V) Indicators

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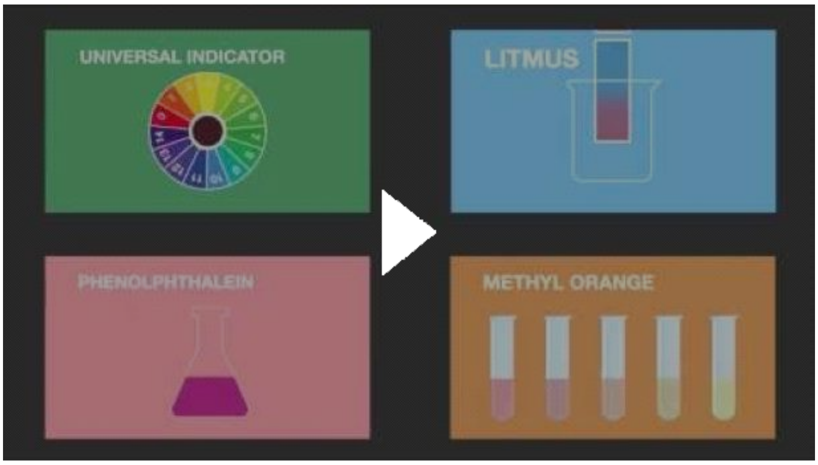
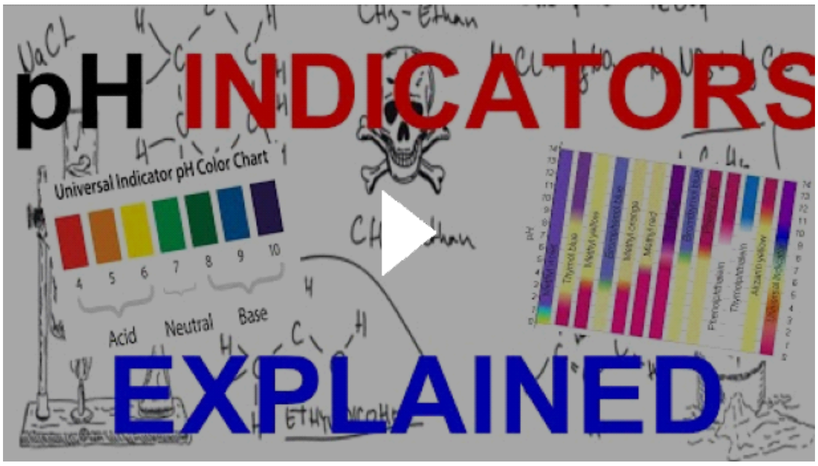
[What are Indicators and how do we use them? | The Chemistry Journey | The Fuse School]



[pH Indicators Explained]



In = Indicator

will be this colour in acid

this colour in base

$HIn_{(aq)} + H_2O_{(l)} \rightleftharpoons In^-_{(aq)} + H_3O^+_{(aq)}$

weak acid

weak base

V) Indicators

Indicators are used to signal the equivalence point (when moles H_3O^+ = moles OH^-) of an acid-base titration using a colour change. An indicator is a solution of a weak organic acid (an acid that contains carbon acid), **HIn**, and its conjugate base, **In**$^-$, at equilibrium. The acid form of the indicator, HIn, is a different colour than the conjugate base form, In$^-$. The following is the general equilibrium for any acid-base indicator:

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

HIn and In$^-$ are different colours.

Let's look at how an indicator equilibrium works in solution using the indicator bromthymol blue:

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

indicator bromthymol blue:

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

For bromthymol blue: yellow (in acid $H_3O^+ > OH^-$) blue (in base $H_3O^+ < OH^-$)

If [HIn] (acid) > [In-] (base), the system favours the left (reactants) side and the solution will be a yellow colour. If [In-] (base) > [HIn], the system favours the right (products) side and the solution will be a blue colour.

What happens to the equilibrium if bromthymol blue is put into an acidic => $H_3O^+ > OH^-$ solution, and what is the resulting solution colour?

$\uparrow [H_3O^+]$ more

① in an acidic sol'n the indicator will always be in "acid form"

$$HIn + H_2O \Leftrightarrow In^- + [H_3O^+] \uparrow$$

yellow green blue

increased in an acidic solution

② $\uparrow [H_3O^+]$ causes an equilibrium shift to the LEFT

③ $[In^-]$ decreases while $[HIn]$ increases

④ If $[HIn] > [In^-]$ the colour has changed to yellow

What will happen if bromthymol blue is put into a basic solution, and what is the resulting solution colour?

$$HIn + H_2O \Leftrightarrow In^- + [H_3O^+] \downarrow$$

yellow blue

OH^- ① → $[OH^-] > [H_3O^+]$

decrease $\uparrow OH^-$: $\downarrow H_3O^+$

② $\downarrow [H_3O^+]$ then equilibrium will shift RIGHT

• FWD rxn rate is faster as RVS rxn rate has decreased

③ $[HIn]$ will decrease : $[HIn] < [In^-]$ ∴ solution will be BLUE

During a titration, pH is constantly changing as base is being added to acid (or *visa versa*). If an indicator such as bromthymol blue is present, it will eventually undergo a colour change due to the continual change in $[H_3O^+]$ and resulting shift of the indicator equilibrium. yellow 6.0 – 7.6 blue

If there is acid in a flask with some bromthymol blue, what colour will it be?

