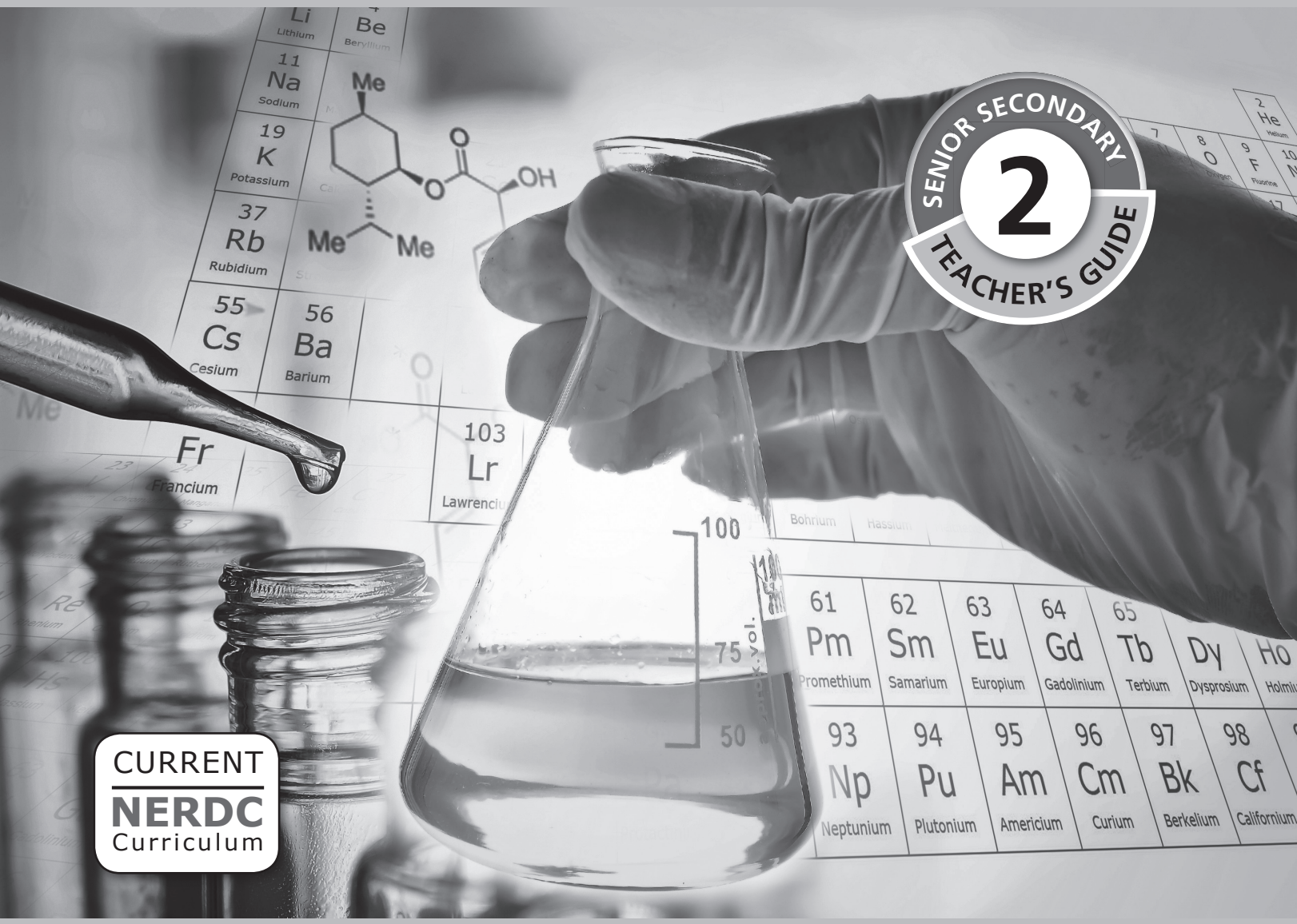
Excellence in Chemistry



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Chemistry

SENIOR SECONDARY 2 TEACHER'S GUIDE

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Introduction

The purpose of the curriculum

The general objectives of the curriculum are to enable students to:

- develop an interest in the subject of chemistry
- acquire basic theoretical and practical knowledge and skills in chemistry
- develop an interest in science, technology and mathematics
- acquire basic STM knowledge and skills
- be positioned to take advantage of the numerous career opportunities offered by chemistry
- be adequately prepared for further studies in chemistry.

The goals

The goals of the reviewed chemistry curriculum ensure that the teaching of chemistry will:

- facilitate a smooth transition in the use of scientific concepts and techniques acquired in the new Basic Science &Technology curriculum with chemistry
- provide students with the basic knowledge in chemical concepts and principles through efficient selection of contents and sequencing
- show chemistry and its inter-relationship with other subjects
- show chemistry and its link with industry, everyday life activities and hazards
- provide a course that is complete for students not proceeding to higher education, while at the same time providing a reasonably adequate foundation for a post-secondary school chemistry course.

Time allocation

To cover this curriculum, the recommended weekly time allocation is 120 minutes each. Students need to do regular revision at home in order to cope with the content and new terminology.

The role of the teacher

One of the principle duties of a Chemistry teacher is to prepare and present good lessons to his or her students. The teacher has to:

- be as well informed as possible on the scheme of work of the subject
- know the aims and objectives of each topic
- select appropriate content materials
- decide on the best methods of presentation such as slide shows, workstations, videos, discussion groups, worksheets, question–answer sessions, debate, and experiments
- gather equipment and other resources required for the activities
- keep informed about new developments and news in chemistry in Nigeria and the rest of the world
- arrange outings and guest speakers from time to time.

To be effective in presentation, the teacher must do a written/typed plan for each lesson. This must include aims, objectives, resources, time frames, content for the lesson, activities, homework, assessment, and ideas and additional worksheets to cater for students requiring extension or learning support (remedial).

Prepare each topic in advance. Many teachers go into the classroom inadequately prepared. It is your responsibility as a Chemistry teacher to actively involve your students in the learning process. It is a proven

fact that students learn far more by doing than by listening.

You should endeavour to apply the scientific method throughout. Science involves being curious and asking questions. Wherever possible ask questions to engage the students and to encourage independent thought processes. Start your lessons by asking the students to write down answers to questions related to your lesson (approximately five). This will settle them into the lesson. You can use different types of questions in your lessons:

- for enabling you to determine prior knowledge on the topic (diagnostic)
- for consolidation of challenging concepts during the lesson
- for stimulation of interest in the subject
- for concluding the lesson.

This will assist you to find out whether students have understood the concepts and terminology in the lesson. It will also highlight any areas that they need to revise at home or for you to revisit in the next lesson.

Teachers must ensure that they do not appear to have favourites in the class, so devise a system to ensure that you ask questions fairly, but be careful not to embarrass weak students if they cannot answer questions.

How to use the book

The purpose of this *Teacher's Guide* is to assist you to use the *Student's Book* and be more thoroughly prepared so that your teaching is be more meaningful to your students. This book supports a hands-on approach and lays a solid foundation for SS2 and SS3.

You need to be familiar with the key features of the *Student's Book*.

The book is divided into 10 topics. Each topic is structured in the following way:

- performance objectives required by the curriculum
- content required by the curriculum
- activities to be completed individually, with a partner, in groups or as a class
- summary of the topic
- key words, which are essential vocabulary for the topic
- revision questions for revision of the topic.

How to use the scheme of work

A scheme of work is defined as the part of the curriculum that a teacher will be required to teach in any particular subject. Its primary function is to provide an outline of the subject matter and its content, and to indicate how much work a student should cover in any particular class. A scheme of work allows teachers to clarify their thinking about a subject, and to plan and develop particular curriculum experiences that they believe may require more time and attention when preparing lessons. The criteria all teachers should bear in mind when planning a scheme of work are continuity in learning and progression of experience. You can add your own notes to the scheme of work provided on pages viii to x.

The scheme of work is sequential. The sequence of the scheme of work is aligned with the textbook. Do not be tempted to jump around. Rather spend time carefully planning the term to ensure that you adhere to the scheme of work.

The year is divided into 3 terms. Each term is divided into 13 weeks. There are 4 topics in Term 1, 3 topics in Term 2 and 3 topics in Term 3. The end of term allows time for revision and an examination. This time frame may vary depending on the planning of your particular school.

The right-hand column gives the number of suggested lessons for each topic. This has been divided according to the content of the topic. These vary from 2 to 5 lessons per topic.

The first lesson is usually an introduction to the topic. Make an effort to make this lesson exciting and informative. You should always explain the meaning of the topic in this lesson, for example: What is chemistry? How can we use chemistry in daily life? What is the scientific method?

The last lesson is allocated to revision. In this lesson you can give the class a revision worksheet, a test or design a fun activity, such as a game or a quiz, to consolidate the topic. Students can also do their own revision by making mind maps, concept maps or other types of summaries. They can also set tests for each other.

It is important to note that these are a suggested number of lessons for the topic. The amount will vary according to the ability of the students in your class and their prior knowledge. Your management of the class will have an enormous influence on your ability to adhere to the time frames. Focus on effective discipline strategies. You will have less discipline issues if you are: punctual, well prepared, follow a plan (write this on the board at the start of the lesson), keep your word (do not make empty threats), consistently adhere to rules, especially rules related to laboratory safety, and strive to make Chemistry an exciting subject.

A teacher of Chemistry is a professional instructor who facilitates, promotes and influences students to achieve the outcomes of the scheme of work. It is the wish of the authors that the students will, at the end of each course in the series (SS1, SS2 and SS3) attain a level of proficiency in chemistry that will equip them for future studies in this field.

Scheme of work

Teaching scheme of work for SS2

Term	Week	Theme	Topic	Performance Objectives (students should be able to:)	Content/Suggested number of lessons	SB page no
Term 1	1	The Chemical World	1. Periodic Table	• Explain the periodic law • Arrange common elements into groups (families) and periods • Distinguish between the families of elements on the Periodic Table • Discuss the changes in the properties of elements down the group and across periods • Discuss the relationship between ionization energy and electron affinity and the properties of elements down the groups and across periods • Explain the diagonal relationship in the properties of elements	1. Periodic Law 2. Blocks of elements: metals, non-metals, metalloids and transition metals 3. Families: s-p-d-f- (according to groups I – VIII, i.e. Group IA – alkali metals, Group IIA – alkali earth and other family names 4. Properties: changes in sizes and change down the group and across periods and accompanying changes in properties 5. Diagonal relationships 6. Ionization energy and electron affinity changes down the group and across the period	1 3 5 13 17 18
	2/3		2. Chemical reactions	• Identify reactants and products of any chemical reaction • Explain the terms reaction time and reaction rate; and the relationship between the two • Explain collision theory with respect to reaction time and reaction rate • Describe the influence of the following on chemical reaction rates: nature of substances; concentration/pressure, temperature and catalysts • Explain endothermic and exothermic reactions • Illustrate by use of a graph, the energy changes in exothermic and endothermic reactions • Write equations for simple equilibrium reactions • State Le Châtelier's principle • Explain the influence of the following factors on the equilibrium of chemical reactions: concentration, temperature, pressure	1. Basic concepts: reactants, products, reaction time and reaction rate 2. Introduction to collision theory 3. Factors affecting the rate of chemical reactions: nature of substances; concentration/ pressure, temperature, catalyst 4. Types of chemical reactions: endothermic and exothermic 5. Chemical equilibrium: introduction using simple equations; Le Châtelier's principles 6. Factors affecting equilibrium of chemical reactions: concentration, temperature, pressure	22 24 26 31 39 43

Term	Week	Theme	Topic	Performance Objectives (students should be able to:)	Content/Suggested number of lessons	SB page no
Term 1	4/5	Chemistry and Environment	3. Mass, volume relationships	• Explain the concepts of the mole, molar, s.t.p., relative densities and relative molar mass (RMM) • Solve problems involving reacting masses and volumes in chemical reactions • State the SI units of various basic quantities	1. Explains the concepts of mole, molar, s.t.p., relative densities and relative molecular mass, and their units 2. Guides students to calculate: masses of reactants and products, number of moles of reacting substances and products 3. SI units of quantities, for example length, mass, volume, etc.	50 59 65
	6/7		4. Acid-base reactions	• Define concentration, in mol·dm^{-3} of solutions • Define standard solutions • Explain relationship between concentrations and volumes of reacting substances • Mathematically express the relationship between the concentration in mol·dm^{-3} and volume of a solution • Carry out acid-base titrations using appropriate indicators • Record correctly titer values to two decimal places • Carry out relevant calculations from titration results	1. Simple acid-base titrations 2. Common indicators and their pH ranges 3. Heat of neutralisation (introductory)	73 71 79
	8/9		5. Water	• Draw and explain the structure of water • Define the following concepts: solute, solvent, solution • Define solubility and state the rules of solubility in water • Explain factors that affect solubility • Explain the causes of hardness of water • Explain the methods employed in the removal of hardness • Explain methods used in purifying water	1. Structure of water 2. Solubility (basic concepts: solute, solvent, solution) 3. Solubility of different substances 4. Factors that affect solubility/ uses of solubility curves 5. Hardness of water and removal of hardness 6. Purification of water 7. Municipal water supply 8. Production of distilled water	83 84 87 89 93 94 95 95
	10		6. Air	• List the constituents of air • List the percentage composition of air • State the properties of air • Draw, label and explain the various zones of flame	1. Air: constituents, percentage composition 2. Properties of air 3. Flame	102 106 109
	11	Revision				
	12/13	Examination				

