SANGAM SKM COLLEGE - NADI
LESSON NOTES – WEEK 1
YEAR 11
CHEMISTRY

STRAND 3	**REACTIONS**
Sub strand 3.2	Types of Reactions
Content Learning Outcome	• Distinguish and describe different types of reactions based on chemical statements and balanced chemical equations.

Chemical Reactions

- Chemical reactions are processes that will cause **change in the properties** of the substances involved.
- Most reactions are chemical changes and are ***irreversible*** and some are ***reversible***.

These chemical reactions are:

1. Combustion 2. Synthesis 3. Decomposition 4. Neutralisation
5. Double Displacement 6. Precipitation 7. Oxidation-Reduction

1. Combustion

- ✓ **Combustion** – is a reaction that occurs when a substance burns in **oxygen** to form compounds called **oxides**.
- ✓ E.g.: $CH_4(g) + 2O_2(g) \longrightarrow CO_2(g) + 2H_2O(l)$
- ✓ **Metals** will burn completely in oxygen to form **metallic oxides**.
- ✓ The oxides are ***ionic*** compounds and are ***basic*** in nature.
- ✓ E.g.: Magnesium + Oxygen → Magnesium oxide ($2Mg(s) + O_2(g) \rightarrow 2MgO(s)$)
- ✓ Combustion of metals may be used to distinguish some common metals as metals burn with distinctive flame.
- ✓ **Non-metals** burn completely in oxygen to form **non-metal oxides**.
- ✓ These oxides are ***molecular*** substances and are ***acidic*** in nature.
- ✓ Most are gases at room temperature.
- ✓ Organic compounds are used as fuels for its high carbon content.
- ✓ **Complete Combustion** burns **completely** in oxygen to produce **carbon dioxide and water**. A lot of **energy is released**.
- ✓ **Incomplete combustion** will form harmful products such as **carbon monoxide, soot (unburnt carbon).** Less **heat is released**.

2. Synthesis

- ✓ **Synthesis** – is where naturally occurring elements combine chemically to form compounds.
- ✓ E.g: $Pb(s) + S(s) \longrightarrow PbS(s)$
- ✓ Two **non-metals** combine, a **covalent** substance is formed.
- ✓ $C(s) + S(s) \rightarrow CS_2(l)$ (Carbon disulphide)
- ✓ **Metals** combine with a **non-metal** to form **ionic compounds**.
- ✓ $Fe(s) + S(s) \rightarrow FeS(s)$ (Iron sulphide)
- ✓ **Formation of oxides.** All combustion of elements is synthesis reaction.

3. Decomposition

- ✓ **Decomposition** - is a process that involves breaking down of a compound.
- ✓ $CaCO_3(s) \rightarrow CaO(s) + CO_2(g)$ (Calcium carbonate → Calcium oxide + Carbon dioxide)
- ✓ Carbonates are decomposed to form **carbon dioxide** and the **oxide of the metal**.

4. Neutralisation

- ✓ **Neutralisation** – is a reaction where **acids** react with **bases** to form **salt and water**.
- ✓ **E**.g: $HCl(aq) + NaOH(aq) \longrightarrow NaCl(aq) + H_2O(l)$

- ✓ This reaction is also known as the **acid-base reaction**.
- ✓ **Carbonates** react with **dilute acids** to form **a salt, water and carbon dioxide**.
- ✓ E.g: $CaCO_3 + 2HCl \rightarrow CaCl_2 + CO_2 + H_2O$

5. Precipitation

- ✓ It is the formation of an ***insoluble*** salt from the mixture of two different ***clear*** solutions.
- ✓ The ***insoluble salt*** formed is the ***precipitate*** (ppt).
- ✓ Some precipitate may settle at the bottom of the test tube; others will form a suspension.

