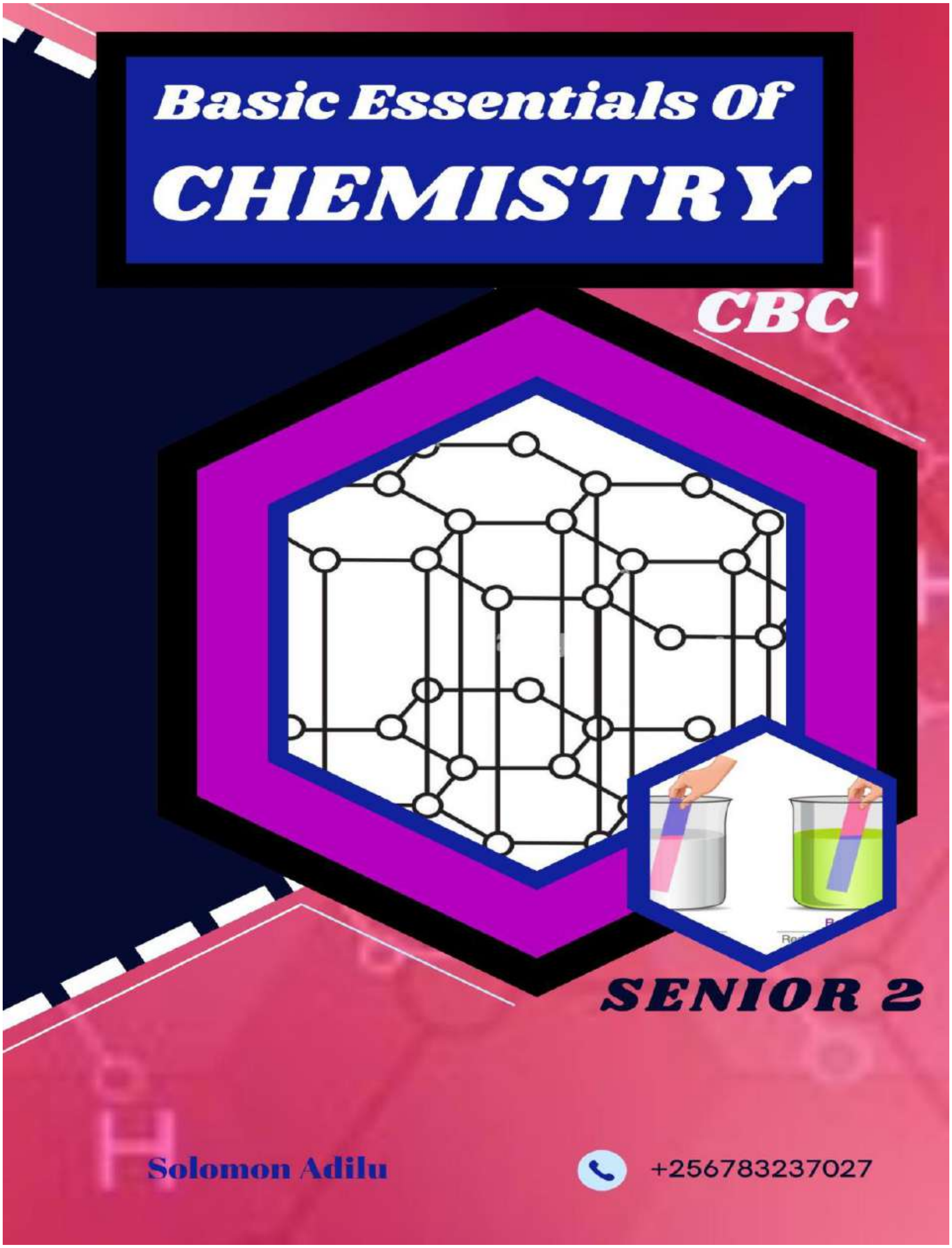
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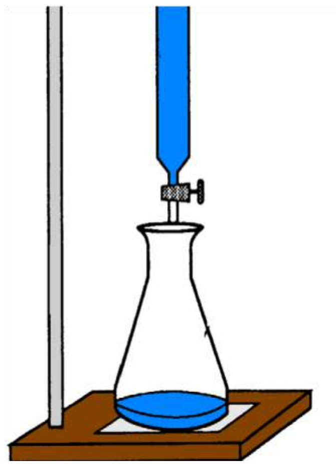

Competency Based Curriculum

TOPIC 1:

ACIDS & BASES

Competency: The learner appreciates the properties and importance of acids, alkalis/bases and salts in everyday life.

Key words

- Acid
- Base
- Alkali
- pH
- Litmus paper
- Indicator
- Hydrogen ion (H+)
- Hydroxide ion (OH-)
- Neutralization

By the end of this topic, the learner should be able to;

- Recognize that locally available materials and substances are either acidic or alkaline(k)
- Understand the concept of pH as a measure of the strength of acids and alkalis(u)
- Understand the reaction between acids and alkalis (u, s)

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1.0 Acids & Bases

Common Acids and Bases

Acids and bases are common substances in everyday life, and play crucial roles in our daily lives. They are often identified by their sour taste (acids) or slippery feel (bases), and can be found in some fruits like ***lemons*** and ***oranges*** (*citric* ***acid***) or soap and Aloe Vera (base). Most are ***safe*** to handle but some are definitely not.

Identifying substances in everyday life that are acidic or alkaline.

Note: The acidity or alkalinity of plant materials can vary depending on factors like species, growth conditions, and part of the plant used. It's always recommended to test specific samples to confirm their pH.

Materials

- ✓ Lemon
- ✓ Raw mango
- ✓ Cabbage
- ✓ Tomatoes
- ✓ Baking powder
- ✓ Aloe Vera
- ✓ Moringa leaves
- ✓ Sour milk

What to do

- ✓ Wash the materials provided
- ✓ Bite each of them, taste them, feel by touching.
- ✓ Cut the lemons using a knife and taste.
- ✓ Wash your mouth every after tasting a fruit.
- ✓ Record your findings in the table

Expected responses

Material	Taste	Feel	Acid / Alkaline
Lemon	Sour	Not slippery	Acid
Raw Mango	Sour	Not slippery	Acid
Cabbage	Sour	Not slippery	Acid
Tomatoes	Sour	Not slippery	Acid
Baking powder	Bitter	Slippery	Alkaline
Aloe Vera	Bitter	Slippery	Alkaline
Banana peelings	Bitter	Slippery	Alkaline
Moringa leaves	Bitter	Slippery	Alkaline
Sour milk	Sour	Not slippery	Acid

Acids

Acids taste sour - the ***Latin*** word for ***sour*** was "***acidus***"

Definition: Substances that, when dissolved in water, produce hydrogen ions (H+).

Characteristics:

- ✓ Typically sour (e.g., lemon juice, vinegar)
- ✓ React with many metals to produce hydrogen gas.
- ✓ Turn blue litmus paper red.
- ✓ Have a pH value less than 7.

Common Acids

Fizzy drinks (*CO_2,carbonic **acid***), ***vinegar*** (*ethanoic **acid***), and even many ***vegetables*** (*ascorbic **acid**, vitamin C*) contain ***acids***. These are examples of weak acids. A more unpleasant example is a ***bee*** sting, ***nettle*** sting or ***ant*** bite (*formic acid*).

Stronger acids are found in car ***batteries*** (*sulphuric acid*) while our ***stomachs*** contain acid (*hydrochloric*) to break down our ***food*** into ***smaller*** molecules.

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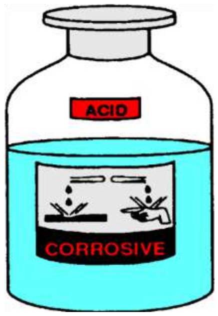


Note:

Acids kill cells. - the most vulnerable part of you is your eyes, because they have living cells on the surface(ensure you wear Safety Glasses when dealing with acids)

Some common acids and their sources

Acid	Sources
Hydrochloric acid	Stomach, industrial processes, cleaning products.
Sulphuric acid	Battery acid, industrial processes, fertilizer
Nitric acid	Fertilizers, explosive, industrial processes
Acetic acid (ethanoic acid)	Vinegar, food, cleaning products
Citric acid	Citrus fruits, soft drinks, food preservatives
Carbonic acid	Carbonated beverages, respiratory system.
Oxalic acid	Spinach, cleaning products, Rhubarb.
Phosphoric acid	Fertilizers, soft drinks, toothpaste.
Tartaric acid	Grapes, wine, baking powder
Lactic acid	Milk, yogurt, muscle tissues
Ascorbic acid	Vitamin C, citrus fruits.

Bases & Alkalis

<u>A base</u> is a metal oxide, hydroxide or carbonate that reacts with an acid to form a salt and water.

Our main source of *bases* are the ***Metal Oxides***.

e.g, *lime* (***calcium oxide***), *soda* (***sodium oxide***), *magnesia* (***magnesium oxide***), *pearl ash* (***potassium oxide***), though, we also use many ***Metal Carbonates***.

e.g, *limestone* (***calcium carbonate***), *marble* (***calcium carbonate***), *baking soda* (***sodium carbonate***) *or potash* (***potassium carbonate***)

Bases are mainly metal oxides and metal carbonates.

All of these *bases* are ***ionic compounds*** and, therefore, ***solids*** at *room temperature*. As ***solids***, they can be directly added to an acid and will ***neutralise*** the acid.

base	**+**	***acid***	→	***salt***	**+**	***water***

Sometimes, however, we prefer to use ***solutions*** of *soluble bases* which we then can call ***alkalis***.

<u>Alkalis</u> are soluble bases that produce hydroxide ions (OH-) when dissolved in water.

<u>Examples</u>

- ✓ Calcium hydroxide ($Ca(OH)_2$)
- ✓ Sodium hydroxide (NaOH)
- ✓ Potassium hydroxide (KOH)
- ✓ Aqueous ammonia (NH_4OH) also known as ammonia solution.

Most ***alkalis*** are made by ***dissolving*** a ***metal oxide*** in ***water*** - though only those in ***Group 1*** are very ***soluble***. (*Data Booklet*). These oxides react with water to produce ***hydroxide ions***, $OH^-_{(aq)}$

Characteristics of Alkalis

- ✓ Typically bitter (e.g., soap solutions)
- ✓ Often feel slippery or soapy to the touch.
- ✓ Turn red litmus paper blue.
- ✓ Have a pH value greater than 7.
- ✓ Some alkalis (Sodium hydroxide and Potassium hydroxide) are deliquescent.

Some common alkalis/bases with their sources

Base	Sources
Sodium hydroxide	Soap, paper, textile industries, Mineral deposits.
Potassium hydroxide	Battery manufacturing, cleaning products, Plant ashes
Calcium hydroxide	Cement, water treatment, Mineral deposits
Magnesium hydroxide	Antacids, laxatives, fireproofing, Seawater
Ammonia	Fertilizers, cleaning products, pharmaceuticals.
Sodium carbonate	Water softening, paper, Glass manufacturing
Sodium bicarbonate	Baking, cleaning products, pharmaceuticals, toothpaste.
Potassium carbonate	Glass manufacturing, soap, pharmaceuticals
Barium hydroxide	Pesticides, oil refining, water treatment.

Other sources may include Milk of magnesia, Aloe Vera plant extract.

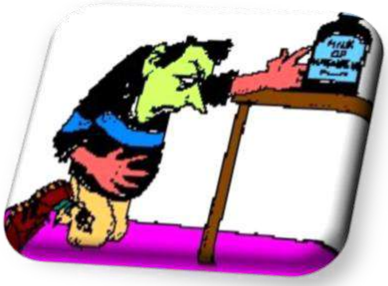
Note: All acids contains at least one hydrogen atom. When in aqueous solution, these hydrogen atoms are released as hydrogen ions making the solution behave like an acid.

All alkalis release hydroxide ions when dissolved in water. It is these hydroxide ions that make the substance behave like an alkali in an aqueous solution.

Because alkalis contain hydroxide ions, toothpaste is used for brushing teeth. ***Toothpaste*** contains a ***base*** to help neutralise the ***acid*** on your ***teeth***, produced by ***bacteria*** that remains.

Most ***soaps*** and ***detergents*** contain bases to help cope with ***greasy*** and ***oily*** stains. Our ***liver*** produces ***bile*** (a *base*) to help break down ***fatty*** foods. ***Farmers*** and ***gardeners*** will spread ***lime*** (*calcium hydroxide*) on the ***soil*** if it is too ***acidic***.

Acidic fumes (SO_2 and CO_2) from ***coal*** burning ***power*** stations are passed through ***lime*** to be neutralised.



The ***human*** stomach produces ***hydrochloric*** acid to help the ***enzyme*** (***catalyst***) *pepsin* to break down ***protein***. Sometimes too much ***acid*** is made and it begins to attack the stomach ***wall*** causing ***pain***. All stomach remedies contain ***bases*** to ***neutralise*** the stomach acid.

Note: All alkalis are bases, but Not all bases are alkalis.

Activity	Int2	SC
Acids and Bases are found everywhere around us. For each of the substances listed below, decide whether they are acids ***(A)*** or bases ***(B)***. *lemon juice* *toothpaste* *lime* *wasp stings* *vitamin C* *Coca* *Cola tomato ketchup* *nettle* *sting* *bleaches baking soda* *stomach* *juices* *detergents*	Q2.	
	S	

Indicators

Imagine you have a secret code that changes colour to reveal a message. Acid-base indicators work similarly!

They are special substances that change color with pH when they come into contact with an acid or a base. This color change helps us identify whether a solution is acidic, basic, or neutral (neither acidic nor basic). An indicator has variable colors in acidic and basic media/solutions.

Did you know that plant extracts can actually act as indicators?

You can even make an indicator at home using red cabbage! It changes color depending on the acidity or basicity of the solution.

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Experiment: Obtaining an Indicator from Local Plant Extracts

Materials

- ✓ Plant material (e.g., red cabbage, hibiscus flowers, turmeric)
- ✓ Mortar and pestle
- ✓ Beaker
- ✓ Filter paper
- ✓ Test tubes
- ✓ Acidic and alkaline solutions (e.g., vinegar, baking soda)

Procedure

- ✓ Grind the plant material into a fine powder using a mortar and pestle.
- ✓ Place the grinded plant material in a beaker and add a small amount of water.
- ✓ Stir the mixture well and let it steep for a few minutes.
- ✓ Filter the mixture through filter paper to obtain a clear extract.

 Red cabbage	 Grinding the leaves	 Effect of indicator extract on acids & bases

Test the extract

- ✓ Divide the extract into two test tubes.
- ✓ Add 2-3 drops of an acidic solution (e.g., lemon juice) to one test tube and an alkaline solution (e.g., baking soda) to the other.
- ✓ Observe any color changes in the extract.

Note:

Many red fruits and vegetables contain natural pigments called **anthocyanins**. These pigments change color depending on the acidity or alkalinity of the solution.

Observations

Acidic Condition (Lemon Juice)

Deep purple color of the red cabbage juice extract shifted to reddish-pink (pH 6-5) and then to deep red when more lemon juice drops are added.

Basic Condition (Baking Soda Solution)

Deep purple color of the red cabbage juice extract shifted to greenish-yellow (pH 8-9)

Conclusion

Red cabbage juice serves as a natural pH indicator, exhibiting color changes in response to acidic and basic conditions.

The color change from purple color towards red or orange clearly indicates an increase in acidity. This demonstrates that lemon juice is an acidic substance due to the presence of citric acid.

Anthocyanins are pH-sensitive pigments. In alkaline solutions, they tend to be blue or purple. As the solution becomes more acidic (due to the addition of lemon juice), the anthocyanins change color to red or orange.

Common commercial acid-base indicators

- ✓ Litmus paper: This is one of the most well-known indicators. It comes in two colors, Red litmus paper turns blue in alkaline solutions, and blue litmus paper turns red in acidic solutions.
- ✓ Phenolphthalein indicator, Colourless in acidic solutions, pink in alkaline solutions.
- ✓ Methyl orange indicator, Red in acidic solutions, yellow in alkaline solutions.
- ✓ Common Acid-Base Indicators:

Experiment: Effects of Substances on Litmus Paper

Safety Precautions

- ✓ Handle acids and bases with care.
- ✓ Avoid contact with your skin and eyes.
- ✓ Work in a well-ventilated area.

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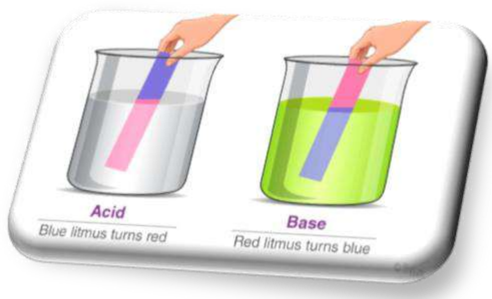
Materials

- ✓ Litmus papers (Red and Blue)
- ✓ Various solutions
  - Lemon juice
  - Vinegar
  - Baking powder
  - Ammonia
  - Diluted liquid soap
  - Tap water

Procedure

- ✓ Measure $3cm^3$ of baking powder solution in a test-tube.
- ✓ Cut small strips of red and blue litmus paper.
- ✓ Dip a strip of red litmus paper into the solution above and observe any color changes. Repeat with blue litmus paper.
- ✓ Record the color changes in a table.
- ✓ Repeat the procedures above using other substances.



Table

[redacted]			
Lemon juice	Red	Red	Acidic
Vinegar	Red	Red	Acidic
Baking powder	Blue	Blue	Alkaline
Ammonia	Blue	Blue	Alkaline
Tap water	Red	Blue	Neutral
Diluted liquid soap	Blue	Blue	Alkaline

Discussion

Litmus paper effectively indicates acidity/alkalinity of solutions. It changes color in response to pH:

- ✓ Acidic (pH < 7): Red
- ✓ Neutral (pH 7): No change
- ✓ Alkaline (pH > 7): Blue

Strength of Acids and Base

The **strength** of **acids** and **bases** can best be determined by obtaining their pH which can

be done using a pH **meter** or using a universal indicator, which displays various colors depending on the nature of the solution, the color is then compared with colors on the **pH scale** and **pH value** read. The pH scale shows the intensity of hydrogen and hydroxide ions in solution, the smaller the pH value, the more acidic the solution is, and the bigger the PH value, the more alkaline the solution is.

PH meter: A pH meter can provide a more accurate measurement of pH.

Remember: The acidity or alkalinity of plant materials can affect their taste, nutritional value, and medicinal properties. Understanding their pH can be helpful in various applications, such as food preparation, traditional medicine, and agriculture.

Universal Indicator

This is mixture of several indicators that change color over a wide pH range. It's available either in solution form or paper form.

It's used to determine the approximate pH of a solution.

The pH scale

Measures the concentration of hydrogen ions in a solution. The pH scale ranges from 0 to 14.

Neutral solutions have a pH which ***equals 7***, pH = *7*.
Neutral solutions have ***equal*** amounts of H^+ ions and OH^- ions

Acid solutions have a pH which is ***less than 7***, pH < *7*.
Acid solutions have ***more*** H^+ ions than OH^- ions, $H^+ > OH^-$

Alkali solutions have a pH which is ***more than 7***, pH > *7*.
Alkali solutions have ***more*** OH^- ions than H^+ ions, $OH^- > H^+$

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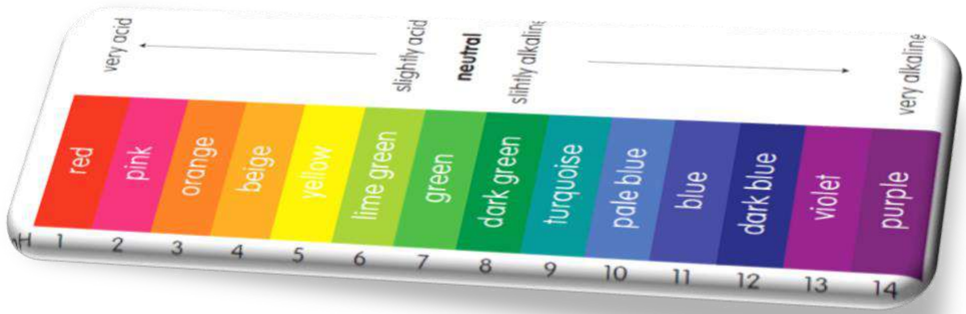
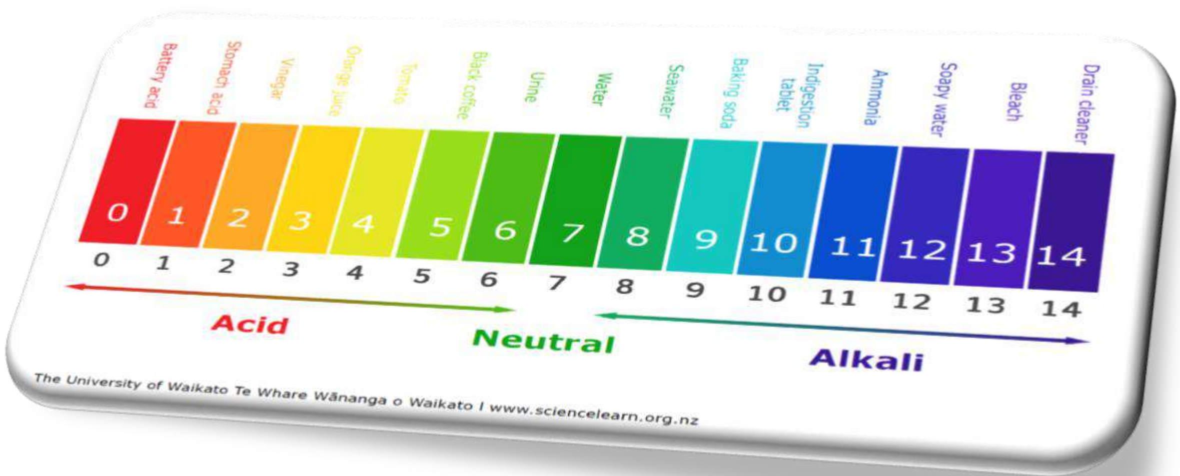
If you ***dilute*** an ***acid*** by adding more ***water***, the pH will ***increase*** towards pH = *7*.

If you ***dilute*** an ***alkali*** by adding more ***water***, the pH will ***decrease*** towards pH = *7*.





Note:

Strong acids completely ionize in aqueous solutions, producing a high concentration of hydrogen ions (H^+). **Examples**: hydrochloric acid (HCl), nitric acid (HNO_3), sulphuric acid (H_2SO_4). The higher the concentration of hydrogen ions the more acidic the solution is and the smaller the pH value.

Weak acids only partially ionize in aqueous solutions, producing a low concentration of hydrogen ions. **Examples**: acetic acid (CH_3COOH), carbonic acid (H_2CO_3), citric acid and their resulting pH value will be smaller but nearing to 7.

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Strong bases: Completely ionize in aqueous solutions, producing a high concentration of hydroxide ions (OH^-). **Examples**: sodium hydroxide (NaOH), potassium hydroxide (KOH), calcium hydroxide ($Ca(OH)_2$).Their pH value will be above 7 nearing to 14.

Weak bases only partially ionize in aqueous solutions, producing a low concentration of hydroxide ions. **Examples:** ammonia (NH_3), carbonate ions (CO_3^{2-}), bicarbonate ions (HCO_3^-). The greater the concentration of hydroxide ions the greater the alkalinity and the higher the pH value.

Experiment: Determining the pH of Solutions

Materials

- ✓ Universal indicator solution
- ✓ Test tubes
- ✓ Dropper
- ✓ Various substances to be tested (e.g., lemon juice, vinegar, baking soda, ammonia, water).

Procedure

- ✓ Measure 3cm^3 of baking powder solution in a test-tube
- ✓ Add a few drops of universal indicator solution to the test tube containing baking powder solution.
- ✓ Observe the color changes in the solutions and record in a table.
- ✓ Repeat the steps above using other substances provided.
- ✓ Use a pH chart or a pH meter to determine the approximate pH based on the color changes.

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Expected Observations

Table

Substance	Color in universal indicator	Approximate pH value
Lemon juice	Red	2-3
Vinegar	Red-Orange	2.4-3.4
Baking soda	Blue - Green	8-9
Ammonia	Blue	11-12
Water	Green	7

Universal indicator will turn red or orange in acidic solutions, blue or purple in alkaline solutions and green or yellow in neutral solutions.

Relationship between Acid Strength and Base Strength

- ✓ The stronger an acid, the weaker its conjugate base.
- ✓ The stronger a base, the weaker its conjugate acid.

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pH of some common substances

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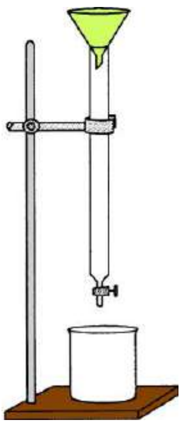
Reactions of Acids

With Alkalis



One of the most characteristic reactions of acids and bases is their interaction with each other, known as a neutralization reaction and with other substances like metals.

Alkalis are solutions which contains hydroxide ions. Most alkalis are made by dissolving a metal oxide in water- though only those in group 1 are Soluble.