Term	Week	Theme	Topic	Performance Objectives (students should be able to:)	Content/Suggested number of lessons	SB page no
Term 2	1/2		7. Hydrogen	• Write and draw the electron configuration of hydrogen • Identify the isotopes of hydrogen and draw their electron configuration • Explain the unique position of hydrogen in the Periodic Table • Describe the laboratory and industrial methods for the preparation of hydrogen • State the physical and chemical properties of hydrogen • State the uses of hydrogen	1. Hydrogen: configuration and possible oxidation numbers; isotopes of hydrogen; unique position of hydrogen in the Periodic Table 2. Laboratory preparation of hydrogen 3. Industrial preparation of hydrogen 4. Physical and chemical properties of hydrogen 5. Uses of hydrogen	113 114 114 117 118
	3/4		8. Oxygen	• Explain the general properties of oxygen group in the Periodic Table • Write and draw the electronic structure of oxygen and explain its bonding capacity • Describe the laboratory and industrial methods for the preparation of oxygen • State the physical and chemical properties of oxygen • List the compounds of oxygen • Explain oxidation as addition of oxygen and give examples of such reactions • State the uses of oxygen	1. General properties of oxygen group 2. The electronic structure and bonding capacity of oxygen 3. Laboratory and industrial preparation of oxygen 4. Physical and chemical properties of oxygen 5. Reaction of oxygen (oxidation) 6. Uses of oxygen	121 123 124 127 127 133
	5/6		9. Halogens	• Write the electronic configuration of the halogens (Cl_2, Br_2, I_2) • State the physical properties of halogens and their gradation down the group • State the chemical properties of halogens and their gradation down the group • List some compounds of halogens • State the uses of halogens • Describe the laboratory preparation of chlorine • Demonstrate the bleaching action of chlorine and adduce the reason for the bleaching action	1. Electronic configuration of halogens 2. Physical properties of halogens and gradation down the group 3. Chemical properties of halogens and gradation down the group 4. Compounds of halogens 5. Uses of halogens (Cl_2, Br_2, I_2) 6. Preparation of chlorine	139 140 141 145 145 147
	7/8		10. Nitrogen	• State the general properties of group VA elements • Explain the laboratory preparation of nitrogen • Explain the industrial preparation of nitrogen from liquid air • List the properties of nitrogen • Outline the uses of nitrogen • List the oxides of nitrogen • Explain the nitrogen cycle • Explain the Haber process for the preparation of ammonia • State the uses of ammonia	1. General properties of the nitrogen family – group VA elements 2. Laboratory preparation of nitrogen 3. Industrial preparation of nitrogen from liquid air 4. Properties of nitrogen 5. Uses of nitrogen 6. Nitrogen cycle 7. Compounds of nitrogen: oxides of nitrogen; ammonia 8. Test for ammonia	152 155 155 156 156 157 158 162

Term	Week	Theme	Topic	Performance Objectives (students should be able to:)	Content/Suggested number of lessons	SB page no
	9/10		11. Sulphur	• State the general properties of group VIA elements • Write and draw the electron configuration of sulphur • Explain the meaning allotropy • Identify the allotropes of sulphur • State the uses of sulphur • Name some common compounds of sulphur • Determine the oxidation states of sulphur in its major compounds • Describe the industrial preparation of H_2SO_4 by the contact process • State the uses of H_2SO_4	1. General properties of group VIA elements 2. Electron structure of sulphur 3. Allotropes of sulphur 4. Uses of sulphur 5. Compounds of sulphur 6. Industrial preparation of H_2SO_4 7. Uses of H_2SO_4	166 167 168 169 169 170 171
	11	Revision				
	12/13	Examination				

Term	Week	Theme	Topic	Performance Objectives (students should be able to:)	Content/Suggested number of lessons	SB page no
Term 3	1	Chemistry and Industry	12. Oxidation-reduction (redox) reactions	• Define oxidation as: addition of oxygen, removal of hydrogen, process of electron loss, process of increase of oxidation number of a substance • Define reduction as the reverse of any of the above processes • Calculate oxidation numbers of elements using a set of arbitrary rules, viz (a) oxidation number of free elements = 0 (b) oxidation number of oxygen in any compound is –2, except in peroxides where it is –1 (c) oxidation number of H is +1, except in hydrides where it is –1 (d) oxidation number of a neutral molecule or compound is zero, for example H_2SO_4 = 0, etc. • Use oxidation numbers to name inorganic compounds, to include number of oxygen atom and water molecules (if hydrated) • Determine the oxidation states number of common elements in their compounds • Define oxidizing and reducing agents in terms of: addition and removal of oxygen and hydrogen respectively; loss and gain off electron; change in oxidation numbers/states • Write and balance redox reactions	1. Oxidation: definitions 2. Reduction: definitions 3. Redox reactions 4. Oxidation numbers of central elements in some compounds 5. Connection of oxidation numbers with IUPAC name 6. Oxidizing and reducing agents 7. Redox equations	174 175 176 176 178 180 185

Term	Week	Theme	Topic	Performance Objectives (students should be able to:)	Content/Suggested number of lessons	SB page no
	2/3		13. The ionic theory	• Explain the difference between: electrovalent and covalent compounds; electrolytes and non-electrolytes • Investigate the movement of ions in solution • Distinguish between strong and weak electrolytes • Rank and explain the position of ions in the electrochemical series • Relate the order of ions in the electrochemical series to their rate of discharge from solution • State the factors affecting the preferential discharge of ions	1. Electrovalent and covalent compounds 2. Electrolytes and non-electrolytes 3. Weak and strong electrolytes 4. Electrochemical series 5. Factors affecting the preferential discharge of ions	193 194 198 198
Term 3	4/5		14. Electrolysis	• Explain the quantitative aspects of electrolysis • Define electrolytes (strong, weak, fused/molten, non-electrolytes), electrolytic and electrochemical cells • Differentiate between strong and weak electrolytes • Illustrate the electrolysis of acidified water, copper(II) sulfates and brines • Identify factors affecting the discharge of ions during electrolysis • Construct electrolytic and electrochemical cells • State Faraday's laws of electrolysis • Calculate amount of substances liberated or deposited at electrodes during electrolysis • Explain the uses of electrolysis in extraction and purification of metals	1. Meaning of electrolysis 2. Terminologies: electrodes; electrolyte; electrolytic cell; electrochemical cells, etc. 3. Electrolysis of acidified water, copper(II) sulfates ($CuSO_4$) and brine 4. Faraday's Laws of electrolysis and the calculations 5. Uses of electrolysis: purification, extraction and electroplating of metals	205 206 211 216 218

Term	Week	Theme	Topic	Performance Objectives (students should be able to:)	Content/Suggested number of lessons	SB page no
Term 4	6/7	Chemistry of life	15. Hydrocarbons	• Explain the structure of carbon and its valency • Define hydrocarbon • Give examples of hydrocarbons and their structure • Explain isomerism and give examples • Explain homologous series as it relates to the physical and chemical properties of hydrocarbons • Distinguish between aliphatic and aromatic hydrocarbons	1. Structure and valency of carbon 2. Hydrocarbon: meaning and examples 3. Homologous series (characteristics and naming – IUPAC) 4. Saturated hydrocarbons: composition and structure 5. Isomerism 6. Unsaturated hydrocarbon (composition and structure)	225 226 228 241 244 238
	8		16. Alcanols	• Relate the structure of alkanes to that of alkanols • Identify –OH as the functional group in alkanols • Explain the increase in boiling point of alkanols, compared with hydrocarbons, in terms of hydrogen bonding • Explain the polar nature of alcohols and its effect on the solubility of substances • Determine solubility of common materials in water and alcohols • Identify di-, tri- and polyhydroxy compounds by their structures and name them appropriately	1. Types of alkanols and their properties 2. Preparation and properties of alkanols 3. Industrial production of alkanols by fermentation 4. Uses of alkanols	250 252 258 259
	9/10		All practicals			
	11		Revision			
	12		Examination			

THEME 1

The chemical world

Topic 1: Periodic Table

Performance objectives

1.1 explain the Periodic Law
1.2 arrange common elements into groups (families) and periods
1.3 distinguish between the families of elements on the Periodic Table
1.4 discuss the changes in the properties of elements down the group and across periods
1.5 discuss the relationship between ionisation energy and electron affinity and the properties of elements down the groups and across periods
1.6 explain the diagonal relationship in the properties of elements.

Introduction

In this topic we will look at how the Periodic Table was developed and what is meant by the Periodic Law. We will see how common elements are arranged into groups (families) and periods in the Periodic Table. The Periodic Table is divided into different blocks of metals, non-metals, metalloids and transition metals and we will look at how these elements display distinguishing characteristics. We will also study the different trends in the Periodic Table regarding atomic and ionic sizes, gting points and boiling points, density, electronegativity, ionisation energy and electron affinity. We will discuss the gradation (trends) in properties of elements down a group and across a period (horizontal and vertical relationships) and also explain the diagonal relationships in properties of elements in the Periodic Table.

Activity 1.1: Periodic Law (SB p. 3)

Resources

Periodic Table of elements, blank periodic table template

Guidelines

Guide students to deduce the periodic law and identify and group elements into families based on shared characteristics.

Students should be able to explain the Periodic Law.

Answers

1. He classified elements into groups of threes.
2. Mendeleev's classification of elements was on the basis of increasing atomic weights while Moseley's classification was on the basis of increasing atomic numbers.
3. The Modern Periodic Law states that the physical and chemical properties of elements are a periodic function of their atomic numbers

Activity 1.2: Blocks of elements (SB p. 5)

Resources

Periodic Table of elements, blank periodic table template

Guidelines

Guide students to deduce the periodic law.

Students should be able to group common elements into families (groups) and periods.

Answers

1. Atomic numbers
2. a) A period is a horizontal row of elements in the periodic table
   b) A group is a vertical column in the periodic table
3.

Element name	Symbol	Group number	Period number	Atomic number
Lithium	Li	1	2	3
Aluminium	Al	3	3	13
Sulphur	S	6	3	16
Helium	He	8	1	2
Calcium	Ca	2	4	20
Silicon	Si	4	3	14
Nitrogen	N	5	2	7
Fluorine	F	7	2	9

4. a) He; Helium
   b) Si; Silicon
   c) B; Boron

Activity 1.3: Investigating the properties of Group I elements (SB p. 5)

Resources

Periodic Table of elements, blank periodic table template

Guidelines

Guide students to arrange elements on the periodic table and identify blocks of elements.

Students should be able to state, with examples, the distinguishing characteristics of the following blocks of elements on the table: metals, non-metals, metalloids, and transition metals

Answers

1. Physical properties
   - They are soft and dull. If you cut the metal they appear shiny but goes dull very quickly.
   - Potassium is soft, Sodium is softer and Lithium is the softest.
   - These metals are softer than iron and copper. On contact with air they immediately appear dull (not shiny anymore).
   - These metals are stored under oil because they are highly reactive with water.
2. Chemical properties
   - They less dense than water because they float
   - Lithium is the least reactive and Potassium is the most reactive
   - Reactivity decreases as you move down the group
   - Red litmus paper turns blue

	Lithium	Sodium	Potassium
Hardness	Softest	Softer	Soft
Reactivity	Least reactive	More reactive than Lithium	More reactive than Lithium and Sodium

3. Predict the properties of caesium:
   a) It will be harder
   b) More reactive. Reactivity increases as you move down the group I in the periodic table.

Activity 1.4: Families of elements (SB p. 10)

Resources

Periodic Table of elements, blank periodic table template

Guidelines

Guide a class discussion on families of elements.

Students should be able to classify families of elements by completing characteristics of the elements in the tables on p. 8 of the SB.

Activity 1.5: Do research (SB p. 10)

Resources

Periodic Table of elements; Internet; pictures from magazines or newspapers

Guidelines

Guide a class discussion on the characteristics of the families of elements.

Students should be able to do individual research on the element of their choice and report back by writing a page about the element.

Answers

Students should read about the names and symbols of the first 20 elements of the Periodic Table.

Each student should choose one element and a write a page about it.

Answers

Element	Group number	Group name	Phase (gas, liquid or solid)	Valency	Conductor (yes or no)	Compare reactivity	Metal/ non-metal
Sodium	1	Alkali metals	Solid	1	Yes	Highly reactive	Metal
Magnesium	2	Alkaline earth metals	Solid	2	Yes	Less reactive than group I metals	Metal
Chlorine	7	Halogens	Gas	1	No	Less reactive than fluorine in group VII	Non-metal
Neon	8	Noble gases	Gas	0	No	Inert	Non-metal

Activity 1.6: Use the Periodic Table (SB p. 10)

Resources

Periodic Table of elements

Guidelines

Guide a class discussion on atomic and ionic sizes.

Students should be able to discuss the changes in the properties of elements down the group and across periods.

Answers

Element symbol	Element name	Atomic number	Metal/ non-metal
Na	Sodium	11	Metal
Ne	Neon	10	Non-metal
B	Boron	5	Semi-Metal
O	Oxygen	8	Non-metal
F	Fluorine	9	Non-metal
Mg	Magnesium	12	Metal
Ar	Argon	18	Non-metal

Activity 1.7: The s-, p-, d- and f-blocks

(SB p. 12)

Resources

Periodic Table of elements

Guidelines

Guide a class discussion on the different blocks of elements on the Periodic Table.

Students should be able to identify the different blocks of elements.

Answers

1. B
2. A
3. C
4. D
5. C
6. B
7. C

Experiment 1.1: Reactivity of Period 3 elements

(SB p. 15)

Resources

Glass trough, beakers, sodium metal, magnesium ribbon, aluminium metal (foil), water

Guidelines

Initiate and guide a class discuss on ionization energy and electron affinity.

Students should be able to perform the experiment and record their results and conclusions.

Answers

The results of the experiment might differ from student to student.

Activity 1.8: Investigate changes in properties

(SB p. 16)

Resources

Periodic Table of elements

Guidelines

Guide a class discussion on the gradation in properties of elements down the groups and across periods.

Students should be able to illustrate the changes in properties of elements down the groups and across periods.

