



European Schools

Office of the Secretary-General of the Board of Governors

Ref. : 2004-D-72-en- 4

Orig. : FR

CHEMISTRY Syllabus (Years 4, 5, 6 & 7)

Approved by the Board of Governors of the European Schools

Meeting on 28, 29 and 30 April 2004 in Parma

Entry into force:

In September 2004: for years 4 and 6

In September 2005: for years 5 and 7.

PREAMBLE

This syllabus presents the programme of chemistry for the fourth year. It does not prescribe an order in which the material should be taught. The teacher decides the order in which the material should be taught the only limitation is that all of the 4th year programme must be taught in the fourth year.

The rubrics «Contents» and «Comments» of the programme indicate the obligatory parts of the programme. The rubric «Suggestions» provides ideas and possible practicals, which the teacher may or may not choose to use. Although the teacher is free to choose the method by which this programme is taught, she/he is encouraged to include the everyday applications of chemistry and experimental work. It is important that the teacher evaluates the risks of carrying out a particular practical for both students and the environment. The students should learn how to work in a laboratory in a way that is both safe and respects the environment.

Regarding the Baccalaureate examinations : "The examinations will normally cover the year 7 syllabus, but will also test knowledge gained in previous years, especially year 6."

MODULE 1 : What is chemistry ?

CONTENTS	COMMENTS	SUGGESTIONS
Chemistry is : ⇒ The study of all matter and how it can be changed into other forms; ⇒ The study of the properties of all matter and how these properties can be exploited.	☞ Study experimentally. ☞ Examples of chemistry in everyday life.	☞ Use different resources (videos, books, CD-rom, ...).

MODULE 2 : The structure of matter

CONTENTS	COMMENTS	SUGGESTIONS																								
<p>2.1.- Elements</p> <ul style="list-style-type: none"> ⇒ Definition: an element is a substance that cannot be broken down into simpler substances by chemical means. ⇒ The symbols of the common elements. <p>2.2.- Structure of the atom</p> <ul style="list-style-type: none"> ⇒ The historical development, considering the important stages, of the concept of the atom and its structure. ⇒ The nucleus and surrounding electrons: The structure of some common elements should be studied. <ul style="list-style-type: none"> ○ particles found in the nucleus (protons, neutrons) and in the space surrounding the nucleus (electrons), their relative mass, their relative charge; ○ the atomic number (or charge number) (Z) ; ○ the mass number (A). 	<p>The 3 states of matter and the kinetic theory of matter are assumed knowledge.</p> <ul style="list-style-type: none"> ☞ A systematic study of the elements is not required at this stage. ☞ The symbols of the elements could be considered from a historical perspective. ☞ Discuss the roles of Democritus, Dalton, Thomson, Rutherford ending with the atomic model of Bohr. This is sufficient at this level. ☞ Show experimentally the existence of different charges in matter. ☞ Particles found in the atom: <table border="1" data-bbox="874 757 1102 1346"> <thead> <tr> <th>Particle</th> <th>Relative mass</th> <th>Relative charge</th> </tr> </thead> <tbody> <tr> <td>Proton</td> <td>1</td> <td>+1</td> </tr> <tr> <td>Neutron</td> <td>1</td> <td>0</td> </tr> <tr> <td>Electron</td> <td>5×10^{-4}</td> <td>-1</td> </tr> </tbody> </table> <ul style="list-style-type: none"> ☞ Compare the parts of the atom to a concrete example to illustrate the size and mass of the constituents of the atom. ☞ $Z = n_{p^+} = n_{e^-}$ ☞ $A = n_{p^+} + n_{n^0}$ ☞ Notation ${}^A_Z X$ 	Particle	Relative mass	Relative charge	Proton	1	+1	Neutron	1	0	Electron	5×10^{-4}	-1	<p>Describe Rutherford's experiment. Use documents, audio-visuals : films, cassettes videos, CD-roms.</p> <p>Movement of a flow of water by a charged rod. Recall the electro-static experiments carried out in integrated science.</p> <p>For information, the absolute values of the masses and charges of the particles in the atom are shown below⁽¹⁾.</p> <table border="1" data-bbox="975 208 1169 723"> <thead> <tr> <th>Particle</th> <th>Mass /kg</th> <th>Charge / C</th> </tr> </thead> <tbody> <tr> <td>Proton</td> <td>$1,673 \times 10^{-27}$</td> <td>$1,602 \times 10^{-19}$</td> </tr> <tr> <td>Neutron</td> <td>$1,675 \times 10^{-27}$</td> <td>0</td> </tr> <tr> <td>Electron</td> <td>$9,109 \times 10^{-31}$</td> <td>$-1,602 \times 10^{-19}$</td> </tr> </tbody> </table> <p>⁽¹⁾ Source : IUPAC, <i>Quantities, Units and Symbols in Physical Chemistry</i>, 2nd Ed., Blackwell Science, 1993 [ISBN 0-632-03683-8].</p>	Particle	Mass /kg	Charge / C	Proton	$1,673 \times 10^{-27}$	$1,602 \times 10^{-19}$	Neutron	$1,675 \times 10^{-27}$	0	Electron	$9,109 \times 10^{-31}$	$-1,602 \times 10^{-19}$
Particle	Relative mass	Relative charge																								
Proton	1	+1																								
Neutron	1	0																								
Electron	5×10^{-4}	-1																								
Particle	Mass /kg	Charge / C																								
Proton	$1,673 \times 10^{-27}$	$1,602 \times 10^{-19}$																								
Neutron	$1,675 \times 10^{-27}$	0																								
Electron	$9,109 \times 10^{-31}$	$-1,602 \times 10^{-19}$																								

CONTENTS	COMMENTS	SUGGESTIONS
<p>⇒ The relative atomic mass (A_r).</p> <p>⇒ The relative molecular mass (M_r)</p> <p>⇒ Isotopes :</p> <ul style="list-style-type: none"> ○ Isotopes (of an element) differ by their mass. The number of protons and electrons are the same but the number of neutrons varies from one isotope to the next. <p>⇒ Electron arrangement in an atom: Bohr's model.</p> <ul style="list-style-type: none"> • Electrons exist in shells (K, L, M...) • Number of electrons in each shell. <p>⇒ Representation of the arrangement of electrons in an atom by using diagrams.</p>	<p>☞ The carbon atom ^{12}C is used as the reference value.</p> <p>☞ Study the isotopes of the following elements: H, C, O, Cl.</p> <p>☞ Limit the study of the electron configuration to the first 20 elements (H → Ca).</p> <p>☞ The importance of the outer shell of electrons to determine the reactivity of an element.</p> <p>☞ Correlation between the number of electrons in the outer shell and the position of the element in the periodic table.</p> <p style="text-align: right;">Group I : X• Group II : X•X• Group III : X•X•X•</p>	<p>Stress the importance of isotopes in medicine: (isotopic tracers, radiotherapy), in biology (isotopic markers and in archeology (carbon 14 dating). A study of radioactivity is not required, it is already studied in the physics course. Rutherford's experiments are an opportunity to recall the ideas of radioactivity .</p> <p>The relationship between the energy of the electrons in a shell and the energy of the atom can be considered. The flame tests of some elements can be used (salts of Cu, Na, K, Ca...) to illustrate the quantised energy of the electrons.</p>

CONTENTS	COMMENTS	SUGGESTIONS
<p>2.3.- Chemical Formulae</p> <ul style="list-style-type: none"> ⇒ The formula of an ionic compound represents the ratio of elements in the compound. ⇒ The formula of a covalent substance represents the number of atoms of each element present in the molecule. ⇒ Writing chemical formulae: <ul style="list-style-type: none"> ○ concept of valency; ○ valency of the elements; ○ existence of radicals; ○ valency of radicals; ○ naming compounds. 	<p>Consider some simple examples that have been seen before.</p> <ul style="list-style-type: none"> ☞ Consider the more important transition elements. ☞ Valency of an element = number of bonds that can be made with an atom of a certain element. ☞ Limit the study of the radicals to the principle groups: nitrate, nitrite, sulphate, sulphite, carbonate, hydroxide, phosphate and ammonium. ☞ Naming compounds made between metals and non-metals. ☞ Naming non-metal compounds. ☞ Naming the more important: acids, bases and salts. 	<p>Use molecular models to show the bonds that exist in a molecule and the geometry of a molecule.</p> <p>Given the abstract nature of this material, it may help to show some of the different uses of these substances or interesting facts about the substances.</p>

