DEPARTMENT OF EDUCATION

GRADE 12

CHEMISTRY

MODULE 2

ACIDS, BASES AND SALTS

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GRADE 12

CHEMISTRY

MODULE 2

ACIDS, BASES AND SALTS

IN THIS MODULE YOU WILL LEARN ABOUT:

12.2.1: COMMON ACIDS AND BASES

12.2.2: STRONG AND WEAK ACIDS AND BASES

12.2.3: DISSOCAIATION CONSTANT AND pH CALCULATION

12.2.4: ACID-BASE TITRATION TO DETECT THE END POINT

Acknowledgement

We acknowledge the contributions of all secondary teachers who in one way or another have helped to develop this Course.

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DIANA TEIT AKIS
PRINCIPAL

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SECRETARY'S MESSAGE

Achieving a better future by individual students and their families, communities or the nation as a whole, depends on the kind of curriculum and the way it is delivered.

This course is a part of the new Flexible, Open and Distance Education curriculum. The learning outcomes are student-centred and allows for them to be demonstrated and assessed.

It maintains the rationale, goals, aims and principles of the national curriculum and identifies the knowledge, skills, attitudes and values that students should achieve.

This is a provision by Flexible, Open and Distance Education as an alternative pathway of formal education.

The course promotes Papua New Guinea values and beliefs which are found in our Constitution, Government Policies and Reports. It is developed in line with the National Education Plan (2005 -2014) and addresses an increase in the number of school leavers affected by the lack of access into secondary and higher educational institutions.

Flexible, Open and Distance Education curriculum is guided by the Department of Education's Mission which is fivefold:

- To facilitate and promote the integral development of every individual
- To develop and encourage an education system that satisfies the requirements of Papua New Guinea and its people
- To establish, preserve and improve standards of education throughout Papua New Guinea
- To make the benefits of such education available as widely as possible to all of the people
- To make the education accessible to the poor and physically, mentally and socially handicapped as well as to those who are educationally disadvantaged.

The college is enhanced through this course to provide alternative and comparable pathways for students and adults to complete their education through a one system, two pathways and same outcomes.

It is our vision that Papua New Guineans' harness all appropriate and affordable technologies to pursue this program.

I commend all the teachers, curriculum writers and instructional designers who have contributed towards the development of this course.

UKE KOMBRA, PhD
Secretary for Education

MODULE 2: ACIDS, BASES AND SALTS

INTRODUCTION

You have learnt about compounds in the previous module. In this module you will study important classes of compounds called **acids, bases and salts**. Some examples of common acids are acetic acid in vinegar and citric acid in lemon juices. The solution used for cleaning toilet bowls and tiles is hydrochloric acid. It is commonly called muriatic acid. An example of a base is sodium hydroxide used in making soaps and drain cleaners. Sodium hydroxide is also called lye or caustic soda which is usually bases. It is almost impossible to prepare a delicious meal without using common salt (sodium chloride, NaCl). The word **salt** describes any metal compounds that can be made from acids.

Learning outcomes

On successful completion of this module, students will:

- demonstrate an understanding of chemical properties of acids and bases.
- investigate the difference between strong and weak acids as well as strong and weak bases discussing their dissociation in water.
- discuss and explain dissociation constant (k_w) of water.
- calculate the pH and pOH of a given solution.
- conduct volumetric analysis (titration) experiments using the different indicators to detect the end point of the titrations.
- demonstrate an understanding of titration and pH curves by plotting them.

Time Frame

Suggested 8 weeks

This unit should be completed within 8 weeks.

If you set an average of three hours per day, you should complete the Unit by the end of the assigned week.

Try to do all the learning activities and compare your answers with the ones provided at the end of the unit. If you do not get a particular exercise right in the first attempt, you should not get discouraged, but instead go back and attempt it again. If you still do not get it right after several attempts then you should seek help from your friend or even your tutor.

DO NOT LEAVE ANY QUESTION UN-ANSWERED.

Terminologies

Before you get into the thick of things, let us make sure you know some of the terminologies that are used throughout this module.

Acid-base titrations	Are based on the neutralization reaction between the analyte and acidic or basic titrant**.**
Alkali	When a base is soluble in water (H_2O), and produces hydroxide ion, (OH^-).
A Lewis acid	Is a substance that can accept a pair of electrons to form a covalent bond.
A Lewis base	Is a substance that can donate a pair of electrons to form covalent bonds.
Amphiprotic	Are substances such as water, which can either accept or donate protons.
Base	Is any metal oxide, metal hydroxide and hydrogen carbonate.
Diprotic acid	Contains two ionisable protons.
Endpoint	**I**is the point at which the titration is complete, as determined by an indicator at the titration endpoint, the quantity of reactant in the titrant.
Hydrated salts	Salts that contain water of crystallization.
Mineral acid	(Inorganic acid) is an acid derived from one or more mineral or inorganic compounds.
Monoprotic acid	Contains one ionisable proton.
Neutralization	Is a process when acids react with bases to form salt and water in the process.
Tripotic acid	Contains three ionisable protons.

12.2.1 COMMON ACIDS, BASES AND SALTS

Mineral and Organic Acids

What are **mineral** and **organic acids**?
Mineral acid (inorganic acid) is an acid derived from one or more mineral or inorganic compounds. All mineral compounds form hydrogen ions when dissolve in water.

Organic acid is an **organic compound** with acidic properties.

The following are the common examples of mineral and organic acid.

Mineral acids	Organic acids
Hydrocholric acid (HCl)	Citric acid ($C_6H_8O_7$) - lemon juices
Sulphuric acid (H_2SO_4)	Gallic acid ($C_7H_6O_5$) - gallnuts and tea leave
Nitric acid (HNO_3)	Lactic acid ($C_3H_6O_3$) - sour milk and yoghurt

The word **acid** comes from the Latin term **acidus**, meaning **sour**. In the 17^{th} Century, a chemist, Robert Boyle, first classified certain substances as either acids or alkalis according to their characteristics.

Whenever you eat an orange or an apple, you are taking in some acids. Our homes are full of acids, bases, and salts. What is the acid found in the vinegar used for cooking sweet and sour fish? What is the acid that is found in your stomach?

However, some bases are corrosive and can harm your skin. When you wash your hands with soap or brush your teeth with toothpaste, you are using a base.

We use **salt** for cooking and for flavouring our food, but not all salts are edible. The salt that can be eaten is **sodium chloride.** There is another salt called **potassium chloride.** It can kill us if it is injected into our blood stream even in small amounts.

What is an acid?
Do you know why oranges and lemons taste sour? What is the common substance found in them that gives them the sour taste?

If juice of orange or lemon is tested with a piece of blue litmus paper, the litmus paper will turn **red.**

The sour taste and the changing of the blue litmus paper to red are just two common properties of acids.

All acids contain hydrogen but not all compounds that contain hydrogen are acid. For example, ammonia (NH_3) and methane (CH_4) contain hydrogen, but they are not acids because they do not produce hydrogen ions in water. All acids produce hydrogen ions, **H^+,** in aqueous solutions. This leads us to the definition of acids.

Acids are substances which produces hydrogen ions, H^+, when dissolves in water.

The table below shows the names of some common acids, their formula and ions produced in aqueous solution.

NAME OF ACID	FORMULA	IONS PRODUCE IN	AN AQUOEUS SOLUTION
Ethanoic acids (found in vinegar)	CH_3COOH	H^+	CH_3COO^-
Hydrochloric acid	HCL	H^+	Cl^-
Nitric acid	HNO_3	H^+	NO_3^-
Sulphuric acid	H_2SO_4	H^+	SO_4^-

Formula of common acids and their ions.

An acid is a substance which produces hydrogen ions in water. From the table above, will you will see how the hydrogen chloride gas dissolves in water, hydrochloric acid is formed.

A solution of hydrochloric acid will ionize as follows: **$HCl(aq) \longrightarrow H^+(aq) + Cl^-(aq)$**

Similarly, an aqueous solution of sulphuric acid will be the separation of ion as follows:

$$H_2SO_4(aq) \rightarrow 2H^+ aq) + SO_4^{2}\ (aq)$$

The table below are the examples of some common uses of acids.

Acids	Uses
Hydrochloric acid (HCl)	Is used in the car industry to remove rust from metals. The body of a motor car is dipped into a solution of hydrochloric acids to remove rust before it is painted.
Sulphuric acid (H_2SO_4)	To make fertilizers and detergents. It is also used in car batteries.
Nitric acid (HNO_3)	For making fertilisers and explosive called **trinitrotoluene (TNT).**
Ethanoic acid (CH_3COOH)	To add flavor to food and for preserving vegetables.

Common acids and their uses.

Now, check what you have just learnt by trying out the learning activity below!

Learning Activities 1 **20 minutes**

Answer the following questions:

1. Define the following acid:

    i. Mineral ______________________________

    ii. Organic ______________________________

2. Give two examples of:

    a. mineral acid
    (i) ______________________________
    (ii) ______________________________

    b. organic acid
    (i) ______________________________
    (ii) ______________________________

3. Write the chemical formula of the following compound.
    (i) Lactic acid
    ____

    (ii) Gallic acid
    ____

Thank you for completing your learning activity 1. Check your work. Answers are at the end of this module.

What is a base?
We have learnt that acids react with bases to form salt and water in the process called **neutralization.**

A **base** is any metal oxide, metal hydroxide and hydrogen carbonate. This means that a base contains either oxides ions, (O^{2-}) or hydroxide ions, (OH^{-}).also a base reacts with an acid to give salts and water only.

Example:
Sodium oxide (Na_2O), sodium hydroxide (NaOH), copper (II) oxides (Cu (II) O). Some bases are soluble in water while others are not. When a base is soluble in water (H_2O), and produces hydroxide ion, (OH^-). It is called an **alkali.**

Alkali is a soluble base that produces hydroxide ions, (OH^-) in water (H_2O).
On the other hand, copper (II) hydroxide, $Cu(OH)_2$ and iron (II) hydroxides, $Fe(OH)_2$are just bases because they are not soluble in water. Therefore, they are unable to produce hydroxide ions in water.

Some common alkalis and their chemical formula

NAME	CHEMICAL FORMULA
Sodium hydroxide	NaOH
Potassium hydroxide	KOH
Calcium hydroxide	$Ca(OH)_2$
Aqueous ammonia	$NH_3(aq)$

We can also define a **base** as a substance that reacts with an acid to give a salt and water only. The general equation for this reaction is:

$$\text{base} + \text{acid} \rightarrow \text{salt} + \text{water}$$

For example:

Copper (II) Oxide + Sulphuric acid → Copper Sulphate + water

$$CuO_{(s)} + H_2SO_{4(aq)} \longrightarrow CuSO_{4(aq)} + H_2O$$

The ionic equation for this reaction is:

$$CuO_{(s)} + 2H^+_{(aq)} \longrightarrow Cu^{2+}_{(aq)} + H_2O_{(l)}$$

In the above reactions, the oxide ions or the hydroxide ions from the bases react with hydrogen ion from the acids to form water.

Uses of alkalis
The table below shows examples of some uses of common alkalis**.**

Alkalis	Uses
Magnesium hydroxide $Mg(OH_2)$	Used in toothpaste to neutralize the acid produced by bacteria, which feed on the sugar in our food.
Antacid tablets	Contain magnesium hydroxide and carbonate, which can neutralise the excess acid in the stomach that causes indigestion.
Ammonia NH_3	Solution is used in window cleaners.
Calcium hydroxide $Ca(OH_2)$	Commonly called slaked lime, is used in agriculture to neutralise excess acids in soil.

Uses of bases and alkalis.

What are salts?

Your tears are salty, your perspiration tastes salty, and your blood tastes salty. This is not because you eat common salt or sodium chloride (NaCl) every day. Other compounds in your body are also known as salts and give that **salty taste**.

Iron sulphate (Fe (II) SO_4) is a salt that helps your blood carry oxygen. Sodium fluoride is a salt that stops your teeth from decaying.

All salts are ionic compounds. They are formed when metallic ion or ammonium ($NH4^+$) replaces one or more hydrogen ions of an acid.

For example:

Zinc sulphate is a salt. It is formed in the following reactions:

Zinc hydroxide + sulphuric acid ⟶ zinc sulphate + water

$$Zn(OH)_{2(s)} + H_2SO_{4(aq)} \longrightarrow ZnSO_{4(aq)} + H_2O_{(l)}$$

The table below shows some examples of salts made from acid base reactions.