Neutralization Reaction

When an acid reacts with a base, they essentially "cancel out" each other's properties. The hydrogen ions (H+) from the acid combine with the hydroxide ions (OH-) from the base to form water (H_2O). The remaining ions form a salt.

General Equation:

Acid + Base → Salt + Water

Experiment: Reactions of Acids and Bases (Neutralization reactions)

Safety Precautions

- ✓ Handle acids and bases with care.
- ✓ Avoid contact with your skin and eyes.
- ✓ Work in a well-ventilated area.
- ✓ Dispose of waste materials properly.

Materials

- ✓ Test tubes
- ✓ Droppers
- ✓ Universal indicator
- ✓ Acidic and alkaline solutions (e.g., hydrochloric acid, lemon juice, sodium hydroxide)
- ✓ Other substances (e.g., baking powder solution)

Procedure

- ✓ Measure 3cm^3 of lemon juice and baking powder solutions into separate test tubes.
- ✓ Add a few drops of universal indicator to each acidic and alkaline solution separately. Observe the color changes to confirm their pH.

- ✓ Measure fresh equal amounts of lemon juice and baking powder solution in a test tube and mix. Add few drops of universal indicator and observe any color changes.
- ✓ Repeat the procedures above using dilute hydrochloric acid and sodium hydroxide.

Expected Observations

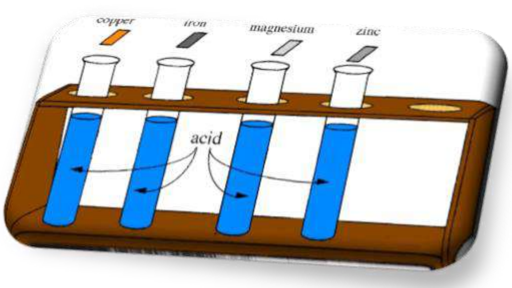
When an acid and a base are combined in equal amounts, they neutralize each other, forming a salt and water. The universal indicator will change color to indicate a neutral pH.

Acids and bases react with each other in neutralization reactions to form salts and water. Lemon juice contains a weak acid (citric acid) which has a pH of about 3. On adding baking powder solution, the resultant solution is neutral in pH and it will have a pH of around 7.

Hydrochloric acid (HCl) reacts with sodium hydroxide (NaOH) to produce a neutral solution of sodium chloride (NaCl) in water.

HCl (acid) + NaOH (base) → NaCl (salt) + H_2O (water)

With Metals



When acids react with metals, they undergo a chemical reaction that typically produces two main products:

- Salt: A salt is an ionic compound formed when a metal cation (positively charged ion) combines with a non-metal anion (negatively charged ion).
- Hydrogen Gas (H_2): This is a colorless, odourless, and flammable gas that is released during the reaction.

The gas produced ***burns*** with a ***squeaky -pop***. This shows that the gas is ***hydrogen.***

Reactive metals are able to force ***hydrogen ions*** to change back to ***hydrogen atoms***. This allows the ***metal*** to take the place of the ***hydrogen*** and form a new ***substance*** called a ***Salt***.

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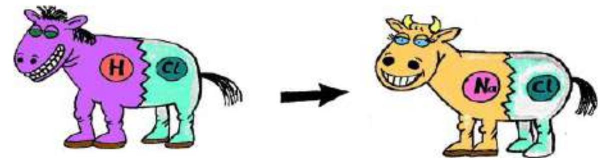
For example, when hydrochloric acid is reacted with sodium metal, The ***sodium ion*** takes the place of the ***hydrogen ion*** to form the salt called ***sodium chloride***.



Each acid has its own salts:-

Hydro***chloric*** acid, HCl → ***chlorides*** e.g. ***sodium chloride***, NaCl, ***Sulphuric*** acid, H_2SO_4 → ***sulpha***tes e.g. ***copper sulphate***, $CuSO_4$, ***nitr***ic acid, HNO_3 → ***nitrates,*** e.g. ***potassium nitrate***, KNO_3

General Reaction:

The general equation for the reaction of an acid with a metal can be represented as:

Acid + Metal → Salt + Hydrogen gas

Experiment: Reaction of Dilute Acids with Metals

Aim: To investigate the reactions of different metals with dilute acids and observe the rate of hydrogen gas evolution.

Materials:

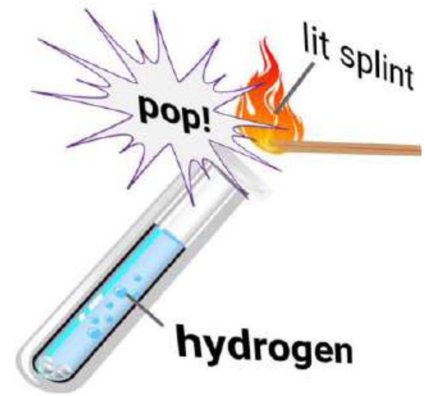
- Dilute hydrochloric acid (HCl)
- Test tubes
- Test tube rack
- Spatula
- Magnesium ribbon (Mg)
- Zinc granules (Zn)
- Iron nails (Fe)
- Lead (Pb)
- Copper wire (Cu)
- Stopwatch
- Lighted splint

Procedure:

- In groups of four, Label four test tubes as Mg, Zn, Fe, and Cu.
- Using a dropper, add 5 mL of dilute HCl to each test tube.
- Add a small piece of each metal (Mg, Zn, Fe, and Cu) to the corresponding test tube.

- Observe the reaction in each test tube, noting the following:

  Rate of reaction: Fast, moderate, or slow.
  Evolution of gas: Presence or absence of bubbles.
  Temperature change: Any heating or cooling of the test tube.
  Test for hydrogen gas: Bring a lighted splint to the mouth of each test tube.
- Repeat the experiment with dilute Sulphuric acid (H_2SO_4) for comparison.



Results Table:

Metal	Dilute hydrochloric acid	Dilute Sulphuric acid
Magnesium (Mg)	Fast reaction, vigorous bubbling, heating, "pop" sound	Fast reaction, vigorous bubbling, heating, "pop" sound
Zinc (Zn)	Moderate reaction, steady bubbling, slight heating, "pop" sound	Moderate reaction, steady bubbling, slight heating, "pop" sound
Iron (Fe)	Slow reaction, slow bubbling, slight heating, "pop" sound	Slow reaction, slow bubbling, slight heating, "pop" sound
Lead (Pb)	Slow reaction, slow bubbling, slight heating, "pop" sound, the reaction stops and doesn't go into completion	Slow reaction, slow bubbling, slight heating, "pop" sound, the reaction stops and doesn't go into completion
Copper (Cu)	No reaction, no bubbles, no temperature change, no "pop" sound	No reaction, no bubbles, no temperature change, no "pop" sound

The reactivity of metals with dilute acids varies. More reactive metals react faster and more vigorously, producing hydrogen gas. A "pop" sound indicates the presence of hydrogen gas.

Not all metals react with all acids. The reactivity of a metal with an acid depends on the position of the metal in the reactivity series. More reactive metals like magnesium, zinc, and iron readily react with dilute acids like hydrochloric acid and Sulphuric acid. Less reactive metals like copper and silver do not react with dilute acids.

For lead, the reaction only takes place slow for a little time and the reaction stops. This is due to formation of a protective layer of an insoluble salt which covers the surface of the metal protecting the metal from further reaction with the acid.

Acids react with metals leading* to corrosion** - ***iron objects, like the ***Forth** Rail Bridge*, Rust much faster these days because of acid rain.

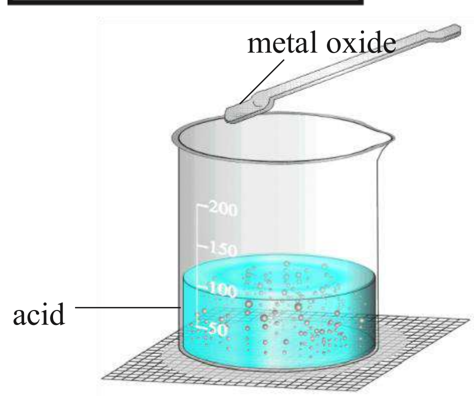
Equation:

Hydrochloric acid + Zinc metal → Zinc Chloride salt + Hydrogen gas

Task;

Write equations (word) for the reaction of acids with other metals

With Metal Oxides



Metal oxides are compounds formed when metals react with oxygen. They typically exhibit basic properties. In ***metal oxides*** the ***metal*** has already formed an ***ion*** and will not ***react*** any more.

The ***oxide ion***, O^{2-}, reacts with the ***hydrogen ions*** in the ***acid*** to form ***water***, H_2O.

This leaves the ***metal ion*** to take the ***hydrogen ions*** place, so again a ***salt*** will be produced.

The general equation for the reaction of a metal oxide with a dilute acid can be represented as:

Metal Oxide + Dilute Acid → Salt + Water

Experiment: Reaction of Dilute Acids with Metal Oxides

Aim: To investigate the reactions of different metal oxides with dilute hydrochloric acid and observe the formation of salts and water.

Materials:

- Dilute hydrochloric acid (HCl)
- Test tubes
- Test tube rack
- Spatula
- Copper(II) oxide (CuO)
- Magnesium oxide (MgO)
- Zinc oxide (ZnO)
- Lead(II) oxide (PbO)
- Dropper
- Universal indicator Paper

Safety precautions

- Lead compounds are toxic. Handle with extreme caution and dispose of properly.
- Always wear appropriate safety gear, such as safety goggles and gloves.

Procedure:

- In groups of three, Label three test tubes as CuO, MgO, and ZnO.
- Place a small amount (approximately 0.5 g) of each metal oxide (CuO, MgO, and ZnO) into the respective labelled test tubes.
- Add 2-3 mL of dilute HCl to each test tube using a dropper.
- Warm the mixture gently.
- Observe the reaction in each test tube, noting the following:

Any changes in the color of the solution or the metal oxide.

Presence or absence of gas bubbles.

After the reaction, dip a piece of universal indicator paper into each solution. Record the observed color and the corresponding pH value.

Results Table:

Metal oxide	Observations	pH of solution
Copper(II) Oxide (CuO)	Black solid dissolves in acid, forming a blue-green solution	2-3 (Acidic)

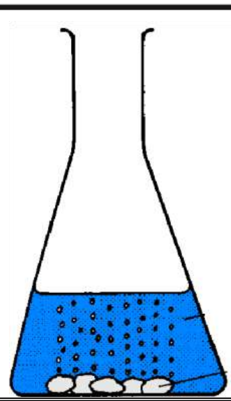
Magnesium Oxide (MgO)	White solid dissolves in acid, forming a colourless solution. Slight heating observed.	2-3 (Acidic)
Zinc Oxide (ZnO)	White solid dissolves in acid, forming a colorless solution.	2-3 (Acidic
Lead(II) Oxide (PbO)	Yellow solid dissolves in acid, forming a colorless solution (lead(II) chloride). Lead(II) chloride is slightly soluble in cold water but more soluble in hot water. This can lead to the formation of a white precipitate of lead(II) chloride as the solution cools down.	2-3 (Acidic)

Note:

The reaction between a basic metal oxide and an acid produces a salt and water, a neutral solution. However, If there's an excess of acid in the reaction, the resulting solution will still be acidic. The indicator paper will show a color corresponding to a pH below 7.

To achieve a truly neutral solution (pH 7), the amounts of metal oxide and acid must be precisely measured and mixed in the correct ratio to ensure complete neutralization. Neither the acid nor the metal oxide should be in excess.

With Metal Carbonates



In ***metal carbonates*** the ***metal*** has already formed an ***ion*** and will not ***react*** any more.

The ***carbonate ion***, CO_3^{2-}, reacts with the ***hydrogen ions*** in the ***acid*** to form ***water***, H_2O and ***carbon dioxide*** gas, CO_2.

This leaves the ***metal ion*** to take the hydrogens place, so again a ***salt*** will be produced.

Acid	*+ metal carbonate*	→	*salt*	*+*	*water*	*+ carbon dioxide*

e.g

Dilute hydrochloric acid + Magnesium Carbonate → Magnesium Chloride+ Water + Carbondioxide

Acids react with carbonates - ***acid*** *rain* is destroying many of the ***marble*** (*calcium carbonate*) ***statues*** in the world

Summary of the Effect of acids on substances

Substance	Effect when Treated with acid
Metals	Acids react with metals producing a salt and hydrogen gas but the reactivity varies. **Metal + Dilute acid →Salt + Water**
Metal Oxides	Reactive metal oxides like Zinc oxide will react with dilute acids, producing a metal salt and water. Less reactive metal oxides, Copper (II) oxide and iron (III) oxide may react with dilute hydrochloric or sulphuric acid, but they may not react with dilute nitric acid. This because nitric acid is a strong oxidizing agent, it oxidizes the metal to its highest oxidation state rather than the metal displacing hydrogen from it. **Metal oxide + Dilute acid → Salt + Water**
Metal Carbonates	The solids dissolve with effervescence of a colorless gas and forms colorless solution. The products are carbon dioxide gas, water, and a metal salt. **Metal Carbonate + Dilute acid → Salt + Water+ Carbondioxide**

Note
In case a clear solution is formed after the reaction of a base and an acid, that means the salt formed is soluble. If a precipitate is formed after the reaction that means the salt formed is an insoluble salt.

Applications of Neutralization Reactions

Neutralization reactions, where an acid and a base react to form a salt and water, have numerous practical applications across various fields. Here are some key examples:

Antacids

Our stomachs naturally contain hydrochloric acid. Hydrochloric acid in our stomach is very crucial in aiding digestion. However, when it accumulates in the stomach, it causes indigestion, burning the lining of the stomach resulting into peptic ulcers.

Antacids contain substances like calcium carbonate, magnesium hydroxide, or aluminium hydroxide. These substances neutralize excess stomach acid, providing relief from heartburn, indigestion, and acid reflux.

Baking powder is used in curing bee stings and ant bites since they contain a base.

Soil Treatment

- ✓ Adjust soil pH to optimize plant growth.

Acidic soils can be neutralized with lime (calcium carbonate) to increase pH. Alkaline soils can be neutralized with sulphur or aluminium sulphate to decrease pH.

Water Treatment

- ✓ Remove impurities from water.
- ✓ Acidic water can be neutralized with lime or sodium hydroxide.
- ✓ Alkaline water can be neutralized with Sulphuric acid or hydrochloric acid.

Chemical Synthesis

- ✓ Produce salts used in various applications. Example: Neutralization of hydrochloric acid with sodium hydroxide produces sodium chloride (table salt).

Food Processing

- ✓ Adjust the pH of food products. Example: Neutralization of acidic fruit juices with sodium hydroxide to improve flavor.

Pharmaceutical Industry

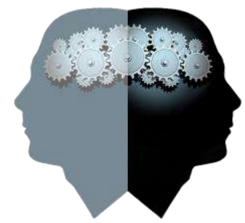
- ✓ Prepare and purify drugs. Neutralization reactions are used in the synthesis of many pharmaceuticals.

Environmental Remediation

- ✓ Neutralize acidic or alkaline pollutants in soil or water.
- ✓ Help restore ecosystems affected by pollution.

These are just a few examples of the many applications of neutralization reactions in our daily lives and industries.

Knowledge Met in this Topic



Common household acids and bases

- ***Acids:*** vinegar, citrus fruits, cola drinks etc.
- ***Bases:*** lime, oven cleaner, bleach, bicarbonate of soda, soap, ammonia

The pH scale

pH scale is a number scale that shows how acidic or alkaline a *solution* is and runs from below 0 to above 14.

- *Universal indicator*, *pH paper* or a *pH meter* can show the pH of a solution.
- ***Acids,*** pH less than 7, $pH < 7$. ***Neutral,*** pH equals 7, $pH = 7$. When *acids* dissolve in water they produce *extra hydrogen ions*, $H^+_{(aq)}$. An acid solution contains more hydrogen ions than hydroxide, $H^+(aq) > OH^-(aq)$.
- ***Alkalis,*** pH more than 7, $pH > 7$. When *bases* dissolve in water they produce *extra hydroxide ions*, $OH^-_{(aq)}$. An alkali solution contains less hydrogen ions than hydroxide, $H^+(aq) < OH^-(aq)$
- Pure water and all *neutral* solutions contain a tiny but *equal* concentration of hydrogen and hydroxide ions, $H^+_{(aq)} = OH^-_{(aq)}$
- ***Diluting*** acids or alkalis will reduce the concentration of either $H^+_{(aq)}$ or $OH^-_{(aq)}$ ions and ***move the pH towards 7***.

Neutralization

- ***Neutralization*** *is a reaction in which a base reacts with an acid to form water. A salt is also formed in this reaction.* In ***Neutralization reaction,*** *the pH of a solution moves towards 7.* Neutralizers which can *be used to neutralize an acid include the following bases (insoluble solids) and alkalis (soluble solutions):*

solid* metal oxides *(base),

metal oxide + an acid ⟶ salt + water

soluble metal oxides / metal hydroxide ***solutions*** (***alkali***) ,

metal hydroxide + an acid ⟶ salt + water

soluble metal carbonates (***alkali***) / ***insoluble*** metal carbonates (***base***).

metal carbonate + an acid ⟶ salt + water + CO_2

Everyday examples of neutralization

- ***Lime*** (***calcium oxide***) is used to ***reduce acidity*** in soil and water.
- Cures for ***acid indigestion*** contain neutralisers such as ***calcium carbonate***.
- ***Wasp stings*** *(base)* can be neutralised using ***vinegar*** *(acid).*
- ***Bee stings*** *and* ***Ant Bites*** *(acid)* can be neutralised using ***baking soda*** *(base).*

End of chapter Scenarios

<u>Item 1:</u>

At Ocokican high school, during a chemistry practical lesson, students were investigating the nature of solutions obtained from different fertilizers used by Tukum residents by identifying their PH values. The residents had complained about poor yields of their crops yet they had used some fertilizers constantly. The results obtained were recorded in the table below.

Solutions	V	W	X	Y	Z
PH values	8.2	7.0	2.0	11.2	5.6

Task

a. Classify the fertilizer solutions above depending on their nature.
   i. Nature of W
   ii. Nature of X
   iii. Nature of V
   iv. Nature of Y
   v. Nature of Z

b. Name the ions that determine the PH values of fertilizer solutions X and Y.
   i. X
   ii. Y

c. During an investigation, sodium carbonate was added to solution X and Z. the students observed that solution X liberated more effervescence than solution Z. As a chemistry student, explain this observation. (include equation for the reaction).

d. Discus the reaction of the following substances with dilute acids (hydrochloric acid) and write the equations of reactions.
   i. Magnesium metal.
   ii. Copper (II) oxide.
   iii. Calcium carbonate
   iv. Copper metal.

Item 2:

The Chemistry Club at Makerere University has received two solutions, A and B, with their behaviour on litmus papers. As a chemistry student, you must categorize these solutions as acidic, basic, or neutral.

Solution A:

- Turns blue litmus paper red

- Turns red litmus paper no change

Solution B:

- Turns red litmus paper blue

- Turns blue litmus paper no change

Task;

a. Categorize solutions A and B as acidic, basic, or neutral based on their behaviour on litmus papers.
b. Suggest examples of substances that behave the same way as solutions A and B.
c. What are the applications of these substances on everyday life.
d. Name the type of reaction that occurs between solution A and solution B.

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Competency Based

Curriculum

TOPIC 2:

SALTS

Competency: The learner appreciates that acids and alkalis form salts.

Key words

- Neutralization
- Solubility
- Insoluble
- Soluble
- Sparingly soluble
- Precipitation
- Crystallization

By the end of this topic, the learner should;

- be familiar with, and be able to carry out neutralization reactions to prepare salts (k, u, s)
- Know and appreciate the uses of common salts in everyday life (k,s)

Salts are ionic compounds formed by the neutralization of an acid and a base. They consist of a cation (positive ion) from the base and an anion (negative ion) from the acid.

A **salt** is a substance formed when either all or part of the ionisable hydrogen are replaced directly or indirectly with a metallic or ammonium ion. Salts are essential in various industries, agriculture and our daily lives.

Common salts and their parent acids

Parent acid	Name of salt formed	Examples of salts the acid forms
Hydrochloric acid, HCI	Chloride salt	• Sodium chloride • Potassium chloride • Ammonium chloride
Nitric acid, HNO_3	Nitrate salt	• Calcium nitrate • Aluminium nitrate
Sulphuric acid, H_2SO_4	Sulphate salt	• Sodium sulphate • Ammonium sulphate • Sodium hydrogen sulphate
Carbonic acid, H_2CO_3	Carbonate salt	• Ammonium carbonate • Magnesium hydrogen carbonate

Natural Sources of Salts

Here are some natural sources of salts and examples of salts found in them

Source	Examples of salts found in them
Seawater	Sodium chloride (NaCl) also called Common table salt Magnesium chloride ($MgCI2$): Used in various industrial processes Potassium chloride (KCI), A common fertilizer
Rock Salt Deposits	Sodium chloride Potassium chloride (KCI), Used as a fertilizer and in medicine
Mineral Springs	Sodium chloride Calcium sulphate ($CaSO_4$): Found in gypsum, used in plaster and drywall Magnesium sulphate ($MgSO4$): Used as a laxative and in bath salts. Calcium carbonate ($CaCO3$), found in chalk, limestone and marble.
Plants	Potassium nitrate ($KNO3$): Found in some plants, used in fertilizers and gunpowder

Chemical Formulae of Salts

A salt is a chemical compound formed by the neutralization reaction of an acid and a base. It consists of a positively charged ion (cation) and a negatively charged ion (anion).

General Formula of a Salt

$Cation^{+} + Anion^{-} \rightarrow Salt$

Examples of Common Salts

- ✓ Sodium chloride (NaCI)
- ✓ Potassium chloride (KCI)
- ✓ Calcium chloride ($CaCI_2$)
- ✓ Magnesium sulphate ($MgSO_4$)
- ✓ Sodium nitrate ($NaNO_3$)
- ✓ Ammonium Sulphate $(NH_4)_2SO_4$

Radicals

A radical is a group of atoms that act as a single unit and carry a charge. They are formed when a molecule gains or loses electrons. Radicals are highly reactive due to the presence of unpaired electrons.

Common Radicals

- ✓ Nitrate (NO_3^{-}): A polyatomic ion with a -1 charge.
- ✓ Sulphate (SO_4^{2-}): A polyatomic ion with a -2 charge.
- ✓ Carbonate (CO_3^{2-}): A polyatomic ion with a -2 charge.
- ✓ Ammonium (NH_4^{+}): A polyatomic ion with a +1 charge.
- ✓ Hydroxide (OH^{-}): A polyatomic ion with a -1 charge.

Note:

A radical is a charged atom or a group of charged atoms that cannot exist on their own. The charge on the radical gives the valency of that radical.

For example: Carbonate (CO_3^{2-}): has a charge of -2, the valency for the carbonate radical is 2.

Valency

Valency is the combining capacity of an element or radical. It represents the number of electrons an atom or radical gains, loses, or shares to form a chemical bond.

Determining Valency

Metals: The valency of a metal is usually equal to the number of electrons it loses to form a cation.

Non-metals: The valency of a non-metal is usually equal to the number of electrons it gains to form an anion.

Example

Sodium (Na) has a valency of 1, as it loses one electron to form Na^{+}.

Chlorine (Cl) has a valency of 1, as it gains one electron to form Cl^{-}.

Valences of some common elements

Elements	Symbol	Valency	Elements	Symbol	Valency
Copper	Cu	1 & 2	**Magnesium**	Mg	2
Iron	Fe	2 & 3	**Aluminium**	Al	3
Lead	Pb	1	**Sodium**	Na	1
Zinc	Zn	2	**Potassium**	K	1
Silver	Ag	1	**Calcium**	Ca	2

Writing Chemical Formulae

To write the chemical formula of a compound, you must know symbol for an element, formula of the radical and their valences

Example 1

Clacium Chloride

Step I Take the symbols along with the valencies of the radicals or elements

Ca^{2+} Cl^{1-}

Step II If the valencies are different irrespective of their sign (+ or -) interchange them and are written in subscript

Ca_1Cl_2

$CaCl_2$ '1' is ignored while writing chemical formula

Formula of the given compound is $CaCl_2$

By understanding radicals and valences, you can accurately predict the chemical formulas of compounds and balance chemical equations.

Naming Salts

The name of a salt is derived from the names of the cation and anion.

Rules for Naming Salts:

1. Identify the cation (positive ion) and anion (negative ion).
2. Use the cation's name followed by the anion's name, modified to end in "-ite", "-ide" or "-ate".

Cation Names:

1. Metals: Use the metal's name (e.g., Sodium, Calcium).
2. Ammonium (NH_4^+): Use "ammonium".

Anion Names:

1. Monatomic anions (single element): Use the ending "-ide" (e.g., Chloride, Oxide).
2. Polyatomic anions (multiple elements): Use the anion's name (e.g., sulphate, Nitrate).

