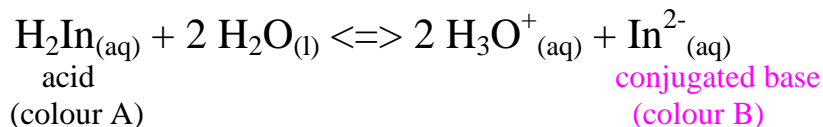


Acid - Base Indicators Phenolphthalein, Methyl Orange, Methyl Red

Acid - Base indicators (also known as pH indicators) are substances which change colour with pH. They are divalent weak acids or bases, involved in protolytic water acid-base equilibrium and form two type coloured ions as Brensted-Lowry protolytic pairs.

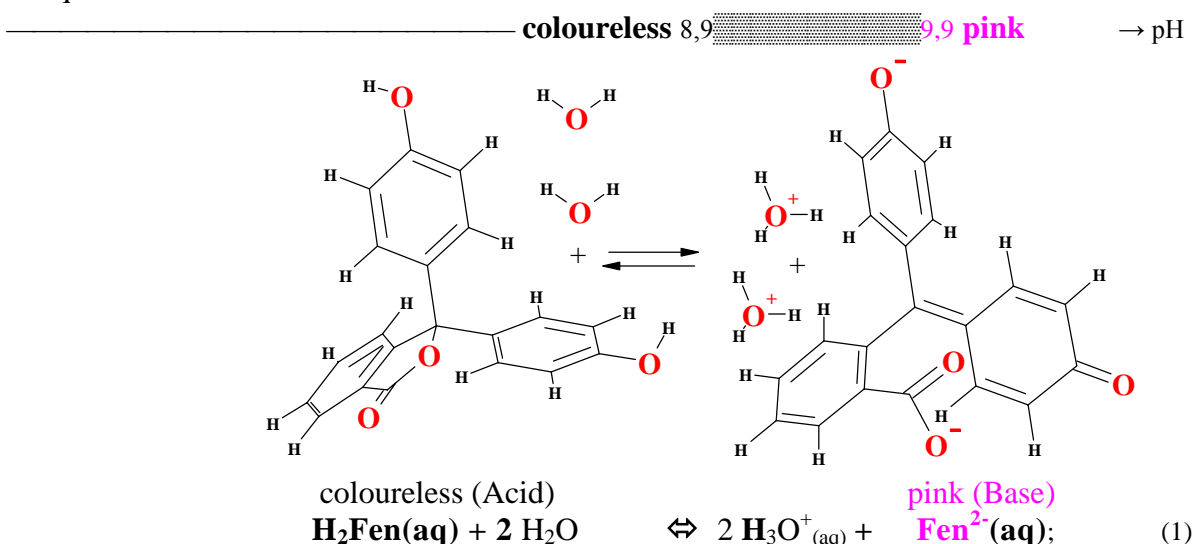
Consider an indicator which is divalent weak acid, with the formula H_2In . At equilibrium, the following equilibrium equation is established with conjugate base In^{2-} of different colour:



The acid and its conjugate base have different colours. At low pH values the concentration of H_3O^+ is high and so the equilibrium position lies to the left. The equilibrium solution has the colour A.

At high pH values, the concentration of H_3O^+ is low - the equilibrium position thus lies to the right and the equilibrium solution has colour B.

Phenolphthalein is an example of an indicator which establishes this type of equilibrium in aqueous solution:



Phenolphthalein (pdb version) is a coloureless, weak acid which dissociates in water forming pink anions. Under acidic conditions, the equilibrium is to the left, and the concentration of the anions is too low for the pink colour to be observed. However, under alkaline conditions, the equilibrium is to the right, and the concentration of the anion becomes sufficient for the pink colour to be observed.

We can apply equilibrium expression to indicator equilibrium - in general for divalent weak acid indicator: $K_{Fen} = ([H_3O^+]^2 \cdot [Fen^{2-}]) / [H_2Fen]$ equilibrium constant. (2)

K_{Fen} is known as the indicator dissociation constant. The colour of the indicator turns from colour A to colour B or vice versa at its turning point. At this point equal stay:

$$[H_2Fen] = [Fen^{2-}]$$

So from equilibrium expression (1,2) is derivate simple equation : $K_{Fen} = [H_3O^+]^2 = 10^{-18.8}$;

$$pK_{Fen} = -\log(10^{-18.8}) = -(-18.8) = 18.8; \text{pH} = -\log(\sqrt{[H_3O^+]}) = \frac{1}{2} \log(10^{-18.8}) = \frac{1}{2} 18.8 = 9.4. \quad (3)$$

The pH of the solution at its turning point is equal to the pK_{Fen} and is the pH at which half of the indicator is in its acid form and the other half in the form of its conjugate base.