Module 3 : Periodic Classification of the elements

CONTENTS	COMMENTS	SUGGESTIONS
<p>⇒ Historical introduction.</p> <p>⇒ Elements in the same group or family have similar properties: group I = alkali metals, group II = alkaline earth metals and group VII = halogens.</p> <p>⇒ Trend of reactivity in a group.</p> <p>⇒ Periodicity of the chemical properties.</p> <p>⇒ Noble gases are inert.</p> <p>⇒ The properties of some of the interesting transition metals.</p>	<p>☞ Study experimentally – Write the corresponding balanced equations.</p> <p>☞ The properties and trends shown by the members of a family or a group.</p> <p>☞ The arrangement of elements in one period is based on the reactivity of the elements.</p> <p>☞ Periodic classification of the elements (MENDELEEV'S Table).</p> <p>☞ Illustrate the inertness of the noble gases by taking examples of their use as an inert 'atmosphere'.</p> <p>Study the properties of some transition metals experimentally.</p>	<p>All the practical work must follow safety rules of the laboratory.</p> <p>Reactivity of the alkali metals with water.</p> <p>Reactivity of the alkaline earth metals with water.</p> <p>Reactivity of the alkaline earth metals with acid (HCl dilué).</p> <p>Do not react the alkali metals with acid.</p> <p>Compare the reactivity of the halogens (Cl_2, Br_2 and I_2) with the solutions of the halides (KCl, KBr, KI) in a fume cupboard; compare the reactivity of Cl_2, Br_2 and I_2 with a metal (Mg) in a fume cupboard.</p> <p>Comparison of the reactivity of the alkali metals and the alkaline earth metals in the same period..</p> <p>E.g. Helium gas is used in air balloons, neon lighting, argon as an inert atmosphere for welding, xenon gas used in car headlights.</p> <p>Illustrate the use of transition metals in the electrical industry (cables, wires), in the chemical industry (catalysts), in the building industry (zinc or copper gutters, lead flashing, nails, screws, rivets... in brass, copper or zinc and in every day life (nickel in coins, chromium used for car fittings, gold or silver jewellery, tin for 'tin' cans, mercury thermometers, tungsten filaments in light bulbs...).</p>

CONTENTS	COMMENTS	SUGGESTIONS
<p>⇒ Periodic Table</p> <p>⇒ Reactivity of the elements depends upon the arrangement of the electrons, particularly the number of electrons in the outer shell, excluding the transition metals.</p>	<p>☞ The periodic table is based on the atomic number (z) and on the physical and chemical properties of the elements.</p> <p>The student must be able to:</p> <ul style="list-style-type: none"> ☐ Work out information concerning the properties of the elements based on the position in the periodic table. ☐ position the metal and non-metals in the periodic table ; ☐ position the noble gases and transition metals in the periodic table. <p>(The knowledge of the properties of metals and non-metals is assumed)</p> <p>There is no requirement to memorise the names: <i>alkali metals, alkaline earth metals, halogens and noble gases.</i></p>	

Module 4 : Chemical Reactions

CONTENTS	COMMENTS	SUGGESTIONS
<p>4.1.- Conservation of Mass</p> <ul style="list-style-type: none"> ⇒ Chemical equations; <ul style="list-style-type: none"> ☞ Write the formulae of the reactants and products; ☞ Balance equations. <p>4.2.- Energy changes in a chemical reaction</p> <ul style="list-style-type: none"> ⇒ exothermic reactions; ⇒ endothermic reactions . <p>⇒ An energy diagram to show the energy of the reactants and products and the energy released or absorbed during the reaction.</p> <p>⇒ Activation energy.</p>	<p>Show experimentally</p> <ul style="list-style-type: none"> ☞ A balanced equation satisfies the "Law of the conservation of mass"; ☞ The type and number of atoms does not change. ☞ Stoichiometric calculations are not required at this stage. <p>☞ Show experimentally</p> <p>☞ Use every day examples.</p> <p>☞ Activation energy can be explained by breaking bonds followed by making bonds.</p>	<p>Experiments: compare the mass before and after the reaction</p> <ul style="list-style-type: none"> ☞ Camera flash ; ☞ Hardening of wall plaster ; ☞ Precipitation reaction – mixing two solutions. <p>☞ Burning alcohol ;</p> <p>☞ Reactions in a calorimeter ;</p> <p>☞ dissolving $\text{NH}_4\text{Cl(s)}$ in water to show an endothermic reaction ;</p> <p>☞ thermal decomposition of glucose or Ag_2O ;</p> <p>☞ Synthesis of iron (II) sulphide ;</p> <p>☞ Reaction between H_2 and O_2 ;</p> <p>☞ Combustion of alkanes.</p> <p>Examples taken from everyday life:</p> <ul style="list-style-type: none"> • It is necessary «to light» a Bunsen burner or a cigarette lighter ; • One must not smoke in a petrol station; • Do not turn on a switch if the room is full of an inflammable vapour.

CONTENTS	COMMENTS	SUGGESTIONS
<p>4.3.- Factors affecting the rate of a reaction</p> <ul style="list-style-type: none"> <input type="checkbox"/> The concentration of the reactants; <input type="checkbox"/> The temperature; <input type="checkbox"/> The surface area of the reactants; <input type="checkbox"/> The presence of a catalyst. 	<p style="text-align: center;">} Study experimentally.</p>	<p><u>Suggestions of experiences:</u></p> <ul style="list-style-type: none"> <input type="checkbox"/> compare the rate of the reaction between oxygen and iron wool in air and pure oxygen. <input type="checkbox"/> compare the rate of the reaction between $\text{Mg(s)} + \text{HCl(aq)}$, considering concentration and temperature ; <input type="checkbox"/> compare the rate of the reaction between $\text{CaCO}_3\text{(s)}$ and HCl(aq), considering temperature, concentration , surface area ; <input type="checkbox"/> consider the effect of catalysing the reaction $\text{Zn(s)} + \text{H}_2\text{SO}_4\text{(aq)}$ with $\text{Cu}^{2+}\text{(aq)}$ <input type="checkbox"/> consider the effect of $\text{MnO}_2\text{(s)}$ on the decomposition of $\text{H}_2\text{O}_2\text{(l)}$.

Module 5 : Chemical Bonds

CONTENTS	COMMENTS	SUGGESTIONS
<p>5.1.- Ionic Bonding</p> <ul style="list-style-type: none"> ⇨ Formation of ions. ⇨ The ionic bond is the electrostatic attraction between ions of opposite charges ⇨ Ionic lattice. ⇨ The formula of an ionic compound shows the ratio of positive ions to negative ions. ⇨ Properties of salts: melting point, hardness, conduction of electricity in the solid state, molten state and solubility in water. ⇨ Pollution of water by salts. 	<ul style="list-style-type: none"> ☞ Schematic representation of the transfer of electrons between two atoms to show the formation of ions. ☞ Illustrate using a model of NaCl. ☞ Study experimentally. ☞ Illustrate by giving some examples of the uses of different salts. ☞ Pollution of fresh water by certain salts: <ul style="list-style-type: none"> - nitrates: the health risks due to the presence of nitrates in drinking water. - phosphates and nitrates : can cause eutrophication. 	<p>Combustion of Mg(s) in O₂(g). Direct synthesis of CuS or FeS. Reaction of a metal with Cl₂(g) in the fume cupboard.</p> <p>☞ Look at some crystals under a binocular microscope.</p> <p>Uses of NaCl: in food, sea water, salting roads in winter); CaCO₃: occurs naturally in water – hard water ; phosphates: used to soften water, to make fertilisers, to make washing powders; nitrates used to make fertilisers.</p>

CONTENTS	COMMENTS	SUGGESTIONS
<p>5.2.- Covalent Bonding</p> <ul style="list-style-type: none"> ⇨ Formation of a covalent bond. ⇨ The 'octet rule'. ⇨ Single, double and triple covalent bonds. ⇨ Electronegativity: considered as the ability of an atom in a covalent bond, to attract the bonding pair of electrons. ⇨ Polarisation of a bond: as a result of the asymmetric distribution of the bonding electrons. ⇨ Polar and non-polar molecules. ⇨ Study some of the properties of covalent properties: melting point, solubility in water, non-conduction of electricity. ⇨ Pollution of fresh water by hydrocarbons. 	<ul style="list-style-type: none"> ☞ Schematic representation of covalent bonding to show one or more pairs of electrons between two atoms. ☞ Examples of molecules: H₂, O₂, Cl₂, N₂, H₂O, CH₄, NH₃, CO₂. ☞ Be able to work out the formula of a covalent compound (as above). ☞ Use molecular models to illustrate the existence of molecules. ☞ To consider the geometry of the molecules: HCl, H₂O and CO₂. But do not consider this subject in detail at this level. ☞ Characterised by being hydrophilic and hydrophobic: H₂O and the hydrocarbons: <ul style="list-style-type: none"> ◦ Polar solvent (H₂O) ; ◦ non-polar solvent (hydrocarbons). ☞ Problems of an oil spillage at sea. 	<p>Show the effect of a charged rod held near a stream of water and then compare this to hexane.</p>

