

ACIDS & BASES

acids and bases are solutions. Therefore, a solid understanding of Solution Chemistry and Equilibrium will guarantee you success in this unit.

I. Defining Acids & Bases:

A. General Characteristics of Acids & Bases:

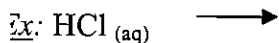
Properties of Acids	Properties of Bases
<p>Ionic compounds which <u>dissociate</u> in H_2O to give a solution that:</p> <ol style="list-style-type: none"> 1. tastes <u>sour</u> 2. conducts <u>electricity</u> 3. causes <u>indicators</u> to change colour <p>Loses the above properties when neutralized with a base. But the ending salt solution will still conduct <u>electricity</u>.</p>	<p>Ionic compounds which dissociate in H_2O to give a solution that:</p> <ol style="list-style-type: none"> 1. Tastes <u>bitter</u> 2. Conducts <u>electricity</u> 3. Causes certain indicators to change colour. 4. Feels <u>slippery</u> 5. Loses the above properties when neutralized with an acid. But resulting salt solution will still conduct electricity.

B. Definitions of Acids & Bases:

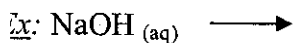
Arrhenius Definition:

Bronsted - Lowry Definition

Acid: an ionic compound that dissociates to give H^+ ions in solution.



Base: an ionic compound that dissociates to give OH^- ions in solution.



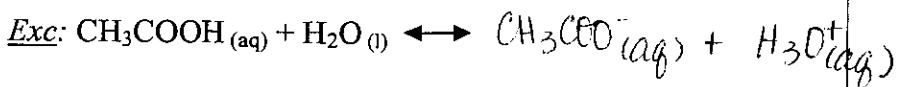
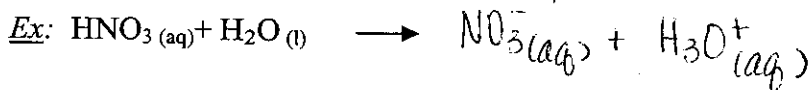
Lewis Acids & Bases

Lewis Acid: any compound that is an e^- pair acceptor.

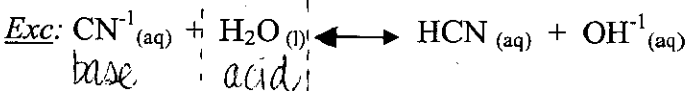
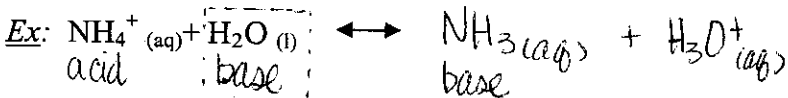
Lewis Base: any compound that is an electron pair donator.

This definition lends itself to describing acids and bases that do NOT have OH^- or H^+ in their chemical makeup.

Acid: species which donate a proton [$H^+_{(aq)}$]-

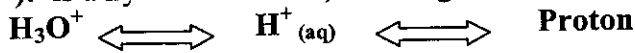


Base: species which receives a proton.



NOTE: In this case, H_2O acts as both a Bronsted-Lowry acid and a base. Therefore, water is considered AMPHIPROTIC or AMPHOTERIC.

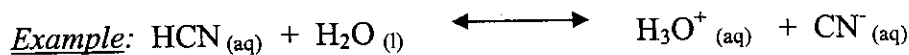
Hydronium ion (H_3O^+): is a hydrated H^+ ion, meaning it is dissolved in water.



II. Relationships between Acids & Bases:

A. **Conjugate Acid - Base pairs:** molecules and ions that differ only by one proton (or H^+).

Conjugate means partner.

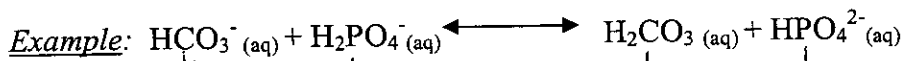


Conjugate Acid of CN^- : HCN

Conjugate Base of HCN : CN^-

Conjugate Acid of H_2O : H_3O^+

Conjugate Base of H_3O^+ : H_2O

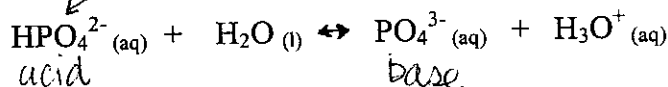
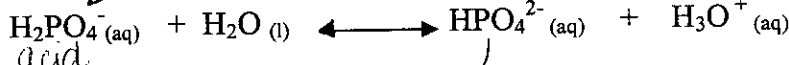
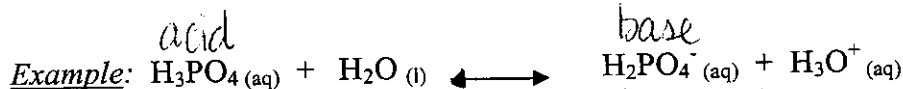


base acid

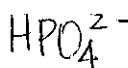
acid

base

B. **Polyprotic Acid:** an acid capable of donating multiple protons (H^+).



• AMPHIPROTIC:

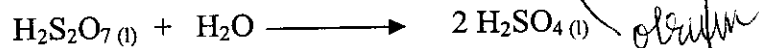
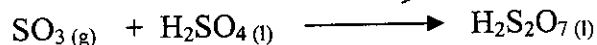


Species can be amphiprotic only if it is able to donate a H^+ AND accept a H^+ .

III. Preparation & Properties of some Common Acids and Bases:

A. Acids:

- **Sulfuric Acid:** simple to make from water, oxygen, and sulfur. No harmful by-products are produced.

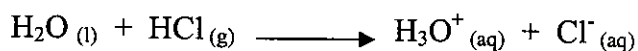
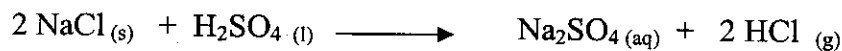


Contact Process.

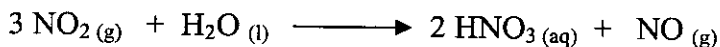
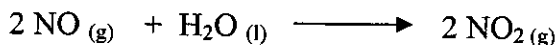
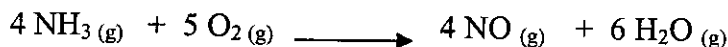
oleum

Properties: Strong acid – Good neutralizer
 Good oxidizing agent
 Great affinity with H₂O and can be used as a desiccating (drying) agent.
 Good electrolyte.

- **Hydrochloric acid:** soluble in all alkali ions.



- **Nitric Acid:** made by the Oswald Process:

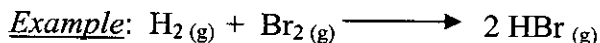


Properties: Has a suffocating odour.

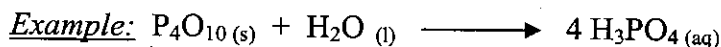
Leaves a yellow stain on skin due to NO₂ being exposed to light.

- **Other Types of Acids:**

Binary Acids: Those that have 2 elements. Can be made by directly combining element with halogen gas.



Other acids like phosphoric acid are made when a covalent oxide and water reacts. (OXYACIDS)



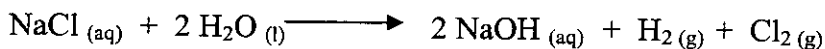
B. Bases:

covalent oxide

- **Sodium Hydroxide:** is in most Drano products. Reacts with fats and proteins to make soap, a substance that is water soluble. Reaction proceeds fastest in warm temperature.

