PUMPEN ACADEMY, AKURE

SS1 CHEMISTRY SECOND TERM E-CLASSNOTE

WAEC/NECO STANDARD LEARNING MATERIAL

Prepared for Online Classroom Learning and Ministry-Standard Documentation

Motto: Ever to be Best

Standard Second Term Scheme of Work

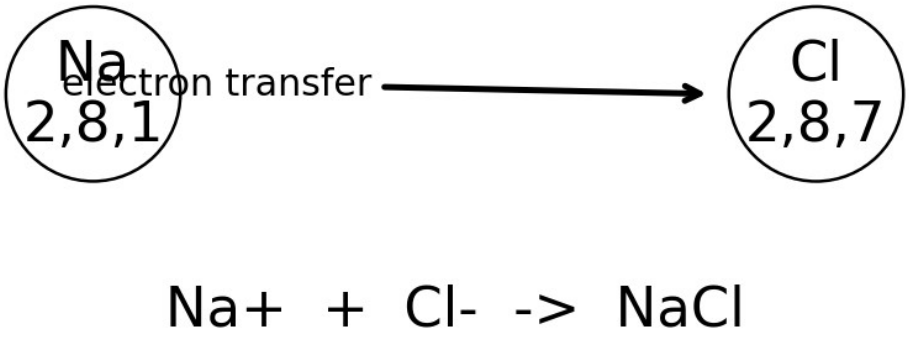
Week	Topic
1	Revision of First Term Work and Introduction to Chemical Bonding
2	Chemical Bonding I - Electrovalent/Ionic Bonding
3	Chemical Bonding II - Covalent and Coordinate Covalent Bonding
4	Chemical Bonding III - Metallic Bonding, Intermolecular Forces and Properties of Bonded Substances
5	Acids, Bases and Salts I - Meaning, Properties and Indicators
6	Acids, Bases and Salts II - Preparation, Types and Uses of Salts
7	Solubility of Substances and Solubility Curves
8	Water I - Sources, Properties, Uses and Treatment of Water
9	Water II - Hardness of Water: Causes, Types and Removal
10	Air and Its Constituents
11	Oxygen - Laboratory Preparation, Properties, Uses and Tests
12	Revision and Examination Focus

This e-classnote is arranged to support classroom teaching, online learning, continuous assessment, practical work and examination preparation. It follows the common Nigerian senior secondary school Chemistry expectations for SS1 learners and provides clear explanations, examples, activities, evaluation questions and assignments.

WEEK 1: REVISION OF FIRST TERM WORK AND INTRODUCTION TO CHEMICAL BONDING

Learning Objectives

1. Recall important first term concepts in Chemistry.
2. Explain atoms, ions, valence electrons and electronic configuration.
3. State why atoms combine with one another.
4. Define chemical bonding.
5. Explain octet and duplet rules.
6. Identify the major types of chemical bonds.



Preview of electron transfer. Valence electrons determine how atoms combine.

Detailed Lesson Content

Revision of First Term Concepts

First term Chemistry introduced students to matter, laboratory safety, elements, compounds, mixtures, separation techniques, atoms, ions, electronic configuration, periodic table, symbols, formulae and valency. These foundations are necessary because second term Chemistry explains how atoms combine and how substances behave.

Atom, Ion and Valence Electrons

An atom is the smallest particle of an element that can take part in a chemical reaction. An ion is a charged particle formed when an atom or group of atoms loses or gains electrons. Valence electrons are the electrons in the outermost shell of an atom; they are mainly responsible for chemical reactions and bonding.

Why Atoms Combine

Most atoms are unstable when they exist alone because their outermost shells are not completely filled. Atoms combine by losing, gaining or sharing electrons in order to attain stable electronic arrangements similar to noble gases. Noble gases such as helium, neon and argon are stable because their outer shells are complete.

Octet and Duplet Rules

The octet rule states that atoms tend to combine so that they achieve eight electrons in their outermost shell. The duplet rule states that some atoms, especially hydrogen, become stable when they have two electrons in their only shell.

Chemical Bonding

Chemical bonding is the force of attraction that holds atoms or ions together in a chemical substance. The major types are ionic/electrovalent bonding, covalent bonding, coordinate covalent bonding and metallic bonding.

Worked Examples / Important Points

- Sodium has electronic configuration 2,8,1 and can lose one electron to become stable.
- Chlorine has electronic configuration 2,8,7 and can gain one electron to become stable.
- Hydrogen shares one electron with another hydrogen atom to form H2.

Class Activities

7. Write the electronic configurations of sodium, magnesium, aluminium, oxygen, chlorine and calcium.
8. Identify the number of valence electrons in each element.

Evaluation Questions

9. Define chemical bonding.
10. What are valence electrons?
11. Why do atoms combine?
12. State the octet rule.
13. Mention four types of chemical bonds.

Assignment

Write the electronic configuration of the first twenty elements and identify their valence electrons.

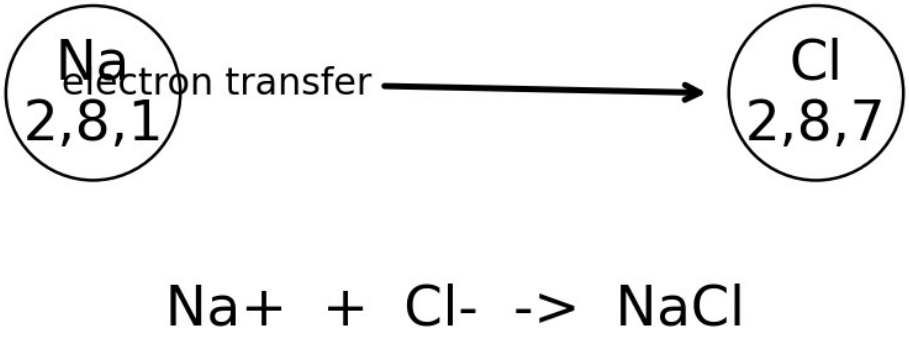
Summary

Atoms combine to attain stable electron arrangements. Chemical bonding is the force that holds atoms or ions together. Stable arrangements are explained using octet and duplet rules.

WEEK 2: CHEMICAL BONDING I - ELECTROVALENT OR IONIC BONDING

Learning Objectives

14. Define ionic/electrovalent bonding.
15. Explain formation of cations and anions.
16. Describe formation of NaCl, MgO, CaCl2 and Al2O3.
17. State properties of ionic compounds.
18. Explain electrical conductivity of ionic compounds.



Ionic bonding involves electron transfer from a metal atom to a non-metal atom.

Detailed Lesson Content

Meaning of Ionic Bond

An ionic or electrovalent bond is a chemical bond formed by the transfer of electron(s) from one atom to another, producing oppositely charged ions which attract each other. Ionic bonding usually occurs between metals and non-metals.

Formation of Ions

Metals lose electrons to form positive ions called cations, while non-metals gain electrons to form negative ions called anions. Sodium forms Na+, magnesium forms Mg2+, calcium forms Ca2+, chlorine forms Cl-, and oxygen forms O2-.

