EXPERIMENT - 4

(a) To determine the pH of the following substances by using a universal indicator solution or pH paper.

(i) Salt solutions

(ii) Acids and bases of different dilutions

(iii) Vegetables and fruit juices

(b) To study the pH change by common ion effect in case of weak acids and weak bases using universal indicator solution or pH paper.

4.1 OBJECTIVES

After performing this experiment, you should be able to :

- *measure the pH of solutions of different substances by using universal indicator solution or pH paper ;*
- *classify various substances as neutral, acidic or alkaline ;*
- *explain the nature of substances on the basis of their pH values ;*
- *observe changes in pH of a weak acid when a compound containing common anion with the weak acid is added to it ;*
- *observe change in pH of weak base when a compound containing common cation with the weak base is added to it ; and*
- *explain the significance of buffer solution.*

4.2 WHAT YOU SHOULD KNOW

Solutions of acids and bases in water have hydronium ions $[H_3O^+]$ and hydroxyl ions $[OH^-]$ in different concentrations. Acidic solutions have $[H_3O^+]$ greater than 10^{-7} mol l^{-1} and alkaline solutions have $[OH^-]$ greater than 10^{-7} mol l^{-1} at 25°C. A neutral solution, or pure water, has $[H_3O^+] = [OH^-] = 10^{-7}$ mol l^{-1}. In expressing the concentration of H_3O^+ in a solution, the negative powers of 10 are involved. This inconvenient use of figures involving negative powers of 10 can be avoided by using a concept of pH scale. The pH of a solution is defined as the negative logarithm to the base 10 of the hydronium ion concentration.

$$pH = -\log_{10}[H_3O^+]$$

The pH of acidic solutions is less than 7, that of basic solution, is greater than 7 and the neutral solutions have pH = 7.

pH of a basic solution can be calculated by using the relation, $pH = 14 - \log_{10} [OH^-]$.

pH is accurately measured by a pH meter, but a reasonably good estimate can be made with the help of universal indicator solution or pH paper. These have characteristic colours depending upon the pH of the solution.

Now you would like to see the pH change of a solution of weak acid or weak base, when their salt is added to the solution. Weak acids and bases do not dissociate completely. An equilibrium exists between the dissociated and undissociated molecules. For example, for a weak acid HA, you can write the equilibrium as follows,

$$HA \rightleftharpoons H^+ + A^-$$

$$K_a = \frac{[H^+][A^-]}{[HA]}$$

$$NaA \longrightarrow Na^+ + A^-$$

This equilibrium is characterised by a constant, K_a. Addition of a salt containing a common ion, say NaA to its solution causes the equilibrium to shift to the left (Le Chattelier's principle). As a consequence, the concentration of hydrogen ions gets reduced and the pH gets increased. You would study this phenomenon (common ion effect) by using acetic acid as the weak acid and sodium acetate as its salt. Similar arguments hold true for the ionization of weak base (e.g. NH_4OH). You can check this by adding NH_4Cl to NH_4OH solution (a weak base and its salt). Such mixtures of weak acid or weak bases with their respective salts are called as Buffer Solution. pH of an acidic buffer can be calculated by using the following equation

$$pH = pK_a + \log \frac{[Salt]}{[Acid]}$$

K_a is the acid dissociation constant and pK_a is the negative logarithm of K_a.

pH for a basic buffer (a weak base and its salt) can be calculated by using the following equation:

$$pH = 14 - pK_b - \log \frac{[Salt]}{[Base]}$$

K_b is the base dissociation constant and pK_b is the negative logarithm of K_b.

pK_a for acetic acid at 25° C = 4.76

pK_b for ammonium hydroxide at 25° C = 4.75

4.3 MATERIALS REQUIRED

(1) Apparatus	(2) Chemicals
Test tubes – 6, Test tube stand, Spatula, Measuring cylinder (25ml) and a Dropper.	Dilute acid and base solutions, Neutral solution (NaCl), Solution of hydrolysable salt like $CuSO_4$, NH_4Cl, CH_3COONa, fruit and vegetable juices,

pH paper or universal indicator solution with colour chart, Acetic acid (O.1M), and Sodium acetate