NaOH has ability to absorb moisture from the air – Thus it is a good desiccating agent.

NaOH is made by electrolysis of NaCl dissolved in water. H₂ collects at the cathode and Cl₂ collects at the anode.

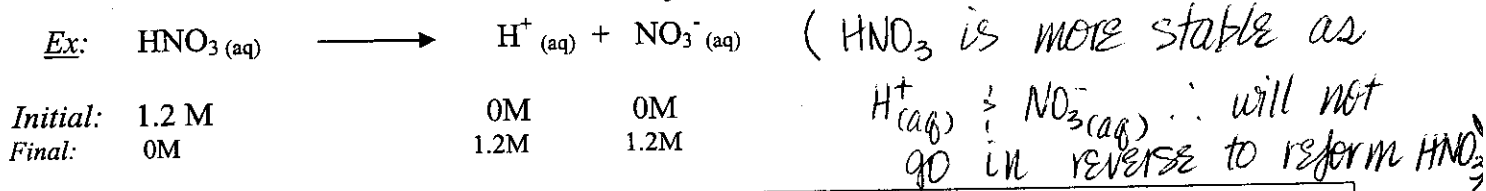


IV. Strengths of Acids & Bases:

A. Strong & Weak Acids:

- Strength of an acid depends on the amount of H^+ ions produced per mole of acid.

1) **Strong Acid:** ionizes ^(100%) completely in water.
 Reaction goes to completion. NO equil^m exists.



Common Strong Acids:	HNO ₃	HBr	HCl	HI	HClO ₄	H ₂ SO ₄
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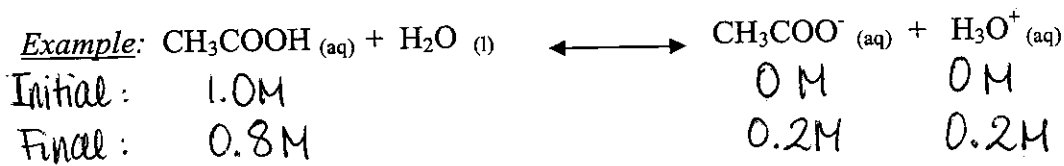
The Leveling Effect:

- Table shows all strong acids dissociating 100% to form $H_3O^+(aq)$. Water is said to have "leveled" all the strong acids to the same strength, they are all essentially sol^{ns} of H_3O^+ - This means that H_3O^+ is the strongest acid that can exist.

ie. ALL the strong acids have **identical strengths**. Their chronological order is purely arbitrary.

- Same can be said for the two strong bases at bottom of table. O^{2-} and NH_2^- cannot exist on their own in water, they immediately react with H_2O to form NH_3 and OH^- in solⁿ - Again, H_2O has "leveled" the strong bases to the same strength, all sol^{ns} of OH^- .

2) **Weak Acids:** do not ionize/dissociate completely in water.
 Equilibrium exists between the ions and the undissociated molecule.

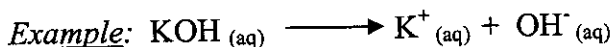


B. Strong & Weak Bases:

1) **Strong Base:** ionizes completely when dissolved in water. Goes to completion.
 Presence of the OH⁻ ion gives solution its basic characteristic.

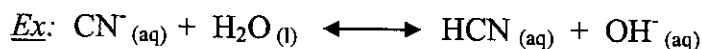
Common Strong Bases: Bases formed from alkali metals are strong.

NaOH	KOH	LiOH	CsOH	NH ₂ ⁻¹	O ⁻²
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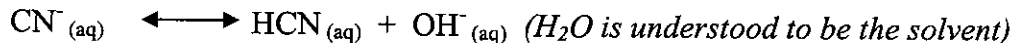


2) **Weak Bases:** do not completely ionize.

Equil^m exists. A large amount of the compound has not dissociated.



NOTE: It is also acceptable to write the above equil^m as:



C. Comparing Strengths of Acids and Bases:

- 1) Left side of "Relative Strength of Acid and Bases" table: species acting as acids
2) Right side of table: species acting as bases

When comparing strengths of species:

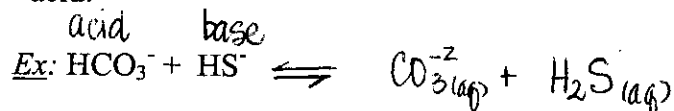
- a) The stronger an acid is, the weaker its conjugate base is & vice versa.

Ex: HIO_3 is strong acid, therefore its conjugate base, IO_3^- is weak. \checkmark
 HIO_3 is a strong acid, so HIO_3 must be a weak base! \times

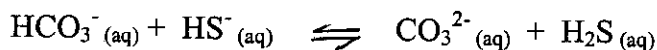
Ex: HCO_3^- and HPO_4^{2-} : both species are amphoteric.
This means they appear on both LEFT and RIGHT sides of the table.

- i) Relative ACID strength: HCO_3^- appears higher than HPO_4^{2-} on left side - HCO_3^- stronger acid.
ii) Relative BASE strength: HPO_4^{2-} appears lower on right side - HPO_4^{2-} stronger base.

NOTE: All acid-base rxns. can be seen as competing for H^+ . The stronger base "wins" and gets the H^+ from the reacting species, which automatically "defaults" as the acid.

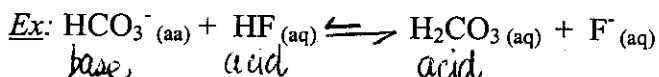


Since HS^- is the stronger base, it "wins" out, and HCO_3^- acts as the acid.



Le Chatelier's Principle & Neutralization Rxn.:

A strong acid will more likely want to donate H^+ than a weaker acid. Therefore, at equil^m, rxn. will favour the direction of the dissociation of the stronger acid.

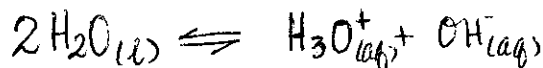


- a) Compare strengths of the acids in the rxn. $\text{HF} > \text{H}_2\text{CO}_3$ - therefore, HF more likely to donate H^+ .

D. Ionization of Water and K_w :

- H_2O is considered a non-electrolyte, meaning it won't conduct electricity. However, a VERY small amount of H_2O molecules do dissociate to form ions.
- $[H_3O^+]$ in pure water is only about $1.0 \times 10^{-7} M$.

This means that aside from the acid or base dissociating, our solvent is also slightly dissociating into ions.



NOTE: The above reaction is in equilibrium. This means that it too has an **Equilibrium Constant**.

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

$$K_{eq} [H_2O]^2 = [H_3O^+][OH^-]$$

$$K_w = [H_3O^+][OH^-] = 1.00 \times 10^{-14}$$

NOTE: $[H_2O]$ is constant within the uncertainty associated with changes of weak electrolytes.

In a **neutral** solution of water, the mole ratio of OH^- to H_3O^+ is 1:1, therefore, there will be as many H^+ as there are H_3O^+ , thus an equal concentrations of each ion.