Reaction Equation

$CuSO_4\ (aq) + 2NaOH\ (aq) \rightarrow Cu\ (OH)_2(ppt) + Na_2SO_4(aq)$

6. Double Displacement

- ✓ **Two** different salt solutions react forming a **clear solution**.
- ✓ The resultant salts formed are both soluble in water.
- ✓ It is termed double displacement as the anions are exchanged between the two cations.
- ✓ **E.g.** Barium chloride + Sodium nitrate → Barium nitrate + Sodium sulphate
  $BaCl_2(aq) + NaNO_3(aq) \rightarrow Ba\ (NO_3)_2(aq) + Na_2SO_4(aq)$

7. Oxidation-Reduction

- ✓ **Oxidation** is the ***gain*** of **oxygen**.
- ✓ **Reduction** is the ***loss*** of **oxygen**.
- ✓ Oxidation and reduction reactions occur simultaneously. As a substance is reduced, the other reactant will be oxidised.
- ✓ The collective term for oxidation-reduction reaction is **Redox reaction.**
- ✓ **E.g.1** in the extraction of metal from metal oxides using carbon, the metal oxide is reduced to the metal and carbon is oxidised to carbon dioxide. $C(s) + 2CuO(s) \rightarrow 2Cu(s) + CO_2(g)$
- ✓ **E.g.2** Iron metal is produced in the Blast Furnace by the reduction of iron (III) oxide by carbon monoxide. Carbon monoxide is oxidised to carbon dioxide.
  $Fe_2O_3(s) + 3CO(g) \rightarrow 2Fe(s) + 3CO_2(g)$

Note: **Oxidation is the loss of hydrogen or electrons. Reduction is the gain of hydrogen or electrons.**

Exercise

1. Copper nitrate when heated, forms copper oxide, nitrogen dioxide and oxygen.

$$\boxed{2Cu(NO_3)_2(s) \rightarrow 2CuO(s) + 4NO_2(g) + O_2(g)}$$

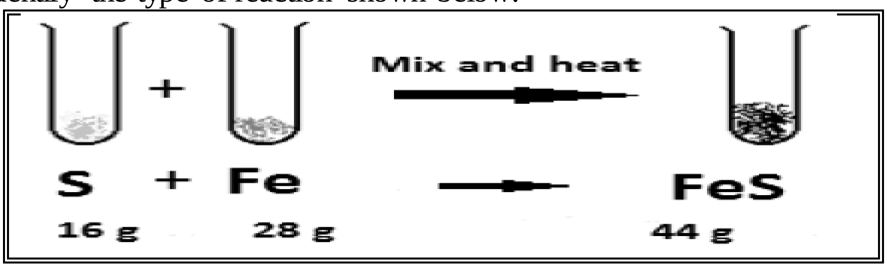
This reaction is an example of ___________________

2. For each reaction below:

i. Write a balanced equation.
ii. Classify the type of reaction.

a) Burning of sulphur
b) Burning of magnesium
c) Formation of ammonia from nitrogen gas and hydrogen gas
d) The reaction of lead and sulphur to form lead sulphide

3. Identify the type of reaction shown below.



SANGAM SKM COLLEGE - NADI
LESSON NOTES – WEEK 2
YEAR 11
CHEMISTRY

STRAND 3	**REACTIONS**
Sub strand 3.2	Types of Reactions
Content Learning Outcome	• Describe the reactions of oxidation and reduction in terms of transfer of atoms and electrons. • Study and write simple oxidation and reduction reactions involving atoms and electrons.

Oxidation-Reduction

- ✓ **Oxidation** is the ***gain*** of **oxygen**. **Reduction** is the ***loss*** of **oxygen**.
- ✓ Oxidation and reduction reactions occur simultaneously. As a substance is reduced, the other reactant will be oxidised.
- ✓ The collective term for oxidation-reduction reaction is **Redox reaction.**
- ✓ **Oxidation** is the **loss** of **hydrogen or electrons**. **Reduction** is the **gain** of **hydrogen or electrons.**

❖ **Example 1: Oxidation reaction**

- Ionisation of metal atoms to positive metal ions **(cations)** is an **oxidation** reaction as electrons are *lost* from its valence shell.
  $Na \rightarrow Na^{+} + e^{-}$
  $Mg \rightarrow Mg^{2+} + 2e^{-}$
  $Al \rightarrow Al^{3+} + 3e^{-}$
  Note: Electron(s) lost will appear on the right side of the equation.

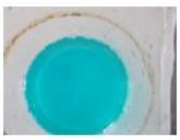
❖ **Example 2: Reduction reaction**

- Non-metals *gaining* electron(s) to form stable negative ions **(anions)** are **reduction** reactions.
  $Cl_2 + 2e^{-} \rightarrow 2Cl^{-}$
  $O_2 + 4e^{-} \rightarrow 2O^{2-}$
  Note: Electron(s) gained will appear on the left side of the equation.