Salt	Chemical formula	Possible Base	reactants acid
Calcium chloride	$CaCl_2$	CaO	HCl
Ammonium sulphate	$(NH_4)_2SO_4$	Aqueous NH_3	H_2SO_4
Copper (II) nitrate	$Cu(NO_3)_2$	$CuCO_3$	HNO_3

Salts made from acid-base reactions

Water of crystallization

Many salts combine with water molecules to form **crystals**. These water molecules are known as **water of crystallisation.**

Salts that contain water of crystallization are called **hydrated salts**. Salts that do not contain water of crystallization are called **anhydrous salts**. Anhydrous salts are powders.

The table below shows the formulae of some anhydrous salts and some hydrated salts. Compare the formulas of anhydrous and hydrated salts.

How do they differ?

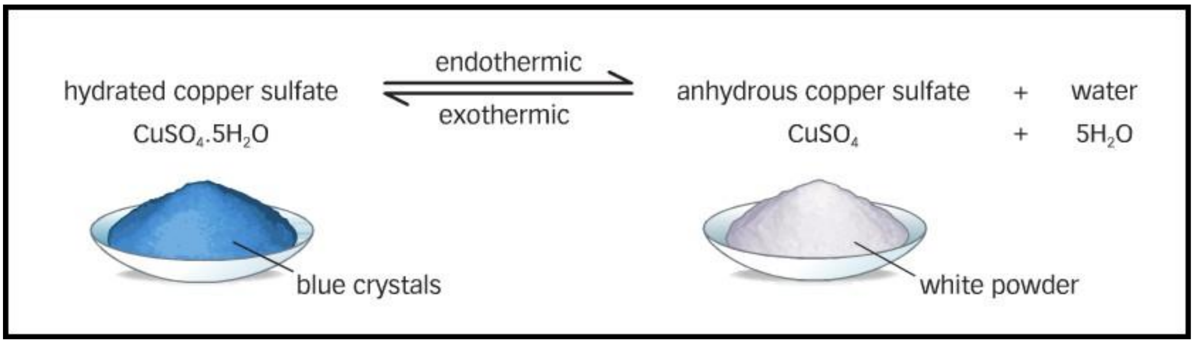
Name of salt	Formula of anhydrous salt	Formula of hydrated salt
Copper (II) sulphate	$CuSO_4$	$CuSO_4.5H_2O$
Magnesium sulphate	$MgSO_4$	$MgSO_4.7H_2O$
Sodium carbonate	Na_2CO_3	$Na_2CO_3.10H_2O$
Zinc sulphate	$ZnSO_4$	$ZnSO_4.7H_2O$

Examples of hydrated and unhydrated salt.

Note: The dot (.) In a formula means that if the substance is heated, everything after the dot will be given off first. For hydrated salts, water will be given off first.

How can water of crystallization be removed from a salt? When a hydrated salt is heated, water of crystallisation is given off.

NOTE: The water of crystallisation gives their colour and shape.



The table below shows the soluble and insoluble salts.

Soluble	Insoluble
All sodium, potassium and ammonium salts	
All nitrates	
Most bromides, chlorides and iodides	Bromides, chlorides and iodides of silver and lead
Most sulphates	Sulphates of barium, calcium and lead
Carbonates, hydroxides of sodium, potassium and ammonium	Most carbonates and hydroxides
Calcium hydroxide is only slightly soluble.	Lead salts are more soluble in hot water.

Common uses of salts.

Now, check what you have just learnt by trying out the learning activity below!

Learning Activity 2

20 minutes

Answer the following questions:

1. Define the following terms:
   (i) Alkali ____
   (ii) Hydrated salt ____
   (iii) An hydrated salt ____

2. Write the chemical formula of the following salt:

	a. anhydrous	b. hydrated
(i) Zinc sulphate	__________	_________
(ii) Copper (II) sulphate	___________	__________
(iii) Magnesium sulphate	__________	_________

Thank you for completing your learning activity 2. Check your work. Answers are at the end of this module.

Arrhenius, Bronsted and Lewis definitions of acids and bases

Arrhenius acids and bases

In 1887 the young Swedish chemist, Svante Arrhenius, proposed a revolutionary way of thinking about acids and bases. He said that acids are compounds containing hydrogen that ionizes to produce **hydrogen ions (H^+)** in aqueous solution. Similarly, he said that bases are compounds that ionize to produce **hydroxide ion (OH^-)** in an aqueous solution.

Some common acids	
Name	Formula
Hydrochloric acid	HCl
Nitric acid	HNO_3
Sulphuric acid	H_2SO_4
Phosphoric acid	H_3PO_4
Acetic acid	CH_3COOH
Carbonic acid	H_2CO_3

Based from the definition, below are of common acids. They are classified according to the numbers of hydrogen present in the compound:

1. **Monoprotic acid** that contains one ionisable proton.
   Example: **Nitric acid (HNO_3)**

2. **Diprotic acid** contain two ionisable protons.
   Example: **Sulphuric acid (H_2SO_4)**

3. **Tripotic acid** that contains **three** ionisable protons.
   Example: **Phosphoric acid (H_3PO_4)**

Arrhenius definition of Acids and Base
Compound that contains Hydrogen and that ionizes to produce hydrogen ions (H^+).
Bases are compound that ionize to produce hydroxide ions (OH^-).

The table below shows the common example of bases.

Some common Bases		
Name	**Formula**	**Solubility in water**
Potassium hydroxide	KOH	high
Sodium hydroxide	NaOH	high
Calcium hydroxide	$Ca(OH)_2$	low

The base with which you are perhaps most familiar is sodium hydroxide (NaOH). This ordinary lye, sodium metals reacts with water to form sodium hydroxide.

$$\text{Sodium + water} \longrightarrow \text{sodium hydroxide + hydrogen}$$
$$2Na + 2H_2O \longrightarrow 2NaOH + H_2$$

Both potassium hydroxide and sodium hydroxide are ionic solids. They dissociate completely into metal ions and hydroxide ions when dissolved in water. Sodium and Potassium are group 1A elements. They named the alkali metals because they react with water to produce alkaline solutions.

Bronsted-lowry acid and bases
One problem with the Arrhenius theory is that it is not comprehensive enough. It does not explain the behavior of all compounds that behave as bases. For example, aqueous solutions of ammonia (NH_3) and sodium carbonate (Na_2CO_3) are basic. Neither of these compounds is a hydroxide.

In 1923 Johannes Bronsted (Danish 1879-1947) and Thomas Lowry (English 1874-1936) independently proposed a new theory. The Bronsted-Lowry theory defines **an acid as a hydrogen-ion donor**. Similarly, Bronsted-Lowry **base is hydrogen-ion acceptor**.

All the acids and bases included in the Arrhenius theory are also acids and bases according to the Bronsted-Lowry theory. Some compounds that were not included in the Arrhenius theory are classified as bases in the Bronsted-Lowry theory.

The behaviour of ammonia as a base is understood using the Bronsted-Lowry theory. Ammonia gas is very soluble in water. When ammonia dissolves it acts as a base. It accepts a hydrogen ion from water.

Ammonia	+	water	$\rightleftharpoons$	ammonium ion	+	hydroxide ion
(Hydrogen ion Acceptor, Bronsted-basic)		(hydrogen-ion donor, Bronsted-Lowry acid)		(makes the solution Lowry base)		

$$NH_3 + H_2O \rightleftharpoons NH_4^+ + OH^-$$

In this reaction ammonia (NH_3) is the hydrogen-ion acceptor. Therefore; it is a Bronsted-Lowry base. Water (H_2O), the **hydrogen-ion donor**, is a Bronsted Lowry acid.
Hydrogen ions are transferred from water to ammonia. This causes the hydroxide-ion concentration to be greater than it is in pure water. As a result, solution of ammonia drives off ammonia gas. This action causes the equilibrium in the ammonia dissolution equation to shift to the left. The ammonium ion, NH4$^+$ gives up hydrogen ion, it acts as an acid. The hydroxide ion accepts H^+, It acts as a base. Overall then, this equilibrium has two acids and two bases.

$$NH_3 + H_2O \rightleftharpoons NH_4^+ + OH^-$$

The Bronsted- Lowry theory is also applicable to acids.

For example:

The ionization of hydrochloric acid in water is shown below:

$$HCl + H_2O \rightleftharpoons H_3O^+ + Cl^-$$

Proton transfer

In this reaction hydrochloric acid (HCl) is hydrogen – ion donor. And water (H_2O) is the hydrogen-ion acceptor.

Note: Water has split personality, sometimes it receives hydrogen ion, and other times donates a hydrogen ion.

The Bronsted- Lowry definition: Acid is a hydrogen-ion donor. Base is hydrogen-ion acceptor.

Lewis acids and bases

A third theory of acids and bases was proposed by Gilbert Lewis (1875-1946). He focused on the donation or acceptance of pair of electrons during a reaction. This concept is more general than either the Arrhenius theory or Bronsted-Lowry theory.

A Lewis acid is a substance that can accept a pair of electrons to form a covalent bond. A Lewis base is a substance that can donate a pair of electrons to form covalent bonds.

A hydrogen ion (Bronsted-Lowry acid) can accept a pair of electrons in forming a bond. A hydrogen ion, therefore, is also a Lewis acid. A substance that accepts a hydrogen ion (Bronsted-Lowry base) must have a pair of electrons available.
(Lewis base).

Consider the reaction of hydrogen (H^+) and hydroxide (OH^-)

$$H^+ + OH^- \longrightarrow H_2O$$

$$\underset{\text{Lewis Acid}}{H^+} + \underset{\text{Lewis base}}{:\ddot{\underset{\cdot\cdot}{O}}-H} \longrightarrow H-O-H$$

A hydroxide ion is a Lewis base. It is also a Bronsted-Lowry base. A hydrogen ion is both a Lewis acid and Bronsted-Lowry theory. It also includes some compounds that are not classified as Bronsted-Lowry acid or base.

Lewis definition:
Substance that can accept a pair of electrons to form a covalent bond is an acid.
Substance that can donate a pair of electrons to form covalent bonds is a base.

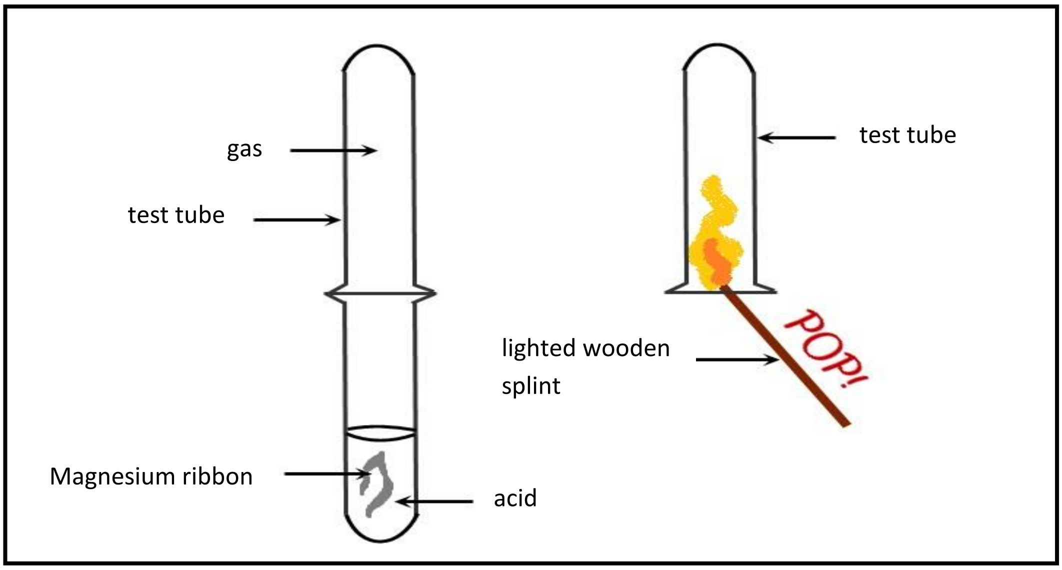
Chemical properties of acid

1. Acids react with metals
   Most diluted acids react with metals that are more reactive than hydrogen in the reactivity series to produce a salt and hydrogen gas. An example is the reaction of dilute sulphuric acid with magnesium to form magnesium sulphate and hydrogen.

