Chemistry Syllabus, Grade 9

General Objectives of Grade 9 Chemistry

To Develop Understanding and Acquire knowledge of:

- Dalton's atomic theory and modern atomic theory,
- Discoveries of the sub-atomic particles,
- The relation ship between the sub-atomic particles and the atomic numbers, mass numbers, atomic masses and isotopes,
- Development of the atomic models and arrangement of the electrons in the atoms,
- Periodic classification of the elements and its importance in studying chemistry,
- The major types of chemical bonding and intermolecular forces,
- Types of chemical reactions and representing chemical reactions by chemical equations and
- The three physical states of matter.

To Develop Skills and Abilities of:

- Writing electron configuration for the elements,
- Drawing structures and make models for some molecules,
- Writing balanced equations and solving related problems,
- Design and conduct simple experiments relevant to their level and
- Applying gas laws' equations to solve gas variables(pressure, temperature and volume).

To Develop the Habit and Attitude of:

- Realizing that properties of substances are attributed to their atomic structures,
- Appreciating problem solving,
- Appreciating that all substances are the results of chemical combination of the atoms and
- Develop personality characteristics such as neatness, exactness, diligence, responsibility and carefulness.

Unit 1: Structure of the atom *(15 periods)*

Unit outcomes: Students will be able to:

- comprehend Dalton's atomic theory and modern atomic theory;
- understand the discovery of the electron and the nucleus;
- know the terms like atomic number, mass number, atomic mass, isotope, energy level, valence electrons and electron configuration;
- understand the Dalton, the Thomson, the Rutherford, the Bohr and the quantum mechanical atomic models;
- develop skills in
  - determining the number of protons, electrons and neutrons of atoms from atomic numbers and mass numbers,
  - calculating the atomic masses of elements that have isotopes,
  - writing the ground state electron configurations of atoms using sub-energy levels and drawing diagrammatic representations of atoms.
- demonstrate scientific inquiry skills: observing, comparing and contrasting, communicating, asking questions, and applying concepts.

Competencies	***Contents***	***Suggested Activities***
Students will be able to: • Describe Dalton's atomic theory	**1. Structure of the atom** **1.1 Atomic theory** *(3 periods)* • Dalton's atomic theory	Students should appreciate that the idea of atoms as the building blocks from which all matter is formed was first suggested by the ancient Greeks although they had no evidence to support this theory. Students should know that in 1808 the scientist John Dalton proposed an atomic theory in which he suggested that: 1. Elements are made of small particles called atoms 1. Atoms can neither be created nor destroyed 2. All atoms of the same element are identical and have the same mass and size 3. Atoms of different elements have different masses and size 4. Atoms combine in small whole numbers to form compounds Students should discuss each point in Dalton's atomic theory and determine whether it is still valid today.
• Describe the modern atomic theory • Compare and contrast Dalton's atomic theory and the modern atomic theory	• Modern atomic theory	Students should appreciate that at the time that Dalton's theory was proposed nothing was known about the internal structure of the atom. As a result of our increasing knowledge about atomic structure we now know that statement #2 and statement #3 are no longer true. Students should attempt to modify Dalton's statements in the light of modern knowledge about the atom. The modified statements could be: 1. Elements are made of small particles called atoms 2. Atoms cannot be created or destroyed during ordinary chemical reactions 3. All atoms of the same element have the same atomic number but may vary in mass number due to the presence of different isotopes 4. Atoms of different elements are different 5. Atoms combine in small whole numbers to form compounds

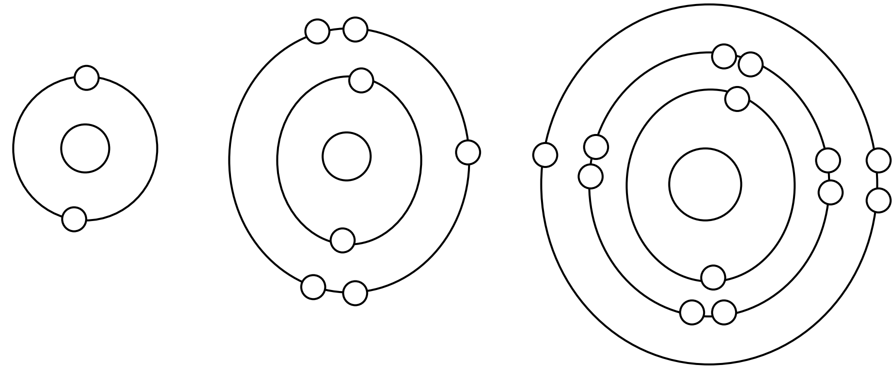
Competencies	Contents	Suggested Activities
		Students could be shown the symbols that Dalton used to represent atoms of different elements and discuss why these symbols are no longer used.
	1.2 Discoveries of fundamental subatomic particles and the atomic nucleus *(3 periods)*	Students should appreciate that our modern understanding of atomic structure came about as a result of the discovery of its component particles. Students should be aware that the electron was discovered by the physicist J.J. Thomson.
• Explain the discovery of electron	• Discovery of the electron	Students should be able to give a brief summary of the famous experiment carried out by Thomson. This could include the following points: • The apparatus consisted of two electrodes in a tube from which most of the air had been removed • When a high potential difference was applied across the electrodes a stream of particles passed between them • Thomson called these particles cathode rays because they came from the cathode • The cathode rays moved straight to the anode, and through a hole in it, to a screen coated with zinc sulphide • The zinc sulphide glowed when struck by the cathode rays • The cathode rays were deflected by both magnetic and electrical fields • The cathode rays were attracted by the +ve electric plate but repelled by the –ve electric plate • Thomson obtained the same results when he changed the gas and the other materials in the tube • A cathode ray consists of a stream of electrons Students should discuss how, from this evidence, Thomson was able to state that: • An electron carries a negative charge. • An electron is a fundamental constituent of all matter
• Explain the discovery of nucleus	• Discovery of the nucleus	Students should be aware that the nature of the nucleus of an atom was discovered by the physicist Rutherford and two of his students; Geiger and Marsden Students should be able to give a brief summary of the famous experiment carried out by them. This could include the following points: • A stream of alpha particles was fired at a piece of thin gold foil • The scattered alpha particles were detected by a small fluorescent screen viewed through a microscope • The screen and microscope could be moved in a circle around the gold foil • Most of the alpha particles went through the gold foil without being deflected • A small proportion of the alpha particles were deflected through small angles

Competencies	***Contents***	***Suggested Activities***
		• A tiny proportion of alpha particles (1 in 20 000) were deflected backwards Students should discuss how, from this evidence, Rutherford was able to state that: • The nucleus of an atom is positively charged • Most of the mass of an atom is in the nucleus • There are large spaces within an atom
• Explain the discovery of neutron	• Discovery of the neutron	Students could research why Rutherford predicted the existence of the neutron and how this was demonstrated by Chadwick in 1932. Students should be able to give a brief summary of the experiment carried out by James Chadwick in 1932 • Bombarded Beryllium atoms with alpha particles and obtained a radiation Showed that this radiation consisted of neutral particles called neutrons
• Write the relative charges of an electron a proton and a neutron • Tell the absolute and relative masses of an electron, a proton and a neutron	**1.3 Composition of an atom and the isotopes** *(3 periods)* • Electrons, protons and neutrons	Students should be able to give the relative mass and charge of each of the fundamental particles in an atom and state in which part of the atom it is found. Students should appreciate that the masses of the sub-atomic particles are very small but they do have absolute values: • Neutron – 1.675×10^{-24}g • Proton – 1.672×10^{-24}g • Electron – 9.11×10^{-28}g Students should understand that the mass of proton is nearly equal to that of a neutron. They should also recognize that the mass of an electron is negligible(too small). Students should appreciate that the nucleus of an atom is composed of protons and neutrons.
• Tell the number of protons and electrons in an atom from the atomic number of the element • Determine the number of neutrons from given values of atomic numbers and mass numbers	• Atomic number and mass number	Students should know that: • The atomic number of an atom is equal to the number of protons in the nucleus • The atomic number is also equal to the number of electrons in a neutral atom • The mass number of an atom is equal to the number of protons + the number of neutrons in the nucleus • The number of neutrons in the nucleus of an atom is equal to the mass number minus the atomic number Students should find the number of protons and electrons in atoms from their atomic numbers. Students should find the atomic numbers and mass numbers of atoms of some elements and use them to calculate the number of neutrons present in each atom.

