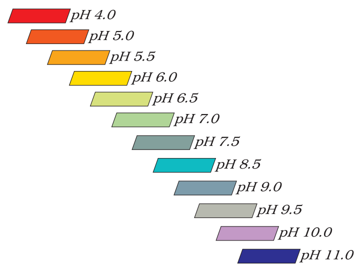
UNIT-5

pH AND pH CHANGE IN AQUEOUS SOLUTIONS



YOU have already performed experiments on dynamic equilibrium between unionised salt and the ions produced by it on dissolving in a solvent. In this unit we will learn about shift in ionic equilibrium between unionised water molecules and H^+ and OH^- ions. The conductivity experiments prove that even pure water ionises to some extent although it has very low conductivity. On this basis it can be concluded that ionic equilibrium exists in pure water also. This ionic equilibrium can be represented as

$$H_2O(l) \rightleftharpoons H^+(aq) + OH^-(aq)$$

Since H^+ ion cannot have independent existence in water because of its positive charge and small ionic radius, a better representation of this equilibrium is

$$2H_2O(l) \rightleftharpoons H_3O^+(aq) + OH^-(aq)$$

This is **self ionisation** of water. Equilibrium constant for this chemical equation can be written as follows:

$$K = \frac{[H_3O^+]\,[OH^-]}{[H_2O]^2}$$

Since water is in large excess, its concentration can be assumed to be constant and combining it with K provides a new constant K_w, which can be written as follows:

$$K_w = [H_3O^+]\,[OH^-]$$

K_w is self ionisation constant of water or simply ionization constant of water. It remains constant at constant temperature. At 25°C value of K_w is 1.0×10^{-14}. Thus, it is quite evident that at a given temperature in any aqueous solution, this product i.e. $[H_3O^+]\,[OH^-]$ remains constant whether acidic, alkaline or neutral in nature. If dissolution of a substance shifts the equilibrium in such a way that at equilibrium the hydronium ion concentration is more than hydroxyl ion concentration then the solution is acidic in nature. If dissolution of a substance shifts the equilibrium in such a way that at equilibrium concentration of OH^- ions is greater than the concentration of hydronium ions, then the solution is alkaline in nature. Thus, concentration of hydronium ion in an aqueous solution can provide information about acidic, basic and neutral nature of the solution. The concentration of H_3O^+ ions in a solution is measured in terms of pH which is defined as the negative logarithm of hydronium ion concentration and is given by the following expression.

$$pH = -\log\,[H_3O^+]$$

At room temperature pH of neutral water is 7. A solution with pH less than this is acidic while the solution with pH greater than this is basic in nature.

EXPERIMENT 5.1

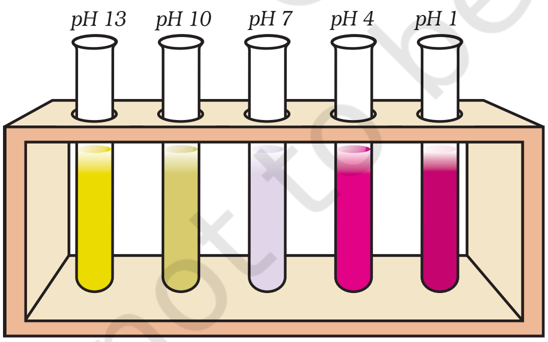
Aim

To determine the pH of some fruit juices.

Theory

Several dyes show different colours at different pH. These act as acid-base indicators. Solution of a mixture of dyes can be used to obtain approximate pH value of a solution. A solution of a mixture of dyes can be obtained to measure pH values from zero to 14. It is called universal indicator. Some universal indicators can measure the pH change of even 0.5. In fact, dyes themselves are weak acids or bases. Colour change occurs as a result of change in the structure of dye due to acceptance or release of protons. Different forms of a dye have different colours and hence, colour change is observed when pH of the solution changes. A standard chart for the colour change of the universal indicator with pH is supplied with the indicator paper or solution and the comparison of observed colour change with the chart provides a good estimate of the pH of the solution.