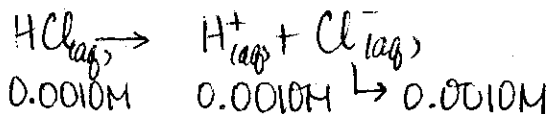
$$[H^+] = [H_3O^+] = 1.0 \times 10^{-7} M$$

The addition of either an acid or base to H_2O will disturb this 1 H^+ :1 OH^- ratio, therefore, resulting in a shift in equilibrium:

$[H_3O^+] > [OH^-]$ - Acidic solution
$[H_3O^+] < [OH^-]$ - Basic solution
$[H_3O^+] = [OH^-]$ - Neutral solution

Ex: What is the $[H^+]$ and $[OH^-]$ for a solⁿ of 0.0010M HCl?

$$[H^+] = 0.0010M$$

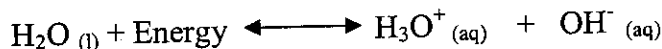


$$K_w = [H^+][OH^-] = 1.0 \times 10^{-14}$$

$$1.0 \times 10^{-14} = (0.0010M)[OH^-]$$

$[OH^-]$ comes from H_2O .

Temperature Dependence of H_2O :



Increase in Temperature: Shifts equilibrium right, and increases concentrations of $[H^+] \neq [OH^-]$. $\therefore K_w$ increases.

Decrease in Temperature: shifts reaction left and K_w decreases.

NOTE: A change of 1 pH value is equivalent to 10X the concentration change of an acid/base.

Ex: What is the pH of a 0.010M nitric acid solⁿ?

$$[H^+] = 0.010M \text{ (Strong acid } \therefore \text{ dissociates completely).}$$

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

$$pH = -\log(0.010M) \Rightarrow pH = 2.00$$

Ex: A solⁿ of ~~phosphoric~~ perchloric acid has a pH of 3.00.

(a) What is $[H^+]$ of the solⁿ?

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

$$\Rightarrow [H^+] = \text{antilog}(-pH)$$

$$[H^+] = \text{antilog}(-3.00) = 0.0010M$$

(b) What is the pOH of the solⁿ and its $[OH^-]$?

$$pH + pOH = 14.00$$

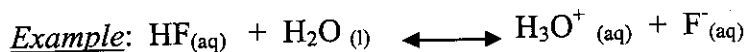
$$3.00 + pOH = 14.00 \Rightarrow pOH = 11.00$$

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

$$[OH^-] = \text{antilog}(-11.00) = 1.00 \times 10^{-11}M$$

F. Weak Acid Equilibrium & K_a :

- For weak acids, ionization is not complete, and \therefore it is NOT possible to use the initial conc. of the acid in the solution.
- For weak acids, the incomplete ionization results in equilibrium.



The equilibrium expression is as follows:

$$K_{eq} = \frac{[H_3O^+][F^-]}{[H_2O][HF]}$$

$$K_{eq} [H_2O] = K_a = \frac{[H_3O^+][F^-]}{[HF]}$$

$$K_a = \frac{[\text{H}_3\text{O}^+][\text{F}^-]}{[\text{HF}]}, \text{ where } K_a - \text{Acid Ionization Constant for Weak Acids.}$$

NOTE: The higher the K_a value, the greater the ionization for the particular acid, \therefore the stronger the acid.
Also remember that as temperature changes, the ability to ionize changes and thus the conditions change.

Example: What pH results when 0.25 mol of acetic acid is dissolved in enough water to make a 1L solution?

1. Write the balanced dissociation equation. $\text{CH}_3\text{COOH}_{(aq)} \rightleftharpoons \text{CH}_3\text{COO}^-_{(aq)} + \text{H}^+_{(aq)}$
0.25M

2. Look up the ionization constant, K_a , for acetic acid:

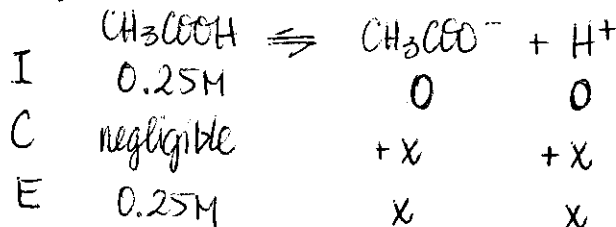
$$K_a = \frac{[\text{CH}_3\text{COO}^-][\text{H}^+]}{[\text{CH}_3\text{COOH}]} = 1.8 \times 10^{-5}$$

3. [solution] = 0.25M

When acetic acid ionizes, a small amount of the molecules dissociates to make the ions. For the acetic acid concentration, this dissociated amount is small enough to be ignored to the change of $[\text{CH}_3\text{COOH}]$. \therefore the concentration of acetic acid can be considered the same @ 0.25M.

5% Rule:

$$K_a = \frac{[\text{H}_3\text{O}^+][\text{CH}_3\text{COO}^-]}{0.25\text{M}} = 1.8 \times 10^{-5}$$



Mole ratio of hydronium ion to acetate ion is 1 - 1. \therefore simply square the values.

$$1.8 \times 10^{-5} = \frac{x^2}{0.25\text{M}} \Rightarrow *x = 0.0021213 = [\text{H}^+]$$

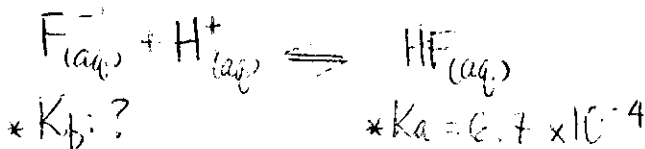
$$\text{pH} = -\log(0.0021213) = 2.67$$

G. K_a & K_b for Conjugate Pairs:

$$K_a \times K_b = K_w$$

Example: Determine K_b for the weak base F^- .

Notice that K_b for F^- is unavailable. But its **conjugate acid HF**, is available, with a K_a value of 6.7×10^{-4} .



$$\begin{aligned} * 1.0 \times 10^{-14} &= K_a \times K_b \\ \frac{1.0 \times 10^{-14}}{6.7 \times 10^{-4}} &= 1.5 \times 10^{-11} \end{aligned}$$

H. Weak Base Equilibrium & K_b :

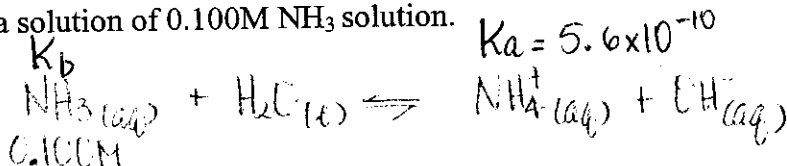
- **Weak Base:** does NOT ionize completely in aqueous solutions. Different bases have varying degrees of ionizability.
- Every weak base also has a corresponding **Base Ionization Constant, K_b** .



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

Example: Determine the pH for a solution of 0.100M NH_3 solution.

1. Balanced equation:



2. Construct an ICE table. (This could also have been done for the weak acid example).

	NH_3	NH_4^+	OH^-
I	0.100M	0 M	0 M
C	negligible	+ x	+ x
E	0.100M	x	x

NOTE!! Because NH_3 is a weak base, we can assume that the ionization is minimal enough that the initial [] is roughly equal to its final [].

The above assumption is true for any weak acid/base if the initial concentration of that acid or base is $1000X >$ than K .