Competencies	***Contents***	***Suggested Activities***
• Explain the terms atomic mass and isotope • Calculate the atomic masses of elements that have isotopes.	• Atomic mass, and isotope	Students should be able to indicate the atomic number and the mass number of an atom of an element by writing the mass number as subscript and the atomic number as superscript , placing them both before the symbol of the element $^{\text{Mass no}}_{\text{atomic no}}\text{SYMBOL OF THE ELEMENT}$ Students should understand that in the atoms of an element: • the number of protons is always the same • the number of neutrons varies slightly Students should know that atoms of the same element which have different numbers of neutrons will have different mass numbers and are called isotopes. Students should appreciate that a natural sample of an element is likely to contain a mixture of two or more isotopes. In determining the atomic mass of the element we must take into account that it is a mixture of isotopes with different mass numbers. Students should be aware that: • Some elements are composed almost entirely of one isotope therefore the atomic mass of the element will be the same as the mass number of the isotope – for example carbon exists as a mixture of the isotopes carbon-12, carbon-13 and carbon-14, but carbon-12 accounts for 98.9% of the element. • Some elements contain significant proportions of different isotopes and an accurate value for the atomic mass is obtained by taking the mass number and percentage of each isotope into account – for example, boron exists as a mixture of the isotopes boron-10, 20%, and boron-11, 80%. Therefore both isotopes make a significant contribution to the atomic mass of boron. The atomic mass is equal to the weighted average mass of its isotopes. Students should calculate the atomic mass of boron using the following method: $10 \times \frac{20}{100} + 11 \times \frac{80}{100} = 2 + 8.8 = 10.8$ Students could apply this method to find accurate values for the atomic masses of other elements such as copper and chlorine.

Competencies	*Contents*	*Suggested Activities*
• Name the five atomic models	**1.4 The atomic models** *(6 periods)*	Students should appreciate from the work already carried out in this unit that scientists picture of the nature of an atom has changed over the years as a result of new discoveries being made. Students should understand the important features of a number of atomic models proposed over the years.
• Describe the Dalton, Thomson and Rutherford Model	• The Dalton atomic model	Dalton model of the atom: • A small hard ball which is indivisible and cannot be destroyed
	• The Thomson atomic model	Thomson model of the atom: • A positively charged solidly sphere • Electrons stuck uniformly in it like plums in a pudding
	• The Rutherford atomic model	Rutherford model of the atom: • A nucleus consisting of positively charged protons • Electrons moving around the nucleus
• State Bohr's Postulates • Describe the Bohr's model	• The Bohr atomic model - Bohr's postulates	Bohr model of the atom: • A nucleus consisting of positively charged protons and neutrons which carry no charge • Electrons moving around the nucleus in a set of orbits • For any atom there is a fixed set of orbits possible • The energy of an electron remains the same as long as it stays in the same orbit • When electrons move between orbits they release or absorb energy • When electrons fall from a higher (excited) state to a lower (ground) state they give out fixed amounts of energy. But, when the electrons jump from a lower to a higher energy state they absorb fixed amounts of energy. Students should appreciate how the picture of the atom became more sophisticated as scientists gained knowledge about it.
• Describe the quantum mechanical model • Describe main energy level and sub energy level	• The quantum mechanical atomic model - Main energy level - Sub-energy level	Students should appreciate that our modern picture of the structure of the atom is based on a branch of science called quantum mechanics. Students should know that quantum mechanics came about because scientists found that the tiny particles in an atom did not obey the classical laws of physics postulated by Sir Isaac Newton. Students should appreciate that in the quantum mechanics atomic model: • Electrons are located in orbitals • An orbital is a volume of space inside which there is a high probability of finding a specified electrons • Orbitals are represented as clouds of charge

Competencies	*Contents*	*Suggested Activities*
		• Energy levels in atoms are arranged in series in which adjacent levels get closer moving up the series • With the exception of the first energy level, each main energy level has sub-levels associated with it Students should know that: • Energy levels are numbers 1, 2, 3, 4, etc. or letters K, L, M, N etc. • Energy sub levels are given the letters s, p, d, f • Each type of sub level has a set number of orbitals: - s – 1 orbital - p – 3 orbitals - d – 5 orbitals - f – 7 orbitals • Each orbital can have a maximum of 2 electrons • The positions of the electrons around the nucleus of an atom in orbitals and sub orbitals is described as the electronic configuration of the atom
• Define the term electronic configuration	- Electronic configuration	Students should know the order in which orbitals are filled and that this corresponds to placing electrons in the lowest energy levels available: 1s, 2s, 2p, 3s, 3p, 4s, 3d. etc Students should note that the 4s orbital is filled before the 3d orbital because it is lower in energy.
• Write the ground state electronic configuration of the elements		Students should be able to write the ground state electron configurations of the elements a suitable form such as: • Hydrogen – $1s^1$ • Helium – $1s^2$ • Lithium – $1s^2$, $2s^1$ etc.
• Draw diagrams to show the electronic configuration of the first 18 elements		Students should be able to draw diagrams to show the electron configurations of the first 20 elements. Students should consider some examples. This could include the diagrammatic representation of the electronic configurations of the elements He, N & Al

Competencies	***Contents***	***Suggested Activities***
		 $_{2}He$: 2 $_{7}N$: 2.5 $_{13}Al$: 2, 8, 3 Students could draw the electronic configurations of other elements.
• Write the electronic configuration of the elements using sub energy levels • Write electronic configuration of the elements using noble gas as a core		Students should write electronic configuration of the elements using sub energy levels. They should also use short cut for electron configuration of elements with large atomic numbers using noble gas as a core.
• Describe valence electrons	- Valence electrons	Students should know that the electrons in the outermost orbital have the highest energy and are called valence electrons. Students should appreciate that valence electrons are: • furthest from the nucleus of the atom • the most easily lost responsible for the chemistry of the element

Assessment

The teacher should assess each student's work continuously over the whole unit and compare it with the following description, based on the Competencies, to determine whether the student has achieved the minimum required level.

Students at minimum requirement level

Students working at the minimum requirement level will be able to: comprehend Dalton's atomic theory and modern atomic theory; explain the discovery of electron, neutron and nucleus; write the relative charges, relative masses and absolute masses of electron, proton and neutron; determine the No of neutrons of an element from given values of atomic number and mass number; explain the terms atomic mass and isotopes; calculate the atomic masses of elements that have isotopes, name the five atomic models and state Bohrs postulates, describe main energy level and sub energy level; write electronic configuration of elements and show them diagrammatically and describe valence electrons.

Students above minimum requirement level

Students working above the minimum requirement level should be praised and their achievements recognized. They should be encouraged to continue working hard and not become complacent.

Students below minimum requirement level

Students working below the minimum requirement level will require extra help if they are to catch up with the rest of the class. They should be given extra attention in class and additional lesson time during breaks or at the end of the day.

Unit 2: Periodic classification of elements *(13 periods)*

Unit outcomes: Students will be able to:

- Understand the periodic classification of the elements.
- Develop the skills of correlating the electron configuration of elements with the periodicity of the elements, predicting the trends of periodic properties of elements in the periodic table.
- Appreciate the importance of classification in chemistry.
- Demonstrate scientific inquiry skills: observing, inferring, predicting, classifying, comparing and contrasting, making models, communicating, measuring, asking questions, interpreting illustrations, drawing conclusion, applying concepts and problem solving.

Competencies	***Contents***	***Suggested Activities***
Students will be able to:	**2. Periodic classification of the elements**	
• Describe periodicity	**2.1 Introduction** *(1 period)*	Students should understand that in classifying elements, scientists were guided by the similarities in chemical properties. The elements are arranged in a table in such a way that elements with similar chemical properties appear at regular intervals or periods. Students could revise early attempts to classify elements such as Dobereiner's Triads and Newland's Octaves.
	2.2 The modern Periodic Table *(5 periods)*	Students should be aware that the modern Periodic Table is based on the work of the Russian chemist Dmitri Mendeleev.
• State Mendeleev's Periodic law • State modern periodic law	• The Periodic law	Students could research the work of Mendeleev, his foresight in leaving a gap between silicon and tin, and his predictions about the properties of the missing elements eg. Germanium. Students should know that the rows of the Periodic Table are called periods. They should understand that elements in the same period:
• Describe period • Describe group	• Groups and periods	• Have the same number of main energy levels • Increase in atomic number by one unit passing across the period • Decrease in metal properties passing from left to right Students should know that the columns of the Periodic Table are called groups. They should understand that elements in the same group: • Have the same number of electrons in the outermost shell • Have similar chemical properties
• Explain the relationship between the electronic configuration and the structure of the modern Periodic Table	• Classification of the elements	Students should compare the electronic configurations of the first 20 elements with their positions in the Periodic Table. They should see that: • Elements whose valence electrons are in s orbital appear to the left-hand side (in Groups 1 and 2) • Elements whose valence electrons are in p orbital appear to the right-hand side (in Groups 3 to 8)

Competencies	***Contents***	***Suggested Activities***
• Describe the three classes of the elements in the modern Periodic Table • Explain the four blocks of the elements related with their electronic configuration in the modern Periodic Table Tell the block of an element from its electronic configuration	- The representative elements (s-and p-blocks) - The transition metals (d- block) -The rare earths (f-block)	Students should understand that this can be used to classify elements in blocks with the Periodic Table e.g. • s-block elements are those in Groups 1 and 2 • p-block elements are those in Groups 3 to 8 Students could write the electronic configurations for elements with atomic numbers 21 to 36 of the Periodic Table. From this they will see that another block of elements emerges in which the valence electrons are in d orbital. Students should appreciate that d-block elements appear between the s-block and p-block elements. Students should appreciate that: • the d-block elements are sometimes called the transition metals • the d-block contains four series of elements corresponding to valence electrons in the 3d, 4d, 5d and 6d Students should be aware that there are two series of f-block elements corresponding to valence electrons in the 4f and 5f orbital. These series are collectively called the rare earths
• Give group names for the main group elements		Students should be aware that some groups have traditional names and others as family of the first member in the group. which are often used in textbooks: Group 1 – alkali metals Group 2 – alkaline earth metals Group 3 - boron family Group 4 - carbon family Group 5 - nitrogen family Group 6 - oxygen family Group 7 – halogens Group 8 – noble gases
• Classify the periods into short, long and incomplete periods • Tell the number of groups and periods in the modern Periodic Table • Tell the number of elements in each period		Students should be aware that some periods are long while others are short and are in incomplete. Students should be able to give examples of each Students should recall the number of elements each period of the Periodic Table. Students should be able to: • Identify an element from its group and period • Identify the group and period of an element from its atomic number • Identify the block to which an element belongs from its electronic configuration. Students should be given information about elements and from the information, identify the element and its position in the Periodic Table.