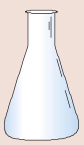
Natural pH Indicators



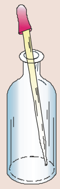
Red cabbage juice has vast pH range. It is a universal indicator of pH in aqueous solution.

The colour of these hydrangeas depends on the pH of the soil in which they grow, If pH of soil is acidic, flowers are blue and in alkaline pH, flowers are pink.

Material Required



- Beakers (100 mL) : Four
- Glass droppers : Four
- Test tubes : Four
- pH chart : One



- Fruit juice : Lemon, orange, apple, pineapple
- pH papers/universal indicator solution : As per need

Procedure

(i) Procure fresh juices of lemon, orange, apple and pineapple in separate beakers of 100 mL capacity each.

(ii) Transfer nearly 2 mL of the fresh juice (20 drops) with the help of a separate dropper for each juice in four different test tubes marked 1, 2, 3 and 4 respectively.

(iii) Add two drops of the universal indicator in each test tube and mix the content of each test tube thoroughly by shaking.

(iv) Match the colour appearing in each test tube with the standard pH chart.

(v) Record your observations in Table 5.1.

(vi) Repeat the experiment using pH papers to ascertain the pH of different juices and match the colour in each case with the one obtained with universal indicator.

(vii) Arrange the pH value of the four juices in increasing order.

Table 5.1 : pH value of different fruit juices

Name of the Juice	Colour with universal indicator	pH	Inference
Lemon			
Orange			
Apple			
Pineapple			

Result

Increasing order of pH value of juices is ____________.

Precautions

(a) Add equal number of drops of universal indicator to equal volumes of solutions in each of the test tubes.
(b) Match the colour of the solution with pH chart carefully.
(c) Store pH papers at a safe place to avoid contact with acidic and basic reagents kept in the laboratory.
(d) Use only fresh juice for the experiment.

Discussion Questions

(i) Out of the four juices, which one is least acidic? Explain.
(ii) If we dilute each of the juices, what effect is likely to be observed on the pH values?
(iii) On mixing any two juices, would the pH alter or remain the same? Verify your answer experimentally.
(iv) How can you ascertain the pH of a soft drink ?

EXPERIMENT 5.2

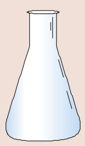
Aim

To observe the variation in pH of acid/base with dilution.

Theory

Hydrogen ion concentration per unit volume decreases on dilution. Therefore, change in pH is expected on dilution of the solution.

Material Required



• Boiling tubes	:	Eight
• Glass droppers	:	Four
• Test tubes	:	As per need

• 0.1 M HCl solution	:	20mL
• 0.1 M NaOH solution	:	20mL
• 0.05 M H_2SO_4 solution	:	20mL
• pH paper/universal indicator	:	As per need

Procedure

(i) Take four boiling tubes and mark them as A, B, C and D. (Fig. 5.1).
(ii) Take 2mL of 0.1 M HCl in boiling tube A.

(iii) Take 2mL of 0.1 M HCl in boiling tube B and add 18 mL water to it and mix thoroughly.

(iv) Take 5mL of dilute HCl solution from boiling tube B in boiling tube C and add 15mL water to it.

NaOH 

H_2SO_4 

HCl 

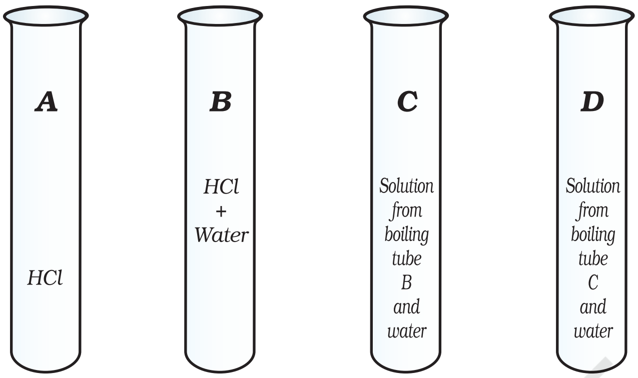


Fig. 5.1 : *Set up for experiment 5.2*

Hazard Warning

- *Never add water to the acid.*
- *For dilution add acid slowly into water.*

(v) Take 5mL of diluted HCl from boiling tube C in boiling tube D and add 15 mL water to it.

