

Chemistry 4, Assessable Prac 1

Semester 2, 2001

Finding $[H^+]$ with indicators.

Joseph Curtis;

Aim:

- 1) To determine the $[H^+]$ of unknown solutions using the pH established by help of a colour scale derived by adding indicators to acids and bases of known molarity.
- 2) To determine the K_a of Acetic acid, using the pH established by help of a colour scale derived by adding indicators to acids and bases of known pH.

Conclusion:

The $[H^+]$ of unknown solution A established by help of a colour scale derived by adding indicators to acids and bases of known molarity is approximately 0.055 moles/L and the pH is approximately 1.26.

The $[H^+]$ of unknown solution B established by help of a colour scale derived by adding indicators to acids and bases of known molarity is approximately 1.82×10^{-13} moles/L and the pH is approximately 12.74.

The K_a of acetic acid using the pH established by help of a colour scale derived by adding indicators to acids and bases of known pH is approximately 1.6×10^{-5} M.

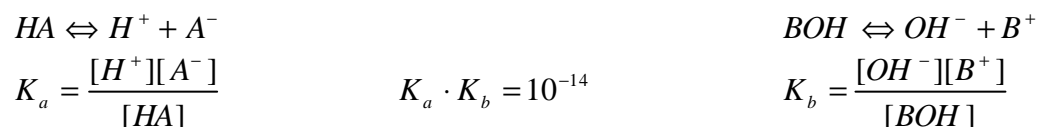
Background:

According to the Lowry-Brønsted concept of acids and bases:

*An **acid** is a substance, which in solution, tends to give up protons (hydrogen ions), and a **base** is a substance which tends to accept protons.*

$[H^+]$ is the concentration of dissociated hydrogen ions in solution when any acid is dissolved in water. Acids which readily part with their hydrogen ions and dissociate in water are called strong acids, such as HCl and HNO_3 . In solution, strong acids almost fully dissociate into hydrogen ions and the conjugate base ions. This means the concentration of the hydrogen ions and the concentration of the acid are almost exactly the same $[H^+] \approx [HA]$. Also for a strong base the concentration of the hydroxide ions approximately equals the concentration of the base itself $[OH^-] \approx [BOH]$.

The equilibrium constant for the ionization of an acid (K_a) or base (K_b) is a measure of acid/base strength. It is merely the equilibrium constant for the ionization reaction of an acid/base with water.



The stronger an acid is the larger the K_a will be. This is because acids that are stronger dissociate more readily and therefore there will be a larger concentration of hydrogen ions in solution. There will be a larger numerator and a smaller denominator and therefore a larger K_a . The above also applies to bases

In reactions involving weak acids the concentration of the acid cannot be used as an approximation of the hydrogen ion concentration in solution because weak acids don't dissociate fully and so the $[H^+]$ will be smaller than the [acid]. In such reactions the K_a of that acid is used to determine the amount of acid that successfully dissociates into hydrogen ions.

Because the values of the concentrations of dissociated hydrogen in water vary greatly between different acids it becomes inconvenient to refer to the strength of an acid or base by the hydrogen ion concentration. To make it easier to measure the strengths of acids the pH scale was introduced. The pH of a solution is defined as the negative of the logarithm (to base 10) of the hydrogen ion concentration.

$$pH = -\log_{10}[H^+] \qquad [H^+] = 10^{-pH}$$

The pH scale ranges from 0 to 14. Pure water and all neutral substances has a pH of 7, acidic substances have a pH lower than 7 and basic substances have a pH of over 7. For

some very strong acids a negative pH is possible likewise with a strong base a pH of over 14 is also possible, however conventionally the scale goes from 0 to 14.

The pOH of a substance can be discovered by getting the negative logarithm (to base 10) of the hydroxide ion concentration. The $\text{pH} = 14 - \text{pOH}$, thus a pH can be found from only knowing the $[\text{OH}^-]$.

Indicators are weak acids or bases whose conjugate acid/base pairs have different colours. Methyl orange is an indicator (represented by HIn) that has its colour change in the acidic range. In acid it is red (HIn), neutral, orange ($\text{HIn} + \text{In}^-$) and in basic solutions it is yellow (In^-).



If there is excess acid there are more hydrogen ions and the equilibrium will shift to the left forming more HIn and will result in red colour. If there is excess base, hydroxide ions will react with H_3O^+ to form water, thus removing the hydrogen ions. This causes the equilibrium to shift to the right, forming more In^- and resulting in a yellow colour. In a neutral solution there will be no excess of acid or base so there will be equal amounts of HIn and In^- . The resulting colour is then orange, a mixture of red and yellow.