Formation of Sodium Chloride

Sodium has electronic configuration 2,8,1 while chlorine has 2,8,7. Sodium transfers its outer electron to chlorine. Sodium becomes Na+ and chlorine becomes Cl-. The attraction between Na+ and Cl- forms sodium chloride, NaCl.

Formation of Magnesium Oxide

Magnesium has configuration 2,8,2 and loses two electrons to form Mg2+. Oxygen has configuration 2,6 and gains two electrons to form O2-. The ions attract to form MgO.

Formation of Calcium Chloride

Calcium has configuration 2,8,8,2 and loses two electrons. Each chlorine atom gains one electron, so two chlorine atoms are required. The formula is CaCl2.

Formation of Aluminium Oxide

Aluminium forms Al3+ while oxygen forms O2-. To balance charges, two Al3+ ions combine with three O2- ions to form Al2O3.

Properties of Ionic Compounds

Ionic compounds are usually crystalline solids. They have high melting and boiling points because strong electrostatic forces hold the ions. Many dissolve in water. They conduct electricity when molten or in aqueous solution because ions are free to move, but they do not conduct in solid state because ions are fixed in position.

Worked Examples / Important Points

- Na -> Na+ + e-
- Cl + e- -> Cl-
- Mg -> Mg2+ + 2e-
- O + 2e- -> O2-

Class Activities

19. Draw dot-and-cross diagrams for NaCl, MgO and CaCl2.
20. Classify NaCl, MgO, CaO and KBr as ionic compounds and explain why.

Evaluation Questions

21. Define ionic bond.
22. What is a cation?
23. What is an anion?
24. Explain the formation of sodium chloride.
25. Why do ionic compounds conduct electricity when molten but not when solid?

Assignment

Draw dot-and-cross diagrams to show the formation of sodium chloride, magnesium chloride, calcium oxide and aluminium oxide.

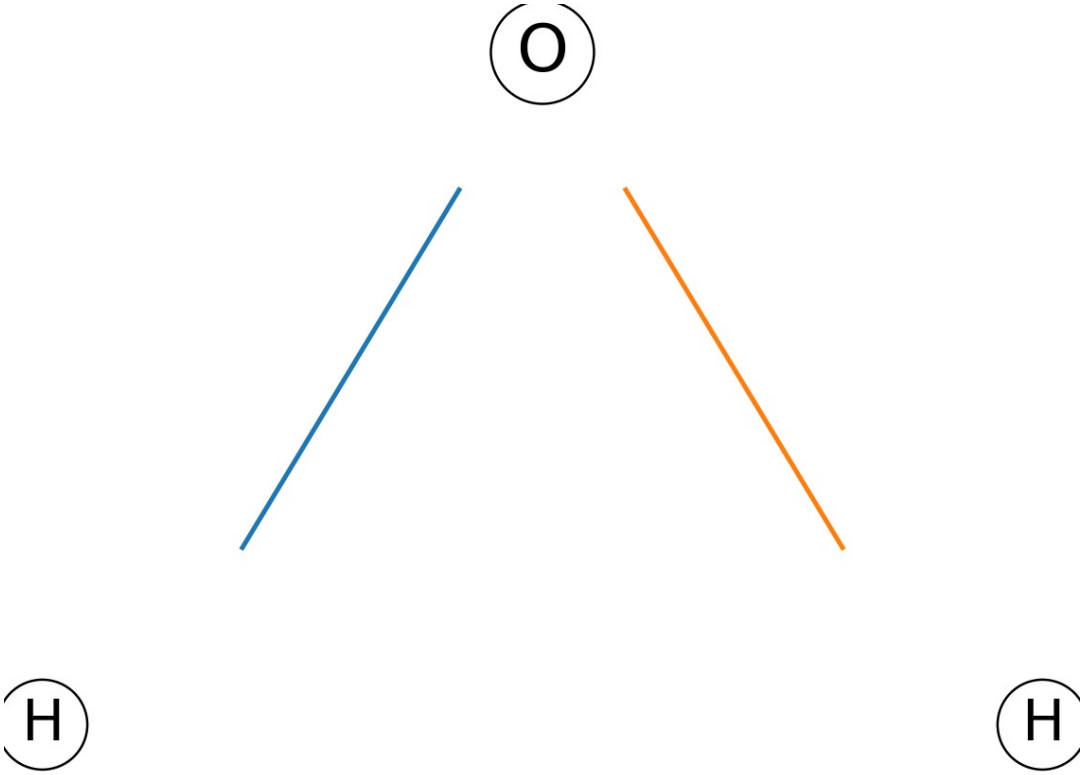
Summary

Ionic bonding involves transfer of electrons. Metals form positive ions and non-metals form negative ions. Oppositely charged ions attract to form ionic compounds with high melting points and electrical conductivity when molten or dissolved.

WEEK 3: CHEMICAL BONDING II - COVALENT AND COORDINATE COVALENT BONDING

Learning Objectives

26. Define covalent bonding.
27. Explain formation of simple covalent molecules.
28. Draw dot-and-cross diagrams for H_2, Cl_2, O_2, H_2O, NH_3 and CH_4.
29. State properties of covalent compounds.
30. Define coordinate covalent bond.
31. Give examples of coordinate bonding.



Water molecule: two shared pairs of electrons

Covalent bonding involves sharing of electron pairs between non-metal atoms.

Detailed Lesson Content

Meaning of Covalent Bond

A covalent bond is a chemical bond formed by sharing electron pair(s) between atoms. It usually occurs between non-metal atoms. Atoms share electrons in order to achieve stable outer shells.

Single, Double and Triple Bonds

A single covalent bond contains one shared pair of electrons, as in H2 and Cl2. A double bond contains two shared pairs, as in O2 and CO2. A triple bond contains three shared pairs, as in N2.

Formation of Hydrogen Molecule

Each hydrogen atom has one electron and needs one more to attain a duplet. Two hydrogen atoms share one pair of electrons to form H-H.

Formation of Water Molecule

Oxygen has six valence electrons and needs two more. It shares one pair of electrons with each of two hydrogen atoms to form H2O.

Formation of Ammonia and Methane

Nitrogen shares electrons with three hydrogen atoms to form NH3. Carbon shares electrons with four hydrogen atoms to form CH4.

Properties of Covalent Compounds

Simple covalent compounds usually have low melting and boiling points, do not conduct electricity, and may be gases, liquids or soft solids. Many are insoluble in water but soluble in organic solvents. Exceptions exist: graphite conducts electricity and sugar dissolves in water.

Coordinate Covalent Bond

A coordinate or dative covalent bond is a covalent bond in which the shared pair of electrons is donated by one atom only. Ammonia donates a lone pair to H+ to form NH4+. Water donates a lone pair to H+ to form H3O+.

Worked Examples / Important Points

- H-H contains one shared pair.
- O=O contains two shared pairs.
- N≡N contains three shared pairs.
- NH3 + H+ -> NH4+.

Class Activities

32. Draw dot-and-cross diagrams for H2, Cl2, O2, H2O, NH3 and CH4.
33. Compare ionic and covalent bonding using NaCl and H2O.