Answers

1. B
2. a) The atomic radius is half the distance between two nuclei in adjacent metal atoms, or half the distance between the nuclei of two atoms in a diatomic molecule.
   b) Ionic size is the distance between the nucleus of the charged particle and the outermost energy shell.
   c) Electronegativity is the ability of an atom in a compound to attract the bonding electron pair(s).
3. The atomic radii of Group I elements is greater than that of Group II. The atomic radius decreases across a period. As we move along the period from left to right, the number of inner electrons remains constant while the nuclear charge increases. The valence electrons are added in the same energy level, and the distance between the electrons and nucleus is the same. With each consecutive electron and proton that is added, the attractive force between the nucleus and electrons becomes stronger and the radius of the atom decreases.

   The ionic radii of the elements in Group I is greater than that of Group II. As you move across a period from left to right of the Periodic Table, the ionic radius decreases for metals forming cations, as the metals lose their outer electron orbitals. Cations (positive ions)

are smaller than their respective atoms. The number of protons is more than the number of electrons and the protons can pull the fewer electrons toward the nucleus more tightly. If the electron that is lost is the only valence electron, so that the electron configuration of the cation is like that of a noble gas, then an entire energy level is lost.

4. a) Hydrogen, Lithium, Sodium and Potassium
   b) Helium, Neon and Argon
   c) Group I elements have the largest atomic radius in a period and the Group VIII elements have the smallest atomic radius. As the atomic numbers increase so does the atomic radius. The atomic radius increases as we go down a group. The number of energy levels increases as we go down a group, and so the size of every subsequent element of a group increases. The filled energy levels between the nucleus and valence electrons shield the valence electrons from the nucleus, and the attractive force between the nucleus and valence electrons is weaker.
   The atomic radius decreases across a period. As we move along the period from left to right, the number of inner electrons remains constant while the nuclear charge increases. The valence electrons are added in the same energy level, and the distance between the electrons and nucleus is the same. With each consecutive electron and proton that is added, the attractive force between the nucleus and electrons becomes stronger and the radius of the atom decreases.
5. a) lithium, sodium, potassium
   b) nitrogen, oxygen, fluorine
   c) iodine, bromine, chlorine
6. The gting and boiling points of the elements in Group I to Group IV increase from left to right in the periodic table. If the attractive forces between the particles in a solid are strong, the gting point will be high; if the forces are weak, the gting point is low. Thus the attractive forces between the particles increases as you move from left to right across a period in the periodic table.
7. a) lithium, beryllium, boron
   b) magnesium, calcium, barium
   c) chlorine, bromine, iodine
   d) sodium potassium, caesium

Activity 1.9: Investigate diagonal relationships

(SB p. 18)

Resources

Periodic Table of elements

Guidelines

Initiate and guide a class discussion on the diagonal relationships.

Students should be able to discuss and describe the diagonal relationships in properties of elements in the periodic table.

Answers

1. C
2. C
3. A
4. B

Activity 1.10: Investigating ionisation energy and electron affinity

(SB p. 20)

Resources

Periodic Table of elements

Guidelines

Initiate and guide a class discussion on ionisation energy and electron affinity.

Students should be able to define the terms ionisation energy and electron affinity.

Answers

1. a) The energy required to remove the outermost (highest energy) electron from a neutral atom in its ground state.
   b) Electron affinity is the energy change that occurs when an electron is

accepted by an atom in the gaseous state to form an anion.

2. a) As you move down a group, first ionization energy decreases. Electrons are further from the nucleus and thus it is easier to remove the outermost one. Inner electrons at lower energy levels essentially block the protons force of attraction toward the nucleus. It therefore becomes easier to remove the outer electron.

   b) As you move across a period, first ionization energy increases. As you move across a period, the atomic radius decreases, that is, the atom is smaller. The outer electrons are closer to the nucleus and more strongly attracted to the centre. Therefore, it becomes more difficult to remove the outermost electron.

3. A
4. D

Revision questions: Answers

(SB p. 21)

1. a) C
   b) Helium (He) has the highest ionization energy because, like other noble gases, helium's valence shell is full. Therefore, helium is stable and does not readily lose or gain electrons.
2. True
   Atomic radius increases from right to left on the periodic table. Therefore, nitrogen is larger than oxygen.
3. a) The element at the top of the one group has some similarities with an element in the next group.
   b) Be and Al or B and Si
4. Bromine (Br)
   In non-metals, gting point increases down a column. Because chlorine and bromine share the same group, bromine possesses the higher gting point.
5. Sulphur (S)
   Note that sulphur and selenium share the same group. Electronegativity increases up a group. This indicates that sulphur is more electronegative than selenium.
6. Most noble gases have full valence shells. Because of their full valence electron shell, the noble gases are extregy stable and do not readily lose or gain electrons.
7. a) Ba
   b) Be
   c) Ca
   d) Sr
   e) Mg
8. Boron (B), Phosphorous (P), Sulphur (S), Fluorine (F) Electron affinity generally increases from left to right and from bottom to top.
9. A. Oxygen (O)
   Periodic trends indicate that atomic radius increases up a group and from left to right across a period. Therefore, oxygen has a smaller atomic radius sulphur.
10. False
    The reasoning behind this lies in the fact that a metal usually loses an electron in becoming an ion while a non-metal gains an electron. This results in a smaller ionic radius for the metal ion and a larger ionic radius for the non-metal ion.

How are you doing?

Take this opportunity to ask learners if there is anything that they do not understand. You can check their understanding by asking them some questions about the information covered in the unit. Explain anything that learners do not understand.

Key words

Periodic Table – an arrangement of the chemical elements based on their atomic numbers and chemical properties: 2

group – a vertical column of the Periodic Table: 3

period – a horizontal row of the Periodic Table: 3

alkali metals – the Group 1 elements: 5

alkaline earth metals – the Group II elements: 5

halogens – the non-metallic elements in Group VII: 5

noble gases – the non-metallic elements in Group VIII: 5

semi-metals (metalloids) – elements with properties between those of metals and non-metals: 4

transition metals – elements between Groups II and III on the Periodic Table: 4

valence electrons – electrons in the outer energy level: 11

ionisation energy – the amount of energy required to remove an electron from a neutral atom: 1

electron affinity – is the ability of an atom to accept one or more electrons: 1

Checklist for Self-evaluation

Theme 1 Topic 1

EVALUATION GUIDE: Student should be able to:

	Criteria	4	3	2	1	Comments
1.1	state the periodic law					
1.2	group common elements into families (groups) and periods					
1.3	state with examples the distinguishing characteristics of different blocks of elements					
1.4	define the terms ionisation energy and electron affinity					
1.5	illustrate changes in properties of elements down the group and across periods					
1.6	describe the diagonal relationships in properties of elements in the periodic table.					

Code for evaluation:			
4 – Very well	3 – Well	2 – Fairly well	1 – Not well at all

Topic 2: Chemical reactions
Performance objectives
2.1 Identify reactants and products of any chemical reaction **2.2** Explain the terms reaction time and reaction rate; and the relationship between the two **2.3** Explain collision theory with respect to reaction time and reaction rate **2.4** Describe the influence of the following on chemical reaction rates: nature of substances; concentration/pressure, temperature and catalysts **2.5** Explain endothermic and exothermic reactions **2.6** Illustrate by use of a graph, the energy changes in exothermic and endothermic reactions **2.7** Write equations for simple equilibrium reactions **2.8** State Le Châtelier's principle **2.9** Explain the influence of the following factors on the equilibrium of chemical reactions: concentration, temperature, pressure

Introduction

In this topic we will start with discussing the basic concepts for chemical reactions, such as reactants, products, reaction time and reaction rate. We will also introduce the collision theory and look at how the collision theory describes chemical reactions. There are different factors that affect chemical reactions and we will be looking at how these factors affect the rate of a reaction. We will also be distinguishing between two types of chemical reactions: endothermic and exothermic reactions. Many chemical reactions do not run to completion and instead reach a chemical equilibrium between the reactant and product particles. We will discuss chemical equilibrium by introducing simple equations and studying Le Châtelier's principle. This principle explains how different factors affect the equilibrium of a chemical reaction.

Activity 2.1: Basic concepts (SB p. 24)

Resources

Pictures from magazines or newspapers of milk, fire, eggs, fruit, dying plants

Guidelines

This activity can be used as a class test. Guide a class discussion on the different types of reactions in our daily lives; eg. Milk going sour, food digesting, fires burning, egg boiling in water, plants and animals dying and decaying, etc.

Students should be able to name the reactants and products of a given chemical reaction, and describe reaction time and reaction rate.

Answers

1. Learners should come up with a variety of reactions.
2. a) Reaction time
   b) Reactants
   c) Reaction rate
   d) Arrow
   e) Products
3. Rotting wood; An egg cooking in boiling water; Milk turning sour
4. A reaction becomes slower as reactants are being used during the reaction.

Activity 2.2: Investigate the collision theory (SB p. 26)

Resources

Student's Book, workbook

Guidelines

This activity can be used as a class test.

Students should be able to explain collision theory by explaining the differences in rates of reaction involving various substances.

Answers

1. D
2. C
3. At higher temperatures the particles have more energy and the collisions are more energetic, resulting in more effective collisions.
4. The minimum amount of energy needed for a reaction to take place is called the activation energy.
5. The molecules do not collide with enough energy to react or the orientation of the molecules is not correct.
6. The orientation of the hydrogen and oxygen molecules is not correct when they collide and thus will they not react to form new molecules.

Experiment 2.1: Investigating the effect of temperature on the rate of reaction

(SB p. 27)

Resources

100 ml conical flask, measuring cylinders (50 cm^3 and 10 cm^3), paper with black cross on it, sodium thiosulphate solution at room temperature and heated to 40 °C and 60 °C, dilute hydrochloric acid, stopwatch or timer

Guidelines

Use a simple experiment to illustrate the effect of temperature on the rate of reaction.

Students should be able to state the effects of increase in temperature on reaction rate.

Answers

The results of the experiment might differ from student to student.

Experiment 2.2: Investigating the effect of concentration on reaction rate

(SB p. 28)

Resources

Beakers, magnesium ribbon, H_2SO_4

Guidelines

Use a simple experiment to illustrate the effect of concentration on the rate of reaction.

Students should be able to state the effects of increase in concentration on reaction rate.

Answers

The results of the experiment might differ from student to student.

Experiment 2.3: Investigating the effect of surface area on reaction rate

(SB p. 29)

Resources

Beaker, dish, grinder, calcium carbonate and dilute hydrochloric acid

Guidelines

Use a simple experiment to illustrate the effect of surface area on the rate of reaction.

Students should be able to determine the effect of surface area on reaction rate.

Answers

The results of the experiment might differ from student to student.

Experiment 2.4: Investigating the effect of a catalyst on reaction rate

(SB p. 30)

Resources

Two test tubes, zinc pieces, hydrochloric acid, copper turnings

Guidelines

Use a simple experiment to investigate the effect of a catalyst on reaction rate.

Students should be able to state the effects of a catalyst on reaction rates.

Answers

The results of the experiment might differ from student to student.

Activity 2.3: Factors affecting rate of reaction (SB p. 31)

Resources

Student's Book, workbook

Guidelines

This activity can be used as a class test.

Students should be able to name the factors affecting reaction rates.

Answers

1. a) B: Mg powder has a greater surface area and thus a higher reaction rate
   b) A: At 60 °C the higher temperature will cause a higher reaction rate
   c) B: At 450 kPa the pressure is higher and thus the reaction rate will also be higher
2. a) $Zn + H_2SO_4 \rightarrow ZnSO_4 + H_2$
   b) Zinc and sulphuric acid
   c) Zinc sulphate and hydrogen gas
   d) The concentration of the reactants is still high and thus the rate of reaction will also be high before t_1.
   e) The concentration of the reactants decrease and thus the rate of reaction will also decrease as the reaction proceeds
   f) Higher
   g) The surface area of zinc could have been increased. A higher concentration of sulphuric acid could have been used. The reaction could have taken place at a higher temperature. A catalyst could have been used.

Experiment 2.5: Investigating the heat involved in a reaction (SB p. 37)

Resources

Beakers, thermometers, stirrers, ammonium chloride, ammonium hydroxide and water

Guidelines

Demonstrate the heat involved in a reaction by performing simple experiment.

Students should be able to note and explain the heat involved in a reaction.

Answers

1. Cold
2. The temperature decreased.
3. The temperature increased.
4. Ammonium chloride in water is an endothermic process and ammonium hydroxide and water is an exothermic process.

Experiment 2.6: Investigating the heat involved in a chemical reaction (SB p. 38)

Resources

Beaker, concentrated H_2SO_4 and sugar

Guidelines

Safety Precaution: when mixing concentrated acid and water, always add the acid to the water; never add water to concentrated acid. Students should be able to state the effects of increase in temperature on reaction rates.

Answers

1. Hot
2. yes
3. black
4. Sugar is a carbohydrate, so when you remove the water from the molecule, you're basically left with elemental carbon.

Activity 2.4: Endothermic and exothermic reactions (SB p. 39)

Resources

Student's Book, workbook

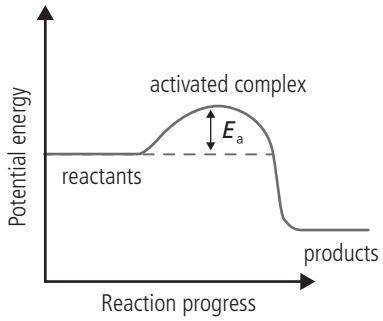
Guidelines

This activity can be used as a class test. Initiate class discussion on the energy relationships in endothermic and exothermic reactions.

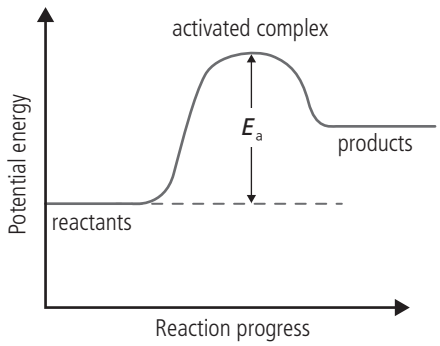
Students should be able to explain the effects of temperature on the equilibrium of chemical reactions.