Common Anion Names:

- ✓ Chlorine: Chloride (Cl^-), Chlorate (ClO_3^-)
- ✓ Oxygen: Oxide (O^{2-})
- ✓ Sulphur: sulphide (S^{2-}), sulphate (SO_4^{2-}), sulphite (SO_3^{2-})
- ✓ Nitrogen: Nitride (N^{3-}), Nitrate (NO_3^-), Nitrite (NO_2^-)
- ✓ Phosphorus: Phosphide (P^{3-}), Phosphate (PO_4^{3-})

Examples:

- ✓ NaCl: Sodium Chloride
- ✓ $CaCO_3$: Calcium Carbonate
- ✓ NH_4NO_3: Ammonium Nitrate
- ✓ $FeSO_4$: Iron(II) sulphate
- ✓ $Cu(OH)_2$: Copper(II) Hydroxide

Roman Numerals:

Use Roman numerals to indicate the oxidation state (Valence) of metals with multiple possible charges.

Examples with Roman Numerals

- ✓ Iron(II) chloride ($FeCl_2$)
- ✓ Copper(II) sulphate ($CuSO_4$)
- ✓ Iron(III) oxide (Fe_2O_3)

Common Salt Names:

- ✓ Table salt: Sodium Chloride (NaCl)
- ✓ Epsom salt: Magnesium sulphate ($MgSO_4$)
- ✓ Baking soda: Sodium Bicarbonate ($NaHCO_3$)

Salts, elements found in it and Chemical formulae

This table provides a list of common salts, their chemical formulas, and the elements that compose them.

Salt Name	Chemical Formula	Elements Present
Sodium Chloride	NaCl	Sodium (Na), Chlorine (Cl)
Potassium Chloride	KCl	Potassium (K), Chlorine (Cl)
Calcium Chloride	$CaCl_2$	Calcium (Ca), Chlorine (Cl)
Magnesium sulphate	$MgSO_4$	Magnesium (Mg), sulphur (S), Oxygen (O)
Sodium Nitrate	$NaNO_3$	Sodium (Na), Nitrogen (N), Oxygen (O)
Ammonium sulphate	$(NH_4)_2SO_4$	Nitrogen (N), Hydrogen (H), sulphur (S), Oxygen (O
Sodium Carbonate	Na_2CO_3	Sodium (Na), Carbon (C), Oxygen (O)
Calcium Carbonate	$CaCO_3$	Calcium (Ca), Carbon (C), Oxygen (O)
Potassium Nitrate	KNO_3	Potassium (K), Nitrogen (N), Oxygen (O)
Copper sulphate	$CuSO_4$	Copper (Cu), sulphur (S), Oxygen (O)

Making Salts

Salt Preparation

Because of the large number of, in particular, metal oxides and carbonates that it is possible to react easily with a number of acids, a whole range of 'new substances' can be made using acid reactions. If we include, ***precipitation*** reactions there are very few compounds that cannot be made quickly and easily.

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Salts are prepared basing on their solubility in water. A salt is either soluble when it dissolves in water or insoluble when it does not dissolve in water.

Soluble salts are prepared in the laboratory by **action of an acid on a base (Neutralization reactions)** or direct combination with a metal. **Insoluble** salts are prepared by **precipitation** or **double decomposition.**

Note:
This is basically a ***Problem Solving*** activity that will test your knowledge of ***acid reactions***, your use of ***solubility tables***, your appreciation of ***practical considerations*** (such as ensuring complete reaction) and your knowledge of ***separation techniques***.

There are 3 parts to salt preparation 1) ***Choice of Reaction***
2) ***Reaction Method***
and 3) ***Separation of Salt produced***

1) ***Possible Reactions***
a) ***Acid*** + (solid) ***Metal***
b) ***Acid*** + (solid) ***Oxide, Hydroxide or Carbonate***
c) ***Acid*** + ***Alkali*** (solution)
d) ***Precipitation*** (solutions, one of which **may** be acid)

2) ***Reaction Methods*** & 3) ***Separation of Salt***

Action of an acid on:

a. Metal Carbonates

When acids react with metal carbonates, they produce a salt, water. ***The reaction produces carbon dioxide gas, which causes the solution to bubble and fizz.***

- ✓ ***General equation***: Acid + metal carbonate → salt + water + carbon dioxide
- ✓ *Examples*:
    - Hydrochloric acid + sodium carbonate → sodium chloride + water + carbon dioxide
    - Sulphuric acid + calcium carbonate → calcium sulphate + water + carbon dioxide

Upon evaporation, the salt will crystallize. The specific salt formed depends on the acid and carbonate used.

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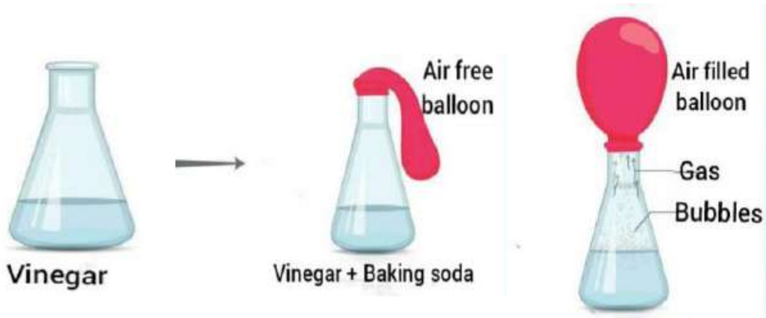
Chemical Reaction Between Vinegar and Baking Soda

Materials

- ✓ Plastic water bottle/Conical flash.
- ✓ Vinegar (acetic acid)
- ✓ Baking soda (sodium bicarbonate, $NaHCO_3$)
- ✓ Balloon
- ✓ Spatula

Procedure



- ✓ Half fill a bottle with vinegar.
- ✓ Place 1-2 spatulas of baking soda into balloon.
- ✓ Stretch balloon over bottle mouth.
- ✓ Quickly release baking soda into vinegar and shake gently .
- ✓ Observe and record

Observations

- ✓ Effervescent/bubbles of a colorless gas (Carbon dioxide) occurs producing a hissing sound.
- ✓ Balloon inflates

Chemical Reaction

Baking soda + Vinegar $\rightarrow$Salt + Carbon dioxide + Water

Note:

Copper (II) nitrate, Magnesium sulphate, Calcium chloride, Calcium nitrate salts can also be prepared by the reaction of acids with carbonates.

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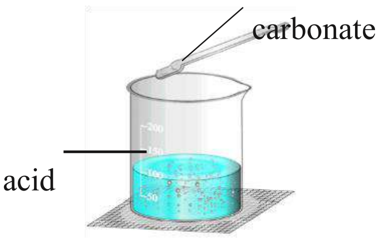
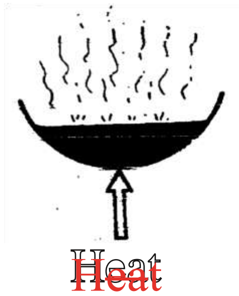
Preparing a Soluble salt

Measure 50cm^3 of an acid using a measuring cylinder and transfer it into a glass beaker.
Add the solid carbonate spatula by spatula, stirring all the time, until there is an obvious layer of unreacted (*excess*) solid lying at the bottom. The acid may need to be heated to speed up the reaction.



Using a filter funnel and a filter paper, filter off the *excess* solid to separate it from the *salt* solution.
The solid trapped in the filter paper can be discarded.

Transfer the filtrate (*salt* solution) into an evaporating dish.
Heat the *salt* solution until all the water has evaporated away leaving solid *salt* powder. If preferred, the solution can be left to evaporate slowly in which case salt crystals will form.

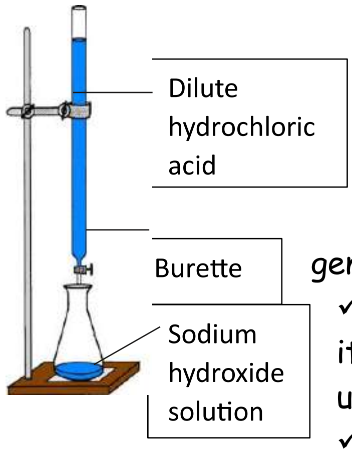


b. Metal hydroxides

Experiment: Preparing Salts by the Reaction of Acids with Hydroxides

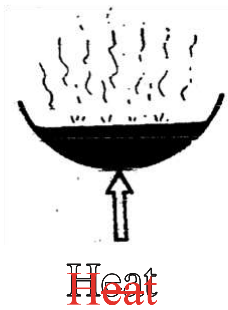
Materials

- ✓ Dilute hydrochloric acid (HCI)
- ✓ Sodium hydroxide solution (NaOH) solution
- ✓ Droppers
- ✓ Conical flask
- ✓ Burette
- ✓ methyl orange indicator



Procedure

- ✓ Using a measuring cylinder, measure 50cm^3 of sodium hydroxide solution into a conical flask.
- ✓ Add a few drops of methyl orange indicator, shake gently and note the color change.
- ✓ Fill the burette with dilute hydrochloric acid and run it into the mixture in the conical flask while shaking the flask until there is a color change.
- ✓ Transfer the contents in a conical flask to an evaporating dish and evaporate the mixture to dryness.
- ✓ Record your observation.
- ✓ Repeat the above steps using dilute hydrochloric acid and potassium hydroxide.



Discussion

When hydrochloric acid (HCl) reacts with sodium hydroxide (NaOH), a neutralization reaction occurs forming a salt and water, the reaction is Exothermic (heat-releasing)

<u>General equation</u>

Hydrochloric acid + Sodium hydroxide → Sodium chloride (common salt) + Water

Conclusion

Salts of sodium, potassium and ammonium can be prepared by the reaction of acids with hydroxides using an indicator which detects the point of neutralization.

The specific salt formed depends on the acid and hydroxide used.

Sodium chloride is the most commonly used salt in our day-to-day life for cooking, preserving foods etc. sodium chloride is found naturally in Lake Katwe in Uganda. Other salts found in Lake Katwe are sodium carbonate and sulphates of sodium and potassium, these salts are removed as impurities during the process of obtaining sodium chloride.

The Process of Salt Extraction from Lake Katwe

Here's a breakdown of the traditional salt extraction process at Lake Katwe:

Natural Evaporation

Sun and Wind: The intense African sun and strong winds naturally evaporate the water from the highly saline lake.

Concentration: As the water evaporates the salt concentration in the remaining brine increases.

Crystallization

When the brine becomes saturated with salt, sodium chloride crystals begin to form on the surface and the bottom of the lake.

Harvesting

Often women, collect the salt crystals that form on the surface of the lake. Workers, usually men, dive into the water to collect salt crystals from the bottom.

Cleaning and Drying

The harvested salt is cleaned to remove any mud, dirt, or other impurities.

The cleaned salt is then dried in the sun to reduce its moisture content.

Processing (Optional)

In some cases, the dried salt may be ground into finer particles for various uses.

c. Metal oxide

Activity: Preparing Magnesium Sulphate (MgSO4) by action of Sulphuric acid on Magnesium Oxide.

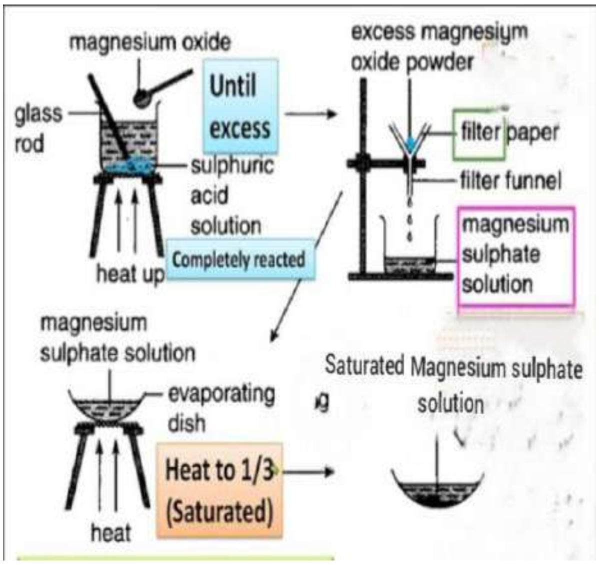
Safety Precautions

- ✓ Handle acids with care.
- ✓ Avoid contact with your skin and eyes.
- ✓ Work in a well-ventilated area.
- ✓ Dispose of waste materials properly.

Materials

- ✓ Dilute sulphuric acid (H_2SO_4)
- ✓ Magnesium oxide powder (MgO)
- ✓ Glass beakers
- ✓ Heat source
- ✓ Tripod stand
- ✓ Filter papers and funnel
- ✓ Droppers

Procedure



- ✓ Using a measuring cylinder, measure 50cm^3 of dilute sulphuric acid solution into a glass beaker. Warm the solution gently to 60^{o}C.
- ✓ Using a spatula, Gradually add magnesium oxide powder to the warmed acid while stirring until the unreacted salts starts to form (a milky solution forms).
- ✓ Filter the mixture that is formed using a filter paper and a funnel to obtain a clear liquid (filtrate)
- ✓ Transfer the contents(filtrate) into an evaporating dish and evaporate some water to obtain a saturated solution
- ✓ Remove the mixture from the heat source and allow it to cool slowly and crystallize.
- ✓ Filter off the crystals and dry them in between the filter papers.

Discussion

When an acid reacts with a metal oxide, the result is a salt and water. This is because metal oxides are bases, which neutralize acids. The specific salt produced depends on the acid and base used.

Chemical Equation

H2SO4 (sulphuric acid) + MgO (magnesium oxide) → MgSO4 (magnesium sulphate) + H2O (water)

Other salts that can be prepared by the action of an acid on an oxide include

Sodium Sulphate, Potassium carbonate, Copper sulphate, Zinc sulphate.

Uses of Magnesium Sulphate

- ✓ Medical applications (e.g., anticonvulsant, laxative)
- ✓ Agriculture (e.g., fertilizer, soil amendment)
- ✓ Industrial processes (e.g., textile manufacturing, paper production)

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Insoluble Salts

Insoluble salts are chemical compounds that do not readily dissolve in water. When these salts are mixed with water, they typically form a solid precipitate. These salts are basically prepared by a method of **precipitation** or **double decomposition.**

Precipitation involves reacting two soluble salts to form a precipitate (insoluble salt) and another soluble salt.

Examples of Insoluble Salts

Silver chloride (AgCl), Lead(II) sulphate (PbSO4), Calcium carbonate (CaCO3), Copper(II) hydroxide (Cu(OH)2)

Experiment: Preparing an Insoluble Salt

Materials

- ✓ Sodium chloride solution (NaCl)
- ✓ Silver nitrate solution ($AgNO_3$)
- ✓ Beakers
- ✓ Filter funnel and filter papers
- ✓ Retort stand

Procedure

- ✓ Using a measuring cylinder, measure $50cm^3$ of silver nitrate solution into a beaker.
- ✓ Add $50cm^3$Sodium chloride solution to the beaker containing sodium chloride solution while continuously stirring.
- ✓ Filter off the mixture.
- ✓ Wash the residue thoroughly with distilled water.

Expected observations

A white precipitate of silver chloride (AgCl) will form, indicating the formation of an insoluble salt.

$$NaCl_{(aq)} + AgNO_{3(aq)} \longrightarrow AgCl(s) + NaNO_{3(aq)}$$

Conclusion

The experiment demonstrates the preparation of an insoluble salt, silver chloride, through a double displacement reaction.

Other insoluble salts can be prepared by combining different solutions of soluble salts, such as lead (II) nitrate and potassium iodide to form lead(II) iodide.

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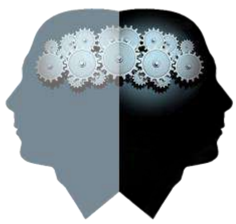
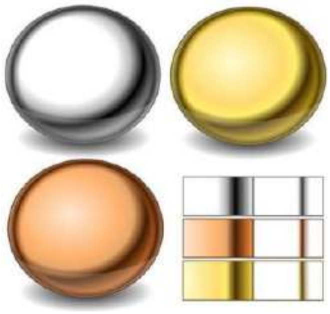
Evaluation of the methods of salt preparation

Reaction	***'Advantages'***	***'Disadvantages'***	***'Suitability'***
solid ***metals***	Easy to ensure 'complete' reaction - excess metal left over at end.	Must filter excess metal. Not all metals reactive enough to react with acid.	Not suitable if salt is insoluble - too difficult to separate from excess solid metal.
solid ***oxides***, ***hydroxides***, ***carbonates***	Easy to ensure 'complete' reaction - excess solid left over at end.	Must filter excess solids. Often need to heat oxides and hydroxides.	Not suitable if salt is insoluble - too difficult to separate from excess solid.
alkali *solutions*	Reaction immediate. No need to filter excess solids.	Difficult to ensure 'exact' neutralisation. Technique may take a very long time.	Very limited choice of alkalis - so limited number of salts can be prepared by this method.
precipitation from *solutions*	Reaction extremely quick.	None really.	Limited to insoluble salts only.

Knowledge Met in this Topic



- ***Salts*** are ionic compounds formed in reactions between acids and bases. A metal ion will have replaced the hydrogen in an acid.
- Hydrochloric acid HCl forms ***chloride*** salts e.g. NaCl, Sulphuric acid H_2SO_4 forms ***sulphate*** salts e.g. CuSO4, Nitric acid HNO3 forms ***nitrate*** salts e.g. KNO3
- ***Soluble Salts*** can be made by reacting ***excess*** *insoluble* metal oxides or metal carbonates with a suitable acid. e.g. At the end of the reaction the mixture is ***filtered*** and the ***soluble salt*** can be ***separated*** by ***evaporation*** to dryness.
- ***Insoluble Salts*** can be made by ***precipitation*** - *two solutions* (two Soluble salts) combine to produce an *insoluble precipitate* of the salt. At the end of the reaction the ***precipitate of the insoluble salt*** can be ***separated by filtration*** and then *dried*.



Competency Based Curriculum

TOPIC 3:

THE PERIODIC TABLE

Competency: The learner investigates the diversity of the elements in the Periodic Table.

Key words

- Atomic number
- Sub-atomic particles
- Atomic mass
- Period
- Group
- Metal
- Non-metal
- Metalloid

By the end of this topic, the learner should be able to;

- Understand that elements can be grouped into metals and non metals and relate the physical properties of metals and non-metals to their uses (k, u, s)
- Know that the Parodic Table is a classification of elements according to their atomic or proton number(k)
- Relate the arrangement of electrons in the first 20 elements to their positions in the Periodic Table (u, s)
- Understand the relationship between the positon of elements in groups and the charge on the ions that they form(u)

➢ PERIODIC TABLE:

The **periodic table** is a **tabular arrangement** of the **chemical elements**, **organized** by their **atomic number**, electron configuration, and recurring chemical properties. It provides a valuable tool for understanding the relationships between elements and predicting their behaviour.

➢ Classification of elements in the periodic table:

Activity: Metal or Non-Metal?

Aim: To enable students to differentiate between metals and non-metals based on their physical properties.

Materials:

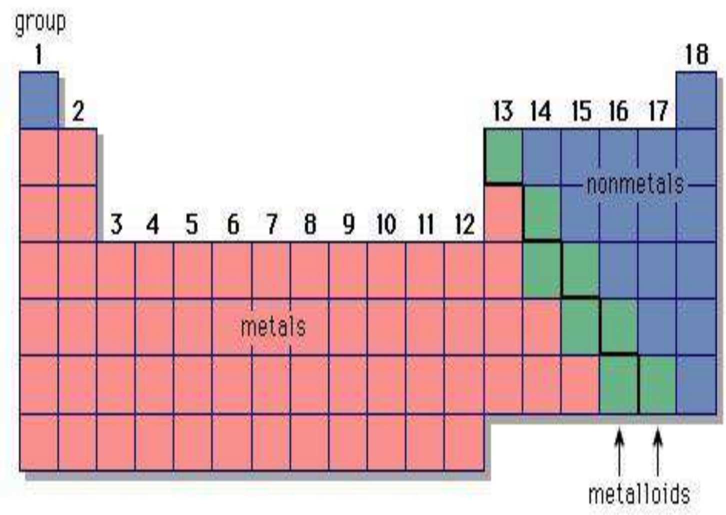
- ✓ Samples of metals (e.g., iron, copper, aluminium)
- ✓ Samples of non-metals (e.g., sulphur, carbon, iodine)
- ✓ Magnets
- ✓ Hammer or screwdriver
- ✓ Hand lenses or magnifying glasses
- ✓ Water
- ✓ Acid (dilute HCl or H_2SO_4)

Procedure

- ✓ Identify the elements provided to you in your groups.
- ✓ Test the physical properties of each sample, recording your observations I.e.
  Attraction to magnets
  Malleability (using hammer or screwdriver)
  Appearance under magnification
  Reaction with water
  Reaction with acid
- ✓ Discuss your findings, comparing the properties of metals and non-metals.

Discussion questions

1. What properties distinguish metals from non-metals?
2. Which properties are unique to metals? To non-metals?
3. How do these properties relate to everyday uses of metals and non-metals?



Elements in the periodic table can be broadly classified into two categories: **metals** and **non-metals**. Metals are generally found on the left side of the periodic table (except hydrogen). Non metals are located on the right side of the periodic table.

METALS:

Metals are elements that typically exhibit the following properties:

- ✓ Shiny and reflective (lustrous)
- ✓ They can be hammered into thin sheets (Malleable)
- ✓ They are Ductile (can be drawn into wires)
- ✓ They are good conductors of heat and electricity.
- ✓ High Density, they are generally dense.
- ✓ They tend to lose electrons to form positive ions (cations).
- ✓ They react with oxygen and water to form oxides, leading to corrosion.

Examples of Metals

- ✓ Alkali Metals: Lithium (Li), Sodium (Na), Potassium (K)
- ✓ Alkaline Earth Metals: Magnesium (Mg), Calcium (Ca), Barium (Ba)
- ✓ Transition Metals: Iron (Fe), Copper (Cu), Gold (Au), Zinc (Zn), Lead (Pb)

NON-METALS:

Non-metals are elements that generally located on the right hand side of the periodic table, lack the metallic properties. They exhibit a wide range of properties:

- ✓ In the upper-right corner of the periodic table
- ✓ Dull, non-reflective appearance (not shinny)
- ✓ Poor conductors of electricity and heat
- ✓ Non-ductile

- ✓ They tend to gain electrons to form negative ions (anions).
- ✓ Many non-metals are relatively unreactive.
- ✓ Some (O, Cl) are gases at room temperature, others (S) are brittle solids (break easily)

Examples of Non-metals:

- ✓ Halogens: Fluorine (F), Chlorine (Cl), Iodine (I)
- ✓ Noble Gases: Helium (He), Neon (Ne), Argon (Ar)
- ✓ Other Non-metals: Carbon (C), sulphur (S), Phosphorus (P)

Note: Some elements, like boron (B) and silicon (Si), have properties of both metals and non-metals and are classified as metalloids or semi-metals.

METALLOIDS ("*SEMI-METAL*"):

- ✓ On the stair-step line that divides the metals from the non-metals
- ✓ Exhibit properties that are intermediate between those of metals and non-metals
- ✓ Important metalloids: **silicon, germanium**

Activity:

Use a black and white copy of the periodic table. On one side, color and label the metals, non-metals, and metalloids. Color and label the groups/families of elements on the other side of your paper. Remember to create a legend.

Relating Physical Properties of Metals to Their Uses

The physical properties of metals make them suitable for a wide range of applications. Here's a breakdown of how specific properties lead to particular uses.

- ✓ **Lustrous**

Metals have a shiny appearance.

Use: Making jewellery, decorative items, and mirrors.

- ✓ **Sonorous**

Metals produce a ringing sound when struck.

Use: Making musical instruments like bells, cymbals, and gongs.

- ✓ **Malleability and Ductility**

Metals can be hammered or rolled into thin sheets

Use: Making utensils, car bodies, and roofing sheets.

Metals can be drawn into wires (Ductile)

Use: Making electrical wires, jewellery, and cables.

- ✓ **Conductivity (Thermal and Electrical)**

Metals have the ability to conduct heat (have high thermal conductivity)

Use: Making cooking utensils, radiators, and heat sinks.

Have high electrical conductivity.

Use: Making electrical wires, cables, and electronic components.

- ✓ **Hardness and Strength**

Metals are generally hard thus resistant to scratching or indentation.

Use: Making tools, machinery parts, and construction materials.

Have a high tensile strength (ability to withstand stress without breaking).

Use: Building bridges, skyscrapers, and vehicles.

Relating Physical Properties of Non-Metals to Their Uses

Non-metals, though diverse in their properties, also find various applications due to their unique characteristics. Here's a breakdown of how their properties relate to their uses.

- ✓ **Poor Conductors of Heat and Electricity**

Low electrical conductivity: Ideal for:

- Semiconductors (e.g., silicon)
- Electronic components like insulators in electrical wiring.

Low thermal conductivity: Suitable for:

- Insulation materials like making pan handles (e.g., graphite, silica)
- Fire-resistant materials

✓ **Brittle and Non-Malleable**

Easy fragmentation so can be used in powdered form for;

- Grinding and polishing materials (e.g., silicon carbide)
- Producing nanomaterials

✓ **Dull Appearance**

Use: Non-reflective coatings and camouflage.

✓ **Low Density**

Use: Used in lightweight materials like plastics and foams.

