HOW DO pH INDICATORS WORK?

pH indicators detect the presence of H^+ and OH^-. They do this by reacting with H^+ and OH^-: they are themselves weak acids and bases. If an indicator is a weak acid and is coloured and its conjugate base has a different colour, deprotonation causes a colour change.

The ratio of the concentration of the indicator, HInd, and its conjugate base, Ind^-, determines the colour we see. This ratio depends on the pK_a of the indicator and the pH according to the Henderson-Hasselbalch equation:

$$\text{pH} = \text{p}K_a + \log\frac{[\text{Ind}^-]}{[\text{HInd}]}$$

Remember: $\log x < 0$ when $x < 1$,
$\log x = 0$ when $x = 1$ and
$\log x > 0$ when $x > 1$

1. If the pH of the solution is *equal* to the pK_a of the indicator, which of the following is true?

   (a) $[\text{Ind}^-] < [\text{HInd}]$ **(b) $[\text{Ind}^-] = [\text{HInd}]$** (c) $[\text{Ind}^-] > [\text{HInd}]$

2. If the pH of the solution is *lower* than the pK_a of the indicator, which of the following is true?

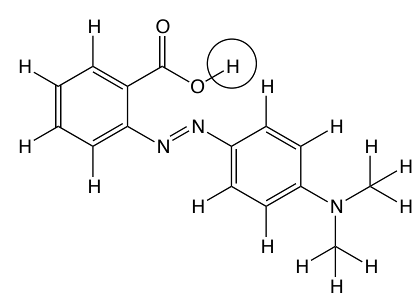
   (a) $[\text{Ind}^-] < [\text{HInd}]$ (b) $[\text{Ind}^-] = [\text{HInd}]$ (c) $[\text{Ind}^-] > [\text{HInd}]$

3. If the pH of the solution is *higher* than the pK_a of the indicator, which of the following is true?

   (a) $[\text{Ind}^-] < [\text{HInd}]$ (b) $[\text{Ind}^-] = [\text{HInd}]$ **(c) $[\text{Ind}^-] > [\text{HInd}]$**

4. If HInd and Ind^- are different colours, at about what pH do you expect the colour of an indicator solution to change? **The pH will be approximately equal to the indicator's pK_a value**

5. The molecular structure of the indicator *methyl red* (pK_a of 5.1) is shown on the right. Circle the H atom that is *not* present in its conjugate base.

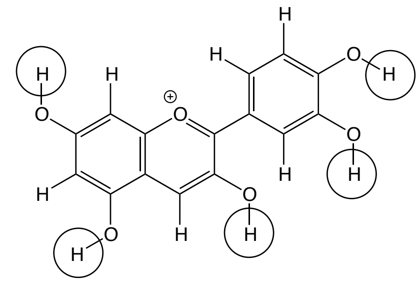


6. The deprotonated form of this indicator is yellow and the protonated form is red. What colour do you **expect** a solution to be when acid is added to a neutral solution containing the indicator? Add your prediction to the first cell in the table below. **Red**

7. Complete the table by adding your predictions and observations. In the final column, indicate the form of the indicator that is present (HInd or Ind^-)

Solution	**Prediction**	**Observation**	**Form present**
Acidic solution containing methyl red		**Red**	**Protonated**
Neutral solution containing methyl red		**Yellow**	**Deprotonated**
Basic solution containing methyl red		**Yellow**	**Deprotonated**

NATURAL pH INDICATORS



The molecule in red cabbage responsible for its colour is an *anthocyanin*. Anthocyanins are a large group of plant pigments that occur in all higher plants including flowers and fruits. They are responsible for the colours of many of our foods, including blueberries to red wine. The structure of the anthocyanin present in red cabbage is shown on the right.

1. Circle one of the H atoms you expect not to be present in the deprotonated form of this anthocyanin.

2. Complete the table below by noting down the colour changes that occurred when acid and base are added to the solution containing the natural indicator.

Indicator	Acidic	Neutral	Basic
Red cabbage	**red / pink**	**purple**	**blue / green**
Red wine			
Tea	**light brown**	**brown**	**dark brown**

3. Predict the colour of 'red cabbage' when it is grown in the following soils:

   (a) acidic soil - **red** (b) neutral soil - **purple** (c) alkaline soil – **blue / green**

4. Red cabbage turns a distasteful blue colour when cooked. What would you add to a saucepan of red cabbage so that it keeps its red colour? **Vinegar (acetic acid)**

5. Iced tea is often much paler in colour than traditional hot tea. What ingredient is added to cause this colour change? **Lemon juice (citric acid)**

Challenge Question

1. Why do we need indicators which change colours at pH values apart from 7.0?

   Hints. Complete the chemical equation below. What is present when enough OH^- has been added to react with *all* of the carboxylic acid? What is the pH of a solution containing this?

   $CH_3COOH(aq) + OH^-(aq) \rightarrow$ **$CH_3COO^-(aq) + H_2O(l)$**

   The solution contains CH_3COO^-. This is the conjugate base of CH_3COOH. It is a base so the pH will be *higher* than 7.0.