❖ **Example 3**

- Displacement reactions of metals are redox reactions. The more **active metal** is **oxidised** to its ions and the **less active metals** is **reduced** from its ionic form (in aqueous solution) to the metal.

  Given below shows iron filings placed in a beaker of copper sulphate solution reduces the copper ions (light blue solution) to copper metal (reddish brown).


Copper sulphate solution


Iron filings


Mixture of Iron and copper sulphate solution


Reddish brown deposits formed

- Oxidation Reaction: $Fe \rightarrow Fe^{2+} + 2e-$
- Reduction Reaction: $Cu^{2+} + 2e- \rightarrow Cu$

❖ **Example 4**

- Displacement reactions of **active metals** with **dilute acid** are **redox reactions**. The **more active metal** is **oxidised** to its **ions** and the hydrogen ions in acid is **reduced** to **hydrogen** gas (bubbles evolved).

- Oxidation Reaction: $Mg \rightarrow Mg^{2+} + 2e-$
- Reduction Reaction: $2H^{+} + 2e \rightarrow H_2$

Exercise

1. A grey iron nail was accidently dropped into a light blue copper sulphate solution. It was observed that the nail rapidly became coated with brown deposits. The following equation describes the reaction that must have taken place.

$CuSO_4 (aq) + Fe(s) \longrightarrow Cu(s) + FeSO_4 (aq)$

i) Identify the brown deposits formed on the nail.

ii) Write the oxidation and reduction half-equations for the above reaction.

Oxidation half-equation

Reduction half-equation

iii) Some redox reactions can also be called displacement reactions and the above reaction is one such example. Explain the reason for this conclusion.

2. Identify the following equations as either oxidation or reduction.

i. $2Cl- \rightarrow Cl_2 + 2e$

ii. $I_2 + 2e- \rightarrow 2I$

iii. $Mg \rightarrow Mg^{2+} + 2e$

iv. $Zn^{2+} + 2e- \rightarrow Zn$

3. The reaction of lead oxide with carbon forms lead metal and carbon dioxide.

i. Write a balanced chemical equation to represent the reaction above.

ii. From the equation, determine which reactant is oxidised and which is reduced.

iii. Explain why the reaction between lead oxide and carbon is called a redox reaction.

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LESSON NOTES – WEEK 3
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STRAND 3	**REACTIONS**
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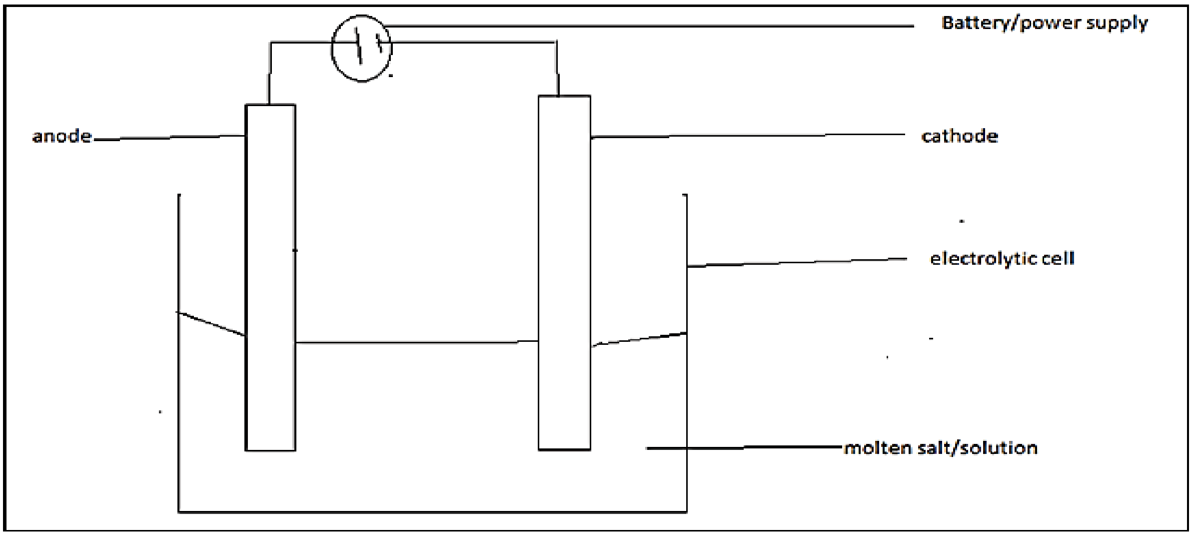
Content Learning Outcome	• Show that electrolysis of molten and aqueous salt experimental set-up involves oxidation and reduction.