The four indicators used in this experiment are methyl orange, orange 1V, indigo carmine and alizarin yellow R. Methyl orange and orange 1V both have colour changes in the acidic range. Indigo carmine and alizarin yellow R both have their colour changes in the basic range.

Method:

Part 1 – The preparation of standard solutions.

1. 8 test tubes were collected and labeled 1-8.
2. Into test tube 1, 5ml 0.1M HCl was added.
3. Into test tube 2, 9mls of water were put and then 1ml of 0.1M HCl from test tube 1 was added.
4. Into test tube 3, 9mls of water were put and then 1ml of 0.01M HCl from test tube 2 was added.
5. Into test tube 4, 9mls of water were put and then 1ml of 0.001M HCl from test tube 3 was added.
6. Test tube 5 was labeled 0.1M and into it were put 2mls of the standard solution from test tube 1.
7. Test tube 6 was labeled 0.01M and into it were poured 2mls from test tube 2.
8. Test tube 7 was labeled 0.001M and into it were poured 2mls from test tube 3.
9. Test tube 8 was labeled 0.0001M and into it were poured 2mls from test tube 4.
10. To test tubes 1-4, 2 drops of methyl orange were added.
11. To test tubes 5-8, 2 drops of orange 1V were added.

12. A table was constructed and the colours were recorded for each indicator against the relative molarities.

The Above procedure was repeated to step 10, except the solutions were prepared in the basic range using 0.1M NaOH instead of 0.1M HCl.

For step 10 instead of using methyl orange, 2 drops of indigo carmine were used.

For step 11 instead of using orange 1V, 2 drops of alizarin yellow R were used.

Step 12 was repeated.

Part 2 – Determining the concentration of the Unknown Solutions.

1. 5mls of unknown solution A were obtained.
2. The pH was tested using litmus paper.
3. The solution was divided between 2 test tubes.
4. If the solution was acidic, 2 drops of methyl orange were added to 1 test tube and 2 drops of orange 1V were added to the other.
5. If the solution was basic, 2 drops of indigo carmine were added to 1 test tube and 2 drops of alizarin yellow R were added to the other.
6. The colours were the compared to the standards and recorded.

This was repeated with solution B.

Part 3 – Determining the $[H^+]$ of a weak acid.

1. 5mls of 0.1M acetic acid were put into a test tube.
2. The sample was divided equally between 2 test tubes.
3. To 1 test tube, 2 drops of methyl orange were added and to the second 2 drops of orange 1V.
4. The colours were compared with the standards and recorded.

Data:

Part 1.

Acidic Range:

Tube Number	Molarity (moles/L)	Colour
1	0.1	Red
2	0.01	Dark pink/red
3	0.001	Pink/red
4	0.0001	Light yellow
5	0.1	Dark purple
6	0.01	Pink
7	0.001	Light orange
8	0.0001	Light yellow/green

Basic Range:

Tube Number	Molarity (moles/L)	Colour
1	0.1	Light green
2	0.01	Very light purple
3	0.001	Very light purple/blue
4	0.0001	Extremely light blue
5	0.1	Light orange
6	0.01	Light orange/yellow
7	0.001	Very light green/yellow
8	0.0001	Very light green

Part 2.**Solution A was acidic.**

With the indicator methyl orange, solution A turned red (colour of tube 1).

With the indicator orange 1V, solution A turned pink (colour of tube 6).

Solution B was basic.

With the indicator indigo carmine, solution B turned light green (colour of tube 1).

With the indicator alizarin yellow R, solution B turned light orange/yellow (colour of tube 6).

Part 3.

With the indicator methyl orange, the acetic acid turned dark pink/red (colour of tube 2).

With the indicator orange 1V, the acetic acid turned light orange/yellow (colour between tubes 7 and 8).

Analysis:

Question 1.

a) The $[H^+]$ and pH values of the solutions in the acidic range.

Tube number	$[H^+]$ moles/L	pH
1	0.1	1
2	0.01	2
3	0.001	3
4	0.0001	4
5	0.1	1
6	0.01	2
7	0.001	3

8	0.0001	4
---	--------	---

The $[H^+]$ values are the same as the concentration of the HCl because it is a strong acid and therefore dissociates fully into $[H^+]$ and Cl.

b) The $[H^+]$ and pH values of the solutions in the basic range.