[HIn] > [In⁻] ∴ yellow

If base is continually added from the buret, what shift results in the equilibrium?

↑[OH⁻] ... and $[H_3O^+]$↓ equilibrium shifts RIGHT (solution yellow → green)

How does this affect [HIn] and $[In^-]$?

[HIn] < [In⁻]
yellow blue

What result will this have on the colour of the solution?

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

yellow blue

- at first [HIn] > [In⁻]; acidic solution; yellow
- as OH⁻ is added, [In⁻] will increase (shift right), [HIn] to decrease
- eventually [HIn] = [In⁻] at the equivalence point where $H_3O^+ = OH^-$
- this is the intermediate colour for the indicator = green
- if more [OH⁻] is added (shift right); [In⁻] > [HIn] ∴ blue [OH⁻] > $[H_3O^+]$

The point at which the colour is an equal mixture of the [HIn] colour and the $[In^-]$ colour is called the **transition point** for the indicator. Another name is the **endpoint**, as this is when a titration would come to an end as the endpoint signals that the equivalence point has been reached. [HIn] = [In⁻] endpoint.

*

The endpoint occurs at different pHs for different indicators, as each indicator has its own unique equilibrium. The acid-base indicator table in the data booklet shows different indicators and the pH range of their colour changes.

selecting the correct indicator to show the equivalence point of a titration is important.

WANT: $\underbrace{\text{transition/endpoint}}_{\text{indicator}} \cong \underbrace{\text{equivalence point}}_{\text{moles } OH^- = H_3O^+ \text{ titration}}$

Most indicators change colour over a range of about 2 pH units. For example, bromthymol blue is yellow at pH 6.0 and below and blue at pH 7.6 and above. From 6.0 to 6.8, it's yellow-green, at 6.8 it's perfect green, and from 6.8-7.6 it's blue-green.

It is very important to be able to distinguish between the two terms **equivalence point** and **endpoint**.

- The equivalence point is the point in the titration where moles of H_3O^+ = moles of OH^-.
- The endpoint is the point in the titration where the colour of the indicator changes. If the indicator is chosen correctly, it will change the colour of the solution at or very near the equivalence point.

Practice Questions:

1. Which of the following indicators is red at pH 13?
   A. Orange IV
   B. Alizarin Yellow
   C. Indigo Carmine
   D. Thymol Blue
2. What colour is a 1.0×10^{-3} M NaOH solution containing the indicator Neutral Red? red 6.8 – 8.0 amber

$[OH^-] = 1.0 \times 10^{-3}$ M

$\therefore pOH = -\log(1.0 \times 10^{-3}) = 3 \quad \therefore pH = 11 \quad \boxed{\therefore \text{ amber}}$

Recall that the general equilibrium equation for an indicator is as follows:

$$HIn_{(aq)} + H_2O_{(l)} \Leftrightarrow In^-_{(aq)} + H_3O^+_{(aq)}$$

Write the K_a equation for the above:

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

indicator [HIn]=[In⁻] will cancel out.

At the endpoint, what is true about [HIn] and $[In^-]$? Therefore, what will the K_a reduce to?

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

So, **at the endpoint** (point of colour change), the $[H_3O^+]$ equals the value of the K_a for the indicator. indicator is a weak organic acid.

and: $-\log K_a = -\log[H_3O^+]$

$pK_a = pH$

It is easy to find the K_a of each indicator (remember, indicators are weak organic acids) using the indicator data table and some simple calculations.

Finding Ka of an Indicator:

i) Find the pH of the endpoint of the indicator using the table.

ii) Use the endpoint pH to find the $[H_3O^+]$ at this point (2^{nd} log(-pH))

iii) At the endpoint, the $[H_3O^+]$ is equal to the K_a

Example: Find the K_a of Orange IV. range 1.4 – 2.8 (red, yellow)

i) $2.8 - 1.4 = \frac{1.4}{2} = 0.7 + 1.4 = 2.1$ middle of range = endpoint. 1sf

$\boxed{pH = 2.1 \text{ @ endpoint}}$

ii) $[H_3O^+] = \text{invlog}^{-pH}$

$[H_3O^+] = \text{invlog}(-2.1)$

$10^{\wedge}(-2.1)$ 1sf.

$[H_3O^+] = 7.943 \times 10^{-3}$

$\boxed{\therefore 8 \times 10^{-3} M}$

iii) Ka? @ the endpoint $[H_3O^+] = K_a$

$\boxed{\therefore K_a = 8 \times 10^{-3}}$

Assignment 6: Indicator Exercises

1. Which of the following chemical indicators has a $K_a = 2.5 \times 10^{-5}$?

   A. methyl orange B. phenolphthalein

   C. thymolphthalein D. bromcresol green

2. Find the K_a of Alizarin Yellow.

3. A weak acid is titrated with a strong base using the indicator phenolphthalein to detect the endpoint. What is the approximate pH at the transition point?

   A. 7.0 B. 8.0 C. 9.0 D. 10.0