C12 - 4.0 - Acids & Bases

$$HCl_{(aq)} + NaOH_{(aq)} \rightarrow NaCl_{(aq)} + H_2O_{(l)}$$

$$Acid \ + \ Base \ \rightarrow \ Salt \ + \ Water$$

$$H^+ + OH^- \rightarrow \mathrm{H_2O}$$

NIE

Salt - The neutralization product which results when an acid and a base react.

Acid - Any substance that releases $H^+_{(aq)}$ in water.
-React with Bases/Are Electrolytes
-Turn Litmus Paper Red
-Taste Sour (Vinegar, Lemon Juice)
-Act on some metals to produce $H_{2(g)}$

Base - Any substance that releases $OH^-_{(aq)}$ in water.
-React with Acids/Are Electrolytes
-Turn Litmus Paper Blue
-Taste Bitter (Baking Soda)
-Feel Slippery

$$Mg_{(s)} + 2HCl_{(aq)} \rightarrow MgCl_{2(aq)} + H_{2(g)}$$

Properties : **Common Uses :**

Sulphuric Acid - H_2SO_4 (Battery Acid)

-Good dehydrating agent
-Strongly exothermic acid when mixed with water
-Concentrated form chars some types of organic material (ie. Sugars)
-Reacts with some metals (Often slowly.)
-Good Electrolyte (Conducts Electricity)
-Concentrated Sulphuric Acid is 98% H_2SO_4 and 2% water. $(18\ M\ H_2\ SO_4)$

-Production of Sulphates
-Manufacturing Fertilizers/Dyes/ Insecticides/ Detergents/Plastics
-Used to absorb water and keep chemicals & nonaqueous solutions free of water
-Used in car batteries as an electrolyte

Hydrochloric Acid - HCl (Muriatic Acid)

-Good electrolyte
-Concentrated solutions have a coking odor
-Reacts with some metals (Often slowly.)
-Concentrated Hydrochloric Acid is 97% HCl and 2% water. (12 $M\ HCL$)

-Production of Chlorides
-Cleaning metal products (removes metal oxides) and bricks.
-Catalyst in some Chemical reactions
-"Stomach acid" is a dilute solution of HCl, which activates a protien-digesting biological catalyst called an "enzyme."
-Removing "Boiler scale", which consists of calcium and magnesium carbonate.

Nitric Acid - HNO_3

-Colours protein yellow
-Very reactive, quickly attacks almost all metals
-Concentrated Nitric Acid is 69% HNO_3 in Water (16 $M\ HNO_3$)

-Production of Nitrates
-Manufacturing Fertilizers/Explosives/Dyes

Sodium Hydroxide - $NaOH$ (Caustic Soda/lye)

-Very corrosive (Caustic) to animal and plant tissue
-Highly exothermic reaction when mixed with water.
-Rapidly "deliquesces"; that is ; absorbs H_2O from air.
-Rapidly absorbs $CO_{2(g)}$ from the air to form carbonates.

$$NaOH_{(s)} + CO_{2(s)} \rightarrow NaHCO_{3(s)}$$

-Sodium salts
-Soap/Cleaning Products ie. Oven/Drain
-Manufacturing Glass/Pulp & Paper/Plastics/Aluminum
-Neutralizing acids during industrial reactions

Potassium Hydroxide - KOH (Caustic Potash)

-Similar to $NaOH$, but melts at a lower temperature.

-Manufacturing liquid soap/Absorbing $CO_{2(g)}$
-Making potassium salts
-Electrolyte in alkaline batteries

Ammonia - NH_3 ($NH_{3(aq)}$ Ammonium Hydroxide)

-Colourless, alkaline, highly toxic, corrosive gas with pungent odor.
-highly soluble in water
-exothermic reaction when dissolved in water

-Manufacturing Nitric Acid/Explosives/Fertilizers/Synthetic fibres
-Refrigeration gas

Acetic Acid - CH_3COOH (5% aq solution called "vinegar")

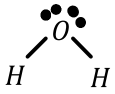
-Non-Electrolyte when concentrated (99-100%, 17M) ; weak electrolyte when diluted
-Only affects highly reactive metals

-Making Acetates
-Food preservation (Pickles)
-Manufacturing Textiles and Plastics

C12 - 4.0 - Acid & Base

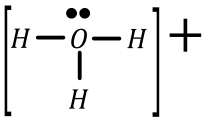
$Proton_{radius} = 2.5 \times 10^{-14}$

Water



$H^+ + H_2O \rightarrow H_3O^+$

"The proton"



"Hydronium Ion" Or "Hydrated Proton"

$HCl_{(g)} \rightarrow H^+_{(aq)} + Cl^-_{(aq)} \longrightarrow HCl_{(g)} + H_2O_{(l)} \rightarrow H_3O^+_{(aq)} + Cl^-_{(aq)}$

Arrhenius - An Acid donates a H^+.
- A Base donates an OH^-.

Bronsted - An Acid is a proton donor.
Lowrey - A Base is a proton acceptor.

(lone electron pair)
Lewis - An Acid is a LEP acceptor.
- A Base is a LEP donor.

