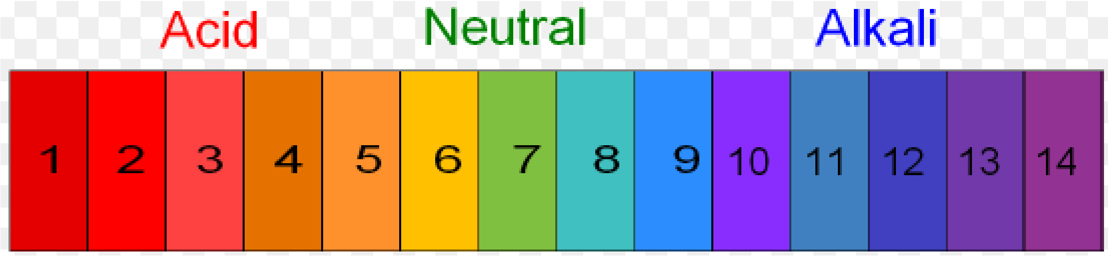
UNIT 5

ACIDS, BASES AND SALTS

PART 2 – TITRATIONS, INDICATORS AND THE PH SCALE



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Key words: acidic, alkaline, neutral, pH, indicator, litmus, phenolphthalein, methyl orange, universal indicator, titration, equivalence point, initial volume, final volume, titre volume, concordancy

Units which must be completed before this unit can be attempted:
Unit 1 – Atomic Structure and the Periodic Table
Unit 2 – Particles, Structure and Bonding
Unit 3 – Amount of Substance
Unit 4 – Introduction to Physical Chemistry

1) The pH scale

Water dissociates very slightly to produce H^+ and OH^- ions: $H_2O \rightleftharpoons H^+ + OH^-$

As a result, all aqueous solutions contain both H^+ and OH^- ions.

Solutions which contain acids contain many more H^+ ions than pure water. As a result, these solutions also contain fewer OH^- ions than pure water, because the extra H^+ ions push the above equilibrium to the left. The higher the concentration of H^+ ions, the lower the concentration of OH^- ions.

Solutions which contain alkalis contain many more OH^- ions than pure water. As a result, these solutions also contain fewer H^+ ions than pure water, because the extra OH^- ions push the above equilibrium to the left. The higher the concentration of OH^- ions, the lower the concentration of H^+ ions.

Any solution in which the concentration of H^+ ions is greater than the concentration of OH^- ions is said to be an **acidic** solution.

Any solution in which the concentration of OH^- ions is greater than the concentration of H^+ ions is said to be an **alkaline** solution.

Any solution in which the concentration of H^+ ions is equal to the concentration of OH^- ions is said to be a **neutral** solution.

The level of acidity or alkalinity of a solution (ie the relative concentrations of H^+ and OH^- ions is measured on a scale called the **pH scale**.

The lower the pH, the more acidic the solution (the higher the concentration of H^+ ions and the lower the concentration of OH^- ions). The most acidic solutions possible have a pH of around -1.

The higher the pH, the more alkaline the solution (the higher the concentration of OH^- ions and the lower the concentration of H^+ ions). The most alkaline solutions possible have a pH of around 15.

Neutral solutions (in which the concentration of H^+ ions is equal to the concentration of OH^- ions) have a pH of 7.

pH	-1	1	3	5	7	9	11	13	15
acidity	strongly acidic		slightly acidic		Neutral	slightly alkaline		Strongly alkaline	
$[H^+]$	very high		quite high		Normal	quite low		very low	
$[OH^-]$	very low		quite low		Normal	quite high		very high	

1 $moldm^{-3}$ HCl has a pH of 0; lemon juice has a pH of approx. 2; vinegar has a pH of approx. 3; rainwater has a pH of approx. 6; pure water has a pH of 7; household bleach has a pH of approx. 12; 1 $moldm^{-3}$ NaOH has a pH of 14.

The pH of a solution can be measured directly using a **pH meter**.