Competencies	***Contents***	***Suggested Activities***
• Predict the period and group of an element from its atomic number • Tell the block and group of an element from its electronic configuration		
	2.3 Periodic properties in the Periodic Table *(6 periods)*	
• Explain the general trends in properties of elements down a group of the Periodic Table	• Periodic properties within a group - Nuclear charge - Atomic radius - Ionization potential - Electron affinity - Electro negativity - Metallic character	Students should appreciate that all of the elements in a group have the same number of electrons in the outer shell, so the elements will have some similarities, but different numbers of main energy levels so they will also have some differences. Students should consider the following characteristics of the elements in Group 1. They should discuss and explain any trends which are apparent. • Nuclear charge – number of protons in the nucleus • Atomic radius – half the distance between the nuclei of adjacent atoms in a substance • Ionization potential – the energy needed to form an ion from an atom by removing valence electron(s) (to form a positive ion) • Electron affinity – the ease with which an atom will accept an electron (to form a negative ion) • Electro negativity- the ability of an atom, when it is in a molecule, to attract or pull electrons towards itself • Metallic character – the tendency to lose electrons. A similar study should be carried out on the elements in other groups. Students should compare the trends in Group 1 or 2 (a group of metals) with those in Group 7 (a group of non-metals). Students could carry out practical work that exemplifies trend in reactivity down a group e.g. • Reaction of Group 2 metals with water • Displacement reactions involving Group 7 halogens and metal halide solutions Students should be able to use their knowledge of the trends in reactivity passing down a group to make predictions about the reactivity of elements. Students should understand whether these properties increase or decrease down a group and able to give reasons.
• Explain the general trends in properties of elements across a period	• Periodic properties within a period	Students should appreciate that all of the elements in a period will have the same main energy level but an increasing number of electrons in the outer shell. Students should consider the following characteristics of the elements in Period 2. They should

Competencies	***Contents***	***Suggested Activities***
of the Periodic Table • Deduce the properties of an element from its position in the Periodic Table • Make a charts to show the trends in properties of elements in the Periodic Table	- Nuclear charge - Atomic radius - Ionization potential - Electron affinity - Electro negativity - Metallic character	discuss and explain any trends which are apparent. • Nuclear charge • Atomic radius • Ionization potential • Electron affinity • Electro negativity • Metallic character A similar study could be carried out on the elements in Period 3. Students should compare the trends across each period. Students should use their knowledge about group and periodic trends to make predictions about the physical and chemical properties of some unfamiliar elements. They could then research the properties of these elements and evaluate how accurate their predictions were. Students should make some form of chart to show the trend in properties of a group or period of elements.
• Describe the advantages of the periodic classification in the study of chemistry	**2.4 Advantages of periodic classification** *(1 period)*	Students should be able to use their knowledge of classification of elements to - read names, symbols, atomic numbers and atomic masses - state period and group of element - deduce number of shells, valence electrons, behaviour of an element and the nature of the oxides of element. Students should appreciate that the Periodic Table puts elements with similar properties into groups with gradual change in reactivity. In periods, some changes are gradual while others are abrupt. Students should be aware that the Periodic Table allows us to make predictions about the properties of elements. This can also be extended to their compounds. Students could be given data on some compounds of elements in the same group and asked to identify any similarities in physical and chemical properties. They could make predictions about similar compounds of the same elements. Students could consider and attempt to explain the diagonal relationship, based on similar chemical properties, between lithium and magnesium.

Assessment

The teacher should assess each student's work continuously over the whole unit and compare it with the following description, based on the Competencies, to determine whether the student has achieved the minimum required level.

Students at minimum requirement level

Students working at the minimum requirement level will be able to: Describe periodicity; state Mendeleev's periodic law and the modern periodic law; describe period and group; explain the relationship between electronic configuration of the elements and the structure of the periodic table; describe the three classes and the four blocks of elements in the periodic table; give group names for the main group elements; classify periods into short, long and incomplete periods; tell the number of groups and periods in the periodic table and the number of elements in each period; predict the period and group of an element from its atomic number; explain the general trends in properties of elements down a group and across a period of the periodic table; deduce the properties of an element from its position in the periodic table and describe the advantages of the periodic classification of the elements.

Students above minimum requirement level

Students working above the minimum requirement level should be praised and their achievements recognized. They should be encouraged to continue working hard and not become complacent.

Students below minimum requirement level

Students working below the minimum requirement level will require extra help if they are to catch up with the rest of the class. They should be given extra attention in class and additional lesson time during breaks or at the end of the day.

Unit 3: Chemical bonding and intermolecular forces *(17 periods)*

Unit outcomes: Students will be able to:

- discuss the formation of ionic, covalent and metallic bonds;
- know the general properties of substances containing ionic, covalent and metallic bonds;
- develop the skills of drawing the electron dot or Lewis structures for simple ionic and covalent compounds;
- understand the origin of polarity within molecules;
- understand the formation and nature of intermolecular forces;
- appreciate the importance of intermolecular forces in plant and animal life;
- demonstrate scientific inquiry skills: observing, predicting, making model, communicating, asking questions, measuring, applying concepts, comparing and contrasting, relating cause and effects.
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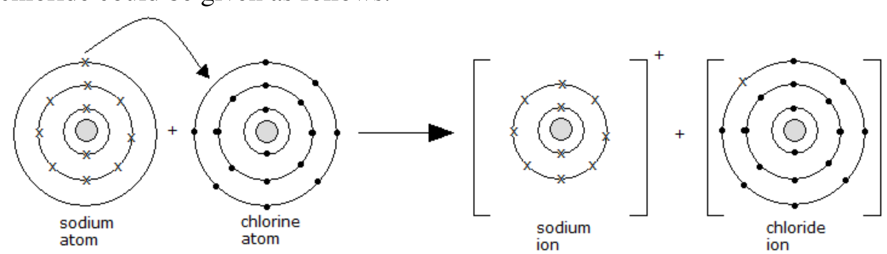
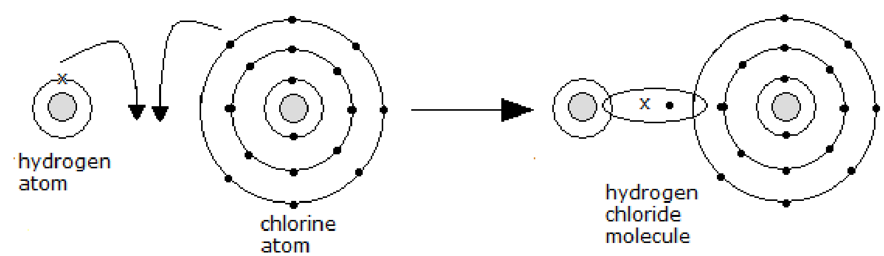
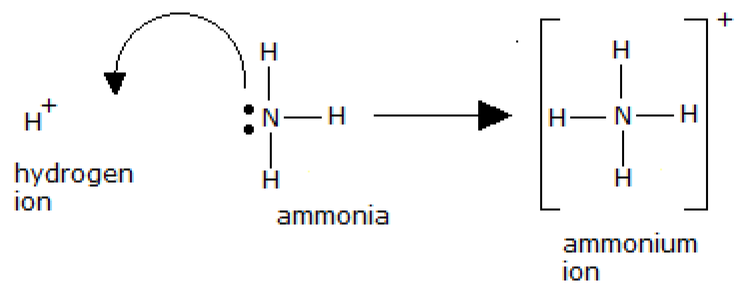
Competencies	***Contents***	***Suggested Activities***
Students will be able to:	**3. Chemical bonding and inter- molecular forces**	
• Define chemical bonding • Explain why atoms form chemical bonds	**3.1 Chemical bonding** *(1 period)*	Students should appreciate that a chemical bond is any force of attraction between two particles. The particles may be atoms, ions or molecules. Students should understand that atoms form bonds in order to achieve a more stable electron arrangement – a full outer shell of electrons. This can be achieved by: • Transferring one or more electrons – ionic bonding • Sharing one or more pairs of electrons – covalent bonding Energy changes occur during bond formation – these usually result in the final substance being more stable. Students should be aware that atoms which already have a full outer shell of electrons – the elements in Group 8 (noble gases) – have little chemistry as they do not readily form bonds.
• Explain the term ion • Illustrate the formation of ions by giving examples • Define ionic bonding	**3.2 Ionic bonding** *(3 periods)* • Formation of ionic bonding	Students should know that an ion is an atom which has lost or gained one or more electrons. • Metals tend to lose electrons to form positively charged ions or cations • Non-metals tend to gain electrons to form negatively charged ions or anions • Hydrogen forms both a cation, H^+ (hydrogen ion), and an anion, H^- (hydride ion) Students should be able to relate the position of elements in the Periodic Table to the normal ion formed by them: • Group 1 – M^+ e.g. Na^+ • Group 2 – M^{2+} e.g. Mg^{2+} • Group 3 – M^{3+} e.g. Al^{3+} • Group 5 – X^{3-} e.g. N^{3-} • Group 6 – X^{2-} e.g. O^{2-}