Acid - A substance that donates a proton to another substance. (Donor - gives away an H^+)
Base - " accepts " from " . (Acceptor - receives an H^+)

Can supply :
Monoprotic Acid - one proton.
Diprotic Acid - two protons.
Triprotic Acid - two protons.
Polyprotic Acid - more than one*

Amphiprotic - Acts as an Acid or a Base $\boxed{H_2PO_4^-, HS^-, HCO_3^-}$

$NH_3 + H_2O \rightleftharpoons NH_4^+ + OH^-$ Water acts as an Acid

$CH_3COOH + H_2O \rightleftharpoons CH_3COO^- + H_3O^+$ Water acts as a Base

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

Conjugate Pair	Conjugate Acid	Conjugate Base
NH_4^+, NH_3	NH_4^+	NH_3
H_3O^+, H_2O	H_3O^+	H_2O
$H_2PO_4^-, HPO_4^{2-}$	$H_2PO_4^-$	HPO_4^{2-}
	Extra Proton	Lacks a Proton

Strong Acid or Base is 100% ionized in solution. $HCl_{(g)} \rightarrow H^+_{(aq)} + Cl^-_{(aq)}$
Weak " less than 100% " . $NH_{3(aq)} + H_2O_{(l)} \rightleftharpoons NH^+_{4(aq)} + OH^-_{(aq)}$

Weak and Strong refer to % of ionization.
Dilute and concentrated refer to molarity of a solution.
$10\ M\ HF_{(aq)}$ is concentrated and weak
$0.001\ M\ HCl$ is dilute and strong.

The Leveling Affect
- All strong acids are 100% dissociated in aqueous solution and are equivalent to solutions of $H_3O^+_{(aq)}$,
-All strong bases " " " " " " $OH^-_{(aq)}$.

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

$K_{H_2O} = [H^+][OH^-] = 1.00 \times 10^{-14}\ @25^oC$

$K_{H_2O} = [H_3O^+][OH^-] = 1.00 \times 10^{-14}\ @25^oC$

Acid Ionization - Reaction of a weak Acid with water. K_a - Acid Ionization Constant.
Base Ionization - " Base " . K_b - Base " .

Acid Ionization	Base Ionization
$NH^+_{4(aq)} + H_2O_{(l)} \rightleftharpoons NH_{3(aq)} + H_3O^+_{(aq)}$	$NH_{3(aq)} + H_2O \rightleftharpoons NH^+_{4(aq)} + OH^-_{(aq)}$

$HReact + Prod^{-1} \rightleftharpoons React^- + HProd$

$$K_{eq} = \frac{K_a(reactant\ acid)}{K_a(product\ acid)} = \frac{[products]}{[reactants]}$$

$HClO_3^- + HS^- \rightleftharpoons CO_3^{2-} + H_2S$

$HClO_3^- \rightarrow CO_3^{2-} + H^+$
$HS^- + H^+ \leftarrow H_2S$ (Stronger)
$HClO_3^- + HS^- \rightleftharpoons CO_3^{2-} + H_2S$

C12 - 4.0 - Acid & Base pH

$pH = -\log_{10}[H_3O^+]$
$pOH = -\log_{10}[OH^-]$ $pH + pOH = 14$ Only decimals are significant.

$10\ mL\ 0.100\ M\ HCl$ Reacts with : $90\ mL\ 0.100\ M\ NaOH$ Find the ph of the mixture.

$HCl + NaOH \rightarrow NaCl + H_2O$

$$[H_3O^+]_{st} = 0.100\ M \times \frac{10\ mL}{100\ mL} = 0.0100\ M$$

$H_3O^+ + OH^- \rightarrow 2H_2O$

$1 : 1 \rightarrow 1$ OH^- is in excess (xs)

$$[OH^-]_{st} = 0.100\ M \times \frac{90\ mL}{100\ mL} = 0.0900\ M$$

$[OH^-]_{xs} = [OH^-]_{st} - [H_3O^+]_{reacted}$ But $[OH^-]_{reacted} = [H_3O^+]_{reacted}$

Since H_3O^+ present reacted $[H_3O^+]_{reacted} = [OH^-]_{st}$

$$\begin{aligned}[OH^-]_{xs} &= [OH^-]_{st} - [H_3O^+]_{st}\\ &= 0.0900 - 0.0100\\ &= 0.0800\ M\end{aligned}$$

$$\begin{aligned}pOH &= -\log_{10}[OH^-]\\ pOH &= -\log_{10}[0.0800]\\ pOH &= 1.097\end{aligned}$$

$$\begin{aligned}pH &= 14 - 1.097\\ pH &= 12.903\end{aligned}$$

$$\begin{aligned}pH &= -\log_{10}[H_3O^+]\\ 12.903 &= -\log_{10}[H_3O^+]\\ 10^{-12.903} &= [H_3O^+]\\ [H_3O^+] &= 1.25 \times 10^{-13}\end{aligned}$$

Note : if H_3O^+ in excess $[H_3O^+]_{xs} = [H_3O^+]_{st} - [OH^-]_{st}$

Excess by moles method :

$$moles\ H^+ = 0.100 \frac{mol}{L} \times 0.0100\ L = 0.00100\ mol$$

$$Total\ Volume = 0.0100 + 0.0900 = 0.100\ L$$

$$moles\ OH^- = 0.100 \frac{mol}{L} \times 0.0900\ L = 0.00900\ mol$$

$$[OH^-]_{xs} = \frac{0.00800\ mol}{0.100\ L} = 0.0800\ M$$

$$moles\ OH^-\ in\ excess = 0.00900 - 0.00100 = 0.00800\ mol$$

Hydrolysis - Of a Salt is Reaction between Water and Cation/Anion/Both contained in the salt to produce an Acidic or Basic solution. (Note : All Salts are considered to ionized in water.)