* K_b of $\text{NH}_3 \Rightarrow$ Conjugate acid of NH_3 is NH_4^+

$$* K_b = \frac{1.0 \times 10^{-14}}{5.6 \times 10^{-10}} = 1.8 \times 10^{-5}$$

$$* 1.8 \times 10^{-5} = \frac{x^2}{0.100M}$$

$$* x = 0.001336M = [\text{OH}^-]$$

$$* K_a = 5.6 \times 10^{-10}$$

$$\text{pOH} = -\log [\text{OH}^-] = 2.87$$

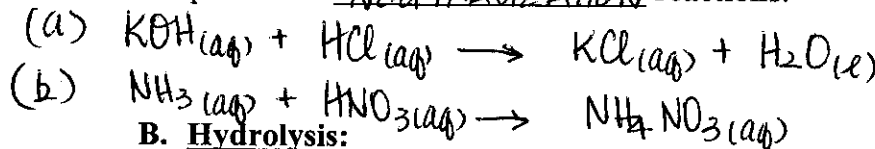
$$\text{pH} = 11.13$$

IV. Typical Reactions of Acids & Bases:

A. Formation of Salts (Neutralization Reactions):

NOTE: Neutralizⁿ Rxns. do not go to equilibrium.

• Salts: are the product of NEUTRALIZATION reactions.



* $KCl(aq)$ & $NH_4NO_3(aq)$ are salts of Neutralization Rxns.

B. Hydrolysis:

Hydrolysis: when the ions of a soluble salt reacts further with H₂O.
 The resultant ~~equilibrium~~ solution may be acidic, or basic, or neutral.

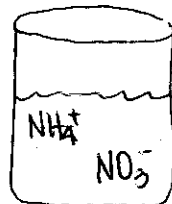
Example: Predict whether or not the salts of KCl, NH₄NO₃, and NaHCO₃ hydrolyze. If they do hydrolyze, determine whether the solution is acidic, basic or neutral.



CATION RULE: Any cation from Family 1 and 2, are ions of strong bases, and will NOT hydrolyze.

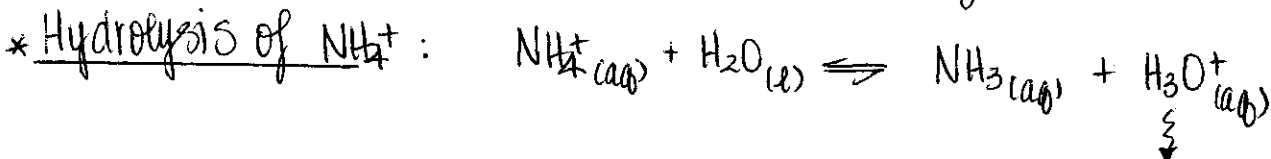
ANION RULE: Conjugate bases of strong acids do NOT hydrolyze. And this is the case of Cl⁻. It is the conjugate of the strong acid HCl.

∴ KCl will not hydrolyze.



NH₄⁺: weak acid - will hydrolyze.

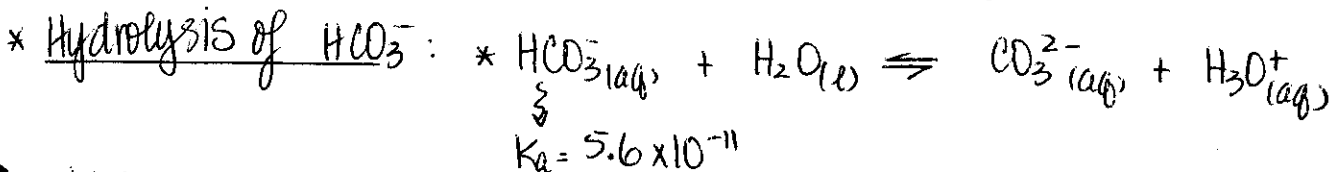
NO₃⁻: ion from dissociation of HNO₃ - strong acid - NO HYDROLYSIS.



* Resulting solⁿ is acidic.

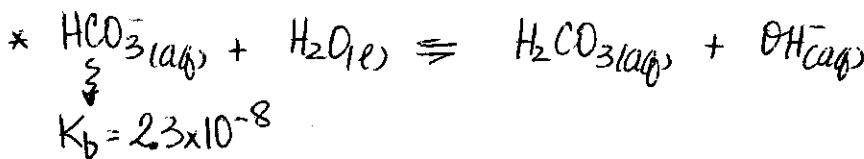


Na⁺: ion from dissociation of NaOH - STRONG BASE - No hydrolysis
 HCO₃⁻: weak species - Hydrolysis. (AMPHIPROTIC)



$K_b > K_a$

Therefore, solⁿ will be basic.



Example: Predict whether the solution of the salt NH_4NO_2 will be acidic, basic, or neutral.

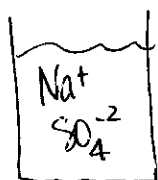
NOTE: Some salts contain BOTH a cation and an anion that will hydrolyze. In such cases, the resulting solution may be acidic, basic, or neutral, depending on the relative extent of the hydrolysis.

- NH_4^+ : will hydrolyze - $\text{NH}_4^+(\text{aq}) + \text{H}_2\text{O}(\text{l}) \rightleftharpoons \text{NH}_3(\text{aq}) + \text{H}_3\text{O}^+(\text{aq})$
- NO_2^- : will hydrolyze - $\text{NO}_2^-(\text{aq}) + \text{H}_2\text{O}(\text{l}) \rightleftharpoons \text{HNO}_2(\text{aq}) + \text{OH}^-(\text{aq})$

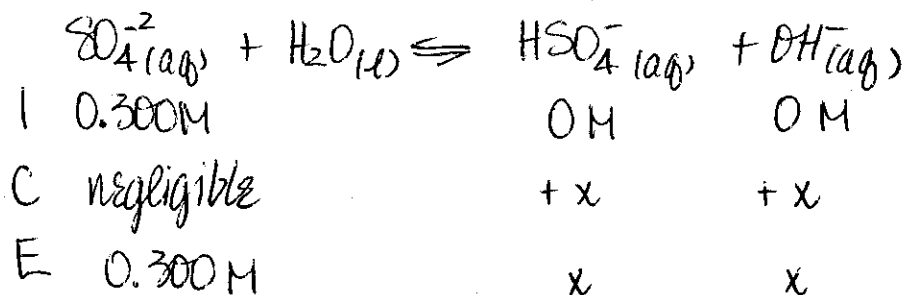
To determine the overall acidity or basicity of the solution, compare K_a to K_b values.

$$\left. \begin{array}{l} * K_a \text{ NH}_4^+ : 5.6 \times 10^{-10} \\ * K_b \text{ NO}_2^- : 2.2 \times 10^{-11} \end{array} \right\} * K_a > K_b \therefore \text{sol.}^{\sim} \text{ will be acidic}$$

Ex. Determine the pH of a 0.300M $\text{Na}_2\text{SO}_4(\text{aq})$ solution.



- * Na^+ : no hydrolysis.
- * SO_4^{2-} : hydrolysis.



$$K_b = 8.3 \times 10^{-13} = \frac{[\text{HSO}_4^-][\text{OH}^-]}{0.300\text{M}}$$

$$8.3 \times 10^{-13} = \frac{x^2}{0.3\text{M}} \Rightarrow * x = 5.0 \times 10^{-7} \text{M} = [\text{OH}^-]$$

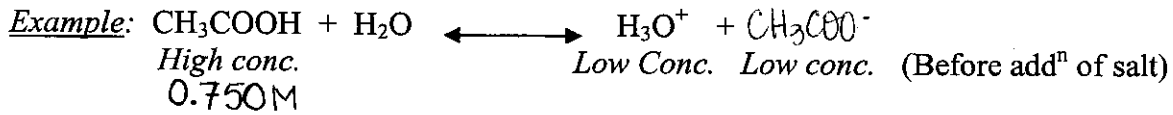
$$* \text{pH} = 7.700$$

V. Buffer Solutions:

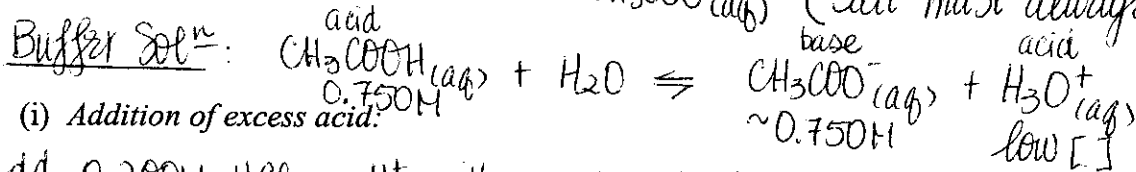
- **Buffer solutions:** equilibrium system that maintains a relatively constant pH when excess amounts of acid or base is added.