Competencies	***Contents***	***Suggested Activities***
• Describe the formation of an ionic bond		• Group 7 – X^- e.g. Cl^- Students could be asked to write the formulas of the ions formed by atoms of different elements. Students should appreciate that the easiest way for a metal to achieve a full outer shell of electrons is by losing electrons e.g. • sodium atom (2.8.1) → sodium ion (2.8) The alternative to this would be for the sodium atom to gain 7 electrons which would be very unstable.
• Give examples of simple ionic compounds		Similarly the easiest way for a non-metal to achieve a full outer shell of electrons is by gaining electrons e.g. • chlorine atom (2.8.7) → chloride ion (2.8.8) The alternative to this would be for the chlorine atom to lose 7 electrons which would require a large amount of energy and produce an unstable ion. Students should be aware that ionic bonding is the result of forces of attraction between oppositely charged ions. Students should be able to show how an ionic bond is formed by the transfer of one or more electrons from a metal atom to a non-metal atom.
• Draw Lewis structures or electron-dot formulas of simple ionic compounds	• Lewis Formulas of ionic compounds	Students should represent the formation of an ionic compound using Lewis structure. An example of dot and cross diagram (Lewis formula) for sodium and chlorine to form sodium chloride could be given as follows.  (2,8,1) (2,8,7) (2,8) (2,8,8) Students should also be able to represent the Lewis structure using symbols. eg. Na + Cl → Na^+ $[Cl]^-$ Students should appreciate that, when forming ions, the name of a non-metal takes the ending –ide e.g. nitride, oxide, chloride. Students could name the ionic compounds formed from given metals and non-metals and draw Lewis structures to show the ionic bonding.

Competencies	***Contents***	***Suggested Activities***
• Explain the general properties of ionic compounds.	• General properties of ionic compounds	Students should discuss general properties of ionic compounds including: • crystalline nature • high melting points and boiling points • ability to conduct an electric current when molten or in aqueous solution • solubility in polar solvents like water Students should be aware that solubility in water in itself is not a proof that a compound is ionic. There are ionic compounds which are effectively insoluble in water, and covalent compounds like glucose which are very soluble.
• Investigate the properties of given samples of ionic compounds.		Students could be given samples of ionic compounds and asked to investigate their properties.
	3.3 Covalent bonding *(8 periods)*	
• Define covalent bonding • Describe the formation of a covalent bond	• Formation of Covalent bond	Students should know that covalent bonds are formed when atoms share pairs of electrons. Students should understand that the bond arises due to the electrostatic attraction between the negative electrons and the positively charged nuclei of the two atoms. Students should appreciate that covalent bonds are generally formed between atoms of non-metals. Students should be able to show how a covalent bond is formed by the sharing of a pair of electrons. One electron is donated by each non-metal atom. The shared pair of electrons is considered to exist in the outer orbital of both of the atoms.
• Draw Lewis structures or electron-dot formulas of simple covalent molecules	• Lewis formula of covalent molecules.	Students should represent the formation of a covalent molecule Lewis structure Symbols or diagrams can be used to illustrate it. Eg. H^x + Cl 

Competencies	***Contents***	***Suggested Activities***
• Give examples of different types of covalent molecules		Students should appreciate that the gaseous elements in Groups 5, 6 and 7 do not exist as atoms but as molecules in which the atoms are joined by covalent bonds e.g. N_2, O_2, Cl_2. Students should practice drawing Lewis structures for common covalent compounds. These could include hydrogen chloride, water, ammonia and methane. Students should appreciate that in some covalent compounds, atoms are joined by double and triple covalent bonds: • C-C e.g. ethane • C=C e.g. ethene • C≡C e.g. ethyne There are also double and triple bonds between atoms in some molecules of gases e.g. • N≡N in nitrogen • O=O in oxygen Students should practice drawing Lewis structures for common covalent compounds in which there are double and triple bonds. Students could research other covalent compounds in which there are double and triple bonds e.g. C=S, C≡N.
• Make models of covalent molecules to show single, double and triple bonds using sticks and balls or locally available materials.		Students should make models to show covalent compounds in which there are single bonds, double bonds and triple bonds, using cocktail sticks and small polystyrene balls or locally available materials.
• Explain polarity in covalent molecules	• Polarity in covalent molecules.	Students should appreciate that the sharing of the pair of electrons in a covalent bond depends on the attraction which each atom has for the electrons. This is related to the electro negativity of the atom. Students should discuss the sharing of the pair of electrons in molecules between two atoms of the same element, or two atoms of similar electro negativity: • the electrons are attracted equally by the two atoms • the electron pair is shared equally Students should discuss the sharing of the pair of electrons in molecules between two atoms which are significantly different in electro negativity e.g. hydrogen and chlorine in hydrogen chloride. • the chlorine atom is more electronegative than the hydrogen atom • the chlorine atom attracts the electrons in the covalent bond more strongly • the pair of electrons in the covalent bond are positioned closer to the chlorine atom than the hydrogen atom

Competencies	***Contents***	***Suggested Activities***
		• the chlorine end of the bond is slightly negative relative to the hydrogen end • the relative charge is shown using the symbols δ^+ and δ^- • the bond is polarised and shown as $H^{\delta+}$ $Cl^{\delta-}$
• Distinguish between polar and non polar covalent molecules		Students should be able to identify polar bonds in covalent compounds and mark their polarity. Students should discuss the structure or different covalent bonds and determine whether they are non-polar or polar, if if they are polar, to what extent. Students could discuss how polarity affects the properties of covalent compounds.
• Define coordinate covalent (dative) bond	• Coordinate covalent bond (dative bond)	Students should appreciate that another type of covalent bond exists in which both of the electrons which form a covalent bond between atoms is donated by one of the atoms. This is a coordinate covalent bond or dative bond.
• Illustrate the formation of coordinate covalent bond using suitable examples		This could include the formation of ammonium ion(NH^+_4) from hydrogen ion (H^+) and a molecule of ammonia(NH_3)  Students should be able to show the formation of coordinate bonds in some familiar substances including: • hydronium ion, H_3O^+ • carbon monoxide, CO Students should understand that a non-bonding pair of electrons is referred to as a lone pair. Students should appreciate the role of lone pairs of electrons in coordinate bonding. Students could name the covalent compounds and draw Lewis structures to show the covalent bonding.
• Explain the general properties of covalent compounds	• General properties of covalent compounds	Students should discuss general properties of covalent compounds including: • liquids or gases, and some are solids. • low melting points and boiling points • do not conduct an electric current when molten or in aqueous solution • solubility in non-polar solvents • insolubility in polar solvents like water
• Investigate the properties of given samples of covalent compounds.		