Answers

1. a) ΔH is negative, exothermic reaction
$2H_2(g) + O_2(g) \rightarrow H_2O(l) + 285{,}8\ kJ$



b) ΔH is positive, endothermic reaction
$H_2O(l) + 40{,}7\ kJ \rightarrow H_2O(g)$



2. ΔH is negative thus this reaction is exothermic. The energy of the reactants is greater than the energy of the products.
3. a) 1: Activation energy 2: Enthalpy change 3: Activated complex
   b) $+100\ kJ \cdot mol^{-1}$
   c) $-150\ kJ \cdot mol^{-1}$
   d) $\Delta H = H_{products} - H_{reactants}$
   $= (-150\ kJ \cdot mol^{-1}) - (+100\ kJ \cdot mol^{-1})$
   $= -250\ kJ \cdot mol^{-1}$
   e) Exothermic reaction, the enthalpy change is negative because the energy of the products is less than the energy of the reactants.

Activity 2.5: Developing K_c expressions (SB p. 41)

Resources

Student's Book, workbook

Guidelines

This activity can be used as a class test. Students should be able to develop K_c expressions.

Answers

1. $K_c = \frac{[SO_3]^2}{[SO_2]^2[O_2]}$
2. $K_c = [HCl]\,[NH_3]$
3. $K_c = \frac{[NO_2]^2}{[N_2O_4]}$
4. $K_c = \frac{[NH_4OH]}{[NH_3]}$
5. $K_c = [CO_2]$

Activity 2.6: Perform equilibrium constant calculations (SB p. 42)

Resources

Student's Book, workbook

Guidelines

This activity can be used as a class test.
Students should be able to perform equilibrium constant calculations.

Answers

1. a) $K_c = \frac{[SO_3]^2}{[SO_2]^2[O_2]}$
   b) $K_c = \frac{[SO_3]^2}{[SO_2]^2[O_2]} = \frac{(5.0)^2}{(3.1)^2(1.9)} = 1.37$

2.

	$H_2(g)$ +	$CO_2(g)$ $\rightleftharpoons$	$H_2O(g)$ +	CO
Mole ratio	1	1	1	1
n initially (mol)	0,8	0,8	0	0
n change (mol)	–0,55	–0,55	+0,55	+0,55
n at equilibrium (mol)	0,25	0,25	0,55	0,55
$c = \frac{n}{V}$ ($mol \cdot dm^{-3}$)	$\frac{0{,}25}{5{,}0}$ = 0,05	$\frac{0{,}25}{5{,}0}$ = 0,05	$\frac{0{,}55}{5{,}0}$ = 0,11	$\frac{0{,}55}{5{,}0}$ = 0,11

$$K_c = \frac{[H_2O][CO]}{[H_2][CO_2]} = \frac{(0.11)(0.11)}{(0.05)(0.05)} = 4.84$$

3.

	$CO(g)$ +	$O_2(g) \rightleftharpoons$	$2CO_2(g)$
Mole ratio	1	1	2
n initially (mol)	5	5	0
n change (mol)	–1,5	–1,5	+3,0
n at equilibrium (mol)	3,5	3,5	3,0
$c = \frac{n}{V}$ (mol·dm^{-3})	$\frac{3.5}{1.0} = 3.5$	$\frac{3.5}{1.0} = 3.5$	$\frac{3.5}{1.0} = 3.5$

$$K_c = \frac{[CO_2]^2}{[CO][O_2]} = \frac{(3.0)^2}{(3.5)(3.5)} = 0.73$$

Activity 2.7: Consider changes in concentration

(SB p. 44)

Resources

Student's Book, workbook

Guidelines

This activity can be used as a class test. Students should be able to explain the effects of concentration on the equilibrium of chemical reactions.

Answers

a) When CO_2 is added to the system, the system will try to decrease the concentration of CO_2. The forward reaction will be favoured, thus the $[CO_2]$ and $[H_2]$ will decrease and $[H_2O]$ and $[CO]$ will increase. The equilibrium will shift to the right.

b) When CO is added to the system, the system will try to decrease the concentration of CO. The reverse reaction will be favoured, thus the $[CO_2]$ and $[H_2]$ will increase and $[H_2O]$ and $[CO]$ will decrease. The equilibrium will shift to the left.

c) When H_2 is removed from the system, the system will try to increase the concentration of H_2. The reverse reaction will be favoured, thus the $[CO_2]$ and $[H_2]$ will increase and $[H_2O]$ and $[CO]$ will decrease. The equilibrium will shift to the left.

d) When H_2O is removed from the system, the system will try to increase the concentration of H_2O. The forward reaction will be favoured, thus the $[CO_2]$ and $[H_2]$ will decrease and $[H_2O]$ and $[CO]$ will increase. The equilibrium will shift to the right.

Activity 2.8: Think about changes in temperature

(SB p. 44)

Resources

Student's Book, workbook

Guidelines

This activity can be used as a class test. Students should be able to explain the effects of temperature on the equilibrium of chemical reactions.

Answers

a) When the temperature in the container is decreased the system will try to counter-act the change by increasing the temperature. A decrease in temperature will favour the exothermic reaction. The forward reaction is the exothermic reaction. The $[SO_2]$ and $[O_2]$ will decrease and the $[SO_3]$ will increase. The equilibrium shifts to the right.

b) When the temperature in the container is increased the system will try to counter-act the change by decreasing the temperature. An increase in temperature will favour the endothermic reaction. The reverse reaction is the endothermic reaction. The $[SO_2]$ and $[O_2]$ will increase and the $[SO_3]$ will decrease. The equilibrium shifts to the left.

Activity 2.9: Consider changes in pressure (SB p. 45)

Resources

Student's Book, workbook

Guidelines

This activity can be used as a class test. Students should be able to explain the effects of pressure on the equilibrium of chemical reactions.

Answers

a) By increasing the pressure in the container, the equilibrium will shift to the side with the least number of moles. The forward reaction will be favoured and the equilibrium shifts to the right.

b) By decreasing the pressure in the container, the equilibrium will shift to the side with the most number of moles. The reverse reaction will be favoured and the equilibrium shifts to the left.

Activity 2.10: Consider factors affecting chemical equilibrium (SB p. 46)

Resources

Student's Book, workbook

Guidelines

This activity can be used as a class test. Students should be able to explain the effects of all the factors such as temperature, concentration and pressure in table form, such as the example on p. 44 of the SB.

Answers

1.

	Reaction	Temperature increased	Temperature decreased	Pressure increased	Pressure decreased
a	$N_2(g) + O_2(g) \rightleftharpoons 2NO(g)$ + heat	left	right	no shift	no shift
b	$2H_2O(g) \rightleftharpoons 2H_2(g) + O_2(g)$, $\Delta H = +241{,}7$ kJ	right	left	left	right
c	$N_2(g) + 2H_2O(g)$ + heat $\rightleftharpoons 2NO(g) + 2H_2(g)$	right	left	left	right
d	$2SO_2(g) + O_2(g) \rightleftharpoons 2SO_3(g)$ + heat	left	right	right	left
e	$N_2O_4(g)$ + heat $\rightleftharpoons 2NO_2(g)$	right	left	left	right
f	$PCl_3(g) + Cl_2(g) \rightleftharpoons PCl_5(g)$ + 21 kJ	left	right	right	left

2. a) exothermic
   b) endothermic
   c) endothermic
   d) exothermic
   e) endothermic
   f) exothermic
3. A catalyst has no effect on the equilibrium position.
4. a) Decreases
   b) Remains the same
   c) Increases
5. a) When the temperature is increased the system will try to counter-act the change by decreasing the temperature. The temperature will decrease when the endothermic reaction is favoured. The endothermic reaction is the reverse reaction. When the reverse reaction is favoured the $[N_2]$ and $[H_2]$ will increase and $[NH_3]$ will decrease. The equilibrium shifts to the left.

b) Adding a catalyst to the reaction mixture has no effect on the equilibrium position. A catalyst increases the rate of both the forward and reverse reaction. Thus the $[NH_3]$ will remain the same.

c) When the pressure is increased the system will try to counter-act the change by decreasing the pressure of the system. The equilibrium will shift to the side with the least number of moles. The forward reaction is favoured, the equilibrium shifts to the right. The $[N_2]$ and $[H_2]$ will decrease and the $[NH_3]$ will increase.

Revision questions: Answers

(SB p. 48)

1. a) A
   b) C
   c) B
2. a) Activated complex
   b) Collision theory
   c) Activation energy
   d) Collision frequency
   e) Reaction rate
3. a) higher
   b) lower
   c) higher
   d) higher
   e) the same
4. a) enthalpy change
   b) reversible reaction
   c) The rate of the forward reaction is equal to the rate of the reverse reaction
   d) exothermic
   e) endothermic
   f) For a system to be able to reach equilibrium a closed system is needed. When working with gases a closed system can be obtained by using a closed container.
5. a) $K_c = \frac{[NO_3]^2}{[NO_2]^2[O_2]}$
   b) $K_c = \frac{[CO][Cl_2]}{[COCl_2]}$
   c) $K_c = \frac{[CO]^2}{[CO_2]}$
   d) $K_c = \frac{[H_2O]^2}{[H_2]^2[O_2]}$
6. a) Homogeneous, the reactants and products are all in the same phase (gas phase).
   b) $K_c = \frac{[HI]^2}{[H_2][I_2]} = \frac{(17.7 \times 10^{-3})^2}{(1.84 \times 10^{-3})(3.13 \times 10^{-3})}$

      $= 3073{,}34$
   c) The increase in pressure will increase the rate of both the forward and reverse reaction equally. The increase in pressure will have no effect on the equilibrium position, because the number of moles of the reactants is equal to the number of moles of the products.
7.

	$N_2(g)$ +	$3H_2(g)$ $\rightleftharpoons$	$2NH_3(g)$
Mole ratio	1	3	2
n initially (mol)	3.0	8.0	0
n change (mol)	–2.0	–6.0	+4.0
n at equilibrium (mol)	1.0	2.0	4.0
c at = $\frac{n}{V}$ (mol·dm^{-3})	$\frac{1.0}{2.0} = 0.5$	$\frac{2.0}{2.0} = 1.0$	$\frac{4.0}{2.0} = 2.0$

$$K_c = \frac{[NH_3]^2}{[N_2][H_2]^3} = \frac{(2.0)^2}{(0.5)(1.0)^3} = 8.0$$

8. a) Le Chatelier's Principle states that a change in any of the factors that determine the position of the equilibrium in a system will cause the system to change in such a manner as to reduce or counteract the effect of the change.
   b) If the temperature in the container was decreased, the system will try to counter-act the change by increasing the temperature. Decreasing the temperature favours the exothermic reaction. The reverse reaction is favoured. The $[CO_2]$ and $[H_2]$ will increase and $[H_2O]$ and $[CO]$ will decrease. The equilibrium shifts to the left.
   c) If the concentration of $CO_2(g)$ in the container was increased, the system will try to counter-act the change by decreasing the concentration of CO_2. The forward reaction will be favoured. The $[CO_2]$ and $[H_2]$ will decrease and $[H_2O]$ and $[CO]$ will increase. The equilibrium will shift to the right.

How are you doing?

Take this opportunity to ask learners if there is anything that they do not understand. You can check their understanding by asking them some questions about the information covered in the unit. Explain anything that learners do not understand.

Key words

chemical reactions – a process which leads to a chemical substance changing into a different substance: 22

reactants – reacting substances that are converted to new substances: 22

products – new substances that are produced in a chemical reaction: 22

reaction rate – the change in concentration of reactants with time or the change in concentration of products with time; describes the speed at which a reaction is taking place: 23

reaction time – how long the reaction is taking place: 23

collision theory – gives the conditions for when a reaction will occur; i.e. particles must collide and only effective collisions will result in a chemical reaction: 24

activation energy – the minimum energy needed for a reaction to take place: 25

exothermic reactions – when heat exit from the system to the surroundings: 31, 71

endothermic reactions – when heat enters into the system from the surroundings: 32

change in enthalpy – the energy of the products minus the energy of the reactants: 32

irreversible reactions – proceed only in the forward direction: 47

reversible reactions – take place in both the forward and backward directions: 47

chemical equilibrium – the rate of the forward and reverse reactions is equal: 39

homogeneous equilibrium – a family of organic molecules in which all the molecules have the same functional group, but different lengths of carbon chains: 228

heterogeneous equilibrium – results from a reversible reaction, involving reactants and products that are in different phases: 40

equilibrium constant (K_c) – the ratio of the product of the product concentrations to the product of the reactant concentrations: 41

Checklist for Self-evaluation

Theme 1 Topic 2

EVALUATION GUIDE: Student should be able to:

	Criteria	4	3	2	1	Comments
2.1	name the reactants and products of a given chemical reaction					
2.2	describe reaction time and reaction rate					
2.3	explain the differences in rates of reaction involving various substances in terms of collision theory					
2.4	name the factors affecting reaction rates					
2.5	state the effects of increase in temperature, concentration/ pressure or a catalyst on reaction rates					
2.6	define endothermic and exothermic reactions and illustrate the energy changes in them					
2.7	state Le Châtelier's principles					
2.8	explain the effects of temperature, concentration and pressure on the equilibrium of chemical reactions					
2.9	write and balance equations for chemical reactions in equilibrium, showing clearly the reactants, and products of the reactions					
2.10	carry out simple calculations on equilibrium constants					
2.11	use simple calculations to illustrate Le Châtelier's principle.					

Code for evaluation:			
4 – Very well	3 – Well	2 – Fairly well	1 – Not well at all

Topic 3: Mass, volume relationships

Performance objectives

3.1 explain the concepts of the mole, molar, STP, relative densities and relative molar mass (RMM)
3.2 solve problems involving reacting masses and volumes in chemical reactions
3.3 state the SI units of various basic quantities.

Introduction

In this topic we will discuss mass and volume relationships. We will study various basic concepts such as the mole, molar volume, relative density, relative molar mass, STP and molar solutions. We will look at SI units of various basic quantities. This topic will also deal with solving problems involving reacting masses and volumes in chemical reactions.