Indicator Range

At a low pH, a weak acid indicator is almost entirely in the $[\text{H}_2\text{Fen}]$ form of acid, the colour of which predominates. As the pH increases - the influence of the colourless of $[\text{H}_2\text{Fen}]$ medium accenting replaces the colour of $[\text{Fen}^{2-}]$ and the equilibrium is pushed to the right. Therefore the intensity of the colour of $[\text{Fen}^{2-}]$ increases. Physiological visible condition for certainly distinct colour is assumed at ratio 10 for conjugated base form:

$$[\text{Fen}^{2-}] / [\text{H}_2\text{Fen}] = 10 ; \text{ and for acid coloured form } [\text{Fen}^{2-}] / [\text{H}_2\text{Fen}] = 1 / 10 = 0.1 \quad (4)$$

$$+K_{\text{Fen}} = [\text{H}_3\text{O}^+]^2 \cdot [\text{Fen}^{2-}] / [\text{H}_2\text{Fen}] = 10^{-18.8} \cdot 10 = 10^{-19.8} ; +\text{pH}_{\text{Fen}} = 19.8 / 2 = 9.9 ;$$

$$-K_{\text{Fen}} = [\text{H}_3\text{O}^+]^2 \cdot [\text{Fen}^{2-}] / [\text{H}_2\text{Fen}] = 10^{-18.8} \cdot 0.1 = 10^{-17.8} ; -\text{pH}_{\text{Fen}} = 17.8 / 2 = 8.9 ;$$

$$\text{pH} = -\log(\sqrt{[\text{H}_3\text{O}^+]^2}) \pm 0.5 = \frac{1}{2} \log(10^{-18.8}) \pm 0.5 = \frac{1}{2} 18.8 \pm 0.5 = 9.4 \pm 0.5 \quad , \quad (5)$$

which is colour change interval for divalent weak acid-base indicator ± 0.5 pH units.

An indicator is most effective if the colour change is distinct and over a low pH range. For divalent acid-base indicators the range is within ± 0.5 of the pK_{In} value: Please see the table below for examples, to the left is a model of the acid form of each indicator - with the colour of the solution at the turning point. Number 2 of involved protons H^+ in equilibrium or 1 Thymol Blue 2nd .

Table pH indicators

Indicator	Colour		pH colour turning point		pH range $\text{pK}_{\text{In}} \text{ or } \text{pK}_{\text{In}}/2$ $-\text{pH} < \text{pK}_{\text{In}} < +\text{pH}$
	Acid	Base	point		
Thymol Blue - 1st change	red	yellow	1.7		1.2 - 2.2
Methyl Orange	red	yellow	3.7		3.2 - 4.2
Bromocresol Green	yellow	blue	4.7		4.2 - 5.2
Methyl Red	yellow	red	5.1		4.6 - 5.6
Bromothymol Blue	yellow	blue	7.0		6.5 - 7.5
Phenol Red	yellow	red	7.9		7.4 - 8.4
Thymol Blue - 2nd change	yellow	blue	8.9		7.9 - 9.9
Phenolphthalein	colourless	pink	9.4		8.9 - 9.9

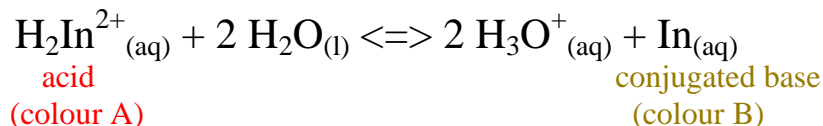
A Universal Indicator is a mixture of indicators which give a gradual change in colour over a wide pH range - the pH of a solution can be approximately identified when a few drops of universal indicator are mixed with the solution.

Indicators are used in titration solutions to signal the completion of the acid-base reaction.

Acid - Base Indicators Methyl Orange

Acid - Base indicators (also known as pH indicators) are substances which change colour with pH. They are divalent weak acids or bases, involved in protolytic water acid-base equilibrium and form two type coloured ions as Brensted-Lowry protolytic pairs.

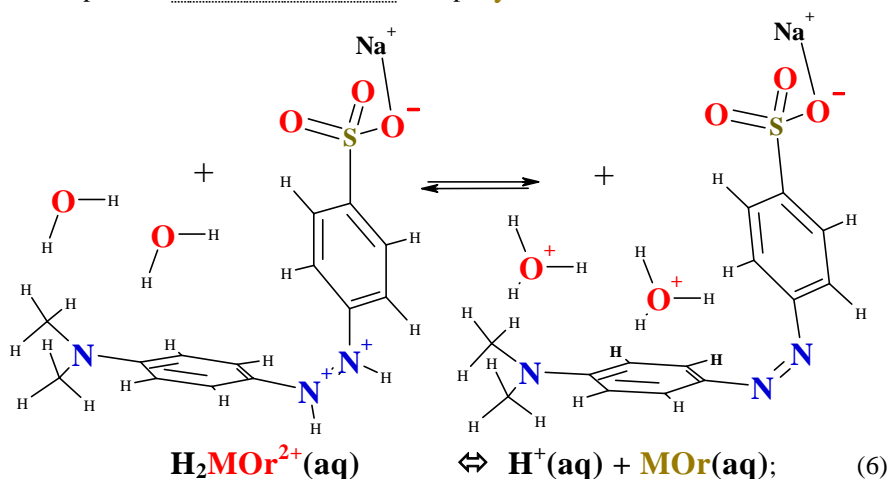
Consider an indicator which is divalent weak acid, with the formula H_2In^{2+} . At equilibrium, the following equilibrium equation is established with its conjugate base In^- :



The acid and its conjugate base have different colours. At low pH values the concentration of H_3O^+ is high and so the equilibrium position lies to the left.