Competencies	***Contents***	***Suggested Activities***
		Students could be given samples of covalent compounds and asked to investigate their properties.
	3.4 Metallic bonding *(1 period)*	
• Explain the formation of metallic bond	• Formation of metallic bond	Students should appreciate that metals consist of a matrix of positively charged ions in a sea of delocalised electrons. Students could discuss why metals are able to lose valence electrons so easily to give this structure. Students should know that metallic bonding is the result of electrostatic attraction between the positively charged metal ions and the negatively charged delocalised electrons. Students could discuss how the strength of the metallic bond is related to the atomic radius of the metal atom and the number of valence electrons which are delocalised. Metallic bonding is strongest where metals have small atomic radii and lose a large number of valence electrons.
• Explain the electrical and thermal conductivity of metals in relation to metallic bonding.	• Properties of metallic bond	Students should appreciate that the ability of metals to conduct both heat and electricity is related to the movement of delocalised electrons. These are able to transfer energy and charge through the metal. Students could discuss how some of the other properties of metals, such as malleability and ductility are related to metallic bonding. Students could demonstrate the difference in strength of metals using locally available metals.
• Make a model to demonstrate metallic bonding.		Students could make a model to demonstrate metallic bonding using small and large balls made of local materials.
	3.5. Intermolecular forces *(4 periods)*	Students should know that intermolecular force is a force of attraction between molecules in a substance.
• Define inter molecular force • Explain hydrogen bonding	• Hydrogen bonding	Students should appreciate that: • a hydrogen bond is an intermolecular force of attraction between a highly polar hydrogen atom on one molecule and a highly electronegative atom on another • hydrogen bonding is the strongest of the intermolecular forces Students should appreciate that: • the O-H bonds in water are polar due to the high electro negativity of oxygen • the bonds can be presented as $H^{\delta+}$- 0^{8-} • water molecules are attracted to each other $H^{\delta+}$- 0^{8-} ... H^{8+} - 0^{8-} • this attraction gives rise to hydrogen bonding

Competencies	***Contents***	***Suggested Activities***
• Explain the effects of hydrogen bond on the properties of substances		Students should discuss the effects of hydrogen bonding on the properties of water including: • high melting point and boiling point • surface tension • formation of a meniscus • capillarity • coolant (in radiators) Students could discuss the effects of hydrogen bonding on other compounds.
• Describe Vander waals forces	• Van der Waals' forces	Students should appreciate that here are always forces of attraction between covalent molecules even if they are polar or not. Van der Waals' forces are weak electrostatic forces that bind both polar and non-polar molecules. They may be: • dipole-dipole forces • induced dipole-induced dipole or dispersion forces
• Explain dipole-dipole force • Give examples of molecules with dipole-dipole forces.	- dipole-dipole forces	Students should understand that a dipole which is two points of equal but opposite charges, separated by a distance. Students should appreciate that diploes exist in polar molecules and oppositely charged ends of dipoles on different molecules are attracted to each other. Students could be asked to give examples of molecules with dipole-dipole forces. This could include interaction between similar molecules like HCl or different molecules like HCl and water.
• Explain dispersion forces • Give examples of molecules in which the dispersion force is important.	- dispersion forces	Students should understand that dispersion forces occur in non-polar molecules. They are the result of temporary fluctuations in electron density. Within any bond, a slight temporary movement of bonding electrons towards one or other of the atoms will cause an induced dipole. This will result in molecules being attracted to each other. Students could be asked to give examples of molecules which contain dispersion forces. This could include molecules like neon(Ne) and methane(CH_4). Students should also recognize that dispersion forces exist in all types of molecules.
• Compare and contrast the three types of inter molecular forces.		Students should identify molecules where there are dipole forces and where there are dispersion forces. Students should describe the similarities and the differences among the hydrogen bond, dipole-dipole force and dispersion (London) force.

Assessment

The teacher should assess each student's work continuously over the whole unit and compare it with the following description, based on the Competencies, to determine whether the student has achieved the minimum required level.

Students at minimum requirement level

Students working at the minimum requirement level will be able to: Define chemical bonding and explain why it is formed between the atoms; explain the ions, ionic bonding and its formation; give examples of simple ionic compounds and draw their Lewis structures, explain the general properties of ionic compounds; investigate the properties of a given sample of ionic compounds; define covalent bonding and describe its formation; give examples of different types of covalent molecules and draw their Lewis structures; make models of covalent compounds; explain polarity of covalent compounds; distinguish between polar and non-polar covalent molecules define coordinate covalent bond and illustrate its formation; investigate properties of a given sample of covalent compounds, explain metallic bonding; explain electrical and thermal conductivity of metals; make model to demonstrate metallic bond, define intermolecular force; explain hydrogen bond and its effect on the properties of substances; describe Vander waals forces, explain dipole-dipole forces and give examples of molecules with dipole - dipole forces, explain dispersion forces and give examples of molecules with dispersion forces and compare and contrast the three types of intermolecular forces.

Students above minimum requirement level

Students working above the minimum requirement level should be praised and their achievements recognized. They should be encouraged to continue working hard and not become complacent.

Students below minimum requirement level

Students working below the minimum requirement level will require extra help if they are to catch up with the rest of the class. They should be given extra attention in class and additional lesson time during breaks or at the end of the day.

Unit 4: Chemical reactions and stoichiometry *(37 periods)*

Unit outcomes: Students will be able to:

- Understand fundamental laws of chemical reactions and know how they are applied.
- Develop skills in writing and balancing chemical equations.
- Understand energy changes in chemical reactions.
- Know types of chemical reactions.
- Develop skills in solving problems based on chemical equations (mass - mass, volume - volume and mass - volume problems).
- Develop skills in determining the limiting reactant, theoretical yield, actual yield and percentage yield.
- Understand oxidation - reduction reactions and analyse redox reactions by specifying the oxidizing agent, the reducing agent, the substance reduced or oxidized.
- Understand the rate of a chemical reaction, the state of a chemical equilibrium and factors affecting them.
- Demonstrate scientific inquiry skills: observing, inferring, predicting, classifying comparing and contrasting, communicating, measuring, asking questions designing experiments, interpreting data, drawing conclusions, applying concepts, relating cause and effect and problem - solving.

Competencies	***Contents***	***Suggested Activities***
Students will be able to:	**4. Chemical reactions and stoichiometry**	
• Define chemical reaction • Give some examples of chemical reactions	**4.1 Introduction** *(1 period)*	Students should appreciate that a chemical reaction is a change that takes place when one or more substances, called reactants, react alone or with each other to produce one or more new substances, called products. Reactants → Products Students should give different examples of chemical reactions. The students could also discuss the examples of changes brought about by the chemical reactions. These could include: • heating sugar • rusting of iron • Fermentation • Souring tella • Digestion of food
• State the law of conservation of mass and illustrate using examples	**4.2 Fundamental laws of chemical reactions** *(2 periods)* • The law of conservation of mass	Students should appreciate that in all types of chemical reactions mass is neither created nor destroyed. Students should be able to quote the law of conservation of mass: 'Matter cannot be created nor destroyed in a chemical reaction' Students could carry out an experiment to prove the law of conservation of mass. For example using the reaction between silver nitrate solution and dilute hydrochloric acid: $HCl(aq) + AgNO_3(aq) \rightarrow AgCl(s) + HNO_3$

Competencies	*Contents*	*Suggested Activities*
• Demonstrate the law of conservation of mass using simple experiments		• Place dilute hydrochloric acid in a conical flask to a depth of about 1 cm • Tie a thread of cotton around the top of a test tube • Half fill the test tube with silver nitrate solution • Place the test tube inside the conical flask so that it is held on a slant by the thread and place a bung in the top of the flask to hold the thread in place • Weigh the conical flask and contents • Tilt the flask so the silver nitrate solution pours into the dilute hydrochloric acid and a white precipitate of silver chloride is produced • Reweigh the conical flask and contents From their experiment students should show that the mass of the products is equal to the mass of the reactants. Students could also use the reaction between barium nitrate and sodium sulphate: $Ba(NO_3)_2(aq) + Na_2SO_4(aq) \rightarrow BaSO_4(s) + 2NaNO_3(aq)$ • Make up solutions containing 2.61 g of barium nitrate and 1.42 g of sodium sulphate • Mix the solutions • Filter off the insoluble barium sulphate and dry it (2.33 g) • Evaporate the water from the filtrate and weigh the residue of sodium nitrate (1.70 g) Students should find that, within the limits of experimental error, the mass of the products equals the mass of the reactants.
• State the law of definite proportion and illustrate using examples • Demonstrate the law of definite proportion using a simple experiment. • State the law of multiple proportion and illustrate using examples.	• The law of definite (constant) proportion • The law of multiple proportion	Students should be able to quote the law of constant or definite proportion: 'The proportion by mass of each element in a pure compound is always the same however the compound is made' Students could carry out an experiment to prove the law of definite proportion. For example: • Make copper(II) oxide by heating copper powder • Make copper(II) oxide by the thermal decomposition of copper(II) carbonate • Take 1 g of each sample of copper(II) oxide • Reduce each sample of copper(II) oxide by heating in a stream of hydrogen • Weigh the copper that remains in each case Students should find that the mass of copper is the same in each case therefore the proportions of copper and oxygen in each sample of copper (II) oxide is the same.
• Describe the conventions used to write chemical equation	**4.3 Chemical equations** *(3 periods)* • Writing chemical equation	Students should appreciate that a chemical equation is a means of describing a chemical reaction: • Qualitatively – by identifying the reactants and products • Quantitatively – by identifying the relative proportion of each reactant and product

