|  |  |  |  |
| --- | --- | --- | --- |
| Each topic has a set of boxes which you can tick to show how well how well you know the topic. This is useful for revision. **Bold text** indicates Higher Level. |  |  |  |
| **6.1 Reaction Rates**  By the end of this section you should be able to | **Good** | **Fair** | **Poor** |
| define rate of reaction  define catalysis  monitor the rate of production of oxygen from hydrogen peroxide, using manganese dioxide as a catalyst  plot reaction rate graphs  interpret reaction rate graphs  distinguish between average and instantaneous rate  calculate instantaneous rate from graphs |  |  |  |
| **6.2 Factors Affecting Rates of Reaction**  By the end of this section you should be able | **Good** | **Fair** | **Poor** |
| explain what is meant by the nature of reactants    describe and explain how concentration, particle size, temperature, nature of reactants, and the presence of a catalyst effects the rate of reaction  describe how to investigate the effect of (i) particle size and (ii) catalysts on reaction rate  explain why dust explosions occur  identify two examples of catalysts produced by living cells (enzymes)  describe catalytic converters in terms of; nature of catalysts, reactions catalysed, environmental benefits and catalyst poisons  investigate the effects on the reaction rate of (i) concentration and (ii) temperature, using sodium thiosulfate solution and hydrochloric acid  describe and explain an experiment to show the oxidation of methanol (methyl alcohol) using a hot platinum or nichrome catalyst  define activation energy  describe and explain the influence of temperature change to changes in reaction rate  draw and interpret reaction profile diagrams  use reaction profile diagrams to explain the influence of catalyst on the rate of reaction  explain the mechanism of catalysis with reference to surface adsorption and intermediate formation theories of catalysis  describe and explain an experiment that demonstrates the oxidation of potassium sodium tartrate by hydrogen peroxide, catalysed by cobalt(II) salts |  |  |  |
| * 1. **The periodic table of elements.**   By the end of this section you should be able to | **Good** | **Fair** | **Poor** |
| describe the periodic table as a list of elements arranged so as to demonstrate trends in their physical and chemical properties  define the term element  associate the first 36 elements with their elemental symbols  distinguish between elements and compounds  state the principle resemblances of elements within each main group, in particular alkali metals, alkaline earth metals, halogens and noble gases  describe the reaction between water and lithium, sodium and potassium having seen the reaction demonstrated  describe by means of a chemical equation the reaction between water and lithium, sodium and potassium having seen the reaction demonstrated  outline the history of the idea of elements, including the contributions of the Greeks, Boyle, Davy and Moseley  outline the contributions of Mendeleev, **Dobereiner, Newlands and Moseley** to the structure of the modern periodic table  compare mendeleev’s periodic table with the modern periodic table  arrange elements in order of relative atromic mass and note differences with modern periodic table |  |  |  |
| **1.2 Atomic Structure**  By the end of this section you should be able | **Good** | **Fair** | **Poor** |
| outline the historical development of atomic theory (outline principles only,  mathematical treatment not required):  *Dalton*: atomic theory;  ***Crookes*: vacuum tubes, cathode rays;**  ***Stoney:* naming of the electron**;  *Thomson*: negative charge of the electron; **e/m for electrons (experimental details not required);**  ***Millikan*: magnitude of charge of electrons as shown by oil drop experiment (experimental details not required);**  *Rutherford***:** discovery of the nucleus as shown by the particle scattering experiment; **discovery of protons in nuclei of various atoms;**  *Bohr*: model of the atom;  ***Chadwick*: discovery of the neutron**.  recall that matter is composed of particles, which may be atoms, molecules or ions  define an atom  appreciate that atoms are minute particles  state the law of conservation of mass  describe, relative mass, relative charge and location of a proton, neutron, and electron in an atom  define atomic number (Z) mass number(*A*)  define relative atomic mass (*A*r) using the C12 scale  define isotope  describe the composition of isotopes using hydrogen and carbon as examples  describe how a mass spectrometer can be used to determine relative atomic mass  describe the principles on which the Mass Spectrometer is based  **explain the fundamental processes that occur in a mass spectrometer**  **calculate the approximate relative atomic masses from abundance of isotopes of given mass number** |  |  |  |
| **1.4 Electronic Structure of Atoms**  By the end of this section you should be able | **Good** | **Fair** | **Poor** |