The equilibrium solution has the colour A. At high pH values, the concentration of H_3O^+ is low - the equilibrium position thus lies to the right and the equilibrium solution has colour B.

Methyl Orange is an example of an indicator which establishes this type of equilibrium in aqueous solution:



$$pH = pK_{MOR} / 2 \pm 0,5 = 3,7 \pm 0,5;$$

Protonate nitrogen atoms in Methyl Orange are the red, weak divalent acid which dissociates in water forming neutral orange molecules. Under acidic conditions, the equilibrium is to the left, and the concentration of neutral molecules is too low for the orange colour to be observed. However, under neutral and alkaline conditions, the equilibrium is to the right, and the concentration of the anion becomes sufficient for the orange colour to be observed.

We can apply equilibrium law to indicator equilibria - in general for a weak acid indicator: $K_{MOR} = ([H_3O^+]^2 \cdot [MOR^-]) / [H_2MOR^{2+}]$ equilibrium constant expression . (7)

K_{MOR} is known as the indicator dissociation constant. The colour of the indicator turns from colour A to colour B or vice versa at its turning point. At this point equal stay:

$$[MOR^-] = [H_2MOR^{2+}]$$

So from equilibrium expression (6,7) is derivate simple equation : $K_{MOR} = [H_3O^+]^2 = 3.98 \cdot 10^{-8}$;

$$pK_{MOR} = -\log(10^{-7.4}) = -(-7.4) = 7.4; \quad pH = -\log(\sqrt{H_3O^+}) = \frac{1}{2} \log(10^{-7.4}) = \frac{1}{2} 7.4 = 3.7 \quad (8)$$

The pH of the solution at its turning point is equal to the pK_{MOR} and is the pH at which half of the indicator is in its acid form and the other half in the form of its conjugate base.

Indicator Range

At a low pH, a weak acid indicator is almost entirely in the $[\text{H}_2\text{MOR}^{2+}]$ form of acid, the colour of which predominates. As the pH increases - the influence of the colour of $[\text{H}_2\text{MOR}^{2+}]$ medium accenting replaces the colour of $[\text{MOR}]$ and the equilibrium is pushed to the right. Therefore the intensity of the colour of $[\text{MOR}]$ increases. Physiological visible condition for certainly distinct colour is assumed at ratio 10 for conjugated base form:

$$[\text{MOR}] / [\text{H}_2\text{MOR}^{2+}] = 10 ; \text{ and for acid coloured form } [\text{MOR}] / [\text{H}_2\text{MOR}^{2+}] = 1 / 10 = 0.1 \quad (9)$$

$$+K_{\text{MOR}} = [\text{H}_3\text{O}^+]^2 \cdot [\text{MOR}] / [\text{H}_2\text{MOR}^{2+}] = 10^{-7.4} \cdot 10 = 10^{-8.4} ; + \text{pH}_{\text{MOR}} = 8.4 / 2 = 4.2 ;$$

$$- K_{\text{MOR}} = [\text{H}_3\text{O}^+]^2 \cdot [\text{MOR}] / [\text{H}_2\text{MOR}^{2+}] = 10^{-7.4} \cdot 0.1 = 10^{-6.4} ; - \text{pH}_{\text{MOR}} = 6.4 / 2 = 3.2 ;$$

$$\text{pH} = -\log(\sqrt{[\text{H}_3\text{O}^+]^2}) \pm 0.5 = \frac{1}{2} \log(10^{-7.4}) \pm 0.5 = \frac{1}{2} 7.4 \pm 0.5 = 3.7 \pm 0.5 \quad , \quad (10)$$

which is colour change interval for divalent weak acid-base indicator ± 0.5 pH units.

An indicator is most effective if the colour change is distinct and over a low pH range. For divalent acid-base indicators the range is within ± 0.5 of the pK_{MOR} value: Please see the table below for examples, to the left is a model of the acid form of each indicator - with the colour of the solution at the turning point. Number 2 of involved protons H^+ in equilibrium or 1 Thymol Blue 1st, 2nd.