Competencies	*Contents*	*Suggested Activities*
		Students should use the following steps to write an equation: 1. Write a word equation 2. Substitute the words by symbols and formulas 3. Balance the equation so that there are equal numbers of atoms of each type of element on each side of the equation
• Balance chemical equations using inspection method • Balance chemical equations using the Least Common Multiple (LCM) method	• Balancing chemical equation	Students should be able to balance an equation: • by inspection method • by the least common multiple method Students should have plenty of practice of writing chemical equations. As far as possible these should relate to reactions which are familiar to students. Examples should increase in difficulty as students become more able.
	4.4 Energy changes in chemical reactions *(3 periods)*	
• Explain energy changes in chemical reactions		Students should appreciate that changes in energy take place during a chemical reaction. Students should understand that enthalpy is a measure of the internal energy of a substance. It is represented by the symbol *H*. The energy change that takes place during a chemical reaction are the result of a
• Define endothermic reaction • Describe endothermic reaction	• Endothermic reaction	Students should know that in an endothermic reaction: • Heat is taken in from the surroundings • The internal energy of the reactants is less than the internal energy of the products • There is a rise in enthalpy therefore the value of ΔH is positive change in enthalpy and is represented as ΔH.
• Define exothermic reaction • Describe exothermic reaction	• Exothermic reaction	Students should know that in an exothermic reaction: • Heat is given out to the surroundings • The internal energy of the reactants is more than the internal energy of the products • There is a fall in enthalpy therefore the value of ΔH is negative Students should appreciate that most of the reactions they will see and carry out in the laboratory will be exothermic.
• Illustrate endothermic and exothermic reactions using diagrams	• Energy diagrams for endothermic and exothermic reactions	Students should be able to draw energy level diagrams to represent exothermic and endothermic reactions. These have the general form:

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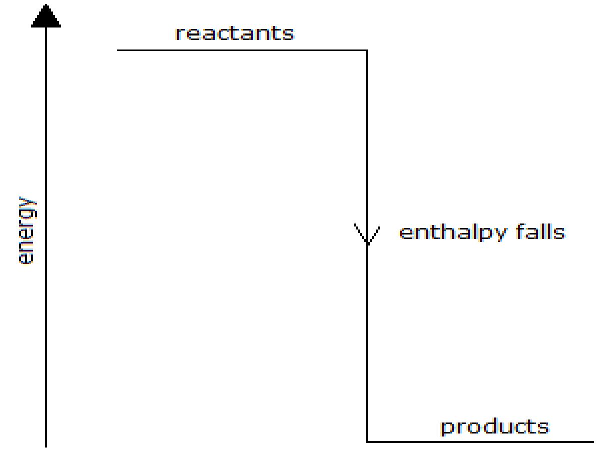
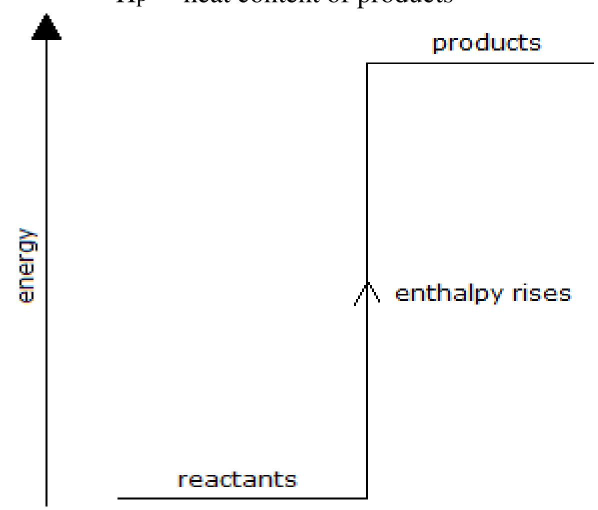
<tr><th>Competencies</th><th>Contents</th><th>Suggested Activities</th></tr>
<tr><td></td><td></td><td>


exothermic reaction
$H_r > H_P$ ∴ ΔH is negative
($\Delta H < 0$)
H_r - heat content of reactants
H_P - heat content of products


endothermic reaction
$H_P > H_r$ ∴ ΔH is positive ($\Delta H > O$)
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Competencies	***Contents***	***Suggested Activities***
• Conduct simple experiment to demonstrate exothermic and endothermic reactions		Students should carry out some exothermic and endothermic reactions and understand that heat is given out or taken in either by direct observation or by measuring temperature rise with a thermometer. These could include: Exothermic • Burning gas • Adding water to anhydrous copper(II) sulphate • Adding a few drops of concentrated sulphuric acid to water • Mixing dilute hydrochloric acid and sodium hydroxide solution • Thermite process Endothermic • Adding solid ammonium nitrate to water • Adding solid potassium nitrate to water
• Describe the importance of chemical changes in the production of new substances and energy	• Importance of chemical changes	Students should appreciate the importance of the energy released in some exothermic reactions. These should include: • Combustion of fuels • Oxidation of glucose during cell respiration
• List the four types of chemical reactions	**4.5 Types of chemical reactions** *(3 periods)*	Students should appreciate that chemical reactions can be classified into types or groups and that all of the reactions within a group have certain similarities.
• Define combination reaction and give examples • Conduct some experiments on combination reactions in groups.	• Combination	Students should know that combination or synthesis reactions involve the reaction of one element with another to form a compound. Students should discuss examples of combination reactions. These could include: • Sodium + chlorine = sodium chloride • Carbon + oxygen = carbon dioxide • Copper + oxygen = copper(II) oxide • Iron + sulphur = iron sulphide • Hydrogen + oxygen = water • Magnesium + oxygen = magnesium oxide Students should discuss the general features of a combination reaction. Students could carry out some combination reactions.

Competencies	*Contents*	*Suggested Activities*
• Define decomposition reaction and give examples • Conduct some experiments on decomposition reactions in groups.	• Decomposition	Students should know that decomposition reactions involve a reactant breaking down to produce two or more products. Students should discuss examples of decomposition reactions. These could include: • Thermal decomposition of Group 1 nitrates to nitrites and oxygen • Thermal decomposition of Group 2 and Transition Metal nitrates to oxides , oxygen and nitrogen dioxide • Thermal decomposition of Group 2 and Transition Metal carbonates to oxides and carbon dioxide Students should discuss the general features of a decomposition reaction. Students could carry out some decomposition reactions.
• Define single displacement reaction and give examples • Conduct some experiments on simple displacement reactions in groups.	• Single displacement	Students should know that single displacement reactions involve a more reactive element displacing a less reactive element from a compound. Students should discuss examples of single displacement reactions. These could include: • Reactive metals with water – displacing hydrogen e.g. sodium + water • Reactive metals with dilute acids – displacing hydrogen e.g. zinc + dilute hydrochloric acid • Metal – metal ion reactions – atoms of a more reactive element displace ions of a less reactive metal from solution e.g. iron + copper(II) sulphate solution • Halogen – halide ion reactions – atoms of a more reactive halide displace ions of a less reactive halide from solution e.g. chloride + potassium bromide solution. Students should discuss the general features of a single displacement reaction. Students could carry out some single displacement reactions.
• Define double decomposition reaction and give examples. • Conduct some experiments on double displacement reactions in groups.	• Double decomposition	Students should know that double decomposition reactions involving reacting solutions of two soluble compounds to form two products, one of which is soluble and one of which is not. The products are separated by filtration. Students should discuss examples of double decomposition reactions. These could include: • Reactions to precipitate silver halides • Reactions to precipitate lead halides • Reactions to precipitate barium sulphate Students should discuss the general features of a double decomposition reaction. Students could carry out some double displacement reactions.
• Deduce mole ratios from balanced chemical equations	**4.6 Stoichiometry** *(10 periods)* • Molar ratios in balanced chemical equations.	The student should revise the terms mole, motor mass and actual mass and related problems studied in grade 8 unit 5. Students should appreciate that stoichiometry is the study of the different amounts of substances which react to give new substances.