Magnesium + sulphuric acid $\longrightarrow$ Magnesium sulphate + hydrogen

$$Mg(s) + H_2O_4(aq) \rightarrow MgSO_4(aq) + H_2(g)$$

To show that hydrogen gas is produced during the reaction, a burning (lighted) splint is placed near the mouth of the test tube. A pop sound is heard and the splint is extinguished.



Test for hydrogen gas

2. Acids react with metal carbonates.
   Many dilute acids react with metal carbonates to produce a salt, water, and carbon dioxide gas.
   For example, dilute hydrochloric acid reacts with calcium carbonate or limestone chips to produce calcium chloride, water, and carbon dioxide.

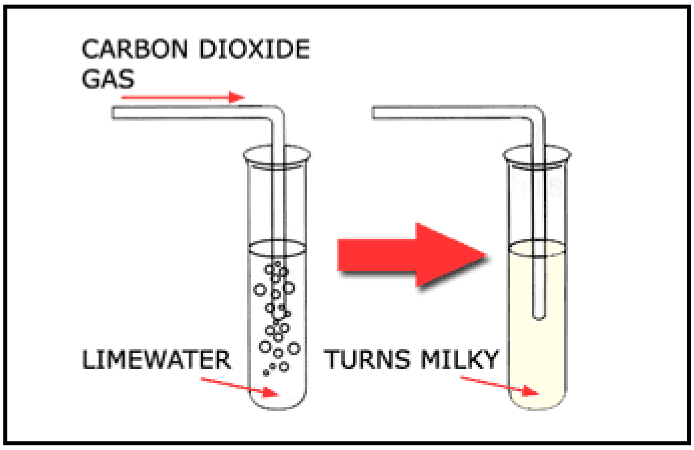
   The equation is shown below.

   acid + metal carbonate $\rightarrow$ salt + water + Carbon dioxide

   Hydrochloric acid + Calcium carbonate $\rightarrow$ Calcium chloride + Water+ Carbon dioxide

   $2HCl_{(aq)} + CACO_{3(s)} \rightarrow CaCl_{2(aq)} + H_2O_{(l)} + CO_{2(gas)}$

To show that carbon dioxide is liberated, the gas is passed through **limewater** using a delivery tube. A white precipitate of calcium carbonate (marble chips) is formed and the limewater turns **milky.**



Test for carbon dioxide gas

3. Acids react with bases.
   Acids react with (metal oxides and metal hydroxide) to form salt and water. For example, the reaction of dilute sulphuric acid with copper(II)oxide to form copper (II) sulphate and water. Hydrochloric acid with sodium hydroxide to form sodium chloride and water. The reaction between acid and base is called **neutralisation.**

(i) Acid + metal oxide → salt + water

Copper (II) oxide + Sulphuric acid → Copper (II) sulphate + Water

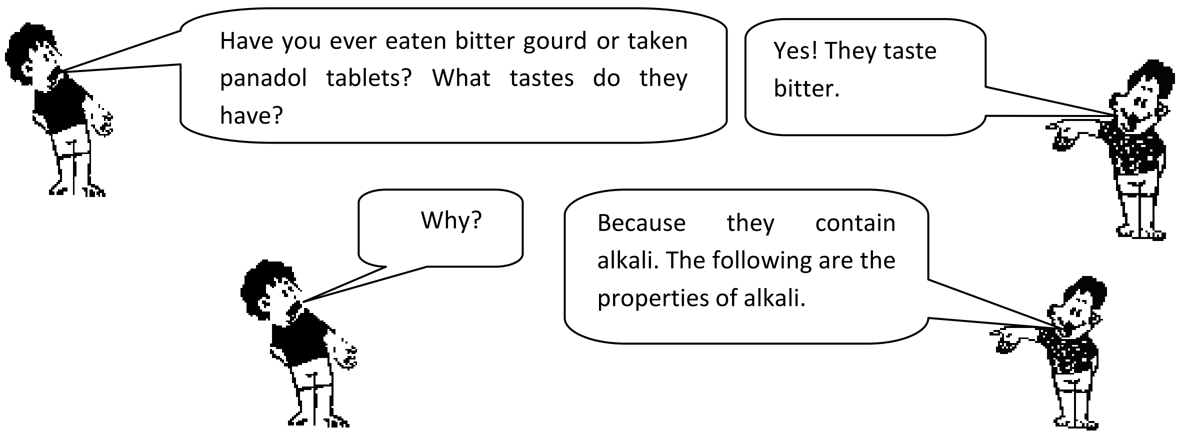
$$CuO_{(s)} + H_2SO_4(_{aq}) \longrightarrow CuSO_4(_{aq}) + H_2O(_{l})$$

(ii) Acid + metal hydroxide → salt + water

Hydrochloric acid + Sodium Hydroxide → Sodium chloride + water

$$HCl + NaOH \rightarrow NaCl + H_2O$$

Chemical properties of base (alkali)



The following are chemical properties of alkalis:

1. Most alkalis have bitter taste.
2. Alkalis turn red litmus paper blue.
3. Alkalis react with acids to form salt and water. This process is called **netralisation**
4. Alkalis react with ammonium salts to produced ammonia gas. If some sodium hydroxide solution is warmed with a little ammonium chloride in a test tube, ammonia gas will be given off and a strong smell can be detected.

For example:

Sodium hydroxide + ammonium chloride $\longrightarrow$ Sodium chloride + water+ Ammonia

$$NaOH_{(aq)} + NH_4Cl_{(s)} \longrightarrow NaCl_{(aq)} + H_2O_{(l)} + NH_{3(g)}$$

Ammonia gas can be detected by its pungent smell (like that of urine). It can be tested by holding a piece of damp red litmus paper near the mouth of the test tube. The gas will turn red litmus paper blue.

This reaction shows that ammonia is an alkaline gas. It dissolves in water to form aqueous ammonia, which is week alkali.

Now, check what you have just learnt by trying out the learning activity below!

Learning Activities 3 **20 minutes**

Answer the following questions:

1. Write the following definition of acid and base by:

   i. Arrhenius ____
   ____

   ii. Bronsted-lowry ____
   ____

   iii. Lewis ____
   ____

2. Define the following type of acid:

   i. Monoprotic ____
   ii. Diprotic ____
   iii. Tripotic ____

3. Write the chemical formula of the following compound and identify whether they are monoprotic, diprotic or tripotic acid.

   i. Acetic ____________ ___________

   ii. Nitric ____________ ___________

   iii. Sulphuric ___________ ___________

   iv. Carbonic ___________ ___________

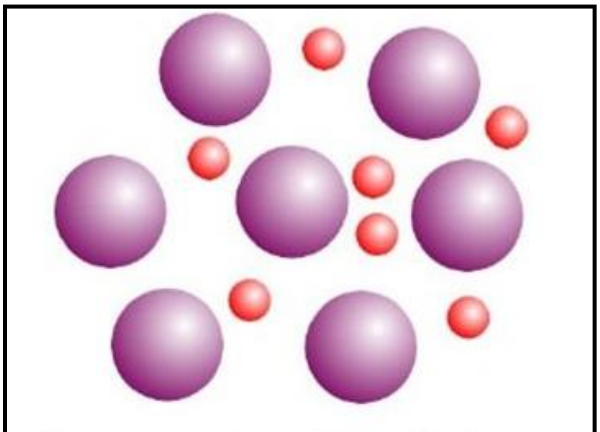
   v. Phosphoric ___________ ___________

Thank you for completing your learning activity 3. Check your work. Answers are at the end of this module.

12.2.2 Strong and weak acids and bases

Is it true that strong acids are acids that burn the skin and weak acids are harmless? The terms `strong' and `weak' are often confused with `concentrated' and `dilute'.

An acid like hydrochloric acid is strong but if it is diluted many times with water, it is not corrosive. On the other hand, a weak acid like ethanoic acid can cause burns to the skin if it is pure.



Strong acids are assumed to dissociate completely when in aqueous solution.

The table below shows some example of strong and weak acid.

Strong acid	Weak acid
Hydrochloric acid (HCl)	Methanoic acid
Sulphuric acid (H_2SO_4)	Ethanoic acid
Nitric acid (HNO_3)	Citric acid

Strong acid is an acid that is completely ionized in water.

For example:

Hydrochloric acid is a strong acid because all its molecules produce hydrogen ions $H^+_{(aq)}$ when it is dissolved in water.

$$HCl_{(aq)} \rightarrow H^+ + Cl_{(aq)}$$

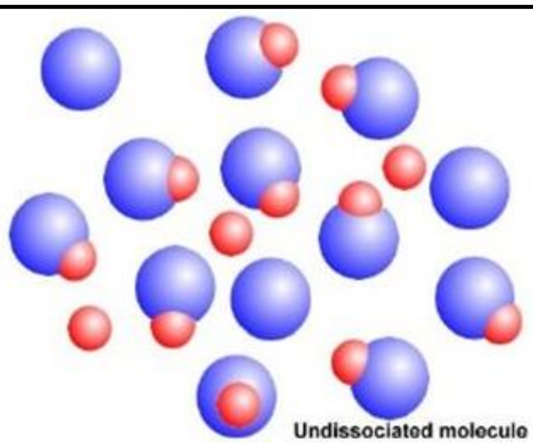
Similarly, sulphuric acid and nitric acid are strong acids because all their molecules will ionize in water to produce hydrogen ions.

$$H_2SO_{4(aq)} \rightarrow H^+_{(aq)} SO_4^{2}{}_{(aq)}$$

$$HNO_{3(aq)} \rightarrow H + _{(aq)} + NO_{3}{}_{(aq)}$$

On the other hand, **ethanoic acid** is classified as a weak acid because only some molecules ionize to produce hydrogen ions when it is dissolved in water.

$$CH_3COOH_{(aq)} \rightarrow H^+_{(aq)} + CH_3COO_{(aq)}$$



Weak acids dissociate only slightly in aqueous solution. The majority of molecules remain undissociated.

Weak acid is an acid that is only partially ionized in water

The table below shows some examples of strong and weak bases.

Strong base	Weak base
potassium hydroxide (KOH)	Sodium cyanide (NaCN)
Sodium hydroxide (NaOH)	Ammonia (NH_3)
Calcium hydroxide (CaOH)	Sodium silicate(Na_2SiO_3)

Strong base completely dissociate with water.

For example: $NaOH \longrightarrow Na^+_{(aq)} + OH^-_{(aq)}$

Weak base dissociate partially in water.

For example: $NH_3 + H_2O \rightleftharpoons NH_4^+{}_{(aq)} + OH^-_{(aq)}$

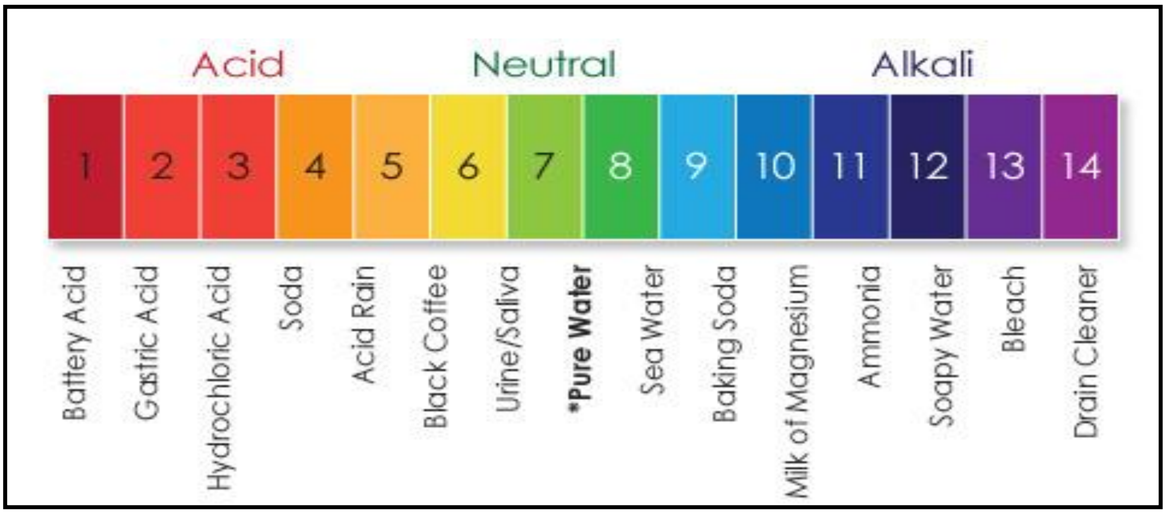
Proton transfer

Water acts as an acid by donating protons to ammonia molecules. Since ammonia is a weak base, only dissociate partially. Few ammonium and hydroxide ions formed.

