Scheme of Work
for the
Academic Year 2017/2018

SUBJECT : Chemistry
CLASS : 10
EXAMINATION : IGCSE
TEACHER IN CHARGE : Gateway College, Colombo

MONTH	UNIT No.	TOPICS	No. of PERIOD	OBJECTIVES
June/July	2 f	Acids, alkalis and titration	12	2.28 describe the use of litmus, phenolphthalein and methyl orange to distinguish between acidic and alkaline solutions lab 2.29 understand how to use the pH scale, from 0–14, can be used to classify solutions as strongly acidic (0–3), weakly acidic (4–6), neutral (7), weakly alkaline (8–10) and strongly alkaline (11–14) 2.30 describe the use of universal indicator to measure the approximate pH value of an aqueous solution 2.31 know that acids in aqueous solution are a source of hydrogen ions and alkalis in a aqueous solution are a source of hydroxide ions 2.32 know that alkalis can neutralise acids 2.33C describe how to carry out an acid-alkali titration
	2 g	Acids, alkalis and salt preparations		2.34 know the general rules for predicting the solubility of ionic compounds in water: • common sodium, potassium and ammonium compounds are soluble • all nitrates are soluble • common chlorides are soluble, except those of silver and lead(II) • common sulfates are soluble, except for those of barium, calcium and lead(II) • common carbonates are insoluble, except for those of sodium, potassium and ammonium • common hydroxides are insoluble except for those of sodium, potassium and calcium (calcium hydroxide is slightly soluble). 2.35 understand acids and bases in terms of proton transfer 2.36 understand that an acid is a proton donor and a base is a proton acceptor 2.37 describe the reactions of hydrochloric acid, sulfuric acid and nitric acid with metals, bases and metal carbonates (excluding the reactions between nitric acid and metals) to form salts 2.38 know that metal oxides, metal hydroxides and ammonia can act as bases, and that alkalis are bases that are soluble in water 2.39 describe an experiment to prepare a pure, dry sample of a

				soluble salt, starting from an insoluble reactant 2.40C describe an experiment to prepare a pure, dry sample of a soluble salt, starting from an acid and alkali 2.41C describe an experiment to prepare a pure, dry sample of an insoluble salt, starting from two soluble reactants 2.42 practical: prepare a sample of pure, dry hydrated copper(II) sulfate crystals starting from copper(II) oxide 2.43C practical: prepare a sample of pure, dry lead(II) sulfate
July	2.d	Reactivity series	12	2.15 understand how metals can be arranged in a reactivity series based on their reactions with: • water • dilute hydrochloric or sulfuric acid. 2.16 understand how metals can be arranged in a reactivity series based on their displacement reactions between: • metals and metal oxides • metals and aqueous solutions of metal salts. 2.17 know the order of reactivity of these metals: potassium, sodium, lithium, calcium, magnesium, aluminium, zinc, iron, copper, silver, gold 2.18 know the conditions under which iron rusts 2.19 understand how the rusting of iron may be prevented by: • barrier methods • galvanising • sacrificial protection. 2.20 understand the terms: in terms of gain or loss of oxygen and loss or gain of electrons. • oxidation • reduction • redox • oxidising agent • reducing agent 2.21 practical: investigate reactions between dilute hydrochloric and sulfuric acids and metals (e.g. magnesium, zinc and iron)
September	1.g	Covalent substances	12	1.49 explain why substances with giant covalent structures are solids with high melting and boiling points 1.50 explain how the structures of diamond, graphite and C60 fullerene influence their physical properties, including electrical conductivity and hardness 1.51 know that covalent compounds do not usually conduct electricity
	1 h	Metallic bonding		1.52C know how to represent a metallic lattice by a 2-D diagram 1.53C understand metallic bonding in terms of electrostatic attractions 1.54C explain typical physical properties of metals, including electrical conductivity and malleability
October/November	1.e	Chemical calculations	12	1.26 calculate relative formula masses (including relative molecular masses) (Mr) from relative atomic masses (Ar) 1.27 know that the mole (mol) is the unit for the amount of a substance 1.28 understand how to carry out calculations involving amount of substance, relative atomic mass (Ar) and relative