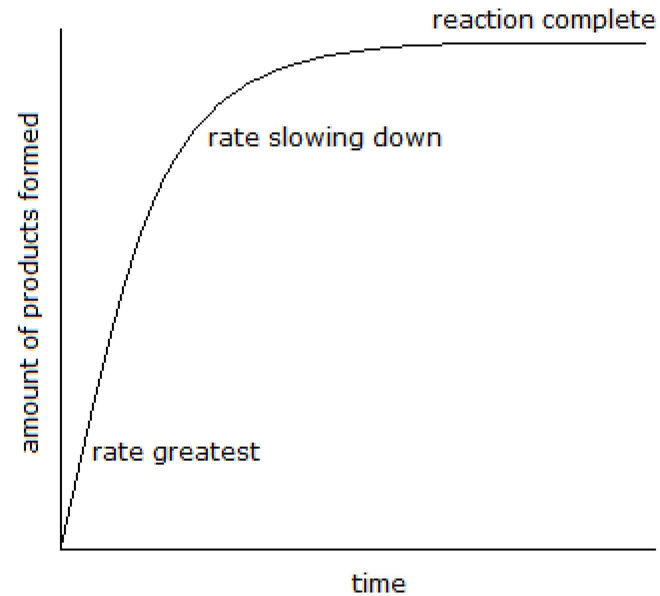
Competencies	***Contents***	***Suggested Activities***
		Students should appreciate the significance of the different numbers of reactant and product particles in a balanced chemical equation. Students should appreciate that a balanced chemical equation indicates the molar quantities of reactants and products involved in a reaction. Students should be able to interpret a balanced chemical equation in terms of ratios of moles of the reactants and products. Students should interpret a variety of balanced equations in this way.
• Solve mass-mass problems based on the given chemical equation	• Mass-mass relationships.	Students should be able to use balanced chemical equations to deduce the ratio of reacting masses and use this to calculate the actual masses. For example to find the mass of AgBr produced from 1 gm of KBr: $KBr(aq) + AgNO_3(aq) \rightarrow KNO_3(aq) + AgBr(s)$ 1 mole 1 mole 1 mole 1 mole 119 g 170 g 101 g 188 g • 119 g of KBr yields 188 g of AgBr • therefore 1.00 g of KBr yields $\frac{1.00 \times 188}{119}$ = 1.58 g of AgBr
• Define molar volume		Students should carry out a number of similar calculations on reacting masses using equations with different mole ratios.
• State Avogadro's principle		Students should appreciate that one mole of gas particles occupies a volume of 22.4 litre at standard temperature and pressure (STP) 0 °C and 101.325 Kpa.
• Solve volume-volume problems based on the given chemical equation	• Volume-volume relationships.	Students should be aware of Avogadro's principle which states that: 'equal volumes of different gases under the same conditions of temperature and pressure contain equal numbers of particles; Students should be able to use balanced chemical equations to deduce the ratio of reacting volumes and use this to calculate the actual volumes. For example to find the volume of HCl produced from 100ml of H_2: $H_2(g) + Cl_2(g) \rightarrow 2HCl(g)$ 1 mole 1 mole 2 moles 22.4 L 22.4 L 44.8 L • 22.4 L of H_2 yields 44.8 L of HCl • 100 ml of H_2 yields $\frac{0.1 \times 44.8}{22.4}$ = 0.2 L = 200 ml of HCl Students should carry out a number of similar calculations on reacting volumes using equations with different mole ratios.

Competencies	***Contents***	***Suggested Activities***
• Solve mass-volume problems based on the given chemical equation	• Mass - volume relationships.	Students should combine their knowledge of reacting masses and molar volume to carry out calculations involving both mass and volume. For example to find the volume of CO_2 produced from 0.50gm of $CaCO_3$ $CaCO_3(s) \rightarrow CaO(s) + CO_2(g)$ 1 mole 1 mole 1 mole 100 g 56 g 44 g or 22.4 L • 100 g of $CaCO_3$ yields 56 g of CaO • 0.50 g of $CaCO_3$ yields $\frac{0.5 \times 56}{100} = 0.28$ g • 100 g of $CaCO_3$ yields 22.4 L of CO_2 • 0.50 g of $CaCO_3$ yields $\frac{0.5 \times 22.4}{100} = 0.112$ L
• Define limiting and excess reactants • Determine limiting and excess reactants of a given chemical reaction • Show that the amount of product of a chemical reaction is based on the limiting reactant	• Limiting and excess reactants.	Students should carry out a number of similar calculations on reacting masses and volumes using equations with different mole ratios. Students should appreciate that in a chemical reaction involving two reactants, the reaction will stop when all of one reactant has been used up no matter how much of the second reactant remains. Students should understand that when two reactants are not in the mole ratio in which they react then one will be the limiting reactant and the other will be in excess. Students should carry out calculations on chemical reactions in which there is a limiting reactant. For example: $Zn(s) + 2HCl(aq) \rightarrow ZnCl_2(aq) + H_2(g)$ 1 mole 2 moles 1 mole 1 mole 5.00 g of zinc is reacted with 3.65g of HCl • 5.00 g of zinc is $\frac{5.00}{65.4} = 0.076$ mol • 3.65g of hydrochloric acid contains 3.65/ 36.5 = 0.1 mol. • The maximum amount of zinc that will react with 0.1 mol dilute hydrochloric acid is 0.050 mol • The dilute hydrochloric acid is the limiting reactant • The zinc is in excess • The amount of zinc that will remain at the end of the reaction is 0.076 – 0.050 = 0.026 mol which is 65.4 x 0.026 = 1.70 g • The amount of zinc chloride formed will be 0.050 mol which is (65.4 + 2 x 35.5) x 0.05 = 6.82 g • The amount of hydrogen gas formed will be 0.050 mol which is 22.4 x 0.050 = 1.12 L at room temperature and standard pressure

Competencies	Contents	Suggested Activities
		Students should carry out a number of similar calculations, identifying the limiting reactant and finding the maximum amount of products that can be formed.
• Define the terms theoretical yield, actual yield and percentage yield	• Theoretical, actual and percentage yields.	Students should appreciate that: • given the equation for a chemical reaction and the quantities of reactants used, it is possible to calculate the yield of products that should be obtained – the theoretical yield • not all of the reactants in a chemical reaction may go to form a desired product; some may be used up in unwanted side reactions. The actual amount of product obtained is – the actual yield • the success of a reaction can be assessed by comparing the ratio of the actual yield to the theoretical yield to give – the percentage yield
• Calculate the percentage yield of a chemical reaction from given information		Students should be able to calculate the percentage yield of a chemical reaction from a given data using the equation: $\text{Percentage yield} = \dfrac{\text{actual yield}}{\text{theoretical yield}} \times 100\%$ Students should appreciate that: • percentage yield will be between 0 – 100% • the higher the percentage yield the more successful the reaction
	4.7 Oxidation and reduction reactions *(5 periods)*	
• Define redox reactions • Define the terms oxidation and reduction in terms of electron transfer	• oxidation • reduction	Students should know that redox reaction is a reaction that involves transfer of electrons. Students should understand that the term redox' is used to describe reduction and oxidation reactions and that these reactions occur together. Students should be able to define oxidation as a loss of electrons and reduction as a gain of electrons. They should also understand that in a redox reaction: • the oxidised species loses one or more electrons • the reduced species gains one or more electrons
• Define oxidation number (oxidation state)	• oxidation number	Students should know that the oxidation number or oxidation state: • is the charge that the atom carries in its compounds • refers to a single atom of the element • has both sign and numerical value
• State oxidation number rules.		Students should state oxidation number rules. Students should be able to deduce the oxidation number of elements in compounds from their formulas

Competencies	***Contents***	***Suggested Activities***
• Determine the oxidation number of an element in a given formula • Determine the oxidation number of an element in a given formula		Students should do simple examples e.g. $MgCl_2$ • oxidation state of Mg = +2 • oxidation state of Cl = -1 And more complex examples where an atom is part of an ion e.g. SO_4^{2-} • oxidation state of O = -2 • total oxidation state of O=4 x -2 = -8 • overall charge on the ion = -2 • oxidation state of S = -2 – (-8) = +6
• Describe the oxidizing and reducing agents	• Oxidizing and reducing agents.	Students should appreciate that an oxidising agent is one which brings about oxidation but is itself reduced. Students should discuss examples of oxidising agents which they may have already met in chemical reactions such as: • chlorine • potassium manganate(VII) • potassium chromate(VI) • potassium dichromate(VI) • sodium chlorate(V) • manganese(IV) oxide Students should appreciate that a reducing agent is one which bring about reduction but is itself oxidised. Students should discuss examples of reducing agents which they may already have met in chemical reactions such as: • carbon • carbon monoxide • hydrogen • iron(II) salts • acidified iodide ions • metals such as sodium, magnesium and zinc • sodium thiosulphate • sodium sulphite
• Analyze a given redox reaction by specifying the substance reduced the substance oxidized, the oxidizing and reducing agents	• Analysing redox reactions	Students should be able to give examples of redox reaction and identify the oxidised and reduced species. They could start with simple reactions e.g. $Cu^{2+}(aq) + Fe(s) \rightarrow Cu(s) + Fe^{2+}(aq)$ • oxidised species is Fe • reduced species is Cu^{2+}

Competencies	***Contents***	***Suggested Activities***
		and progress to more difficult examples e.g. e.g. $SO_3^{2-}(aq) + H^+ (aq) + MnO_4^- (aq) \rightarrow SO_4^{2-}(aq) + H_2O(l) + Mn^{2+} (aq)$ • oxidised species is SO_3^{2-} (S oxidised from +4 to +6) • reduced species is MnO_4^- (Mn reduced from +7 to +2)
• Distinguish between redox and non-redox reactions	- Differences between redox and non - redox reactions	Students should appreciate that a chemical reaction can only be described as a redox reaction if one species is reduced and another species is oxidised. There are many reactions in which oxidation and reduction do not occur e.g. in double decomposition reactions: $Pb(NO_3)_2(aq) + 2KI(aq) \rightarrow PbI_2(s) + 2KNO_3(aq)$ Students should find the oxidation states of the metal ions in both the reactants and products and satisfy themselves that no oxidation or reduction has taken place.
• Define a reaction rate	**4.8 Rate of chemical reaction and chemical equilibrium** *(10 periods)* • Reaction rate	Students should understand that the rate of a reaction is the rate at which reactants are converted to products. Students could discuss different ways in which the rate of a reaction can be monitored. These could include: • Change of colour • Volume of gas evolved • Amount of precipitate formed • Loss or gain of mass
• Describe a reaction rate using graphs		Students should be able to draw a graph showing how the rate of a chemical reaction changes over time and relate the rate to the slope of the graph.