A. Buffer Preparation & Action:

- **Buffers:** can be prepared by combining appreciable amounts of a weak acid/base with its conjugate ion.



Salt to be added: 0.750M $\text{NaCH}_3\text{COO}_{(aq)}$ (Salt must always be soluble)



NOTE: Acids will always seek out bases to react w/

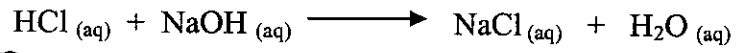
* Add 0.200M HCl: H^+ will react w/ the CH_3COO^- in the buffer solⁿ. This $\downarrow [\text{CH}_3\text{COO}^-]$ to 0.550M - This causes buffer equil^m to slightly shift right, but not enough.

to drastically $\Delta [\text{H}_3\text{O}^+]$ \therefore pH remains relatively the same.

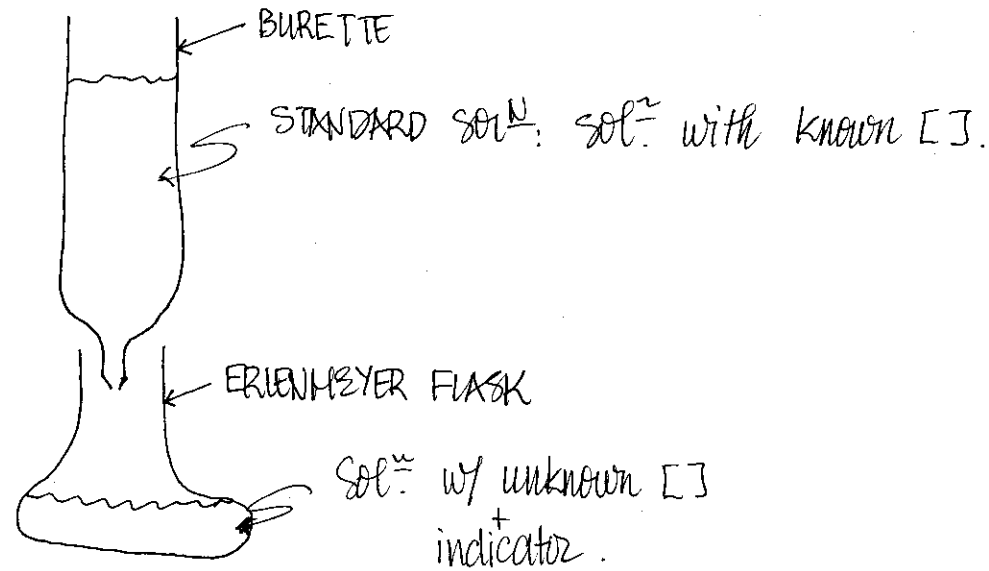
(ii) Addition of excess Base:

* Add 0.200M NaOH: OH^- will react w/ CH_3COOH - This lowers $[\text{CH}_3\text{COOH}]$ to 0.550M, this will cause for a shift to the left, but not drastic enough to significantly alter $[\text{H}_3\text{O}^+]$, \therefore pH remains relatively the same.

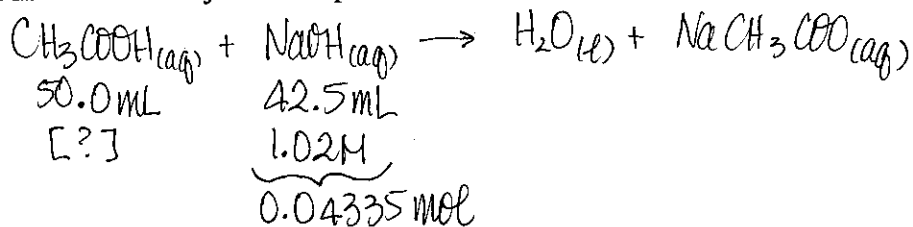
VI. Titration:



- **Equivalence Point:** the point at which the mol ratio of reactants in the beaker is equivalent to the ratio described by the balanced equation.
- In a rxn. of a **strong acid & strong base**, there is a dramatic change in pH at the equivalence point - This fact aides in the finding of the equivalence point - Thus **titration** is born.
- **Titration:** uses indicators to "cue" the equivalence point. They all change colour at the equivalence point and are thus used for that purpose. Titrations can be used to find out the concentration of an unknown acid/base.
- **Standard Solutions:** a solution of known concentration.



Example: It is found that 42.5 mL of 1.02M NaOH have been added to 50.0 mL of vinegar when the phenolphthalein in the solution just turns pink. What is the [vinegar]?



* $0.04335\text{ mol NaOH} \times \frac{1\text{ CH}_3\text{COOH}}{1\text{ NaOH}} = 0.04335\text{ mol CH}_3\text{COOH}$ was originally in the Erlenmeyer flask.

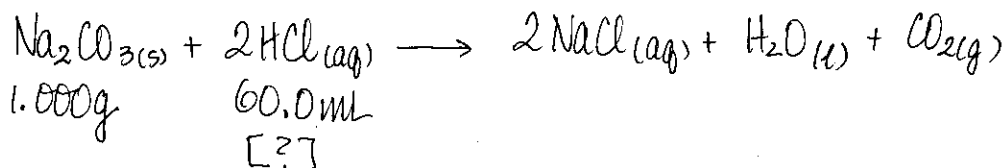
* $0.04335\text{ mol CH}_3\text{COOH} \div 0.0500\text{ L} = 0.867\text{ M CH}_3\text{COOH}$.

(i) **Primary Standards:** when a highly pure substance of known composition is used to react with a substance to be standardized.



A known mass of sodium carbonate is added to water and an indicator is added. This primary standard is then titrated with the HCl solution. The balanced equation allows you to determine the moles of HCl in solution and hence, the [HCl].

Example: A solution of HCl is standardized using pure $\text{Na}_2\text{CO}_3_{(s)}$ as a primary standard. What is the [HCl] solution if 60.00 mL of this solution are needed to titrate a 1.000g sample of sodium carbonate?



* $1.000\text{ g Na}_2\text{CO}_3 \times \frac{\text{mol}}{106\text{ g}} = 0.00943396\text{ mol Na}_2\text{CO}_3 \times \frac{2\text{ HCl}}{1\text{ Na}_2\text{CO}_3} = 0.01886\text{ mol HCl}$

* $\frac{0.0188679\text{ mol Na}_2\text{CO}_3}{0.06000\text{ L}} = 0.3145\text{ M HCl}$.

↳ This solution of HCl will eventually be used as a standard solⁿ for a subsequent titration.