Activity 3.1: Performing calculations involving the mole concept (SB p. 53)

Resources

Student's Books, workbook

Guidelines

Give the students practice questions. This activity can be used as a class test.

Students should be able to calculate the masses, moles and volumes in chemical reactions.

Answers

1. a) $M(MgSO_4 \cdot 7H_2O) = (24,3) + (32,1) + (4 \times 6) + (7 \times 2 \times 1) + (7 \times 16) = 246,4\ g \cdot mol^{-1}$
   b) $M(K_2Cr_2O_7) = (2 \times 39,1) + (2 \times 52) + (7 \times 16) = 294,2\ g \cdot mol^{-1}$
   c) $M(K_4Fe(CN)_6) = (4 \times 39,1) + (55,8) + (6 \times 12) + (6 \times 14) = 368,2\ g \cdot mol^{-1}$
2. a) $m = nM = (3)(63,5) = 190,5\ g$
   b) $M(CO_2) = (12) + (2 \times 16) = 44\ g \cdot mol^{-1}$
   $m = nM = (0,25)(44) = 11\ g$
   c) $M(N_2) = (2 \times 14) = 28\ g \cdot mol^{-1}$
   $m = nM = (0,02)(28) = 0,56\ g$
3. a) $M(O_2) = (2 \times 16) = 32\ g \cdot mol^{-1}$
   $n = \frac{m}{M} = \frac{6,4}{32} = 0,2\ mol$
   b) $M(NH_4NO_3) = (14) + (4 \times 1) + (14) + (3 \times 16) = 80\ g \cdot mol^{-1}$
   $n = \frac{m}{M} = \frac{480}{80} = 6\ mol$
   c) $n = \frac{m}{M} = \frac{2\,000}{55,8} = 35,8\ mol$

Activity 3.2: Perform calculations on molar volume (SB p. 54)

Resources

Student's Books, workbook

Guidelines

Guide students to calculate the relative densities of substances.

Give the students practice questions. This activity can be used as a class test.

Students should be able to calculate the masses, moles and volumes in chemical reactions.

Answers

1. a) $V = nV_M = (1,5)(22,4) = 3,6\ dm^3$
   b) $V = nV_M = (0,01)(22,4) = 0,224\ dm^3$
   c) $V = nV_M = (3)(22,4) = 67,2\ dm^3$
2. a) $n = \frac{V}{V_M} = \frac{50}{22,4} = 2,23\ mol$
   number of atoms $= n \times N_A = (2,23)(6,02 \times 10^{23}) = 1,34 \times 10^{24}$ atoms
   b) $n = \frac{V}{V_M} = \frac{10}{22,4} = 0,45\ mol$
   number of nucleons $= n \times N_A = (0,45)(6,02 \times 1023) = 2,71 \times 10^{23}$

Activity 3.3: Perform calculations with molality and molarity

(SB p. 56)

Resources

Student's Books, workbook

Guidelines

Guide students to calculate molality and molarity.

Give the students practice questions. This activity can be used as a class test.

Students should be able to calculate the masses, moles and volumes in chemical reactions.

Answers

1. $\text{Molality} = \frac{\text{moles of solute (mol)}}{\text{mass of solvent (kg)}}$

   $= \frac{0{,}75\ \text{mol}}{2{,}5\ \text{kg}} = 0{,}3\ \text{mol}\cdot\text{kg}^{-1}$

2. $n = \frac{m}{M} = \frac{30}{119} = 0{,}25\ \text{mol}$

   $\text{Molality} = \frac{\text{moles of solute (mol)}}{\text{mass of solvent (kg)}}$

   $= \frac{0{,}25\ \text{mol}}{0{,}75\ \text{kg}} = 0{,}33\ \text{mol}\cdot\text{kg}^{-1}$

3. $\text{Molality} = \frac{\text{moles of solute (mol)}}{\text{mass of solvent (kg)}}$

   $= \frac{0{,}79\ \text{mol}}{0{,}5\ \text{kg}} = 1{,}58\ \text{mol}\cdot\text{kg}^{-1}$

4. moles of solute = molality × mass of solvent = (0,195)(1,5) = 0,29 mol

5. $M(Na_2O) = (2 \times 23) + (16) = 62\ \text{g}\cdot\text{mol}^{-1}$

   $c = \frac{m}{MV} = \frac{(10)}{(62)(0{,}2)} = 0{,}806\ \text{mol}\cdot\text{dm}^{-3}$

6. $M(MgCl_2) = (24{,}3) + (2 \times 35{,}5)$

   $= 95{,}3\ \text{g}\cdot\text{mol}^{-1}$

   $m = cMV = (1{,}0)(95{,}3)(0{,}1) = 9{,}53\ \text{g}$

7. $M(NaCl) = (23) + (35{,}5) = 58{,}5\ \text{g}\cdot\text{mol}^{-1}$

   $m = cMV = (0{,}1)(58{,}5)(0{,}5) = 2{,}925\ \text{g}$

Activity 3.4: Perform calculations with relative density and relative molar mass

(SB p. 54)

Resources

Student's Books, workbook

Guidelines

Guide students to calculate the relative densities of substances.

Give the students practice questions. This activity can be used as a class test.

Students should be able to calculate the masses, moles and volumes in chemical reactions.

Answers

1. a. $M(H_2SO_4) = (2 \times 1) + (32{,}1) + (4 \times 16) = 98{,}1\ \text{g}\cdot\text{mol}^{-1}$

   b. $M(Na_2CO_3) = (2 \times 23) + (12) + (3 \times 16) = 106\ \text{g}\cdot\text{mol}^{-1}$

   c. $M(CH_2CH_2) = (12) + (2 \times 1) + (12) + (2 \times 1) = 28\ \text{g}\cdot\text{mol}^{-1}$

   d. $M((NH_4)_3PO_4) = (3 \times 14) + (3 \times 4 \times 1) + (31) + (4 \times 16) = 149\ \text{g}\cdot\text{mol}^{-1}$

2. $\rho = \frac{m}{M} = \frac{0{,}051}{75 \times 10^{-6}} = 680\ \text{kg/m}^3$

3. $RD = \frac{\rho\text{mercury}}{\rho\text{water}} = \frac{13\ 534\ \text{kg/m}^3}{1\ 000\ \text{kg/m}^3} = 13{,}5$

4. Copper is 8,96 times more dense than water. Density of copper $= 8{,}96 \times 1000 = 8\ 960\ \text{kg/m}^3$

Activity 3.5: Perform calculations involving mass and volume

INDIVIDUAL (SB p. 56)

Resources

Student's Books, workbook

Guidelines

Guide students to calculate the volumes of gas in chemical reactions.

Give the students practice questions. This activity can be used as a class test.

Students should be able to calculate the masses, moles and volumes in chemical reactions.

Answers

1. $2KClO_3 \rightarrow 2KCl + 3O_2$

   $n(KClO_3) = \frac{m}{M} = \frac{29,4}{122,6} = 0,24$ mol

   Mole ratio:
   $KClO_3 : O_2$
   2 : 3
   0,24 mol : $\frac{3}{2}(0,24) = 0,36$ mol
   Mass of O_2:

   $m(O_2) = nM = (0,36)(32) = 11,52$ g

   Volume of O_2:

   $V(O_2) = nV_m = (0,36)(22,4) = 8,064\ dm^3$

2. a) $2SO_2(g) + O_2(g) \rightarrow 2SO_3(g)$

   $n(SO_2) = \frac{m}{M} = \frac{96}{64} = 0,5$ mol

   Mole ratio:
   $SO_2 : SO_3$
   2 : 2
   1 : 1
   1,5 mol : 1,5 mol
   $m(SO_3) = nM = (1,5)(80) = 120$ g

   b) $V(SO_3) = nV_m = (1,5)(22,4) = 33,6\ dm^3$

   c) Mole ratio:
   $SO_2 : O_2$
   2 : 1

   1,5 mol : $\frac{1}{2}(1,5) = 0,75$ mol

   $V(O_2) = nV_m = (0,75)(22,4) = 16,8\ dm^3$
   Volume of O_2 used is $16,8\ dm^3$

   d) Number of moles of O_2 left at the end of reaction:
   $n(O_2) = 2$ mol – 0,75 mol = 1,25 mol
   $V(O_2) = nV_m = (1,25)(22,4) = 28\ dm^3$
   Total volume of gas in the container at the end of the reaction:
   Total volume = $V(O_2) + V(SO_3)$
   $= 28\ dm^3 + 33,6\ dm^3 = 61,6\ dm^3$

   e) number of molecules of $SO_3 = n \times N_A$
   $= (1,5)(6,02 \times 10^{23}) = 9,03 \times 10^{23}$

3. a) $2Li + 2H_2O \rightarrow 2LiOH + H_2$

   b) $n(Li) = \frac{m}{M} = \frac{2}{7} = 0,29$ mol

   Mole ratio:
   $Li : H_2$
   2 : 1
   0,29 mol : $\frac{1}{2}(0,29) = 0,145$ mol

   Mass of H_2:

   $m(H_2) = nM = (0,145)(2) = 0,29$ g

   Volume of H_2:

   $V(H_2) = nV_m = (0,145)(22,4)$
   $= 3,248\ dm^3$

4. a) $2Mg + O_2 \rightarrow 2MgO$

   b) $n(Mg) = \frac{m}{M} = \frac{20}{24,3} = 0,82$ mol

   Mole ratio:
   Mg : MgO
   2 : 2
   1: 1
   0,82 mol : 0,82 mol
   $n(MgO) = nM = (0,82)(40,3)$
   $= 33,05$ g

   c) Mole ratio:
   Mg: O_2
   2 : 1

   0,82 mol : $\frac{1}{2}(0,82) = 0,41$ mol

   $V(O_2) = nV_m = (0,41)(22,4)$
   $= 9,184\ dm^3$

Activity 3.6: Match SI units of quantities

(SB p. 67)

Resources

Student's Books, workbook

Guidelines

Guide students to deduce the SI units of quantities.
Give the students practice questions. This activity can be used as a class test.

Students should be able to derive SI units of quantities.

Answers

1. C
2. E
3. D
4. A
5. B

Revision questions: Answers

(SB p. 68)

1. a) B
   b) C
   c) D
   d) B
2. a) Standard solution
   b) Mol
   c) Molar mass
   d) Water of crystallisation
   e) Molarity
   f) Molar volume
3. a) $O_2 + 2H_2 \rightarrow 2H_2O$
   b) $n(O_2) = \frac{V}{V_M} = \frac{0{,}03}{22{,}4} = 0{,}0013$ mol
   Mole ratio:
   $O_2 : H_2O$
   1 : 2
   0,0013 mol : (2 × 0,0013) = 0,0026 mol
   $V(H_2O) = nV_M = (0{,}0026) + (22{,}4)$
   $= 0{,}05824\ dm^3$
   c) $n(O_2) = \frac{V}{V_M} = \frac{0{,}03}{22{,}4} = 0{,}0013$ mol
   d) $m = nM = (0{,}0013)(32) = 0{,}0416$ g
4. a) $\text{Molality} = \frac{\text{moles of solute (mol)}}{\text{mass of solvent (kg)}}$
   $= \frac{3\ \text{mol}}{0{,}5\ \text{kg}} = 6\ \text{mol·kg}^{-1}$
   b) $n = \frac{m}{M} = \frac{12}{171{,}3} = 0{,}07$ mol
   $\text{Molality} = \frac{\text{moles of solute (mol)}}{\text{mass of solvent (kg)}}$
   $= \frac{0{,}07\ \text{mol}}{2\ \text{kg}} = 0{,}035\ \text{mol·kg}^{-1}$
   c) $\text{Molality} = \frac{\text{moles of solute (mol)}}{\text{mass of solvent (kg)}}$
   $= \frac{6\ \text{mol}}{0{,}2\ \text{kg}} = 30\ \text{mol·kg}^{-1}$
5. a) Convert 250 ml $= \frac{250}{1\,000}\ dm^3 = 0{,}25\ dm^3$
   b) A standard solution is a solution of known concentration.
   c) $M(LiCl) = (7) + (35{,}5) = 42{,}5\ \text{g·mol}^{-1}$
   $m = cMV = (2)(42{,}5)(0{,}25) = 21{,}25$ g
6. a) Air, Wood, Ice, Ethanol, Methanol
   b) Aluminium, Lead, Gold
   c) Methanol
   d) $RD = \frac{\rho_{ice}}{\rho_{water}} = \frac{930\ \text{kg/m}^3}{1\,000\ \text{kg/m}^3} = 0{,}93$
7. a) $M(H_2SO_4) = (2 \times 1) + (32{,}1) + (4 \times 16) = 98{,}1\ \text{g·mol}^{-1}$
   b) $M(NH_3) = (14) + (3 \times 1) = 17\ \text{g·mol}^{-1}$
   c) $M(KMnO_4) = (39{,}1) + (54{,}9) + (4 \times 16) = 158\ \text{g·mol}^{-1}$
   d) $M(FeCl_3) = (55{,}8) + (3 \times 35{,}5) = 162{,}3\ \text{g·mol}^{-1}$
   e) $M(Mg(OH)_2) = (24{,}3) + (2 \times 16) + (2 \times 1) = 58{,}3\ \text{g·mol}^{-1}$
8. $Ca + 2HCl \rightarrow CaCl_2 + H_2$
   $n(Ca) = \frac{m}{M} = \frac{15}{40} = 0{,}375$ mol
   Mole ratio:
   $Ca : CaCl_2$
   1: 1
   0,375 mol : 0,375 mol
   $m(CaCl_2) = nM = (0{,}375)(111)$
   $= 41{,}625$ g
9. $CaCO_3 + 2HCl \rightarrow CaCl_2 + CO_2 + H_2O$
   $n(HCl) = \frac{m}{M} = \frac{15{,}2}{36{,}5} = 0{,}416$ mol
   Mole ratio:
   $2HCl : CaCO_3$
   2: 1
   0,416 mol: $(\frac{1}{2} \times 0{,}416) = 0{,}208$ mol
   $M(CaCO_3) = (40) + (12) + (3 \times 16)$
   $= 100\ \text{g·mol}^{-1}$
   $m(CaCO_3) = nM = (0{,}208)(100) = 20{,}8$ g
10. $O_2 + 2CO \rightarrow 2CO_2$
    $V(CO_2) = 20\ dm^3$
    $n(CO_2) = \frac{V(CO_2)}{V_m} = \frac{20}{22{,}4} = 0{,}89$ mol
    Mole ratio:
    $CO_2 : O_2$
    2 : 1
    0,89 mol : $\frac{1}{2}(0{,}89) = 0{,}445$ mol
    $V(O_2) = nV_m = (0{,}445)(22{,}4) = 9{,}968\ dm^3$
11. $CaCO_3 \rightarrow CaO + CO_2$
    $n(CO_2) = \frac{V(CO_2)}{V_m} = \frac{11{,}2}{22{,}4} = 0{,}5$ g
    Mole ratio:
    $CO_2 : CaCO_3$
    1 : 1
    0,5 mol: 0,5 mol

$m(CaCO_3) = nM = (0,5)(100) = 50$ g

How are you doing?