Evaluation Questions

34. Define covalent bond.
35. Explain the formation of water.
36. State five properties of covalent compounds.
37. What is a coordinate covalent bond?
38. Give two examples of coordinate covalent bonding.

Assignment

Draw dot-and-cross diagrams for chlorine, oxygen, water, ammonia and methane.

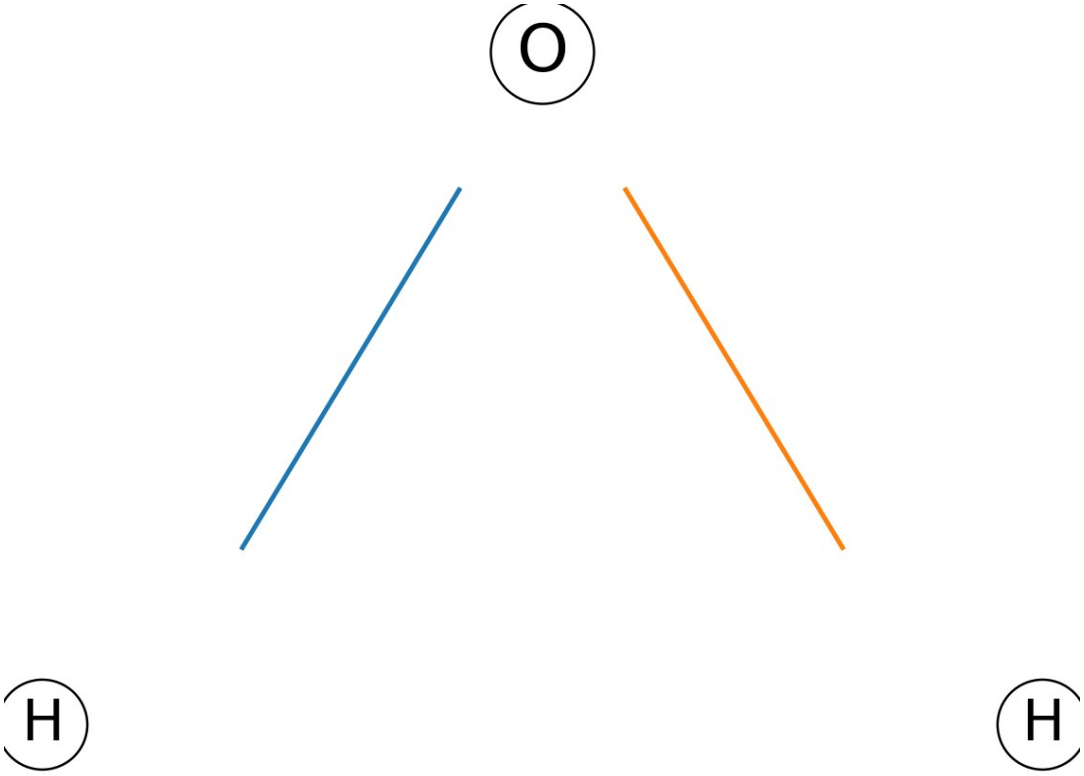
Summary

Covalent bonding involves sharing of electrons between non-metals. Coordinate bonding occurs when one atom donates both electrons in the shared pair. Covalent substances usually do not conduct electricity and often have low melting points.

WEEK 4: CHEMICAL BONDING III - METALLIC BONDING AND INTERMOLECULAR FORCES

Learning Objectives

39. Define metallic bonding.
40. Explain the structure of metals.
41. Relate metallic bonding to properties of metals.
42. Differentiate intermolecular and intramolecular forces.
43. Compare ionic, covalent and metallic substances.



Water molecule: two shared pairs of electrons

Intramolecular bonds hold atoms within substances, while intermolecular forces act between molecules.

Detailed Lesson Content

Metallic Bonding

Metallic bonding is the force of attraction between positive metal ions and a sea of delocalized electrons. In metals, valence electrons are free to move throughout the structure. The positive metal ions are held together by attraction to these mobile electrons.

Properties of Metals

Metals conduct electricity because delocalized electrons move and carry charge. Metals conduct heat because electrons transfer energy. Metals are malleable and ductile because layers of metal ions can slide over one another without breaking the metallic bond. Metals are shiny because free electrons reflect light. Many metals have high melting points because metallic bonding is strong.

Intermolecular Forces

Intermolecular forces are weak forces of attraction between molecules. They are different from intramolecular forces such as ionic, covalent and metallic bonds which hold atoms or ions together within a substance.

Hydrogen Bonding and Van der Waals Forces

Hydrogen bonding occurs when hydrogen is bonded to highly electronegative atoms such as oxygen, nitrogen or fluorine. It explains the relatively high boiling point of water. Van der Waals forces are weaker attractions found between molecules, especially simple covalent molecules.

Giant Covalent Structures

Diamond and graphite are forms of carbon. Diamond is very hard and does not conduct electricity because each carbon is bonded to four others. Graphite is soft and conducts electricity because each carbon is bonded to three others and has delocalized electrons.

Worked Examples / Important Points

- Copper conducts electricity due to delocalized electrons.
- Diamond is hard because of a giant covalent network.
- Graphite conducts electricity due to delocalized electrons.

Class Activities

44. Compare sodium chloride, water and copper under type of bond, particles present, conductivity and melting point.

Evaluation Questions

45. Define metallic bonding.
46. What are delocalized electrons?
47. Why do metals conduct electricity?
48. Differentiate intermolecular and intramolecular forces.
49. Why does graphite conduct electricity?

Assignment

Write a detailed note on how metallic bonding explains the physical properties of metals.

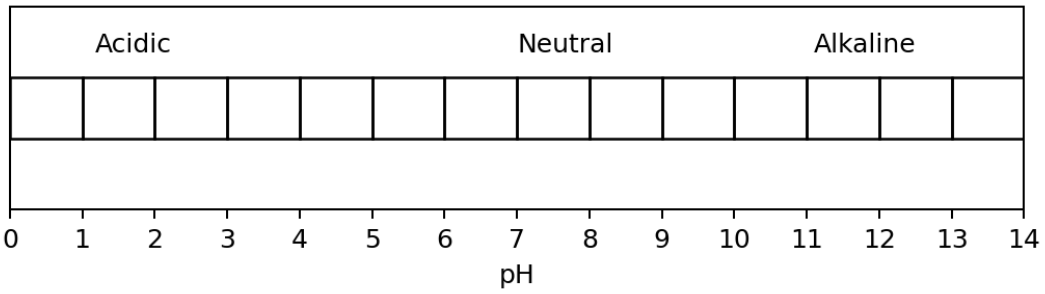
Summary

Metallic bonding explains conductivity, malleability, ductility and lustre of metals. Intermolecular forces act between molecules, while intramolecular forces are bonds within substances. Diamond and graphite are giant covalent structures with different properties.

WEEK 5: ACIDS, BASES AND SALTS I - MEANING, PROPERTIES AND INDICATORS

Learning Objectives

50. Define acids, bases and alkalis.
51. State properties of acids and bases.
52. Identify common acids and bases.
53. Explain indicators and colour changes.
54. Explain neutralization and pH.



The pH scale shows acidic, neutral and alkaline regions.