| define and explain energy levels in atoms  describe the organization of particles in atoms of elements numbers 1-20  classify the first twenty elements in the periodic table on the basis of the number of outer electrons  list the numbers of electrons in each main energy level in atoms of elements nos. 1–20  **describe and explain the emission spectrum of the hydrogen atom using the Balmer series in the emission spectrum as an example**  **describe and explain the absorption spectrum**  use flame tests to provide evidence that energy is absorbed or released in discrete units when electrons move from one energy level to another  **explain how flame tests provide evidence that energy is absorbed or released in discrete units when electrons move from one energy level to another**  relate energy levels in atoms to everyday applications such as sodium street lights and fireworks  **discuss the uses of atomic absorption spectrometry (AAS) as an analytical technique**  **illustrate how line spectra provide evidence for energy levels**  **use a spectroscope or a spectrometer to view emission spectra of elements**  **define and explain energy sub-levels**  **state the Heisenberg uncertainty principle**  **state the dual wave-particle nature of the electron (mathematical treatment not required)**  **define and explain atomic orbitals**  **describe the shapes of s and p orbitals**  **build up the electronic structure of the first 36 elements**  **derive the electronic configurations of ions of s- and p block elements only**  **describe the arrangement of electrons in individual orbitals of p-block atoms**  define and explain atomic radius  explain the general trends in values of atomic radii (covalent radii only)   * down a group * across a period (main group elements only)   define and explain first ionisation energy  **explain the general trends in first ionisation energy values:**   * **down a group** * **across a period (main group elements) and**   **explain the exceptions to the general trends across a period**  **define and explain second and successive ionisation energies**  **describe how second and successive ionisation energies provide evidence for energy levels**  **recognise the relationship and trends in successive ionisation energies of an individual element**  explain how chemical properties of elements depend on their electronic structure  explain how atomic radius, screening effect and nuclear charge account for general trends in properties of elements in groups I and VII |  |  |  |
| **1.5 Oxidation and Reduction**  By the end of this section you should be able | **Good** | **Fair** | **Poor** |
| define oxidation and reduction in terms of electron transfer  use simple examples , e.g. Na with Cl2, Mg with O2, Zn with Cu2+ to describe oxidation and reduction in terms of electron transfer  apply knowledge of oxidation and reduction to explain the rusting of iron  define oxidising agent and reducing agent  arrange the electrochemical series of metals in order of their ease of oxidation (reactions, other than displacement reactions, not required)  carry out an experiment to show that halogens act as oxidising agents(reactions with bromides, iodides, Fe2+ and sulfites; half equations only required)  carry out an experiment to demonstrate the displacement reactions of metals (Zn with Cu2+, Mg with Cu2+)  explain what happens at each electrode during the electrolysis of:   * copper sulfate solution with copper electrodes * acidified water with inert electrodes   **(half equations only required)**  **describe and account for the observations of what happens at each electrode during the electrolysis of (teacher demo)**   * **aqueous sodium sulfate (using universal indicator)** * **aqueous potassium iodide (using phenolphthalein indicator) with inert electrodes (half equations only required)**   describe the extraction of copper by displacements using scrap iron  describe and explain ionic movement as observed during teacher demonstration  describe the following electrolytic processes: purification of copper, chrome and nickel plating. Give one everyday application of chrome and nickel plating e.g. cutlery |  |  |  |
| **2.1 Chemical Compounds**  By the end of this section you should be able to | Good | Fair | Poor |
| understand that compounds can be represented by chemical formulas  relate the stability of noble gasses to their electron configurations  describe bonding and valency in terms of the attainment of a stable electronic structure  state the octet rule  explain its limitations  use the octet rule to predict the formulas of simple binary compounds of the first 36 elements (excluding d-block elements) binary compounds of the first 36 elements (excluding d-block elements) and the hydroxides, carbonates, nitrates, hydrogencarbonates, sulfites and sulfates of these elements (where such exist).  recognise that Cu, Fe, Cr and Mn have variable valencies  relate the uses of helium and argon to their chemical unreactivity |  |  |  |
| **2.2 Ionic Bonding**  By the end of this section you should be able | Good | Fair | Poor |
| define ion, positive ion, negative ion  appreciate the minute size of ions  explain ionic bonding in terms of electron transfer  represent ionic bonds using dot and cross diagrams  describe the structure of a sodium chloride crystal having reviewed models  associate ionic substances with their characteristics    outline two uses of ionic materials in everyday life  test for anions in aqueous solutions: chloride, carbonate, nitrate, sulfate, phosphate, sulfite, hydrogencarbonate |  |  |  |
| **2.3 Covalent Bonding**  By the end of this section you should be able to | Good | Fair | Poor |
| define molecule  appreciate the minute size of molecules    explain covalent bonding in terms of the sharing of pairs of electrons (Single, double and triple covalent bonds)  represent covalent bonds in molecules using dot and cross diagrams  distinguish between sigma and pi bonding  distinguish between polar and non-polar covalent bonding  test a liquid for polarity using a charged plastic rod  give examples of polar and non-polar materials in everyday life (two examples in each case)  associate covalent substances with their characteristics  test the solubility of ionic and covalent substances in different solvents |  |  |  |
| **2.4 Electronegativity**  By the end of this section you should be able | Good | Fair | Poor |
| define electronegativity  recognise the trends in electronegativity values down a group and across a period  explain the general trends in electronegativity values   * down a group * across a period.   relate differences in electronegativity to polarity of bonds  predict bond type using electronegativity differences |  |  |  |