				formula mass (Mr) 1.29 calculate reacting masses using experimental data and chemical equations 1.30 calculate percentage yield 1.31 understand how the formulae of simple compounds can be obtained experimentally, including metal oxides, water and salts containing water of crystallization 1.32 know what is meant by the terms empirical formula and molecular formula 1.33 calculate empirical and molecular formulae from experimental data 1.34C understand how to carry out calculations involving amount of substance, volume and concentration (in mol/dm3) of solution 1.35C understand how to carry out calculations involving gas volumes and the molar volume of a gas (24 dm3 and 24 000 cm3 at room temperature and pressure (rtp)) 1.36 practical: know how to determine the formula of a metal oxide by combustion (e.g. magnesium oxide) or by reduction (e.g. copper(II) oxide)
January	1.h	Electrolysis	12	1.55C understand why covalent compounds do not conduct electricity 1.56C understand why ionic compounds conduct electricity only when molten or in aqueous solution 1.57C know that anion and cation are terms used to refer to negative and positive ions respectively 1.58C describe experiments to investigate electrolysis, using inert electrodes, of molten compounds (including lead(II) bromide) and aqueous solutions (including sodium chloride, dilute sulfuric acid and copper(II) sulfate) and to predict the products 1.59C write ionic half-equations representing the reactions at the electrodes during electrolysis and understand why these reactions are classified as oxidation or reduction 1.60C practical: investigate the electrolysis of aqueous solutions
Januban/February	2.e	Extraction and uses of metals	8	2.22C know that most metals are extracted from ores found in the Earth's crust and that unreactive metals are often found as the uncombined element 2.23C explain how the method of extraction of a metal is related to its position in the reactivity series, illustrated by carbon extraction for iron and electrolysis for aluminium 2.24C be able to comment on a metal extraction process, given appropriate information detailed knowledge of the processes used in the extraction of a specific metal is not required 2.25C explain the uses of aluminium, copper, iron and steel in terms of their properties the types of steel will be limited to low-carbon

				(mild), high-carbon and stainless 2.26C know that an alloy is a mixture of a metal and one or more elements, usually other metals or carbon 2.27C explain why alloys are harder than pure metals
February	2.c	Gases in the atmosphere	8	2.9 know the approximate percentages by volume of the four most abundant gases in dry air 2.10 understand how to determine the percentage by volume of oxygen in air using experiments involving the reactions of metals (e.g. iron) and non-metals (e.g. phosphorus) with air 2.11 describe the combustion of elements in oxygen, including magnesium, hydrogen and sulfur 2.12 describe the formation of carbon dioxide from the thermal decomposition of metal carbonates, including copper(II) carbonate 2.13 know that carbon dioxide is a greenhouse gas and that increasing amounts in the atmosphere may contribute to climate change 2.14 practical: determine the approximate percentage by volume of oxygen in air using a metal or a non-metal
March	2.h	Chemical test	8	2.44 describe tests for these gases: • hydrogen • oxygen • carbon dioxide • ammonia • chlorine. 2.45 describe how to carry out a flame test 2.46 know the colours formed in flame tests for these cations: • Li+ is red • Na+ is yellow • K+ is lilac • Ca2+ is orange-red • Cu2+ is blue-green. 2.47 describe tests for these cations: • NH4+ using sodium hydroxide solution and identifying the gas evolved • Cu2+, Fe2+ and Fe3+ using sodium hydroxide solution. 2.48 describe tests for these anions: • Cl–, Br– and I– using acidified silver nitrate solution • SO42– using acidified barium chloride solution • CO32– using hydrochloric acid and identifying the gas evolved. 2.49 describe a test for the presence of water using anhydrous copper(II) sulfate 2.50 describe a physical test to show whether a sample of water is pure