(ii) **Equivalence Point vs. Transition Point:**

- **Equivalence Point:** refers to the end point of the titration.
- **Transition Point:** refers to the point where the indicator changes colour.
- Ideally, the end and equivalence point should be reached at the same time.

(a) Strong Acid/Strong Base Titration Curves:

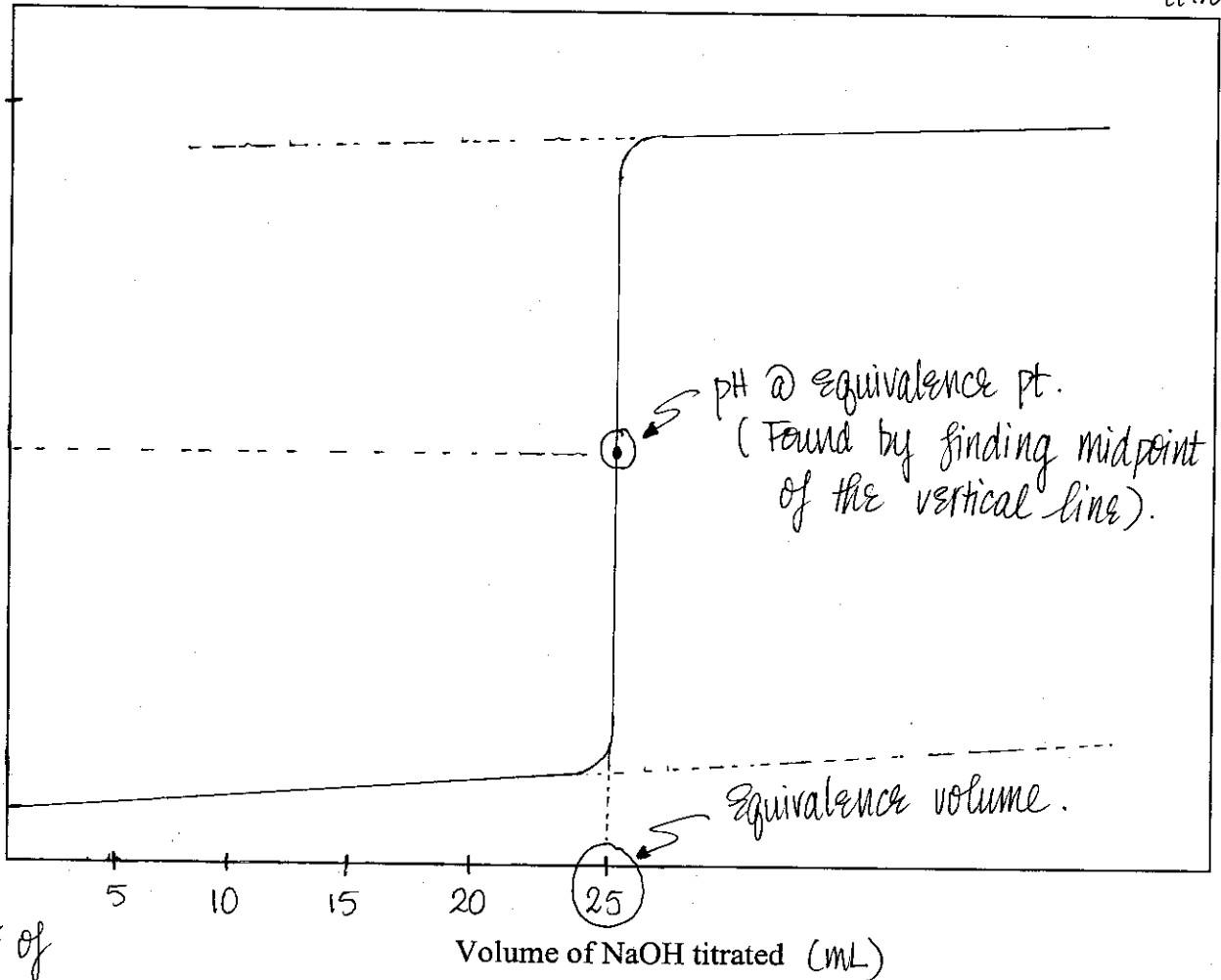
NaOH + HCl titration curve (Indicator: Phenolphthalein)

* Bad choice b/c transition pt. is higher than that of \rightarrow trans. pt. for this titration.

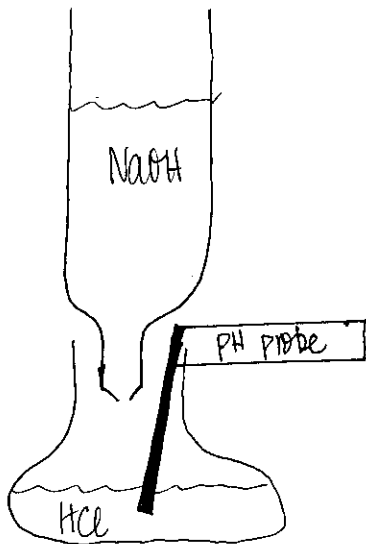
* pH of NaOH once titration is overshoot.

* Neutral b/c @ equiv. pt. salt does not do hydrolysis.

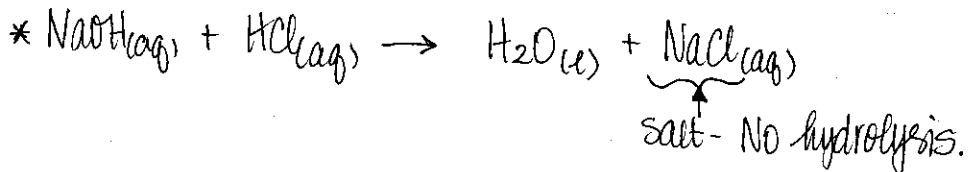
* pH of HCl before addⁿ of NaOH.



- For a strong acid/strong base titration, the equivalence point is always pH = 7. This means that an indicator whose transition point is 7 should ideally be used.