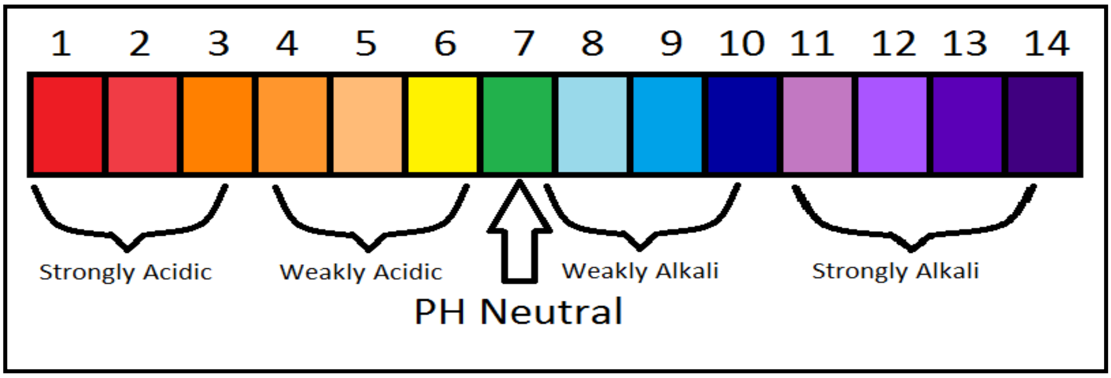
To find out whether an acid is strong or weak, we can measure the concentration of the hydrogen ions in the solution with an **electrical pH meter**. It is a more reliable and accurate method of measuring pH than using the universal indicator.

A pH meter can measure the concentration of hydrogen ions in a solution.

The pH Scale
The concentration of the hydrogen ions in a solution can be measured with a set of numbers or scale called **pH scale.** pH is the potential of hydrogen.



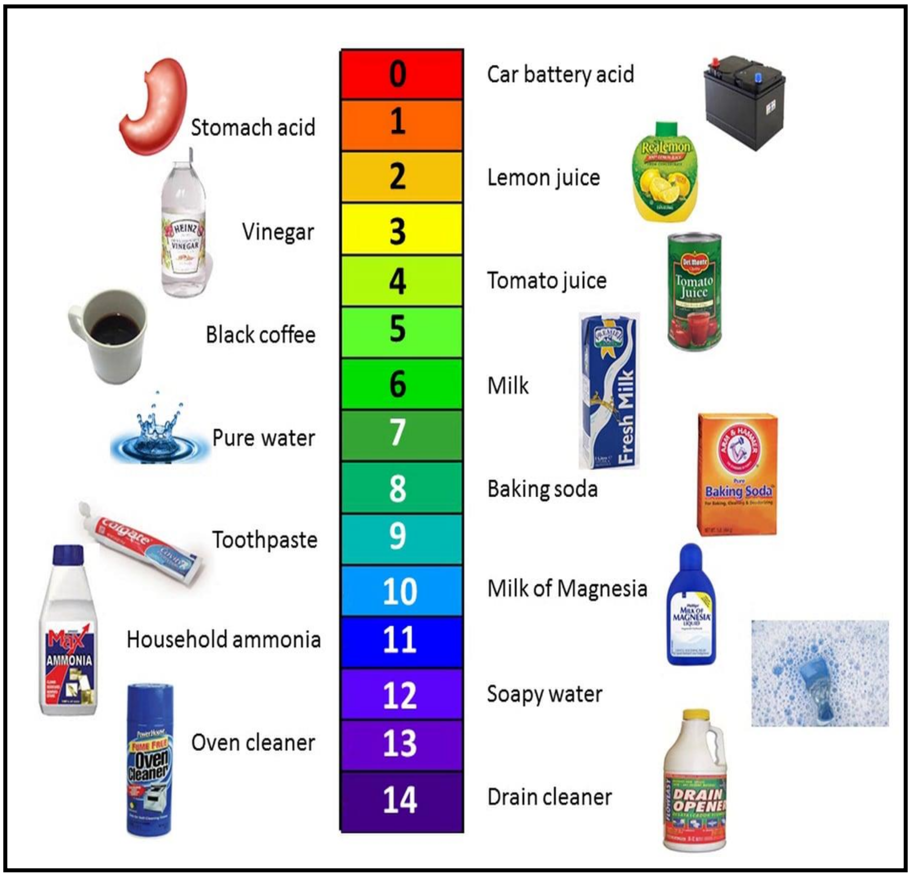
The pH of a solution is related to the concentration of hydrogen ions and hydroxide ions.



The pH of a solution is a measure of how acidic or alkaline the solution is. The pH scale ranges from 0 to 14. **The lower the pH, the more acidic the solution** is. It contains a **higher concentration of hydrogen ions**. **A pH of 7** is neutral. Pure water has a pH of 7. **The higher the pH, the more alkaline** the solution is, that is, it contains a higher concentration of hydroxide ions.

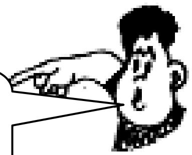
Do you know the pH of some common household items?

The diagram below shows some common substances and their approximate pH.

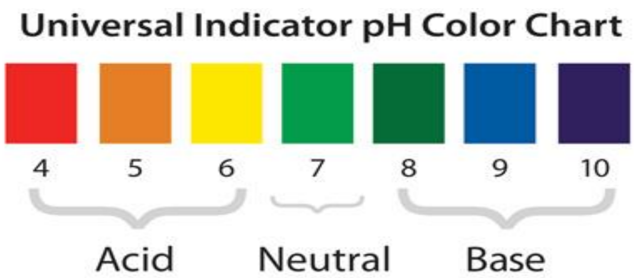


pH of some common substances

Universal indicator



1. Chemical compound called an indicator.
2. pH probe attached to a data logger.

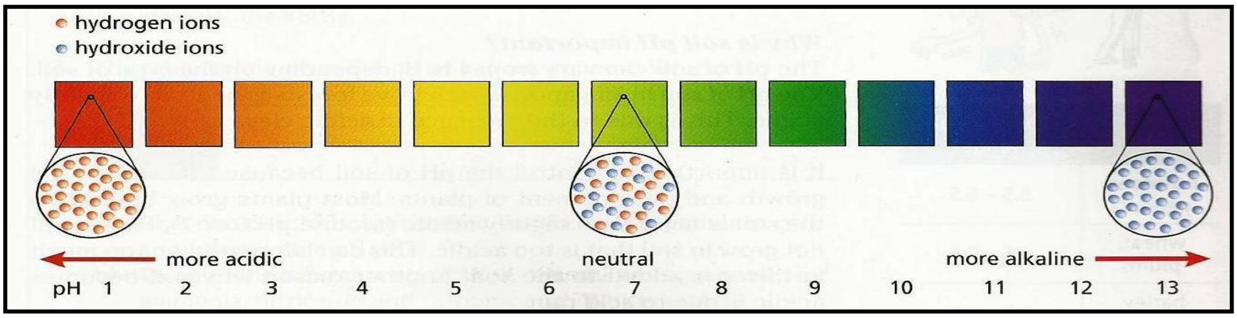


pH probe and pH logger

Universal indicator contains a mixture of dyes. It gives different colour in solutions of different pH.

The **pH** of a solution is calculated based on the number of hydrogen ions or hydroxide ions. Acids with a smaller pH value have higher concentration of hydrogen ions. Alkaline solutions with a larger pH value have a higher concentration of hydroxides ions.

Look at the diagram below.



pH scale

Hence, pH can be used to compare the strength of acids and alkalis of the same concentration. For example 0.1 mol/dm^3 of hydrochloric acid has a ph of 1. It is a stronger acid than 0.1 mol/dm^3 of ethanoic acid, which has a pH of 3.

Besides universal indicator, there are many other coloured indicators that are commonly used in titrations. Each indicator has one colour in an acidic solution and another colour in an alkaline solution. It will change its colour at a certain pH value.

Some important indicators are listed below.

Indicator	Colour in acidic solution	pH range at which indicator changes colour.	Colour in alkaline solution
Methyl orange	red	3-5	yellow
Screened methyl orange	violet	3-5	green
litmus	red	5-8	blue
Bromothymol blue	yellow	6-8	blue
phenolphthalein	colourless	8-10	pink

Colour changes of some indicators.

Note: Apart from the indicators above, there are some naturally indicator that can be used. Example: Hibiscus flower and purple cabbage.

Now, check what you have just learnt by trying out the learning activity below!

Learning Activities 4

20 minutes

Answer the following questions.

1. Define the following:

   i. Strong Acid ____________________

   ii. Weak acid ____________________

   iii. pH ____________________

2. Mary tested some solutions with a universal indicator paper. She wrote their pH numb numbers as **1, 5, 7 and 14.** But she forgot to write the names and symbols of the solutions. Can you help her write the correct symbols that match the pH number of the correct solutions?

Solution tested	pH	Chemical formula
Vinegar (acetic acid)		
Sulphuric acid		
Distilled water		
Sodium Hydroxide		

3. Identify the following compound as a strong, weak acid or strong, weak base:

   i. Citric ____________

   ii. Ammonia ____________

iii. Hydrochloric acid _____________

iv. Potassium hydroxide _____________

Thank you for completing your learning activity 4. Check your work. Answers are at the end of this module.

Dilute and Concentrated Acids and Bases

The word **concentrated** and **dilute** tell how much of an acid or a base is dissolved in a solution. The concentration of a solution is a measure of the quantity (amount) of dissolved **solute** in a given **volume of solution**. And often measured in moles per litre, calculated using the following equation:

$$C = \frac{n}{v}$$

where c = concentration (mol/L)
n = number of moles dissolved solute (mol)
v = volume of solution (L)

For example:

Calculate the concentration of a solution that contains 0.24 mol of hydrochloric acid (HCl) dissolved in 2.0L of solution.

By definition, molarity measures the amount of solute dissolved per litre of solution. If **0.24 mol** is dissolved in **2.0L** of solution then **0.12mol** dissolved in 1.0L of solution.

Alternatively, we can use the relationship:

$$n = c \times V \text{ or can be written as: } c = \frac{n}{V}$$

$$\text{So } c\,(HCl) = \frac{n(HCl)}{V(HCl)}$$

$$= \frac{0.24\text{mol}}{2.0\text{L}}$$

$$\mathbf{= 0.12\ mol/L\ or\ 0.12M}$$

Dilution using molarity

Dilution was discussed in previous module. Similar dilution calculations can be done using molarity. A useful mathematical relationship can be derived from the process of dilutions.

$$n_1 = c_1V_1 = n_2 = c_2V_2$$

where: n = number of mol

c = concentration per mol

V = volume of solution

For example:

The concentration of $\mathbf{H^+}_{(aq)}$ ion in pure gastric juices is 0.1M. 10ml sample of gastric juice is added to 90ml of water. Calculate the concentration of $H^+_{(aq)}$ in the diluted solution.

Solution:

$c_1V_1 = c_2V_2$ since we are looking for c_2 change the subject of your formula

$$c_2 = \frac{c_1V_1}{V_2}$$

$$c_2 = \frac{0.1M \times 10ml}{100ml}$$

$\mathbf{C_2 = 0.01M}$

Now, check what you have just learnt by trying out the learning activity below!

Learning Activities 5 **30 minutes**

Answer the following questions:

1. Calculate the concentration (molarity) of the following solutions:

   i. 0.50L of solution, which contains 0.24mol of glucose molecules

ii. 0.20L of solution, which contains 0.010 mol of sodium chloride

2. Calculate the concentration of each of the following diluted solutions:

   i. 10ml of water is added to 5ml of 1.2M HCl

   ii. 1.0L of water is added to 3.0L of 0.10M HCl

Thank you for completing your learning activity 5. Check your work. Answers are at the end of this module.

Amphoteric and amphiprotic substances

Amphoteric are solid substances that show:

- basic properties by reacting with acids
- acidic properties by reacting with bases.

Example:

Aluminium hydroxide and zinc oxide are amphoteric properties.