Tube number	$[OH^-]$ moles/L	$[H^+]$ moles/L	pH
1	0.1	10^{-13}	13
2	0.01	10^{-12}	12
3	0.001	10^{-11}	11
4	0.0001	10^{-10}	10
5	0.1	10^{-13}	13
6	0.01	10^{-12}	12
7	0.001	10^{-11}	11
8	0.0001	10^{-10}	10

c) Determine the $[H^+]$ of the unknown solutions.

Unknown solution A.

In methyl orange, the colour of the unknown solution A was the same as tube 1. In orange 1V, the colour of the unknown solution A was the same as tube 6. Therefore the $[H^+]$ of the unknown solution A is between that of tube 1 and that of tube 6. The $[H^+]$ of the unknown solution A is approximately 0.055 moles/L, the pH of the solution is 1.26.

Unknown solution B.

In indigo carmine, the colour of the unknown solution B was the same as in tube 1. In the alizarin yellow R, the colour of the solution B was the same as in tube 6. Therefore the $[OH^-]$ of the solution was between that of tubes 1 and 6. The $[OH^-]$ of the unknown solution B is approximately 0.055 moles/L, the $[H^+]$ of the solution is 1.82×10^{-13} moles/L, the pH is 12.74.

Question 2.

a) Determine the $[H^+]$ of the weak acid.

In methyl orange, the solution of the weak acid (acetic acid) turned the colour of tube 2. In the orange 1V, the colour of the acetic acid was between those of tubes 7 and 8. Therefore, the $[H^+]$ of the acetic acid solution is between 0.01M and 0.001M, and is approximately 0.00125M, the pH of the acid is approximately 2.9.

b) Calculate the equilibrium constant for acetic acid.

$[H^+]$ of the acetic acid \approx 0.00125M

$$CH_3COOH \rightleftharpoons CH_3COO^- + H^+$$

Moles	1	1	1
Conc. at start	0.1	0	0
Conc. at equil	0.09875	0.00125	0.00125

$$K_a = \frac{[CH_3COO^-][H^+]}{[CH_3COOH]}$$

$$K_a = \frac{(0.00125)^2}{0.09875}$$

$$K_a \approx 1.6 \times 10^{-5} M$$

The equilibrium constant (K_a) for acetic acid is approximately equal to $1.6 \times 10^{-5} M$.

QUESTIONS

1.

- When the salt sodium acetate is dissolved in the 0.1M acetic acid solution, the salt dissociates to form CH_3COO^- and Na^+ . The excess of the acetate ion in the solution reacts with the hydrogen ions to produce more acetic acid, thus shifting the equilibrium constant to the left. This would cause the $[H^+]$ to decrease because the H^+ have reacted with the acetate ions forming more acid.
- When the sodium hydroxide solution is added to the acetic acid solution, it dissociates to form Na^+ and OH^- . The OH^- then react with the hydrogen ions in the solution to produce water. Because some of the products (H^+) have been taken away (by the OH^-), the equilibrium shifts to the right to relieve that stress and the hydrogen ion concentration in the solution increases.
- When the hydrogen ion concentration in the solution increases because the equilibrium shifts to the right, the pH of the solution then decreases. The higher the $[H^+]$, the lower the pH.

Errors:

When preparing the standard solutions, if the concentrations were incorrect or if one failed to correctly dilute the acids and bases to the intended molarities. This error would cause major problems with the optimal and satisfactory completion of the rest of the experiment, because all of the tested solutions from which the result data is extrapolated use the standard solutions as a guide to determine the hydrogen ion concentration, or in the case of the basic scale, the hydroxide ion concentration. Performing this error would immediately invalidate the experiment and one would have to re-prepare these solutions correctly to get any valid answers.

Comparing the colour of a test solution incorrectly to that of the standard solutions would obviously give you a wrong answer for the $[H^+]$ or the $[OH^-]$ of that solution being tested. Using some kind of spectrometer to measure the colours would be a much more effective method of comparing colours than the human eye which can make a mistake so easily and destroy the results for the rest of the experiment.

If a slightly different amount of indicator were added to a different test tube, the shades of the resulting colour would be different. Or if a greater volume of solution was used, the indicator would be less concentrated in that solution and therefore a lighter shade, making discerning the correct shade of colour rather difficult, resulting in errors in the calculations of the results.

Because of the fact that the scale of the standard solutions is only very limited, with a large space between the gradations, if an acid is somewhere between two different colours, it may look somewhat like both of them and be rather difficult to place on the scale. A visual colour scale is really not the best way to find the hydrogen ion concentration of a substance, or its pH, it has so many possible errors and it would be much more accurate to use a pH meter or some other accurate device for an experiment like this.