

Acid - base indicators

Indicators are Substances those can change their color when added to solutions of acids or bases. Indicators have different colours in solution of different pH. pH of the solution can be determined by adding one or two drops of the indicator solution.

Indicators and their color in different medium.

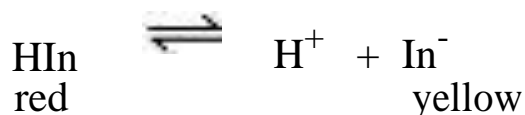
Indicators	Color in acidic medium	Color in alkaline medium
Phenolphthalein	colorless	pink
Methyl orange	reddish orange	yellow
Methyl red	red	yellow
Thymol blue	red	yellow

Indicators used in different types of acid base titration

Nature of acid base	Example	Suitable indicators
Strong acid- strong base	HCl – Na ₂ CO ₃	Most of the indicators
weak acid- strong base	CH ₃ COOH-NaOH	Phenolphthalein
Strong acid- weak base	HCl-NH ₄ OH	Methyl orange, Methyl red
weak acid- weak base	CH ₃ COOH- NH ₄ OH	No suitable indicators

How an acid-base indicator works:

Let us explain the indicator action by taking example of methyl orange. Methyl orange is a weak acid and gives the following ionization equilibrium in solution.



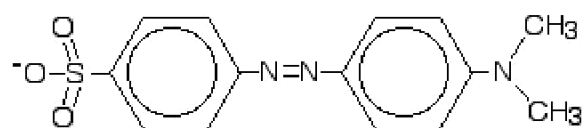
In accordance with the law of mass action, where K_{in} is the dissociation constant of the indicator and is called the Indicator constant.

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

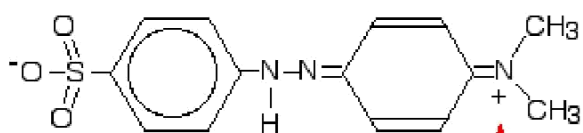
The anion In^- is **yellow** and the **nonionised form HIn** is **red**. If an acid is added to the solution, the hydrogen ion concentration, $[\text{H}^+]$, in the equilibrium expression increases. To maintain K_{in} constant, the equilibrium shifts to the left. Thereby the concentration of $[\text{In}^-]$ is reduced and the concentration of $[\text{HIn}]$ increases so that the solution is red.

On the other hand, upon addition of a base to the solution, H^+ ions are removed as H_2O by reacting with OH^- ions of the base. This shifts the equilibrium to the right, resulting in the increase of In^- ions that are yellow. Thus in acid solution the non-ionised HIn molecules predominate and the solution is red, while in basic solution In^- ions are in excess and the solution is yellow.

the yellow form of methyl orange



the red form of methyl orange

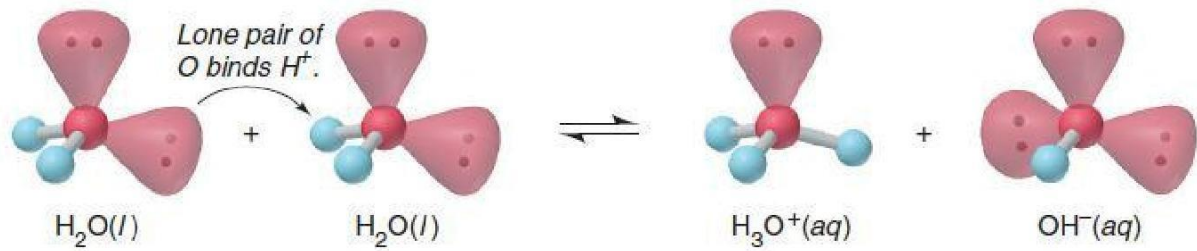


This hydrogen ion is attached here.

Notice the positive charge on the nitrogen.

