Module 1: **Atoms and Reactions**

* + - 1.1.1 - **Atoms**

|  |  |
| --- | --- |
| **Atomic Structure** | |
| Modern development of the structure of the atom; the changing accepted view of the structure of the atom |  |
| Describe protons, neutrons and electrons in terms of relative charge and relative mass |  |
| Describe the distribution of mass and charge within an atom; |  |
| Describe the contribution of protons and neutrons to the nucleus of an atom, in terms of atomic (proton) number and mass (nucleon) number; |  |
| Deduce the numbers of protons, neutrons and electrons in an atom/ion |  |
| Explain the term *isotopes* as atoms of an element with different numbers of neutrons and different masses; |  |
| **Relative Masses** | |
| Define the terms *relative isotopic mass* and *relative atomic mass*, based on the 12C scale |  |
| Calculate the relative atomic mass of an element given the relative abundances of its isotopes |  |
| Use the terms *relative molecular mass* and *relative formula mass* and calculate values from relative atomic masses. |  |

Targets –

What went well -

* 1.1.2 – **Moles and equations acids**

|  |  |
| --- | --- |
| **The Mole** | |
| Explain amount of substance, the mole and avagadro’s constant |  |
| Define and use the term *molar mass* (units g mol–1) as the mass per mole of a substance |  |
| **Empirical And Molecular Formulae** | |
| Explain the term *empirical formula* as the simplest whole number ratio of atoms of each element present in a compound |  |
| Explain the term *molecular formula* as the actual number of atoms of each element in a molecule |  |
| Calculate empirical and molecular formulae, using composition by mass and percentage compositions |  |
| **Chemical Equations** | |
| Construct balanced chemical equations for reactions studied and for unfamiliar reactions given reactants and products |  |
| **Calculation of reacting masses, mole concentrations and volumes of gases** | |
| Carry out calculations, using amount of substance in mol, involving: mass, gas volume, and solution volume and concentration |  |
| Deduce stoichiometric relationships from calculations |  |
| Use the terms *concentrated* and *dilute* as qualitative descriptions for the concentration of a solution |  |

Targets –

What went well -

* 1.1.3 – **Acids**

|  |  |
| --- | --- |
| **Acids** | |
| Explain that an acid releases H+ ions in aqueous solution |  |
| State the formulae of the common acids: hydrochloric, sulfuric and nitric acids |  |
| State that common bases are metal oxides, metal hydroxides and ammonia |  |
| State that an alkali is a soluble base that releases OH– ions in aqueous solution |  |
| state the formulae of the common alkalis: sodium hydroxide, potassium hydroxide and aqueous ammonia |  |
| **Salts** | |
| Explain that a salt is produced when the H+ ion of an acid is replaced by a metal ion or NH4+ |  |
| Describe the reactions of an acid with carbonates, bases and alkalis, to form a salt |  |
| Explain that a base readily accepts H+ ions from an acid: eg OH– forming H2O; NH3 forming NH4+ |  |
| Explain the terms *anhydrous*, *hydrated* and *water of crystallisation* |  |
| Calculate the formula of a hydrated salt from given percentage composition, mass composition or experimental data |  |
| Perform acid–base titrations and carry out structured titrations |  |

Targets –

What went well -

* 1.1.4 – **Redox**

|  |  |
| --- | --- |
| **Oxidation Number** | |
| Apply rules for assigning oxidation number to atoms in elements, compounds and ions |  |
| Describe the terms *oxidation* and *reduction* in terms of:  electron transfer and changes in oxidation number |  |
| Use a Roman numeral to indicate the magnitude of the oxidation state of an element, when a name may be ambiguous, eg nitrate(III) and nitrate(V) |  |
| Write formulae using oxidation numbers |  |
| **Redox Reactions** | |
| Explain that metals generally form ions by losing electrons with an increase in oxidation number to form positive ions |  |
| Explain that non-metals generally react by gaining electrons with a decrease in oxidation number to form negative ions |  |
| Describe the redox reactions of metals with dilute hydrochloric and dilute sulfuric acids |  |
| Interpret and make predictions from redox equations in terms of oxidation numbers and electron loss/gain |  |

Targets –

What went well -

Module 2: **Electrons, Bonding and Structure**

**1.2.1 – Electron Strucuture**

|  |  |
| --- | --- |
| **Ionization Energies** | |
| Define the terms *first ionisation energy* and *successive ionisation energy* |  |
| Explain that ionisation energies are influenced by nuclear charge, electron shielding and the distance of the outermost electron from the nucleus |  |
| Predict from successive ionisation energies of an element using the number of electrons in each shell of an atom and the group number |  |
| **Electrons: electronic energy levels, shells,** (d) **sub-shells, atomic orbitals, electron configuration** | |
| State the number of electrons that can fill the first four shells | |
| Describe an orbital as a region that can hold up to two electrons, with opposite spins (PAULI) |  |
| Describe the shapes of s and p orbitals |  |
| State the number of: orbitals making up s-, p- and d-sub- shells, electrons that occupy s-, p- and d-sub- shells |  |
| Describe the relative energies of s-, p- and d- orbitals for the shells 1, 2, 3 and the 4s and 4p orbitals |  |
| Deduce the electronic configurations of atoms and ions up to the mass number of 36 |  |
| Classify the elements into s, p and d blocks |  |