Table pH indicators

Indicator	Colour		pH colour turning pH range	
	Acid	Base	point	pK_{In} or $\text{pK}_{\text{In}}/2$ $-\text{pH} < \text{pK}_{\text{In}} < +\text{pH}$
Thymol Blue - 1st change	red	yellow	1.7	1.2 - 2.2
Methyl Orange	red	yellow	3.7	3.2 - 4.2
Bromocresol Green	yellow	blue	4.7	4.2 - 5.2
Methyl Red	yellow	red	5.1	4.6 - 5.6
Bromothymol Blue	yellow	blue	7.0	6.5 - 7.5
Phenol Red	yellow	red	7.9	7.4 - 8.4
Thymol Blue - 2nd change	yellow	blue	8.9	7.9 - 9.9
Phenolphthalein	colourless	pink	9.4	8.9 - 9.9

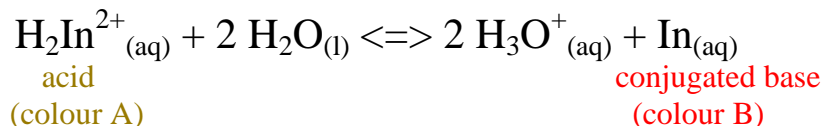
A Universal Indicator is a mixture of indicators which give a gradual change in colour over a wide pH range - the pH of a solution can be approximately identified when a few drops of universal indicator are mixed with the solution.

Indicators are used in titration solutions to signal the completion of the acid-base reaction.

Acid - Base Indicators Methyl Red

Acid - Base indicators (also known as pH indicators) are substances which change colour with pH. They are divalent weak acids or bases, involved in protolytic water acid-base equilibrium and form two type coloured ions as Brensted-Lowry protolytic pairs.

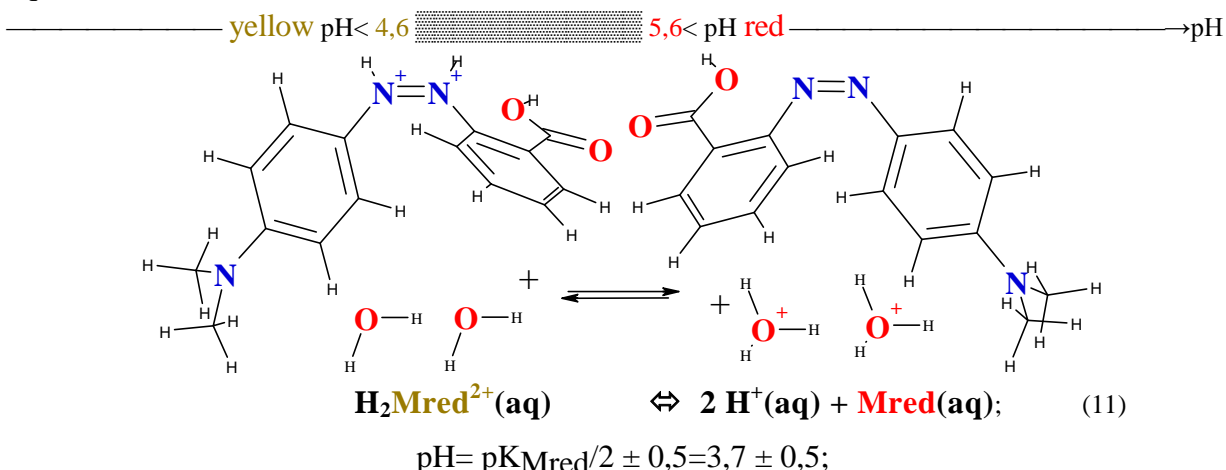
Consider an indicator which is divalent weak acid, with the formula H_2In^{2+} . At equilibrium, the following equilibrium equation is established with its conjugate base In^- :



The acid and its conjugate base have different colours. At low pH values the concentration of H_3O^+ is high and so the equilibrium position lies to the left.

The equilibrium solution has the colour A. At high pH values, the concentration of H_3O^+ is low - the equilibrium position thus lies to the right and the equilibrium solution has colour B.

Methyl Red is an example of an indicator which establishes this type of equilibrium in aqueous solution:



Protonate nitrogen atoms in Methyl Red are the yellow, weak divalent acid which dissociates in water forming red molecules. Under acidic conditions, the equilibrium is to the left, and the concentration of the anions is too low for the red colour to be observed. However, under neutral and alkaline conditions, the equilibrium is to the right, and the concentration of the anion becomes sufficient for the red colour to be observed.

We can apply equilibrium law to indicator equilibrium - in general for a weak acid indicator:

$$K_{Mred} = ([H_3O^+]^2 \cdot [Mred]) / [H_2Mred^{2+}] \text{ equilibrium constant expression .} \quad (12)$$

K_{Mred} is known as the indicator dissociation constant. The colour of the indicator turns from colour A to colour B or vice versa at its turning point. At this point equal stay:

$$[Mred] = [H_2Mred^{2+}]$$

So from equilibrium expression (11,12) is derivate simple equation : $K_{Mred} = [H_3O^+]^2 = 10^{-10,2}$;
 $pK_{Mred} = -\log(10^{-10,2}) = -(-10,2) = 5,1$; $pH = -\log(\sqrt{[H_3O^+]^2}) = -\frac{1}{2} \log(10^{-10,2}) = -\frac{1}{2} \cdot -10,2 = 5,1$. (13)

The pH of the solution at its turning point is equal to the pK_{Mred} and is the pH at which half of the indicator is in its acid form and the other half in the form of its conjugate base.