Module 6 : Introduction to organic chemistry

CONTENTS	COMMENTS	SUGGESTIONS
<p>6.1.- Introduction</p> <p>⇒ Compare the properties of certain inorganic substances (water, oxides, salts) to organic substances (solvents, plastic materials).</p> <p>⇒ The diversity of organic compounds.</p> <p>6.2.- Alkanes</p> <p>⇒ Methane</p> <p>⇒ Fractional distillation of crude oil gives rise to a family of hydrocarbons containing simple bonds C-C and consisting only of C and H: the alkanes.</p> <p>⇒ Naming.</p> <p>⇒ Molecular formula.</p>	<p>Compare the properties of inorganic substances with organic substances. :</p> <ul style="list-style-type: none"> - electrical conduction ; - solubility ; - melting point/ boiling point. <p>All organic compounds contain carbon and usually hydrogen. The carbon atom is able to bond to itself, to make chains.</p> <p>Shape of the methane molecule.</p> <p>Importance of methane: there are many examples of natural sources of methane - biomethane, major constituent of natural gas.</p> <p>Identification, description and comparison of the constituents resulting from the fractional distillation of crude oil.</p> <p>Up to C₈</p>	<p>Show the risk associated with the critical mixture of methane and air.</p>

CONTENTS	COMMENTS	SUGGESTIONS
<p>⇒ Structural formula</p> <p>⇒ Combustion of an alkane: complete and incomplete.</p> <p>⇒ The role of methane and carbon dioxide in the greenhouse effect.</p>	<p><u>Illustrated here by ethane</u></p> <p>Molecular formula: C_2H_6</p> <p>structural formula:</p> <ul style="list-style-type: none"> • showing all the bonds in the molecule. <p>e.g.</p> $\begin{array}{c} H & H \\ & \\ H - C & - C - H \\ & \\ H & H \end{array}$ <ul style="list-style-type: none"> • showing only certain bonds in the molecule <p>e.g. CH_3CH_3 or CH_3-CH_3</p> <ul style="list-style-type: none"> ○ No calculations are required based on the combustion of an alkane. ○ Formation of soot and carbon monoxide from incomplete combustion. ○ The health risks associated with carbon monoxide and soot. 	<p>Illustrate the greenhouse effect by:</p> <ul style="list-style-type: none"> • Comparing the temperature of the inside of a vessel made from different materials eg glass, plastic, pyrex. • Comparing the temperature of the inside of a vessel filled with different gases eg air, carbon dioxide, methane.

CONTENTS	COMMENTS	SUGGESTIONS
<p>⇒ Cracking of long-chain hydrocarbons</p> <p>6.3.- Alkenes</p> <p>⇒ Addition reaction of an alkene.</p> <p>⇒ Polymerisation: limited to the formation of polyethene.</p> <p>⇒ Saturated fats and unsaturated fats</p>	<p>Cracking results in the formation of unsaturated hydrocarbons, containing a double bond C=C, formed exclusively from C and H: the alkenes.</p> <p>Limit to the addition of bromine water, Br₂ (aq). The mechanism is not required.</p> <p>Be able to write a balanced equation and show the structure of the product.</p> <p>A test to show unsaturation: using bromine water.</p> <p>Be able to write a balanced equation but the mechanism is not required.</p> <p>Plastic materials.</p> <p>Know the meaning of the terms 'saturated fat' and 'unsaturated fat'.</p> <p>The structure of fats / oils is not required.</p>	<p>Give the significance of the product CFC since it is implicated in the depletion of the ozone layer.</p> <p>Discuss the environmental problems of plastic waste and possible solutions.</p> <p>Consider the biological importance of saturated and unsaturated fats.</p>

PREAMBLE

This syllabus presents the programme of chemistry for the fifth year it does not prescribe an order in which the material should be taught. The teacher decides the order in which the material should be taught, the only limitation is that the entire 5th year programme must be taught in the fifth year. The examination at the end of the fifth year is harmonised and it is essential that the entire programme has been taught. The 5th year programme must be coordinated to facilitate the writing of the harmonised paper.

The rubrics «Contents» and «Comments» of the programme indicate the obligatory parts of the programme. The rubric «Suggestions» provides ideas and possible practicals, which the teacher may or may not choose to use. Although the teacher is free to choose the method by which this programme is taught, she/he is encouraged to include the everyday applications of chemistry and experimental work. It is important that the teacher evaluates the risks of carrying out a particular practical for both the students and the environment. The students should learn how to work in a laboratory in a way that is safe and respects the environment.

MODULE 1 : Redox reactions

CONTENTS	COMMENTS	SUGGESTIONS
<p>1.1 Redox reactions.</p> <p>1.1.1.- Definition of redox limited to the loss or gain of oxygen</p> <p>⇒ Oxidation as the gain of oxygen.</p> <p>⇒ Reduction as a loss of oxygen.</p> <p>⇒ Redox can be considered as loss and gain of oxygen.</p>	<p>Introduce ideas through showing simple experiments.</p> <p>Consider the corrosion of iron.</p>	<p>Combustion (oxidation) of Mg(s) , of Fe(s), of C(s) or of S₈(s) in O₂(g).</p> <p>Thermal decomposition of Ag₂O(s) ; Reactions CuO(s) + Fe(s) ; Mg(s) + H₂O(l) ; Fe₂O₃(s) + Al(s) ; Mg(s) + CO₂(g)</p>

5th YEAR SYLLABUS

CONTENTS	COMMENTS	SUGGESTIONS
<p>1.1.2.- General definition of redox in terms of loss and gain of electrons.</p> <p>⇒ Redox is defined as a loss and gain of electrons.</p> <p>⇒ Definition of oxidising agent and reducing agent in terms of electron transfer.</p> <p>⇒ To be able to write half-ionic equations for redox reactions and the overall ionic equations.</p> <p>⇒ Reactivity series of metals.</p> <p>⇒ To consider in detail, an application of a redox reaction.</p>	<p>Oxidation: loss of electrons. Reduction: gain of electrons.</p> <p>Oxidising agent: accepts electrons. Reducing agent: donates electrons.</p> <p>Redox reactions involving complex ions or radicals are not required.</p> <p>Classify the following metals according to their ability to act as oxidising agents: Ag, Cu, Zn, Fe, Al and Mg. Derive this experimentally from the 'displacement reactions' for example: $Zn(s) + Cu^{2+}(aq) \rightarrow Zn^{2+}(aq) + Cu(s)$.</p> <ul style="list-style-type: none"> ○ Blast furnace: Outline of how a blast furnace works and its use to extract iron from its ore. Only consider the reduction of iron III oxide by carbon monoxide (even though in the modern blast furnace this is not the only reduction reaction, since methane is added which supplies hydrogen). 	<p>Illustrated by simple reactions such as :</p> $Zn(s) + Cu^{2+}(aq)$ $I^{-}(aq) + Cl_2(aq)$ $Fe(s) + H^{+}(aq)$ $Fe(s) + S(s)$ <p>The idea of oxidation number may be introduced..</p> <p>Application: Reduction of metal oxides by a more reactive metal or by carbon.</p> <p>Consider everyday examples: corrosion of metals and protection against corrosion, photosensitivity of silver halides, extraction of certain metals, incompatibility of joining certain metals e.g. soldering certain metals.</p>

5th YEAR SYLLABUS

CONTENTS	COMMENTS	SUGGESTIONS
<p>1.2.- The Daniell Cell as an example of an electrochemical cell.</p> <p>⇒ Experimental study of the Daniell cell.</p>	<p>Identification of the different components of the Daniell cell.</p> <p>Measurement of the electro-motive force between the two electrodes.</p> <p>Identification of the reaction that takes place in the Daniell cell and provides the source of electrons.</p> <p>The ionic-half equations for the reactions of oxidation and reduction at each electrode.</p> <p>Identification of each electrode where oxidation and reduction takes place, in relation to the polarity of the electrodes.</p> <p>Principle of the working of an electrochemical cell: a redox reaction where the electrons are transferred from the negative electrode (site of oxidation) towards the positive the electrode (site of reduction), via an external circuit.</p>	

