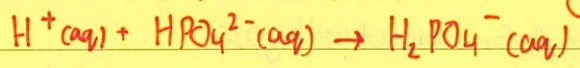
Buffers

Example Q: Describe how a phosphate buffer $[H_2PO_4^-(aq) \rightleftharpoons HPO_4^{2-}(aq) + H^+(aq)]$ minimises the effect of the addition of (a) a strong base, (b) a strong acid

(a) OH⁻ reacts with H₂PO₄⁻ to form HPO₄²⁻ and H₂O, minimising the effect of this addition



(b) H⁺ reacts with HPO₄²⁻ to form H₂PO₄⁻, minimising the effect of this addition



Kw, Ka and Kb

Example Q: Why does an acidic solution still have OH⁻ ions?

[OH⁻] cannot go to 0 because there is an equilibrium present such that $[OH^-] = \frac{K_w}{[H^+]}$, thus [OH⁻] ≠ 0

General Tips + Notes

- 1) ALWAYS WRITE EQUATIONS
- 2) Refer to the reactions in explaining how buffer solutions resist change
- 3) Always divide moles by volume for everything
- 4) State, in a titration question, the acid base combo in terms of strength
- 5) To find proportions of salt and acid → use Ka

e.g. 9.50 pH Buffer required with NH₃ and NH₄Cl

$$K_b = \frac{[NH_4^+][OH^-]}{[NH_3]}$$

$\frac{K_b}{[OH^-]} = \frac{[NH_4^+]}{[NH_3]}$ → find ratio, then deduce vol needed, can get directly if concs are already equal

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

$$pOH = pK_b + \log\left(\frac{[salt]}{[base]}\right)$$

$$4.50 = 4.25 + \log\left(\frac{[salt]}{[base]}\right) \Rightarrow \log\left(\frac{[salt]}{[base]}\right) = \log(0.25) = \frac{10^{-0.25}}{0.62} = 5.62 \times 10^{-1}$$