Indicator Range

At a low pH, a weak acid indicator is almost entirely in the $[\text{H}_2\text{Mred}^{2+}]$ form of acid, the colour of which predominates. As the pH increases - the influence of the colour of $[\text{H}_2\text{Mred}^{2+}]$ medium accenting replaces the colour of $[\text{Mred}]$ and the equilibrium is pushed to the right. Therefore the intensity of the colour of $[\text{Mred}]$ increases. Physiological visible condition for certainly distinct colour is assumed at ratio 10 for conjugated base form:

$$[\text{Mred}]/[\text{H}_2\text{Mred}^{2+}] = 10 ; \text{ and for acid coloured form } [\text{Mred}]/[\text{H}_2\text{Mred}^{2+}] = 1 / 10 = 0.1 \quad (14)$$

$$+K_{\text{MOR}} = [\text{H}_3\text{O}^+]^2 \cdot [\text{Mred}] / [\text{H}_2\text{Mred}^{2+}] = 10^{-10.2} \cdot 10 = 10^{-9.2} ; + \text{pH}_{\text{MOR}} = 9.2 / 2 = 4.6 ;$$

$$- K_{\text{MOR}} = [\text{H}_3\text{O}^+]^2 \cdot [\text{Mred}] / [\text{H}_2\text{Mred}^{2+}] = 10^{-10.2} \cdot 0.1 = 10^{-11.2} ; - \text{pH}_{\text{MOR}} = 11.2 / 2 = 5.5 ;$$

$$\text{pH} = -\log(\sqrt{[\text{H}_3\text{O}^+]^2}) \pm 0.5 = \frac{1}{2} \log(10^{-10.2}) \pm 0.5 = \frac{1}{2} 10.2 \pm 0.5 = 5.1 \pm 0.5 \quad , \quad (15)$$

which is colour change interval for divalent weak acid-base indicator ± 0.5 pH units.

An indicator is most effective if the colour change is distinct and over a low pH range. For divalent acid-base indicators the range is within ± 0.5 of the pK_{MOR} value: Please see the table below for examples, to the left is a model of the acid form of each indicator - with the colour of the solution at the turning point. Number 2 of involved protons H^+ in equilibrium or 1 Thymol Blue 2nd .

Table pH indicators

Indicator	Colour	pH colour		
		turning point	pH range	
	Acid	Base	pK_{In} or $\text{pK}_{\text{In}}/2$	$-\text{pH} < \text{pK}_{\text{In}} < +\text{pH}$
Thymol Blue - 1st change	red	yellow	1.7	1.2 - 2.2
Methyl Orange	red	yellow	3.7	3.2 - 4.2
Bromocresol Green	yellow	blue	4.7	4.2 - 5.2
Methyl Red	yellow	red	5.1	4.6 - 5.6
Bromothymol Blue	yellow	blue	7.0	6.5 - 7.5
Phenol Red	yellow	red	7.9	7.4 - 8.4
Thymol Blue - 2nd change	yellow	blue	8.9	7.9 - 9.9
Phenolphthalein	coloureless	pink	9.4	8.9 - 9.9

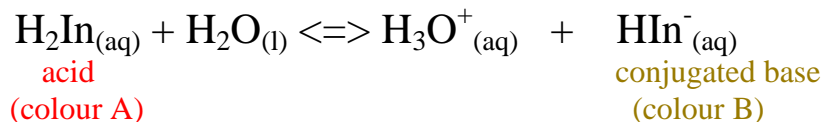
A Universal Indicator is a mixture of indicators which give a gradual change in colour over a wide pH range - the pH of a solution can be approximately identified when a few drops of universal indicator are mixed with the solution.

Indicators are used in titration solutions to signal the completion of the acid-base reaction.

Acid - Base Indicators Thymol Blue 1.7, Thymol Blue 8.9

Acid - Base indicators (also known as pH indicators) are substances which change colour with pH. They are weak acids or bases, involved in protolytic water acid-base equilibrium and form two type coloured ions as Brensted-Lowry protolytic pairs.