4.4 HOW TO PERFORM THE EXPERIMENT

(a) For determining the pH of a given solution proceed as follows :

Take six test tubes and label them 1 to 6. If you are using universal indicator solution, then take 3 - 4 ml each of the test solution (about one fourth of test tube) into separate labelled test tubes. To all the test tubes, add 4 to 5 drops of the universal indicator solution and observe the appearance of colour, if any. If you are using pH paper, add 1– drop each of the test solutions with the help of a dropper on a pH paper strip. (Alternatively, you may dip the pH paper strips into each of the test solution). Observe the colour. Compare the colour of the solution in the test tubes or on the pH papers with the standard colour chart as provided to you to find the pH. Record your observations in Table 4.1 given in section 4.6. Classify the solutions as neutral, acidic or alkaline.

(b) To study the pH change due to common ion effect proceed as follows :

Take 5 labelled test tubes and put them in a test tube stand. Take 5 ml of given acetic acid solution in each of the labelled test tubes. Add 0.5 g, 1.0 g, 1.5 g and 2.0 g of sodium acetate to test tube No. 2, 3, 4 and 5 respectively and dissolve to get clear solutions. Add 4-5 drops of the universal indicator solution to each of the test tubes. In case of pH paper, add 1-2 drops of the solutions to pH papers separately with the help of separate dropper. (Alternatively, you may dip the pH paper strips into each of the test solution). Compare the colours in test solution or on pH paper developed with the colour chart and record the corresponding pH values, in Table 4.2 of section 4.6. Calculate the pH of each buffer solution you have prepared using the appropriate equation and list these in the table 4.2.

You may try the same procedure given above at (b) to study the pH in case of basic buffer using ammonium hydroxide as a base and ammonium chloride as its salt.

4.5 PRECAUTIONS

1. Compare the colour which appears on pH paper with the standard colour chart carefully.
2. Don't add too much of the indicator solution. You may follow the instruction given on the label of the indicator bottle.

4.6 OBSERVATIONS

Table 4.1 : pH values of different test solutions

S. No.	Solution	pH value
1.		
2.		
3.		
4.		
5.		
6.		

Table 4.2 : pH values of different combinations of CH_3COOH and CH_3COONa

S.No.	Vol. of acetic acid (ml)	Amount of sodium acetate (g)	pH Observed	Calculated pH
1.	5.0	0.0		
2.	5.0	0.5		
3.	5.0	1.0		
4.	5.0	1.5		
5.	5.0	2.0		

4.7 CONCLUSIONS

(a) pH of solution 1 is ....and therefore, it is acidic/basic/neutral

pH of solution 2 is ....and therefore, it is acidic / basic / neutral

pH of solution 3 is ....and therefore, it is acidic / basic neutral

pH of solution 4 is ....and therefore, it is acidic / basic / neutral

pH of solution 5 is ....and therefore it is acidic / basic / neutral

pH of solution 6 is ....and therefore, it is acidic/ basic/ neutral

(b) pH of acetic acid solution increases / decreases / remains unchanged on adding incresing amounts of sodium acetate to it.

4.8 CHECK YOUR UNDERSTANDING

1. *Calculate the pH of 10^{-2} M HCl solution and 10^{-2} M NaOH solution?*

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2. *The pH of an aqueous sodium chloride solution is 7. Explain, What does it mean?*

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3. *Why does copper sulphate solution (a salt solution) has a pH other than 7 ?*

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4.9 NOTE FOR THE TEACHERS

For acetic acid– sodium acetate experiment, dilute 1 ml of glacial acetic acid with water to get 100 ml solution in a measuring cylinder.

4.10 CHECK YOUR ANSWERS

Ans.1. pH = 2 for 10^{-2} M HCl

pH = 12 for 10^{-2} M NaOH

Ans.2. NaCl is a salt of strong acid (HCL) and a strong base (NaOH). In solution it dissociates completely to give sodium and chloride ions which do not hydrolyse. In aqueous solutions the hydrogen ions come only from water and have a concentration of 10^{-7} M, hence the pH=7.

Ans.3. $CuSO_4$ solution will have a pH which is less than 7, because in water solution, Cu^{2+} ion will hydrolyse to give strong acid (H_2SO_4) which makes the solution acidic in nature.