2) Indicators

(a) Theory of Indicators

Most indicators are weak acids in which the acid and its anion have different colours:

$$\underset{\text{Colour 1}}{HA} \rightleftharpoons H^+ + \underset{\text{Colour 2}}{A^-}$$

Some indicators are weak bases in which the base and its cation have different colours:

$$\underset{\text{Colour 1}}{BH^+} \rightleftharpoons H^+ + \underset{\text{Colour 2}}{B}$$

At low pH, the H^+ ions suppress the dissociation of the indicator and the indicator will appear as colour 1. At high pH, the $[H^+]$ is low so the indicator dissociates to produce more H^+ and the indicator will appear as colour 2.

Indicators therefore appear as one colour in low pH, and a different colour in high pH.

The pH range over which an indicator changes colour varies from indicator to indicator:

Indicator	Colour 1 (low pH)	Colour 2 (high pH)	pH of colour change
Methyl orange	Red	Yellow	3 – 5
Litmus	Red	Blue	5 – 8
Phenolphthalein	colourless	Pink	8 - 10

(b) Effect of acids and bases on indicators

The effect of acids and bases on indicators depends on the exact pH of the solution, and the pH at which the indicator changes colour:

Strongly acidic solutions with a pH of 3 or below will cause all three indicators to appear as Colour 1.
Weakly acidic solutions with a pH of 5 – 6 will cause methyl orange to appear as Colour 2 but will cause litmus and phenolphthalein to appear as colour 1.
Neutral solutions (pH = 7) will cause methyl orange to appear as Colour 2, litmus to appear as an intermediate colour (or whatever colour it was before) but will cause phenolphthalein to appear as colour 1.
Alkaline solutions with a pH of 10 or more will cause all three indicators to appear as Colour 2.

Single indicators, therefore, cannot give you the exact pH of a solution. They will simply tell you whether the pH is above or below a certain number.
The best indicator to use if you simply want to determine whether a solution is acidic, alkaline or neutral is LITMUS. This is because litmus changes colour at a pH of around 7. It is usually used soaked in paper, as **litmus paper**. Litmus paper can be red or blue:
Red litmus paper contains litmus in Colour 1 form.
Blue litmus paper contains litmus in Colour 2 form.

Acidic solutions will turn blue litmus paper red but will have no effect on red litmus paper.
Alkaline solutions will turn red litmus paper blue but will have no effect on blue litmus paper.
Neutral solutions will have no effect on either type of litmus paper.

(c) Use of indicators in simple chemical tests to classify substances as acids, bases or salts

(i) Acids:

There are two simple chemical tests for an acid in aqueous solution:

- It will turn blue litmus paper red
- It will fizz when sodium carbonate powder is added

If the substance is insoluble in water, it can be classified as an acid if it dissolves in an alkali.

(ii) Alkalis:

There is one simple chemical test for an alkali in aqueous solution:

- It will turn red litmus paper blue

If the substance is insoluble in water, it can be classified as a base if it dissolves in an acid.

(iii) Salts:

A substance can be classified as a salt if:

- An aqueous solution of the substance conducts electricity (showing that it an acid, base or salt), and
- It has no effect on either red or blue litmus paper (showing that it is not an acid or an alkali)

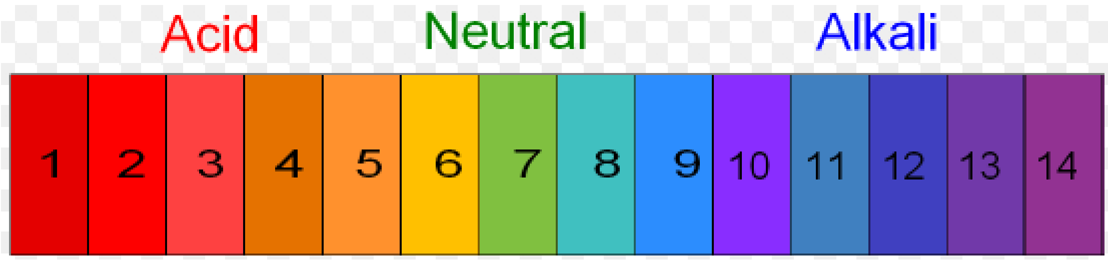
This test does not work on salts which undergo significant hydrolysis.