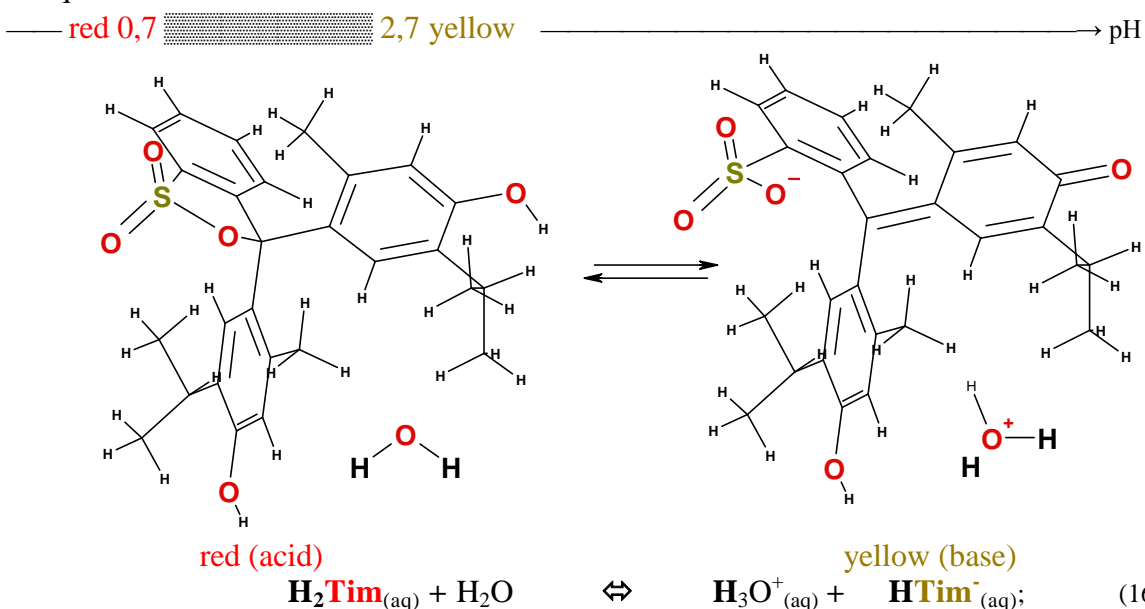
Consider an indicator which is weak acid, with the formula H_2In . At equilibrium, the following equilibrium equation is established with conjugate base HIn^- of different colour:



The acid and its conjugate base have different colours. At low pH values the concentration of H_3O^+ is high and so the equilibrium position lies to the left. The equilibrium solution has the colour A.

At high pH values, the concentration of H_3O^+ is low - the equilibrium position thus lies to the right and the equilibrium solution has colour B.

Thymol Blue 1.7 is an example of an indicator which establishes this type of equilibrium in aqueous solution:



Thymol Blue 1.7 (pdb version) is a red, weak acid which dissociates in water forming yellow anions. Under acidic conditions, the equilibrium is to the left, and the concentration of the anions is too low for the yellow colour to be observed. However, under alkaline conditions, the equilibrium is to the right, and the concentration of the anion becomes sufficient for the yellow colour to be observed.

We can apply equilibrium expression to indicator equilibrium - in general for weak acid indicator:

$$K_{Tim1} = ([H_3O^+] \cdot [HTim^-]) / [H_2Tim] \text{ equilibrium constant.} \quad (17)$$

K_{Tim1} is known as the indicator dissociation constant. The colour of the indicator turns from colour A to colour B or vice versa at its turning point. At this point equal stay:

$$[H_2Tim] = [HTim^-]$$

So from equilibrium expression (16,17) is derivate simple equation : $K_{Tim1} = [H_3O^+] = 10^{-1.7}$; $pK_{Tim1} = -\log(10^{-1.7}) = -(-1.7) = 1.7$; $pH = -\log([H_3O^+]) = -\log(10^{-1.7}) = 1.7$ (18)

The pH of the solution at its turning point is equal to the pK_{Tim1} and is the pH at which half of the indicator is in its acid form and the other half in the form of its conjugate base.

Indicator Range

At a low pH, a weak acid indicator is almost entirely in the $[\text{H}_2\text{Tim}]$ form of acid, the colour of which predominates. As the pH increases - the influence of the red colour of $[\text{H}_2\text{Tim}]$ medium accenting replaces the colour of $[\text{HTim}^-]$ and the equilibrium is pushed to the right. Therefore the intensity of the colour of $[\text{HTim}^-]$ increases. Physiological visible condition for certainly distinct colour is assumed at ratio 10 for conjugated base form:

$$[\text{HTim}^-] / [\text{H}_2\text{Tim}] = 10 ; \text{ and for acid coloured form } [\text{HTim}^-] / [\text{H}_2\text{Tim}] = 1 / 10 = 0.1 \quad (19)$$

$$+K_{\text{Tim1}} = [\text{H}_3\text{O}^+] \cdot [\text{HTim}^-] / [\text{H}_2\text{Tim}] = 10^{-1.7} \cdot 10 = 10^{-0.7} ; +\text{pH}_{\text{Tim1}} = 0.7 ;$$

$$-K_{\text{Tim1}} = [\text{H}_3\text{O}^+] \cdot [\text{HTim}^-] / [\text{H}_2\text{Tim}] = 10^{-1.7} \cdot 0.1 = 10^{-2.7} ; -\text{pH}_{\text{Tim1}} = 2.7 ;$$

$$\text{pH} = -\log([\text{H}_3\text{O}^+]) \pm 1 = -\log(10^{-1.7}) \pm 1 = 1.7 \pm 1 \quad , \quad (20)$$

which is colour change interval for mono protic weak acid-base indicator ± 1 pH units.