Detailed Lesson Content

Acids

An acid is a substance that produces hydrogen ions, H+, when dissolved in water. Examples include hydrochloric acid, sulphuric acid, nitric acid, ethanoic acid and carbonic acid.

Properties of Acids

Acids have sour taste, turn blue litmus red, react with reactive metals to produce hydrogen, react with carbonates to produce carbon dioxide, neutralize bases to form salt and water, have pH less than 7 and conduct electricity in solution because they contain ions.

Bases and Alkalis

A base is a substance that neutralizes an acid to form salt and water. An alkali is a base that dissolves in water to produce hydroxide ions, OH-. All alkalis are bases, but not all bases are alkalis. Sodium hydroxide and potassium hydroxide are alkalis; copper(II) oxide is a base but not an alkali.

Properties of Bases and Alkalis

Bases are bitter, feel soapy or slippery, turn red litmus blue and neutralize acids. Alkalis have pH greater than 7 and conduct electricity in solution.

Indicators and pH

An indicator changes colour depending on whether a solution is acidic or alkaline. Litmus, methyl orange, phenolphthalein and universal indicator are common indicators. pH less than 7 is acidic, pH 7 is neutral and pH greater than 7 is alkaline.

Neutralization

Neutralization is the reaction between an acid and a base to produce salt and water. For example, HCl + NaOH -> NaCl + H2O. Neutralization is used in treating acidic soil with lime, treating indigestion with antacids and controlling acidic industrial waste.

Worked Examples / Important Points

- Zn + 2HCl -> ZnCl2 + H2
- HCl + NaOH -> NaCl + H2O
- CaCO3 + 2HCl -> CaCl2 + H2O + CO2

Class Activities

55. Test lemon juice, vinegar, soap solution and salt solution with litmus paper.
56. Record the colour changes and classify each as acidic, neutral or alkaline.

Evaluation Questions

57. Define acid.
58. Define base and alkali.
59. State five properties of acids.
60. What is an indicator?
61. State the colour of methyl orange in acid and alkali.
62. Define neutralization.

Assignment

Write five examples of acids and five examples of bases. State one use of each.

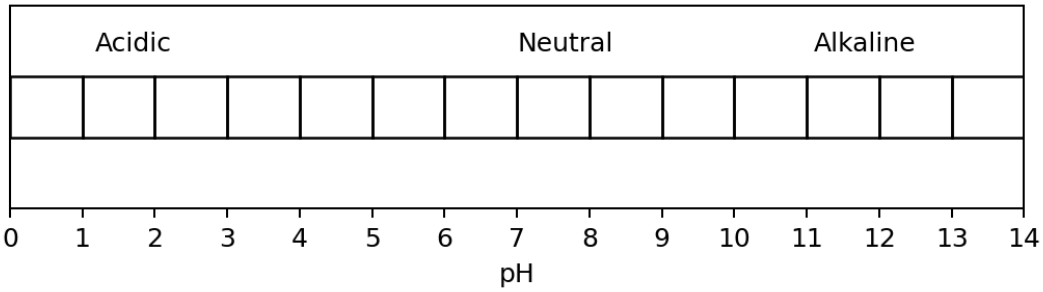
Summary

Acids produce H+ in water, while alkalis produce OH-. Indicators are used to identify acids and alkalis. Neutralization produces salt and water.

WEEK 6: ACIDS, BASES AND SALTS II - PREPARATION, TYPES AND USES OF SALTS

Learning Objectives

63. Define salt.
64. Explain how salts are formed.
65. Name common salts.
66. Describe methods of preparing salts.
67. Differentiate soluble and insoluble salts.
68. State uses of common salts.



Salts can be prepared by neutralization, reaction with metals or precipitation.

Detailed Lesson Content

Meaning of Salt

A salt is a compound formed when the replaceable hydrogen ion of an acid is replaced wholly or partly by a metal ion or ammonium ion. Salts are commonly formed when acids react with bases, metals or carbonates.

Naming Salts

Hydrochloric acid forms chlorides, sulphuric acid forms sulphates, nitric acid forms nitrates, carbonic acid forms carbonates and ethanoic acid forms ethanoates.

Types of Salts

Normal salts are formed when all replaceable hydrogen ions are replaced, for example NaCl and K2SO4. Acid salts still contain replaceable hydrogen, such as NaHCO3 and NaHSO4. Basic salts contain hydroxide or oxide groups due to incomplete neutralization.

Solubility Rules

All sodium, potassium and ammonium salts are soluble. All nitrates are soluble. Most chlorides are soluble except silver chloride and lead(II) chloride. Most sulphates are soluble except barium sulphate, lead(II) sulphate and calcium sulphate which is slightly soluble. Most carbonates are insoluble except sodium, potassium and ammonium carbonates.

Preparation of Soluble Salts

Soluble salts can be prepared by reacting an acid with a metal, an insoluble base, a carbonate or an alkali. For example, copper(II) sulphate is prepared by reacting copper(II) oxide with warm dilute sulphuric acid, filtering excess oxide, concentrating the filtrate and cooling to crystallize.

Preparation of Insoluble Salts

Insoluble salts are prepared by precipitation. Two soluble salt solutions are mixed and the insoluble product forms as precipitate. For example, BaCl2 + Na2SO4 -> BaSO4 + 2NaCl.

Uses of Salts

Sodium chloride is used as table salt and for food preservation. Calcium carbonate is used in cement, chalk and toothpaste. Copper(II) sulphate is used as a fungicide and laboratory reagent. Sodium hydrogen carbonate is used in baking powder, antacids and fire extinguishers.

Worked Examples / Important Points

- $CuO + H_2SO_4 \rightarrow CuSO_4 + H_2O$
- $Zn + H_2SO_4 \rightarrow ZnSO_4 + H_2$
- $BaCl_2 + Na_2SO_4 \rightarrow BaSO_4 + 2NaCl$

Class Activities

69. Name the salts formed from hydrochloric, sulphuric, nitric, carbonic and ethanoic acids.

Evaluation Questions

70. Define salt.
71. Name the salt formed from sulphuric acid.
72. State three methods of preparing salts.
73. Describe the preparation of copper(II) sulphate crystals.
74. What is precipitation?

Assignment

Write balanced equations for preparing zinc sulphate, sodium chloride, copper(II) sulphate and barium sulphate.