MODULE 2 : Acids and alkalis

CONTENTS	COMMENTS	SUGGESTIONS
<p>2.1.- ARRHENIUS definition</p> <p>⇒ Acid: a substance which dissociates in aqueous solution, liberating hydrogen ions, $H^+(aq)$.</p> <p>⇒ Alkali: a substance which dissociates in aqueous solution liberating hydroxide ions, $OH^-(aq)$.</p>	<p>At this level, do not use the term oxonium ion $H_3O^+(aq)$.</p> <p>Do not consider the definition of a base by Brønsted or Lewis. At this stage, $NH_3(aq)$ will be treated as $NH_4OH(aq)$.</p> <div style="border: 1px solid black; padding: 10px; text-align: center;"> <p>This module should be studied from an experimental approach in all aspects.</p> </div>	<p>Study experimentally some everyday life acids and alkalis: vinegar, products for unblocking pipes, detergents, products used for descaling, indigestion tablets, aspirin.</p>
<p>2.2.- Dissociation of acids and alkalis in aqueous solution.</p> <p>⇒ Equation to represent the dissociation of an acid in aqueous solution: $HX(aq) \rightarrow H^+(aq) + X^-(aq)$.</p> <p>⇒ Equation to represent the dissociation of an alkali in aqueous solution: $MOH(aq) \rightarrow M^+(aq) + OH^-(aq)$.</p>	<p>Use experimentally, a solution of an indicator or universal indicator paper.</p> <p>At this level, pH should be considered as a simple comparison of the acidity or alkalinity of an aqueous solution.</p> <p>pH calculations are not required.</p> <p>The theory of strong/weak acids/alkalis should not be studied here. This is studied in the 7th year. At this level, limit the study to an experimental approach; to observe that acids and alkalis can be more or less acidic / alkaline.</p>	
<p>2.3-Acidity of a solution – pH scale</p> <p>⇒ Use of indicators, as shown by colour changes, to show if a solution is acidic, neutral or alkaline.</p> <p>⇒ pH in relation to hydrogen ion concentration, $[H^+(aq)]$.</p> <p>⇒ pH scale.</p> <p>⇒ Show experimentally the existence of weak and strong acids and alkalis.</p>		

5th YEAR SYLLABUS

CONTENTS	COMMENTS	SUGGESTIONS
<p>2.4.- Chemical properties of acids</p> <p>⇒ Chemical properties of dilute acids:</p> <ul style="list-style-type: none"> <input type="checkbox"/> reaction with certain metals; <input type="checkbox"/> reaction with certain metal oxides; <input type="checkbox"/> reaction with alkalis; <input type="checkbox"/> reaction with certain salts: carbonates. <p>2.5.- Acid Rain</p> <p>⇒ Acid rain :</p> <ul style="list-style-type: none"> <input type="checkbox"/> nature; <input type="checkbox"/> causes; <input type="checkbox"/> consequences; <input type="checkbox"/> remedies. 	<p>Study these reactions experimentally. Students should be able to write balanced equations to represent these reactions.</p> <p>Metal + acid → salt + Hydrogen</p> <p>Metal oxide + acid → salt + water</p> <p>Alkali + acid → salt + water</p> <p>Acid + carbonate → salt + water + carbon Dioxide</p> <p><i>Students should know the tests for the gases carbon dioxide and hydrogen.</i></p> <p>The above reactions are used to prepare soluble salts. Insoluble salts can be prepared by the addition of two salt solutions, the resulting insoluble salt forming as a precipitate. Ex. : $\text{AgNO}_3(\text{aq}) + \text{NaCl}(\text{aq}) \rightarrow \text{AgCl}(\text{s}) + \text{NaNO}_3(\text{aq})$.</p> <p>Formation of acidic oxides : $\text{CO}_2, \text{SO}_2, \text{NO}_x$.</p> <p>Reaction of acidic oxides with H_2O leading to the formation of acids.</p> <p>Effects: augmentation of the natural acidity of water, effects on animals, flowers, soil and buildings.</p> <p>Remedies: cleaning waste from a power station, catalytic conversion of exhaust fumes, reduction of the consumption of fossil fuels.</p>	<p>Do not add an alkali metal to an acid.</p> <p>Measure the pH of different mineral waters and soil.</p>

Module 3 : Quantitative Chemistry

CONTENTS	COMMENTS	SUGGESTIONS
<p>3.1.- The molar mass and the mole</p> <ul style="list-style-type: none"> ⇒ Concept of the mole and its definition. ⇒ Relative atomic mass and relative molecular mass. ⇒ Avogadro's constant. ⇒ Simple calculations based on the mole and relative atomic / molecular mass. <p>3.2.- Concentration</p> <ul style="list-style-type: none"> ⇒ Definition. ⇒ Units. ⇒ Preparation of solutions of a given concentration. ⇒ Concentration calculations. <p>3.3.- Acid-base titration</p> <ul style="list-style-type: none"> ⇒ Titrate a solution of a strong monoprotic acid against a solution of a strong monobase (or vice versa) in the presence of an acid-base indicator to identify the equivalence point. 	<p>The definition of the mole must include a reference to the Avogadro's constant.</p> <p>The unit for molar mass is g mol^{-1}.</p> <p>Use this formula and apply it to chemical reactions.</p> <div style="border: 1px solid black; padding: 5px; width: fit-content; margin: 10px auto;"> $n(\text{mol}) = \frac{m(\text{g})}{M(\text{g mol}^{-1})}$ </div> <p>Use the formula</p> <div style="border: 1px solid black; padding: 5px; width: fit-content; margin: 10px auto;"> $c(\text{mol dm}^{-3}) = \frac{n(\text{mol})}{V(\text{dm}^3)}$ </div> <p>The student should know the experimental procedure for titration.</p> <p>Carry out titrations in the laboratory.</p> <p>To try calculations that use data taken from a titration to determine the concentration of an unknown solution.</p>	<p>The relative abstract character of these ideas means that students often find these ideas difficult. As a result, students should be introduced to these ideas slowly, using suitable analogies.</p>

Module 4 : Organic chemistry

CONTENTS	COMMENTS	SUGGESTIONS
<p>4.1.- Organic compounds containing oxygen</p> <p>⇒ <u>Alcohols:</u></p> <ul style="list-style-type: none"> □ Production of ethanol by fermentation. □ Identification of the functional group. □ Nomenclature of primary alcohols. <p>□ Physical and chemical properties:</p> <ul style="list-style-type: none"> ☞ solubility in water; ☞ combustion; ☞ oxidation of ethanol to ethanoic acid. <p>□ Existence of polyalcohols: glycerol and glucose (cyclic form)</p>	<p>Be able to write the balanced equation: $C_6H_{12}O_6(aq) \rightarrow 2 C_2H_5OH(aq) + 2CO_2(g)$</p> <p>Until C₄.</p> <p>Only naming of primary alcohols required. The classification of alcohols as primary, secondary and tertiary is not required.</p> <p>To know that solubility decreases with increasing carbon chain length. No explanation is required.</p> <p>There is no requirement to explain that an aldehyde is the intermediate in the oxidation of an alcohol to a carboxylic acid. Simple equation required:</p> $CH_3CH_2OH_{(l)} \xrightarrow{[oxidant]} CH_3COOH_{(aq)}$ <p>Stereochemistry of glucose is not required. The structure of glucose does not have to be remembered.</p>	<p>Show the fermentation reaction in the laboratory. Be aware that fermentation is an important biological reaction.</p> <p>Compare the combustion of a range of alcohols comparing the flame and energy released.</p> <p>Laboratory practical the oxidation of ethanol to ethanoic acid.</p> <p>Mention the importance of glycerol as a constituent of fats and glucose as the main source of energy in respiration.</p>

5th YEAR SYLLABUS

CONTENTS	COMMENTS	SUGGESTIONS
<p>⇒ <u>Carboxylic acids and their derivatives</u> :</p> <ul style="list-style-type: none"> □ Identification of the functional group; □ Nomenclature of monocarboxylic acids; □ Chemical properties : <ul style="list-style-type: none"> ☞ Acid character; ☞ Esterification: reaction with an alcohol. □ Fatty acids: □ Soaps: preparation and structure. 	<p>Until C₄.</p> <p>State the weak acid character of the carboxylic acids without extending into an explanation based on the partial dissociation of weak acids into ions.</p> <p>No mention of the reversible nature of esterification is required.</p> <p>Carboxylic acids with a high number of carbon atoms. Difficult to dissolve in water</p> <p>Hydrolysis of an ester by an alkali.</p> <p>Soaps are salts (potassium or sodium) of fatty acids.</p> <p>The soap molecule has one end that is polar (hydrophilic) and another end that is non-polar (hydrophobic) and this structure allows the soap to clean.</p> <p>Only a word equation is required: Fat + sodium hydroxide → glycerol + sodium salt of the fatty acid (soap)</p>	<p>Refer to the odours and tastes of esters, for example the smell of fruits</p> <p>Mention the biological importance of fatty acids and in particular unsaturated fatty acids.</p> <p>Show experimentally the preparation of a soap and the properties of detergents.</p>

5th YEAR SYLLABUS

CONTENTS	COMMENTS	SUGGESTIONS
<ul style="list-style-type: none">□ Amino acids☞ General formula.	General formula of an amino acid Formulae of glycine and alanine.	Define a protein as a polypeptide with biological functions (enzymes, hormones, cell structure). The structure of a polypeptide or a protein could be shown. The biological importance of amino acids and proteins could be discussed.