Solution	Aluminium hydroxide Al(OH)3	Zinc oxide, ZnO
Acidic	$Al(OH)_3 + 3H^+ \rightarrow Al_3^+ + 3H_2O$	$ZnO + 2H^+ \rightarrow Zn_2^+ + H_2O$
Basic	$Al(OH)_3 + OH \quad [Al(OH)_4]$	$ZnO + 2OH \quad + H_2O \rightarrow [Zn(OH)_4]^2$

Amphiprotic are substances such as water, which can either accept or donate protons.

Other substances that exhibit this behaviour are the hydrogen carbonate ions and the hydrogen sulfate ions.

Amphiprotic substances contain a:

- Hydrogen atom with a polar bond to a non-metal atom, so they can act as proton donors or acids.
- Non-metal atom with a lone pair of electrons, so they can act as proton acceptors or bases. Substances that neither accept nor donate protons are said to be **neutral.**

Water is an amphiprotic molecule. Water contains hydrogen atoms linked by polar bonds to oxygen and lone pairs of electrons. A water molecule can act as both a proton donor and a proton acceptor**.**

- **Acting as an acid**, water is a proton donor:

$$H_2O \rightarrow H^+ + OH$$

- **Acting as a base**, water is a proton acceptor:

$$H_2O + H^+ \rightarrow H_3O^+$$

Example of amphiprotic ions:

Hydrogen carbonate and hydrogen sulphate are amphiprotic ions.

Amphiprotic ion	Hydrogen carbonate HCO_3^-	Hydrogen sulphate HSO_4^-
Acting as an acid	$HCO_3 \quad + H_2O \rightarrow H_3O \ ^- + CO_3^{\ 2}$	$HSO_4 \quad + H_2O \rightarrow H_3O + + SO_4^{\ 2}$
Acting as a base	$HCO_3^{\ +} \quad + H_2O \rightarrow OH \quad + H_2CO_3$	$HSO_4^{\ -} \quad + H_2O \rightarrow OH \quad + H_2SO_4$
Dominant reaction in solution	$HCO_3 \quad + H_2O \rightarrow OH^- \quad + H_2CO_3$	$HSO_4^{\ -} \quad + H_2O \rightarrow H_3O^+ + SO_4^{\ 2}$
Nature of solution	basic	acidic

Now, check what you have just learnt by trying out the learning activity below!

Learning Activities 6 **20 minutes**

Answer the following questions:

1. Give two example of water acting as:

   i. An acid ____

   ii. A base ____

2. Classify the following substances as basic, amphoteric or amphiprotic. Give reasons for your answers, including equations.

   i. $Zn(OH)_2$

   ii. NaOH

Thank you for completing your learning activity 6. Check your work. Answers are at the end of this module.

12.2.3 Dissociation constant and pH calculation

Calculation of acid and base constant.
Strong acid are completely ionised in aqueous solution.

For example: Hydrochloric acid → Hydrogen ion + hydroxide ion

$$HCl \rightarrow H^+ + Cl \text{ (100\% ionised)}$$

Weak acids ionised only partially in aqueous solution. The ionisation of acetic acid, a typical weak acid, is not complete.

$$HC_3COOH + H_2O \rightarrow H_3O^+ + C_2H_3O^2$$

Over 99% of acetic acid molecules exist in solution as undissociated covalent molecules. Less 1% are ionised at any instant.

The equilibrium-constant expression for ionisation of acetic acid can be written from preceding equation.

$$K_a = \frac{[H_3O+]x[C_2H_3O^2\ \ ]}{[HC_3COOH]x[H_2O]}$$

For dilute solutions, the concentration of water is constant. It can be combined with k_{eq} to give an acid dissociation constant.

An acid dissociation constant (k_a) is the ratio of the concentration of the dissociated form of an acid to the undissociated form. The dissociated form includes both hydrogen ion and the anion.

$$K_a x[H_2O] = K_a \frac{[H^+]x[C_2H_3O_2\ \ ]}{[HC_3COOH]}$$

The acid dissociation constant, for acetic acid at 25°C is $1.8x10^{-5}$

A weak acid has a small K_a value. A larger value of K_a means that dissociation, or ionisation, of the acid is more complete.

Just as there are strong acids and weak acids, there are also strong bases and weak bases.
Strong bases dissociate completely into metal ions and hydroxide ions in aqueous solution.
Weak bases do not dissociate completely in aqueous solution.
Aqueous ammonia is an example of weak base.

$$NH_3 + H_2O \rightarrow NH_4^+ + OH$$

Approximately 99% of the ammonia is un-ionised. Only 1% is in form of NH_4^+ and OH^- are available.

The base dissociation constant, K_b, is the ratio of the dissociated form of a base to the undissociated form.

It indicates the degree of dissociation. The base dissociation constant for ammonia is $1.8x10^{-5}$

$$K_b = \frac{[NH_4^+]x[OH\ \ ]}{[NH_3]} = 1.8x10\ \ 5$$

Example of calculating dissociation constant

1. A 0.100M solution of acetic acid is only partially ionised. Using a measure of pH, the $[H^+]$ is calculated as $1.34x10^{-3}$M. What is the acid dissociation constant of acetic acid?

Solution:
Write the equation for the ionisation of acetic acid.

$$CH_3COOH_{(aq)} \rightleftharpoons H^+ + C_2H_3O^{2-}$$

Each molecule of $HC_2H_3O_2$ that ionises gives a H^+ and an $C_2H_3O_2$ ion. Therefore, at equilibrium the $[H+] = [C_2H_3O^2\ ] = 1.34x10^{-3}$M. The equilibrium concentration of $HC_2H_3O_2$ is the initial concentration changed by ionisation of the acid. (0.100-0.00134) M = 0.09866M.

The equilibrium now values are substituted into the expression for K_a:

$$K_a = \frac{[H+]x[C_2H_3O^2\ ]}{[CH_3COOH]}$$

$$= \frac{(1.34x10^3\ )x(1.34x10^3\ )}{0.0987}$$

$$\underline{= 1.82x10^{-5}}$$

Hydrogen ion from water
Water molecules are highly polar. (Even at a room temperature they are in continuous motion). Occasionally, the collisions between water molecules are energetic enough that hydrogen ions are transferred from one water molecule to another.

A water molecule that loses hydrogen ions becomes a negatively charged Hydroxide ion (OH^-).

A water molecule that gains hydrogen ions becomes a positively charged hydronium ion (H_3O^+).

$$H_2O + H_2O \rightarrow H_3O^+ + OH- -$$

The reaction in which two water molecules react to give ions is the self-ionisation of water. This can be written as a simple dissociation.

$$H_2O \rightleftharpoons H^+ + OH^-$$

Self-ionisation of water occurs to very small extent. In pure water at 25°C, the concentration of hydrogen ions and the concentration of hydroxide ions are produced in a 1:1 ratio.

Chemist $[H^+]$ describes pure water as neutral. In fact, any aqueous solution in which hydrogen concentration [and the hydroxide concentration $[OH^-]$ are 1.0x10-^{7} mol/L is described as a neutral solution.

In any aqueous solution the $[H^+]$ and the $[OH^-]$ are interdependent. If the $[H^+]$ increases then the $[OH^-]$ decreases or the other way around.

For aqueous solutions, the product of the hydrogen ion concentration, $[H^+]$, and the hydroxide ion concentration, $[OH^-]$, is equal to $1.0x10^{-14}$ $(mol/L)^2$ as shown below:

$$[H^+] \quad x\ [OH^-] = 1.0x10^{-14}(mol\ /L)^{\ 2}$$

The product of the concentrations of the hydrogen ions, and hydroxide ions in water is K_w the ionisation product constant for water**.**

$$\mathbf{K_w} = [H^+]x\ [OH^-] = 1.0x10^{-14}(mol\ /L)^{\ 2}$$

In acidic solution the $[H^+]$ is greater than the $[OH^-]$, therefore, the $[H^+]$ of acidic solution is greater than $1.0x10^{-7}$ mol/L

In basic solution the $[H^+]$ is less than the $[OH^-]$. Therefore, the $[H^+]$ of a basic solution is less than $1.0x10^{-7}$ mol/L basic solution is also known as alkaline solution.

Example:

1. If hydrogen concentration $[H^+]$= $1.0x10^{-5}$ mol/L, is the solution acidic, basic, or neutral?
   The hydrogen concentration $[H^+]$ is $1.0x10^{-5}$ mol/L. Because this is greater than $1.0x10^7$ mol/L the solution is **acidic.**

2. What is the hydroxide concentration $[OH^-]$ of this solution?
   We know that K_w = $[H^+]x\ [OH^-]$= $1.0x10^{-14}$(mol /L) or $10-^{14}$M

$$\textbf{Therefore } [OH^-]\ \frac{1.0x10^{\ 14}\ mol/L}{1.0x10^{\ 5}\ mol/L} = 1.0x10^{\ 9}\ mol/L$$

Now, check what you have just learnt by trying out the learning activity below!

Learning Activities 7 **20 minutes**

Answer the following questions:

1. Calculate the hydrogen ion concentration of 2.0M ethanoic acid solution. K_a for acetic acid is $1.8x10^{-5.}$

2. 0.200M solution of a weak acid has a hydrogen concentration $[H^+]$ of $9.86x10^{-4}$. What is the pH of this solution?

Thank you for completing your learning activity 7. Check your work. Answers are at the end of this module.

---

pH and pOH calculation

The pH concept

The expression of hydrogen-ion concentration in moles per litre is seldom used. A more widely used system is the ph scale. It was proposed in 1909 by a Danish scientist Soren Sorensen (1868-1939).

The pH of a solution is the negative logarithm of hydrogen-ion concentration.

$$pH = -\log[H^+]$$

In a neutral solution the $[H^+]= 1x10^{-7}$ mol/L. The pH of neutral solution is 7

Solution:

$pH = -\log [H^+]$

$pH = -\log (1.x10-7)$

$= -(\log1 + \log 10-7$

$= -(0.0 + (-7))$

$= 7.0$

In a similar way, the pOH of a solution equals the negative logarithm of the hydroxide ion concentration.

$$\mathbf{pOH = -\log[OH^-] \text{ or } pOH = 14 - pH}$$

Example:

What is the pH of a 0.08M solution of Sodium hydroxide?

Solution:

$pOH = -\log[OH^-]$

or $pOH = 14 - pH$

$= -\log[0.08]$

$= 1.09$

$pOH = 14 - 1.09 = 12.91$

Now, check what you have just learnt by trying out the learning activity below!

Learning Activities 8

30 minutes

Answer the following questions:

1. Calculate the pH of the following solutions :

   (i) 0.1M of HCl solution

   (ii) 0.01M of NaOH

   (iii) 0.005M H_2SO_4

Thank you for completing your learning activity 8. Check your work. Answers are at the end of this module.

12.2.4 Acid-base titration to detect the end point

Volumetric titration

Titration is a common laboratory method of quantitative/chemical analysis that can be used to determine the concentration of a known reactant (**analyte**). The basis of the method is a chemical reaction of a **standard solution (titrant)** with a solution of an analyte.

The **analyte (*A*)** is a solution of the substance whose concentration is unknown and sought in the analysis. The **titrant (*T*)** is a solution in which the concentration of a solute is precisely known. Because volume measurements play a key role in titration, it is also known as **volumetric analysis**.

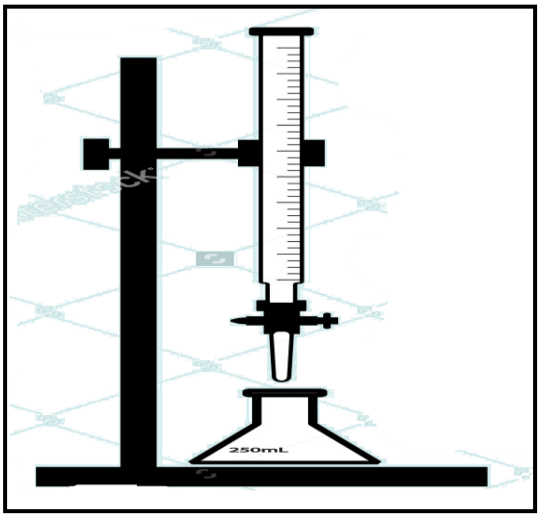
Usually it is the **volume of the titrant** required to react with a given quantity of an analyte that is precisely determined during a titration.