✓ **Reactivity**

Reactive Non-metals: Used in various chemical reactions, such as chlorine for water purification and sulphur for fertilizers, Chemical synthesis (e.g., oxygen, nitrogen), Catalysts (e.g., carbon nanotubes)

Non-reactive Non-metals: Used as inert gases in various applications, such as helium in balloons and neon in neon signs.

Specific Examples:

Carbon:

- ✓ Diamond: Extremely hard, used in cutting tools and jewellery.
- ✓ Graphite: Soft and slippery, used in pencils and lubricants.

Sulphur: Used in the production of sulphuric acid, a key industrial chemical.

Phosphorus: Used in fertilizers and matches.

Chlorine: Used in water purification and as a disinfectant.

Key Groups and Their Properties:

Group 1 (Alkali Metals): Highly reactive metals that readily lose one electron.

Group 2 (Alkaline Earth Metals): Reactive metals that lose two electrons.

Group 17 (Halogens): Highly reactive non-metals that gain one electron.

Group 18 (Noble Gases): Nonreactive gases with a full outer electron shell.

The structure of the atom

When scientists started exploring matter, they realized that matter can be divided into smaller and still smaller particles. They called the smallest particle an 'atom'.

The name 'atom' was derived from the Greek word ***'atomos'***, meaning 'indivisible'.

They discovered that the 'atom' maintains its chemical identity through all chemical and physical changes.

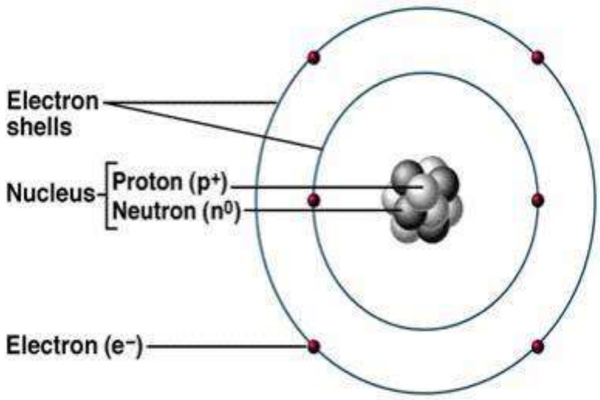
Nature of Atom

At present we know that the atom is the smallest particle of an element.

It is made up of sub-atomic particles like electrons, protons and neutrons.

Atoms of one type of element differ from those of the other due to different number of sub-atomic particles.

The protons and neutrons are in the nucleus (centre) of the atom and the electrons orbit round the outside in shells (energy levels or layers).



The energy levels are counted from centre outwards as 1st, 2nd, etc.

Summary of sub-atomic particles

Particle	Relative mass	Electric charge	Comments
Proton	1	+1 (positive)	In the nucleus (a nucleon)
Neutron	1	0 (zero)	In the nucleus (a nucleon)
Electron	$^1/_{1850}$	-1 (negative)	Arranged in energy levels or shells around the nucleus (see later)

Notice that the number of protons and that of electrons are always equal in neutral atoms.

The total number of protons in the nucleus of an atom contributes to Atomic number while the total number of both protons and neutrons in the nucleus of an atom is referred to as mass number/Atomic mass.

Atomic Number and Mass Number

The nuclei of atoms are made up of protons and neutrons. These two components of the nucleus are referred to as nucleons.

Since an atom is electrically neutral, the number of protons in the nucleus is exactly equal to the number of electrons.

This number is the atomic number given by the symbol Z.

Atomic number represents the number of protons in an atom.

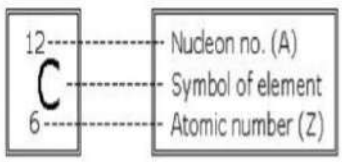
As atoms are electrically neutral, an atom contains as many electrons as it has protons.

The total number of protons and neutrons present in one atom of an element is known as its mass number.

Mass number (A) = number of protons (Z) + number of neutrons (n)

It can also be stated that:

Mass number (A) = atomic number (Z) + number of neutrons (n)



Symbols

The mass/nucleon number (A) is written as a superscript on the top-left corner of the symbol of the atom. The atomic number(Z) is written as a subscript on the bottom-left corner.

Example:

The symbol represents an atom of sodium whose atomic mass is 23 and atomic number is 11. Calculate the number of protons, electrons and neutrons.

Atomic number Z =11
Atomic mass A = 23
No. of protons = Z = 11
No. of electrons = 11

No. of neutrons = A - Z
23 - 11 = 12

Activity: Relating the atomic number to the position of an element in the periodic table

What to do:

1. Observe the periodic table and state the general pattern of atomic number and atomic mass in rows and columns.
2. Focus on the atomic numbers of 4 consecutive elements and deduce on the arrangement of elements in the periodic table.

Expected Responses

Both atomic number and atomic mass increases across the period, atomic number increase by one, while mass number doesn't have a definite number of increase.

Atomic Number is a unique identifier for each element.

Elements are arranged in the periodic table in increasing order of their atomic numbers. This means that as you move from left to right across the table, and from top to bottom, the atomic number increases.

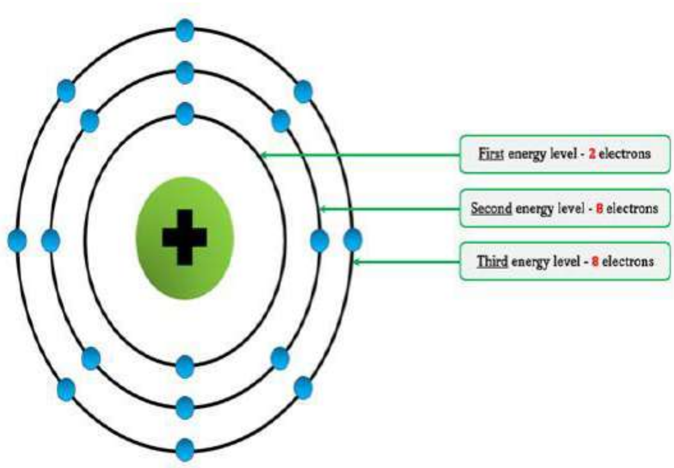
The position of an element in the periodic table is directly related to its electron configuration, which in turn determines its chemical properties. Elements in the same group (vertical column) have similar chemical properties because they have the same number of valence electrons (electrons in the outermost shell).

Electron energy levels in atoms

1. How are electrons arranged in an atom?
2. Determine the electronic structure of the first 20 elements.

The arrangement of electrons in an atom's energy levels determines its chemical properties and, consequently, its position in the periodic table.

Electrons are arranged in energy levels or shells around the nucleus. The shells are numbered starting from the inner most shell near the nucleus as 1, 2,3 and so on.



1st energy level (one next to the nucleus) can only take up a maximum of 2 electrons

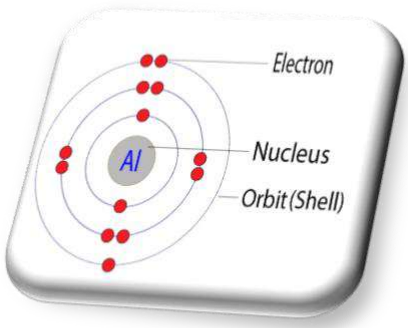
2nd energy level (one that follows the first one) takes a maximum of 8 electrons and so the 3rd, 4th...

Note: The electrons are not accommodated in a given shell unless the inner shells are filled up.

The arrangement of these electrons in the various shells/orbits/energy levels of an atom of the element is known as **electronic configuration.**

Valence electrons are the outermost electrons of an atom. They determine an element's chemical reactivity. Elements in the same group have the same number of valence electrons.

Example 1: Aluminum atom (Atomic/proton number 13 thus has 13 electrons)



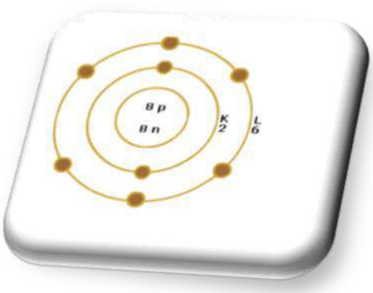
Steps:

The first 2 electrons will go to the 1st shell

The next shell takes a maximum of 8 electrons.

In this way 2 + 8 = 10 electrons have been accommodated. The next 3 electrons will go to the next Shell.

The **electronic Structure/Configuration** will be **2:8:3**



Example 2: Oxygen atom (Atomic/proton number 8, thus has 8 electrons)

Steps:

- ✓ The first 2 electrons will go to the 1st shell.
- ✓ The next shell takes a maximum of 8 electrons but we are left with only 6 electrons, so they will be all accommodated in this 2nd shell.

The **electronic configuration** of oxygen will be **2:6**

The electronic Structure/Configuration of the first 20 elements.

Element	Symbol	Shell number (n)			
		1	2	3	4
Hydrogen	H	1			
Helium	He	2			
Lithium	Li	2	1		
Beryllium	Be	2	2		
Boron	B	2	3		
Carbon	C	2	4		
Nitrogen	N	2	5		
Oxygen	O	2	6		
Fluorine	F	2	7		
Neon	Ne	2	8		
Sodium	Na	2	8	1	
Magnesium	Mg	2	8	2	
Aluminium	Al	2	8	3	
Silicon	Si	2	8	4	
Phosphorous	P	2	8	5	
Sulphur	S	2	8	6	
Chlorine	Cl	2	8	7	
Argon	Ar	2	8	8	
Potassium	K	2	8	8	1
Calcium	Ca	2	8	8	2

Food for thoughts

1) Draw the electronic structure of the atoms of the elements in the above table.
2) Identify the elements with fully filled outermost shells.

The periodic table of elements

The periodic table is organized into rows and columns.
The **horizontal rows** are called **periods** and are assigned numbers according to the number of electron shells/energy levels that are filled by electrons. All elements in the same period have the same number of electron shells /energy levels but the number of electrons occupying the last shell increase from left to right i.e. from one to eight. The periods are written in roman numbers as 1, 2, 3, etc. As one moves from left to right in a given period, the chemical properties of the elements slowly change.

Columns in the periodic table are called **groups** represented by roman numerals as (I), (II), (VI), (VIII) etc.. Groups move from **top-down**.

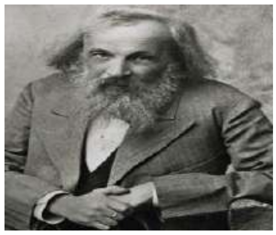
The number of shells increases down a group. However, the number of electrons in the last shell (valence Electrons) of each element in the same group is the same.

Elements in a given group in the periodic table share many similar chemical and physical properties.

For your own understanding

Dmitri Mendeleev – mid 1800's

-proposed a table for 70 elements based on increasing mass and similar properties



Dmitri Mendeleev is credited as being the Father of the modern periodic table.

In 1869 he arranged the 50 or so known elements in order of atomic number, Z, putting elements with similar properties in the same vertical group, and leaving gaps for unknown elements, yet to be discovered.

When the elements were later discovered, they were found to have the properties predicted by Mendeleev's table.

Henry Moseley – 1913

-determined the atomic number of elements and arranged the table in order of increasing atomic number

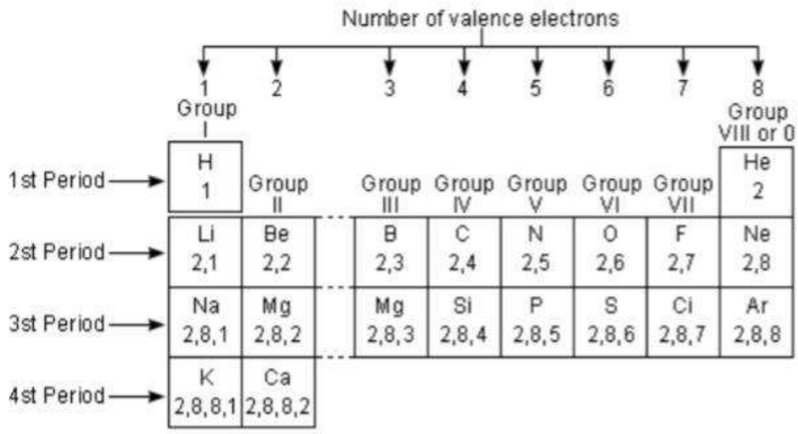
The First Twenty Elements

1. **Considering the first 20 elements, determine the group and period to which each element belongs.**

Atomic Number	Symbol	Name	Group	Period
1	H	Hydrogen	I	1
2	He	Helium	VIII	1
3	Li	Lithium	I	2
4	Be	Beryllium	II	2
5	B	Boron	III	2
6	C	Carbon	IV	2

7	N	Nitrogen	V	2
8	O	Oxygen	VI	2
9	F	Fluorine	VII	2
10	Ne	Neon	VIII	2
11	Na	Sodium	I	3
12	Mg	Magnesium	II	3
13	Al	Aluminium	III	3
14	Si	Silicon	IV	3
15	P	Phosphorous	V	3
16	S	Sulphur	VI	3
17	Cl	Chlorine	VII	3
18	Ar	Argon	VIII	3
19	K	Potassium	I	4
20	Ca	calcium	II	4

The Periodic Table of the First Twenty Elements



Note:

Hydrogen, the simplest element, occupies a unique position in the periodic table. Its placement has often been a subject of debate due to its distinct properties.

Hydrogen has only one electron in its orbital. This simple configuration allows it to exhibit properties similar to both alkali metals and halogens.

When it loses an electron, it forms a hydrogen cation (H^+), similar to alkali metals. When it gains an electron, it forms a hydride ion (H^-), similar to halogens.

Due to its dual nature, hydrogen is often placed in two positions, Group 1 (Alkali Metals) since they have Similar outer electronic configuration (one valence electron) and Can lose an electron to form a cation.

Group 17 (Halogens), Can gain an electron to complete its outer shell, Forms diatomic molecules (H_2).

In conclusion, hydrogen's unique position in the periodic table reflects its ability to exhibit properties of both metals and non-metals. While it's often placed with alkali metals, its versatility and distinctive behaviour make it a truly unique element.

Relating Group Position to Ion Charge

The group number of an element in the periodic table often indicates the number of valence electrons and, consequently, the charge of the ion it typically forms.

Atoms always tend to lose, gain or share electrons during chemical reactions.

By the loss or gain of electrons a neutral atom is changed to an ion. **Ions** are charged atoms or a group of atoms. In other words, ions are particles formed by atoms by the donation or acceptance of electrons. The ions formed are stable since they have fully filled outer shells.

Most of these atoms try to attain the configurations of either neon (2,8) or argon (2,8,8).

For group VIII elements, their atoms are completely filled with 8 electrons and stable and therefore, don't participate in chemical reactions.

IONS:

ION: an atom or group or atoms that has a positive or negative charge

Recall…an atom is electrically neutral because it has equal # of protons (+) & # of electrons (-)

Positive & negative ions form when electrons are transferred between atoms!. This results into imbalance in the number of protons and electrons. Atoms lose or gain electrons to attain a stable electronic structure equal to that of noble gases.

The losing or gaining of Electrons depends on the number of valence Electrons, atoms with less that 4 valence electrons find it easier to lose these electrons during chemical reactions.

Those with more than 4 electrons find it easier to gain than to lose their valence Electrons.

The charge on the cation (positively charged ion) indicates the number of electrons donated. In an anion (negatively charged ion), it shows the number of electrons gained.

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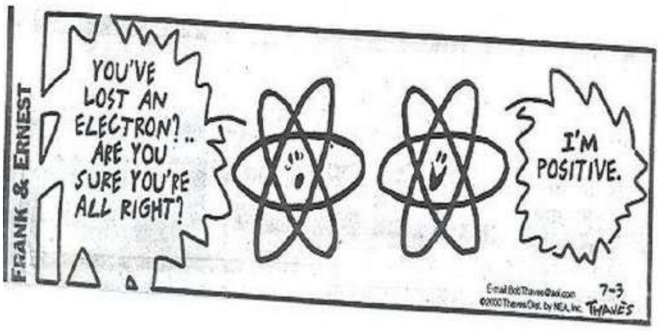
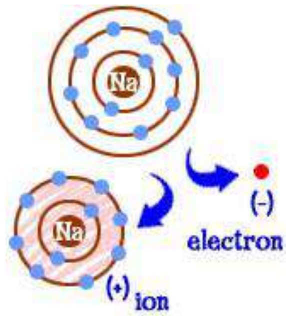
Formation of positively charged ions:



joke

- Two atoms walk into a bar.
One atom stops and says to the other, "I think I just lost an electron."
- The second atom asks "Are you sure?"
- The first atom replies, "I'm positive!"

Atoms of **METALS** tend to form <u>positive ions</u> by <u>losing 1 or more</u> electrons from their valence (outermost) shell. This lose of electrons causes electron (negative Charge) deficiency causing positive charges (protons) to be greater than negative charges, this makes the atom to have an overall positive charge.



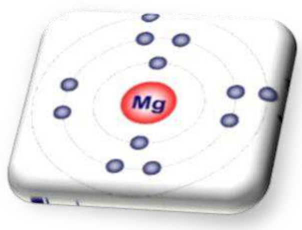
An ion with a positive charge is called a **<u>CATION</u>**. example: SODIUM ($Na \rightarrow Na^+$)
Atoms, particularly metals, tend to lose electrons to achieve a stable electron configuration, often resembling the electron configuration of a noble gas. **CATIONS** (positive ions) have empty valence shells, they are **<u>SMALLER</u>** than their neutral atom.

Activity: "Formation of Positive Ions"

Aim: To understand how positively charged ions (cations) are formed through the loss of electrons.

What to do:

1. Draw the atomic structure of the element, I.e Magnesium and write its electronic configuration..



E.C:-2:8:2

2. Identify the number of electrons in the outermost energy level of Magnesium atom.
Magnesium has two (2) electrons in its valence shell/outermost shell.

3. Determine how many electrons the atom needs to lose to form a positively charged ion.

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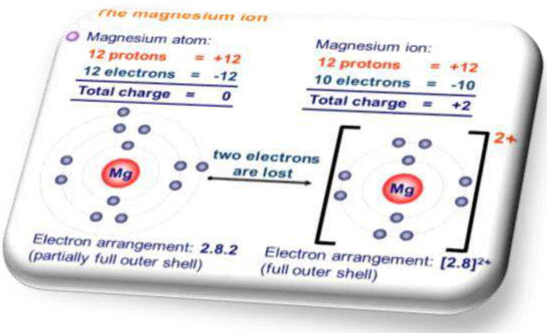
Always atoms (particularly metals) lose electrons from their outermost shells. Magnesium needs to lose 2 electrons from its outermost energy level to attain a stable structure.



4. Write the resulting ion's symbol, electronic configuration and charge.
Mg^{2+}, the charge of Magnesium ion is +2
E.C:-2:8

Independent Practice (15 minutes): Work in pairs to complete a simple worksheet with elements' symbols and atomic numbers.

1. Identify the elements that tend to lose electrons to form positively charged ions.
2. Write the resulting ion's symbol and charge for each element.

Whole-Class Discussion (10 minutes): As a class, review your findings and discuss any challenges or misconceptions.

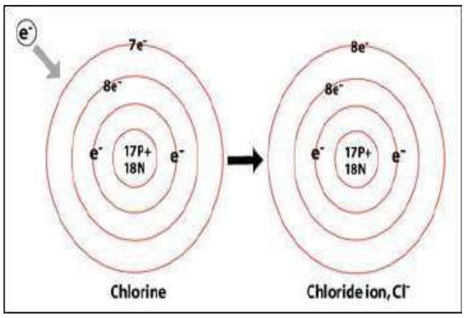
Formation of negatively charged ions:

Atoms of **NONMETALS** tend to form negative ions by gaining 1 or more e(& thus filling their outermost energy level)

An ion with a negative charge is called an **ANION**.

Example:

CHLORINE (Cl Cl^-)



ANIONS (negative ions) have full valence shells: they are **LARGER** than their neutral atom.

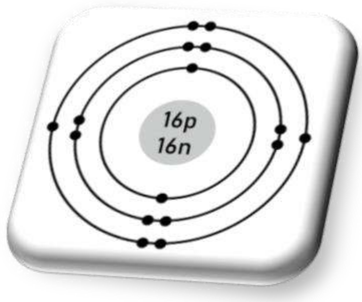
When an atom gains one or more electrons, it becomes negatively charged.

Activity: "Formation of negative Ions"

Aim: To understand how negatively charged ions (anions) are formed through the gain of electrons.

What to do:

1. Draw the atomic structure of the element, i.e Sulphur and write its electronic configuration.



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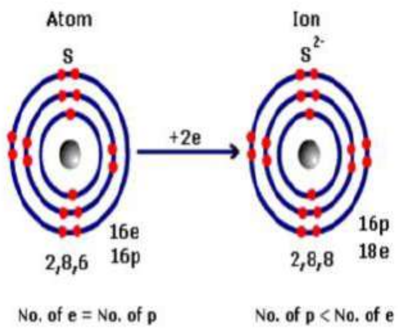
2. Identify the number of electrons in the outermost energy level.

Sulphur atom has six (6) electrons in its outer energy level.

3. Determine how many electrons the atom needs to gain to form a positively charged ion.



Non-metallic atoms find it easier to gain electrons to fill their outer shells. This increases the number of electrons as compared to protons making it have more negative charges than positive charges. Sulphur lacks two (2) electrons for it to fill its shell. Therefore, it will need to gain two (2) electrons.

4. Write the resulting ion's symbol, electronic configuration and charge.

S^{2-},the charge for the sulphide ion is -2.

E.C:-2:8:8

Independent Practice (15 minutes): Work in pairs to complete a simple worksheet with elements' symbols and atomic numbers.

1. Identify the elements that tend to lose electrons to form positively charged ions.
2. Write the resulting ion's symbol and charge for each element.

Whole-Class Discussion (10 minutes): As a class, review your findings and discuss any challenges or misconceptions.

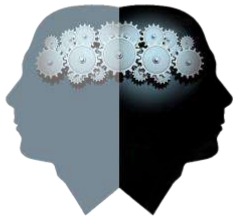
In summary

Group I, group II, group III elements gain stability by lose of the outermost electrons during chemical reactions.

Group V, VI, VII gain stability by gain of electrons.

Group VIII elements are stable and do not participate in chemical reactions.

Knowledge Met in this Topic



- ✓ Periodic Table is the arrangement of elements by atomic number and chemical properties. It consists of groups (vertical columns) and periods (horizontal rows).
- ✓ Elements are substances consisting of one type of atom. They are placed in the periodic table in groups and periods based on the number of electrons/increasing atomic number. Elements in the periodic table are categorized as Metals and Non-metals.
- ✓ Metals are found on the left side of the periodic table, they are malleable, Ductile, good Conductors of Heat and electricity and Non-metals are found on the right hand side of the periodic table, they are neither malleable nor ductile.
- ✓ Atomic Number is the number of protons in an atom's nucleus. Atomic mass is the number of both protons and neutrons in the nucleus of an atom.
- ✓ Metals form positive ions (cations) by loss of Valence electrons and non-metals form negative ions (anions) by gain of electrons.

End of chapter Scenarios

Item 1:

The reaction of oxygen and charcoal liberates heat energy essential for domestic purposes such as cooking.

However, when the reaction occurs in a poorly ventilated room, a toxic substance is formed following the equation below.

$2C(s) + O_2(g) \rightarrow 2CO(g)$

Task:

Using knowledge of chemistry;

a. Explain the category of the atoms in the product above.
b. Guide Trevor on:
   i. writing the electronic configuration of one of the atoms in the product.
   ii. The group and period of the chosen atom.
   iii. Formula of the ion the chosen atom forms.

Item 2:

The Uganda Aerospace Industry is developing new materials for aircraft and spacecraft. Elements exhibit unique properties based on their classification as metals or non-metals. The materials available are Aluminium (Al), Carbon (C), Oxygen (O), Iron (Fe), Nitrogen (N), Copper (Cu), Silicon (Si), Hydrogen (H).

Task;

As a materials scientist, classify elements into metals and non-metals to determine their suitability.

Item 3:

A local manufacturing plant produces various products, including electronics, furniture, and packaging materials. The plant receives a shipment of unknown substances for use in production.

Substances:

1. Substance J: Density 8.5 g/cm^3, conducts electricity, malleable
2. Substance K: Density 2.8 g/cm^3, brittle, transparent
3. Substance L: Density 10.2 g/cm^3, ductile, reactive with acids
4. Substance M: Density 1.9 g/cm^3, soft, porous

Task;

As a quality control specialist, identify and analyse these substances to determine their category, and potential uses.

Item 4:

A certain art man was given to make a school bell for a certain school. After his work, he tested the bell by ringing it but couldn't give a deep and ringing sound. He became confused and couldn't know where he made a mistake.

Task;

As a learner of chemistry;

I. Identify the problem that the art man made while choosing the substance for use.
II. Predict on the properties of the substance used by the art man.
III. Advise him on the impact of the substance on the environment.

Item 5:

Two boys from Kiyala High School were found fighting and they were given to dig a new rubbish pit. While digging the pit, they came across some substances from underground. On analysis, some substances gave a deep and ringing sound when struck while others could neither give a deep nor a ringing sound. The learners want to know more about the substance they found.

Task;

As a chemistry learner, help the boys to;

I. Give the categories of the substances they found.
II. Take them through the suitability of the substances they found.
III. Advise them on whether the substances have impacts on the environment or not.