Take this opportunity to ask learners if there is anything that they do not understand. You can check their understanding by asking them some questions about the information covered in the unit. Explain anything that learners do not understand.

Key words

mole is the amount of substance: 50

molar mass – the mass of one mole of a substance; calculated by adding the relative atomic masses of the elements making up the substance: 51

formula mass – the sum of the relative atomic masses of a molecule: 52

molar volume – one mole of any gas occupies a volume of 22.414 dm^3 or 22.414 litres at STP: 54

concentration – the amount of solute present in a given quantity of solvent: 55

molality (molal concentration) - the amount of substance in a specified amount of mass (in kg) of the solvent: 55

molarity (molar concentration) – the number of moles of solute in 1 litre (1 dm^3) of solution: 55

standard solution – a solution of accurately known concentration: 55

relative density – the ratio of the density of a given substance with respect to the density of a standard substance (reference), whose density is already known under specific conditions: 57

relative molar mass (RMM) – is the sum of the mass of all the atoms present in the molecule of a compound: 58

stoichiometric calculations – calculations to find the masses or volumes of reactants or products: 59

SI units – the international system of units for measurement: 65

Checklist for Self-evaluation

Theme 1 Topic 3

EVALUATION GUIDE: Student should be able to:

	Criteria	4	3	2	1	Comments
3.1	define the mole, molar mass, molar volume, relative density, relative molar mass; molar solutions and s.t.p.					
3.2	compute the number of moles and molar masses of reactants and products from chemical equations of chemical reactions					
3.3	calculate the masses and volume of substances in hypothetical chemical reactions					
3.4	use appropriate units for expressing their answers					
3.5	derive the SI units of given quantities.					

Code for evaluation:			
4 – Very well	3 – Well	2 – Fairly well	1 – Not well at all

THEME 2

Chemistry and environment

Topic 4: Acid-base reactions

Performance objectives

4.1 define concentration in mol·dm^{-3} of solutions
4.2 define standard solutions
4.3 explain the relationship between concentrations and volumes of reacting substances
4.4 mathematically express the relationship between the concentration in mol·dm^{-3} and volume of a solution
4.5 carry out acid-base titrations using appropriate indicators
4.6 record correctly titre values to two decimal places
4.7 carry out relevant calculations from titration results.

Introduction

Many substances that we encounter in our homes every day can be classified either as acids or bases. Aspirin, for example, contains acetylsalicylic acid, while milk of magnesia is magnesium hydroxide, which is a base.

Before we look at acid-base reactions in more detail, here is a short summary of the background knowledge you need for this topic.

Experiment 4.1: Colours of indicators

(SB p. 72)

Resources

Dilute solutions of acids (HCl, H_2SO_4, HNO_3 or acetic acid (CH_3COOH)), dilute solutions of bases (NaOH, NH_3, Na_2CO_3 or $NaHCO_3$), indicator solutions (solutions of litmus, methyl orange, bromothymol blue, phenolphthalein and universal indicator), test tube stand with five test tubes, medicine dropper

Guidelines

Guide the students to identify different types of solutions by their colour.

Students should be able to define and give examples of solutions.

Answers

No questions to answer

Experiment 4.2: Making a standard solution

(SB p. 74)

Resources

200 ml volumetric flask, balance, watch glass, small funnel, NaCl crystals, wash bottle with distilled water

Guidelines

Guide the students to prepare standard solutions.

Display and explain the apparatus necessary for carrying out simple titrations.

Students should be able to prepare standard solutions.

Answers

No questions to answer

Activity 4.2: Test yourself

(SB p. 77)

Resources

Student's Book, workbook

Guidelines

This activity can be done as a class test.

Students should be able to carry out relevant calculations from titre values.

Answers

1. a) $Na_2CO_3 + 2HNO_3 \rightarrow 2NaNO_3 + H_2O + CO_2$

   b) $c_b = \frac{m}{MV} = \frac{1{,}06\ g}{(106{,}0\ g \cdot mol^{-1})(0{,}2\ dm^3)}$

   $= 0{,}05\ mol \cdot dm^{-3}$

   $\frac{n_a}{n_b} = \frac{c_a V_a}{c_b V_b}$

   $\therefore \frac{2}{1} = \frac{c_a \times 40\ cm^3}{0{,}05\ mol \cdot dm^{-3} \times 25\ cm^3}$

   $\therefore c_a = 0{,}06\ mol \cdot dm^{-3}$

2. a) Hydrochloric acid: strong acid; sodium hydroxide: strong base

   b) Bromothymol blue

   c) End point of the reaction

   d) $HCl + NaOH \rightarrow NaCl + H_2O$

   e) $\frac{n_a}{n_b} = \frac{c_a V_a}{c_b V_b}$

   $\therefore \frac{1}{1} = \frac{c_a \times 25{,}0\ cm^3}{0{,}12\ mol \cdot dm^{-3} \times 18.8\ cm^3}$

   $\therefore c_a = 0{,}09\ mol \cdot dm^{-3}$

   f) 50 cm^3 and 250 cm^3 of NaOH solution contain the same number of moles of NaOH, because only water is added in the dilution.

   Number of moles of NaOH in 50 cm^3 water = number of moles of NaOH in 250 cm^3 water

   $n_1 = N_2$

   $c_1V_1 = c_2V_2$

   $0{,}21\ mol \cdot dm^{-3} \times 0{,}05\ dm^3 = c_2 \times 0{,}25\ dm^3$

   $\therefore c_2 = \frac{0{,}12\ mol \cdot dm^{-3} \times 0{,}05\ dm^3}{0{,}25\ dm^3}$

   $= 0{,}024\ mol \cdot dm^{-3}$

3. a) Weak acid

   b) pH higher than 7

   c) Phenolphthalein

   d) $(COOH)_2 + 2KOH \rightarrow K_2(COO)_2 + 2H_2O$

   e) $\frac{n_a}{n_b} = \frac{c_a V_a}{c_b V_b}$

   $\therefore \frac{1}{2} = \frac{0{,}12\ mol \cdot dm^{-3} \times 20{,}0\ cm^3}{c_b \times 25\ cm^3}$

   $\therefore c_b = 0{,}19\ mol \cdot dm^{-3}$

Experiment 4.3: Acid-base titrations

(SB p. 77)

Resources

Burette and 25 cm^3-pipette or two swift pipettes, retort stand, glass beaker or conical flask, spatula, funnel, watch glass, volumetric flask, distilled water, 0.5 $mol \cdot dm^{-3}$ sodium hydroxide solution, phenolphthalein solution, oxalic acid crystals ($(COOH)2 \cdot 2H_2O$), white vinegar, basic drain cleaner, mass meter (balance)

Guidelines

Demonstrate the process of acid-base titrations.

Students should be able to neutralise acids with bases.

Answers

Observations and calculations

Part 1

$c = \frac{m}{MV}$

$= \frac{1.57\ g}{126.0\ g \cdot mol^{-1} \times 0.025\ dm^3}$

$= 0.50\ mol \cdot dm^{-3}$

Part 2

$2NaOH(aq) + (COOH)_2(aq) \rightarrow Na_2(COO)_2(aq) + 2H_2O(l)$

From the stoichiometric equation we have 1 mol $(COOH)_2$ reacts with 2 mol NaOH.

The mol ratio is $\frac{\text{number of mol of acid}}{\text{number of mol of base}} = \frac{1}{2}$, so the number of moles of acid must be $\frac{1}{2}$ that of the base.

$2NaOH(aq) + (COOH)_2(aq) \rightarrow$
$Na_2(COO)_2(aq) + 2H_2O(l)$

$$\frac{n_a}{n_b} = \frac{c_a V_a}{c_b V_b}$$

$$\frac{1\ \text{mol}}{2\ \text{mol}} = \frac{0.50\ \text{mol}\cdot\text{dm}^{-3} \times V_a}{c_b \times 0.025\ \text{dm}^3}$$

$$\therefore c_b = \frac{2\ \cancel{\text{mol}} \times 0.50\ \text{mol}\cdot\text{dm}^{-3} \times V_a\ \cancel{\text{cm}^3}}{1\ \cancel{\text{mol}} \times 0.025\ \cancel{\text{dm}^3}}$$

$$= x\ \text{mol}\cdot\text{dm}^{-3}$$

Part 3

To calculate the percentage of acetic acid in vinegar:

$$\frac{\text{titration volume of acetic acid}}{0.025\ \text{dm}^3\ \text{vinegar}} \times 100\%$$

Part 4

To calculate the percentage of sodium hydroxide in the drain cleaner:

$$\frac{\text{titration volume of drain cleaner}}{0.025\ \text{dm}^3\ \text{drain cleaner}} \times 100\%$$

Revision questions: Answers

(SB p. 82)

1. A strong base dissociates completely in water to form a strong alkali; whereas a weak base dissociates partially in water to form a weak alkali.

   A concentrated alkaline solution contains a large amount of base (solute) and a small amount of water (solvent), whereas a dilute alkaline solution contains a lot of water.
2. An indicator is an organic molecule that has a different colour at different pH values. Examples of indicators are litmus, bromothymol blue, phenolphthalein and methyl orange.
3. Neutralisation occurs at the equivalence point of the reaction where equal molar amounts of the acid and base have reacted according to the molar ratio in the chemical equation.

   Titration is the technique of slowly adding an acid to a base or vice versa in a neutralisation reaction.
4. a) $2NaOH(aq) + H_2SO_4(aq) \rightarrow Na_2SO_4(aq) + 2H_2O(l)$

   b) $\frac{n_a}{n_b} = \frac{c_a V_a}{c_b V_b}$

   $\frac{1 \text{ mol}}{2 \text{ mol}} = \frac{c_a \times 30 \text{ cm}^3}{0.18 \text{ mol·dm}^{-3} \times 70 \text{ cm}^3}$

   $c_a = 0.21 \text{ mol·dm}^{-3}$

   c) $m = cVM$

   $= 0.18 \text{ mol·dm}^{-3} \times 0.070 \text{ dm}^3 \times 40 \text{ g·mol}^{-1}$

   $= 0.504 \text{ g NaOH}$
5. a) $c = \frac{m}{MV} = \frac{14 \text{ g}}{56 \text{ g·mol}^{-1} \times 1 \text{ dm}^3}$

   $= 0.25 \text{ mol·dm}^{-3}$

   b) $2KOH(aq) + H_2SO_4(aq) \rightarrow K\ SO\ (aq) + 2H_2O(l)$

   $\frac{n_a}{n_b} = \frac{c_a V_a}{c_b V_b}$

   $\frac{1 \text{ mol}}{2 \text{ mol}} = \frac{0.10 \text{ mol·dm}^{-3} \times Va}{0.25 \text{ mol·dm}^{-3} \times 30 \text{ cm}^3}$

   $V_a = 37.5 \text{ cm}^3$

   c) The end point will be neutral between a strong acid and a strong base.
6. $2HCl(aq) + MgO(s) \rightarrow MgCl_2(aq) + H_2O(l)$

   mol HCl:

   $na = caVa$

   $= 0.25 \text{ mol·dm}^{-3} \times 0{,}040 \text{ dm}^3$

   $= 0.01 \text{ mol}$

   2 mol HCl reacts with 1 mol MgO (from the balanced equation)

   0.01 mol HCl reacts with $\frac{0.01}{2} = 0.005$ mol MgO

   Mass MgO:

   $mb = nbMb$

   $= 0.005 \text{ mol} \times 40.0 \text{ g·mol}^{-1}$

   $= 0.20 \text{ g}$

How are you doing?

Take this opportunity to ask learners if there is anything that they do not understand. You can check their understanding by asking them some questions about the information covered in the unit. Explain anything that learners do not understand.

Key words

pH – the measurement of a substance's acidity, neutrality or alkalinity (known as 'bases'):

indicator – a compound that has a different colour at different pH values: 72

base – a substance that neutralises an acid: 70

acid – a sour substance that turns litmus red; it has a pH below 7: 70

titration – technique of accurately neutralising an acid with an alkali or vice versa to determine the unknown concentration of an acid or a base: 75

alkali – the Group 1 elements: 5

concentration – the amount of solute present in a given quantity of solvent: 55

neutralisation – neutralisation reaction a reaction between an acid and a base; products can include a salt, water and carbon dioxide: 71

standard solution – a solution of accurately known concentration: 55

heat of neutralisation – heat of the reaction the change in energy that results when reactants form products in a chemical reaction: 80

Checklist for Self-evaluation

Theme 2 Topic 4

EVALUATION GUIDE: Student should be able to:

	Criteria	4	3	2	1	Comments
4.1	define and give examples of solutions					
4.2	prepare a 1,0 $mol \cdot dm^{-3}$, 0,5 dm^3 solution of a named acid and base					
4.3	carry out the acid-base titration using appropriate indicator					
4.4	record titre values					
4.5	perform the calculations required.					