A sample of hydrochloric acid added to a large volume of water becomes more dilute, but is still strong acid.

Acids and Bases

(i) A strong acid reacts with a strong base to form a neutral (pH=7) solution.
(ii) A strong acid reacts with a weak base to form an acidic (pH<7) solution.
(iii) A weak acid reacts with a strong base to form a basic (pH>7) solution.

A weak acid reacts with a weak base. The endpoint solution is basic if the base is stronger and acidic if the acid is stronger. If both are of equal strength then the endpoint pH is neutral.



Example of titration set up

Titration procedure:

1. A known volume ***VA*** of the analyte (acid) is placed in a titration flask.

2. The burette is filled by a standard solution (titrant, base) of known concentration ***cT*** (M).

3. Before the titration is started, 1-3 drops of indicator (phenolphthalein) is placed in the titration flask with the analyte (acid). The chosen indicator must be one colour when the solution is acidic.

4. A base solution is then slowly added from the burette, drop by drop.

5. The titration continues, drop by drop, until the indicator suddenly achieves the intermediate color (weak pink) between that of the acid and the colour of the base (fuchsia). At that point the titration ceases.

6. The point at which the system is neither acidic nor basic is referred to as the endpoint. The endpoint will correspond to a perfect stoichiometric relationship between the acid and the base.

7. Once the endpoint has been reached, the burette must be read. The bottom of the meniscus line determines the quantity of the base ***VT*** that was required to reach the endpoint.

8. Once the titration is completed, the final calculation is done.

Acid - base titration

In normal titration, it is required to carry out at least one rough and two accurate titrations. If no consistent volumes are obtained, the average of at least three readings should be used.

For example:

The table below shows how the results are to be recorded and the mean volume of titre calculated.

Titration number	1	2	3	4
Final reading of burette /cm^3	25.2	24.8	33.3.6	24.9
Initial reading of burette/cm^3	0.0	0.0	7.4	0.1
Volume of Noah used/cm^3	25.2	24.8	25.9	24.8
Best titration results (✓)	✓	✓	✓	

Recording the results of titration

Average volume of NaOH used in the titration = 24.8 cm^3.

Example 1: **Titration of a known acid with alkali**

Find the concentration of a solution of sulphuric acid by titrating 25.0 cm^3 portions of acid against a standard solution of 0.100 mol/dm^3 sodium hydroxide. Use phenolphthalein as the indicator. The equation for the reaction is shown below.

$$H_2SO_4 + 2NaOH \longrightarrow Na_2SO_4 + 2H_2O$$

Solution:

Supposed the results obtained are as shown in the table above, average volume of sodium hydroxide used = 24.8 cm^3

From the equation, $\frac{\text{number of moles of sulphuric acid}}{\text{number of moles of sodium hydroxide}} = \frac{1}{2}$

$$\frac{\text{volume of } H_2SO_4 \text{ x concentration of } H_2SO_4}{\text{number of NaOH x concentration of NaOH}} = \frac{1}{2}$$

$$\frac{\frac{25.0}{1000} \text{ x concentration of } H_2SO_4}{\frac{24.8}{1000} \text{ x } 0.100 \text{ mol/dm}^3} = \frac{1}{2}$$

Concentration of H_2SO_4 = $\frac{1 \text{ x } 24.8 \text{ x } 0.100 \text{ mol/dm}^3}{2 \text{ x } 25.0}$

= 0.0496 mol/dm³

Example 2:

What is the concentration of a solution of HCl of which 25ml would be needed to neutralize 12.5ml of 2.0M NaOH solution?

Let the concentration of HCl = c mole/L
The volume of HCl = 25.00ml
Concentration of NaOH = 2.0 mol/L
Volume of bases added = 12.5ml

$$NaOH + HCl \longrightarrow NaCl + H_2O$$

One mole of NaOH reacts with one mole of HCl to give one mole NaCl and one mole of water.

The number of moles of NaOH = $\frac{2.0\text{M x } 12.5}{1000}$

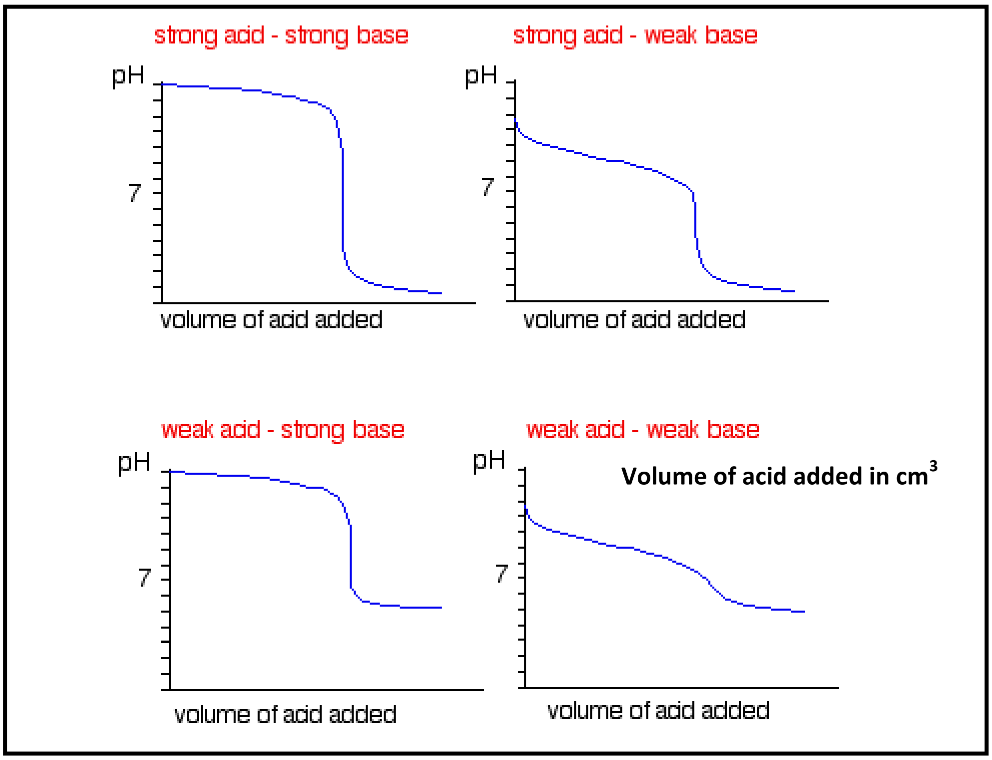
The number of moles of HCl $= \dfrac{C \times 25.0}{1000}$

$$\frac{2.0 \times 12.5}{1000} = \frac{C \times 25.0}{1000}$$

C = 1.0 moles/litre

The change of pH during titration (Titration curve)

As an acid is added to a base or a base added to acid, the pH changes. The amount of acid/ base added against the pH can be plotted to obtain titration curves. The shape of these curves depends on what is added from the burette and whether the acid / base is strong or weak.



Acid –base titration curves

Conductrimetric titration

The principle of **conductometric titration** is based on the fact that during the titration, one of the ions is replaced by the other and invariably these two ions differ in the ionic conductivity with the result that conductivity of the solution varies during the course of titration. The equivalence point may be located graphically by plotting the change in conductance as a function of the volume of titrant added.

In order to reduce the influence of errors in the conductometric titration to a minimum, the angle between the two branches of the titration curve should be as small as possible. If the angle is very obtuse, a small error in the conductance data can cause a large deviation. The following approximate rules will be found useful.

- The smaller the conductivity of the ion which replaces the reacting ion, the more accurate will be the result. Thus, it is preferable to titrate a silver salt with lithium chloride rather than with HCl. Generally, cations should be titrated with Lithium salts and anions with acetates as these ions have low conductivity

- The larger the conductivity of the anion of the reagent which reacts with the cat-ion to be determined, or vice versa, the more acute is the angle of titration curve.

- The titration of a slightly ionized salt does not give good results, since the conductivity increases continuously from the commencement. Hence, the salt present in the cell should be virtually completely dissociated; for a similar reason, the added reagent should also be strong electrolyte.

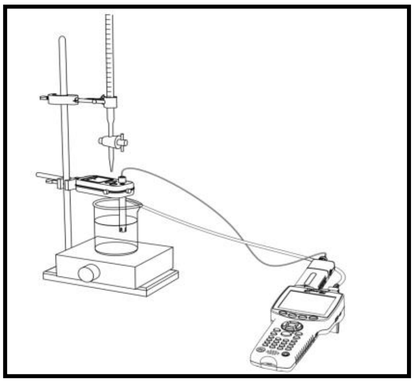
- Throughout a titration the volume of the solution is always increasing, unless the conductance is corrected for this effect, non linear titration curves result. The correction can be accomplished by multiplying the observed conductance either by total volume (V+V') or by the factor (V+ V')/V, where V is the initial volume of solution and V' is the total volume of the reagent added. The correction presupposes that the conductivity is a linear function of dilution, this is true only to a first approximation.

- In the interest of keeping V small, the reagent for the conductometric titration is ordinarily several times more concentrated than the solution being titrated (at least 10-20 times). A micro burette may then be used for the volumetric measurement. The main advantages to the conductometric titration are its applicability to very dilute and coloured solutions and to system that involve relative incomplete reactions. For example, which neither a potentiometric, nor indicator method can be used for the neutralization titration of phenol (Ka = 10–10) a conductometric endpoint can be successfully applied.

Application:
Acid-base titration, especially at trace levels. Relative precision better than 1% at all levels. There are also few disadvantages with this technique. Conductance is a non-specific property, the concentration of other electrolytes can be troublesome.

The electrical conductance of a solution is a measure of its currents carrying capacity and therefore, determined by the total ionic strength. It is a non-specific property and for these reason direct conductance measurements are of little use unless the solution contains only the electrolyte to be determined or the concentrations of other ionic species in the solution are known. Conductometric titrations, in which the species in the solution are converted to non-ionic by neutralization, or precipitation.

Equipment set-up



Conductometric titration

1) Set up the equipment as shown in the picture.
   - Use a right-angle clamp to attach the Drop Counter to the support rod. Above the Drop Counter, use a burette or three-finger clamp to affix a burette to the support rod. Ensure the burette valve is closed.
   - Place a magnetic stirrer at the base of the setup. Place the beaker you will be using for the titration on the stirrer.
   - Lower the Drop Counter until it rests just above the beaker. Then lower the burette so that the valve is about one inch above the Drop Counter opening.

2) Use a burette clamp to attach a 50mL burette to the support rod with the tip about 1 inch above the opening in the drop counter. Ensure the burette stopcock is closed.

3) Fill the burette.
   • Remove the beaker from the magnetic stirrer and replace with a waste beaker.
   • Rinse your burette with 1-2 mL of titrant. Discard as directed. Then, fill burette with ~50 mL of titrant, making sure not to fill past the top volume gradation.

Typical Conductometric Titration Curves are:

(i) Strong Acid with a Strong Base, example HCl with NaOH.
(ii) Weak Acid with a Strong Base, example acetic acid with NaOH.
(iii) Strong Acid with a Weak Base, example sulphuric acid with dilute ammonia.
(iv) Weak Acid with a Weak Base.
(v) Mixture of a Strong Acid and a Weak Acid vs. a Strong Base or a Weak Base.
(vi) Displacement (or Replacement) Titrations.

Now, check what you have just learnt by trying out the learning activity below!

Learning Activities 9 **30 minutes**

Answer the following questions.

1. A 25ml solution of 0.5M NaOH is titrated until neutralized into a 50ml sample of HNO_3. What was the concentration of the HNO_3?

2. What volume of 0.075M HNO_3 is required to neutralize 100ml of 0.01M $Ca(OH)_2$ solutions?

3. Calculate the volume of 0.20M HCl needed to neutralise 25.0cm^3 of 0.10M NaOH.

4. A student found that 25.0cm^3 of 0.80M sulphuric acid was needed to neutralise 20.0 cm^3 of sodium hydroxide solution. What is the concentration of the sodium hydroxide solution?

Thank you for completing your learning activity 9. Check your work. Answers are at the end of this module.

REVISE WELL USING THE MAIN POINTS ON THE NEXT PAGE.

SUMMARY

You will now revise this module before doing Assessment 6. Here are the main points to help you revise. Refer to the module topic if you need more information.

Common acids and bases
The word `acid' comes from Latin term `acidus', means **sour.** In 17th century, a chemist, **Robert Boyle**, first classified certain substances as either acids, or alkalis according to their characteristics.