Targets –

What went well -

**1.2.2 Bonding and Structure**

Targets –

What went well -

|  |  |
| --- | --- |
| **Ionic Bonding** | |
| Describe the term *ionic bonding* as electrostatic attraction between oppositely- charged ions |  |
| Construct ‘*dot-and-cross*’ diagrams, to describe ionic bonding |  |
| Predict ionic charge from the position of an element in the Periodic Table |  |
| State the formulae for the following ions: NO3, CO32–, SO42– and NH4+ |  |
| **Covalent bonding and dative covalent (coordinate) bonding** | |
| Describe the term *covalent bond* as a shared pair of electrons |  |
| Construct ‘*dot-and-cross*’ diagrams to describe: single covalent bonding; Multiple covalent bonding; dative covalent; and other molecules/ions |  |
| **The shapes of simple molecules and ions** | |
| Explain that the shape of a simple molecule is determined by repulsion between electron pairs surrounding a central atom |  |
| State that lone pairs of electrons repel more than bonded pairs (lone pair reduces angle by 2.5) |  |
| Explain the shapes of, and bond angles in, molecules and ions with up to six electron pairs (including lone pairs) surrounding a central atom  BF3 (trigonal planar); CH4 and NH4+ (tetrahedral); SF6 (octahedral); NH3(pyramidal); H2O (non Linear); CO2 (Linear) |  |
| **Electronegativity and bond polarity** | |
| Describe the term *electronegativity* as the ability of an atom to attract the bonding electrons in a covalent bond |  |
| Explain that a permanent dipole may arise when covalently-bonded atoms have different electronegativities, resulting in a polar bond |  |
| **Intermolecular forces** | |
| Describe intermolecular forces based on permanent dipoles, as in hydrogen chloride, and induced dipoles (van der Waals’ forces), as in the noble gases |  |
| Describe *hydrogen bonding*, including the role of a lone pair, between molecules containing –OH and –NH groups |  |
| Describe and explain the anomalous properties ofwater resulting from hydrogen bonding, eg: the density of ice compared with water, and its relatively high freezing point and boiling point |  |
| **Metallic bonding** | |
| Describe *metallic bonding* as the attraction of positive ions to delocalised electrons |  |
| **Bonding and physical properties** |  |
| Describe structures as: giant ionic lattices; giant covalent; giant metallic; simple molecular lattices |  |
| Describe, interpret and/or predict physical properties, including melting and boiling points, electrical conductivity and solubility in terms of structures of particles and bonding types |  |

**Module 3 – The Periodic Table**

**1.3.1 - Periodicity**

|  |  |
| --- | --- |
| **The Structure Of the Periodic Table** | |
| Describe the Periodic Table in terms of the arrangement of elements: by increasing atomic (proton) number, in periods showing repeating trends in physical and chemical properties, In groups having similar physical and chemical properties; |  |
| Describe *periodicity* in terms of a repeating pattern across different periods |  |
| Explain that atoms of elements in a group have similar outer shell electron configurations, resulting in similar properties |  |
| **Periodicity of physical properties of elements** | |
| Describe and explain the variation of the first ionisation energies of elements shown by: a general increase across a period, in terms of increasing nuclear charge, a decrease down a group in terms of increasing atomic radius and increasing electron shielding outweighing increasing nuclear charge; |  |
| Describe the variation of electron configurations, atomic radii melting/boiling points in Periods 2 & 3 |  |
| Explain variations in melting and boiling points in terms of structure and bonding |  |
| Interpret data on electron configurations, atomic radii, first ionisation energies, melting points and boiling points to demonstrate periodicity. |  |

Targets –

What went well -

**1.3.2 – Group 2**

|  |  |
| --- | --- |
| **Redox Reactions of group 2 metals** | |
| Describe the redox reactions of the Group 2 elements Mg - Ba with oxygen and water |  |
| Explain the trend in reactivity of Group 2 elements down the group due to the increasing ease of forming cations, in terms of atomic size, shielding and nuclear attraction; |  |
| **Reactions of group 2 compounds** | |
| Describe the action of water on oxides of elements in Group 2 and state the approximate pH of any resulting solution |  |
| Describe the thermal decomposition of the carbonates of elements in Group 2 and the trend in their ease of decomposition |  |
| Interpret and make predictions from the chemical and physical properties of Group 2 elements and compounds |  |
| Explain the use of Ca(OH)2 in agriculture to neutralise acid soils; the use of Mg(OH2) in some indigestion tablets as an antacid |  |

Targets –

What went well -

**1.3.3 – Group 7**

|  |  |
| --- | --- |
| **Characteristic physical properties** | |
| Explain, in terms of van der Waals’ forces, the trend in the boiling points of Cl2 Br2 and I2 |  |
| **Redox reactions and trends in reactivity of Group 7 elements and their compounds** | |
| Describe the redox reactions, including ionic equations, of the Group 7 elements Cl2, Br2 and I2 with other halide ions, in the presence of an organic solvent, to illustrate the relative reactivity of Group 7 elements |  |
| Explain the trend in reactivity of Group7 elements down the group from the decreasing ease of forming negative ions, in terms of atomic size, shielding and nuclear attraction |  |
| Describe the term *disproportionation* as a reaction in which an element is simultaneously oxidised and reduced, illustrated by: the reaction of chlorine with water as used in water purification, the reaction of chlorine with cold, dilute aqueous sodium hydroxide, as used to form bleach |  |
| Interpret and make predictions from the chemical and physical properties of the Group 7 elements and their compounds |  |
| Contrast the benefits of chlorine use in water treatment (killing bacteria) with associated risks (hazards of toxic chlorine gas and possible risks from formation of chlorinated hydrocarbons) |  |
| **Characteristic reactions of halide ions** | |
| Describe the precipitation reactions, including ionic equations, of the aqueous anions Cl–, Br– and I– with aqueous silver ions, followed by aqueous ammonia |  |
| Describe the use of the precipitation reactions above as a test for different halide ions |  |

Targets –

What went well -