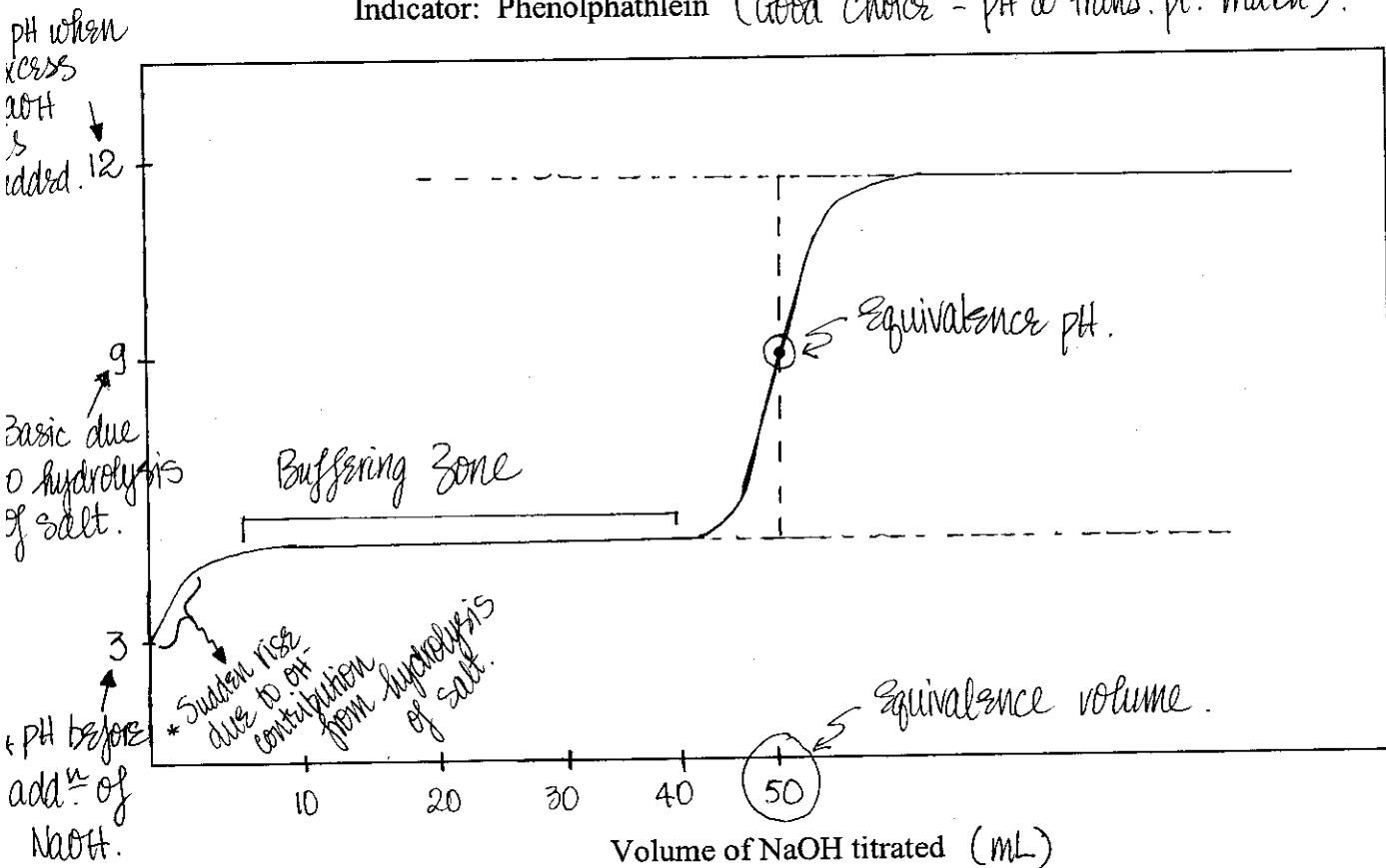
* As base is added, the pH probe will monitor the change in $[H^+]$ in the Erlenmeyer flask, thus pH will raise over time, due to a continual decrease in $[H^+]$, until all H^+ ions have been neutralized by the NaOH. - This is the EQUIVALENCE PT.



(b) Weak Acid/Strong Base Titration Curves:

CH₃COOH + NaOH titration curve:

Indicator: Phenolphthalein (Good choice - pH @ trans. pt. match).

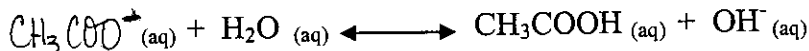


NOTE: A weak acid is capable of demonstrating buffering action when a base is added. This buffer action is what causes for the gradual rise in pH in this region of the curve.

At the Equivalence Point: The solution is not neutral!



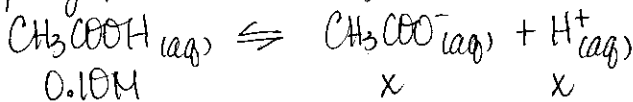
The salt sodium acetate has been formed at the equivalence point. Because it is in an aqueous environment, it will dissociate to give Na⁺ and CH₃COO⁻. The CH₃COO⁻ ion will hydrolyze in the aqueous environment to give the following ions:



- It is because of the OH⁻ ions produced that causes the solution to be slightly basic.

Example: Determine the pH at the following points when 50.0mL of 0.10M acetic acid is titrated with 0.10M NaOH.

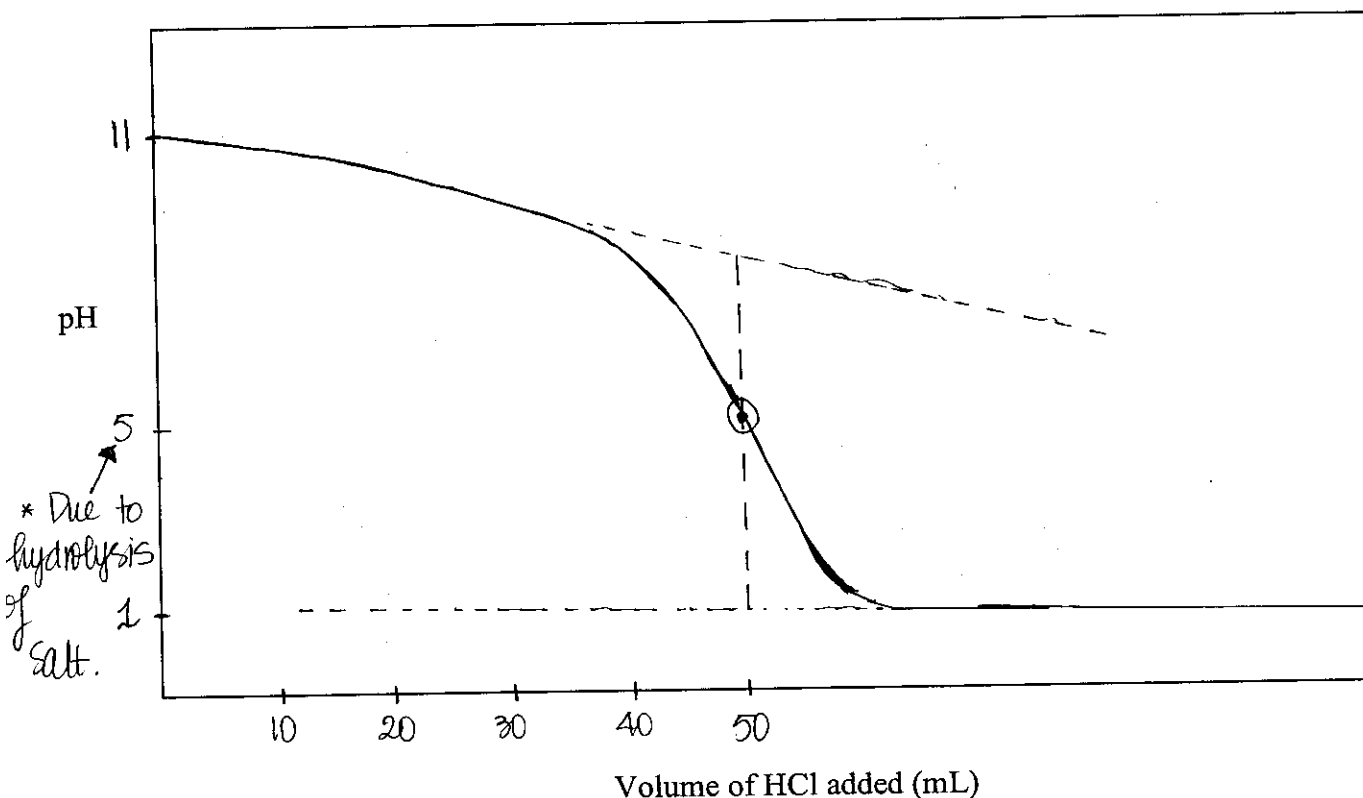
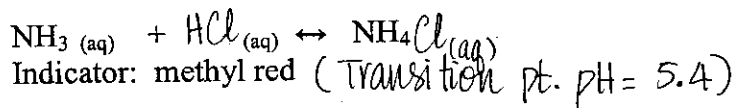
(a) The pH of the solution of acetic acid when no NaOH is yet added. (So, we are essentially just measuring pH of the acid alone).