In 1887, a Swedish scientist named **Svante Arrhenius** proposed that water could dissolve many

Mineral and organic acids
Many naturally occurring acids, such as citric acid found in oranges are sometimes known as **organic acids.**

All acids produce hydrogen ions, **H^+,** in aqueous solutions. This leads us to the definition of **acid as a substance which produces hydrogen ions, H^+, when dissolved in water.**

The following are some uses of the common acids:
- Hydrochloric acid is used in the car industry to remove rust from metals. The body of a motor car is dipped into a solution of hydrochloric acids to remove rust before it is painted.
- Sulphuric acid is used to make fertilizers and detergents. It is also used in car batteries.
- Nitric acid is used for making fertilisers and explosive called trinitrotoluene (TNT).
- Ethanoic acid is used to add flavor to food and for preserving vegetables.
- Citric acid is used in the making of `fruit salts'.

Common properties of bases
A base is any metal oxide or hydroxide. This means that a base contains either oxides ions, O^{2-}, or hydroxide ions, OH^-. Or a base is an oxide or hydroxide of metal. A base reacts with an acid to give salts and water only.

Some uses of alkalis are as follows:
- Sodium hydroxide and potassium hydroxide are used for making soaps and detergent.
- Magnesium hydroxide is used in toothpaste to neutralize the acid produced by bacteria, which feed on the sugar in our food.
- Antacid tablets contain magnesium hydroxide and carbonate, which can neutalise the excess acid in the stomach that causes indigestion.
- Ammonia solution is used in window cleaners.

- Calcium hydroxide, commonly called slaked lime, is used in agriculture to neutalise excess acids in soil.

Common acids and bases

Arrhenius, Bronsted – Lowry and Lewis definition of Acid and base.In 1887 the young Swedish chemist Svante Arrhenius proposed a revolutionary way of thinking about acids and bases. He said that acids are compounds containing hydrogen that ionize to yield hydrogen ions (H^+) in aqueous solution.

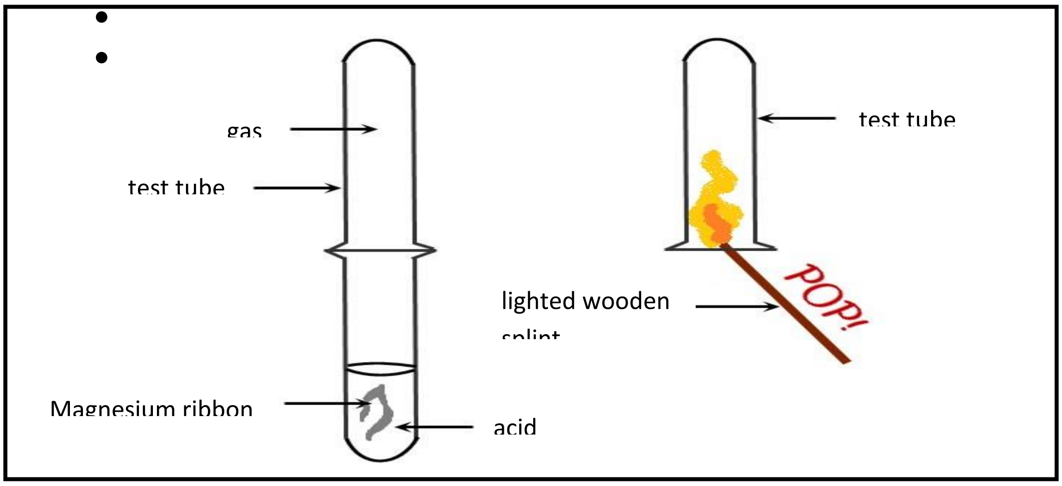
Similarly, he said that bases are compounds that ionize to yield hydroxide ion (OH^-) in an aqueous solution. that contains one ionisable proton are called monoprotic acid. Sulphuric acid (H_2SO_4) and any acids that contain two ionisable protons are called diprotic acid. Phosphoric acid (H_3PO_4) and any acids that contains three ionisable proton is tripotic acid.

In 1923 Johannes Bronsted (Danish 1879-1947) and Thomas Lowry (English 1874-1936) independently proposed a new theory. The Bronsted-Lowry theory defines **an acid as a hydrogen-ion donor**. Similarly, Bronsted-Lowry **base is hydrogen-ion acceptor**. All the acids and bases included in the Arrhenius theory are also acids and bases according to the Bronsted-Lowry theory.

A Lewis acid is a substance that can accept a pair of electrons to form a covalent bond. A Lewis base is a substance that can donate a pair of electrons to form covalent bonds.

Chemical properties of Acids and bases

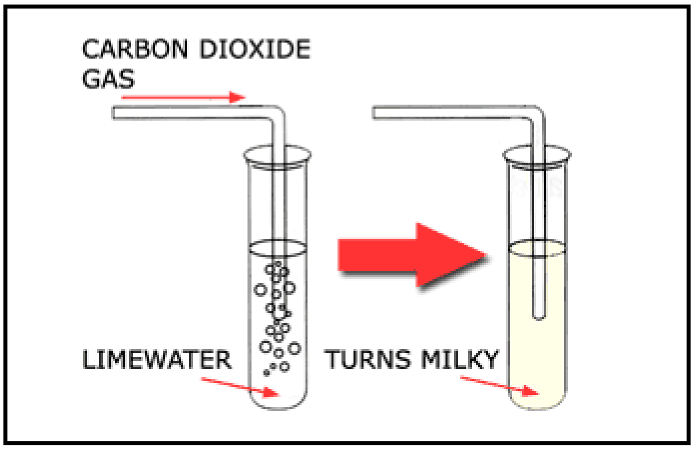
- Acids react with metals.
  Most dilute acids react with metals which are more reactive than hydrogen in the reactivity series to produce a salt and hydrogen gas.
- Test for hydrogen gas.
  To show that hydrogen gas is produced during the reaction, a burning (lighted) splint is placed near the mount of the test tube (see the figure below). A pop sound is heard and the splint is extinguished.



Test for hydrogen gas

- Acids react with carbonates.
  Many diluted acids react with carbonates to produce a salt, water and carbon dioxide gas. For example, dilute hydrochloric acid reacts with calcium carbonate or limestone chips to produce calcium chloride, water and carbon dioxi**d**e.

- Test for Carbon dioxide gas.
  To show that carbon dioxide is liberated, the gas is passed through **limewater** using a delivery tube. A white precipitate of calcium carbonate is formed and the limewater turns **milky.**



Test for carbon dioxide gas

- Acids reacts with bases.
  All acids react with bases (metal oxides or hydroxides) to form salt and water.
- The reaction between acid and base is called **neutralisation.**

Chemical properties of alkalis:

- Alkalis turn red litmus paper blue.
- **Neutralisation** – Alkalis react with acids to form salt and water.
- Alkalis react with ammonium salts to produced ammonia gas. If some sodium hydroxide solution is warned with a little ammonium chloride in a test tube, ammonia gas will be given off and a strong smell can be detected.

Test for ammonia gas

Ammonia gas can be detected by its pungent smell. (Like that of urine).

Strong and weak acids and bases

Is it true that strong acids are acids that burn the skin and weak acids are harmless? The terms **strong** and **weak** are often confused with **concentrated** and **dilute**.

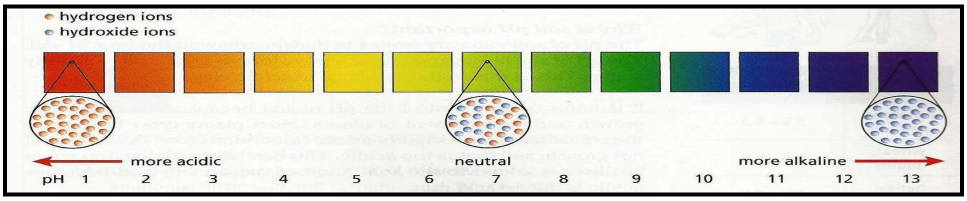
- **A strong acid is an acid which is completely ionized in water.**
- **A weak acid is an acid that is only partially ionized in water.**

Other weak acids include carbonic acid, sulphurous acid, citric acid, tartaric acid, and all other organic acids such as carboxylic acids.

The pH Scale
The concentration of the hydrogen ions in a solution can be measured with a set of numbers or scale called **pH scale.**

Universal indicator
a. A chemical compound called an indicator.
b. A pH probe attached to a data logger.

Universal indicator contains a mixture of dyes. It gives different colour in solutions of different pH.



Dilute and concentrated acids and bases

The word **concentrated** and **dilute** tell how much of an acid or a base is **dissolved in a solution.** These terms refer to the number of moles of the acid or base in a given volume.

The word strong and weak refers to the extent of **ionization of** an **acid or base**.

Amphoteric and Amphiprotic substances
Amphiprotic are substance such as water, which can either accept or donate protons.

Amphiprotic substances contain a:

- Hydrogen atom with a polar bond to a non-metal atom, so they can act as protondonors or acids.
- Non-metal atom with a lone pair of electrons, so they can act as proton acceptors or bases. Substances that neither accept nor donate protons are said to be **neutral.**

Water is an amphiprotic molecule. Water contains hydrogen atoms linked by polar bonds to oxygen and lone pairs of electrons. A water molecule can act as both a proton donor and a proton acceptor**.**

Amphoteric are solid substances that show:
- basic properties by reacting with acids.
- acidic properties by reacting with bases.

Dissociation constant and ph calculation
Calculation of acids and base constant
Strong acids are completely ionised in aqueous solution. **Weak acids** ionised only slightly in aqueous solution. The ionisation of acetic acid, a typical weak acid, is not complete.A weak acid has a small K_a value. A larger value of K_a means that dissociation, or ionisation, of the acid is more complete.

Strong bases dissociate completely into metal ions and hydroxide ions in aqueous solution. Weak bases do not dissociate completely in aqueous solution. The base dissociation constant, K_b, is the ratio of the dissociated form of a base to the undissociated form.
The equilibrium now values are now substituted into the expression for $K_{a.}$

A water molecule that loses hydrogen ions becomes negatively charged. Hydroxide ion (OH^-). A water molecule that gains a hydrogen ions becomes a positively charge hydronium ion (H_3O^+).Chemist describes pure water as neutral. Any aqueous solution in which $[H^+]$ and the $[OH^-]$ are 1.0x10-^{7} mol/L is describe as a neutral solution.

In any aqueous solution the $[OH^-]$ and the $[OH^-]$ are interdependent. If the $[H^+]$ increases then the $\left[OH^-\right]$ decreases. Or the other way around. Le Chatelier's principle applies here.

$$\mathbf{K_w = [H^+] \times [OH^-] = 1.0\times10^{-14}(mol\ /L)^{\ 2}}$$

In acidic solution the $[H^+]$ is greater than the $[OH^-]$, therefore the $[H^+]$ of acidic solution is greater than 1.0×10^{-7} mol/L.

In basic solution the $[H^+]$ is less than the $[OH^-]$. Therefore the $[H^+]$ of a basic solution is less than 1.0×10^{-7} mol/L basic solution is also known as **alkaline solution.**

pH and pOH calculation
The pH concept
The pH of a solution is the negative logarithm of hydrogen-ion concentration.

$$pH = - \log[H^+]$$

In a similar way, the pOH of a solution equals the negative logarithm of the hydroxide ion concentration.

$$\mathbf{pOH= - \log[OH^-] \text{ or } pOH = 14 – pH}$$

Acid-base titration to detect the end point
Volumetric titration
Titration is a common laboratory method of quantitative, chemical analysis that is used to determine the concentration of a known reactant (**analyte**). The basis of the method is a chemical reaction of a **standard solution (titrant)** with a solution of an analyte.The **analyte (*A*)** is a solution of the substance whose concentration is unknown and sought in the analysis.

Endpoint is the point at which the titration is complete, as determined by an **indicator** at the titration endpoint, the quantity of reactant in the titrant.

Acid-base titrations are based on the neutralization reaction between the analyte and acidic or basic titrant.

Neutralization is a chemical reaction, also called a **water forming reaction**,
which an acid and a base or alkali (soluble base) react and produce a salt and water(H2O).
acid + base → salt + water

- A strong acid reacts with a strong base to form a neutral (pH=7) solution.
- A strong acid reacts with a weak base to form an acidic (pH<7) solution.
- A weak acid reacts with a strong base to form a basic (pH>7) solution.

