Calculate the pH of the solution formed by mixing 20.0 mL of 0.500 M HCl with 30.0 mL of 0.300 M NaOH.

$$[HCl] = \frac{(20mL)(.5M)}{50mL} = 0.2M$$
$$\parallel$$
$$[H_3O^+]$$

$$[NaOH] = \frac{(30mL)(.3M)}{50mL} = 0.18M$$
$$\parallel$$
$$[OH^-]$$

extra $[H_3O^+] = 0.02M$

$pH = -\log[H_3O^+]$

$$\boxed{pH = 1.699}$$

NaCN

dissociation

$NaCN \longrightarrow \cancel{Na^+} + CN^-$

hydrolysis

$CN^- + H_2O \rightleftharpoons HCN + OH^-$

Kb expression

$$K_b = \frac{[HCN][OH^-]}{[CN^-]}$$

$$K_b = \frac{K_w}{K_a}$$

# Acid-Base Indicators

- indicators are used to determine the pH of a solution

| |
|---|
| **indicator** = a weak organic acid or base that has different colours for their conjugate acid and base forms |

- indicators are often indicated by the symbol HIn d (acid form) and $In^-$ (base form)
- since an indicator is a weak acid or base, the following equilibrium can be written:

$$\underset{\text{yellow}}{HIn} + H_2O \rightleftarrows \underset{\text{red}}{In^-} + H_3O^+$$

- when an indicator is placed into an acid solution, the excess $H_3O^+$ causes a shift in the indicator's equilibrium

$$\underset{\text{yellow}}{HIn} + H_2O \rightleftarrows \underset{\text{red}}{In^-} + H_3O^+$$

| |
|---|
| An indicator will be in its **conjugate acid (HIn)** form in **highly acidic solutions**. |

- when an indicator is placed into a base solution, the $OH^-$ reacts to decrease the $[H_3O^+]$ and cause a shift in the indicator's equilibrium

$$\underset{\text{yellow}}{HIn} + H_2O \rightleftarrows \underset{\text{red}}{In^-} + H_3O^+$$

| |
|---|
| An indicator will be in its **conjugate base ($In^-$)** form in **highly basic solutions**. |

- the colour of an indicator depends on the relative concentrations of the conjugate acid and base forms of the indicator

| |
|---|
| **transition point** = indicator is mid-way through its colour change and $[HIn] = [In^-]$ |

- consider the following indicator equilibrium:

$$HIn + H_2O \rightleftarrows In^- + H_3O^+$$

$$Kin = \frac{[H_3O^+][In^-]}{[HIn]}$$

- at the transition point $[HIn] = [In^-]$, so

$$Kin = \frac{[H_3O^+][\cancel{In^-}]}{[\cancel{HIn}]} = [H_3O^+]$$

- in addition,

$$-\log Kin = -\log[H_3O^+]$$

$$pKin = pH$$

| |
|---|
| At the transition point of any indicator, the following relationships exist: |
| $[HIn] = [In^-]$ $\quad Kin = [H_3O^+]$ $\quad pKin = pH$ |

- the Ka of an indicator (Kin) can be calculated using the pH range over which the indicator changes colour

Using indicator table to calculate the Kin.

ex. What is the Kin for phenolpthalein?

phenolpthalein

8.2-10.0

$\frac{8.2+10}{2} = 9.1$

transition point

$HIn + H_2O \rightleftharpoons In^- + H_3O^+$

$Kin = \frac{[\cancel{In^-}][H_3O^+]}{[\cancel{HIn}]}$

$Kin = [H_3O^+]$

$Kin = [H_3O^+] = \text{antilog} -9.1$

$= 8 \times 10^{-10}$

Using indicators to determine the pH of a solution.

ex. What is the approximate pH range of a solution that changes methyl red → yellow and neutral red → red?

methyl red

4.8 - 6.0

red yellow

neutral red

6.8 - 8.0

red amber

$6.0 < \text{solution} < 6.8$

**Thymol Blue**

➪ appears twice on table

• diprotic acid ⇒ colour change each time loses H+ (proton)

thymol blue (Tb) = weak acid ($H_2Tb$)

1st ionization: $H_2Tb + H_2O \rightleftharpoons H_3O^+ + HTb^-$

**(red)** **(yellow)**

2nd ionization: $HTb^- + H_2O \rightleftharpoons H_3O^+ + Tb^{2-}$

**(yellow)** **(blue)**

| | | |
|---|---|---|
| Thymol Blue: | 1.2 - 2.8 | **red - yellow** |
| | 8.0 - 9.6 | **yellow - blue** |

<u>Universal Indicators</u>

➪ mix of indicators to get several colour changes

## Predicting the colours of indicators at various pH.

ex. A mixture of the indicators methyl orange, phenol red and thymol blue is added to a pH 5.0 and pH 8.0 solution. What is the colour of the mixture at pH 5.0 and 8.0?