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Competency Based Curriculum

TOPIC 4:

CARBON IN THE ENVIRONMENT

Competency: The learner investigates the diversity of carbon compounds in the environment.

Key words

- Briquettes
- Fossil fuels
- Sustainability
- Renewable
- Green house gases
- Allotropes
- Water hardness

By the end of this topic, the learner should be able to;

- Understand how and why carbon compounds are used as fuels (k,u)
- Know and appreciate the difference between renewable and non-renewable fuels and understand that non-renewable fuels are not sustainable (k,u)
- Know and appreciate the impact of burning carbon- based fuels on the environment (k,u)
- Understand the processes of making charcoal, but recognize that the use of charcoal as a fuel is cheap, efficient , and sustainable only if it is made from wood that can be regrown easily (u, s)
- Know and appreciate the physical properties and uses of carbon dioxide (k,u)
- Understand how the increase in carbon dioxide in the air can cause the atmosphere and the oceans to get warmer(u)
- Understand what greenhouse gases are, where they come from, and how they are affecting climate(u)
- Understand the origin of hard water in limestone areas and investigate how it can be softened (u, s)
- Understand how the properties and uses of the allotropes of carbon relate to their structures (u)

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Carbon is a vital non metallic element located in group IV and period 2 of the periodic table.

- ✓ Has <u>Atomic number</u>: **6** and <u>Atomic mass</u>: **12.011**
- ✓ <u>Symbol</u>: **C**
- ✓ <u>Electron configuration:</u> **2:4**
- ✓ Solid (graphite, diamond) at room temperature
- ✓ Density: 2.26 g/cm^3 (graphite), 3.52 g/cm^3 (diamond)
- ✓ Tetravalent (forms 4 bonds)

Activity : Simple Carbon Cycle Drawing

- ✓ Draw a large circle on the board. This is our "Carbon Circle."
- ✓ Brainstorm with the students the different places carbon can be found (e.g., in plants, animals, the air, the ground, the ocean). Write these as labels around the circle.
- ✓ Using different colored markers, draw arrows connecting the different carbon locations. Explain what each arrow represents (e.g., photosynthesis: arrow from air to plants; eating: arrow from plants to animals; decomposition: arrow from plants/animals to ground; respiration: arrow from plants/animals to air).
- ✓ Discuss the processes that move carbon between these locations.

Occurrence

- ✓ Found in all living organisms
- ✓ Present in fossil fuels (coal, oil, gas)
- ✓ Component of minerals (limestone, dolomite)
- ✓ Atmospheric carbon dioxide

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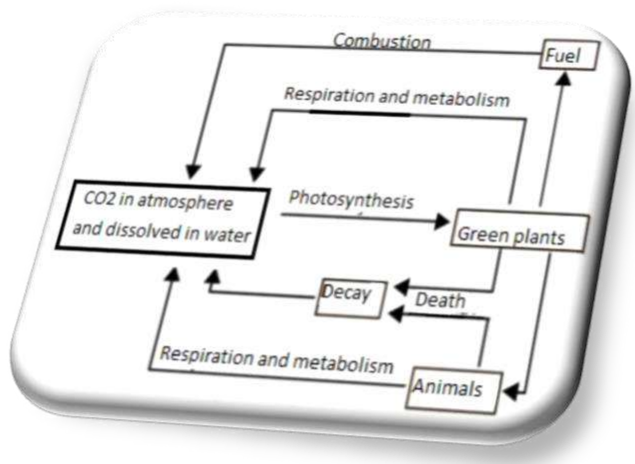
Carbon Cycle

The carbon cycle is the natural process by which carbon is exchanged between the atmosphere, oceans, land, and living organisms. It helps regulate the Earth's temperature.

Processes involved in the carbon cycle



- ✓ Organisms release carbon dioxide into the atmosphere through respiration. Carbondioxide can also be released into the atmosphere by volcanic activity. The Plants absorb carbon dioxide from the atmosphere and use it to produce glucose.
- ✓ When organisms die, their remains decompose and in the process releasing carbon dioxide into the atmosphere, the dead decayed organisms are subjected to high pressure over millions of years and are converted into fossil fuels which contain carbon components, these fossil fuels when burnt, they release carbon dioxide into the atmosphere.
- ✓ Ocean-atmosphere exchange; Carbon dioxide is exchanged between the ocean and the atmosphere.
- ✓ Sedimentation; Carbon is stored in sediments as carbonates.

Human Impact on the Carbon Cycle

The carbon cycle is a natural process essential for life, circulating carbon dioxide (CO_2) through atmosphere, oceans, land, and living organisms. However, human activities have significantly disrupted this cycle, leading to an imbalance with serious consequences. Here's how:

1. Burning Fossil Fuels:

We burn coal, oil, and natural gas for energy in cars, factories, and power plants. These fossil fuels were formed over millions of years from the remains of plants and animals.

Burning them releases large amounts of carbon dioxide (CO_2) into the atmosphere, a greenhouse gas that traps heat and contributes to global warming and climate change.

Mitigation:

- ✓ Transition to Renewable Energy: Shift from fossil fuels to renewable energy sources like solar, wind, hydro, and geothermal. This is the most crucial step.
- ✓ Improve Energy Efficiency: Use less energy through better insulation in buildings, more efficient appliances, and smarter transportation systems.
- ✓ Carbon Capture and Storage (CCS): Capture CO2 emissions from industrial sources and store them underground, preventing them from entering the atmosphere (still under development and has limitations).

2. Deforestation:

Cutting down forests for agriculture, urbanization, and other purposes. Trees absorb CO_2 during photosynthesis. When we cut them down, this carbon is released back into the atmosphere, further increasing greenhouse gas concentrations. Additionally, deforestation reduces the Earth's capacity to absorb CO_2.

Mitigation:

- ✓ Reforestation and Afforestation: Planting new trees in deforested areas and establishing forests in areas that were previously not forested.

- ✓ Sustainable Forest Management: Practices that ensure forests are managed for long-term health and carbon sequestration, including selective logging and preventing forest fires.
- ✓ Reduce Consumption of Wood Products: Use less wood and paper, and choose products made from recycled materials or sustainably sourced wood.

3. Land Use Changes:

Converting natural ecosystems like forests and grasslands into agricultural land or urban areas. These changes can reduce the amount of carbon stored in the soil and vegetation. Agriculture practices can also release greenhouse gases like methane and nitrous oxide.

- ✓ Sustainable Agriculture: Practices that reduce greenhouse gas emissions from agriculture, such as no-till farming, cover cropping, and improved manure management and farming methods that minimize soil disturbance and maintain soil organic matter, which stores carbon.
- ✓ Protecting Natural Ecosystems: Preventing the conversion of forests, grasslands, and wetlands to other land uses.

4. Industrial Processes:

Certain industrial processes, like cement production, release CO_2 as a by-product. These processes contribute to the overall increase in atmospheric CO_2.

Consequences of these impacts:

- ✓ Climate Change: Increased greenhouse gases trap more heat, leading to rising global temperatures, changes in weather patterns, sea level rise, and more frequent extreme weather events.
- ✓ Ocean Acidification: The ocean absorbs a significant portion of atmospheric CO2, which makes it more acidic. This harms marine life, especially shellfish and coral reefs.
- ✓ Disruption of Ecosystems: Changes in temperature and precipitation patterns can disrupt ecosystems, affecting plant and animal species.

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Mitigation:

- ✓ Developing Alternative Materials: Finding substitutes for cement and other carbon-intensive materials.
- ✓ Carbon Capture and Utilization (CCU): Capturing CO2 from industrial processes and using it to create other products, such as fuels or building materials.
- ✓ Improving Industrial Efficiency: Reducing energy consumption and waste in industrial processes.

Carbon Compounds as Fuels

Carbon is a versatile element that forms a wide range of compounds, many of which are excellent fuels. The abundance of carbon in the Earth's crust and its ability to form stable bonds with other elements make it an ideal component of fuels.

A **fuel** is any substance that can be burned to produce heat or power. Or any substance that can be consumed to release energy and perform work.

Fuels contain stored chemical energy that can be released through a combustion reaction (burning).

Fuels exist in various forms, including Solid (Coal, wood, charcoal), Liquid (Gasoline, diesel, oil), Gas (Natural gas, propane, methane)

Burning of Carbon-Based Fuels

Burning carbon-based fuels releases energy which can be inform of heat and sometimes light, but also emits harmful pollutants and greenhouse gases.

Combustion Process

Fuel + Oxygen → Heat + Light + Products (CO_2, H_2O, etc.)

Limited Oxygen (Incomplete Combustion)

Carbon + Oxygen → Carbon Monoxide + Water

Excess Oxygen (Complete Combustion)

Carbon + Oxygen → Carbon Dioxide + Water

Common carbon-based fuels

Food for thoughts

a. Identify the common carbon based fuels used in our day to day life.
b. Discuss the major use of these fuels.

- ✓ Firewood
- ✓ Coal; A solid fossil fuel formed from the remains of plants.
- ✓ Natural gas; a gaseous fossil fuel composed primarily of methane.
- ✓ Petroleum; A liquid fossil fuel formed from the remains of marine organisms. It consists of paraffin, petrol , diesel etc.
- ✓ Biofuels; Fuels derived from biomass, such as ethanol and biodiesel.

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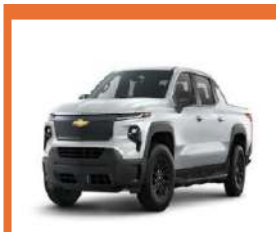
Main uses of carbon based fuels (charcoal, paraffin, petrol, firewood, and diesel)

 Diesel	 Paraffin	 Charcoal	 Petrol	 Firewood
Vehicle fuel (trucks, buses), Industrial equipment, Generators, Construction machinery, Marine fuel, Heating oil, Agricultural equipment.	Cooking fuel (stoves, heaters), Lighting (lanterns, candles), Industrial solvents, Cosmetics (moisturizers, lotions), Fuel additives, Lubricants.	Cooking fuel (barbecues, braais), Water filtration, Air purification, Skincare products, Art supplies (drawing, painting), Odors control, Industrial applications (metal production)	Vehicle fuel (cars, trucks), Airplane fuel, Industrial machinery, Generators, Cleaning solvents, Paint thinners, Chemical production.	Heating (homes, fireplaces), Cooking (wood-fired ovens), Camping fuel, Paper production, Pulpwood, Furniture manufacturing, Smoking meats.

Comparative Analysis

Fuel	Energy Density	Environmental Impact
Charcoal	Low	Moderate
Paraffin	Medium	High
Petrol	High	High
Firewood	Low	Moderate
Diesel	High	High

Why carbon compounds are used as fuels

- ✓ High energy content; Carbon compounds, especially hydrocarbons, store a significant amount of energy in their chemical bonds. When these bonds are broken during combustion, a large amount of heat and light is released.
- ✓ Abundance; Carbon-based fuels, such as coal, natural gas, and petroleum, are abundant and widely available.
- ✓ Easy to handle and transport; Many carbon compounds are liquids or gases at room temperature, making them easy to handle and transport.
- ✓ Versatile applications; Carbon-based fuels can be used for various purposes, including heating, cooking, transportation, and electricity generation

Renewable vs. Non-Renewable Fuels

Renewable sources replenish naturally within a human timescale, while non-renewable sources are finite and take millions of years to form.

Activity Title: Fuel Frenzy: Renewable vs. Non-Renewable
Learning Objectives: By the end of this activity, learners will be able to:

- ✓ Define renewable and non-renewable energy sources.
- ✓ Identify examples of renewable and non-renewable fuels.
- ✓ Classify given fuels as renewable or non-renewable.
- ✓ Explain the advantages and disadvantages of using each type of fuel.
- ✓ Discuss the environmental impact of fuel choices.

Materials:

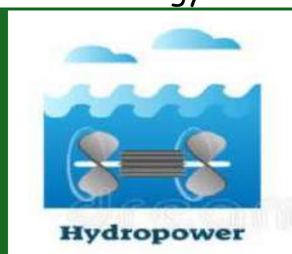
- ✓ Large sheets of paper or a whiteboard
- ✓ Marker
- ✓ Pictures or cut-outs of various fuels (coal, oil, natural gas, solar panels, wind turbines, hydroelectric dams, biomass, geothermal, etc.) You can also use drawings or simple labels if pictures are unavailable.

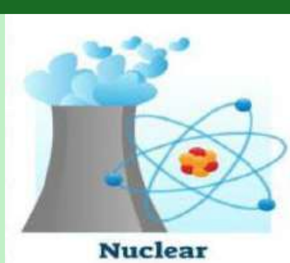
- ✓ Two large containers or boxes labelled "Renewable Fuels" and "Non-Renewable Fuels"
- ✓ Real samples of some fuels (e.g., small pieces of coal, wood, etc.) for observation (ensure safety precautions are taken).

Procedure:

<u>Ask probing questions.</u>

What is energy?", "Where does energy come from?", "What do we use energy for?".





- ✓ In groups, brainstorm and list as many examples of fuels as you can think of. Think beyond just car fuel – consider fuels used for cooking, heating, electricity generation, etc.
- ✓ Share your list with the class.
- ✓ Individually, choose a fuel picture and place it in the correct container (Renewable or Non-Renewable).

Probing questions

Why do you think this fuel is renewable/non-renewable?", "What are some of its uses?", "Does it have any impact on the environment?". Justify your choices.
What are the benefits of using renewable energy?", "What are the drawbacks of relying on non-renewable energy?", "Why is it important to consider the environmental impact of our fuel choices?".

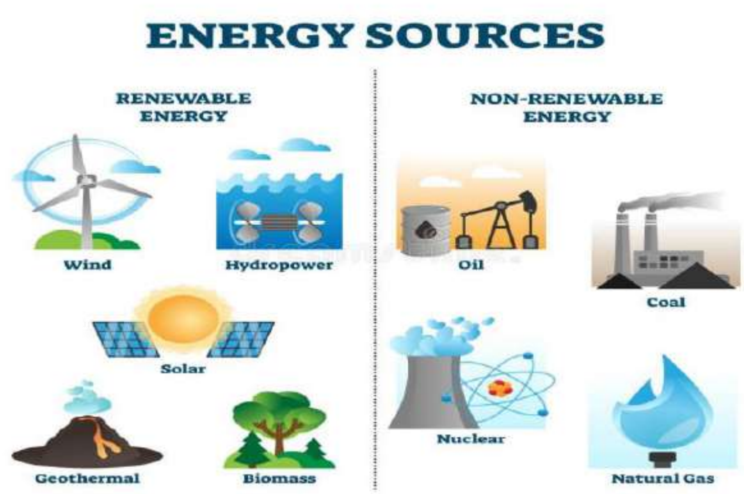
Renewable fuels; are energy sources that can be replenished naturally within a human lifespan. These fuels are naturally replenished at a rate comparable to or faster than their rate of consumption.

Renewable Fuels include;

- ✓ Solar energy; harnessing sunlight to generate electricity.
- ✓ Wind energy; Converting wind power into electricity.
- ✓ Hydropower; using the force of flowing water to generate electricity.
- ✓ Biofuels; Fuels derived from organic matter, such as ethanol and biodiesel.
- ✓ Geothermal energy; utilizing heat from the Earth's interior to generate electricity.

Key Characteristic: Sustainable, with minimal environmental impact.

Non-renewable fuels; are finite resources that cannot be replenished at a rate that keeps pace with consumption. These fuels are formed over millions of years from the remains of dead organisms.



Non-Renewable Fuels include;

✓ Fossil fuels; Coal, oil, and natural gas, formed from the remains of ancient organisms.
✓ Nuclear energy; using nuclear fission to generate electricity.

Key Characteristic: Limited supply, contribute significantly to climate change due to greenhouse gas emissions.

Did you know that according to current consumption rate.

General Estimates

- ✓ Coal reserves could last for over 100 years. However, this assumes no significant increase in coal use.
- ✓ Oil reserves are estimated to last for approximately 50-60 years at current consumption rates.
- ✓ Natural Gas reserves are estimated to last for approximately 50-60 years at current consumption rate.

Comparison

Feature	Renewable Fuels	Non-Renewable Fuels
Availability	Abundant and constantly replenished	Limited and finite
Environmental impact	generally lower emissions	Contribute to climate change

Economic impact	Can create jobs and reduce dependence on foreign oil	Can lead to economic dependence and price fluctuations
Technological development	require ongoing research and development	Well-established technology

Fuel sustainability

Sustainability is the ability to meet the needs of the present without compromising the ability of future generations to meet their own needs. In the context of fuel use, it encompasses three key pillars:

1.**Environmental Sustainability:** Minimizing environmental impacts like pollution (air, water, soil), greenhouse gas emissions, and resource depletion.

2.**Social Sustainability:** Ensuring equitable access to energy, promoting social justice, and considering the impacts on local communities.

3.**Economic Sustainability:** Ensuring the long-term viability of energy systems, including affordability and economic growth.

Making Fuel Use in the Locality More Sustainable

- ✓ Transition to Renewable Energy Sources like Solar, wind, hydro, geothermal, and biomass.
- ✓ Improve Energy Efficiency by encouraging the use of energy-efficient lighting, appliances, and transportation, Implement energy-efficient building codes and regulations.
- ✓ Reduce Reliance on Fossil Fuels i.e Phase out the use of highly polluting fuels but promote the use of cleaner-burning fuels like natural gas (if available) as a transitional fuel.

- ✓ Encourage cycling and walking i.e create safe and accessible cycling and walking paths, promote the use of electric vehicles
- ✓ Carbon Capture and Storage: Capture and store carbon emissions from fuel use to reduce greenhouse gas emissions, and utilize waste materials as fuel sources through waste-to-energy technologies.

Future Outlook

The future of fuel sustainability lies in a diverse energy mix that includes renewable energy sources, energy efficiency measures, and advanced technologies like carbon capture and storage. By transitioning to cleaner and more sustainable fuels, we can reduce our reliance on fossil fuels, mitigate climate change, and ensure a secure energy future.

Environmental Impacts of Burning Carbon-Based Fuels

Burning carbon-based fuels, also known as fossil fuels, has significant environmental impacts, primarily due to the release of greenhouse gases and other pollutants. Here's a breakdown of the key effects:

Climate Change:

The most significant impact is the contribution to climate change. When fossil fuels are burned, they release large amounts of carbon dioxide (CO_2) into the atmosphere, which is the primary greenhouse gas. Greenhouse gases trap heat in the atmosphere, leading to global warming and climate change. This results in rising temperatures, changes in weather patterns, sea level rise, and more frequent extreme weather events.

Mitigation:

- ✓ Shift to Renewable Energy: This is the biggest one. Solar, wind, hydro, geothermal, and biomass energy sources produce far fewer greenhouse gases.

- Improve Energy Efficiency: Using less energy in the first place reduces the need to burn fossil fuels. This means better insulation in homes, more efficient appliances, and smarter transportation systems.
- Carbon Capture and Storage (CCS): Technologies that capture CO_2 emissions from power plants and industrial sources, then store it underground. Still under development, but promising.
- Reduce Methane Emissions: Methane is a potent greenhouse gas, and leaks from natural gas infrastructure are a major source. Better leak detection and repair are crucial.

Air Pollution:

Burning fossil fuels releases various air pollutants, including, Sulphur dioxide (SO_2), which contributes to acid rain and respiratory problems, Nitrogen oxides (NOx), which contribute to smog and respiratory issues, Particulate matter, which can cause respiratory and cardiovascular problems, Ground-level ozone, a major component of smog, which can irritate the lungs.

Mitigation:

- Cleaner Fuels: Switching to natural gas (which burns cleaner than coal or oil) or using fuels with lower sulphur content.
- Pollution Control Technologies: Scrubbers in power plants to remove SO_2, catalytic converters in cars to reduce NOx, and filters to capture particulate matter.
- Reduce Vehicle Use: Promoting public transportation, cycling, and walking reduces the number of cars on the road and their associated emissions.

Water Pollution:

Fossil fuel extraction and transportation can lead to water pollution through oil spills, leaks, and runoff. Fracking, a method of extracting natural gas, can contaminate groundwater with chemicals and methane.

Acid rain, caused by SO_2 and NOx emissions, can acidify lakes and streams, harming aquatic life.

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Mitigation:

- ✓ Stricter Regulations: Preventing oil spills from tankers and offshore drilling platforms.
- ✓ Responsible Fracking Practices: Ensuring proper well construction and waste management to prevent groundwater contamination.
- ✓ Treating Wastewater: Removing pollutants from industrial wastewater before it's released into waterways.

Ocean Acidification:

The ocean absorbs a significant portion of the CO_2 released from burning fossil fuels. This absorption leads to ocean acidification, which can harm marine life, particularly shellfish and coral reefs.

Mitigation:

- ✓ Reduce CO_2 Emissions: The most direct way to address ocean acidification is to reduce the amount of CO_2 released into the atmosphere from burning fossil fuels. This ties back to the climate change solutions.
- ✓ Explore Geoengineering: Some scientists are exploring ways to remove CO_2 from the atmosphere or reduce ocean acidity directly, but these are still experimental and may have unintended consequences.

Other Environmental Impacts:

- ✓ Habitat destruction due to mining and drilling for fossil fuels.

Reclamation and Restoration: Restoring areas damaged by mining or drilling.

Minimize Habitat Destruction: Careful planning of mining and drilling operations to reduce their footprint.

- ✓ Noise pollution from drilling and transportation activities.

Noise Reduction: Using quieter technologies and implementing noise barriers..

Process of making charcoal

Definition: Charcoal is a lightweight, porous, and black solid material obtained from the incomplete combustion of organic materials, such as wood, bamboo, or coconut shells.

It has a high surface area and Porosity, good adsorption capacity, high thermal insulation and Chemical reactivity.

Composition: Charcoal primarily consists of:

- ✓ Carbon (85-98%)
- ✓ Volatile matter (2-15%)
- ✓ Ash (1-5%)
- ✓ Moisture (1-5%)

Types of Charcoal

- ✓ Wood charcoal
- ✓ Bamboo charcoal
- ✓ Coconut shell charcoal
- ✓ Activated charcoal
- ✓ Bio char (Charcoal made from organic wastes)

Uses:

- ✓ Fuel for cooking and heating

- ✓ Water filtration and purification
- ✓ Air purification
- ✓ Medical applications (e.g., detoxification)
- ✓ Industrial applications (e.g., metal production)
- ✓ Art and crafts

Interesting Facts:

- ✓ Charcoal is an ancient material (used by Egyptians and Greeks)
- ✓ Charcoal can be used as a natural fertilizer
- ✓ Activated charcoal can trap toxins and odors
- ✓ Charcoal production supports local economies

Process of making wood charcoal

Charcoal production involves the thermal decomposition of organic materials, such as wood, in the absence of oxygen. This process is called pyrolysis.

- ✓ The organic material is prepared by cutting it into small pieces and stacking it into piles or kilns.
- ✓ The piles or kilns are ignited using a small amount of flammable material. As the temperature rises, the organic material undergoes pyrolysis, breaking down into charcoal, tar(viscous, dark liquid mixture of hydrocarbons), and gases.
- ✓ Once the pyrolysis process is complete, the charcoal is allowed to cool before being removed from the piles or kilns.

Charcoal is a valuable fuel source and is also used in a variety of other applications, such as: Barbecue grilling, Cooking, Metallurgy, Water purification, Art and crafts.

Note: The specific methods and techniques used for charcoal production may vary depending on the region and the type of organic material being used.

Benefits of Charcoal Burning

Charcoal burning, while contributing to deforestation and air pollution, offers several benefits:

- ✓ Renewable energy; Charcoal is a renewable energy source, as it can be produced from sustainably managed forests.
- ✓ Versatile fuel; Charcoal can be used for cooking, heating, and industrial processes.
- ✓ Traditional craft; Charcoal production is a traditional craft in many cultures, providing livelihoods and preserving cultural heritage.
- ✓ Soil amendment; Charcoal can be used as a soil amendment to improve soil fertility and water retention.
- ✓ Water filtration; Charcoal can be used to filter water, removing impurities and contaminants.

Environmental Impacts of Charcoal Burning

Charcoal burning, while providing benefits, also has significant environmental impacts:

Greenhouse Gas Emissions:

The process of charcoal production and burning releases greenhouse gases, including carbon dioxide (CO_2), methane (CH_4), and nitrous oxide (N_2O), which contribute to global warming and climate change.

Traditional charcoal production methods are often inefficient, meaning a significant portion of the wood's energy is lost as emissions rather than being converted to charcoal.

Air Pollution:

Charcoal burning releases various air pollutants, including particulate matter, carbon monoxide, and volatile organic compounds (VOCs), which can contribute to respiratory problems and other health issues. These pollutants can also contribute to the formation of smog, which further degrades air quality.

Soil Degradation:

Deforestation for charcoal production can lead to soil erosion, as the removal of trees exposes the soil to wind and rain. This erosion can result in the loss of topsoil and essential nutrients, making it difficult for vegetation to regenerate.

Biodiversity Loss:

The clearing of forests for charcoal production destroys habitats for countless plant and animal species, leading to a decline in biodiversity. This loss of biodiversity can disrupt ecological balance and have cascading effects on entire ecosystems.