If the substance is insoluble, there is no simple test to classify it as a salt.

(d) Universal Indicator

Universal indicator is a mixture of several different indicators, each of which changes colour at a different pH. As a result, universal indicator appears at a different colour for very pH unit between 3 and 11.

At pH values of 3 and below, universal indicator appears red.
Between the pH values of 3 and 5, universal indicator appears orange.
Between the pH values of 5 and 6, universal indicator appears yellow.
Between the pH values of 6 and 8, universal indicator appears green.
Between the pH values of 8 and 9, universal indicator appears blue.
Between the pH values of 9 and 11, universal indicator appears indigo.
Between the pH values of 11 and above, universal indicator appears violet.



Universal indicator is most useful for finding out the approximate pH of aqueous solutions. It can be used in paper form or in solution form. In most cases, a colour chart is also provided and the colour of the indicator can be compared to the chart to deduce the pH of the solution. It is not an accurate method of determining pH, but is very simple and easy to use.

Using universal indicator to measure pH is an example of **colorimetry** – the use of colour to make measurements.

pH meters measure pH more accurately, but they are not as simple or convenient to use.

Practical: Determine the pH value of various solutions by colorimetry

Your teacher will have prepared for you five different solutions, labelled A, B, C, D and E, each with a different pH.

1) Using a dropping pipette, place around 1 cm^3 of solution A into a test tube
2) Add three drops of universal indicator
3) Compare the colour to the colour chart and hence determine the approximate pH of the solution
4) Repeat Steps 1 – 3 using solutions B, C, D and E

If you are using universal indicator paper, you will need around 10 cm^3 of each solution (use a measuring cylinder) and will need to use small beakers instead of test tubes. Dip the universal indicator paper into each solution, compare the colour to the colour chart and hence determine the approximate pH of the solution.

Practical: Determine the pH of different soil samples

1) Take 20 g of the soil to be tested and place it in a 100 cm^3 beaker.
2) Add 40 cm^3 of water to the beaker using a measuring cylinder.
3) Leave the beaker for around 30 minutes, stirring every few minutes.
4) If using universal indicator paper, dip the UI paper into the mixture, compare the colour to the colour chart and hence determine the approximate pH of the solution. If using UI solution, fold and insert a piece of filter paper into the funnel, and pour the mixture through the funnel into a boiling tube. When around 10 cm^3 of filtrate has been collected, add a few drops of UI solution, compare the colour to the colour chart and hence determine the approximate pH of the solution.
5) Repeat for any other soil samples you need to test.

3) Acid-Base Titrations

(a) Carrying out a titration

A **titration** is a quantitative experiment used to find out what exact volume of one solution is required to react completely with a fixed volume of another solution in a flask. It is the most common example of **volumetric analysis**. From the titration result, you can deduce the concentration of the alkali if the concentration of the acid is known, or vice versa.

The most common titrations are reactions between acids and alkalis. In most cases, acid is added gradually to a fixed volume of alkali.

During a titration, you need to know when the reaction is complete. In acid-base reactions, the best way to do this is to add a small quantity of an indicator. When the indicator changes colour, the reaction is complete.

(i) Choosing a suitable indicator

During titrations, the pH of the solution in the flask will change gradually until the exact amount of acid has been added to neutralise the alkali. This point is called the **equivalence point** of the titration. At the equivalence point, the pH changes sharply. For the indicator to work, it needs to change colour during the equivalence point of the titration.

The exact change in pH during the equivalence point of titrations depends on whether the acids and bases are strong or weak, as follows:

Type of titration	pH change at equivalence point
Strong acid - strong base	11 to 3
Weak acid - strong base	11 to 7
Strong acid - weak base	7 to 3

Weak acid – weak base reactions do not lead to a sudden pH change at the equivalence point and so are not suitable for analysis by titration.