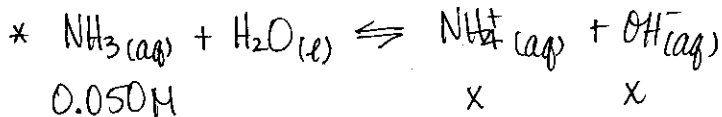
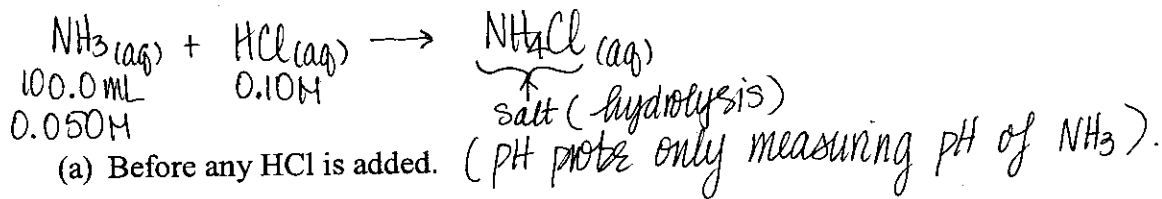
$$* 1.8 \times 10^{-5} = \frac{x^2}{0.10M}$$

$$* x = 0.00134 \Rightarrow \text{pH} = 2.87$$

(c) Strong Acid/Weak Base Titration:



Example: Determine the pH at the following points of the titration between 100.0mL of 0.050M NH_3 with 0.10M HCl.



$$K_b = 1.79 \times 10^{-5} = \frac{x^2}{0.050\text{M}}$$

* $x = 9.449 \times 10^{-4} \text{ M} = [\text{OH}^-]$
 * $\text{pOH} = 3.02 \Rightarrow \text{pH} = 10.97$

(b) After 25.0mL of HCl is added.
 ↳ This is the mid-point of the titration, since total volume of HCl needed is 50.0mL.

$$\therefore K_a = [H_3O^+]$$

Example: Determine the K_a for the chemical indicator phenolphthalein.

Phenolphthalein pH range: 8.2 – 10.6

Assume the transition point lies in between these two pH values.

$$* \text{ pH @ TRANSITION PT. : } \frac{8.2 + 10.0}{2} = 9.1$$

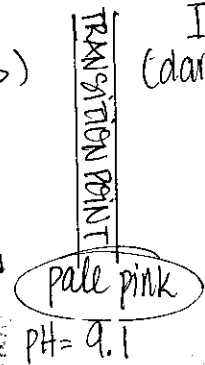
$$* K_a = [H_3O^+] \text{ @ TRANSITION PT.}$$

$$\therefore K_a = \text{antilog}(-9.1) = 7.9 \times 10^{-10}$$

Phenolphthalein:

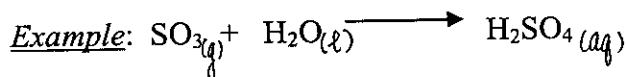
HI_n (colourless) I_n^- (dark pink)

This means phenolphthalein will begin to change colour when a pH = 9.1 is detected in solⁿ.

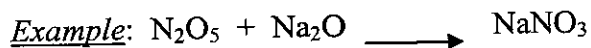
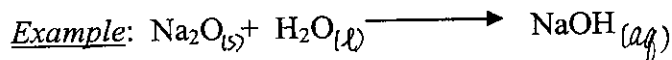


VIII. Acidic & Basic Anhydrides:

- **Acidic Anhydrides:** an oxygen – containing covalent compound that reacts with water to produce an acidic solution.



- **Basic Anhydrides:** an oxygen containing ionic compound that reacts with water to give a basic solution.



Application of Oxides – Acid Rain:

- Normal rainwater has a pH close to 6 due to the natural reaction of H₂O with CO₂ gas from the environment, forming dilute carbonic acid:
- **Acid Rain:** has a pH below 5.6 due mainly to reaction of the rainwater with gaseous oxides of industrial processes:

Industrial Process	Reaction with Rain Water	Resulting Acidic Sol ⁿ :
<p>1) Combustion of <u>fossil</u> fuels:</p> $C_{(s)} + O_{2(g)} \rightarrow CO_{2(g)}$ <p><u>octane</u></p> $2 C_8H_{18(l)} + 25 O_{2(g)} \rightarrow 16 CO_2 + 18 H_2O$ $C_2H_5OH + 3 O_2 \rightarrow 2 CO_2 + 3 H_2O$ <p><u>ethanol</u></p>	<ul style="list-style-type: none"> • $CO_{2(g)} + H_2O_{(l)} \rightarrow H_2CO_{3(aq)}$ 	<p><u>carbonic acid</u></p> <ul style="list-style-type: none"> • $H_2CO_{3(aq)} \rightleftharpoons HCO_3^-(aq) + H^+(aq)$
<p>2) Roasting of Sulfur Ores and Coal Fired Power Stations:</p> $2 ZnS_{(s)} + 3 O_{2(g)} \rightarrow 2 ZnO_{(s)} + 2 SO_{2(g)}$ <p>NOTE: SO₂ is also produced industrially by the Contact process used to make <u>sulfuric</u> acid:</p> $S_{(s)} + O_{2(g)} \rightarrow SO_{2(g)}$ $SO_{2(g)} + O_{2(g)} \rightarrow SO_{3(g)}$	<ul style="list-style-type: none"> • $SO_{2(g)} + H_2O_{(l)} \rightarrow H_2SO_{3(aq)}$ • $SO_{3(g)} + H_2O_{(l)} \rightarrow H_2SO_{4(aq)}$ 	<p><u>sulfurous acid</u></p> <p><u>sulfuric acid</u></p> <ul style="list-style-type: none"> • $H_2SO_{3(aq)} \rightleftharpoons HSO_3^-(aq) + H^+(aq)$ • $H_2SO_{4(aq)} \rightleftharpoons HSO_4^-(aq) + H^+(aq)$
<p>3) Incomplete combustion Inside Engines at High Temperatures:</p> $N_{2(g)} + O_{2(g)} \rightarrow 2 NO_{(g)}$ $2 NO_{(g)} + O_{2(g)} \rightarrow 2 NO_{2(g)}$	<ul style="list-style-type: none"> • $NO_{(g)} + H_2O_{(l)} \rightarrow HNO_{2(aq)}$ • $NO_{2(g)} + H_2O_{(l)} \rightarrow HNO_{3(aq)}$ 	<p><u>nitrous acid</u></p> <p><u>nitric acid</u></p> <ul style="list-style-type: none"> • $HNO_{2(aq)} \rightleftharpoons NO_2^-(aq) + H^+(aq)$ • $HNO_{3(aq)} \rightarrow NO_3^-(aq) + H^+(aq)$

Negative Outcomes of Acid Rain:

- 1) Buildings or statues made of solid CaCO₃ (limestone) will gradually dissolve:
- 2) Leach metals from soils into ground water and lakes and rivers, poisoning aquatic life.
(eg. Al)
- 3) Can disrupt photosynthesis in plants. Low acid concentrations will slow down the process of photosynthesis and concentrated levels of acid will kill the plant.