Spectator Ions - The conjugates of strong Acids and Bases.

$NaOH \rightarrow Na^+ + OH^-$ $Na^+ + H_2O \neq NaOH + H^+$ Na^+ is a spectator

$HCl \rightarrow H^+ + Cl^-$ $Cl^- + H_2O \neq HCl + OH^-$ Cl^- is a spectator

Determining the behavior of a salt in water :
-Determine ions produced when salt dissociates. (Discard any spectators from further consideration.)
-Ions will react as acids id on the left side of the table and/or bases if they are on the right.

$NaCl_{(s)} \rightarrow Na^+_{(aq)} + Cl^-_{(aq)}$ Neutral, no Hydrolysis

$NaHC_2O_4 \rightarrow Na^+_{(aq)} + H^+ + C_2O^{-1}_{4(aq)}$ $\neq Chem\ 12$

$NH_4Cl_s \rightarrow NH^+_{4(aq)} + Cl^-_{(aq)}$ Cl^- is a spectator

NH_4^+ is found on the acid side of the Table so,

$NH_4^+ + H_2O \rightleftharpoons H_3O^+ + NH_3$ Hydrolysis Reaction

H_3O^+ is produced, therefore solution is Acidic.

$NaF_{(s)} \rightarrow Na^+_{(aq)} + F^-_{(aq)}$ Na^+ is a spectator

F^- is found on the base side of the Table so,

$F^- + H_2O \rightleftharpoons HF + OH^-$

OH^- is produced, therefore solution is Basic.

$NaHC_2O_4 \rightarrow Na^+_{(aq)} + HC_2O^{-1}_{4(aq)}$ Na^+ is a spectator $HC_2O_4^{-1}$ is Amphiprotic

$HC_2O_4^{-1}$ as an Acid

$HC_2O^-_{4(aq)} + H_2O_{(l)} \rightleftharpoons H_3O^+ + C_2O^{2-}_{4(aq)}$

$K_a(HC_2O_4^-) = 6.4 \times 10^{-5}$

$HC_2O_4^{-1}$ as a Base

$HC_2O^-_{4(aq)} + H_2O_{(l)} \rightleftharpoons H_2C_2O_{4(aq)} + OH^-_{(aq)}$

$$K_b(HC_2O_4^{-1}) = \frac{K_w}{K_a(H_2C_2O_4)} = \frac{1.0 \times 10^{-14}}{5.9 \times 10^{-2}} = 1.7 \times 10^{-13}$$

$K_a > K_b$; $HC_2O_4^{-1}$ is an Acidic Solution.

C12 - 4.0 - Acid & Base

Find pH $0.500\ M\ CH_3COOH$

$Mole\ Ratio\ -1:1:1:1$

$\frac{mol}{L}$	CH_3COOH	$+$	H_2O	$\rightleftharpoons$	CH_3COO^-	$+$	H_3O^+
$Start$	0.500				0		0
$+\Delta$	$-x$				$+x$		$+x$
$=Eq$	$(0.500-x)$				x		x

$$K_a=\frac{[CH_3COO^-][H_3O^+]}{[CH_3COOH]}$$

$$1.8\times10^{-5}=\frac{x^2}{(0.500-x)}$$ See Table

Assume $(0.500-x)=0.500$

$$1.8\times10^{-5}=\frac{x^2}{0.500}$$ Multiply

$$9\times10^{-5}=x^2$$ Square Root Both Sides

$$x=3.0\times10^{-3}$$

$\neq Quadform$
Assume weak Acid
is sufficiently Weak

$$[H_3O^+]=3.0\times10^{-3}$$

$$pH=-\log_{10}[3\times10^{-3}]$$
$$pH=2.52$$

Maximum of 2 sig figs in final answer. See Table.

$0.50\gg3.0\times10^{-3}$ $0.50-x=0.50-3.0\times10^{-3}\approx0.50$ $Assume:\ 0.50-x\approx0.50$

$$\%\ Dissociation=\frac{[H_3O^+]_{eq}}{[HA]_{st}}=\frac{3.0\times10^{-3}}{0.5}=0.006$$ $0.006=0.60\%<5\%$

Find K_a $pH=1.70$ $0.100\ M\ weak\ acid\ HA$

$Mole\ Ratio\ -1:1:1:1$

$\frac{mol}{L}$	HA	$+$	H_2O	$\rightleftharpoons$	H_3O^+	$+A^-$
$Start$	0.100				0	0
$+\Delta$	-0.0200				$+0.0200$	$+0.0200$
$=Eq$	0.080				0.0200	0.0200

$$pH=-\log_{10}[H_3O^+]$$
$$1.7=-\log_{10}[H_3O^+]$$
$$[H_3O^+]=0.01995=0.0200$$

$$K_a=\frac{[H_3O^+][A^-]}{[HA]}=\frac{(0.0200)^2}{0.080}=5.0\times10^{-3}$$

$[HA]=0.080\ M$

Find Mass NH_4Cl to produce $1.50L$ pH 4.75.