Competencies	***Contents***	***Suggested Activities***
		
• Carry out an experiment to illustrate the relative rate of reactions		Students should carry out experiments to study how the rate changes during a chemical reaction. These could include: • Metal + dilute acid e.g. zinc + dilute sulphuric acid • Metal carbonate + dilute acid e.g. calcium carbonate + dilute hydrochloric acid In both of these types of reaction the rate can be monitored by: • Measuring the loss of mass over time • Measuring the volume of gas produced over time
• List the preconditions for a chemical reaction to occur • Explain how collision, activation energy and proper orientation of reactants cause a chemical reaction to occur	• Preconditions for a chemical reaction	Students should be able to list the preconditions for a chemical reaction to take occur – collision, activation energy and proper orientation

Competencies	***Contents***	***Suggested Activities***
• List factors that affect rate of chemical reaction	• Factors affecting rate of chemical reaction	Students should appreciate that in order for a chemical reaction to take place the reacting particles must collide with each other with sufficient force – an effective collision. Students should discuss how providing reacting particles with activation energy: • Increases the kinetic energy of the particles • Increases the chances that collisions between particles will be effective Students should appreciate that the rate of a chemical reaction depends on: • Nature of reactants • Temperature • Concentration or pressure • Surface area • Use of a catalyst
• Explain the effect of changes in temperature, concentration or pressure and surface area on the rate of a chemical reaction • Explain the effect of catalysts on the rate of chemical reaction		Students should discuss these factors in terms of collision theory. This could include: • Temperature – an increase in temperature increases the kinetic energy of the reacting particles so there is a greater chance of collisions being effective • Concentration or pressure – increasing the concentration (of reacting solutions) or pressure (of reacting gases) increases the chances of particles colliding with each other as there will be more particles per unit volume • Surface area – increasing the surface area of a solid reactant increases the chances of collisions with the other reactant • Use of a catalyst – a catalyst reduces the activation energy so particles collisions can be effective at lower temperatures
• Carry out an activity on how the factors affect the rate of chemical reaction		Students could investigate the effects of each of these factors experimentally e.g. • Temperature – metal-acid reaction using acid of different temperatures • Concentration – metal-acid reaction or metal carbonate-acid reaction using acid of different concentrations • Surface area – zinc-acid reaction using granulated zinc and zinc powder; calcium carbonate-dilute hydrochloric acid using chippings and powdered calcium carbonate • Catalyst – decomposition of hydrogen peroxide using manganese(IV) oxide
• Define the terms reversible reaction and irreversible reaction.	• Reversible and irreversible reactions	Students should appreciate that some chemical reactions are difficult or impossible to reverse whereas others can be reversed. Students should discuss examples of irreversible and reversible reactions.
• Define chemical equilibrium	• Chemical equilibrium	Students should understand that where a reaction is reversible, as soon as any product is formed, the reverse reaction will begin unless the products are removed.

Competencies	***Contents***	***Suggested Activities***
• Describe the characteristics of chemical equilibrium		Forward reaction Reactants $\rightleftharpoons$ Products Reverse or backward reaction Students should appreciate that: • The symbol '$\rightleftharpoons$' is used to show that a reaction is reversible • At some point in time the rate of the forward reaction becomes equal to the rate of the backward reaction • At this point the concentration of each reactant and each product remains constant • The reaction is said to be in equilibrium • This is described as dynamic equilibrium since the reactions are still taking place even though the concentrations of the reactants and products remain constant. Students should be aware of the conventional use of square brackets [] to show the concentration of a species in a reversible reaction.
• Write the expression for equilibrium constant of a reversible reaction	• Equilibrium constant	Students should know that for a reversible reaction represented by: $aA + bB \rightleftharpoons cC + dD$ • The ratio of the equilibrium concentrations of products raised to the power of their coefficient in a balanced equation is constant at a given temperature • This is called the equilibrium constant and has the symbol K_c • $K_c = \frac{[C]^c\,[D]^d}{[A]^a\,[B]^b}$ Students should carry out calculations using given data to find the equilibrium constant for reaction at equilibrium.
• State Le Châtelier's principle	• Factors that affect chemical equilibrium	Students should be able to state Le Châtelier's principle: 'If a condition is changed, the position of equilibrium will move in such a way as to oppose the change and restore the original equilibrium conditions'
• Use Le Châtelier's principle to explain the effect of changes in temperature, pressure and concentration of reaction at equilibrium		Students should discuss how this principle accounts for the effects of changes in temperature and pressure (concentration) in reactions including: • $N_2O_4(g) \rightleftharpoons 2NO_2(g)$ • $H_2(g) + I_2(g) \rightleftharpoons 2HI(g)$ And in industrial processes including: • The Haber process $N_2(g) + 3H_2(g) \rightleftharpoons 2NH_3(g)$ • The Contact process $2SO_2(g) + O_2(g) \rightleftharpoons 2SO_3(g)$

Assessment

The teacher should assess each student's work continuously over the whole unit and compare it with the following description, based on the Competencies, to determine whether the student has achieved the minimum required level.

Students at minimum requirement level
Students working at the minimum requirement level will be able to: Define chemical reactions and give examples; state and demonstrate the law of conservation of mass and the law of definite proportions, write chemical equations and balance them using inspection and LCM methods; explain energy changes in chemical reactions, conduct simple experiments on endothermic and exothermic reactions; describe the importance of chemical changes; define combination, decomposition, single displacement, double displacement reactions and give examples for each; conduct experiments on combination; decomposition, simple displacement and double displacement reactions; define molar volume, limiting reactant, excess reactant, theoretical yield, actual yield and percentage yield, deduce molar ratios from balanced chemical equations and solve problems on mass - mass, volume - volume and mass - volume relationships, determine the limiting reactant of a chemical reaction and show that the amount of product is based on the limiting reactants; calculate the percentage yield of a chemical reaction; define redox reactions, oxidation ,reduction and oxidation number, state oxidation number rules and determine the oxidation number of an element; describe the oxidizing and the reducing agents, distinguish between redox and non -redox reactions; analyze a given redox reaction; define and describe a reaction rate; list and explain the preconditions for a chemical reaction to occur; list and explain factors that affect the rate of a chemical reaction; carryout an activity of factors that affect the rate of a reaction; define the terms reversible reactions, irreversible reactions and chemical equilibrium; describe the characteristics of chemical equilibrium and write an expression for its equilibrium constant, state Lechattlers principle and use it to explain the effect of the factors that affect chemical equilibrium.

Students above minimum requirement level
Students working above the minimum requirement level should be praised and their achievements recognized. They should be encouraged to continue working hard and not become complacent.

Students below minimum requirement level
Students working below the minimum requirement level will require extra help if they are to catch up with the rest of the class. They should be given extra attention in class and additional lesson time during breaks or at the end of the day.

Unit 5: The physical states of matter *(20 periods)*

Unit outcomes: Students will be able to:

- Understand the kinetic molecular theory and properties of the three physical states of matter.
- Know the behaviour of gases by using the variables volume, temperature, pressure and number of moles.
- Know terms like ideal gas, diffusion, evaporation, boiling, condensation, vapor pressure, boiling point, molar heat of vaporization, molar heat of condensation, melting, fusion, sublimation, melting point, freezing point, molar heat of fusion, molar heat of solidification.
- Understand gas laws.
- Develop skills in solving problems to which the gas laws apply.
- Perform activities to illustrate gas laws.
- Carryout experiments to determine the boiling points of liquids and the melting point of solids.
- Demonstrate an experiment to show phase changes.
- Demonstrate scientific inquiry skills: observing, predicting, comparing and contrasting, measuring, interpreting data, drawing conclusion, applying concepts and making generalizations.

Competencies	***Contents***	***Suggested Activities***
Students will be able to:	**5. The physical states of matter**	
• Name the three physical states of matter. • Give examples for each of the three physical states of matter	**5.1 Introduction** *(1 period)*	Students should know that there are three physical states of matter: solids, liquids and gases. Students should also be aware that, although it is seldom found in everyday life, there is the fourth state known as the plasma state. Students should give everyday examples of substances that occurin each of these states.
• State kinetic theory of matter	**5.2 Kinetic theory and properties of matter** *(3 periods)* • Kinetic theory of matter	Students should know the kinetic theory of matter: • All matter is composed of particles which are in constant motion • The particles possess kinetic energy (movement energy) and potential energy • The difference between the three states of matter are due to their energy contents and the motion of the particles
• Explain the properties of the three physical states of matter in terms of kinetic theory	• Properties of matter	Students should know that in a solid: • Particles are in fixed positions • Particles are able to vibrate but not translate Students should discuss how the kinetic theory explains the properties of solids e.g. fixed shape, can't be compressed Students should know that in a liquid: • The particles have more energy than in a solid • The particles are not in fixed positions but able to move • The particles are in continual constant motion • The particles are closely packed but less so than in a solid