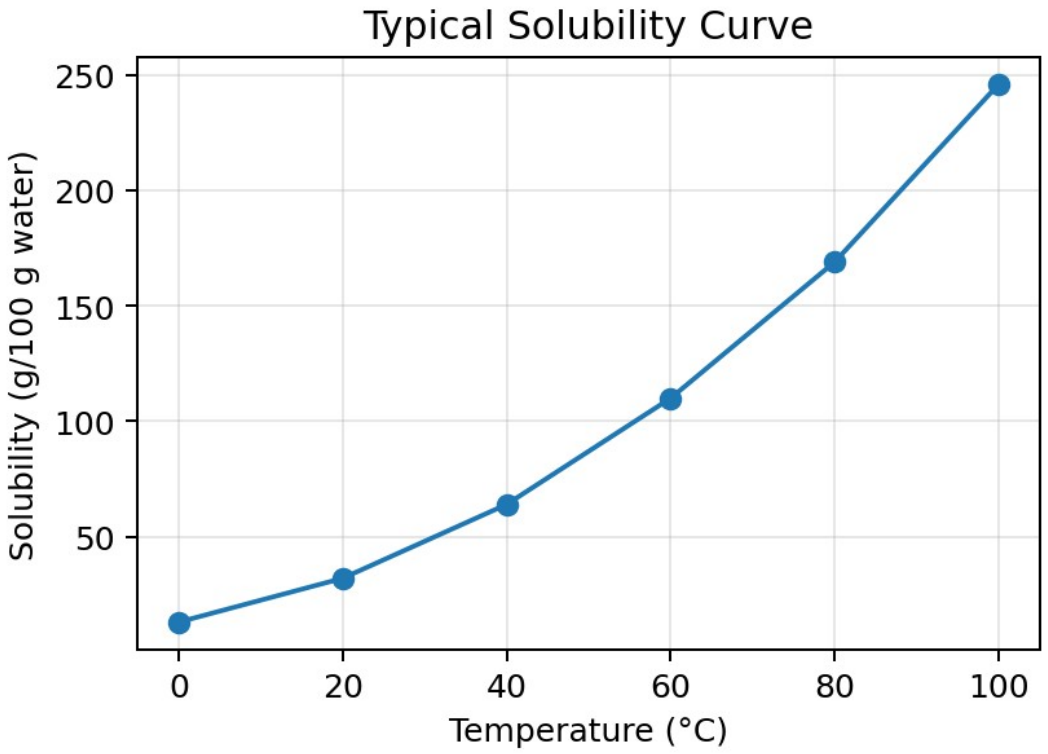
Summary

Salts are formed when hydrogen ions in acids are replaced by metal or ammonium ions. Soluble salts are prepared by reactions followed by crystallization, while insoluble salts are prepared by precipitation.

WEEK 7: SOLUBILITY OF SUBSTANCES AND SOLUBILITY CURVES

Learning Objectives

75. Define solute, solvent and solution.
76. Define solubility.
77. Differentiate saturated, unsaturated and supersaturated solutions.
78. State factors affecting solubility.
79. Interpret solubility curves.
80. Solve simple solubility problems.



A solubility curve shows how solubility changes with temperature.

Detailed Lesson Content

Solute, Solvent and Solution

A solute is the substance that dissolves. A solvent is the liquid that dissolves the solute. A solution is a uniform mixture formed when a solute dissolves in a solvent. In salt solution, salt is the solute and water is the solvent.

Solubility

Solubility is the maximum amount of solute that can dissolve in a given amount of solvent at a particular temperature. It is often expressed as grams of solute per 100 g of water.

Types of Solutions

A saturated solution contains the maximum amount of solute it can dissolve at a particular temperature. An unsaturated solution contains less solute and can still dissolve more. A supersaturated solution contains more solute than normally dissolves at that temperature and is unstable.

Factors Affecting Solubility

Temperature, nature of solute, nature of solvent and pressure affect solubility. For most solids, solubility increases with temperature. For gases, solubility usually decreases as temperature increases but increases with pressure.

Rate of Dissolving

Heating, stirring and crushing increase the rate at which a solute dissolves. These may not increase the final solubility, but they help the solute dissolve faster.

Solubility Curves

A solubility curve is a graph showing solubility against temperature. A point on the curve represents a saturated solution. A point below the curve is unsaturated, while a point above the curve is supersaturated.

Crystallization and Solubility

When a hot saturated solution cools, its solubility may decrease and excess solute separates out as crystals. This principle is used in crystallization.

Worked Examples / Important Points

- If 40 g dissolves at 60°C and 25 g dissolves at 30°C, crystals formed on cooling = 40 - 25 = 15 g.
- If 50 g dissolves in 100 g water at 40°C, solubility = 50 g per 100 g water at 40°C.

Class Activities

81. Dissolve sugar or salt in cold water and hot water and compare the rate of dissolving.

Evaluation Questions

82. Define solute.
83. Define solvent.
84. What is solubility?
85. Differentiate saturated and unsaturated solutions.
86. State four factors affecting solubility.
87. What is a solubility curve?

Assignment

A salt has solubility of 60 g per 100 g water at 80°C and 35 g per 100 g water at 30°C. Calculate the mass of crystals formed on cooling.

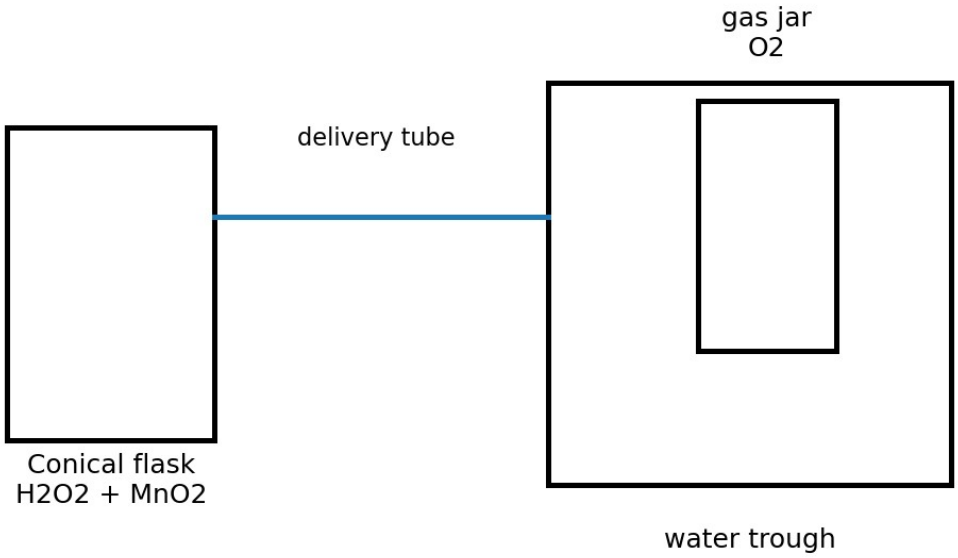
Summary

Solubility is the maximum amount of solute that dissolves in a solvent at a given temperature. Temperature, pressure and the nature of solute and solvent affect solubility. Solubility curves help predict crystallization.

WEEK 8: WATER I - SOURCES, PROPERTIES, USES AND TREATMENT OF WATER

Learning Objectives

88. State sources of water.
89. Describe physical and chemical properties of water.
90. State uses of water.
91. Explain water treatment.
92. State tests for water.



Oxygen is collected over water because it is only slightly soluble in water.

Laboratory and industrial processes often depend on pure or treated water.

Detailed Lesson Content

Introduction and Occurrence

Water is one of the most important substances on earth. Its formula is H2O, meaning that one molecule contains two hydrogen atoms and one oxygen atom. Water occurs as solid ice, liquid water and gaseous water vapour.

Sources of Water

Sources include rainwater, rivers, streams, lakes, wells, boreholes, springs, sea water, ponds, dams and reservoirs.