The two indicators most widely used in titrations are methyl orange and phenolphthalein, which change colour over the following pH ranges:

Indicator	Colour 1 (low pH)	Colour 2 (high pH)	pH of colour change
Methyl orange	Red	Yellow	3 – 5
Phenolphthalein	Colourless	Pink	8 - 10

Note that:

- In strong acid - strong base titrations, the pH changes from 11 - 3 at the equivalence point; both methyl orange and phenolphthalein change colour within this pH range and so both indicators are suitable for use in these titrations.

- In weak acid - strong base titrations, the pH changes from 11 to 7 at the equivalence point; phenolphthalein changes colour within this range but methyl orange does not; **methyl orange is therefore not suitable for titrations involving weak acids** (only phenolphthalein is suitable)

- In strong acid - weak base titrations, the pH changes from 7 to 3 at the equivalence point; phenolphthalein does not change colour within this range but methyl orange does; **phenolphthalein is therefore not suitable for titrations involving weak bases** (only methyl orange is suitable)

(ii) Preparing a fixed volume of solution 1

To prepare accurately a fixed volume of one solution, the best method is to use a pipette. A **pipette** is a piece of glassware designed to deliver a fixed volume (usually 25.0 cm^3) from one container to another. In most titrations, the fixed volume of solution is the alkali, and it is usually measured out into a conical flask ready for the titration to take place.

- Pour around 100 cm^3 of the alkali into a beaker
- Use a pipette with a pipette filler to suck the alkali into the pipette until it is just above the mark; remove the pipette filler and place your thumb over the top of the pipette; release the pressure gradually until the base of the meniscus lies on the mark; then increase the thumb pressure to fix the solution at that level
- Still applying pressure with your thumb, move the pipette out of the alkali solution and over a clean, empty 250 cm^3 conical flask; remove your thumb and allow the pipette to empty itself naturally
- Once the solution has stopped draining out of the pipette, dip the tip of the pipette into the solution in the conical flask for a few seconds and then remove it
- Place the empty pipette to one side
- Add a few drops of your chosen indicator to the alkali. The indicator should appear as Colour 2 (the expected colour of the indicator in alkaline solution)

(iii) Adding solution 2 gradually to solution 1 until the reaction is complete

A **burette** is a piece of glassware designed to measure how much of a solution has been delivered. It does not contain a measurable volume of solution, it is possible to measure the volume in the burette before any solution has been added, and then the volume in the burette after the reaction is complete. The difference between these two volumes is the volume of solution delivered.

In most titrations, the acid is added to the burette; this is because acids, unlike alkalis, do not leave residue behind after evaporation and this makes the burette easier to clean.

- Use a boss to attach a clamp to a stand
- Clamp a burette into position
- Ensure that the burette is clean and that the tip of the burette is closed
- Using a funnel, add the acid solution gradually to the burette until it is around half-full
- Open the tap of the burette and allow acid to escape until the tip of the burette is completely full (there should be no air bubbles), then close the tip again
- Add more acid into the burette, using the funnel, until the level of solution in the burette is in between the 1 cm^3 and the 5 cm^3 mark; record this volume accurately as the “initial volume”
- Remove the funnel from the burette
- Place the conical flask containing 25 cm^3 of solution 1 and the indicator and place it below the burette; the rim of the conical flask should be just below the tip of the burette
- Open the tap of the burette and add the solution gradually into the conical flask, swirling the conical flask continuously; after around 10 cm^3 has been added, close the tip gradually until the solution is dripping slowly into the conical flask at the rate of one drop per second; keep swirling the mixture and continue until the indicator changes colour (into Colour 1); then close the tap
- Record the new volume of solution accurately as the “final volume”

(iv) Achieving concordancy

The initial and final volume of solution used should be recorded in a table similar to the one below. The **titre volume** is the difference between the initial and final volume.

	Titration 1	Titration 2	Titration 3	Titration 4	Titration 5
Final volume (cm^3)					
Initial volume (cm^3)					
Titre volume (cm^3)					
Concordant? (Y or N)	-				

The titration should then be repeated. The conical flask should be rinsed, a new fixed volume of solution 1 needs to be prepared and added to the conical flask, and a few drops of indicator should be added again.