An indicator is most effective if the colour change is distinct and over a low pH range. For mono protic acid-base indicators the range is within ± 1 of the pK_{In} value: Please see the table below for examples, to the left is a model of the acid form of each indicator - with the colour of the solution at the turning point. One of involved proton H^+ in equilibrium for Thymol Blue 1st or 1 for Thymol Blue 2nd .

Table pH indicators

Indicator	Colour	pH colour		
		Acid	Base	turning pH range
			point	pK_{In} or $\text{pK}_{\text{In}}/2$
				$-\text{pH} < \text{pK}_{\text{In}} > +\text{pH}$
Thymol Blue - 1st change	red	yellow	1.7	1.2 - 2.2
Methyl Orange	red	yellow	3.7	3.2 - 4.2
Bromocresol Green	yellow	blue	4.7	4.2 - 5.2
Methyl Red	yellow	red	5.1	4.6 - 5.6
Bromothymol Blue	yellow	blue	7.0	6.5 - 7.5
Phenol Red	yellow	red	7.9	7.4 - 8.4
Thymol Blue - 2nd change	yellow	blue	8.9	7.9 - 9.9
Phenolphthalein	coloureless	pink	9.4	8.9 - 9.9

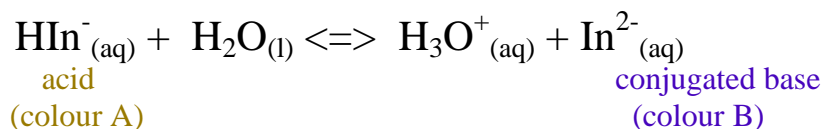
A Universal Indicator is a mixture of indicators which give a gradual change in colour over a wide pH range - the pH of a solution can be approximately identified when a few drops of universal indicator are mixed with the solution.

Indicators are used in titration solutions to signal the completion of the acid-base reaction.

Acid - Base Indicators Thymol Blue 8.9

Acid - Base indicators (also known as pH indicators) are substances which change colour with pH. They are weak acids or bases, involved in protolytic water acid-base equilibrium and form two type coloured ions as Brensted-Lowry protolytic pairs.

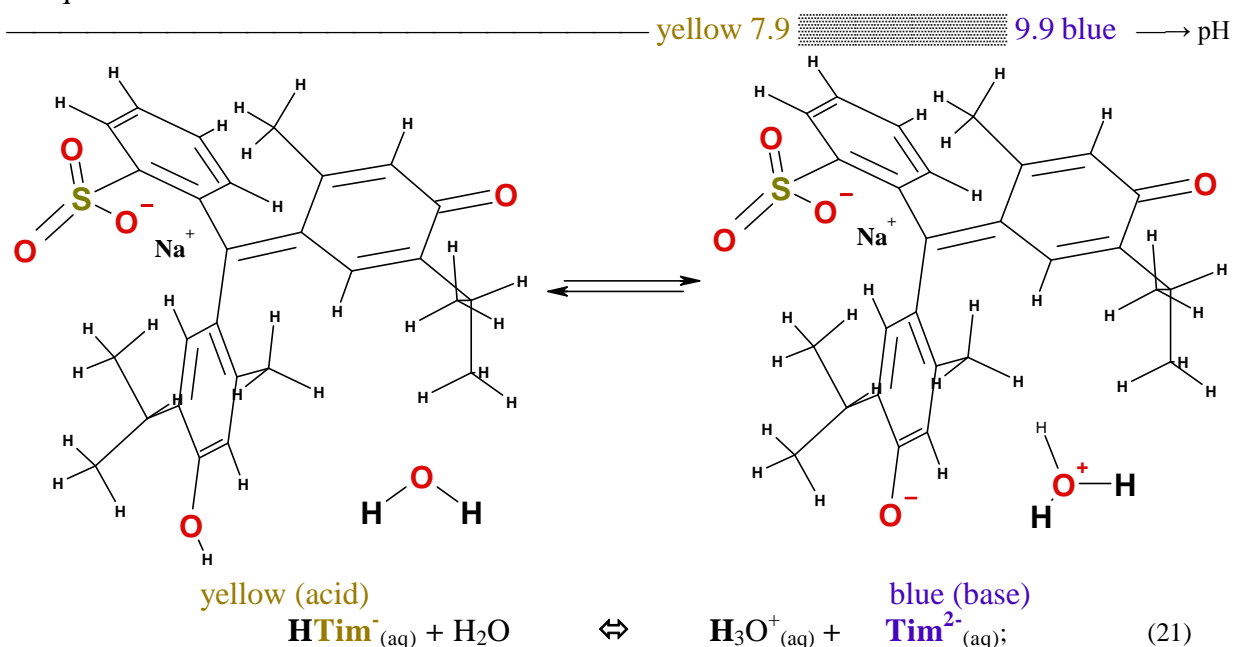
Consider an indicator which is weak acid, with the formula HIn^- . At equilibrium, the following equilibrium equation is established with conjugate base In^{2-} of different colour:



The acid and its conjugate base have different colours. At low pH values the concentration of H_3O^+ is high and so the equilibrium position lies to the left. The equilibrium solution has the colour A.