Physical Properties of Pure Water

Pure water is colourless, odourless, tasteless, neutral to litmus, has pH 7, freezes at 0°C, boils at 100°C at normal atmospheric pressure, has density about 1 g/cm3 at 4°C and is a good solvent.

Chemical Properties of Water

Sodium reacts vigorously with water to form sodium hydroxide and hydrogen: 2Na + 2H2O -> 2NaOH + H2. Calcium oxide reacts with water to form calcium hydroxide: CaO + H2O -> Ca(OH)2. Carbon dioxide dissolves in water to form carbonic acid. Water can be decomposed by electrolysis into hydrogen and oxygen.

Uses and Importance

Water is used for drinking, cooking, washing, bathing, irrigation, transport, hydroelectric power, cooling machines, fire fighting, laboratory work, digestion and photosynthesis.

Impurities and Treatment

Natural water may contain sand, mud, microorganisms, dissolved salts, organic matter, industrial chemicals and sewage. Treatment involves screening, sedimentation, filtration, chlorination and sometimes distillation. At home, water can be purified by boiling, filtration, chlorination and allowing particles to settle.

Tests for Water

Anhydrous copper(II) sulphate is white and turns blue in the presence of water. Dry cobalt chloride paper is blue and turns pink in the presence of water.

Worked Examples / Important Points

- $CO_2 + H_2O \rightarrow H_2CO_3$
- $CaO + H_2O \rightarrow Ca(OH)_2$
- $2H_2O \rightarrow 2H_2 + O_2$ by electrolysis.

Class Activities

93. Mention sources of water in the community and classify them as likely safe or unsafe for drinking.

Evaluation Questions

94. State five sources of water.
95. Give five physical properties of pure water.
96. Why is water called a universal solvent?
97. State the stages in water treatment.
98. State two tests for water.

Assignment

Write a detailed note on how water is treated before it becomes safe for drinking.

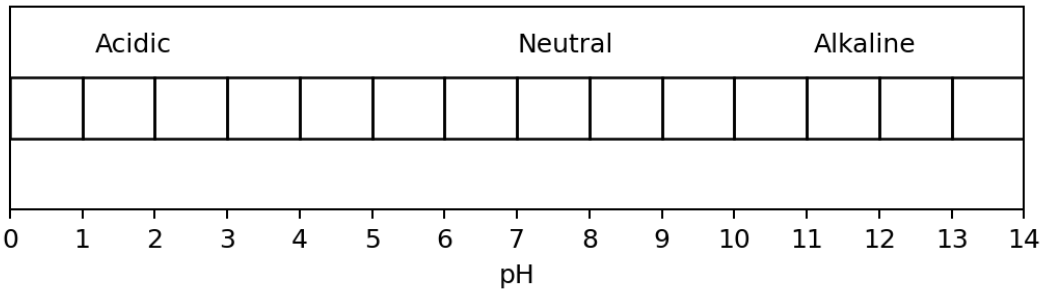
Summary

Water is essential for life and industry. Natural water may contain impurities, so it must be treated before use. Pure water is neutral and can be tested using anhydrous copper(II) sulphate or cobalt chloride paper.

WEEK 9: WATER II - HARDNESS OF WATER

Learning Objectives

99. Define hard and soft water.
100. State causes of hardness.
101. Differentiate temporary and permanent hardness.
102. Explain methods of removing hardness.
103. State advantages and disadvantages of hard water.



Hardness affects washing, boilers and domestic water use.

Detailed Lesson Content

Soft and Hard Water

Soft water lathers easily with soap. Hard water does not lather easily with soap because it contains dissolved calcium and magnesium ions. These ions react with soap to form scum.

Causes of Hardness

Hardness is caused by dissolved salts such as calcium hydrogen carbonate, magnesium hydrogen carbonate, calcium sulphate, magnesium sulphate, calcium chloride and magnesium chloride.

Temporary Hardness

Temporary hardness is caused by calcium and magnesium hydrogen carbonates. It can be removed by boiling. For example, Ca(HCO3)2 -> CaCO3 + H2O + CO2. The insoluble calcium carbonate can be filtered off.

Permanent Hardness

Permanent hardness is caused by calcium and magnesium sulphates or chlorides. It cannot be removed by boiling. It can be removed by adding washing soda, ion exchange or distillation.

Removal by Washing Soda

Washing soda, Na2CO3, removes calcium and magnesium ions by forming insoluble carbonates. For example, CaSO4 + Na2CO3 -> CaCO3 + Na2SO4.

Advantages and Disadvantages

Hard water supplies calcium needed for bones and teeth and may taste better. However, it wastes soap, forms scum, causes scale in kettles and boilers, blocks pipes and reduces efficiency of boilers.

Testing Hardness

Shake the water sample with soap solution. Soft water forms lather easily; hard water forms scum and lathers poorly.

Worked Examples / Important Points

- Temporary hardness: $Ca(HCO_3)_2$ and $Mg(HCO_3)_2$.
- Permanent hardness: $CaSO_4$, $MgSO_4$, $CaCl_2$ and $MgCl_2$.

Class Activities

104. Test rainwater, tap water and well water with soap solution and compare the amount of lather.

Evaluation Questions

105. What is hard water?
106. What is soft water?
107. State two causes of hardness.
108. Differentiate temporary and permanent hardness.
109. How can permanent hardness be removed?
110. State four disadvantages of hard water.

Assignment

Explain how washing soda removes permanent hardness of water using chemical equations.

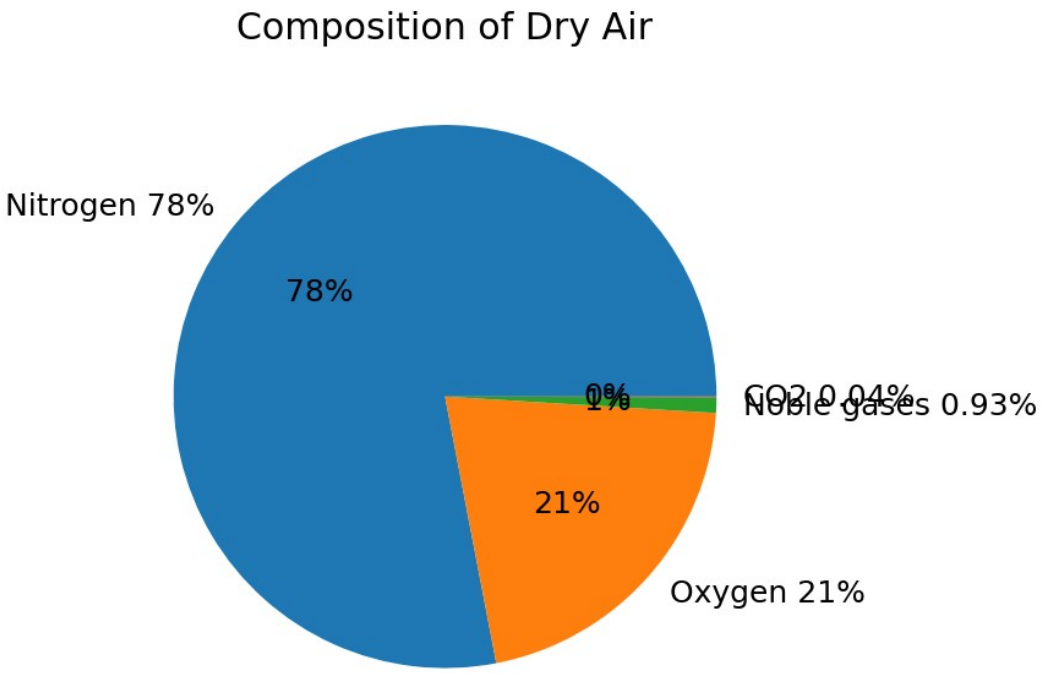
Summary

Hard water contains calcium and magnesium ions and does not lather easily with soap. Temporary hardness is removed by boiling, while permanent hardness is removed by washing soda, ion exchange or distillation.