The burette should be topped up until the initial volume is between 1 cm^3 and 5 cm^3 and the experiment should be repeated. The new titre volume should be recorded.

If the second titre volume is within 0.1 cm^3 of the first titre volume, the two values are said to be **concordant** and the experiment can be stopped. If not, the experiment needs to be repeated until two concordant values are obtained. The average of the concordant values is the official titre volume for the experiment.

(v) special techniques for good accuracy

Technique	Reason
Rinse the burette with acid before use	Ensure that any water or other substance left in the burette is washed out
Ensure the tap is full of liquid before use (with no air bubbles)	Ensure that the volume change is exactly the same as the volume of liquid delivered
Record all burette readings to 2 dp	Ensure that all readings are made to the nearest drop
Ensure the reading is taken from the base of the meniscus	Ensure that the correct volume delivered is calculated
Allow the pipette to empty under gravity	Ensure that the pipette delivers the correct volume (not too much)
Touch the surface of the liquid with the pipette tip after emptying	Ensure that the pipette delivers the correct volume (not too little)
Add only two drops of indicator	Because indicators are acidic and will affect the equivalence point
Add the acid dropwise at the end point	Ensure that the equivalence point is seen as it happens
Rinse the top of the conical flask with distilled water during the titration	Ensure that any acid or alkali is washed into the conical flask
Rinse the conical flask with distilled water after use	Ensure there are no moles of acid or alkali left behind

(b) Calculating concentrations and other quantities from titration results

Once the titre volume has been determined, other desired quantities can be determined using techniques from the amount of substance topic. The simplest quantity to determine is the concentration of the unknown solution (acid or base). Other quantities which can be determined from titration results include:

- The molar masses of acids and bases and water of crystallization
- The solubility of acids and bases;
- The percentage purity of acids and bases

Practical: Determine the concentration of a solution of NaOH by titration

1) Add 1.0 moldm^{-3} hydrochloric acid to a burette and record the initial volume
2) Use a pipette and filler to transfer 25.0 cm^3 of the sodium hydroxide solution into a conical flask
3) Add a few drops of methyl orange indicator (it should appear yellow)
4) Add the HCl solution to the NaOH solution until the indicator changes colour (from yellow to pink); record the final volume and deduce the titre volume
5) Repeat the titration until concordancy is achieved
6) Hence determine the concentration of the sodium hydroxide solution

Here is a worked example of a simple calculation based on a titration result:

Question:
In a titration 28.3 cm^3 of a 0.10 moldm^{-3} solution of NaOH was required to react with 25 cm^3 of a solution of H_2SO_4. What was the concentration of the H_2SO_4 solution?

Worked answer:
Equation: $H_2SO_4 + 2NaOH \rightarrow Na_2SO_4 + 2H_2O$
Moles of NaOH used in titration = 28.3/1000 x 0.1 = 2.8 x 10^{-3}
2:1 ratio so moles of H_2SO_4 present in conical flask = 2.8 x 10^{-3}/2 = 1.4 x 10^{-3}
so concentration of H_2SO_4 = 1.4 x 10^{-3}/25 x 1000 = 0.056 moldm^{-3}.

Practical: Determine the percentage purity of a sample of vinegar by titration

$$CH_3COOH + NaOH \rightarrow CH_3COONa + H_2O$$

1) A solution of vinegar is labelled as 62.3 gdm^{-3}. Pipette 25 cm^3 of this vinegar sample into a 250 cm^3 volumetric flask
2) Make the volume in the flask up to the mark, shaking continuously to ensure that the contents are thoroughly mixed
3) Transfer some of this diluted acid into a burette and record the initial volume
4) Pipette a 25.0 cm^3 sample of the sodium hydroxide solution into a conical flask and add a few drops of phenolphthalein indicator (the solution should turn pink)
5) Add the diluted vinegar solution to the NaOH solution until the indicator changes colour (from pink to colourless); record the final volume and deduce the titre volume
6) Repeat the titration until concordancy is achieved
7) Hence determine the concentration of the diluted vinegar solution
8) Hence determine the concentration of the original vinegar solution
9) Hence determine the percentage purity of the original vinegar solution
10) Explain why phenolphthalein was used as the indicator for this titration