| **2.5 Shapes of Molecules and Intermolecular Forces**  By the end of this section you should be able | Good | Fair | Poor |
| describe the shapes of simple molecules  use appropriate modeling techniques to illustrate molecular shape  explain the basis for electron pair repulsion theory  use electron pair repulsion theory to explain the shapes of molecules for up to four pairs of electrons around the central atom refer to bond angles (Shapes of molecules with pi bonds not to be considered)  explain the relationship between symmetry and polarity in a molecule (dipole moments not required)  describe and distinguish between intramolecular bonding and intermolecular forces (van der Waals’, dipole-dipole, hydrogen bonding)  describe the effects of intermolecular forces on the boiling point of covalent substances  relate the differences in boiling points of H2 and O2 , C2H2 and HCHO and of H2O and H2S to the effect of intermolecular forces |  |  |  |
| **2.6 Oxidation Numbers**  By the end of this section you should be able | Good | Fair | Poor |
| define oxidation number, oxidation state  define oxidation and reduction in terms of change of oxidation numbers  state the rules for oxidation numbers (exclude peroxides, except for hydrogen peroxide)  calculate oxidation numbers of transition metals in their compounds and of other elements  use oxidation numbers in nomenclature of transition metal compounds  give an example of an oxidising and a reducing bleach |  |  |  |
| 3 Stoichiometry, Formulas and Equations | | | |
| **3.1 States of Matter**  By the end of this section you should be able to | Good | Fair | Poor |
| describe the motion of particles in solids, liquids and gases  explain diffusion  demonstrate diffusion (Graham’s law not required) using simple chemicals |  |  |  |
| **3.3 The Mole**  By the end of this section you should be able | Good | Fair | Poor |
| define the Avogadro constant  define the mole  calculate relative molecular mass from relative atomic masses  define s.t.p  define molar volume at s.t.p., molar mass, relative molecular mass (Mr)  convert moles to grams, litres and number of particles  convert grams, litres and number of particles to moles  convert moles to number of atoms of a molecular species  explain how a mass spectrometer can be used to determine relative molecular mass(Mr) (limited to simple treatment interpretation of mass spectra not required )  determine the relative molecular mass of a volatile liquid using suitable apparatus |  |  |  |
| **3.4 Chemical Formulas**  By the end of this section you should be able | Good | Fair | Poor |
| define empirical formula , molecular formula  calculate empirical formulas given the percentage composition by mass  calculate empirical formulas given the masses of reactants and products  calculate molecular formulas given the empirical formulas and the relative molecular masses  (examples should include simple biological substances, such as glucose and urea)  calculate percentage composition by mass  define structural formula  deduce, describe and explain structural formulas (simple examples) |  |  |  |
| **3.5 Chemical Equations**  By the end of this section you should be able | Good | Fair | Poor |
| recall that chemical equations are used to represent chemical reactions  construct chemical equations  balance simple chemical equations  balance redox equations (ionic equations only – ignore spectator ions)  perform calculations based on balanced equations using the mole concept (calculations in g and kg. Calculations may involve masses and volumes)  perform calculations based on balanced equations involving excess of one reactant (calculations in g and kg. Calculations may involve masses and volumes) |  |  |  |
| 4 Volumetric Analysis | | | |
| **4.1 Concentration of Solutions**  By the end of this section you should be able to | Good | Fair | Poor |
| define solution  define concentration  define molarity  express concentration of solutions in  mol-1(molarity), g l -1 and also in % (w/v), % (v/v), % (w/w)  appreciate the everyday use of % v/v e.g. in alcoholic beverages  calculate molarity from concentration in grams per litre and vice versa  calculate number of moles from molarity and volume  perform simple calculations involving percentage concentrations  calculate the effect of dilution on concentration  apply knowledge of concentrations of solutions to everyday examples  describe how colour intensity can be used as an indicator of concentration  Define a primary standard and a standard solution  prepare standard solution of sodium carbonate |  |  |  |
| **4.3 Volumetric Analysis**  By the end of this section you should be able | Good | Fair | Poor |
| identify appropriate apparatus used in volumetric analysis  use correct titrimetric procedure when carrying out titrations  solve volumetric problems, using the formula method  solve volumetric problems from first principles  carry out a titration between hydrochloric acid and sodium hydroxide solutions and use this titration to make a sample of sodium chloride (OL only)  standardise a hydrochloric acid solution using a standard solution of sodium carbonate  calculate the relative molecular mass of a compound and of the amount of water of crystallisation in a compound from titration data (balanced equations will be given in all volumetric problems)  determine the concentration of ethanoic acid in vinegar  determine the amount of water of crystallisation in hydrated sodium carbonate  carry out a potassium manganate(VII)/ammonium iron(II) sulfate titration  determine the amount of iron in an iron tablet  carry out an iodine/thiosulfate titration  determine the percentage (w/v) of hypochlorite in bleach |  |  |  |