# Measuring the pH of solutions

In this activity, you will make a number of solutions of acids and an alkali and measure their pH. Before you take the measurements, you should calculate the expected pH.

Finally, you will make mixtures of acids and alkalis and test the effect on pH.

## Health and safety

Wear eye protection.

0.05 mol dm–3 NaOH(aq) causes skin irritation and serious eye irritation.

## Procedure

### Making up solutions:

1. To make 0.01 mol dm–3 and 0.001 mol dm–3 solutions of HC*l*(aq):  
   - Transfer 1 cm3 of the 0.1 mol dm–3 solution of HC*l*(aq) from Tube 1 into a measuring cylinder and make up to 10 cm3 with distilled water.  
   - Transfer this solution to a boiling tube and label this tube ‘2’.  
   - Dilute 1 cm3 of the solution from Tube 2 in the same way. Transfer the resulting solution to a boiling tube and label this tube ‘3’.
2. Repeat the steps above for the NaOH(aq) solution in Tube 4. This will give you a 0.005 mol dm–3 solution in Tube 5, and a 0.0005 mol dm–3 solution in Tube 6.
3. Repeat the steps for the CH3COOH solution in Tube 7. This will give you a 0.1 mol dm–3 solution in Tube 8, and a 0.001 mol dm–3 solution in Tube 9.

### Calculating pH:

1. Calculate the expected pH for the solutions 1–9. Remember:   
   - pH = –log[H+]  
   - for an alkali: [H+] = 1 × 10–14/[OH–]  
   - for a weak acid: [H+] = √(*K*a × [acid])  
   - *K*a for ethanoic acid = 2 × 10–5Record your calculated pH values in the table at the end of this worksheet.
2. What two assumptions are you making in the calculations for ethanoic acid?

### Measuring the pH:

1. Now measure the pH of Tubes 1–9, and record the values in the table at the end of this worksheet.  
   How close are your measured values to the calculated pH values?   
   Which are more likely to be correct, the calculated values or the measured values?  
   What could have caused errors?
2. What do you notice about the way the pH of strong acids and alkalis changes with dilution compared with the pH of the weak acid when it is diluted? Can you explain the difference?

### Mixing solutions:

1. Mix the contents of Tube 2 with those of Tube 5.   
   Predict the pH before you measure it. To do this, calculate the amount (in mol) of acid in Tube 2, and the amount of alkali in Tube 5. Then work out how much acid is left after neutralisation. Finally, you can work out the concentration in 18 cm3 of solution.
2. Mix the contents of Tube 4 with those of Tube 7.  
   The sodium hydroxide should neutralise half the ethanoic acid, so you will have a solution containing ethanoic acid (a weak acid) with sodium ethanoate. This will behave as a buffer solution.  
   Calculate the amount of sodium ethanoate (salt) and ethanoic acid (acid) in the 18 cm3 of solution. Then calculate the pH using [H+] = *K*a[acid]/[salt].
3. Finally, test your buffer solution to see how well it can resist changes to pH compared to   
   0.001 mol dm–3 hydrochloric acid (Tube 3) (pH 3).  
   Add a few drops of universal indicator to both tubes.  
   Add 0.05 mol dm–3 sodium hydroxide dropwise to each tube. Record how many drops it takes for the pH to rise above 7.

### Results table

|  |  |  |  |  |
| --- | --- | --- | --- | --- |
| **Tube** | **Contents** | **Concentration / mol dm–3** | **Calculated pH** | **Measured pH** |
| 1 | Hydrochloric acid | 0.1 |  |  |
| 2 | Hydrochloric acid | 0.01 |  |  |
| 3 | Hydrochloric acid | 0.001 |  |  |
| 4 | Sodium hydroxide | 0.05 |  |  |
| 5 | Sodium hydroxide | 0.005 |  |  |
| 6 | Sodium hydroxide | 0.0005 |  |  |
| 7 | Ethanoic acid | 0.1 |  |  |
| 8 | Ethanoic acid | 0.01 |  |  |
| 9 | Ethanoic acid | 0.001 |  |  |
| 2+5 | Hydrochloric acid |  |  |  |
| 7+4 | Ethanoic acid + sodium ethanoate | |  |  |
| **Tube** | **Drops of 0.05 mol dm–3 sodium hydroxide needed to raise pH above 7** | | | |
| 7+4 |  | | | |
| 3 |  | | | |