Note : Solutions of CO_2
$H_2CO_{3(aq)}+H_2O_{(l)}\rightleftharpoons HCO_3^-+H_3O^+$
$SO_{2(g)}\approx H_2\ SO_3$

$NH_4Cl\rightarrow NH_4^++Cl^-$ Cl^- is a spectator

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

$\frac{mol}{L}$	NH_4^+	$+$	H_2O	$\rightleftharpoons$	H_3O^+	$+$	NH_3
$Start$	x				0		0
$+\Delta$	-1.78×10^{-5}				$+1.78\times10^{-5}$		$+1.78\times10^{-5}$
$=Eq$	$(x-1.78\times10^{-5})$				1.78×10^{-5}		1.78×10^{-5}

$$pH=-\log_{10}[H_3O^+]$$
$$4.75=-\log_{10}[H_3O^+]$$
$$[H_3O^+]=1.78\times10^{-5}$$

$$K_a=\frac{[H_3O^+][NH_3]}{[NH_4^+]}$$ $x-1.78\times10^{-5}\approx x$

See Table

$$5.6\times10^{-10}=\frac{(1.78\times10^{-5})^2}{x}$$ Cross Multiply

$$x=0.56578$$ $[NH_4^+]=[NH_4Cl]=0.56578$ $mass\ NH_4Cl=0.565\ mol\times\frac{53.5\ g}{mol}=45g$

Find pH $0.10\ M\ NaCN$

$NaCN\rightarrow Na^++CH^-$ Na^+ is a spectator

$\frac{mol}{L}$	CN^-	$+$	H_2O	$\rightleftharpoons$	HCN	$+\ OH^-$
$Start$	0.100				0	0
$+\Delta$	$-x$				$+x$	$+x$
$=Eq$	$0.100-x$				x	x

$$K_b(CN^-)=\frac{K_w}{K_a(HCN)}=\frac{1.0\times10^{-14}}{4.9\times10^{-10}}=2.04\times10^{-5}$$

$$K_b=\frac{[HCN][OH^-]}{[CN^-]}$$ $0.100-x\approx0.100$

Multiply

$$2.04\times10^{-5}=\frac{x^2}{0.10}$$ Square Root Both Sides

$$x=1.43\times10^{-3}$$ $[OH^-]=1.43\times10^{-3}\ M$

$$pOH=-\log_{10}[OH^-]$$
$$pOH=-\log_{10}[1.43\times10^{-3}]$$
$$pOH=2.845$$

$$pH=14-pOH$$
$$pH=14-2.845$$
$$pH=11.15$$

C12 - 4.0 - Acid & Base

Find $K_b(A^-)$ $0.50\ M\ of\ a\ weak\ acid\ HA.$ $pOH = 10.64$

$pH = 14 - pOH$ $pH = -\log_{10}[H_3O^+]$
$pH = 14 - 10.64$ $3.36 = -\log_{10}[H_3O^+]$
$pH = 3.36$ $[H_3O^+] = 4.37 \times 10^{-4}$

$\frac{mol}{L}$	HA	$+$	H_2O	$\rightleftharpoons$	H_3O^+	$+$	A^-
$Start$	0.50				0		0
$+\Delta$	-4.37×10^{-4}				$+4.37 \times 10^{-4}$		$+4.37 \times 10^{-4}$
$= Eq$	$0.50 - 4.37 \times 10^{-4}$				4.37×10^{-4}		4.37×10^{-4}

$0.50 - 4.37 \times 10^{-4} \approx 0.50$

$$K_a(HA) = \frac{[HCN][OH^-]}{[CN^-]} = \frac{(4.37 \times 10^{-4})^2}{0.50} = 3.81 \times 10^{-7} \quad K_b(A^-) = \frac{K_w}{K_a(HA)} = \frac{1.0 \times 10^{-14}}{3.81 \times 10^{-7}} = 2.6 \times 10^{-8}$$

$aA + bB \rightarrow cC + dD$ $\frac{moles\ A}{moles\ B} = \frac{a}{b}$ Equivalence Point Parameters
-Concentration/Volume of Acid/Base/Mol Ratio

Find $[H_2SO_4]$ $H_2SO_4 + 2NaOH \rightarrow Na_2SO_4 + H_2O$

$$\boxed{mmol = millimol \quad 1\ M = \frac{1\ mol}{L} = \frac{1\ mmol}{mL}}$$

Equivalence Point Reached @ $23.10\ mL\ 0.2055\ M\ NaOH$ $25\ mL\ H_2SO_4$

$$\boxed{c = \frac{n}{V}}$$

$$moles\ NaOH = 0.2055\frac{mmol}{mL} \times 23.10\ mL = 4.7471\ mmol$$

$$[H_2SO_4] = 4.7471\ mmolNaOH \times \frac{1mol\ H_2SO_4}{2\ mol\ NaOH} = 2.3735\ mmol \times \frac{1}{25\ mL} = 0.09494\ M$$

Indicator - A weak organic acid or base with different colours for its conjugate acid and base forms.