Water Resources:

Deforestation can disrupt local water cycles, leading to changes in rainfall patterns and reduced water availability.

Charcoal production sites can also contribute to water pollution through runoff and soil erosion.

Health Impacts:

Exposure to air pollutants from charcoal burning can cause or exacerbate respiratory problems, such as asthma and bronchitis.

Making Briquettes from Waste Organic Materials

Briquettes can also be made from waste organic materials, such as sawdust, agricultural residues, and food waste. This process can help reduce waste and create a sustainable fuel source.

Specific Types of Waste Organic Materials for Briquettes

- ✓ Agricultural residues; Sawdust, rice husks, corn cobs, and coconut shells.
- ✓ Food waste; Fruit and vegetable peels, coffee grounds, tea leaves, and eggshells.
- ✓ Municipal solid waste; Paper, cardboard, and yard waste.
- ✓ Industrial waste; Biomass from agricultural and food processing industrial

Materials

- ✓ Waste organic materials (e.g., sawdust, rice husks, coconut shells)
- ✓ A binding agent e.g., starch from cassava flour.
- ✓ Water
- ✓ A mixing device (e.g., hammer mill, blender)
- ✓ A briquette machine or press
- ✓ A drying rack or oven

Procedure

- ✓ Waste organic materials are collected and prepared by drying, grinding, and sometimes composting.
- ✓ The prepared materials are mixed with binders, such as starch or molasses, to improve their shape and strength.
- ✓ The mixture is moulded into briquettes using a hydraulic press.
- ✓ The briquettes are dried to remove excess moisture.

The resulting briquettes can be used as a fuel source for cooking, heating, or industrial processes.

Benefits of making briquettes from waste organic materials include

- ✓ Waste reduction; Reduces the amount of organic waste going to landfills.
- ✓ Sustainable fuel; Provides a renewable and sustainable fuel source.

- ✓ Improved soil health; composting waste organic materials can improve soil health.
- ✓ Reduced greenhouse gas emissions; using briquettes made from waste organic materials can help reduce greenhouse gas emissions compared to burning fossil fuels.

However, it's important to note that the quality and performance of briquettes made from waste organic materials may vary depending on the specific materials used and the production process.

Chemistry of Incomplete Combustion

Incomplete combustion happens when there isn't enough oxygen available for a fuel to burn completely. This is in contrast to complete combustion where there's plenty of oxygen, and the fuel reacts fully to produce only carbon dioxide (CO_2) and water (H_2O). The reactants are still a fuel (like methane, CH_4, organic material like wood) and oxygen O_2, limited.

This process involves:

Pyrolysis

Heat breaks chemical bonds in the organic material, releasing volatile gases, Carbon-rich residues (soot, Char (C)) and liquids).

Organic material → Volatile gases (CO, CH_4, VOCs) + Carbon-rich residues (soot, Char (C))

Char is the solid residue, which is the charcoal.

Partial Oxidation

Because there's not enough oxygen to go around, the carbon in the fuel can't fully react to form CO_2. Instead, you get a mix of products, including:

- ✓ Carbon monoxide (CO) formation:

$C + O \rightarrow CO$

- ✓ Methane (CH_4) formation
  $C + 2H_2 \rightarrow CH_4$
- ✓ Volatile organic compounds (VOCs) formation
  $C_6H_{12}O_6 \rightarrow C_2H_4O + CO + H_2O$

Cracking and Reforming

High temperatures break complex molecules:

- ✓ Cracking
  Large molecules → Smaller molecules (e.g., alkanes, alkenes)
- ✓ Reforming
  Smaller molecules → CO, H_2, CH_4

Carbon Dioxide

Molecular formula: **CO_2**

Molecular weight:**44.01 g/mol**

Density: **1.96 kg/m³ (gas), 1.57 g/cm³ (solid)**

Category: **An acidic oxide and a covalent compound formed by mutual sharing of Valence electrons between two non-metal atoms (carbon and oxygen).**

Laboratory test: Can be identified using a burning splint and lime water ($Ca(OH)_2$).

Activity 1: The burning splint test

Aim: To demonstrate that carbon dioxide extinguishes a flame, and to use this property to identify the gas.

Materials:

- ✓ Source of Carbon Dioxide (React Baking Soda and Vinegar)
- ✓ Test tube
- ✓ Burning Splint
- ✓ Lighter or Matches

Procedure:

- ✓ Put some baking soda in the container, and slowly add vinegar. The fizzing will produce carbon dioxide gas.
- ✓ Hold the opening of the container producing the CO_2 near the mouth of the test tube to collect the gas.
- ✓ Light the splint and let it burn for a few seconds.
- ✓ Quickly insert the glowing splint into the test tube containing the collected gas.
- ✓ Observe what happens to the glowing splint.

Discussion:

It extinguishes a burning splint and forms a white precipitate with lime water. The glowing splint is extinguished because carbon dioxide does not support combustion (burning). This is a characteristic property of carbon dioxide and can be used to identify it as compared to how a glowing splint relights in the presence of oxygen.

Activity 2: The Classic Limewater Test

Aim: To demonstrate the presence of carbon dioxide through the reaction with limewater.

Materials:

- ✓ Limewater (calcium hydroxide solution) - Prepared by dissolving calcium hydroxide in water and filtering the solution. It should be clear.
- ✓ Test tubes or small clear bottles

- ✓ Straws (for blowing) or a source of CO_2 (e.g., baking soda and vinegar in a separate container)
- ✓ Rubber stoppers or Parafilm to seal the test tubes

For your understanding

Carbon dioxide is a gas we exhale and is also produced by various chemical reactions.

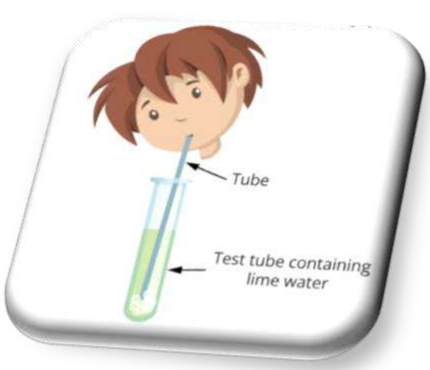
A chemical test is like a detective's tool for scientists! It's a way to find out what substances are present in a sample by observing how they react with other chemicals.

Different substances have unique ways of reacting with other chemicals. These reactions can produce visible changes, like a change in color, the formation of a solid (precipitate), or the release of a gas.

Scientists use these reactions as clues to identify the substances they are testing for. Just like a detective looks for clues at a crime scene, scientists look for specific changes in a chemical test.

Procedure

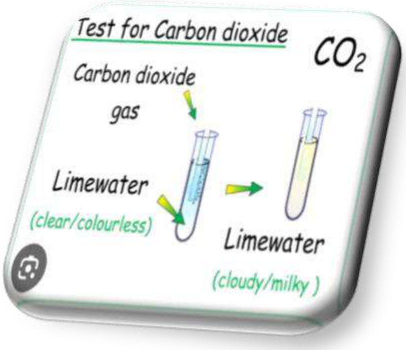
Alternative 1



- ✓ Pour a small amount of limewater into each test tube.
- ✓ Gently blow through a straw into the limewater in one test tube for a minute or two. Do not suck the limewater back into your mouth.

Alternative 2

- ✓ In a separate container, mix baking soda and vinegar to produce CO_2 gas.



- ✓ Direct the gas into the test tube containing limewater (you can use tubing or just hold the opening of the CO_2 generating container near the test tube opening).
- ✓ Observe what happens to the limewater in the test tube you're blowing into or that's exposed to the baking soda/vinegar reaction.

Observation

The limewater appears milky.

Carbon dioxide reacts with calcium hydroxide to form calcium carbonate, which is an insoluble white precipitate, making the limewater appear milky. The formation of a white precipitate is due to the formation of an insoluble calcium carbonate.

The limewater test is a common method for detecting carbon dioxide.

Properties of carbondioxide

- ✓ Colorless, odorless gas
- ✓ Soluble in water (forms carbonic acid)
- ✓ Non flammable gas
- ✓ Reacts with bases (forms carbonates)
- ✓ Stable at room temperature

Sources

- ✓ Atmosphere (0.03%)
- ✓ Fossil fuel combustion (60%)
- ✓ Deforestation and land-use changes (15%)
- ✓ Industrial processes (10%)
- ✓ Natural sources (volcanic eruptions, respiration)

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Some useful properties of carbon dioxide

- **Low Melting and Boiling Point hence;**

1. Used as a refrigerant in refrigeration systems, air conditioning systems, and heat pumps due to its low boiling point.

2. Used as a cooling agent in fire extinguishers, due to its ability to rapidly expand and cool the surrounding area.

- **Does Not Conduct Electricity thus;**

1. Used as an electrical insulation gas in high-voltage electrical equipment, such as switchgear and transformers.

2. Used as a laser cutting gas, due to its ability to provide a clean, non-conductive cutting environment.

- **Exists as Solid, Liquid, and Gas,** Carbon dioxide exists as a solid (dry ice) at low temperatures, and is used for cooling, refrigeration, and special effects.
- **Non-flammable** Making it safe for use in applications like welding, fire extinguishing, and chemical processing.
- **High solubility in water hence** useful for applications like carbonated beverages and oil recovery.

Laboratory Preparation of Carbon Dioxide

Aim: To prepare carbon dioxide gas in the laboratory using marble chips and dilute hydrochloric acid.

Hypothesis: Marble chips (calcium carbonate) react with dilute hydrochloric acid, to produce carbon dioxide gas.

Variables

Independent variable- Concentration of hydrochloric acid

Dependent variable- Volume of carbon dioxide gas produced

Controlled variables- Mass of marble chips, temperature, and pressure

Risks

Chemical splashes- Hydrochloric acid can cause skin and eye irritation.

Eye damage- Hydrochloric acid can cause eye damage.

Respiratory problems- Inhaling carbon dioxide gas can cause respiratory problems.

Mitigation

Wear protective gear- Wear gloves, safety goggles, and lab coats.

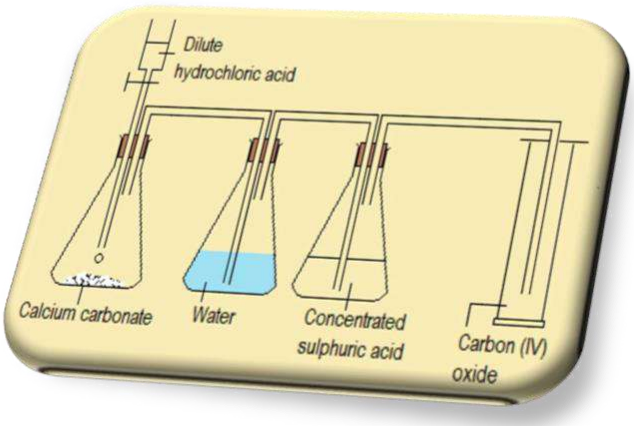
Handle chemicals carefully- Use tongs or pipettes to handle hydrochloric acid.

Clean up spills immediately- Use paper towels or a cleaning solution to clean up spills.

Conduct the experiment in a well-ventilated area- Prevent inhalation of carbon dioxide gas.

Materials

- ✓ Marble chips or limestone
- ✓ Dilute hydrochloric acid
- ✓ Flat bottomed flask
- ✓ Delivery tube
- ✓ Thistle funnel
- ✓ Gas jar
- ✓ Wash Bottles



Procedure

- ✓ Place marble chips or limestone in the Flat bottomed flask
- ✓ Use the delivery tubes and connect Flat bottomed flask to a wash bottle containing water or potassium hydrogen carbonate solution and to another was bottle containing concentrated sulphuric acid, it's then connected to a gas jar to collect the gas.
- ✓ Add dilute hydrochloric acid to the Flat bottomed flask using a thistle funnel.
- ✓ Observe the reaction and the formation of carbon dioxide gas.
- ✓ Lower a lightened splint into a gas jar containing the collected gas.

Discussion

Calcium carbonate (CaCO3) from the marble chips reacts with hydrochloric acid (HCl) to form calcium chloride (CaCl2), carbon dioxide (CO2) gas, and water (H2O), the gas moves to a wash bottle containing water. The purpose of passing the gas through water is to absorb any acid traces that did not react. It's then passed to another wash bottle containing concentrated Sulphuric acid to absorb water or dry the gas.

The gas is then collected in the gas jar by downward delivery since it's denser than air. The carbon dioxide gas produced can be collected and used for various experiments or demonstrations.

When a lightened splint is introduced in to the jar, the splint is extinguished since Carbondioxide is a non flammable gas and doesn't support burning.

Equation

$CaCO_3$ (marble chips) + 2HCl (hydrochloric acid) → $CaCl_2$ (calcium chloride) + CO_2 (carbon dioxide) + H_2O (water)

Alternative methods

Heating sodium bicarbonate

$NaHCO_3(s) \rightarrow Na_2CO_3(s) + CO_2(g) + H_2O(l)$

Carbon Dioxide's Impact on Global Warming

Carbon dioxide (CO_2) plays a crucial role in regulating Earth's temperature. Excessive CO_2 emissions lead to global warming as seen below.

The Greenhouse Effect

- ✓ CO_2 traps infrared radiation
- ✓ Prevents heat escape into space
- ✓ Retains warmth in atmosphere

Ocean Warming

- ✓ CO_2 absorbs heat (93% of excess heat) leading to thermal expansion (sea-level rise) and ocean acidification (decreased pH)

Climate Effects

The greenhouse effect and ocean warming may result into;

- ✓ Extreme weather events thus droughts and heatwaves
- ✓ Heavy precipitation
- ✓ Sea-level rise

These climate effects will in turn lead to;

- ✓ Ecosystem disruption
- ✓ Biodiversity loss
- ✓ Food security threats

Mitigation Strategies

- ✓ Reduce fossil fuel emissions
- ✓ Renewable energy sources
- ✓ Carbon capture and storage
- ✓ Energy efficiency
- ✓ Sustainable land-use practices

Greenhouse Gases(GHGs)

What are Greenhouse Gases?

Greenhouse gases are gases in the Earth's atmosphere that trap heat from the sun. This trapped heat warms the planet, making it habitable. However, when these gases are in excess, they absorb and re-emit infrared radiation from the Earth's surface, contributing to the greenhouse effect and so global warming. This warming has significant effects on the climate.

Some examples of greenhouse gases include:

- ✓ **Carbon dioxide (CO_2)**

  A colorless, odorless gas that is a natural part of the Earth's atmosphere. Human activities have increased the amount of CO_2 in the atmosphere.

- ✓ **Methane (CH_4)**

A primary component of natural gas that is more than 25 times more effective at trapping heat than CO_2. Methane emissions come from livestock and waste management, landfills, and the production and transportation of fossil fuels.

- ✓ **Nitrous oxide (N_2O)**

  Also known as laughing gas, nitrous oxide is a potent greenhouse gas that can last for over 100 years in the atmosphere. Human activities that release nitrous oxide include the use of nitrogen fertilizers and animal waste on farmland.

- ✓ **Water vapor**

  The most abundant greenhouse gas on Earth, responsible for about half of the greenhouse effect. As temperatures increase, more water evaporates and remains in the lower atmosphere, absorbing radiation and pushing it down to the Earth's surface.

- ✓ **Hydrofluorocarbons (HFCs)**

  Artificial compounds that are used in air conditioners, refrigerators, and other industrial processes. HFCs are hundreds to thousands of times more potent than carbon dioxide.

Other greenhouse gases include Ozone, Chlorofluorocarbons (CFCs),sulphur hexafluoride and nitrogen trifluoride

Effects of Greenhouse Gas Emissions:

Greenhouse Gas	Sources	Effects	Mitigation Strategies
Sulphur dioxide (SO2)	Fossil fuel combustion	Causes acid rain which results to corrosion, reduction	Catalytic converters on gas exhaust pipes/ chimneys.

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		in soil pH affecting crop yields.	
Carbon Dioxide (CO_2)	Fossil fuel combustion, deforestation, land-use changes	Global warming which results to climate change hence floods, droughts.	Transition to renewable energy, energy efficiency, carbon capture, afforestation.
Methane (CH_4)	Agriculture, natural gas production, waste management	Global warming thus climate change hence Sea-level rise, ecosystem disruption	Sustainable agriculture, waste reduction, methane capture
Nitrous Oxide (N_2O)	Agriculture, industrial processes, fossil fuel combustion	Global warming, climate change hence Ocean acidification, biodiversity loss	Efficient fertilizer use, catalytic converters on gas chimneys.
Fluorinated Gases (F-gases)	Air conditioning, refrigeration, electrical equipment	Extreme weather events, ecosystem disruption	Phase-down under Montreal Protocol, alternative technologies
Ozone (O_3)	Atmospheric reaction of nitrogen oxides and VOCs	Respiratory problems, crop damage	Reduce nitrogen oxides and VOCs emissions

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Chlorofluorocarbons (CFCs)	Refrigeration, air conditioning, propellants like insecticide sprays.	Ozone depletion, global warming hence Ozone layer damage, climate change	Phase-out under Montreal Protocol, alternative technologies

Note

Global Warming: Increased **greenhouse gas** concentrations **trap** more **heat**, **leading** to **rising global temperatures** resulting into **Climate Change**, Changes in weather patterns, including more frequent and intense heatwaves, droughts, floods, and storms.

Rising Sea Levels: Melting glaciers and ice sheets contribute to sea-level rise, threatening coastal cities and islands.

Ocean Acidification: Increased CO_2 absorption by oceans leads to acidification, **harming** marine **ecosystems**.

Disruption of Ecosystems: Climate change alters ecosystems, **affecting biodiversity** and food chains.

Human Health Impacts: Heat-related illnesses, respiratory problems, and waterborne diseases can increase.

Hardness of water

Hard water is water that contains high levels of dissolved minerals, primarily calcium and magnesium ions. These minerals can react with soap to form a soap scum, which can make it difficult to lather and rinse the scum also leaves stains on clothing.

Hard water can also leave mineral deposits on surfaces, such as plumbing fixtures and appliances.

Soft water is water that has low concentrations of minerals like calcium and magnesium, and is often compared to hard water, which has high levels of these minerals.

Soft water can occur naturally in areas where runoff and drainage basins are formed by rough, impermeable, calcium-poor rocks. It can also be created through a water softening process, which can result in water with high levels of sodium and bicarbonate ions.

	Soft water	**Hard water**
Mineral content	Low in calcium and magnesium	High in calcium and magnesium
Examples	Rainwater	Groundwater
Effects	Lathers easily with soap, leaves dishes clean and shiny	Deposits calcium carbonate on pipes, making soap and detergents less effective

Sources of soft and hard water

Hard water

Lakes, wells, rivers, swamp water.

Soft water

Rain water.

How Rainwater Becomes Hard Water

Rainwater, as it falls from the sky, is naturally soft. This means it has low mineral content. Rain water combines with carbondioxide in the atmosphere forming carbonic acid.

Water + Carbondioxide ⟶ Carbonic acid

As this water containing carbonic acid makes its way through the soil and rock layers, it picks up dissolved minerals, primarily calcium carbonate(limestone) and magnesium sulphate. The dissolved minerals increase the water's hardness.

Carbonic acid + Calcium Carbonate ⟶ Calcium Hydrogen Carbonate

Effects on Soap

- ✓ Reduced lather formation: Calcium and magnesium ions interfere with soap's ability to create lather.
- ✓ Soap scum formation: Minerals react with soap, creating insoluble deposits.
- ✓ Decreased cleaning effectiveness: Hard water ions reduce soap's ability to emulsify oils.
- ✓ Increased soap consumption
- ✓ Spotting on surfaces

Hard Water and Soap Reaction

When soap is added to water, it dissolves and breaks down into its constituent ions. Hard water contains high levels of minerals (calcium and magnesium), which react with soap forming a less Soluble substance which starts to precipitate out of solution, forming a sticky, insoluble substance called soap scum.

Chemical Reaction

Calcium/Magnesium ions ($Ca2+/Mg2+$) + Soap (fatty acid salts) → Insoluble precipitate (soap scum)

The soap scum formed reduces the amount of soap available for lather formation. As a result, the soap solution becomes less effective at creating a rich lather.

Despite the formation of soap scum, some soap ions are still available to interact with water and form lather. However, the lather may not be as rich or abundant as it would be in soft water.

Causes of hardness in water

Hardness in water is caused by the presence of dissolved minerals, particularly calcium and magnesium. These minerals can react with soap to form a precipitate, making it difficult to create a rich lather.

There are several causes of hardness in water, including:

- ✓ **Geological Sources:** Hardness in water can come from geological sources, such as rocks and soil. Calcium and magnesium ions can leach into groundwater from limestone, dolomite, and other mineral-rich rocks.
- ✓ **Mineral Dissolution:** Hardness in water can also come from the dissolution of minerals in water. For example, calcium carbonate (limestone) can dissolve in water to form calcium ions.
- ✓ **Agricultural Runoff:** Hardness in water can also come from agricultural runoff. Fertilizers and other agricultural chemicals can contain calcium and magnesium ions, which can run off into waterways and contribute to hardness.
- ✓ **Industrial Wastewater:** Hardness in water can also come from industrial wastewater. Some industries, such as mining and manufacturing, can release calcium and magnesium ions into waterways.
- ✓ **Atmospheric Deposition:** Hardness in water can also come from atmospheric deposition. Calcium and magnesium ions can be deposited into waterways through atmospheric processes, such as precipitation and dry deposition.

Types of Hardness in Water

There are two main types of hardness in water:

Temporary Hardness

Temporal hardness, also known as carbonate hardness or temporary hardness, measures water's calcium and magnesium bicarbonate content.

<u>Sources of Water with Temporary Hardness</u>

- ✓ Springs
- ✓ Well water
- ✓ Groundwater
- ✓ Surface water (lakes, rivers)
- ✓ Municipal water supplies (in affected regions)

Removal Methods

- ✓ Boiling (heat decomposition)
- ✓ Lime precipitation (chemical treatment)
- ✓ Ion exchange (water softeners)
- ✓ Reverse osmosis

Permanent hardness

Permanent hardness is caused by the presence of calcium and magnesium sulphates, chlorides, and nitrates in water.

Permanent hardness measures water's non-carbonate mineral content.

Sources of Water with permanent Hardness

- ✓ Groundwater (aquifers with high mineral content)
- ✓ Surface water (lakes, rivers with high mineral runoff)
- ✓ Industrial processes (mining, manufacturing)
- ✓ Agricultural activities (fertilizers, pesticides)
- ✓ Municipal water supplies (in affected regions)

Removal Methods

- ✓ Ion exchange (water softeners)
- ✓ Reverse osmosis
- ✓ Distillation
- ✓ Chemical treatment (lime precipitation)

Effects of Hardness in Water

- ✓ Scaling: Hardness in water can cause scaling, which is the formation of a hard, mineral-rich deposit on surfaces like kettles, plumbing pipes, taps.
- ✓ Spotting on surfaces
- ✓ Reduced Soap Effectiveness: Hardness in water can reduce the effectiveness of soap and detergent since it reacts with soap to form a soap scum, which can make it difficult to lather and rinse
- ✓ Increased Energy Consumption: Hardness in water can increase energy consumption by reducing the efficiency of appliances.
- ✓ Appliance Damage: Hardness in water can cause damage to appliances, such as water heaters and dishwashers.

Softening Hard Water

Hard water, rich in minerals like calcium and magnesium, can cause various inconveniences. Water softening techniques are employed to reduce the mineral content, making it more suitable for domestic and industrial use.

The method of softening depends on the type of hardness in water

An experiment to demonstrate the softening of hard water

Materials

- ✓ Borehole water
- ✓ Rain water
- ✓ Soap solution
- ✓ Limewater

Procedure

- ✓ Measure out $50cm^3$ of borehole water using a measuring cylinder.
- ✓ Add a few drops of soap solution to a sample of borehole water while shaking.
- ✓ Observe and note what happens.
- ✓ Repeat the above steps using rain water.

Identify which of the two water samples is hard water.

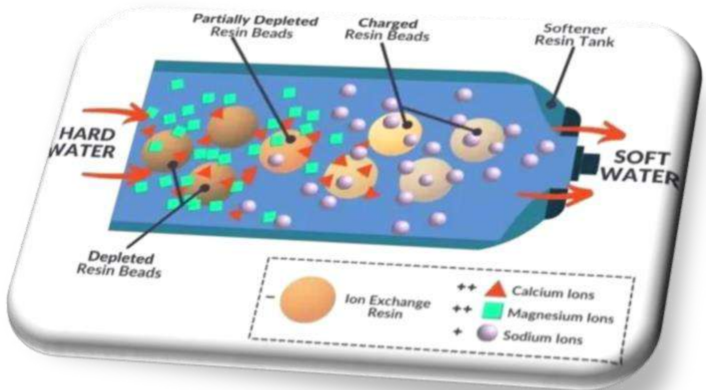
Note; Hard water forms soap scum due to the reaction between calcium and magnesium ions and soap.

- ✓ Measure out $50cm^3$ of hard water using a measuring cylinder.
- ✓ Add limewater to a sample of hard water and add a few drops of soap solution while shaking.
- ✓ Observe the formation of a precipitate and note the amount of soap used to form lather
- ✓ Repeat the above steps using rain water.