Auto-ionization of Water

-reaction in which two like molecules react to give ions.



$$K_c = \frac{[\text{H}_3\text{O}^+][\text{OH}^-]}{[\text{H}_2\text{O}]^2}$$

Because the concentration of ions formed is very small, the concentration of H_2O remains essentially constant, about 56 M at 25°C. So you can write-

$$\underbrace{[\text{H}_2\text{O}]^2}_{\text{constant}} K_c = [\text{H}_3\text{O}^+][\text{OH}^-]$$

Ionic product, $K_w = [\text{H}_3\text{O}^+][\text{OH}^-]$

At 25°C,

$$K_w = 1.0 \times 10^{-14}$$

$$\text{So, } K_w = [\text{H}_3\text{O}^+][\text{OH}^-] = 1.0 \times 10^{-14}$$

$$\text{Since, } [\text{H}^+] = [\text{H}_3\text{O}^+]$$

$$\text{Now, } K_w = [\text{H}^+][\text{OH}^-]$$

In pure water, $[\text{H}^+] = [\text{OH}^-]$

$$\text{So, } K_w = [\text{H}^+][\text{H}^+] = 1.0 \times 10^{-14}$$

$$[\text{H}^+] = 10^{-7}$$

$$[\text{OH}^-] = 10^{-7}$$

When, $[\text{H}^+] = [\text{OH}^-]$neutral

$[\text{H}^+] > [\text{OH}^-]$acidic

$[\text{H}^+] < [\text{OH}^-]$ basic

At body temperature (37 °C),

$$K_w = 2.5 \times 10^{-14}$$

pH scale

- 'pH of a solution is the negative logarithm(to the base 10) of the hydrogen ion (strictly speaking H_3O^+) concentration (in mol L^{-1})'.

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

$$\text{pOH} = -\log[\text{OH}^-]$$

From the ionic product of water, $[\text{H}^+][\text{OH}^-] = 1.0 \times 10^{-14}$

$$\log[\text{H}^+] + \log[\text{OH}^-] = -14$$

$$(-\log[\text{H}^+]) + (-\log[\text{OH}^-]) =$$

$$14 \quad \text{pH} + \text{pOH} = 14$$

pH = 7neutral

solution pH >7basic

solution pH < 7acidic

solution

Problem In an art restoration project, a conservator prepares copper-plate etching solutions by diluting concentrated HNO_3 to 2.0 M , 0.30 M , and 0.0063 M HNO_3 . Calculate $[\text{H}_3\text{O}^+]$, pH , $[\text{OH}^-]$, and pOH of the three solutions at 25°C .

Plan We know from its formula that HNO_3 is a strong acid, so it dissociates completely; thus, $[\text{H}_3\text{O}^+] = [\text{HNO}_3]_{\text{init}}$. We use the given concentrations and the value of K_w at 25°C (1.0×10^{-14}) to find $[\text{H}_3\text{O}^+]$ and $[\text{OH}^-]$ and then use them to calculate pH and pOH .

Solution Calculating the values for 2.0 M HNO_3 :

$$[\text{H}_3\text{O}^+] = 2.0\text{ M}$$

$$\text{pH} = -\log [\text{H}_3\text{O}^+] = -\log 2.0 = -0.30$$

$$[\text{OH}^-] = \frac{K_w}{[\text{H}_3\text{O}^+]} = \frac{1.0 \times 10^{-14}}{2.0} = 5.0 \times 10^{-15}\text{ M}$$

$$\text{pOH} = -\log (5.0 \times 10^{-15}) = 14.30$$

Calculating the values for 0.30 M HNO_3 :

$$[\text{H}_3\text{O}^+] = 0.30\text{ M}$$

$$\text{pH} = -\log [\text{H}_3\text{O}^+] = -\log 0.30 = 0.52$$

$$[\text{OH}^-] = \frac{K_w}{[\text{H}_3\text{O}^+]} = \frac{1.0 \times 10^{-14}}{0.30} = 3.3 \times 10^{-14}\text{ M}$$

$$\text{pOH} = -\log (3.3 \times 10^{-14}) = 13.48$$

Calculating the values for 0.0063 M HNO_3 :

$$[\text{H}_3\text{O}^+] = 6.3 \times 10^{-3}\text{ M}$$

$$\text{pH} = -\log [\text{H}_3\text{O}^+] = -\log (6.3 \times 10^{-3}) = 2.20$$

$$[\text{OH}^-] = \frac{K_w}{[\text{H}_3\text{O}^+]} = \frac{1.0 \times 10^{-14}}{6.3 \times 10^{-3}} = 1.6 \times 10^{-12}\text{ M}$$

$$\text{pOH} = -\log (1.6 \times 10^{-12}) = 11.80$$