$Hln + H_2O \rightleftharpoons ln^- + H_3O^+$ When an indicator is put into an acid, the excess of H_3O^+ shifts the Eq.
yellow red (Le'Chatelier's Principle)

$Hln + H_2O \rightarrow ln^- + H_3O^+$ An indicator is in its conjugate acid form when in highly acidic solutions
Causing the solution to turn yellow

Acidic solutions
$[Hln] > [ln^-]$ More yellow molecules than red

$Hln + H_2O \underset{\leftarrow}{\longrightarrow} ln^- + H_3O^+$ In basic solutions $[H_3O^+]$ is very low so that equilibrium shifts to te ln^- side and the solution surns red
red

Basic Solutions
$[ln^-] > [Hln]$ More red molecules than yellow An indicator is in its conjugate bass form when in highly basic solutions

Neutral Solutions
$[Hln] = [ln^-]$ Orange Solution
(End/Transition Point)

$$K_a = \frac{[ln^-][H_3O^+]}{[Hln]} = [H_3O^+] \qquad K_a = [H_3O^+]$$

The $[H_3O^+]$ at which an indicator changes colour equals the value of K_a for the indicator.

$$pK_a = -logK_a \qquad -logK_a = -\log[H_3O^+] \qquad pK_a = pH$$

An indicator is at midpoint of its colour change when the pH of the solution equals the K_a of the indicator.

Alizarin Yellow $10.1 - 12$ $If\ Aliz^-$ is red, find Alizarian Yellow in $1 \times 10^{-5}\ M\ NaOH.$

$HAliz + H_2O \rightleftharpoons Aliz^- + H_3O^+$
yellow red

$pH = 11 : [HAliz] = [Aliz^-]$
$pH < 11 : [HAliz] > [Aliz^-]$
$pH > 11 : [HAliz] < [Aliz^-]$

$$[OH^{-1}] = 1 \times 10^{-5}\ M \qquad K_a(H^+) = \frac{K_w}{K_b(OH^-)} = \frac{1 \times 10\text{^} - 14}{1 \times 10^{-5}} = 1 \times 10^{-9}$$

$$[H^+] = 1 \times 10^{-9}$$

$pH = -\log_{10}[H_3O^+]$
$pH = -\log_{10}[1 \times 10^{-9}]$
$pH = 9$

Alizarin Yellow is YELLOW

C12 - 4.0 - Titration of strong Acid with strong Base

Find pH mixing 50.0 mL 0.150 M $NaOH$ and 50.0 mL 0.200 M HCl.

$$(0.0500)(0.150) = \boxed{0.00750\, mol\, NaOH}$$

$$(0.0500)(0.200) = \boxed{0.0100\, mol\, HCl}$$

$$0.0100 - 0.00750 = \boxed{0.00250\, mol\, HCl\, xs}$$

$$[H_3O^+] = \frac{0.0025}{0.100} = \boxed{0.025\, M}$$

$$pH = -\log[0.025]$$

$$pH = \boxed{1.6}$$

Find pOH mixing 75.0 mL 0.200 M HBr and 225.0 mL 0.150 M KOH.

$$(0.0750)(0.200) = \boxed{0.0150\, mol\, HBr}$$

$$(0.22500)(0.150) = \boxed{0.03375\, mol\, HCl}$$

$$0.03375 - .0150 = \boxed{0.01875\, mol\, HCl\, xs}$$

$$[OH^-] = \frac{0.01875}{0.300} = \boxed{0.0625\, M}$$

$$pOH = -\log[0.0625]$$

$$\boxed{pOH = 1.204}$$

Find pH mixing 5.00 mL 0.100 M sulfuric acid and 12.0 mL 0.100 M $NaOH$.

$$(0.00500)(0.100) \times \frac{2H^+}{1H_2SO_4} = \boxed{0.00100\, mol\, H^+}$$

$$(0.0120)(0.100) = \boxed{0.00120\, mol\, OH^-}$$

$$0.00120 - 0.00100 = \boxed{0.00020\, mol\, OH^-\, xs}$$

$$pH = 14 + \log\left(\frac{0.00020}{0.0170}\right)$$

$$pH = \boxed{12.07}$$

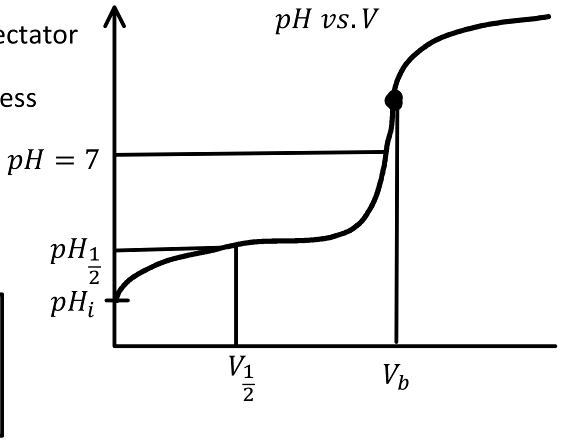
C12 - 4.0 - Titration of weak Acid with strong Base

$\boxed{i = initial}$

$CH_3COOH + NaOH \rightarrow Na^+ + CH_3COO^- + H_2O$ Na^+ is a spectator

$CH_3COOH + H_2O \rightleftharpoons CH_3COO^- + H_3O^+$ CH_3COOH is in excess

$All\ NaOH$ used up



25.0 mL Acetic Acid CH_3COOH of initial pH 2.2 is titrated with 28.8 mL $NaOH$ to equivalence point.
Therefor at 14.4 mL to point of titration with pH = 3.2.