WEEK 10: AIR AND ITS CONSTITUENTS

Learning Objectives

111. Define air.
112. State the composition of air.
113. Explain why air is a mixture.
114. Describe experiments showing oxygen, carbon dioxide and water vapour in air.
115. Explain air pollution and control.



Dry air is mainly nitrogen and oxygen.

Detailed Lesson Content

Meaning of Air

Air is a mixture of gases surrounding the earth. It is part of the atmosphere and is essential for life. Air is matter because it has mass and occupies space.

Composition of Air

Dry air contains about 78% nitrogen, 21% oxygen, 0.93% noble gases and about 0.04% carbon dioxide. It may also contain variable water vapour, dust and pollutants.

Why Air Is a Mixture

Air is a mixture because its components are not chemically combined, its composition can vary, each gas retains its properties, the gases can be separated by physical methods and air has no fixed chemical formula.

Importance of Constituents

Nitrogen is used in ammonia and fertilizers. Oxygen supports respiration and combustion. Carbon dioxide is used by plants for photosynthesis and in fire extinguishers. Noble gases are used in bulbs, balloons and advertising signs.

Experiments

A candle burning under an inverted gas jar over water goes out after oxygen is used up and water rises, showing oxygen is about one-fifth of air. Passing air through limewater turns it milky, showing carbon dioxide. A cold dry glass exposed to air becomes wet outside, showing water vapour.

Combustion and Rusting

Combustion is burning in oxygen to produce heat and light. Rusting is the slow reaction of iron with oxygen and water to form hydrated iron(III) oxide. Rusting can be prevented by painting, oiling, greasing, galvanizing, electroplating, alloying and plastic coating.

Air Pollution

Air pollution is the release of harmful substances into the air. Pollutants include carbon monoxide, sulphur dioxide, nitrogen oxides, smoke, dust and unburnt hydrocarbons. Sources include vehicles, factories, bush burning, refuse burning and gas flaring. Effects include respiratory diseases, acid rain, global warming and damage to plants and buildings.

Worked Examples / Important Points

- Ca(OH)2 + CO2 -> CaCO3 + H2O (limewater turns milky).
- Rusting requires oxygen and water.

Class Activities

116. List five sources of air pollution in the environment and suggest ways to reduce them.

Evaluation Questions

117. What is air?
118. State the composition of dry air.
119. Why is air a mixture?
120. How can you show air contains carbon dioxide?
121. What is rusting?
122. State four methods of preventing rusting.

Assignment

Write a detailed note on air pollution, its causes, effects and control.

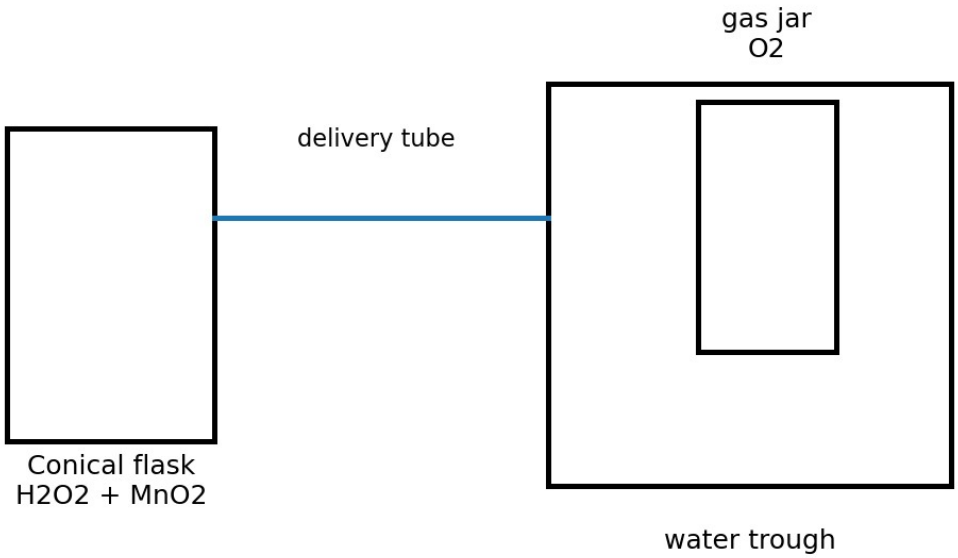
Summary

Air is a mixture of gases, mainly nitrogen and oxygen. It supports life and combustion. Air pollution harms living things and the environment and must be controlled through clean practices and environmental laws.

WEEK 11: OXYGEN - LABORATORY PREPARATION, PROPERTIES, USES AND TESTS

Learning Objectives

123. State the occurrence of oxygen.
124. Prepare oxygen in the laboratory.
125. State physical and chemical properties of oxygen.
126. Test for oxygen gas.
127. State uses of oxygen.
128. Explain oxides.



Oxygen is collected over water because it is only slightly soluble in water.

Oxygen can be prepared by decomposing hydrogen peroxide with manganese(IV) oxide catalyst.

Detailed Lesson Content

Occurrence of Oxygen

Oxygen occurs free in air as O2 and combined in water, oxides, minerals, rocks and living organisms. It makes up about 21% of air by volume.

Laboratory Preparation

Oxygen is prepared by decomposition of hydrogen peroxide using manganese(IV) oxide as catalyst: 2H2O2 -> 2H2O + O2. The gas is collected over water because it is only slightly soluble in water.

Catalyst

A catalyst is a substance that changes the rate of a chemical reaction without being used up at the end. Manganese(IV) oxide speeds up the decomposition of hydrogen peroxide.

Alternative Preparation

Oxygen can also be prepared by heating potassium trioxochlorate(V), KClO3, in the presence of manganese(IV) oxide: 2KClO3 -> 2KCl + 3O2.

Test for Oxygen

Oxygen relights a glowing splint. This is the confirmatory test for oxygen gas.

Physical Properties

Oxygen is colourless, odourless, tasteless, slightly soluble in water, neutral to litmus, slightly denser than air and condenses to a pale blue liquid at very low temperature.

Chemical Properties

Oxygen supports combustion and reacts with many elements to form oxides. Magnesium burns with a bright white flame to form magnesium oxide: 2Mg + O2 -> 2MgO. Carbon burns to form carbon dioxide: C + O2 -> CO2. Sulphur burns with a blue flame to form sulphur dioxide: S + O2 -> SO2.