Conclusion: Limewater reacts with calcium and magnesium ions in hard water to form insoluble calcium carbonate and magnesium hydroxide precipitates, effectively softening the water.

Methods of Water softening

Lime-soda process



<u>Principle:</u> This chemical process involves adding lime (calcium hydroxide) and soda ash (sodium carbonate) to hard water.

<u>Reaction:</u> The added chemicals react with the calcium and magnesium ions, forming insoluble **precipitates** that can be removed through sedimentation and filtration.

<u>Advantages:</u> Effective for high hardness levels.

<u>Disadvantages</u>: Requires careful chemical dosing and sludge disposal.

Ion exchange process

<u>Principle:</u> This method uses ion exchange resins, which are synthetic materials with a high capacity to exchange ions.

<u>Process:</u> Hard water is passed through a bed of ion exchange resin. The resin exchanges its sodium ions for the calcium and magnesium ions in the water, softening it.

<u>Advantages:</u> Effective for a wide range of hardness levels, easy to operate.

<u>Disadvantages:</u> Requires periodic regeneration with a brine solution.

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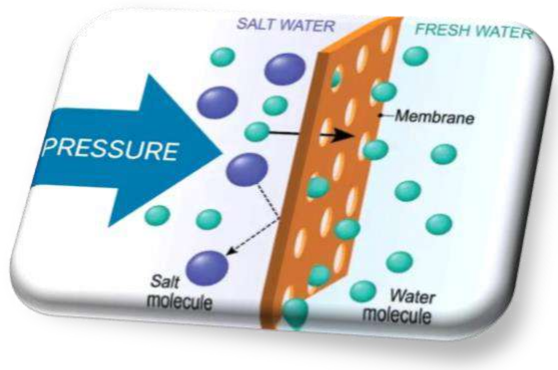
Reverse Osmosis



Principle: This process involves forcing water through a semi-permeable membrane under pressure.

Process: The membrane allows water molecules to pass through but blocks larger mineral ions.

Advantages: Produces high-quality water, removes a wide range of contaminants.

Disadvantages: Relatively high energy consumption, produces wastewater.

Distillation

Principle: This process involves boiling water and condensing the steam to produce pure water.

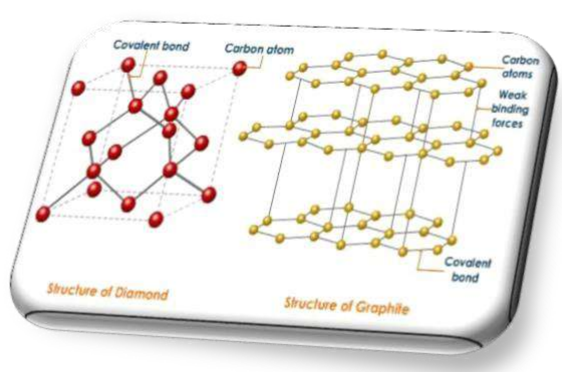
Process: Hard water is heated to boiling, and the steam is collected and condensed.

Advantages: Produces very pure water.

Disadvantages: Energy-intensive and time-consuming.

Allotropes of Carbon: Properties, Uses, and Structural Relationships

Allotropes are different forms of the same element that have distinct physical and chemical properties due to variations in their atomic arrangement. Carbon, a versatile element, exhibits several allotropes, each with unique characteristics.



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Diamond

Structure

A rigid, three-dimensional network of carbon atoms strongly bonded covalently to four other carbon atoms in a tetrahedral arrangement.

Diamond is a crystalline form of pure carbon, differing from graphite and fullerenes.

Occurrence

- ✓ Formed through high-pressure (45 kbar) and high-temperature (2,000°C) processes
- ✓ Found in volcanic pipes (kimberlite) and alluvial deposits

Properties of diamond

- ✓ Hardest known natural material.
- ✓ High refractive index.
- ✓ High electrical resistivity (Poor conductor of electricity).
- ✓ Has a tetrahedral structure
- ✓ High thermal conductivity
- ✓ It's transparent

Uses of diamond

- ✓ Jewelry and ornamentation
- ✓ Industrial cutting and drilling tools
- ✓ Industrial abrasives.

Graphite

Crystalline allotrope of carbon with layers of carbon atoms arranged in hexagonal rings, weakly held together by Van der Waals forces.

- ✓ Hexagonal lattice structure
- ✓ Carbon atoms bonded in trigonal planar arrangement
- ✓ Weak van der Waals forces between layers

Occurrence

- ✓ Found in metamorphic rocks
- ✓ Often associated with quartz, feldspar, and mica

Properties of graphite

- ✓ Soft and slippery.
- ✓ High thermal conductivity.
- ✓ Good conductor of electricity.
- ✓ Have hexagonal layers.
- ✓ It's opaque.

Uses

- ✓ Writing and art supplies (pencils, charcoal)
- ✓ Lubricants and bearings
- ✓ Electrical contacts and electrodes
- ✓ Nuclear reactors (moderator)

Other Allotropes

- ✓ Amorphous carbon
- ✓ Graphene
- ✓ Fullerene

Amorphous Carbon

Amorphous carbon is a non-crystalline form of carbon that lacks a regular, repeating structure (has a disordered structure). It is also known as non-crystalline carbon or glassy carbon. Examples are charcoal and coal.

Properties

- ✓ Amorphous carbon is typically black or dark grey in color.
- ✓ Relatively inert, but can react with certain chemicals, such as acids and bases.
- ✓ Has a high thermal conductivity, making it useful for thermal management applications.
- ✓ Semiconductor, with a resistivity that can vary depending on the specific material.
- ✓ Has a high mechanical strength, making it useful for structural applications.

Applications of Amorphous Carbon

- ✓ Used in electronic devices such as batteries, fuel cells, and solar cells.
- ✓ Used in energy storage devices such as super capacitors and batteries.
- ✓ Biomedical applications such as implants, biosensors, and drug delivery systems.
- ✓ Environmental applications such as water purification, air filtration, and soil remediation.

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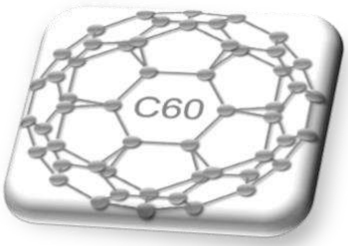
Fullerenes Buckminsterfullerene

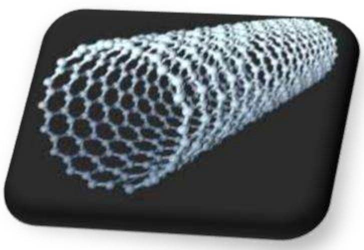
Fullerenes are a class of carbon molecules with a cage-like structure.

Structure



Spherical: The most common fullerene is C_{60}, often referred to as a "buckyball." It consists of 60 carbon atoms arranged in a spherical structure with 20 hexagons and 12 pentagons.

Other Shapes: Fullerenes can also form cylindrical structures known as carbon nanotubes.



Properties

- ✓ High Stability: Fullerenes are remarkably stable due to their strong carbon-carbon bonds.
- ✓ Superconductivity: Some fullerenes exhibit superconducting properties at low temperatures.
- ✓ Unique Electronic Properties: They have unique electronic properties that make them useful in various applications.
- ✓ High Strength and Rigidity: Fullerenes are incredibly strong and rigid.

Applications

- ✓ Materials Science: Fullerenes can be used to create new materials with enhanced properties, such as increased strength and conductivity.
- ✓ Electronics: They have potential applications in electronics, including organic solar cells and transistors.
- ✓ Medicine: Fullerenes are being explored for their potential use in drug delivery and medical imaging.
- ✓ Lubricants: Their unique structure makes them excellent lubricants.

Note: Fullerenes can be functionalized to target specific cells or tissues, allowing for more precise and efficient drug delivery.

They can be designed to release drugs in a controlled manner, reducing the risk of side effects and improving treatment outcomes. Fullerenes can be used as contrast agents in medical imaging techniques such as MRI and CT scans, allowing for better visualization of tumors and other diseases.

Fullerenes can be functionalized with fluorescent molecules, enabling optical imaging of cells and tissues.

Relationship between Structure of the allotropes of carbon and their Properties

Diamond; The strong covalent bonds in diamond contribute to its hardness and high melting point, this makes diamond to be used for cutting glass and making drilling devices.

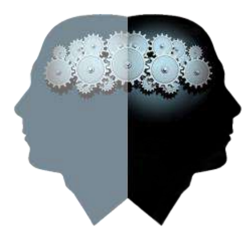
Its durability and shinny appearance makes it to be used for making jewellery and ornaments.

Graphite; The layered structure of graphite allows the layers to slide over each other, making it soft and slippery, hence used as a lubricant.

The delocalized electrons in the layers move throughout the layers allowing it to conduct electricity hence graphite is used in electrolysis in blast furnace and batteries.

In conclusion, the diverse properties and uses of carbon allotropes are directly related to their unique atomic arrangements. Understanding these structural relationships provides valuable insights into the potential applications of carbon-based materials.

Knowledge Met in this Topic



- ✓ Carbon cycle: Exchange between atmosphere, oceans, land, and living organisms.
- ✓ Carbon forms: Organic, inorganic, fossil fuels, and atmospheric CO_2.
- ✓ Environmental impact of carbondioxide Climate change, ocean acidification, and ecosystem disruption.
- ✓ Human activities that raise carbondioxide levels: Deforestation, burning fossil fuels, and land use changes. Mitigation strategies: Carbon capture, renewable energy, sustainable practices.
- ✓ Greenhouse gases (GHGs) include CO_2, CH_4, N_2O, and fluorinated gases. They are gases that trap heat from the earth resulting into global warming, this warming can cause Rising temperatures, extreme weather (Droughts, floods, heatwaves), Sea-

level rise i.e Oceans expanding, ice melting, Ecosystem disruption (Biodiversity loss, migration).

- ✓ Water hardness: Measure of calcium and magnesium ions. Temporary hardness, Carbonate-based, removable by boiling. Permanent hardness, non-carbonate minerals, requires chemical treatment.
- ✓ Allotropes are various forms in which an element exists without change of physical state. Carbon allotropes include Diamond (Crystalline, hardest substance), graphite(Soft, conductive, and lubricative), fullerenes(Buckminsterfullerene (C60), nanotubes), and amorphous carbon (Charcoal, soot, glassy carbon)

End of chapter Scenarios

Item 1:

Mitchell visited her uncle who burns charcoal to earn a living. When her uncle explained to her the steps through which charcoal is made, she was worried of the environment impacts that will arise if her uncle continues with the practice.

Task:

As a chemistry learner, make a write up highlighting the causes of Mitchell's worries. In the write up, include the following:

a. The category of fuel which charcoal is and its composition.
b. The category and composition of the resource being affected.
c. The consequences of Mitchell's uncle activities on the resource.
d. Benefits of charcoal as a fuel.

Item 2:

Carbon dioxide (CO_2) is a greenhouse gas that plays a significant role in regulating Earth's temperature. Human activities, such as burning fossil fuels, have increased the concentration of CO_2 in the atmosphere, leading to global warming and climate change.

Task;

As a young scientist, you are tasked with investigating the impact of human activities on carbon dioxide levels in the atmosphere and proposing solutions to mitigate climate change.

Item 3:

A manufacturing company uses charcoal as a reducing agent in metal production. As a process engineer, you must investigate charcoal production, determine its composition, uses, and environmental impacts.

Charcoal Production Process:

- Raw material: Wood waste

- Temperature: 500°C

- Reaction time: 2 hours

Task;

1. Explain the process of making charcoal.

2. List industrial uses of charcoal.

3. Discuss environmental impacts of charcoal burning

Item 3:

A local energy company wants to transition to sustainable energy sources. As an energy consultant, you must evaluate various fuels and categorize them as renewable or non-renewable, discuss environmental impacts, and suggest alternatives.

Fuels:

- ✓ Charcoal
- ✓ Solar Energy
- ✓ Natural Gas
- ✓ Biomass
- ✓ Coal
- ✓ Kerosene
- ✓ Hydroelectricity power
- ✓ Diesel

Task:

For each fuel:

1. Categorize as renewable or non-renewable

2. Discuss environmental impacts of fuels.

3. Suggest alternative sustainable energy sources

Item 4:

A chemical manufacturing company uses carbon dioxide in various processes. As a process engineer, you must investigate carbon dioxide's properties and applications.

Task:

1. List physical and chemical properties of carbondioxide.

2. Discuss industrial applications of carbondioxide.

3. Analyze global carbon dioxide emissions (sources, trends) and environmental impacts of carbondioxide levels in the atmosphere.

Item 5:

A manufacturing company uses water in various industrial processes, including cleaning, cooling, and chemical reactions. The company's water supply comes from a local river. However, the company has experienced scaling issues in its equipment and pipes, which has reduced efficiency and increased maintenance costs.

Task:

As a water quality specialist, analyze the water to determine its type, identify the ions responsible for its characteristics, and recommend water treatment methods.

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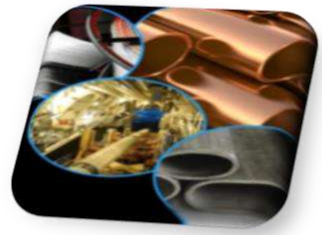
Competency Based Curriculum

TOPIC 5:

REACTIVITY SERIES



Competency: The learner evaluates data on reactivity in order to arrange metallic elements according to their reactivity.

Key words

- Reactivity series
- Displacement
- Alloys
- Blow pipe

By the end of this topic, the learner should be able to;

- Appreciate that metals vary in their chemical reactivity and can be arranged in a reactivity series (k, u,s)
- Understand that alloys are mixtures of a metal with other metals and/ or non-metals and compare the properties of common metals with their alloys (u, s)

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Metal or non-metal items surround us, and we must distinguish between the two. We must first comprehend the characteristics that distinguish metal from non-metal and their reactivity.

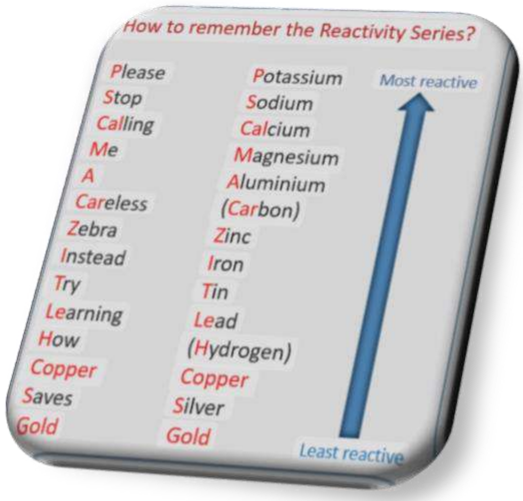
The reactivity series is a table that provides information on the reactivity of various metals. It may be used to predict whether one metal can displace another in a metal reaction and metal reactivity to water and acids. The reactivity series examples can also help you determine if a displacement reaction is single or double.

All metals and non-metals have unique characteristics, which may be used to identify whether a metal is more reactive in air, water, or acid.

The reactivity series ranks metals according to their reactivity, with the most reactive metals at the top and the most minor reactive metals at the bottom. As a result, a metal **reactivity series** is a list of **metals arranged** in order of their reactivity. Metals higher in the series are more reactive than those lower in the series and will always displace them from their compounds. Chemical reactivity of metals is related to their ability to form positive ions.

Metals' reactivity is **caused** by either their **electronic structure** or their imprecise **outer orbitals**. Metals lose electrons, resulting in positively charged ions. Metals with higher atomic numbers are more reactive because their electrons are farther away from the positively charged nucleus. As a result, they can easily be lost.

Reactivity Series of Metals



Note: This table includes hydrogen, which is a non-metal but is often included in the reactivity series for comparison purposes.

Carbon which is a non metal is also included just to predict the outcome of the reactions involving metal oxides and carbon and helps in extraction of metals from their ores.

Metals that React with cold water

General Reaction:
Metal + Water → Metal Hydroxide + Hydrogen Gas

Safety Precautions

- ✓ Wear protective gloves and eye protection
- ✓ Perform the experiment in a well-ventilated area
- ✓ Do not touch or handle potassium or sodium directly
- ✓ Use a Small piece of potassium and sodium metal

Materials:

- ✓ Trough or beaker
- ✓ Cold water
- ✓ Potassium metal
- ✓ Sodium metal

Procedure

- ✓ Pour water in a trough or beaker
- ✓ Prepare a small piece of potassium metal using a knife. And drop it on the water in a beaker using a test tube holder.
- ✓ Observe the reaction
- ✓ Use a blue and red litmus paper to test the solution in the beaker
- ✓ Repeat the above steps using sodium metal.

Observations

Vigorous reaction; Potassium reacts explosively with water. It floats, melts into a silvery ball that darts on the surface of water, producing a hissing sound and a purple flame. The reaction will produce hydrogen gas, which is flammable and Potassium hydroxide (KOH) will be formed in solution. When the solution is tested with a litmus paper, the red litmus paper turns blue while a blue litmus paper remains blue. This is due to the formation of alkaline potassium hydroxide solution.

$2K(s) + 2H_2O(l) \rightarrow 2KOH(aq) + H_2(g)$

Sodium react violently with water, producing a hissing sound, the sodium metal floats and melts into a silvery ball that darts on the surface of the water. The heat produced in the reaction may not be sufficient enough to produce a flame as potassium does, but if so it burns with a yellow flame. This indicates that potassium is more reactive as compared to sodium. The solution formed turns a red Litmus paper blue indicating an alkaline solution.

$2Na(s) + 2H_2O(l) \rightarrow 2NaOH(aq) + H_2(g)$

Conclusion

Potassium and sodium are highly reactive metals that react vigorously with cold water to produce potassium hydroxide and sodium hydroxide respectively and hydrogen gas. The reaction is exothermic, releasing heat and energy.

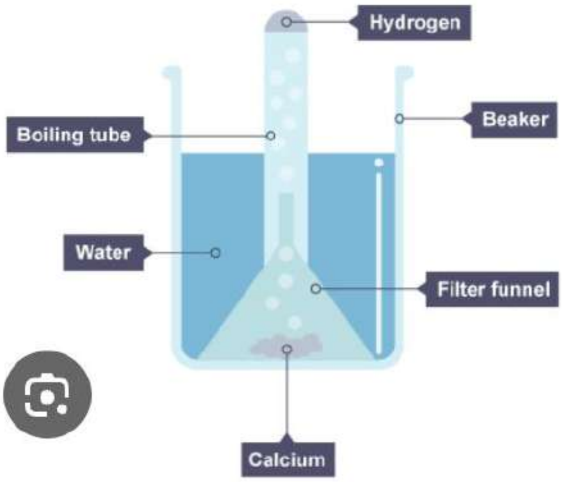
Calcium reacts readily with water.

calcium + water → calcium hydroxide + hydrogen

$Ca(s) + 2H_2O(l) \rightarrow Ca(OH)_2(aq) + H_2(g)$

Calcium hydroxide is slightly *soluble* in water so once the solution is saturated, it starts to become milky as solid calcium hydroxide appear.

The apparatus in the picture is used to react calcium with water and collect the gas. This would not be used with sodium or potassium as they float and the reaction is too vigorous.



The reaction is exothermic, meaning it releases heat.

The reaction is less vigorous than the reaction of alkali metals (like sodium or potassium) with water and the heat generated by the reaction is not sufficient to ignite the hydrogen gas.

Calcium hydroxide is a weak base and is also known as slaked lime.

Metals that react with steam

Many metals, especially those that are less reactive than alkali metals and alkaline earth metals do not react with cold water but can react with steam at high temperatures. This reaction typically produces metal oxides and hydrogen gas.

General Reaction:

Metal + Steam → Metal Oxide + Hydrogen Gas

<u>Magnesium with Steam</u>

Aim: An experiment to investigate the effect of steam on magnesium metal.

Hypothesis: Magnesium reacts with steam producing hydrogen gas.

Safety Precautions

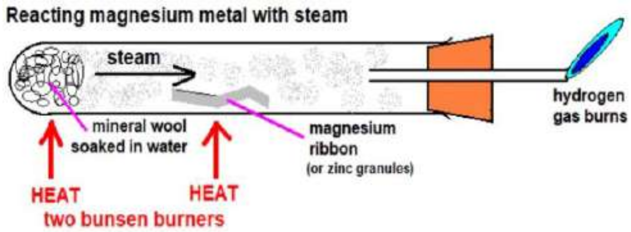
- ✓ Wear protective gloves and eye protection.
- ✓ Perform the experiment in a well-ventilated area
- ✓ Do not touch or handle magnesium directly
- ✓ Use a small amount of magnesium

Materials

- ✓ Small piece of magnesium metal
- ✓ Crucible
- ✓ Bunsen burner
- ✓ Crucible tongs
- ✓ Test tube
- ✓ Water

Procedure

- ✓ Soak the mineral wool with water and load into the boiling tube.
- ✓ Clean the magnesium ribbon with sand paper and insert it in the middle of boiling tube.
- ✓ Add the bung fitted with glass tubing. The end of the tubing should protrude at least 2 cm from the rubber to enable the evolved hydrogen from the demonstration to be lit.
- ✓ Clamp the tube horizontally, heat the magnesium ribbon Strongly and the wool gently.



Observations

The grey solid glows red hot producing a bright flame and a white smoke. The reaction will produce hydrogen gas, which is flammable and white powder of Magnesium oxide (MgO) will be formed.

Conclusion

Magnesium reacts with steam to produce magnesium oxide and hydrogen gas. The reaction is exothermic and produces a bright flame. The metal oxides produced are often solid compounds.

Note; The reactivity of metals with steam can vary depending on factors such as the temperature, the pressure of the steam, and the surface area of the metal.

Metal	Reaction with water	Equation
Magnesium (Mg)	Reacts slowly with cold water but reacts readily with steam, producing magnesium oxide (MgO) and hydrogen gas (H_2).	$Mg(s) + H_2O(g) \rightarrow MgO(s) + H_2(g)$
Aluminium (Al)	Reacts very slowly with cold water due to the formation of a protective oxide layer on its surface. However, it reacts with steam to produce aluminium oxide (Al_2O_3) and hydrogen gas (H_2).	$2Al(s) + 3H_2O(g) \rightarrow Al_2O_3(s) + 3H_2(g)$
Zinc (Zn)	Reacts very slowly with cold water but reacts with steam to produce zinc oxide (ZnO) and hydrogen gas (H_2)	$Zn(s) + H_2O(g) \rightarrow ZnO(s) + H_2(g)$
Iron (Fe)	Reacts very slowly with cold water but reacts with steam to produce iron(III) oxide (Fe_2O_3) and hydrogen gas (H_2)	$2Fe(s) + 3H_2O(g) \rightarrow Fe_2O_3(s) + 3H_2(g)$

Metals that do not react with water or steam

- ✓ Copper (Cu)
- ✓ Silver (Ag)
- ✓ Gold (Au)

Note:

The reactivity of a metal with water is determined by its position in the reactivity series. More reactive metals react more vigorously with water.

The general products of the reaction between a metal and water are a metal hydroxide and hydrogen gas. Some less reactive metals react only with steam to form a metal oxide and hydrogen gas.

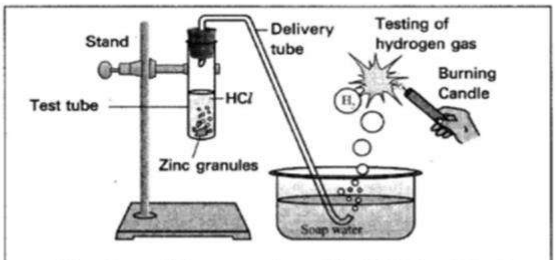
Metals below hydrogen in the reactivity series do not react with water or steam.

Reactions of metals with acids

Metals above hydrogen in the reactivity series react with dilute acids to produce hydrogen gas and a salt.

Investigating the effect of acids on Metals

Aim : To show the reaction of acids with metals.
Required Materials : 1. Test tube 2. Delivery tube 3. Glass trough 4. candle 5. Soap water 6. Dil. HCl 7. Zinc granules.



Reaction of Zinc granules with dil.HCl and testing hydrogen gas by a burning candle

Procedure :

1. Set the apparatus as shown in figure.
2. Take about 10 ml of dilute HCl in a test tube and add a few zinc granules to it.
3. We will observe the formation of gas bubbles on the surface of zinc granules.
4. Pass the gas being evolved through the soap water.
5. Gas filled bubbles are formed in the soap solution which rise into the air.

Observations

Reactive metals (e.g., zinc, iron, these metals will react with acids to produce hydrogen gas and a metal salt.

Less reactive metals (e.g., copper , these metals will not react with dilute acids under normal conditions)

General reaction

Metal + Acid → Salt + Hydrogen gas

For example:

$Zn(s) + 2HCl(aq) \rightarrow ZnCl_2(aq) + H_2(g)$

Highly Reactive Metals (K, Na, Ca, Mg)

These metals readily lose electrons to form ions, reacting vigorously with dilute HCl.