Half way to the equivalence point (See Graph) half of initial acid $[CH_3COOH]$ is neutralized to its conjugate base $[CH_3COO^-]$.

$$\boxed{[CH_3COOH] = [CH_3COO^-]}$$

$$K_a = \frac{[CH_3COO^-][H_3O^+]_{\frac{1}{2}}}{[CH_3COOH]}$$

$$\boxed{K_a = [H_3O^+]_{\frac{1}{2}}} \quad \boxed{pK_a = pH_{\frac{1}{2}}}$$

let x = amount of CH_3COOH which dissociates

let $A = [CH_3COOH]_i$

$\frac{mol}{L}$	CH_3COOH	$+\ H_2O \rightleftharpoons$	CH_3COO^-	$+ H_3O^+$
Start	A		0	0
$+\Delta$	$-x$		$+x$	$+x$
$= Eq$	$A - x$		x	x

$$K_a = \frac{[CH_3COO^-][H_3O^+]}{[CH_3COOH]}$$

$$K_a = \frac{[H_3O^+]^2}{[CH_3COOH]}$$

At Equivalence Point
$x = [CH_3COO^-] = [H_3O^+]$

$$K_a = \frac{x^2}{A - x}$$

$$\boxed{K_a = \frac{[OH^-]^2}{[Acid]_{eq}}}$$

$$A - x = \frac{x^2}{K_a}$$

$$A - x = \frac{(6.31 \times 10^{-3})^2}{6.31 \times 10^{-4}}$$

$$A - x = 0.0631$$
$$A = 0.0631 + x$$
$$A = 0.0631 + 6.31 \times 10^{-3}$$
$$A = 0.0694$$

$$\boxed{pK_a = pH_{\frac{1}{2}} = 3.2}$$

$$pH = -\log_{10}[H_3O^+]$$
$$3.2 = -\log_{10}[H_3O^+]$$
$$\boxed{[H_3O^+] = 6.31 \times 10^{-4}}$$
$$\boxed{K_a = [H_3O^+] = 6.31 \times 10^{-4}}$$

$$pH_i = -\log_{10}[H_3O^+]$$
$$2.2 = -\log_{10}[H_3O^+]$$
$$\boxed{[H_3O^+] = 6.31 \times 10^{-3}\ M}$$

$$\boxed{[CH_3COOH] = 0.0694\ M}$$

$$moles\ CH_3COOH = 0.0694 \frac{mmol}{mL} \times 25\ mL = \boxed{1.74\ mmol = mol\ NaOH}$$

$$[NaOH] = \frac{1.74\ mmol}{28.8\ mL} = \boxed{0.0602\ M}$$

$pH > 7$ at Equivalence Point therefor $\boxed{\text{Basic}}$ Use a test with $\boxed{high\ pH}$

Titration or 25 mL propanoic acid C_2H_5COOH $(Initial\ pH = 2.95)$ $with\ 23.8\ mL\ 0.100\ M\ NaOH$ to equivalence point and pH = 4.87 at 11.9 mL at point of titration.

$C_2H_5COOH + NaOH \rightarrow C_2H_5COO^- + Na^+$

$C_2H_5COOH + H_2O \rightarrow C_2H_5COO^- + H_3O^+$

$$\boxed{pK_b = p[OH]_{\frac{1}{2}} = 1.35 \times 10^{-5}}$$

let x = amount of C_2H_5COOH which dissociates
let $B = [C_2H_5COOH]_i$

$\frac{mol}{L}$	C_2H_5COOH	$+\ H_2O \rightleftharpoons$	$C_2H_5COO^-$	$+ H_3O^+$
Start	B		0	0
$+\Delta$	$-x$		$+x$	$+x$
$= Eq$	$B - x$		x	x

$$pOH = -\log_{10}[OH^-]$$
$$4.87 = -\log_{10}[OH^-]$$
$$\boxed{[OH^-] = 1.35 \times 10^{-5}\ M}$$

$$pOH = -\log_{10}[OH^-]$$
$$2.95 = -\log_{10}[OH^-]$$
$$\boxed{[OH^-] = 1.12 \times 10^{-3}\ M}$$