At high pH values, the concentration of H_3O^+ is low - the equilibrium position thus lies to the right and the equilibrium solution has colour B.

Thymol Blue 8.9 is an example of an indicator which establishes this type of equilibrium in aqueous solution:



Thymol Blue 8.9 (pdb version) is a **yellow** weak acid which dissociates in water forming **blue** anions. Under acidic conditions, the equilibrium is to the left, and the concentration of the anions is too low for the **yellow** colour to be observed. However, under alkaline conditions, the equilibrium is to the right, and the concentration of the anion becomes sufficient for the **yellow** colour to be observed.

We can apply equilibrium expression to indicator equilibrium - in general for weak acid indicator:

$$K_{\text{Tim}2} = ([\text{H}_3\text{O}^+] \cdot [\text{Tim}^{2-}]) / [\text{HTim}^-] \text{ equilibrium constant.} \quad (22)$$

$K_{\text{Tim}2}$ is known as the indicator dissociation constant. The colour of the indicator turns from colour A to colour B or vice versa at its turning point. At this point equal stay:

$$[\text{HTim}^-] = [\text{Tim}^{2-}]$$

So from equilibrium expression (21,22) is derivate simple equation : $K_{\text{Tim}2} = [\text{H}_3\text{O}^+] = 10^{-8.9}$; $\text{p}K_{\text{Tim}2} = -\log(10^{-8.9}) = -(-8.9) = 8.9$; $\text{pH} = -\log([\text{H}_3\text{O}^+]) = -\log(10^{-8.9}) = 8.9$ (23)

The pH of the solution at its turning point is equal to the $\text{p}K_{\text{Tim}2}$ and is the pH at which half of the indicator is in its acid form and the other half in the form of its conjugate base.

Indicator Range

At a low pH, a weak acid indicator is almost entirely in the $[\text{HTim}^-]$ form of acid, the colour of which predominates. As the pH increases - the influence of the yellow of $[\text{HTim}^-]$ medium accenting replaces the colour of $[\text{Tim}^{2-}]$ and the equilibrium is pushed to the right. Therefore the intensity of the colour of $[\text{Tim}^{2-}]$ increases. Physiological visible condition for certainly distinct colour is assumed at ratio 10 for conjugated base form:

$$[\text{Tim}^{2-}] / [\text{HTim}^-] = 10; \text{ and for acid coloured form } [\text{Tim}^{2-}] / [\text{HTim}^-] = 1 / 10 = 0.1 \quad (24)$$

$$+K_{\text{Tim}2} = [\text{H}_3\text{O}^+] \cdot [\text{Tim}^{2-}] / [\text{HTim}^-] = 10^{-8.9} \cdot 10 = 10^{-7.9}; \text{ + pH}_{\text{Tim}2} = 7.9;$$

$$-K_{\text{Tim}2} = [\text{H}_3\text{O}^+] \cdot [\text{Tim}^{2-}] / [\text{HTim}^-] = 10^{-8.9} \cdot 0.1 = 10^{-9.9}; \text{ - pH}_{\text{Tim}2} = 9.9;$$

$$\text{pH} = -\log([\text{H}_3\text{O}^+]) \pm 1 = -\log(10^{-8.9}) \pm 1 = 8.9 \pm 1, \quad (25)$$

which is colour change interval for mono protic weak acid-base indicator ± 1 pH units.

An indicator is most effective if the colour change is distinct and over a low pH range. For mono protic acid-base indicators the range is within ± 1 of the pK_{In} value: Please see the table below for examples, to the left is a model of the acid form of each indicator - with the colour of the solution at the turning point. One of involved proton H^+ in equilibrium for 1 Thymol Blue 2nd.

Table pH indicators

Indicator	Colour		pH colour turning pH range	
	Acid	Base	point	pK_{In} or $\text{pK}_{\text{In}}/2$ $-\text{pH} < \text{pK}_{\text{In}} < +\text{pH}$
Thymol Blue - 1st change	red	yellow	1.7	1.2 - 2.2
Methyl Orange	red	yellow	3.7	3.2 - 4.2
Bromocresol Green	yellow	blue	4.7	4.2 - 5.2
Methyl Red	yellow	red	5.1	4.6 - 5.6
Bromothymol Blue	yellow	blue	7.0	6.5 - 7.5
Phenol Red	yellow	red	7.9	7.4 - 8.4
Thymol Blue - 2nd change	yellow	blue	8.9	7.9 - 9.9
Phenolphthalein	coloureless	pink	9.4	8.9 - 9.9

A Universal Indicator is a mixture of indicators which give a gradual change in colour over a wide pH range - the pH of a solution can be approximately identified when a few drops of universal indicator are mixed with the solution.

Indicators are used in titration solutions to signal the completion of the acid-base reaction.