1. Equation: $2M (s) + 2HCl (aq) \rightarrow 2MCl (aq) + H_2 (g)$
2. Characteristics:
   - ✓ Vigorous effervescence (bubbling)
   - ✓ Hydrogen gas produced
   - ✓ Metal chloride formed

Moderately Reactive Metals (Al, Zn, Fe)

These metals react slowly, losing electrons to form ions:

1. Equation: $2M (s) + 2HCl (aq) \rightarrow 2MCl (aq) + H_2 (g)$
2. Characteristics:
   - Slow effervescence
   - Hydrogen gas produced
   - Metal chloride formed

Less Reactive Metals (Cu)

Copper does not react with dilute HCl due to its low reactivity:

1. Equation: No reaction
2. Characteristics:
   - No effervescence
   - No hydrogen gas produced
   - No metal chloride formed

Conclusion

The reactivity of metals with acids depends on the metal's position in the reactivity series. More reactive metals will react with acids to produce hydrogen gas and a metal salt, while moderately reactive metals will react moderately with the acid.

Metals below hydrogen (less reactive metals) in the reactivity series do not react with dilute acids.

Note

- ✓ Reactivity decreases down the reactivity series.
- ✓ More reactive metals react vigorously.
- ✓ Less reactive metals react slowly or not at all.
- ✓ Hydrogen gas produced in most reactions.

Reactions of metals with oxygen

Experiment

Aim: To investigate the effect of burning metals in air(oxygen).

Hypothesis: Metals burn in air producing oxides.

Safety Precautions

- ✓ Wear protective gloves and eye protection
- ✓ Perform the experiment in a well-ventilated area
- ✓ Do not touch or handle metals directly
- ✓ Use a small amount of metal

Materials

- ✓ Various metals (e.g., magnesium, zinc, copper, iron)
- ✓ Crucible
- ✓ Bunsen burner
- ✓ Crucible tongs
- ✓ Test tube

Procedure

- ✓ Place a small piece of metal in the crucible
- ✓ Heat the crucible using a Bunsen burner until the metal starts to glow
- ✓ Allow the metal to cool and observe the appearance

Observations

Reactive metals (e.g., magnesium, zinc, these metals will react with oxygen to form metal oxides, which are often powdery solids.

Less reactive metals (e.g., copper, iron, these metals may react with oxygen more slowly or not at all, depending on the conditions.

Example

Magnesium with oxygen

$2Mg(s) + O_2(g) \rightarrow 2MgO(s)$

Discussion

The reactivity of a metal with oxygen generally increases as you move up the reactivity series. Meaning, Metals higher in the series are more reactive than those lower down.

Many metals react with oxygen to form metal oxides. Potassium and sodium are soft metal which are easily cut exposing a shiny surface which changes to dull rapidly. The change from shiny to dull is called ***tarnishing***.

The two react immediately with oxygen to form metal oxides. For example, Potassium burns with a lilac flame when heated in air forming potassium oxide. Sodium reacts with oxygen to form sodium oxide.

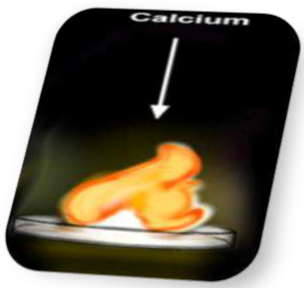
potassium + oxygen ➞ potassium oxide

$4K(s) + O_2(g) \rightarrow 2K_2O(s)$

Characteristics:

- ✓ Vigorous combustion
- ✓ Spontaneous ignition
- ✓ High heat and light

When calcium burns in oxygen, it produces calcium oxide and a brick red flame.



The chemical reaction for this process is:

$2\ Ca\ (s) + O_2\ (g) \rightarrow 2\ CaO\ (s) + \Delta$

This reaction is an example of a combustion reaction, and it's also exothermic, meaning it gives off heat.

Calcium oxide, also known as quicklime or burnt lime, is an important ingredient in cement. It reacts vigorously with water, sometimes releasing enough heat to ignite combustible materials. Inhaling, touching, or getting calcium oxide in your eyes or skin can cause severe irritation, coughing, sneezing, and burns in the respiratory canal.

Moderately reactive metals

React with oxygen after heating. For example, Magnesium reacts readily in air burning with a white light forming magnesium oxide.

Magnesium + oxygen → magnesium oxide

$2Mg(s) + O2(g) \rightarrow 2MgO(s)$

Iron reacts slowly with oxygen in moist air to form rust (iron(III) oxide).

$4Fe(s) + 3O_2(g) \rightarrow 2Fe_2O_3(s)$ (Iron(III) oxide)

Aluminium: Reacts with oxygen to form a thin, protective layer of aluminium oxide (Al_2O_3) on its surface, preventing further oxidation.

Characteristics

- ✓ Slow combustion
- ✓ Requires heat
- ✓ Forms oxide layer

The reactivity series ranks metals by how easily they react. Metals at the top of the reactivity series are the most reactive and are the easiest to oxidize.

The reaction of metals with oxygen is an oxidation reaction because the metals gain oxygen. A more reactive metal can displace a less reactive metal from a compound

Types of Oxides formed

- ✓ Basic oxides: React with acids (K_2O, Na_2O).
- ✓ Amphoteric oxides: React with acids and bases (Al_2O_3, ZnO).
- ✓ Neutral oxides: Do not react (Cu_2O).

Displacement in solutions

The reactivity series of metals is a crucial tool for understanding how metals interact with each other and with other substances. A key concept related to this series is the idea of **displacement reactions**, which occur when a more reactive metal displaces a less reactive metal from its compound. This competition for ions and oxygen is a fundamental aspect of these reactions.

Displacement Reactions are reactions where a **more reactive metal** replaces a **less reactive metal** in a compound.

Competition for Ions: When a more reactive metal is placed in a solution containing ions of a less reactive metal, the more reactive metal will tend to lose electrons and form its own ions, while the less reactive metal ions gain electrons and are reduced to the elemental form.

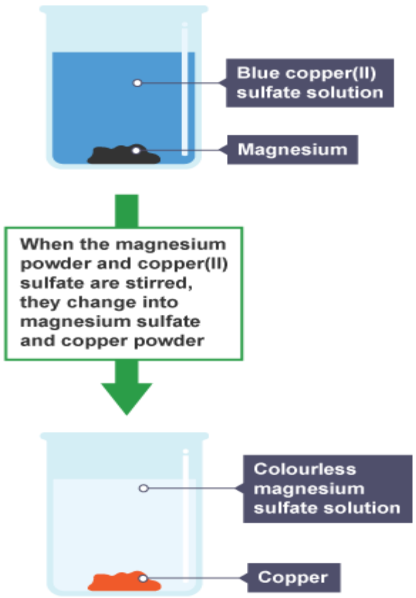
Competition for Oxygen: In reactions involving metal oxides, a more reactive metal can displace a less reactive metal from its oxide by combining with the oxygen.

Examples
Displacement of Copper from Copper(II) Oxide by Zinc

Reaction: $Zn(s) + CuO(s) \rightarrow ZnO(s) + Cu(s)$

Zinc is more reactive than copper. When zinc is heated with copper(II) oxide, zinc atoms react with oxygen from copper(II) oxide to form zinc oxide, while copper metal is displaced.

Magnesium is **more reactive** than **copper**. It **displaces** copper from copper(II) sulphate solution.



magnesium + copper(II) sulphate→ magnesium sulphate + copper

$Mg(s) + CuSO_4(aq) \rightarrow MgSO_4(aq) + Cu(s)$

In this displacement reaction:

- the copper coats the magnesium
- the solution's blue colour fades as blue copper(II) sulphate is replaced by colourless magnesium sulphate solution.

General Rules

A more reactive metal can displace a less reactive metal from its compound in solution or in the solid state when heated. The reactivity series provides a predictive framework for determining whether a displacement reaction will occur.

Applications

Extraction of Metals: The concept of displacement reactions is used in the extraction of metals from their ores. For example, more reactive metals like carbon or aluminium can be used to reduce metal oxides to their elemental forms.

Corrosion: Displacement reactions contribute to the corrosion of metals. For example, iron in contact with a more reactive metal like zinc can be protected from corrosion, as the zinc will preferentially react with oxygen and water.

In Summary

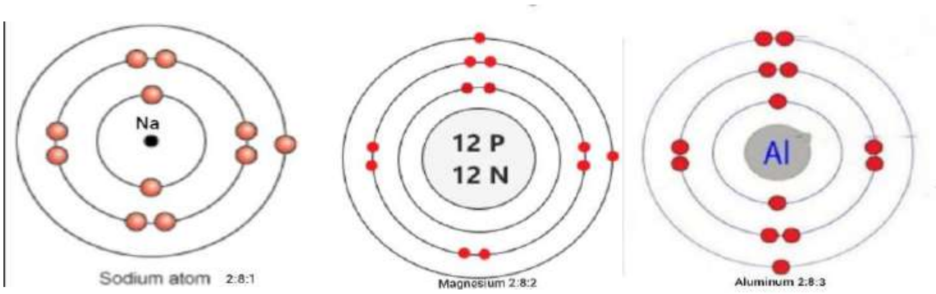
The reactivity series provides a valuable framework for understanding the competitive interactions between metals, particularly in terms of their ability to displace each other from compounds and compete for oxygen. These concepts have significant implications in various fields, including metallurgy, chemistry, and materials science.

Reactivity of Metals and Electron Structure

The reactivity of an element, as reflected in its position within the reactivity series, is directly linked to its **electronic structure**, particularly the arrangement of electrons in its outermost shell (valence shell). It depends on the number of valence electrons, the outermost electrons in an atom.

Metals, especially those high in the reactivity series (like alkali and alkaline earth metals), have a strong tendency to lose their valence electrons. This is because they have relatively few valence electrons (often 1 or 2), these valence electrons are less tightly held by the nucleus thus can be easily lost forming positive ions

Example: Sodium (Na) has one valence electron. It readily loses this electron to form a stable Na+ ion, making it highly reactive as compared to Magnesium and Aluminium.



Their reactivity makes them useful in various applications, such as:

Sodium: Used in street lamps, manufacturing of compounds like sodium hydroxide.

Potassium: Employed in fertilizers and the production of certain drugs.

Alkaline Earth Metals (Group 2): Have two valence electrons. They are also reactive but less so than alkali metals. Their uses include:

Magnesium: Utilized in lightweight alloys, fireworks, and as a reducing agent.

Calcium: Essential component of bones and teeth, used in cement and fertilizers.

Aluminium is quite a reactive metal. Aluminium forms a thin oxide layer with oxygen which prevents further reaction and makes it slightly unreactive, this makes it to be used in items such as kitchen utensils and door handles.

Metals with more valence electrons for **example transition metals**, are less reactive as they are less likely to lose their valence electrons.

Transition Metals exhibit a wider range of reactivity due to more complex electron configurations and the involvement of inner electrons in bonding. This versatility in reactivity leads to their diverse applications:

Iron: Used in steel production, construction, and various industrial applications.

Copper: Employed in electrical wiring, plumbing, and the production of alloys like brass and bronze.

Less Reactive Metals: Have more tightly held valence electrons, making them less likely to lose electrons and thus less reactive. Their resistance to corrosion makes them suitable for certain applications:

Gold: Used in jewellery, dentistry, and electronics due to its resistance to corrosion and its attractive appearance.

Silver: Employed in jewellery, photography, and as a catalyst in certain chemical reactions.

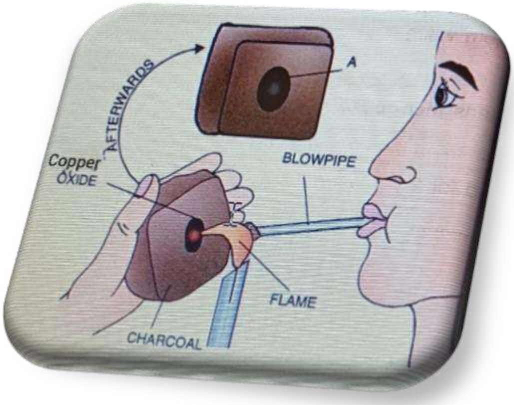
The size of an atom in addition determines how easily an atom loses electrons. The larger the size of an atom the more reactive an element is as the outer electrons are less tightly held by the nucleus and can be easily lost to form positive ions making the element reactive, this explains why potassium is more reactive than sodium yet they both have one Valence electron.

In summary, the reactivity of a metal is influenced by its tendency to lose or gain electrons to achieve a stable electron configuration. Metals with fewer valence electrons and those with large atomic size are more likely to lose electrons and are therefore more reactive.

In essence: The reactivity of an element, as determined by its electronic structure, plays a pivotal role in shaping its potential uses. Highly reactive metals find applications where their reactivity is advantageous, such as in batteries and as reducing agents. Less reactive metals are often utilized in applications where resistance to corrosion and chemical stability are crucial.

Extracting Copper from Copper Oxide Using a Charcoal Block and Blowpipe
This process demonstrates the reduction of copper oxide to copper metal using carbon as a reducing agent.

Procedure



✓Grind a small amount of copper oxide into a fine powder.
✓Make a small depression in a charcoal block.
✓Place the powdered copper oxide in the depression and then wet it with a drop of water
✓Direct the flame of a Bunsen burner using a blowpipe onto the copper oxide powder.
✓The blowpipe concentrates the flame onto a small area, providing the necessary heat for the reaction.

Observations

The black copper oxide powder gradually changes color and Small, shiny beads of copper metal begin to appear on the charcoal block.

A reddish-brown deposit of copper metal may also form around the depression.

Word Equation

Copper(II) oxide + Carbon → Copper + Carbon dioxide

Why More Reactive Metals Cannot Be Extracted This Way

More reactive metals like those in Groups 1 and 2 (alkali and alkaline earth metals) and aluminium cannot be effectively extracted using a charcoal block and blowpipe for the following reasons:

High Reactivity: These metals are highly reactive and readily react with oxygen in the air, forming stable oxides.

Stronger Reducing Agent Required: To extract these metals, a stronger reducing agent than carbon is needed. Carbon is not a strong enough reducing agent to overcome the strong affinity of these metals for oxygen.

Formation of Stable Oxides: Even if reduction occurs initially, the highly reactive metals would immediately react with oxygen in the air to reform their oxides.

Gold as an Uncombined Metal

Gold is found naturally in its elemental (uncombined) form due to its low reactivity.

Gold is a very unreactive metal, meaning it does not readily react with other elements or compounds in the environment. It is often found in alluvial deposits, which are formed by the accumulation of sediments carried by water. These deposits can contain small particles of native gold.

Due to its uncombined state, gold can be extracted from alluvial deposits through physical methods like panning or sluicing, which separate the denser gold particles from lighter sediments.

In Summary

The extraction of copper from copper oxide using a charcoal block and blowpipe demonstrates the principle of reduction. The reactivity of metals plays a crucial role in their extraction methods. While this technique is suitable for less reactive metals like copper, more reactive metals require stronger reducing agents and specialized extraction processes. Gold, being a highly unreactive metal, can be found in its native form and extracted using physical methods.

Alloys vs. Pure Metals

Pure Metals: Consist of a single element, such as pure iron, gold, or copper. They possess a uniform and consistent atomic structure.

Alloys: Are mixtures of two or more elements, typically metals, but can also include non-metals. The composition of an alloy can be varied to achieve specific properties. They are often created to improve the properties of the pure metals.

Examples include:

Brass: An alloy of copper and zinc, applied in making locks, hinges, door knobs.

Steel: An alloy of iron and carbon (with varying amounts of other elements like chromium, nickel, and manganese), used in car parts, construction materials.

Bronze: An alloy of copper and tin, applied in making statues, bells, bearings.

Types of Alloys

Ferrous alloys; Contain iron as the main component. Examples include steel, cast iron, and stainless steel.

Non-ferrous alloys; Do not contain iron as the main component. Examples include brass, bronze, and aluminium alloys.

Composition, Properties and Uses of Alloys

The chemical composition of an alloy significantly influences its properties, which in turn determine its suitability for specific applications as seen below

Alloy	Composition	Properties	Use
Steel	Iron (98%), Carbon (2%)	Strong, durable, malleable	Construction, machinery, tools
Brass	Copper (70%), Zinc (30%)	Malleable, ductile, corrosion-resistant, conductive, has a golden attractive color.	Musical instruments, Plumbing, electrical components, decorative objects
Duralumin	Aluminium (94%), Copper (4%), Manganese (0.5%), Magnesium (1.5%)	Lightweight, strong, corrosion-resistant, hard, Malleable	Aircraft, automotive parts, construction, pressure cookers
Bronze	Copper (90%), Tin (10%)	Resistant to wear, Antimicrobial Corrosion-resistant	Art, industrial bearings, architecture
Solder	Tin (60%), Lead (40%)	Low melting point, High electrical and thermal conductivity, Flexible and resistant to cracking, High Wettability (ability to flow and adhere to metal surfaces), corrosion resistant	Electronics i.e Connecting components on PCBs. Bonding metal sheets
Stainless steel	Iron (70%), Chromium (30%), Nickel	Corrosion-resistant, high strength, durable	Cutlery, cookware, medical equipment

Key differences between metals and their alloys

Feature	Pure metals	Alloys
Composition	Single element	Mixture of two or more elements
Strength and Hardness	Generally softer and more ductile (capable of being drawn into wires)	Often exhibit enhanced strength, hardness, and durability compared to their constituent pure metals. This is due to: ○ The presence of different-sized atoms in the alloy disrupts the regular crystal structure of the pure metal, hindering the movement of dislocations (defects in the crystal lattice) that cause deformation. ○ In some alloys, the atoms of one element dissolve in the crystal structure of another, creating internal stresses that increase strength.
Corrosion Resistance	Can be susceptible to corrosion (chemical degradation) due to their reactivity with the environment.	Often exhibit improved corrosion resistance. For example, stainless steel (an alloy of iron, chromium, and nickel) is highly resistant to rusting due to the formation of a protective chromium oxide layer on its surface.
Melting Point	Have specific melting points.	Typically have lower melting points than some of their constituent metals. This can be advantageous in certain manufacturing processes.

Note:

Strength and Hardness: Alloys are generally stronger and harder than pure metals due to the disruption of the crystal lattice by different types of atoms.

Corrosion Resistance: Alloys can be designed to have better corrosion resistance than pure metals by incorporating elements that form protective oxide layers.

Ductility and Malleability: Alloys can be more ductile and malleable or less so, depending on the specific elements combined.

Cost: Alloys can be more cost-effective than pure metals, especially when using less expensive elements to achieve desired properties.

Discussion items.

Item 1: A Galvanized Iron Bucket

Galvanized iron buckets are often used for storing water. Why is iron coated with zinc to make these buckets?

Response: Zinc is more reactive than iron. When a galvanized iron bucket is exposed to water, the zinc layer reacts with oxygen and water to form a protective zinc oxide layer. This sacrificial layer prevents the iron from coming into contact with oxygen and water, thus preventing rust formation.

Item 2: Copper Vessels and Silverware

Why do copper vessels and silverware tarnish over time?

Response: Copper and silver are both reactive metals, though less reactive than iron. When exposed to air and moisture, they react with oxygen and sulphur compounds to form tarnish, which is a mixture of metal oxides and sulphides. This tarnish can be removed using various cleaning methods.

Item 3: Sacrificial Anodes

Ships often have blocks of zinc attached to their hulls. What is the purpose of these zinc blocks?

Response: These zinc blocks act as sacrificial anodes. Zinc is more reactive than the iron in the ship's hull. When the ship is submerged in water, the zinc reacts with water and oxygen, corroding instead of the iron hull. This protects the ship's hull from corrosion.

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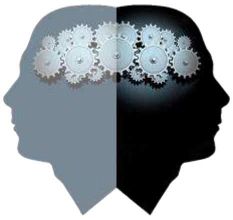
Knowledge Met in this Topic



- ✓ Reactivity series is A list of metals arranged in order of decreasing reactivity.
- ✓ Metals higher in the series are more reactive than those lower down i.e In the reactivity series above, Potassium (K) is most reactive and Gold (Au) is least.
- ✓ Metals that easily lose electrons are highly reactive because their outer energy level electrons are less tightly held by the positively charged nucleus. Metals with small atoms and those with many outer electrons hardly lose their Valence electrons making them less reactive and they form protective oxides.

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THE METAL REACTIVITY SERIES

Metals can be ordered according to their reactivities; the table below shows a selection of common metals and their reactivities with water, air, and dilute acids. A more reactive metal will displace a less reactive metal from a compound.

METAL NAME & SYMBOL	REACTION WITH COLD WATER *Produces metal hydroxide & hydrogen*	REACTION WITH STEAM *Produces metal oxide & hydrogen*	REACTION WITH AIR/OXYGEN *Produces metal oxide*	REACTION WITH DILUTE ACIDS *Produces metal salt & hydrogen*	EXTRACTION METHOD
POTASSIUM (K)	VIOLENT REACTION	VIOLENT REACTION	REACTS READILY	VIOLENT REACTION	ELECTROLYSIS OF MOLTEN METAL ORE
SODIUM (Na)	STRONG REACTION	VIOLENT REACTION	REACTS READILY	VIOLENT REACTION	ELECTROLYSIS OF MOLTEN METAL ORE
CALCIUM (Ca)	MODERATE REACTION	VIOLENT REACTION	REACTS READILY	VIOLENT REACTION	ELECTROLYSIS OF MOLTEN METAL ORE
LITHIUM (Li)	MODERATE REACTION	STRONG REACTION	REACTS READILY	VIGOROUS REACTION	ELECTROLYSIS OF MOLTEN METAL ORE
MAGNESIUM (Mg)	VERY SLOW REACTION	STRONG REACTION	SLOW REACTION	VIGOROUS REACTION	ELECTROLYSIS OF MOLTEN METAL ORE
ALUMINIUM (Al)	NO REACTION	MODERATE REACTION	SLOW REACTION	MODERATE REACTION	ELECTROLYSIS OF MOLTEN METAL ORE
(Carbon)					
ZINC (Zn)	NO REACTION	MODERATE REACTION	REACTS WHEN HEATED	MODERATE REACTION	METAL ORE SMELTED WITH CARBON
IRON (Fe)	NO REACTION	REVERSIBLE REACTION	REACTS WHEN HEATED	MODERATE REACTION	METAL ORE SMELTED WITH CARBON
NICKEL (Ni)	NO REACTION	SLOW REACTION	REACTS WHEN HEATED	SLOW REACTION	METAL ORE SMELTED WITH CARBON
TIN (Sn)	NO REACTION	NO REACTION	REACTS WHEN HEATED	SLOW REACTION	METAL ORE SMELTED WITH CARBON
LEAD (Pb)	NO REACTION	NO REACTION	REACTS WHEN HEATED	SLOW REACTION	METAL ORE SMELTED WITH CARBON
(Hydrogen)					
COPPER (Cu)	NO REACTION	NO REACTION	REACTS WHEN HEATED	NO REACTION	HEAT OR PHYSICAL EXTRACTION
MERCURY (Hg)	NO REACTION	NO REACTION	REVERSIBLE REACTION	NO REACTION	HEAT OR PHYSICAL EXTRACTION
SILVER (Ag)	NO REACTION	NO REACTION	NO REACTION	NO REACTION	HEAT OR PHYSICAL EXTRACTION
GOLD (Au)	NO REACTION	NO REACTION	NO REACTION	NO REACTION	HEAT OR PHYSICAL EXTRACTION
PLATINUM (Pt)	NO REACTION	NO REACTION	NO REACTION	NO REACTION	HEAT OR PHYSICAL EXTRACTION

- ✓ Displacement reactions involve More reactive metals displacing less reactive ones from their Compounds.

- ✓ More reactive metals react more vigorously with water, producing metal hydroxides and hydrogen gas, and react more vigorously with acids, producing metal salts and hydrogen gas.
- ✓ Metals below hydrogen in the reactivity series can not displace hydrogen from its Compounds that is why copper can not react with water under normal conditions.
- ✓ Metals below carbon can be extracted from their Oxides using charcoal.
- ✓ Metals that are reactive readily form compounds while less reactive metals are less likely to form compounds, this is why metals lower in the series (Silver and gold) are found as free elements.
- ✓ The reactivity series is relevant in electrochemical processes like batteries and electroplating.
- ✓ An alloy is a uniform mixture of two or more elements usually metal and metal or non-metal.
- ✓ Alloys are generally stronger and harder than pure metals due to the disruption of the crystal lattice by different types of atoms, they can be designed to have better corrosion resistance than pure metals by incorporating elements that form protective oxide layers.

End of chapter Scenarios

Item 1:

John and Emily are engineers designing automotive parts. John uses copper for a heat sink, while Emily uses a copper-zinc mixture(brass). Analyse their choices.

Task;

As a chemist;

1. Identify the category of the material used by John and Emily.
2. Give the suitability of the materials they used.
3. Discuss applications of the materials in automotive engineering.

Item 2:

Michael and Sarah are aerospace engineers designing aircraft components. Michael uses aluminium for a structural frame, while Sarah uses an aluminium-lithium mixture. However, Michaels' components are always found of not lasting for long.

Task:

As a young chemist;

1. Explain the problem with Michael's choice
2. Give the suitability of the materials used by the two engineers.
3. Evaluate the materials used by the engineers.

All the best from Tr.solomon
This book is still under review
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Thanks