Code for evaluation:			
4 – Very well	3 – Well	2 – Fairly well	1 – Not well at all

Topic 5: Water
Performance objectives
5.1 draw and explain the structure of water **5.2** define the following concepts: solute, solvent, solution **5.3** define solubility and state the rules of solubility in water **5.4** explain factors that affect solubility **5.5** explain the causes of hardness of water **5.6** explain the methods employed in the removal of hardness **5.7** explain methods used in purifying water.

Introduction

The water molecule and the forces between the molecules determine its chemical properties. Water is a universal solvent and dissolves ionic compounds and other polar molecules. Solutions play an important role in our everyday lives and also in the transport of nutrients through the soil.

Activity 5.1: Explore solutions, solutes and solvents

(SB p. 85)

Resources

Pictures from magazines and newspapers of various everyday solutions such as tea, coffee, fizzy drinks, and solvents such as water, ethanol, turpentine, acetone (nail polish remover), methylated spirits, benzene and vinegar

Guidelines

This activity can be done as a class test.

Guide the students on a class discussion about the different types of solutions and solvents they might use at home.

Students should be able to list and explain the different types of solutions.

Answers

1. A solution is a homogeneous mixture of two or more substances.
2. The solvent is the substance in which you want to dissolve another substance. The solute is the substance that dissolves in the solvent.
3. The solution can be separated through physical processes, such as evaporation and crystallization.
4. ethanol, acetone, turpentine, paint thinners, vinegar and methylated spirits

Experiment 5.1: Prepare solutions

(SB p. 86)

Resources

Two glass beakers, 2 spatulas, table salt, copper(II) sulphate, water

Guidelines

Perform an experiment to help students determine the solubility of substances.

Answers

Students should be able to make a dilute, concentrated and saturated solution.

Experiment 5.2: Investigate the solubility of salts

(SB p. 88)

Resources

Test tubes, spot-test tiles or micro-chemical sets, medicine droppers, silver nitrate ($AgNO_3$), barium nitrate ($Ba(NO_3)_2$), solutions of the following salts: Group 1 (sodium chloride (NaCl), sodium bromide (NaBr), sodium iodide (NaI), sodium carbonate (Na_2CO_3), sodium nitrate ($NaNO_3$)), Group 2 (sodium sulphate (Na_2SO_4), magnesium sulphate ($MgSO_4$), sodium carbonate (Na_2CO_3), sodium chloride (NaCl))

Guidelines

This experiment may be carried out with small quantities of solutions in test tubes, or drop by drop with the aid of a medicine dropper and a spot-test tile.

There are also micro-chemical sets available. Use whatever resources are available to your school.

Students should be able to present their findings in a table such as the one in the example on p. 88 of the SB.

Answers

	Group 1				
	NaCl	NaBr	NaI	Na_2CO_3	$NaNO_3$
$AgNO_3$	AgCl	AgBr	AgI	Ag2CO_3	No reaction
	Group 2				
	Na_2SO_4	$MgSO_4$	Na_2CO_3	NaCl	$NaNO_3$
$Ba(NO_3)_2$	$BaSO_4$	$MgSO_4$	Na_2SO_4	No reaction	No reaction

Ag_2CO_3 and $BaCO_3$ dissolve in concentrated HNO_3.

Experiment 5.2: Investigate the solubility of sugar (SB p. 89)

Resources

Sugar, a beaker, jug or mug, water, source of heat

Guidelines

Guide the students to dissolve the sugar in the water.

Students should be able to explain the factors that affect solubility.

Answers

Conclusion: More sugar dissolves faster in hot water than in cold water.

Experiment 5.3: Investigate the effect of heat on solubility (SB p. 89)

Resources

Sugar cubes, two glass beakers, hot water, ice cold water, stirring spoon

Guidelines

Guide the students to dissolve the sugar cubes in the water.

Students should be able to explain the effect of heat on solubility.

Answers

Conclusion: More sugar blocks dissolve in the hot water than in the cold water.

Experiment 5.4: Investigate solubility (SB p. 90)

Resources

Iodine crystals (I_2) (molecular compound), water (H_2O) and ethanol (C_2H_5OH) (polar solvents), carbon tetrachloride (CCl_4) and ether ($C_2H_5OC_2H_5$) (non-polar solvents), four test tubes with stoppers in test tube stand

Guidelines

Guide the students to determine the solubility of molecular and ionic substances in polar and non-polar solutions.

Students should be able to explain the factors that affect solubility.

Answers

Conclusion: We can conclude that polar and ionic substances dissolve in polar solvents and non-polar substances (molecular substances) dissolve in non-polar solvents. The general rule can be summarised as "like dissolves like".

Activity 5.3: Answer questions about solubility (SB p. 92)

Resources

Student's Books, workbook

Guidelines

This activity can be used as a class test.

Answers

1. Similar types of substances will dissolve in each other, e.g. molecular compounds dissolve in non-polar liquids and ionic compounds dissolve in polar liquids.
2. a) water/ether, water/carbon tetrachloride
   b) water/ethanol, carbon tetrachloride/ether
   c) iodine/water, potassium manganate(VII)/carbon tetrachloride
   d) iodine/carbon tetrachloride, potassium manganate(VII)/water
3. ethanol
4. Oil is a polar liquid and water is non-polar
5. a) 23 °C
   b) i) 95 g/100 g water
      ii) 4 g/100 g water
      iii) 86 g/100 g water
   c) $Ce_2(SO_4)_3$
   d) i) $AgNO_3$
      ii) KNO_3
   e) 48 °C
   f) No
   g) 50 °C
   h) The solubility of $Ce_2(SO_4)_3$ decreases as the temperature is increased.

Experiment 5.5: Remove hardness from water

(SB p. 94)

Resources

Six glass beakers, Buchner flask, filter paper, dilute hydrochloric acid, washing soda (Na_2CO_3), spatula, temporary hard water, permanently hard water

Guidelines

Guide a class discussion on the different methods to remove hardness from water.

Demonstrate to the class one removal of hardness in water.

Students should be able to carry out an experiment to remove hardness by boiling the water and adding washing soda.

Answers

Conclusion: The calcium and magnesium ions in permanent hard water react with washing soda (sodium carbonate) to produce insoluble calcium and magnesium carbonates. The sodium ions that are left in the water, are soluble and harmless.

Activity 5.4: Investigate a water treatment plant

(SB p. 97)

Resources

A local water treatment plant

Guidelines

This activity is a class outing.

Lead the students on an excursion to water works to observe the water distillation process.

Students should be able to observe, record, draw and explain the process of obtaining distilled water in the water works.

Answers

The results of the excursion might differ from student to student.

Activity 5.5: Build a sand filter

(SB p. 97)

Resources

A container for the filter (any suitable container, for example a 2ℓ plastic cooldrink bottle or a large tin that can hold about 2ℓ of liquid), fine sand, coarse sand or gravel, small pebbles.

Guidelines

Guide a class discussion on other methods of purifying water.

Students should be able to make an inexpensive water purifier from the materials provided.

Answers

a) The schmutzdecke is a biological layer that forms on the surface of the water that filters slowly through a sand filter. The biological layer is also called a biofilm

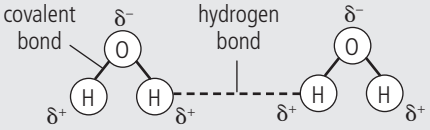
and contains bacteria, fungi, protozoa, rotifer and aquatic insect larvae. These live organisms use the organic material in the dirty water as food and remove the impurities. To maintain the biofilm, the water level must be kept constant during periods where no water is added. The outlet pipe must be a standpipe that maintains the standing level to a constant depth above the top of the filtering sand.

b) Activated carbon is charcoal that is very porous and light. Pumice is volcanic rock that formed from frothy lava. It is also very porous and light. The large surface area that is available for adsorption and chemical reactions remove many gases and other toxins in the water.

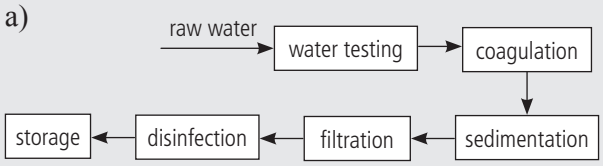
Revision questions: Answers

(SB p. xxx)

1. a) A b) B c) A d) D e) C
2. a) solution
   b) insoluble
   c) solubility graph
   d) boiling
   e) distilled water
3. a)



   b) The polar water molecule can attract and dissolve other polar and ionic substances. The forces between the polar substances and ionic bonds are similar in strength.
   c) Ionic substances e.g. table salt, $KMnO_4$ and KCl, polar liquids e.g. ethanol and methanol
   d) Molecular substances e.g. I_2, non-polar liquids e.g. CCl_4, ether, kerosene
   e) The hydrogen bonds between water molecules allow for large open spaces between the molecules. This makes the ice structure less dense than water, where the water molecules can be closer together.
4. a) table salt
   b) water
   c) dissolution/ dissolving
   d) salt solution
   e) physical change, the solution can be separated through physical separation techniques
5. a) B, dissolution of sugar increases with temperature
   b) B, more sugar dissolves at a higher temperature
   c) A = B, the temperature does not affect the amount of salt that dissolves in the water
   d) D, more solute can dissolve at 50 °C.
6. a) $KClO_4$
   b) Li_2SO_4
   c) KNO_3
   d) 10 g
   e) 48 g
   f) No, only 32 g will dissolve and 2 g will be in excess
7. a)



   b) Chlorine kills the bacteria
      A coagulant brings the dirt particles together to form a floc
      Lime stabilises the pH
   c) Filtration
   d) Activated charcoal removes odours and tastes in the water.
   e) It is tested to make sure all harmful bacteria and viruses (pathogens) are removed thus making it suitable for human consumption.

How are you doing?

Take this opportunity to ask learners if there is anything that they do not understand. You can check their understanding by asking them some questions about the information covered in the unit. Explain anything that learners do not understand.

Key words

hydrogen bond – a force between the hydrogen atom on one molecule and a F, N or O atom on another molecule: 84

covalent – a bond in which two orbitals overlap and the atoms share an electron pair: 83

solubility – the maximum amount of solute that can dissolve in a given amount of solvent at a specific temperature: 85

solution – a mixture in which a liquid, gas or solid is evenly mixed in a liquid (and sometimes a gas): 83

super saturated solution – a solution that contains more solute, or dissolved material, than it would under normal conditions: 86

solubility – the maximum amount of solute that can dissolve in a given amount of solvent at a specific temperature: 85

precipitate – an insoluble solid that separates from a solution: 86

dissolution – the opposite of mixing a solution, whereby the solvent is extracted from the solution, leaving the original solute: 85

solubility curves – a graphic representation of the variation with changing temperature of the solubility of a given substance in a given solvent: 91

purification – to remove an unwanted substance: 95

distillation – the process of heating a liquid so that the gases inside it evaporate and condense, and are extracted: 95

Checklist for Self-evaluation

Theme 2 Topic 5

EVALUATION GUIDE: Student should be able to:

	Criteria	4	3	2	1	Comments
5.1	draw and explain the structure of water					
5.2	define the following concepts: solute, solvent, solution					
5.3	define solubility and state the rules of solubility in water					
5.4	explain factors that affect solubility					
5.5	explain the causes of hardness of water					
5.6	explain the methods employed in the removal of hardness					
5.7	explain methods used in purifying water.					

Code for evaluation:			
4 – Very well	3 – Well	2 – Fairly well	1 – Not well at all

Topic 6: Air
Performance objectives
6.1 list the constituents of air **6.2** list the percentage composition of air **6.3** state the properties of air **6.4** draw, label and explain the various zones of flame.

Introduction

The atmosphere that surrounds the Earth is composed of a mixture of gases that we call air. When the air moves from one place to another, we say the wind is blowing. The atmosphere supports life on Earth. We breathe air, as do many animals and plants. When we breathe in air, our bodies take in oxygen. When we breathe out, we give out carbon dioxide and water vapour. This process is called respiration. Air also supports the wings of birds and aeroplanes and makes it possible for them to fly.

Experiment 6.1: Determine percentage oxygen in air (SB p. 105)

Resources

10 cm^3 measuring cylinder, beaker, iron sponge (steel wool), vinegar, water

Guidelines

Perform an experiment using rust to show the composition of air.

Students should be able to watch teacher's demonstration to state the percentage composition of air, and write a report on it.

Answers

Observation and conclusion: As the iron sponge starts to rust, it uses the oxygen in the air. As the pressure drops inside the measuring cylinder, water will fill the space left by the partial vacuum. The iron sponge rusts according to the following equation:

$4Fe + 3O_2 \rightarrow 2Fe_2O_3$

As the oxygen in the air in the measuring cylinder reacts with iron to form iron oxide, the volume of the trapped air decreases and water enters the measuring cylinder. This change in volume is equal to the volume of oxygen used in the reaction.

You can read off the volume of oxygen used on the side of the measuring cylinder. It should be about 2 cm^3. If you started off with 10 cm^3 of water, the final volume will be 8 cm^3. It means that the percentage oxygen in air is 0%.

Experiment 6.2: Determine percentage oxygen in air (SB p. 106)

Resources

Two 100 cm^3 gas syringes, retort stands and clamps, hard glass (pyrex) connecting tube fitted into rubber connectors, Bunsen burner, copper turnings, wooden splint

Guidelines

Perform an experiment using oxidation to show the composition of air.

Students should be able to watch the teacher's demonstration of the experiment to state the percentage composition of air, and be able to write a report on it.

Answers

Observation and conclusion: Copper turnings react with oxygen in the air to form black copper(II) oxide. The amount of gas in a closed system will decrease during the oxidation process as the oxygen is used up. The balanced equation is:

$2Cu(s) + O_2(g) \rightarrow 2CuO(s)$

We can calculate the percentage oxygen in the air from the difference in volume in the syringes before and after the oxidation process.