$$K_b = \frac{x^2}{(B - x)}$$

$$B - x = \frac{[1.12 \times 10^{-3}]^2}{1.35 \times 10^{-5}}$$

$$B - x = 0.0933$$
$$B = 0.0933 + x$$
$$B = 0.0933 + 0.00112$$
$$B = 0.0944$$

$$\boxed{[C_6H_5NH_2] = 0.0944\ M}$$

C12 - 4.0 - Titration of weak Base with strong Acid

$\boxed{i = initial}$

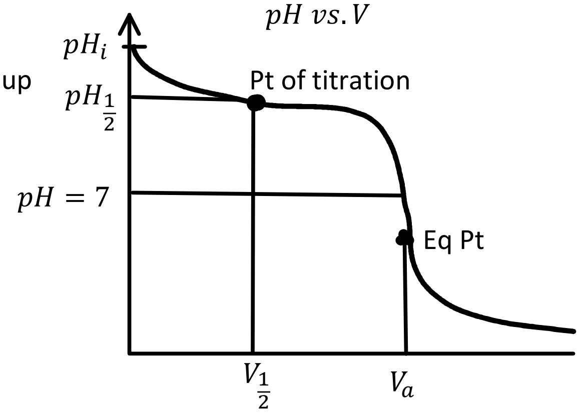
$NH_3 + HCl \rightarrow NH_4^+ + Cl^-$ Cl^- is a spectator

$NH_3 + H_2O \rightleftharpoons NH_4^+ + OH^-$ $\boxed{NH_3 \text{ is in excess}}$ All HCl is used up

25.0 $mL\ NH_3$ (a weak base) of initial pH 11.5 is titrated with 19.20 $mL\ HCl$ to equivalence point.
Therefor at 9.60 mL to point of titration with pH 10.8.

Half way to the equivalence point (See Graph) half of initial base $[NH_3]$ is neutralized to its conjugate acid $[NH_4^+]$.

$\boxed{[NH_3] = [NH_4^+]}$



$$K_b = \frac{[NH_4^+][OH^-]_{\frac{1}{2}}}{[NH_3]}$$

$\boxed{K_b = [OH^-]_{\frac{1}{2}}}$ $\boxed{pK_b = pOH_{\frac{1}{2}}}$

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

$$K_b = \frac{[OH^-]^2}{[NH_3]}$$

At Equivalence Point
$x = [NH_4^+] = [OH^-]$

$$K_b = \frac{x^2}{B - x}$$

$$\boxed{K_b = \frac{[OH^-]^2}{[Base]_{eq}}}$$

$$B - x = \frac{x^2}{K_b}$$

$$B - x = \frac{(3.2 \times 10^{-3})^2}{6.3 \times 10^{-4}}$$

$B - x = 0.0158\ M$
$B = 0.0158 + x$
$B = 0.0158 + 6.3 \times 10^{-4}$
B = 0.0165 M

$[NH_3] = 0.0165\ M$

$let\ x = amount\ of\ NH_3\ which\ dissociates$
$let\ B = [NH_3]_i$

$\frac{mol}{L}$	NH_3	$+\ H_2O \rightleftharpoons$	NH_4^+	$+\ OH^-$
$Start$	B		0	0
$+\Delta$	$-x$		$+x$	$+x$
$= Eq$	$B - x$		x	x

$pK_b = p[OH^-]_{\frac{1}{2}} = 3.2$

$pOH = 14 - pH$
$pOH = 14 - 10.8$
$pOH = 3.2$

$pOH = -\log_{10}[OH^-]$
$3.2 = -\log_{10}[OH^-]$
$[OH^-] = 6.3 \times 10^{-4}\ M$

$K_b = [OH^-] = 6.3 \times 10^{-4}$

$pOH = 14 - pH$
$pOH = 14 - 11.5$
$pOH = 2.5$

$pH = -\log_{10}[OH^-]$
$2.5 = -\log_{10}[OH^-]$
$[OH^-] = 3.2 \times 10^{-3}$

$$moles\ NH_3 = 0.0165 \frac{mmol}{mL} \times 25\ mL = 0.413\ mmol = mol\ HCl$$

$$[HCl] = \frac{0.413\ mmol}{19.2 -\ mL} = 0.0215\ M$$

$pH < 7$ at Equivalence Point therefor Acidic Use a test with low pH

Titration or organic base aniline $C_6H_5NH_2$ $(Initial\ pH = 8.72)\ with$ 16.3 mL 0.100 $M\ HCl$ to equivalence point and pH = 4.63 at 8.16 mL at point of titration.

$C_6H_5NH_2 + HCl \rightarrow C_6H_5NH_3^+ + Cl^-$

$C_6H_5NH_2 + H_2O \rightarrow C_6H_5NH_3^+ + OH^-$

$let\ x = amount\ of\ C_6H_5NH_2\ which\ dissociates$
$let\ B = [C_2H_5NH_2]_i$

$\frac{mol}{L}$	$C_6H_5NH_2$	$+\ H_2O \rightleftharpoons$	$C_6H_5NH_3^+$	$+\ OH^-$
$Start$	B		0	0
$+\Delta$	$-x$		$+x$	$+x$
$= Eq$	$B - x$		x	x

$pK_b = p[OH]_{\frac{1}{2}} = 4.26 \times 10^{-10}$

$pOH = 14 - 4.63 = 9.37$

$pOH = -\log_{10}[OH^-]$
$9.37 = -\log_{10}[OH^-]$
$[OH^-] = 4.26 \times 10^{-10}\ M$