(vi) Cut a pH paper into small pieces and spread these on a clean glazed tile.

(vii) Take out some solution from boiling tube A with the help of a dropper and pour one drop on one of the pieces of pH paper kept on the glazed tile. Compare the colour of the pH paper with the standard chart.

(viii) Similarly test the pH of solutions of boiling tubes B, C, and D respectively and record your results as in Table 5.2.

(ix) Calculate the hydrogen ion concentration of solution B, C and D.

(x) Take out 1mL of solution from each boiling tube and transfer in separate test tubes. Add 2 drops of universal indicator to each of these test tubes. Shake the test tubes well and match the colour of these solutions with the standard pH chart to estimate the pH.

(xi) Similarly observe the change in pH of 0.05 M H_2SO_4 and 0.1M NaOH solution with dilution as detailed in steps (i) to (ix) above.

(xii) Record your observations in Table 5.2.

(xiii) Compare the result obtained by using universal indicator paper and that obtained by using universal indicator solution.

Table 5.2 : pH change on dilution

Boiling tube	HCl		H_2SO_4		NaOH	
	Colour	pH	Colour	pH	Colour	pH
A						
B						
C						
D						

Result

(i) Concentration of solutions of test tube B, C and D are ____________.

(ii) Write your conclusion about the variation of pH with dilution.

Precautions

(a) Add equal number of drops of the universal indicator to equal amounts of solution in each of the boiling tubes.

(b) Match the colour of the solution with pH chart carefully.

Discussion Questions

(i) What trend is observed in the variation of pH with dilution for acidic as well as for basic solutions?

(ii) How do you explain the results of variation in pH with dilution?

(iii) If any two acidic solutions (say A and C) are mixed, what would happen to the pH of the mixture? Verify your answer experimentally.

(iv) For each acidic solution, whether we use HCl or H_2SO_4, pH is same to a reasonably good extent, even though HCl is 0.1M, and H_2SO_4 is 0.05M. How do you explain this result?

(v) Will the pH of 0.1M acetic acid be the same as that of 0.1M hydrochloric acid? Verify your result and explain it?

EXPERIMENT 5.3

Aim

To study the variation in pH by common ion effect in case of weak acids and weak bases.

Theory

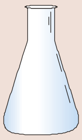
It is a known fact that the ionisation in the case of either a weak acid or a weak base is a reversible process. This can be represented as:

(1) $HA \rightleftharpoons H^+ + A^-$ (weak acid)

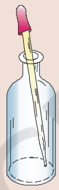
(2) $BOH \rightleftharpoons B^+ + OH^-$ (weak base)

The increase in concentration of A^- ions in case (1) and that of B^+ ions in case (2) would shift the equilibrium in the reverse direction thereby decreasing the concentration of H^+ ions and OH^- ions in cases (1) and (2) respectively so as to maintain the constancy of equilibrium constant *K*. This change either in H^+ ion concentration or OH^- ion concentration brings a change in the pH of the system, which can be judged either with the help of a pH paper or by using a universal indicator solution.

Material Required



• Beakers (100 mL)	:	Four
• Pipettes (25 mL)	:	Two
• Test tubes	:	Four
• pH chart	:	One



• Sodium ethanoate	:	2 g
• Ammonium chloride	:	2 g
• Ethanoic acid (1.0 M)	:	50 mL
• Ammonia solution (1.0 M)	:	50 mL
• pH paper and universal indicator	:	As per need

Procedure

(i) Take four 100 mL beakers and mark them as A, B, C and D.