KNOWLEDGE OF THE MATERIAL from YEAR 4 & 5 SYLLABUS IS ASSUMED

CONTENT	COMMENTS	SUGGESTIONS
<p>1.- ELECTRONIC STRUCTURE OF THE ATOM AND THE PERIODIC TABLE</p> <p>1.1.- Electron shells and subshells Emission spectrum of an atom.</p> <p>Quantisation of energy in the atom, introduction of quantum number n (principle quantum number)</p> <p>Representations of the electronic configuration of an element by electron shells and subshells (Bohr's Atomic Model). Hund's rule.</p> <p>1.2.- Orbital Model Electronic orbitals: s, p, d, f, introduction of quantum number l (orbital shape) and quantum number m (number of orbitals for each value of l).</p> <p>Pauli Exclusion Principle</p> <p>Schematic representation of the possible electronic configuration</p>	<p>The students should understand the difference between line emission spectra and a continuous spectrum. To know that all elements have a line emission spectrum that is unique to the element. To be able to explain the line emission spectrum of hydrogen.</p> <p>First ionisation energies of the elements and successive ionisation energy of an element.</p> <p>Shells 1, 2, 3, ... or K, L, M, \dots Subshells s, p, d, f.</p> <p>The shape and orientation of s and p orbitals are required. The shape and orientation of d and f orbitals are not required.</p> <p>An orbital represents a region of space in which there is a high probability of finding an electron.</p> <p>Spin of the electron, introduction of quantum number s.</p> <p>Example: the electron configuration of nitrogen ($Z = 7$)</p> <p style="margin-left: 40px;">Shell K: $\boxed{\uparrow\downarrow}$ $1s^2$</p> <p style="margin-left: 40px;">Shell L: $\boxed{\uparrow\downarrow}$ $2s^2$ $\boxed{\uparrow}$ $2p_x$ $\boxed{\uparrow}$ $2p_y$ $\boxed{\uparrow}$ $2p_z$</p>	<p>Historical approach.</p> <p>Flame tests for elements of groups I and II.</p> <p>Heisenberg's Uncertainty Principle.</p>

CONTENT	COMMENTS	SUGGESTIONS
<p>1.3.- Periodicity in the third period and trends within groups</p> <p>Periodicity and trends of</p> <ul style="list-style-type: none">- radii of atoms and ions;- ionisation energies;- electron affinities;- electronegativities. <p>1.4.- Periodicity of the chemical properties of oxides.</p>	<p>Periodicity of the properties of the 3rd period elements (Na → Ar) and trends within groups I, II and VII.</p> <p>Periodicity of the chemical properties of the oxides of elements of the third period (excluding transition elements): from Na to Cl</p>	

CONTENT	COMMENTS	SUGGESTIONS
<p>2.- CHEMICAL BONDS</p> <p>2.1.- Covalent bonds Formation of molecular orbitals. σ and π molecular orbitals. sp, sp^2 and sp^3 hybrid orbitals</p> <p>Use of the VSEPR theory to predict the shape of simple molecules and ions.</p> <p>Use of the difference in electronegativity to predict bond polarity. Be able to predict if a molecule is polar or non-polar.</p>	<p>Molecular orbitals are considered as a result of the overlap of atomic orbitals.</p> <p>Polar and non-polar molecules in relation to:</p> <ul style="list-style-type: none"> • bond polarity; • shape (symmetry) of the molecule. 	
<p>Coordinate bonds</p>	<p>H_3O^+, NH_4^+ and $Cu(NH_3)_4^{2+}$</p>	<p>Electronic structure and geometry of some complex ions.</p>
<p>2.2.- Ionic bonds Formation of ions leading to ionic bonding. Crystal lattice.</p> <p>Possible covalent character of some ionic bonds (polarisation of an ion).</p>	<p>Limit to NaCl. Use the crystal model to show that each ion is surrounded by ions carrying an opposite charge. Coordination numbers are not necessary.</p> <p>The fact that some covalent compounds have some ionic character and vice versa should be considered.</p>	<p>Eg LiCl, $BeCl_2$, $AlCl_3$</p>
<p>2.3.- Comparison of properties of covalent and ionic compounds Comparison of solubility, conductivity and melting points for ionic and covalent compounds.</p>		

CONTENT	COMMENTS	SUGGESTIONS
<p>2.4.- Metallic bonds</p> <p>2.5.-Polymorphism The existence of various allotropic forms for some elements.</p>	<p>The description of a metallic bond as an array of cations in a 'sea' of electrons.</p> <p>The importance of metallic bonds in conductivity of metals.</p> <p>Allotropic forms of carbon: comparison of the structure of graphite, diamond and fullerenes (C₆₀).</p>	

The concepts of mole, Avogadro's number, relative molecular mass, molecular molar mass are prerequisites for the present module.		SUGGESTIONS
CONTENT	COMMENTS	
3.- INTERMOLECULAR BONDS Van der Waals bonds. Permanent dipole-permanent-dipole interactions. Hydrogen bonds. The influence of inter-molecular bonds on: <ul style="list-style-type: none"> • melting points; • boiling points; • solubility in water and in non-polar solvents. 	Van der Waals forces between non-polar molecules Permanent dipole-permanent-dipole interactions between polar molecules. Use H ₂ O & NH ₃ to illustrate substances that are hydrogen bonded. Comparison of the boiling points of hydrides of group IV, V, VI and VII elements.	The existence of hydrogen bonding in DNA and proteins could be mentioned.
4.-IDEAL GAS LAW Definition of an ideal gas. The use of the ideal gas law $pV = nRT$ to calculate : <ul style="list-style-type: none"> • the molar volume of an ideal gas at different temperatures; • the molecular molar mass of gaseous compounds and volatile organic substances. 	Reminder: SI units must be used within this area of work.	The laws of Boyle-Mariotte, Avogadro-Ampère and Gay-Lussac.

CONTENT	COMMENTS	SUGGESTIONS
<p>5.- ENERGY IN CHEMISTRY</p> <p>5.1.- Conservation of energy The total amount of energy is conserved, although energy can be converted from one form to another.</p> <p>5.2.- Enthalpy change Definition of the enthalpy of a system (H). Enthalpy change or enthalpy (heat) of reaction (ΔH). Calculation of enthalpy change using Hess' Law.</p> <p>5.3.- Entropy change Entropy change (ΔS) accompanying a transformation (chemical reaction or change of state).</p> <p>5.4.- Spontaneity of a transformation Factors affecting the spontaneity of a reaction.</p>	<p>Energy can be converted from one form to another, but the total amount of energy in a system isolated from its surrounding always remains constant (taking mass as a form of energy).</p> <p>The enthalpy of a system cannot be measured; only enthalpy change can be determined. Enthalpy change under standard conditions (ΔH^\ominus). Different types of enthalpy change:</p> <ul style="list-style-type: none"> • enthalpy of combustion; • enthalpy of formation; • lattice enthalpy; • bond enthalpy; • enthalpy of solution; • enthalpy of neutralization; • Born-Haber cycle. <p>Entropy as a measure of the degree of disorder of a system. Entropy change under standard conditions (ΔS^\ominus). Qualitative approach of the sign of the entropy change based on the comparison of the degree of (dis)order of a system before and after a transformation.</p> <p>The fact that not only ΔH determines the spontaneity of a reaction should be stressed. The role of ΔS should be mentioned. The Gibbs free energy function (G) and the relationship: $\Delta G = \Delta H - T \Delta S$ should be considered, but only from a qualitative point of view.</p>	<p>Experimental approach (dissolving hydrated and anhydrous copper sulphate). Calculations based on experimental data.</p> <p>The use of a data book should be encouraged.</p>