Electrolysis

- Redox is commercially used in a process called electrolysis.
- Electrolysis is the decomposition of an electrolyte by passing an electric current through it.
- An **electrolyte** is a molten salt or solution that conducts electricity.
- Electrolysis is carried out in an electrolytic cell, as shown below.



❖ The components of an electrolytic cell are:

1. **Electrolyte** – molten or solutions of ionic compounds. The mobile/free ions are the carriers of electric current.
   Examples include: $NaCl_{(l)}$, $NaCl_{(aq)}$, $H_2O_{(l)}$, $MgCl_{2(aq)}$, $CuSO_{4(aq)}$.
2. **Batteries/Direct Current, DC power supply** – source of current, creates or Discharge ions in the electrolyte. The electrode potential should be large enough to drive the reactions.
3. **Electrodes** – connects batteries/DC power supply to electrolyte. The two types are **anode** (positively charged) and **cathode** (negatively charged). Electrodes are usually inert or unreactive and a conductor of electricity. A common electrode is carbon (graphite) as it is inert and a conductor. Less reactive metals such as copper, iron and zinc, are used in electroplating.

- The electrolytic cell is a complete circuit. Electrons move from the anode to the cathode.
- As electrons moves away from the anode, the anode becomes positive and as electrons are deposited on the cathode, the cathode becomes negative.

Electrolysis of a salt solution

- Electrolysis of an aqueous solution depends on the concentration, ions present and nature of the electrodes.
- Water could be oxidised or reduced. Its presence complicates the electrolysis of aqueous solutions.
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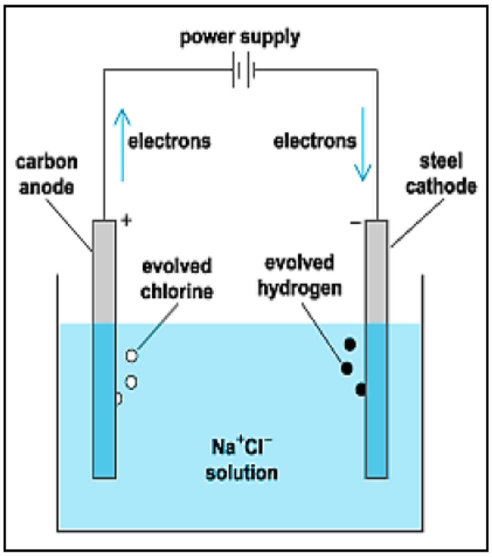
  - ❖ **Anode Reaction (Oxidation)**
    - ✓ The anions that are easily oxidised instead of water include chloride ion (Cl^-), bromide ion (Br^-) and iodide ion (I^-).
    - ✓ **Polyatomic ions** are not discharged. An example of such ions include sulphate ion (SO_4^{2-}). Instead, oxygen gas is evolved due to oxidation of water.
  - ❖ **Cathode Reaction (Reduction)**
    - ✓ A solution that contains cation of an element below aluminium in the Activity Series will be reduced instead of water.
    - ✓ Cations of salt solutions of very active elements will not be reduced as water is reduced.
  - ❖ Given below shows the electrolysis of brine, concentrated $NaCl_{(aq)}$; it is an important industrial application as it produces much needed chlorine. The **salt solution** contains the **electrolyte**, **sodium ions, chloride ions and water.**
  - ❖ **Water is reduced** at the ***cathode*** instead of sodium ions as it was **easier** to reduce; at the ***anode***, **chloride ion** was easier to **oxidise** than water.

Anode Reaction

$2Cl^-_{(aq)} \rightarrow Cl_{2\,(g)} + 2e^-$

Greenish yellow gas



Cathode Reaction

$2H_2O_{(l)} + 2e^- \rightarrow H_{2(g)} + 2OH^-_{(aq)}$

Note: *Hydroxide ion formed (OH^-) is picked by sodium ion in the solution producing sodium hydroxide, NaOH)*

Exercise

1. In the electrolysis of sodium chloride solution, the cathode reaction can be best described as
   A. oxidation of water.
   B. reduction of water.
   C. reduction of sodium.
   D. oxidation of chlorine.

2. State four main components of an electrolysis set-up.

3. Why is graphite used as an electrode?