$pOH = 14 - 8.72 = 5.28$

$pOH = -\log_{10}[OH^-]$
$5.28 = -\log_{10}[OH^-]$
$[OH^-] = 5.25 \times 10^{-6}\ M$

$$K_b = \frac{x^2}{(B - x)}$$

$$B - x = \frac{[5.25 \times 10^{-6}]^2}{4.26 \times 10^{-10}}$$

$B - x = 0.0647$
$B = 0.0647 + x$
$B = 0.0647$

$\boxed{x\ is\ neglible}$

$[C_6H_5NH_2] = 0.0647\ M$

C12 - 4.0 - Buffers

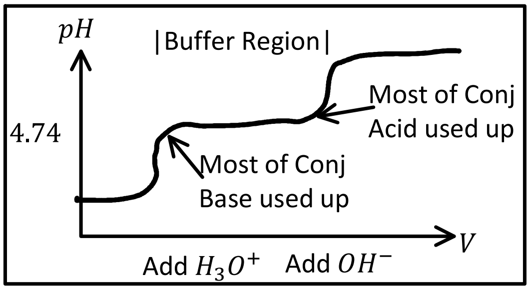
Buffer - A solution containing appreciable amounts of a weak acid and it's conjugate weak base.

$$\underset{1.0\,M}{CH_3COOH} + H_2O \rightleftharpoons \underset{1.0\,M}{CH_3COO^-} + H_3O^+$$

$$K_a = \frac{[CH_3COO^-][H_3O^+]}{[CH_3COOH]} \qquad [CH_3COOH] = [CH_3COO^-]$$

$$K_a = [H_3O^+]$$

$$K_a = 1.8 \times 10^{-5} \qquad \boxed{pH = pK_a}$$



When equal concentrations of a weak acid and it's conjugate base are added to water, the resulting buffer will equal the pK_a value of the weak base.

$$CH_3COOH + H_2O \rightleftharpoons CH_3COO^- + H_3O^+ \qquad K_a = 1.8 \times 10^{-5}$$

$$[CH_3COOH] = [CH_3COO^+] = 1.0\,M$$

$$K_a = \frac{[CH_3COO^-][H_3O^+]}{[CH_3COOH]}$$

$$K_a = \frac{[1.0][H_3O^+]}{[1.0]} \qquad pH = -\log_{10}[H_3O^+]$$

$$pH = -\log_{10}[1.8 \times 10^5]$$

$$K_a = [H_3O^+] = 1.8 \times 10^5 \qquad pH = 4.74$$

After diluting tenfold

$$[CH_3COOH] = [CH_3COO^+] = 0.10\,M$$

$$K_a = \frac{[CH_3COO^-][H_3O^+]}{[CH_3COOH]}$$

$$K_a = \frac{[0.10][H_3O^+]}{[0.10]}$$

$$K_a = [H_3O^+]$$

Diluting a Buffer has no effect on pH

$1\ mol\ CH_3COOH$
$1\ mol\ CH_3COO^-$
$1\ L$

If $0.10\ mol\ H_3O^+$ added

Le'Chatalier

Equilibrium shifts to
-use up $0.1\ mol\ CH_3COO^-$
-produce $0.1\ mol\ CH_3COOH$

$$\underset{\text{Acid}}{CH_3COOH} + H_2O \rightarrow \underset{\text{Base}}{CH_3COO^-} + H_3O^+$$

$\leftarrow$

Original Ratio	Ratio after adding H_3O^+
$1\ mol\ Acid$	$1.1\ mol\ Acid$
$1\ mol\ Base$	$0.9\ mol\ Base$

$$K_a = \frac{[CH_3COO^-][H_3O^+]}{[CH_3COOH]}$$

$$1.8 \times 10^{-5} = \frac{[0.9][H_3O^+]}{[1.1]}$$

$$[H_30^+] = 2.2 \times 10^{-5}$$

$$ph = pK_a$$

$$pH = -\log_{10}[H_3O^+]$$

$$pH = -\log_{10}[2.2 \times 10^{-5}]$$

$$pH = 4.66$$

If $0.10\ mol\ OH^-$ added

Le'Chatalier

Equilibrium shifts to
-use up $0.1\ mol\ CH_3COH$
-produce $0.1\ mol\ CH_3COO^-$

$$\underset{\text{Acid}}{CH_3COOH} + OH^- \rightleftarrows \underset{\text{Base}}{CH_3COO^-} + H_3O^+$$

Original Ratio	Ratio after adding H_3O^+
$1\ mol\ Acid$	$0.9\ mol\ Acid$
$1\ mol\ Base$	$1.1\ mol\ Base$

$$K_a = \frac{[CH_3COO^-][H_3O^+]}{[CH_3COOH]}$$

$$1.8 \times 10^{-5} = \frac{[1.1][H_3O^+]}{[0.9]}$$

$$[H_30^+] = 1.47 \times 10^{-5}$$

$$ph = pK_a$$

$$pH = -\log_{10}[H_3O^+]$$

$$pH = -\log_{10}[1.47 \times 10^{-5}]$$

$$pH = 4.83$$