(ii) Transfer 25 mL of 1M ethanoic acid in beaker 'A' and 25 mL of (1M) ammonia solution in beaker 'B'.

(iii) Similarly transfer 25 mL of (1 M) ethanoic acid in beaker 'C' and 25 mL of (1.0 M) ammonia solution in beaker 'D'. Now add 2 g sodium ethanoate in beaker 'C' and dissolve it. Likewise add 2 g of ammonium chloride in beaker 'D' and dissolve it by shaking the content of the beaker thoroughly.

(iv) Take approximately 2 mL (20 drops) of the solution from beakers A, B, C and D respectively into test tubes marked as 1, 2, 3 and 4.

(v) In each of the test tubes add 2 drops of universal indicator solution. Shake the content of the test tubes well and match the colour in each case with the standard pH chart.

(vi) Record your observations as given in Table 5.3.

(vii) Compare pH of the solution in test tubes 1 and 3 and record the change in pH.

(viii) Similarly compare pH of the solution in test tubes 2 and 4 and record the change in pH.

Ammonia solution 

Ethanoic acid 

Ammonium chloride 

Table 5.3 : Comparison of pH of acid/base and its buffer

Sl. No. of Test Tube	Composition of the system	Colour of pH paper	pH
1	CH_3COOH in water		
2	NH_4OH (NH_3 in water)		
3	$CH_3COOH + CH_3COONa$		
4	NH_4OH+NH_4Cl		

Result

(a) pH of acetic acid is _______

(b) pH of buffer of acetic acid and sodium acetate is ________ than acetic acid.

(c) pH of ammonia solution is _______

(d) pH of buffer of ammonia solution and ammonium chloride is ________ than the ammonia solution.

(e) Common ion effect ________ ionization of acid/base.

Precautions

(a) Try only weak acid/weak base and its salt for the study of the common ion effect.

(b) Handle the bottle of ammonium hydroxide with care.

(c) Add equal number of drops of the universal indicator in each of the test tubes.

(e) Store pH papers at a safe and dry place.

Discussion Questions

(i) The addition of sodium acetate to acetic acid increases the pH whereas, the addition of NH_4Cl to aqueous NH_3 solution (NH_4OH) decreases the pH of the system. How do you explain these observations?

(ii) Suggest suitable replacement for CH_3COONa for system 3 and NH_4Cl for system 4.

(iii) Suggest other pairs of weak acid and its salt and weak base and its salt to carry out the present investigations.

(iv) In salt analysis/mixture analysis, point out the situations where the variation in pH is carried out by common ion effect.

(v) How do buffer solutions resist change in the pH? Explain with a suitable example.

EXPERIMENT 5.4

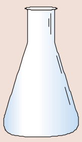
Aim

To study the change in pH during the titration of a strong acid with a strong base by using universal indicator.

Theory

It is assumed that strong acids and strong bases are completely dissociated in solution. During the process of neutralisation, H^+ ions obtained from the acid combine with the OH^- ions produced by base and form water. Therefore, when a solution of strong acid is added to a solution of strong base or vice versa, the pH of the solution changes. As the titration proceeds, initially there is slow change in the pH but in the vicinity of the equivalence point there is very rapid change in the pH of the solution.

Material Required



- Burette : One
- Beakers (250 mL) : Two
- Conical flask (100 mL) : One
- Dropper : One
- pH chart : One

- Hydrochloric acid (0.1 M) : 25 mL
- Sodium Hydroxide solution (0.1 M) : 50 mL
- Universal indicator : As per requirement

Procedure

(i) Take 25 mL hydrochloric acid solution (0.1 M) in 100 mL conical flask.

(ii) Add five drops of universal indicator solution to it.

(iii) Add (0.1 M) sodium hydroxide solution from the burette as given in Table 5.4.

(iv) Shake the content of the flask well after each addition of sodium hydroxide solution. Note the colour of the solution in the conical flask each time and find out the pH by comparing its colour with the pH chart.

(v) Note down your observations in Table 5.4.

(vi) Plot a graph of pH vs. total volume of NaOH added.