CONTENT	COMMENTS	SUGGESTIONS
<p>6.- CHEMICAL KINETICS AND EQUILIBRIA</p> <p>6.1.- The factors which determine the rate of a chemical reaction.</p> <p>6.2.-Collision Theory and Transition State Theory Energy (enthalpy) diagram of the reagents and products in the course of a chemical reaction.</p> <p>Collision theory and activation energy</p> <p>Maxwell-Boltzmann Distribution</p> <p>Effect of catalysts on rate of the reaction.</p>	<p>Temperature, concentration and surface area of the reagents, presence of a catalyst.</p> <p>Use the collision theory and concept of activation energy to explain the effect on the rate of a reaction if temperature, concentration or surface area is change.</p> <p>Catalysts lower the activation energy for the reaction</p>	<p>Study the rate of a reaction between HCl and Mg</p> <p>Comparison of the reaction rate of the hydrolysis of urea</p> <ul style="list-style-type: none"> - by an acid; - by urease. <p>(NB: the reaction rate can be estimated by the change in conductivity of the solution).</p>
<p>6.3.- Reversible reactions Knowledge of dynamic equilibria for homogeneous and heterogeneous reactions. The law of chemical equilibrium.</p> <p>The importance of the numerical value of K_c</p>	<p>Introduction of equilibrium constant K_c. Experimental determination of an equilibrium constant. Calculations based on the law of equilibrium.</p>	<p>Esterification.</p>
<p>6.4.- Factors influencing equilibria The effect of the following factors on the position of equilibrium and K_c within the system: Le Chatelier's Principle</p> <ul style="list-style-type: none"> • temperature, • concentration, • catalyst • pressure 	<p>A variation in temperature affects K_c and modifies the position of equilibrium. A change in concentration alters the position of equilibrium without changing K_c. A change in pressure alters the position of equilibrium, for gaseous systems, without changing K_c. The addition of a catalyst has no effect on the position of the equilibrium, but it causes the system to reach equilibrium more rapidly than it otherwise would. No calculation of K_p is required.</p>	<p>Industrial applications: synthesis of $\text{NH}_3(\text{g})$ and $\text{H}_2\text{SO}_4(\text{l})$.</p>

CONTENT	COMMENTS	SUGGESTIONS
<p>7.- ORGANIC CHEMISTRY</p> <p>7.1.- Properties of hydrocarbons</p> <p>7.1.1.- Alkanes, alkenes</p> <p>Origin. Naming. Structural formula.</p> <p>Structural and geometric isomerism.</p> <p>Physical properties.</p> <p>Chemical properties:</p> <ul style="list-style-type: none"> • alkanes - combustion • alkanes - free radical substitution. • alkenes - electrophilic addition. 	<p>Naming : C₁ (or C₂) to C₁₀. A structural formula can show:</p> <ul style="list-style-type: none"> • all the bonds <p>E.g.:</p> $\begin{array}{ccccccc} & & \text{H} & & \text{H} & & \text{H} \\ & & & & & & \\ \text{H} & - & \text{C} & - & \text{C} & - & \text{C} & - & \text{H} \\ & & & & & & \\ & & \text{H} & & \text{O-H} & & \text{H} \end{array}$ <ul style="list-style-type: none"> • the bonds present between the carbon atoms in the chain. E.g. CH₃ -CH(OH)-CH₃ • the arrangement of the atoms or groups of atoms in the molecule E.g. CH₃CH(OH)CH₃ <p>Limit geometric isomers to cis/trans isomers of alkenes.</p> <ul style="list-style-type: none"> • solubility in different solvents (polar and non-polar); • melting and boiling points. <p>Reaction of a liquid alkane with bromine solution.</p> <p>Electrophilic addition to alkenes:</p> <ul style="list-style-type: none"> • addition of bromine to an alkene; • mechanism of electrophilic addition to alkenes; <p>The importance of the inductive effect to the addition to asymmetric alkenes (the explanation of Markovnikov's empirical rule).</p>	<p>Fractional distillation of crude oil.</p> <p>Polymerisation of alkenes [this topic has been studied in the 4th and 5th year course].</p>

CONTENT	COMMENTS	SUGGESTIONS
<p>7.1.2.- Aromatic compounds Structure of benzene.</p>	<p>Explain the structure of benzene on the basis of physical and chemical evidence:</p> <ul style="list-style-type: none"> • all C-C bonds are of equal length (between that of a single bond and of a double bond); • the delocalisation of π bonding electrons increases the stability of the aromatic ring (as shown by experimental data); • benzene does not readily undergo addition reactions. • Halogenation (use of halogen carriers, e.g. $AlCl_3$, $FeBr_3$); • nitration of benzene. 	<p>For example, using heats of combustion and heats of hydrogenation.</p>
<p>7.2.- Determination of the structure of an organic substance Empirical and molecular formulae. Methods used to elucidate molecular structures.</p>	<p>Determination of the empirical formula.</p> <ul style="list-style-type: none"> • Use of mass spectrometry (parent ion and fragmentation patterns), • IR spectroscopy (O-H, C-H, C=O, C-O and the importance of the fingerprint region); <p>Low resolution 1H NMR to determine molecular structures.</p>	<p>Combustion methods. Use of computer analysis of spectra.</p>

CONTENT	COMMENTS	SUGGESTIONS
<p>1.- ACIDS AND BASES</p> <p>1.1.- Brønsted definition Acids and bases according to Brønsted.</p> <p>Conjugate acid/base (acid/base pair)</p> <p>Amphoteric substances.</p>	<p>Acids as proton donors and bases as proton acceptors. Equation for ionization of acid and bases :</p> $HA(aq) + H_2O(l) \rightarrow H_3O^+(aq) + A^-(aq)$ $B(aq) + H_2O(l) \rightarrow BH^+(aq) + OH^-(aq)$ <p>$HA_{(aq)}/A^-(aq)$; $BH^+_{(aq)}/B_{(aq)}$</p> <p>Amphoteric substance: can work as an acid (proton donor) or a base (proton acceptor), depending on the conditions.</p> <p>E.g.: $HCO_3^-(aq) + H_2O(l) \rightarrow H_3O^+(aq) + CO_3^{2-}(aq)$ $HCO_3^-(aq) + H_2O(l) \rightarrow OH^-(aq) + H_2CO_3(aq)$ [NB: $H_2CO_3(aq)$ or $CO_2(aq)$]</p> $H_2O(l) + H_2O(l) \rightarrow H_3O^+(aq) + OH^-(aq)$	<p>e.g. HX/X^- (X=halogen), NH_4^+/NH_3, H_3O^+/H_2O, H_2O/OH^-.</p> <p>e.g. HCO_3^-, HS^-, HSO_3^-, HSO_4^-, $H_2PO_4^-$, HPO_4^{2-}.</p>
<p>Ionic dissociation of water and the amphoteric nature of water.</p> <p>Ionic product of water.</p> <p>1.2.- pH and the measurement of pH Definition of pH. Determining pH :</p> <ul style="list-style-type: none"> acid/base indicators; pH-meter. 	<p>Ionic product of water K_w, pK_w. K_w is temperature dependent. The value of K_w at 25°C (298 K) must be known.</p> <p>A pH indicator is an acid/base pair where the acidic and basic forms have different colours in solution. Relationship between the colour change and the pK_a of the indicator.</p> <p>No theory or technical details required.</p>	<p>Experimental demonstration of the pH differences between solutions of strong acids and bases and weak acids and bases at the same concentration. The effect of the concentration of the strong (or weak) acid (or base) on the pH of a solution. Determination of the acid constant of ethanoic acid.</p> <p>Experiments using a pH-meter</p>
<p>1.3.- Relative strength of acids and bases Strong acids and bases. Weak acids and bases.</p>	<p>Relative strength of inorganic and organic monoprotic acids and monobases by reference to the extent of ionization in solution. Explanation based upon the structural formula (bond energies, electronegativities, inductive effects) and upon experimental evidence.</p>	

CONTENT	COMMENTS	SUGGESTIONS
<p>Ionization constant of weak acids and bases.</p> <p>Calculations involving K_a, K_b, concentration and pH of weak monoacids or monobases.</p> <p>1.4-pH of aqueous solutions</p> <p>Calculation of pH of aqueous solutions.</p> <ul style="list-style-type: none"> • Acids • bases 	<p>K_a and K_b, pK_a and pK_b; the relationship between K_a and K_b (or pK_a and pK_b) of an acid/ base pair.</p> <p>Classification of acids and bases according to their K_a or K_b value.</p> <p>Only the use of approximation in calculations involving K_a or K_b is required.</p> <p>Calculation of pH of solutions of strong mono/diprotic acids and strong bases.</p> <p>Calculations of pH of solutions of weak monoprotic acids and weak bases.</p> <p>Calculations on a mixture of acids or a mixture of bases are not required.</p>	

