Investigation: Use of a pH meter and universal indicator in measuring pH

Please note

- A full risk assessment should be carried out prior to commencing this experiment.
- Personal safety equipment should be worn.
- Chemicals should be disposed of safely and with due regard to any environmental considerations.

Aim

- To learn how to measure pH using pH meters and universal indicator paper.
- To compare the strengths of different acids by measuring their pH at the same concentration.
- To examine the effect of dilution on the pH of ethanoic acid, CH_3COOH.

Introduction

The pH scale is widely used to describe acid and base properties. It is a measure of the hydrogen ion concentration in solution, expressed as:

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

The hydrogen ion concentration in the solution of an acid depends on both:

- the strength of the acid (which is a measure of the extent of its dissociation) and
- the concentration of the acid (which is a measure of the ratio of moles of solute to volume of solution).

This investigation should help you to clarify the difference between the two pairs of opposite terms used to describe acids and bases:

- strong versus weak and
- concentrated versus dilute

We will measure pH using both pH meters and universal indicator paper. pH meters need to be calibrated before each use, using buffers that span the pH range of study.

Universal indicator paper is paper that has been soaked in a mixture of acid/base indicators, which give a distinct range of colours across the pH range 1–14. The colour observed can be converted to pH by reading off a chart, usually provided with the universal indicator paper.

Before starting these experiments make sure that you understand the principles and practice of calibration of the pH meter using two different buffer solutions. You must use this process before and between the two experiments. Please treat the electrode with great respect and make sure it is *never left dry*!

Pre-lab questions

1 Explain the difference between:

a *strong* acid and a *concentrated* acid
a *weak* acid and a *dilute* acid.

2 Explain why it is important to keep the temperature constant when comparing the pH of different solutions.

Part 1

Method

1 Prepare 50 cm^3 of 0.010 mol dm^{-3} solutions of the following acids from the given solids:

sodium dihydrogenphosphate, $NaH_2PO_4.H_2O$

2 Prepare 50 cm^3 of 0.010 mol dm^{-3} solutions of the following acids from the given 0.10 mol dm^{-3} solutions:

ethanoic acid, CH_3COOH
hydrochloric acid, HCl

3 Measure the pH of each solution at a recorded temperature using the pH meter and universal indicator paper.

Results

0.010 mol dm^{-3} acid	pH from pH meter	pH from universal indicator paper
sodium dihydrogenphosphate, $NaH_2PO_4.H_2O$		
ethanoic acid, CH_3COOH		
hydrochloric acid, HCl		

Conclusion

- List three acids in order of decreasing acid strength.
- Are the results using the pH meter and universal indicator paper consistent with each other?
- Comment on the advantages and disadvantages of the two methods of measuring pH.

Part 2

1 Starting with the solution of ethanoic acid of 0.10 mol dm^{-3}, prepare a serial dilution of 100 cm^3 ethanoic acid at each of the following concentrations:

0.010 mol dm^{-3}
0.0010 mol dm^{-3}
0.00010 mol dm^{-3}

2 Use the pH meter to measure the pH of each of these solutions. Consider carefully the order in which you are going to take the readings and any other precautions to improve the accuracy. Record your results in the table.

Results

Concentration of acid / mol dm^{-3}	pH of solutions of ethanoic acid
0.00010	
0.0010	
0.010	
0.10	

Conclusion

- Comment on the effect of dilution on the pH of ethanoic acid.

For consideration

1 Compare the pH of ethanoic acid at each concentration with the *calculated* pH of a solution of hydrochloric acid of the same concentration at 298 K.

2 What does this imply about the extent of dissociation in the two acids?

3 Sketch a graph to show the *difference* between the pH of the two acids as their concentration decreases.

4 What does this indicate about the effect of dilution on dissociation of acids? Can you explain this observation in terms of equilibrium theory?

Equipment list

Chemicals / materials

sodium dihydrogenphosphate, $NaH_2PO_4.H_2O$
0.100 mol dm^{-3} ethanoic acid, CH_3COOH
0.100 mol dm^{-3} hydrochloric acid, HCl

Apparatus (per group of students)

balance
10 cm^3 pipette and filler
4 × 100 cm^3 volumetric flasks
pH meter
universal indicator paper and colour identification chart

Appendix:

Using a pipette to dilute 0.1 M sodium hydroxide to 0.01 M sodium hydroxide

Pipettes are indispensable for dealing accurately with small quantities of liquids.

To transfer 10ml of the above **sodium chloride** solution with a 10ml pipette, take the following step:

1. Obtain a suitable amount of the liquid to be measured in a beaker. Pipetting directly from the reagent bottle risks contaminating the entire contents of the bottle.
2. Use a pipette suction bulb to draw a small amount of the liquid into the pipette and use it to rinse the inside of the pipette thoroughly. Discard the rinse liquid in the sink.
3. Use a pipette suction bulb to draw liquid from the beaker until the meniscus is well above the index line on the pipette.

   Detailed steps:

   Remove the pipette suction bulb, and quickly place your index finger over the top end of the pipette. Place the rinsed pipette into the stock solution in the 50ml beaker and draw up the solution until it is past the level marked on the pipette. Remove the pipette from the stock solution. Suspend the pipette within the confines of a beaker. Allow the solution to escape from the pipette drop by drop until the bottom of the meniscus (when viewed at eye level) sits just on the level marked on the pipette. Wipe off the lower stem and tip of pipette using a small piece of filter paper.
4. Transfer 10 ml 0.1mol/L of sodium chloride solution into a 100ml volumetric flask.

   Detailed steps:

   Position the pipette so that the tip of the pipette meets the inside neck of the clean 100ml volumetric flask at an angle. Allow the solution in the pipette to escape and run down the neck of the volumetric flask into the body of the flask.

   (Pipettes are usually designed to deliver a known volume, this means that the small drop of solution remaining in the tip of the pipette is supposed to be there, so don't try to blow it out.)
5. Carefully fill the volumetric flask with distilled water until the bottom of the meniscus just touches the index line. Use an eyedropper to add the last ml or so of distilled water.

   Stopper the volumetric flask and invert it repeatedly until the solution is thoroughly mixed.
6. Clean the pipette inside and out with a thin stream of tap water.