Monoprotic acids contain one acidic hydrogen. Titration curve of strong monoprotic acid with strong base.

Polyprotic acids contain more than one acidic hydrogen.

pH indicators are generally very complex organic molecules (frequently weak acids or bases).

Titration procedure:

1. A known volume ***VA*** of the analyte (acid) is placed in a titration flask.
2. The burette is filled by a standard solution (titrant, base) of known concentration ***cT*** (M).
3. Before the titration is started, 1-3 drops of indicator (phenolphthalein) is placed in the titration flask with the analyte (acid). The chosen indicator must be one colour when the solution is acidic.
4. A base solution is then slowly added from the burette, drop by drop.
5. The titration continues, drop by drop, until the indicator suddenly achieves the intermediate colour (weak pink) between that of the acid and the color of the base (fuchsia). At that point, the titration ceases.
6. The point at which the system is neither acidic nor basic is referred to as the endpoint. The endpoint will correspond to a perfect stoichiometric relationship between the acid and the base.
7. Once the endpoint has been reached, the burette must be read. The bottom of the meniscus line determines the quantity of the base ***VT*** that was required to reach the endpoint.
8. Once the titration is completed, the final calculations can be done.

Conductrimetric titration

The principle of conductometric titration is based on the fact that during the titration, one of the ions is replaced by the other and invariably these two ions differ in the ionic conductivity with the result that conductivity of the solution varies during the course of titration.

- The smaller the conductivity of the ion which replaces the reacting ion, the more accurate will be the result. Ions have low conductivity.
- The larger the conductivity of the anion of the reagent which reacts with the cation to be determined, or vice versa, the more acute is the angle of titration curve.
- The titration of a slightly ionized salt does not give good results, since the conductivity increases continuously from the commencement.
- Throughout a titration, the volume of the solution is always increasing, unless the conductance is corrected for this effect, non linear titration curves result.

Typical Conductometric Titration Curves are:

1. **strong acid with a strong base, example HCl with NaOH.**
   Before NaOH is added, the conductance is high due to the presence of highly mobile hydrogen ions.

2. **weak acid with a strong base, example acetic acid with NaOH.**
   Initially the conductance is low due to the feeble ionization of acetic acid.

3. **strong acid with a weak base, example sulphuric acid with dilute ammonia**.
   Initially the conductance is high and then it decreases due to the replacement of H^+.

4. **weak acid with a weak base**.
   The nature of curve before the equivalence point is similar to the curve obtained by titrating weak acid against strong base. After the equivalence point, conductance virtually remains same as the weak base, which is being added is feebly ionized and, therefore, is not much conducting.

5. **mixture of a strong acid and a weak acid vs. a strong base or a weak base**.
   In this curve there are two break points.

6. **displacement (or Replacement) titrations**.
   When a salt of a weak acid is titrated with a strong acid, the anion of the weak acid is replaced by that of the strong acid. Weak acid itself is liberated in the undissociated form.

pH value	Substance	Acid or Alkali
1	Hydrochloric aci, sulphuric acid and nitric acid	Acid
2	Lime juice and lemon	Acid
3	Vinegar, apple and pineapple	Acid
4	Soda water, carbonated drinks , banana and tomato	Acid
5	Tea	Acid
6	Piped water and fresh milk	Acid
7	Distilled water and sodium chloride solution	Neutral
8	Soap water and baking soda (baking powder)	Alkali
9	Toothpaste	Alkali
10	Milk of magnesia	Alkali
11	Washing soda	Alkali
12	Lime water (calcium hydroxide solution)	Alkali
13	Sodium hydroxide solution and potassium	Alkali
14	Hydroxide solution	Alkali

Acidic, neutral and alkaline substances used in everyday life.

NOW YOU MUST COMPLETE ASSESSMENT 2 AND RETURN IT TO THE PROVINCIAL CENTRE CO-ORDINATOR.

ANSWERS TO LEARNING ACTIVITIES 1- 9

Learning Activity 1

1. i. Mineral acid – derived from one or more mineral or inorganic compound.
   ii. organic acid – organic compounds with acidic properties.

2. i. Hydrochloric acid
   Nitric acid
   ii. nitric acid
   gallic acid

3. i. Lactic acid – $C_3H_6O_3$
   ii. gallic acid – C_7H_6O

Learning Activity 2

1. i. Soluble base that produces hydroxide ions.(OH^-)
   ii. Salts that contain water of crystallisation.
   iii. Salts that do not contain water of crystallisation.

2.

		Anhydrous	hydrated
i.	Zinc sulphate	$ZnSO_4$	$ZnSO_4.7H_2O$
ii.	Copper sulphate	$CuSO_4$	$CuSO4.5H_2O$
iii.	Magnesium sulphate	$MgSO_4$	$MgSO_4.7H_2O$

Learning Activity 3

1. i. Arrhenius acid is a compound that contains hydrogen and that ionises to form (H^+). Bases are compound that ionises to produce hydroxide ion (OH^-)

   ii. Bronsted - Lowry acid is a hydrogen ion donor. Base is hydrogen ion acceptor.

   iv. Lewis – substance that accepts a pair of electron to form covalent bond is an acid. Substance that can donate a pair of electron to form covalent bond is a base.

2. i. Contain one ionisable proton.
   ii. Contains two ionisable protons.
   iii. Contains three ionisable protons.

3. Chemical symbol — type
   i. CH_3COOH — tripotic
   ii. HNO_3 — monoprotic
   iii. H_2SO_4 — diprotic
   iv. H_2CO_3 — diprotic
   v. H_3PO_4 — tripotic

Learning Activity 4

1. i. Strong acid
   ii. weak acid
   iii. pH
2.

Solution tested	pH	symbol
Vinegar (acetic acid)	5	CH_3COOH
Sulphuric acid	1	H_2SO_4
Distilled water	7	H_2O
Sodium Hydroxide	14	NaOH

3. i. weak acid
   ii. weak base
   iii. strong acid
   iv. strong base

Learning Activity 5

1. i. $C = \frac{n}{v} = \frac{0.24}{0.50} = 0.48M$

   ii. $C = \frac{0.010}{0.20} = 0.050M$

2. i. $C = \frac{n}{v} \quad = C_1V_1 = C_2V_2$

   $1.2 \times 5 = C_2 \times 15 = 0.4M$

   ii. $C_1V_1 = C_2V_2$

   $C_2 = \frac{C_1V_1}{V_2} = \frac{1.0x3.0}{4.0} = 0.75M$

Learning Activity 6

1. i. $H_2O \longrightarrow H^+ + OH^-$
   ii. $H_2O + H^+ \longrightarrow H_3O^+$
2. i. $Zn(OH)_2$ is amphoteric – reacts acid and base.
   ii. NaOH basic ionise in water to produced OH^-.

Learning Activity 7

1 $1.8x10^{-5} = \frac{x^2}{1.999}$

$X^2 = 3.599$

$= \mathbf{6.0 \times 10^{-3}}$

2. pH $= - \log[H^+]$

$= - \log [9.86x10-4]$

$= \mathbf{3.006}$

Learning Activity 8

1. i. pH= - log [0.1]

pH = 1

ii. pH = - log [0.01]

pH = 2

pOH = 14 -2

pH **= 12**

iii. pH = - log [0.005]

= 2.3

Learning Activity 9

1. $M_1V_1 = M_2V_2$

$M_Ax50 = 0.5 \times \frac{25}{50} = \mathbf{0.25M}$

2. $M_AV_A = M_BV_B$

$M_A = 0.075 \, xV_A = 0.07 \times 100$

$V_A = 13.33 \times 2 =$ **26.6 ml**

3. $M_AV_A = M_BV_B$

$$V_A = \frac{0.10 \times 25}{0.20}$$

$$\mathbf{V_A = 12.5\ cm^3}$$

4. $H_2SO_4 + 2NaOH \longrightarrow Na_2SO_4 + 2H_2O$

$2(M_AV_A) = M_BV_B$

$2(25 \times 0.80) = M_B \times 20$

$$M_B = \frac{40}{20}$$

$$\mathbf{= 2M}$$

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FODE PROVINCIAL CENTRES CONTACTS

PC NO.	FODE PROVINCIAL CENTRE	ADDRESS	PHONE/FAX	CUG PHONE (COORDINATOR)	CUG PHONE (SENIOR CLERK)
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2	BUKA	P. O. Box 154, Buka	9739838	72228108	72229073
3	CENTRAL	C/- FODE HQ	3419228	72228110	72229050
4	DARU	P. O. Box 68, Daru	6459033	72228146	72229047
5	GOROKA	P. O. Box 990, Goroka	5322085/5322321	72228116	72229054
6	HELA	P. O. Box 63, Tari	73197115	72228141	72229083
7	JIWAKA	c/- FODE Hagen		72228143	72229085
8	KAVIENG	P. O. Box 284, Kavieng	9842183	72228136	72229069
9	KEREMA	P. O. Box 86, Kerema	6481303	72228124	72229049
10	KIMBE	P. O. Box 328, Kimbe	9835110	72228150	72229065
11	KUNDIAWA	P. O. Box 95, Kundiawa	5351612	72228144	72229056
12	LAE	P. O. Box 4969, Lae	4725508/4721162	72228132	72229064
13	MADANG	P. O. Box 2071, Madang	4222418	72228126	72229063
14	MANUS	P. O. Box 41, Lorengau	9709251	72228128	72229080
15	MENDI	P. O. Box 237, Mendi	5491264/72895095	72228142	72229053
16	MT HAGEN	P. O. Box 418, Mt. Hagen	5421194/5423332	72228148	72229057
17	NCD	C/- FODE HQ	3230299 ext 26	72228134	72229081
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19	RABAUL	P. O. Box 83, Kokopo	9400314	72228118	72229067
20	VANIMO	P. O. Box 38, Vanimo	4571175/4571438	72228140	72229060
21	WABAG	P. O. Box 259, Wabag	5471114	72228120	72229082
22	WEWAK	P. O. Box 583, Wewak	4562231/4561114	72228122	72229062

FODE SUBJECTS AND COURSE PROGRAMMES

GRADE LEVELS	SUBJECTS/COURSES
Grades 7 and 8	1. English
	2. Mathematics
	3. Personal Development
	4. Social Science
	5. Science
	6. Making a Living
Grades 9 and 10	1. English
	2. Mathematics
	3. Personal Development
	4. Science
	5. Social Science
	6. Business Studies
	7. Design and Technology- Computing
Grades 11 and 12	1. English – Applied English/Language& Literature
	2. Mathematics – General / Advance
	3. Science – Biology/Chemistry/Physics
	4. Social Science – History/Geography/Economics
	5. Personal Development
	6. Business Studies
	7. Information & Communication Technology

REMEMBER:

- For Grades 7 and 8, you are required to do all six (6) subjects.
- For Grades 9 and 10, you must complete five (5) subjects and one (1) optional to be certified. Business Studies and Design & Technology – Computing are optional.
- For Grades 11 and 12, you are required to complete seven (7) out of thirteen (13) subjects to be certified.

Your Provincial Coordinator or Supervisor will give you more information regarding each subject and course.

Notes: You must seek advice from your Provincial Coordinator regarding the recommended courses in each stream. Options should be discussed carefully before choosing the stream when enrolling into Grade 11. FODE will certify for the successful completion of seven subjects in Grade 12.

GRADES 11 & 12 COURSE PROGRAMMES			
No	**Science**	**Humanities**	**Business**
1	Applied English	Language & Literature	Language & Literature/Applied English
2	General / Advance Mathematics	General / Advance Mathematics	General / Advance Mathematics
3	Personal Development	Personal Development	Personal Development
4	Biology	Biology/Physics/Chemistry	Biology/Physics/Chemistry
5	Chemistry/ Physics	Geography	Economics/Geography/History
6	Geography/History/Economics	History / Economics	Business Studies
7	ICT	ICT	ICT

CERTIFICATE IN MATRICULATION STUDIES		
No	**Compulsory Courses**	**Optional Courses**
1	English 1	**Science Stream:** Biology, Chemistry and Physics
2	English 2	**Social Science Stream:** Geography, Intro to Economics and Asia and the Modern World
3	Mathematics 1	
4	Mathematics 2	
5	History of Science & Technology	

REMEMBER:

You must successfully complete 8 courses: 5 compulsory and 3 optional.