CONTENT	COMMENTS	SUGGESTIONS
<p>1.5.- Acid-base Titrations Titration of</p> <ul style="list-style-type: none"> • a strong monoprotic acid with a strong monobase; • a strong monobase with a weak monoprotic acid; • a weak monoprotic acid with a strong monobase. <p>Equivalence point of the titration.</p> <p>Relationship between pH at the equivalence point of the titration and the choice of indicator.</p> <p>pH at half-equivalence point and its relationship with the pK_a of the weak acid or pK_b of the weak base.</p> <p>Plotting and using titration curves to derive:</p> <ul style="list-style-type: none"> • the pH at equivalence point; • the half equivalence point; • the initial pH. <p>1.6.- Behaviour of ionic compounds in solution Dissolving salts in water.</p> <p>Hydrolysis of salts. The importance of hydrolysis in acid/base titration (pH at equivalence point).</p> <p>1.7.- Buffer solutions Preparation of an acidic buffer solution. Equations to illustrate how an acidic buffer works. pH calculation of a buffer solution made from a monoprotic acid and its conjugate base.</p>	<p>Students should complete at least one titration, involving an acid-base indicator and a pH meter. The data collected should be plotted to show the titration curve.</p> <p>Salts dissolved in water are completely dissociated. Acid/base reactions of ions with water (hydrolysis). Acid/base behaviour of salts in solution (only the qualitative approach is required).</p> <p>A buffer solution is efficient in the pH range $pK_a \pm 1$. If the concentration of the salt is equal to the concentration of the acid, the pH of the resulting buffer is equal to the pK_a of the acid.</p>	<p>Titration of HCO_3^- (aq) from mineral water against hydrochloric acid.</p> <p>The case of weak polyacids and weak polybases could be treated.</p> <p>The use of a computer to process data, make calculations and draw graphs should be encouraged.</p>

CONTENT	COMMENTS	SUGGESTIONS
<p>2.- ELECTROCHEMISTRY</p> <p>2.1.- Redox reactions Redox reaction as electron transfer Determination of the oxidation numbers of elements in molecules and ions.</p> <p>Ionic and molecular redox reactions. Half-equations of redox reactions using oxidation numbers. Use of standard redox potentials</p> <p>2.2.- Electrochemical cells and standard electrode potential Conventional representation of an electrochemical cell. The standard hydrogen electrode Definition of the redox potential of a redox couple.</p> <p>Redox couple</p> <p>Use of standard redox potentials.</p> <p>Calculation of the emf of a cell under standard conditions Use of standard redox potentials.</p> <p>2.3.- Applications 2.3.1.- Redox titrations Calculations based on redox titrations (the redox couples and conditions will be given).</p>	<p>Students are expected to construct half-ionic equations for redox reaction when given the relevant couples and conditions. Students are expected to be able to reproduce the half equation of the reduction of MnO_4^-(aq) to Mn^{2+}(aq) in acidic conditions.</p> <p>Students are not expected to memorise the diagram of the standard hydrogen cell, but they must be able to label it.</p> <p>The difference between the electrical potential of the couple under standard conditions and the standard hydrogen electrode. E.g. Ag^+ (aq)/Ag (s), Cu^{2+} (aq)/Cu^+ (aq) Measurement of the emf of a couple under standard conditions against the standard hydrogen electrode.</p> <p>To predict the reaction between two redox couples when all the required information is given.</p> <p>In the calculations, the electrode potentials are considered to be equal to standard redox potentials. The Nernst equation is not required.</p> <p>The conditions, colour of the reagents and products, colour change occurring at the equivalence point should be known for the redox titration using acidified MnO_4^-(aq) as the oxidizing agent. The colour change when starch is used to determine the end-point of a redox titration, involving iodine, should be known.</p>	<p>Action of an acid on a metal [e.g. $\text{Mg(s)} + \text{HCl(aq)}$, ...] Action of a metal on a metal ion [e.g. $\text{Fe(s)} + \text{Cu}^{2+}$(aq), ...] Fe^{2+}(aq) + MnO_4^- + H_3O^+(aq) H_2O_2(l) + $\text{MnO}_4^-(\text{aq})$ [+ $\text{H}_3\text{O}^+(\text{aq})$] Oxalic acid + $\text{MnO}_4^-(\text{aq})$ $\text{I}_2(\text{aq}) + \text{S}_2\text{O}_3^{2-}(\text{aq})$ Ethanol + $\text{MnO}_4^-(\text{aq})$ or $\text{Cr}_2\text{O}_7^{2-}(\text{aq}) + \text{H}_3\text{O}^+(\text{aq})$...</p> <p>Experimental measurement of the standard electrode potential of a redox couple using a reference electrode.</p> <p>Redox titrations of: Fe^{2+}(aq) with MnO_4^-(aq) in an acidic solution H_2O_2(l) with $\text{MnO}_4^-(\text{aq})$ in an acidic solution Oxalic acid with $\text{MnO}_4^-(\text{aq})$ $\text{I}_2(\text{aq})$ with $\text{S}_2\text{O}_3^{2-}(\text{aq})$.</p>

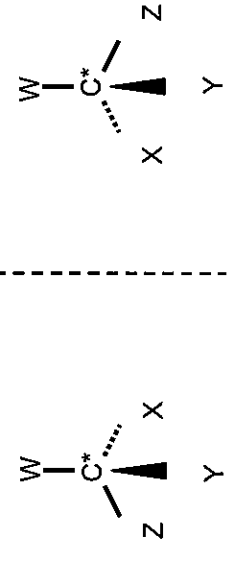
CONTENT	COMMENTS	SUGGESTIONS
<p>2.3.2.- The Chemical Principles of electrolysis</p> <p>2.3.3.- Electrolysis as an industrial process</p> <p>Prediction of the nature of the products obtained at each electrode.</p> <p>Calculations on electrolysis.</p> <p>2.3.4.- Cells Principles of cells and accumulators. The lead-acid accumulator Simple fuel cells</p> <p>2.3.5.- Corrosion as a redox process The half equation for the oxidation of iron in the process of corrosion. Protection of a metal from corrosion (electrochemical protection).</p>	<p>Carry out the electrolysis of simple molten and aqueous salts. Be able to draw and label a diagram to show an electrolysis experiment. Know that electrolysis is not a spontaneous chemical reaction. Electrolysis will only occur if the ions in the electrolyte are free to move. The positive ion moves towards the negative electrode. The negative ion moves towards the positive electrode. Be able to write ionic-half equations for the reaction that takes place at each electrode. As illustrated by the electrolysis of molten NaCl, aqueous NaCl, Al₂O₃ (dissolved in cryolite), and the refining of impure copper (only the redox principle of the transfer of copper from the positive electrode to the negative electrode is required). No technical details of any industrial process are required.</p> <p>Qualitative study of the factors which influence the nature of the ion discharged at one electrode, i.e.:</p> <ul style="list-style-type: none"> • standard redox potential of the relevant couple. • concentration of the ion in the solution as illustrated by a concentrated solution of sodium chloride compared to a dilute solution of sodium chloride. • nature of the electrode as illustrated by the use of Platinum, graphite, copper in the electrolysis of copper (II) sulphate]; <p>Using Faraday's Laws.</p> <p>The interpretation of reactions at the electrodes, but equations do not have to be memorised.</p> <p>The interpretation of reactions at the electrodes, but equations do not have to be memorised.</p> <p>$\text{Fe(s)} \rightarrow \text{Fe}^{2+}(\text{aq}) + 2\text{e}^{-}$</p> <p>Protection of iron from corrosion by making an electrochemical cell using Mg(s) as a reducing agent.</p>	<p>Electrolysis of: PbBr₂ (l), CuCl₂ (aq), CuBr₂ (aq)</p> <p>Galvanization and copper electroplating could be considered.</p>

CONTENT	COMMENTS	SUGGESTIONS
<p>3- ORGANIC CHEMISTRY</p> <p>3.1.- Alcohols Monohydric alcohols : primary, secondary and tertiary alcohols, isomers, nomenclature(up to C₆). Examples of polyhydric alcohols.</p> <p>3.1.1.- Physical properties of alcohols In relation to their structure.</p> <p>3.1.2.- Chemical properties of alcohols</p> <ul style="list-style-type: none"> • related to the breaking of the O-H bond; • related to the breaking of the C-O bond; • oxidation of primary and secondary alcohols; • comparison of the acidic properties of phenol and alcohols. <p>3.2.- Aldehydes and ketones</p> <p>3.2.1.- Nomenclature up to C₆, isomers.</p> <p>3.2.2.- Physical properties In relation to their structure.</p> <p>3.2.3.-Chemical properties</p> <ul style="list-style-type: none"> • oxidation of aldehydes; • distinguishing between aldehydes and ketones by oxidation; 	<p>For each of the following classes of organic compounds, the specified physical and chemical properties should be considered.</p> <p>Ethane-1,2-diol and propane-1,2,3-triol</p> <ul style="list-style-type: none"> • solubility in polar and non-polar solvents • melting point • boiling point <p>Reaction with sodium Reaction of esterification, including the mechanism and its evidence by ¹⁸O.</p> <p>Reaction with HBr (mechanism required). Dehydration (with H₂SO₄ or Al₂O₃) to form an ether or an alkene.</p> <p>Primary alcohol → aldehyde → carboxylic acid Secondary alcohol → ketone</p> <p>Acidic strength of phenols versus alcohols. Influence of inductive and mesomeric effects on the acidic properties of phenols and alcohols.</p> <p>Isomers of aldehydes and ketones, isomeric relationships</p> <ul style="list-style-type: none"> • solubility in polar and non-polar solvents • melting point • boiling point <p>Aldehydes can be easily oxidized, ketones cannot be oxidised without breaking the carbon chain. Aldehydes and ketones can distinguished using Fehling's solution. Half-equations : 2Cu²⁺(aq) + 2 OH⁻(aq) + 2e⁻ → Cu₂O(s) + H₂O(l) RCHO(aq) + 3OH⁻(aq) → RCOO⁻(aq) + 2H₂O(l) + 2 e⁻</p>	<p>Reaction of the isomers of butan-1-ol with sodium. Esterification of ethanol with ethanoic acid.</p> <p>Oxidation of the isomers of butan-1-ol with acidified MnO₄⁻(aq).</p> <p>Comparison of the physical and chemical properties of ethanal and propanone.</p>