Hydrochloric acid

Sodium hydroxide

Table 5.4 : pH change during the neutralisation of 25 mL of HCl (0.1 M) with NaOH (0.1 M) solution

Sl. No.	Volume of NaOH added in lots (mL)	Total volume of NaOH added to the solution in flask (mL)	pH
1.	0	0	
2.	12.5	12.5	
3.	10.0	22.5	
4.	2.3	24.8	
5.	0.1	24.9	
6.	0.1	25.0	
7.	0.1	25.1	
8.	0.1	25.2	
9.	0.1	25.3	
10.	0.1	25.4	
11.	0.5	25.9	

Precautions

(a) To get good results perform the reaction with solutions of strong acid and strong base of same concentration.

(b) Handle the bottle of acid and base with care.

(c) Use small amount of indicator.

Result

Write down your result on the basis of data.

Discussion Questions

(i) What trend of pH change will you observe in the neutralisation of strong acid with strong base?

(ii) Do you expect the same trend of pH change for neutralisation of weak acid (acetic acid) with a strong base (sodium hydroxide)?

(iii) In which pH range should the indicator show colour change if the hydrochloric acid is to be neutralised by sodium hydroxide? Give answer after looking at the graph of the experiment.

(iv) Explain how does the study of pH change help in choosing the indicator for neutralisation reaction.

EXPERIMENT 5.5

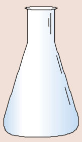
Aim

To study pH of solutions of sodium chloride, ferric chloride and sodium carbonate.

Theory

Salts of strong acid and strong base form neutral solutions while salts of weak acid/base and strong base/acid are basic and acidic respectively in nature. Salts of weak acid/base with strong base/acid are hydrolysed in water while salts formed by neutralization of strong acid and strong base do not hydrolyse in solution. You have already learnt about this in your chemistry textbook.

Material Required



- Boiling tubes : Three
- Test tubes : Three
- Glass droppers : Three

- pH paper/universal indicator : As per need
- 0.1 M NaCl solution : As per need
- 0.1 M $FeCl_3$ solution : As per need
- 0.1 M Na_2CO_3 solution : As per need

Procedure

(i) Take three boiling tubes and mark them as A, B and C.

(ii) Take 20 mL of 0.1 M solution(s) of NaCl, $FeCl_3$ and Na_2CO_3 in boiling tubes A, B and C respectively.

(iii) Cut the pH paper in small pieces and spread the pieces on a clean glazed tile.

(iv) Test the pH of the solution in boiling tubes A, B and C as in the experiment 5.1.

(v) Arrange three clean test tubes in a test tube stand.

(vi) Number the test tubes as 1, 2 and 3.

(vii) Pour 4 mL solution from boiling tube A in each of the test tubes.

(viii) Add 5 mL, 10 mL and 15 mL water in the test tubes 1, 2 and 3 respectively.

(ix) Note the pH of the solutions of test tubes 1, 2 and 3 with the help of pH paper and universal indicator.

(x) Repeat the experiment with the solutions of boiling tubes B and C.

(xi) Note your results in tabular form as in Table 5.5.

Table 5.5 : pH of NaCl, $FeCl_3$ and Na_2CO_3 solutions of different concentrations

Solution	pH of solution		
	Test tube 1	Test tube 2	Test tube 3
NaCl			
$FeCl_3$			
Na_2CO_3			

Result

Write down the result on the basis of your observations.

Precautions

(a) Use freshly prepared solutions.
(b) Do not leave the bottles open after taking out salt.
(c) Use separate and clean test tube for each solution.
(d) Store the pH paper at a safe and dry place.

Discussion Questions

(i) Why are $FeCl_3$ and Na_2CO_3 solutions not neutral?

(ii) Why are the salts of strong acid and strong base not hydrolysed? Explain.

(iii) How is the phenomenon of hydrolysis useful in salt analysis?

(iv) What is the effect of dilution on pH of salt solution? Verify and explain your results.