Practical: Determine the relative formula mass, and hence water of crystallisation, of hydrated sodium carbonate, $Na_2CO_3.xH_2O$, by titration

$$\mathbf{Na_2CO_3 + 2HCl \rightarrow 2NaCl + CO_2 + H_2O}$$

1) Weigh out approximately 3.5 g of hydrated sodium carbonate into a weighing boat. Record the exact mass of the solid.
2) Dissolve the solid in approximately 100 cm^3 of distilled water.
3) Transfer this solution into a 250 cm^3 volumetric flask, and make the volume in the flask up to 250 cm^3 by using washings from the beaker; shake continuously to ensure that the contents are thoroughly mixed
4) Pipette a 25.0 cm^3 sample of this solution into a conical flask and add a few drops of methyl orange indicator (the solution should turn yellow)
5) Fill a burette with 0.1 $moldm^{-3}$ HCl and record the initial volume
6) Add the HCl solution to the sodium carbonate solution until the indicator changes colour (from yello to pink); record the final volume and deduce the titre volume
7) Repeat the titration until concordancy is achieved
8) Hence determine the moles of sodium carbonate present in the conical flask, and in the original volumetric flask
9) Hence determine the molar mass of the hydrated sodium carbonate sample
10) Hence calculate the value of x, given that the formula of sodium carbonate is $Na_2CO_3.10H_2O$

Test Your Progress: Topic 5 Exercise 4

1. Succinic acid has the formula $(CH_2)_n(COOH)_2$ and reacts with dilute sodium hydroxide as follows:
   $(CH_2)_n(COOH)_2 + 2NaOH \rightarrow (CH_2)_n(COONa)_2 + 2H_2O$
   2.0 g of succinic acid were dissolved in water and the solution made up to 250 cm^3. This solution was placed in a burette and 18.4 cm^3 was required to neutralise 25 cm^3 of 0.1 $moldm^{-3}$ NaOH. Deduce the molecular formula of the acid and hence the value of n.
2. Sodium carbonate exists in hydrated form, $Na_2CO_3.xH_2O$, in the solid state. 3.5 g of a sodium carbonate sample was dissolved in water and the volume made up to 250 cm^3. 25.0 cm^3 of this solution was titrated against 0.1 $moldm^{-3}$ HCl and 24.5 cm^3 of the acid were required. Calculate the value of x given the equation:
   $Na_2CO_3 + 2HCl \rightarrow 2NaCl + CO_2 + H_2O$
3. 25 cm^3 of a sample of vinegar (CH_3COOH) was pipetted into a volumetric flask and the volume was made up to 250 cm^3. This solution was placed in a burette and 13.9 cm^3 were required to neutralise 25 cm^3 of 0.1 $moldm^{-3}$ NaOH. Calculate the molarity of the original vinegar solution and its concentration in gdm^{-3}, given that it reacts with NaOH in a 1:1 ratio.
4. 2.5 g of a sample of impure ethanedioic acid, $H_2C_2O_4.2H_2O$, was dissolved in water and the solution made up to 250 cm^3. This solution was placed in a burette and 21.3 cm^3 were required to neutralise 25 cm^3 of 0.1 $moldm^{-3}$ NaOH. Given that ethanedioic acid reacts with NaOH in a 1:2 ratio, calculate the percentage purity of the sample.
5. A toilet cleaner containing sodium hydrogensulphate, $NaHSO_4$ is believed to have been contaminated. 5.678 g of the sample were dissolved in water and the solution was made up to 250 cm^3. This solution was placed in a burette and 23.1 cm^3 of it were required to neutralise 25 cm^3 of 0.1 $moldm^{-3}$ sodium hydroxide.
   Calculate the percentage purity of the sample given that the species react in a 1:1 ratio.