CONTENT	COMMENTS	SUGGESTIONS
<ul style="list-style-type: none">• Nucleophilic addition of the carbonyl group.• Reduction of the carbonyl group	<p>Reaction with CN^-, including the mechanism. (This reaction should not be carried out practically). Limited to knowing that a ketone is reduced to a 2° alcohol and an aldehyde is reduced to a 1° alcohol.</p>	

CONTENT	COMMENTS	SUGGESTIONS
<p>3.2.4.-Carbohydrates</p> <ul style="list-style-type: none"> • Monosaccharide: glucose; • Disaccharide: maltose; • Polysaccharide: starch. 	<p>Open chain and ring structure: given one structure, the students should be able to explain how the other one forms. Solubility in water. Presence of an aldehyde group in the molecule in the open chain form. Oxidation of glucose using Fehling's solution.</p> <p>Condensation of two molecules of glucose to form the disaccharide maltose.</p> <p>Understand that the polymerisation of glucose to form starch is an example of a condensation reaction. The balanced equation is not required but the repeating unit should be represented.</p>	<p>The importance of glucose in the production of cellular energy.</p> <p>Silver mirror test or Fehling's solution with various carbohydrates.</p> <p>Complex between starch and iodine in aqueous solution.</p> <p>Reference to cellulose and glycogen.</p>

CONTENT	COMMENTS	SUGGESTIONS
<p>3.3.- Carboxylic acids</p> <p>3.3.1.- Nomenclature, isomers.</p> <p>3.3.2.- Physical properties In relation to their structure.</p> <p>3.3.3.-Chemical properties</p> <ul style="list-style-type: none"> • acidic character and strength. • esterification. <p>3.3.4.- Aspirin</p> <p>3.4.- Optical isomerism Chirality of a molecule.</p>	<p>Naming of mono- and dicarboxylic acids up to C₆. Monocarboxylic acids of C₁₆ and C₁₈.</p> <ul style="list-style-type: none"> • solubility in polar and non-polar solvents; • melting point; • boiling point. <p>Stabilisation of the carboxylate ion by mesomerism. Influence of the inductive effects of substituents on the strength of carboxylic acids and comparison of their K_a (pK_a).</p> <p>Synthesis and hydrolysis of an ester.</p> <p>Formation of polyesters. Obtained by the condensation of a diol and a dicarboxylic acid. The balanced equation is not required, but the repeating unit should be represented.</p> <p>Molecular structure of aspirin: the students are not expected to memorise the structure, but they should be able to identify the functional groups present in the molecule. Aspirin shares the properties of both carboxylic acids and esters.</p> <p>Presence of an asymmetric carbon, i.e. a carbon atom bonded to four different substituents (chiral molecule). A molecule with one asymmetric carbon (C*) can be represented in a 3-dimensional way, for example:</p> <div style="text-align: center;"> <p style="text-align: center;"> $\begin{array}{c} \text{W} \\ \\ \text{C}^* \\ \vdots \\ \text{Z} \end{array} \begin{array}{c} \text{X} \\ \text{Y} \end{array}$ </p> </div>	<p>Fatty acids (saturated and unsaturated).</p> <p>Formation of butyl ethanoate.</p> <p>Physiological properties of aspirin.</p> <p>butan-2-ol. Lactic acid.</p> <p>Biological importance of chirality (amino-acids, active sites of enzymes, membrane receptors...).</p>

CONTENT	COMMENTS	SUGGESTIONS
Enantiomers.	<p>Enantiomers are non superimposable mirror images of the same molecule.</p> 	
Effect of optically active compounds on polarized light.	<p>Optically active compounds rotate the plane of plane polarised light.</p> <p>Enantiomers rotate the plane of plane polarised light by the same amount but in opposite directions.</p>	
Racemic mixture.	<p>An equimolar mixture of the two enantiomers of a substance.</p>	

CONTENT	COMMENTS	SUGGESTIONS
<p>3.5.- Fats and oils .</p> <p>3.5.1.- Nomenclature..</p>	<p>Fats and oils are esters of propane-1,2,3-triol (glycerol) and long chain saturated and unsaturated fatty acids.</p> <p>Naming of fatty acids, limited to C₁₆ and C₁₈.</p>	<p>Dry cleaning.</p>
<p>3.5.2.- Physical properties of saturated and unsaturated fats and oils.</p>	<ul style="list-style-type: none"> • solubility in polar and non-polar solvents; • melting point; <p>Explain the difference in melting points of fats and oils in terms of their degree of saturation/unsaturation.</p>	<p>Preparation of margarine.</p>
<p>3.5.3.-Chemical properties</p> <ul style="list-style-type: none"> • hydrogenation; • iodination; • saponification. 	<p>Hydrogenation of unsaturated fats. The student should be able to write a balanced equation.</p> <p>Iodine number to determine the degree of unsaturation of a fat.</p> <p>Saponification of fats to produce soaps. The student should be able to write a balanced equation.</p> <p>Cleaning properties of soaps. Know that the chemical structure of a soap molecule - a long chain molecule with a polar (hydrophilic) and non-polar (hydrophobic) ends – and be able to explain why a soap cleans.</p>	<p>Preparation of a soap.</p> <p>Demonstrate practically.</p>
<p>Comparison of synthetic detergents and soaps</p>	<ul style="list-style-type: none"> • Chemical structure; • sensitivity to the hardness of water; • effects on the environment; • biodegradability. 	<p>Demonstrate practically.</p>

CONTENT	COMMENTS	SUGGESTIONS
<p>3.6.- Amines 3.6.1.- Nomenclature, isomers.</p> <p>3.6.2.-Chemical properties</p> <ul style="list-style-type: none"> • basic character of the amine group; • diazotization. <p>3.6.3.- Formation of the amide bond (peptide bond)</p> <ul style="list-style-type: none"> • formation of amides. <p>3.7.- Amino acids 3.7.1.- Nomenclature. Optical isomerism</p> <p>3.7.2.- Physical properties In relation to their structure.</p> <p>3.7.3.-Chemical properties</p> <ul style="list-style-type: none"> • amphoteric character; formation of the zwitterions. • formation of peptides and polypeptides; • Primary structure of proteins. 	<p>Naming of primary amines up to C₆. Be able to distinguish between primary, secondary and tertiary amines.</p> <p>Restricted to primary and secondary amines. Comparison of the basic strength of primary and secondary amines, ammonia and phenylamine in aqueous solution. Influence of substituents on the basic strength of amine (inductive and mesomeric effects) to explain the difference in K_b (pK_b) values.</p> <p>Azo dyes: reaction between diazonium chloride and phenol or 2-naphthol (β-naphthol) to form diazo dyes. Equations are not required. Diazo dyes are coloured substances because their molecules show extended conjugation of double bonds (responsible for the absorption of light in the visible part of the spectrum).</p> <p>The formation of the amide (or peptide) bond. Polyamides: obtained by a condensation reaction of a diamine and a dicarboxylic acid. The balanced equation is not required, but the repeating unit should be represented.</p> <p>Structure of a 2-amino acid (α-amino acid). IUPAC and trivial names of both glycine and alanine. All 2-amino acids exist as enantiomers except for glycine.</p> <ul style="list-style-type: none"> • solubility in polar and non-polar solvents; • melting point; <p>Reaction of amino acids with water, acids and bases to illustrate the amphoteric nature of amino acids.</p> <p>Students should know the meaning of "isoelectric point of an amino acid", but no further development is required.</p> <p>Condensation of amino acids (limited to three amino acids). The order of the sequence of amino acids in a protein.</p>	<p>Making nylon.</p> <p>The secondary, tertiary and quaternary structures of proteins: Their importance in the biological function of proteins and enzymes. Lab: Hydrolysis of aspartame.</p>