Oxides and Uses

An oxide is a compound of oxygen and another element. Metallic oxides are usually basic, non-metal oxides are usually acidic, some oxides are neutral and some are amphoteric. Oxygen is used in respiration, hospitals, welding, steel making, rockets, diving, space travel and sewage treatment.

Worked Examples / Important Points

- 2H2O2 -> 2H2O + O2
- 2Mg + O2 -> 2MgO
- C + O2 -> CO2
- S + O2 -> SO2

Class Activities

129. Observe the teacher demonstrate oxygen preparation and test the gas with a glowing splint.

Evaluation Questions

130. State three sources of oxygen.
131. How is oxygen prepared in the laboratory?
132. What is the function of MnO2?
133. Why is oxygen collected over water?
134. What is the test for oxygen?
135. State five uses of oxygen.

Assignment

Write equations for the reaction of oxygen with magnesium, carbon, sulphur, calcium and iron.

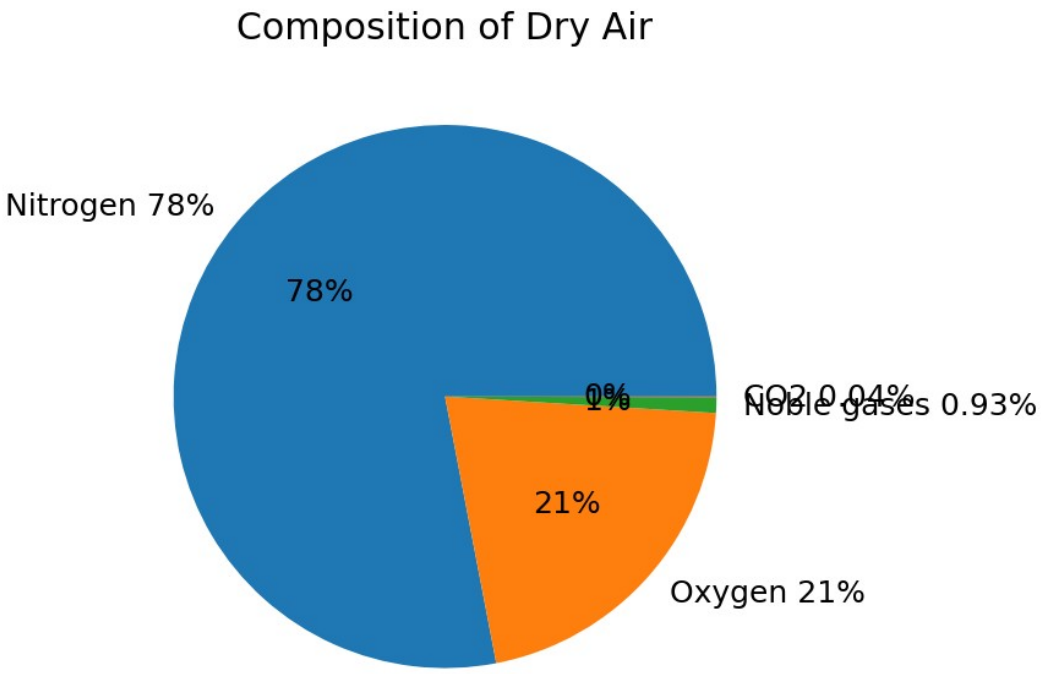
Summary

Oxygen occurs in air and many compounds. It can be prepared from hydrogen peroxide using manganese(IV) oxide. Oxygen supports combustion, relights a glowing splint and is used in respiration, medicine, welding and industry.

WEEK 12: REVISION AND EXAMINATION FOCUS

Learning Objectives

136. Revise all topics taught in second term.
137. Answer WAEC/NECO-standard questions.
138. Identify key examination areas.
139. Avoid common mistakes.
140. Prepare for theory, objective and practical questions.



Revision connects all second term concepts for examination success.

Detailed Lesson Content

Topics for Revision

Revise chemical bonding; ionic, covalent, coordinate and metallic bonding; acids, bases and salts; indicators and pH; preparation and uses of salts; solubility and solubility curves; water treatment; hardness of water; air and constituents; oxygen preparation, properties and uses.

Important Definitions

Students should master definitions of chemical bond, ionic bond, covalent bond, coordinate covalent bond, metallic bond, acid, base, alkali, salt, solubility, saturated solution, hard water, soft water, catalyst and oxide.

Important Diagrams

Students should practise diagrams for NaCl electron transfer, water covalent bonding, preparation of oxygen, solubility curve, pH scale and composition of air.

Important Equations

Practise balancing equations: Zn + 2HCl -> ZnCl2 + H2; HCl + NaOH -> NaCl + H2O; CaCO3 + 2HCl -> CaCl2 + H2O + CO2; 2H2O2 -> 2H2O + O2; 2Mg + O2 -> 2MgO; C + O2 -> CO2; S + O2 -> SO2; Ca(HCO3)2 -> CaCO3 + H2O + CO2.

Common Mistakes

Avoid calling covalent bonding transfer of electrons; saying solid ionic compounds conduct electricity; confusing alkali with base; saying oxygen burns instead of oxygen supports burning; forgetting that air is a mixture; writing wrong formulae; and failing to balance equations.

Worked Examples / Important Points

- Objective practice: Ionic bonding involves transfer of electrons.
- Theory practice: Explain the formation of sodium chloride using electron transfer.
- Practical practice: Oxygen relights a glowing splint.

Class Activities

141. Students should answer revision questions individually and discuss corrections in class.

Evaluation Questions

142. Define chemical bonding and state three types.
143. Explain formation of sodium chloride.
144. Differentiate ionic and covalent compounds.
145. Describe preparation of copper(II) sulphate.
146. Explain temporary and permanent hardness.
147. Describe laboratory preparation of oxygen.

Assignment

Prepare for the second term examination by revising all notes, equations, diagrams and class activities.

Summary

Second term Chemistry explains how atoms combine, how acids and bases behave, how salts are prepared, how substances dissolve, and the chemistry of water, air and oxygen.

Appendix A: Quick Reference Tables

Indicator	Acidic Solution	Neutral Solution	Alkaline Solution
Blue litmus	Turns red	No change	Blue
Red litmus	Red	No change	Turns blue
Methyl orange	Red	Orange	Yellow
Phenolphthalein	Colourless	Colourless	Pink
Universal indicator	Red/Orange/Yellow	Green	Blue/Purple

Substance / Concept	Key Test or Point
Hydrogen gas	Burns with a pop sound.
Carbon dioxide	Turns limewater milky.
Oxygen gas	Relights a glowing splint.
Water	White anhydrous copper(II) sulphate turns blue; blue cobalt chloride paper turns pink.
Hard water	Forms scum and does not lather easily with soap.

Bond Type	How Formed	Common Particles	Electrical Conductivity

Ionic	Transfer of electrons	Positive and negative ions	Conducts when molten or in aqueous solution
Covalent	Sharing of electrons	Molecules or giant atoms	Usually does not conduct
Metallic	Attraction between positive metal ions and delocalized electrons	Metal ions and free electrons	Conducts in solid and molten states