

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.

$$\text{pH} = 1.699$$

$$[\text{HCl}] = \frac{(20\text{mL})(0.5\text{M})}{50\text{mL}} = 0.2\text{M}$$

$$[\text{NaOH}] = \frac{(30\text{mL})(0.3\text{M})}{50\text{mL}} = 0.18\text{M}$$

$$\text{extra } [\text{H}_3\text{O}^+] = 0.02\text{M}$$

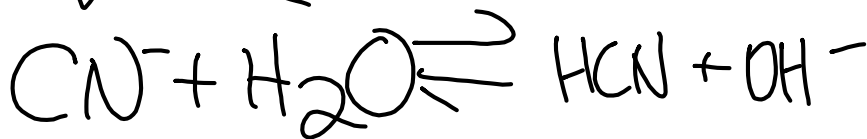
$$\text{pH} = -\log[\text{H}_3\text{O}^+]$$

NaCN

dissociation



hydrolysis



$K_b$  expression

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

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

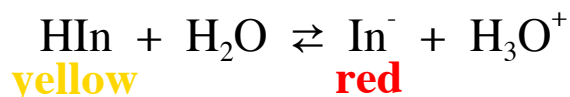
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## Acid-Base Indicators

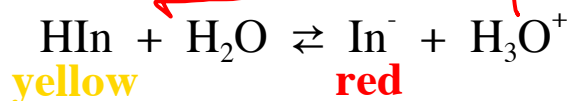
- indicators are used to determine the pH of a solution

**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  $\text{HIn}$  (acid form) and  $\text{In}^-$  (base form)
- since an indicator is a weak acid or base, the following equilibrium can be written:

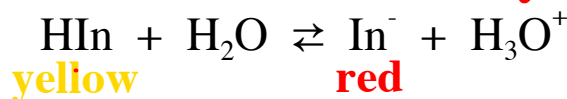


- when an indicator is placed into an acid solution, the excess  $\text{H}_3\text{O}^+$  causes a shift in the indicator's equilibrium



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

- when an indicator is placed into a base solution, the  $\text{OH}^-$  reacts to decrease the  $[\text{H}_3\text{O}^+]$  and cause a shift in the indicator's equilibrium

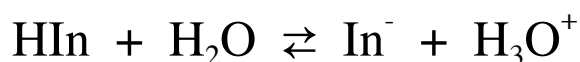


An indicator will be in its **conjugate base (In<sup>-</sup>)** form in **highly basic solutions**.

- 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:



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

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

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

- in addition,

$$-\log K_{in} = -\log[H_3O^+]$$

$$pK_{in} = pH$$

At the transition point of any indicator, the following relationships exist:

$$[HIn] = [In^-] \quad K_{in} = [H_3O^+] \quad pK_{in} = pH$$

- the  $K_a$  of an indicator ( $K_{in}$ ) can be calculated using the pH range over which the indicator changes colour

Using indicator table to calculate the  $K_{in}$ .

ex. What is the  $K_{in}$  for phenolphthalein?

phenolphthalein  
8.2 - 10.0

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

Transition Point



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

$$K_{in} = [H_3O^+]$$

$$K_{in} = [H_3O^+] = \text{antilog } -9.1 \\ = 8 \times 10^{-10}$$

Using indicators to determine the pH of a solution.

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

methyl red  
4.8 - 6.0  
red yellow

neutral red  
6.8 - 8.0  
red amber

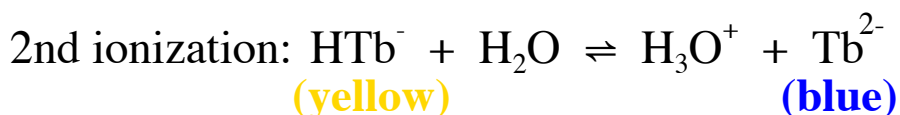
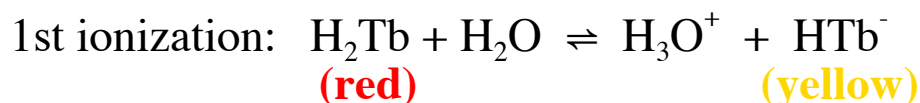
$$6.0 < \text{solution} < 6.8$$

## Thymol Blue

⇨ appears twice on table

- diprotic acid ⇒ colour change each time loses H<sup>+</sup> (proton)

thymol blue (Tb) = weak acid (H<sub>2</sub>Tb)



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

## Universal Indicators

⇨ 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?