The difference in volume in syringe A should be 20 cm^3. If you started with 100 cm^3, the final reading on the syringe should be about 80 cm^3. This is in accordance to the percentage oxygen in the air, which is 21%.

Activity 6.1: Demonstration to prove that air occupies space (SB p. 107)

Resources

Tissue, plastic cup, large container of water

Guidelines

Do a simple demonstration to prove that air occupies space by showing how a tissue can be submerged in water without getting wet. Students should be able to state the properties of air.

Answers

Conclusion: The tissue stays dry inside the cup. The water cannot enter the cup, because it is already full with air that occupies space.

Experiment 6.3: Proving that air has mass and determining its density (SB p. 107)

Resources

Two balloons, bicycle pump, large bucket (that the inflated balloon will fit into), water, balance, measuring cylinder, ruler, string, needle

Guidelines

Perform a simple experiment to prove that air has mass and determine its density.

Students should be able to prove that air has mass.

Answers

Students may have different answers to each other, depending on the volume of the inflated balloon.

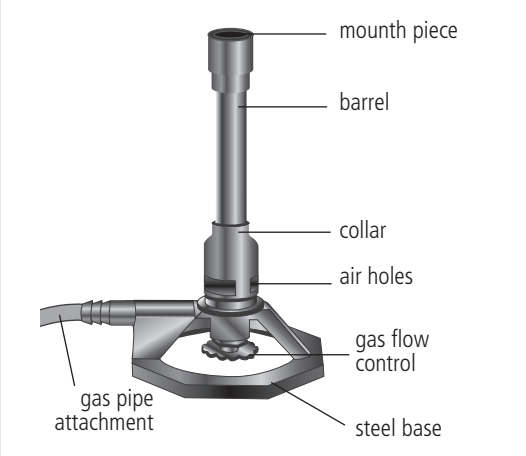
Revision questions: Answers

(SB p. 112)

1. a) B b) D c) B d) A
2. a)

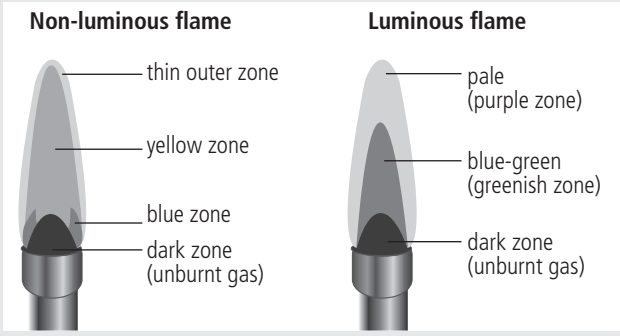
Constituents	Percentage
Nitrogen	78%
Oxygen	21%
Trace gases	1%

   b) Argon, carbon dioxide, neon, methane, helium, krypton, hydrogen, xenon
   c) Water vapour is a variable gas and it percentage composition in the atmosphere is not always the same.
3. Air can be compressed, is invisible, has mass and volume, exerts a pressure in all directions, can vary in temperature, is a mixture and its composition is constant, the amount of water vapour can vary.
4.



Gas pipe attachment: allows gas to flow into the burner
Steel base: heavy to keep the burner steady and prevent it from falling over
Gas flow control: increasing or decreasing the gas flow by turning the dial
Collar with air holes: holes are closed to change from luminous flame to non-luminous flame
Barrel: channels the gas
Mouth piece: contains the flame

5. A luminous flame is produced when the air holes of the Bunsen burner are closed, it is yellow, produces smoke and burns black, is unsteady, tall and elongated, not very hot, does not make a noise, used for warmth
   A non-luminous flame is produced when the air holes are open, does not produce smoke, is very hot, is steady, short and cone-shaped, blue in colour, makes a noise, used for high heat
6.



How are you doing?

Take this opportunity to ask learners if there is anything that they do not understand. You can check their understanding by asking them some questions about the information covered in the unit. Explain anything that learners do not understand.

Key words

luminous – producing or seeming to produce light: 110

non-luminous – not capable of producing light, but can be capable of reflecting light from another source: 110

combustion – a class of chemical reaction that occurs when hydrocarbon reacts with oxygen to produce carbon: 102

Checklist for Self-evaluation

Theme 2 Topic 6

EVALUATION GUIDE: Student should be able to:

	Criteria	4	3	2	1	Comments
6.1	list the constituents of air					
6.2	state the percentage composition of the various constituents of air					
6.3	state the properties of air					
6.4	draw the flame and label the different parts.					

Code for evaluation:			
4 – Very well	3 – Well	2 – Fairly well	1 – Not well at all

Topic 7: Hydrogen
Performance objectives
7.1 write and draw the electron configuration of hydrogen **7.2** identify the isotopes of hydrogen and draw their electron configuration **7.3** explain the unique position of hydrogen in the Periodic Table **7.4** describe the laboratory and industrial methods for the preparation of hydrogen **7.5** state the physical and chemical properties of hydrogen **7.6** state the uses of hydrogen.

Introduction

During the sixteenth century, Paracelsus, a Swiss-German physician, noted that a flammable gas was formed when iron reacted with sulphuric acid. However, he did not realise that the gas was a pure substance. In 1766, Henry Cavendish determined that the flammable material was a distinct substance.

He produced hydrogen by having various metals react with acids. He also found that hydrogen produced water when it combusted in air. It was Lavoisier who named the gas "hydrogen", which is derived from the Greek word for "water-former".

Activity 7.1: Test your knowledge of hydrogen

(SB p. 114)

Resources

Periodic Table, Student's Book and workbook

Guidelines

This activity can be used as a class test.

Guide students to write and draw the electron configuration of hydrogen, name the isotopes of hydrogen and write their electron configuration, and explain the position of hydrogen on the Periodic Table.

Students should be able to: Write and draw the electron configuration of a hydrogen atom, identify the isotopes of hydrogen and draw models of the atoms, and explain the unique position of hydrogen in the Periodic Table.

Answers

1. one electron in a 1s orbital: 1s1
2. Three isotopes: Protium: 1H, Deuterium: 2H, Tritium: 3H
3. Hydrogen can react with metals and non-metals. With metals it shows a negative oxidation state of −1 similar to the halogens of group 17, and in reaction with the non-metals it shows a +1 oxidation state similar to the alkali metals of group 1. Its position can be either as part of group 1 or group 17.

Experiment 7.1: Prepare hydrogen gas

(SB p. 115)

Resources

Test tube, conical flask with stopper and glass tube outlet, zinc metal granules (Zn), dilute hydrochloric acid (HCl) or sulphuric acid (H_2SO_4), wooden splint and match.

Guidelines

Set up the apparatus for the laboratory preparation of hydrogen.

Safety rules: Be very careful, because a mixture of hydrogen and air is explosive. Extinguish all open flames and prepare just enough hydrogen for a positive test of hydrogen.

Students should be able to record observations and draw apparatus for the laboratory preparation of hydrogen.

Answers

Observations and conclusions:

The zinc reacts with the acid to release bubbles of hydrogen.

Hydrogen is less dense than air and will collect at the top of the test tube.

The equations for the reaction are:

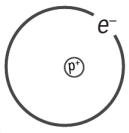
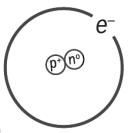
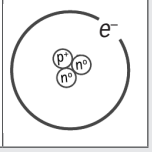
$Zn(s) + 2HCl(aq) \rightarrow ZnCl_2(aq) + H_2(g)$

$Zn(s) + H_2SO_4(aq) \rightarrow ZnSO_4(aq) + H_2(g)$

Revision questions: Answers

(SB p. 119)

1. a) C b) A c) D d) D e) B
2.

	Hydrogen (99.985%)	Deuterium (0.015%)	Tritium (radio-active)
	H	H	H
protons	1	1	1
neutrons	0	1	2
			

3. a) −1
   b) +1
   c) −1
   d) +1
4. a) $2Li(s) + 2H_2O(l) \rightarrow H_2(g) + 2LiOH(aq)$
   b) $Mg(s) + H_2O(g) \rightarrow MgO(s) + H_2(g)$
   c) $Mg(s) + H_2SO_4(aq) \rightarrow MgSO_4(aq) + H_2(g)$
   d) $Zn(s) + 2HCl(aq) \rightarrow H_2(g) + ZnCl_2(aq)$
   e) $3Fe(s) + 4H_2O(g) \rightarrow Fe_3O_4(s) + 4H_2(g)$
5. a) Reaction between zinc and dilute hydrochloric acid
   b) Reaction between steam and iron
   Steam-reforming of natural gas or propane
   Electrolysis
   Reaction between steam and carbon
6. The hydrogen is very pure, it is energy intensive, making it expensive
7. a) $3H_2(g) + N_2(g) \rightarrow 2NH_3(g)$
   b) $2H_2(g) + O_2(g) \rightarrow 2H_2O(l)$
   c) $H_2(g) + Cl_2(g) \rightarrow 2HCl(g)$
8. • as reducing agent in the reduction of metal oxides
   • to process crude oil into refined fuels
   • to purify fuels and remove sulphur
   • to produce ammonia for agricultural fertiliser
   • to make cyclohexane and methanol
   • in the hydrogenation of oil to make margarine
   • as a protective atmosphere for making flat glass sheets and silicon chips
   • as a fuel for oxy-hydrogen blowpipes
   • with oxygen as rocket fuel
   • in fuel cells

How are you doing?

Take this opportunity to ask learners if there is anything that they do not understand. You can check their understanding by asking them some questions about the information covered in the unit. Explain anything that learners do not understand.

Key words

isotopes – an atom of an element that has the same number of protons, but a different number of neutrons to one or more of that element's other atoms: 113

oxidation number – the number of charges the atom would have in a molecule, if electrons were transferred completely 114

downward displacement of water – a method used in chemistry for collecting gases; the water is displaced (pushed out), of where it is by a particular chemical reaction that produces a gas and the pressure drives the water away: 114

Checklist for Self-evaluation

Theme 2 Topic 7

EVALUATION GUIDE: Student should be able to:

	Criteria	4	3	2	1	Comments
7.1	describe the laboratory preparation of hydrogen					
7.2	list four methods for the industrial preparation of hydrogen					
7.3	state the physical and chemical properties of hydrogen					
7.4	list the uses of hydrogen.					

Code for evaluation:			
4 – Very well	3 – Well	2 – Fairly well	1 – Not well at all

Topic 8: Oxygen
Performance objectives
8.1 explain the general properties of the oxygen group in the Periodic Table **8.2** write and draw the electron structure of oxygen and explain its bonding capacity **8.3** describe the laboratory and industrial methods for the preparation of oxygen **8.4** state the physical and chemical properties of oxygen **8.5** list the compounds of oxygen **8.6** explain oxidation as addition of oxygen and give examples of such reactions **8.7** state the uses of oxygen.

Introduction

Group 16 of the Periodic Table of elements form the oxygen group and is also called the chalcogens.

The rest of this topic deals with various aspects of oxygen.

Activity 8.1: Answer questions about some members of the oxygen family

(SB p. 122)

Resources

Periodic Table, Student's Book and workbook

Guidelines

This activity can be used as a class test.

Explain the general properties of the oxygen group in the periodic table.

Students should be able to answer questions about some members of the oxygen family.

Answers

1. Photoconductivity is when the electrical conductivity of an element increases when light is shone on it. Both selenium and tellurium display this property.
2. six
3. A compound that contains oxygen in the −1 oxidation state.
4. Metal: polonium
   Metalloid: tellurium
   Non-metals: oxygen, sulphur and selenium
5. −2
6. oxygen

Experiment 8.1: Preparation of oxygen

(SB p. 124)

Resources

Student's Book, workbook, diagrams on p. 125 of SB, and laboratory apparatus and reagents as listed below for Method 1 and Method 2.

Guidelines

Method 1:
Bunsen burner, hard glass (Pyrex) test tube with stopper, retort stand with test tube clamp, delivery tube, water trough, gas cylinders, beehive shelf, potassium chlorate ($KClO_3$), manganese(IV) dioxide (MnO_2), balance

Method 2:
250 ml conical flask (Erlenmeyer flask) with stopper, retort stand with clamp, thistle tube, delivery tube, water trough, gas cylinders, beehive shelf, measuring cylinders, water, 3% hydrogen peroxide (H_2O_2), manganese(IV) dioxide (MnO_2), balance

Guide students to follow the safety rules exactly.

Regulate the heat applied to the mixture carefully. The decomposition process is exothermic and heat will be given off. Remove the flame when the gas is given off too violently. Under certain conditions, too much heat can lead to an explosion!

Students should be able to observe the experiment on the laboratory preparation of oxygen, and record their observations correctly in their workbook.

Answers

Observations and Conclusions

When $KClO_3$ and MnO_2 are mixed and heated, the $KClO_3$ decomposes to liberate O_2.

H_2O_2 decomposes to form H_2O and O_2 gas.

The MnO_2 acts as catalyst and does not decompose or take part in the reaction. It may be recovered by dissolving the potassium chloride from the residue with water.

Activity 8.2: Answer questions about the preparation of oxygen (SB p. 126)

Resources

Test tube, conical flask with stopper and glass tube outlet, zinc metal granules (Zn), dilute hydrochloric acid (HCl) or sulphuric acid (H_2SO_4), wooden splint and match.

Guidelines

Set up the apparatus for the laboratory preparation of hydrogen.

Safety rules: Be very careful, because a mixture of hydrogen and air is explosive. Extinguish all open flames and prepare just enough hydrogen for a positive test of hydrogen.

Students should be able to record observations and draw apparatus for the laboratory preparation of hydrogen.

Answers

1. A catalyst is a chemical compound that speeds up a reaction without undergoing permanent change in the process.
2. The glowing splint will burst into flames.
3. $H_2O_2(aq) \rightarrow H_2(g) + O_2(g)$
4. Magnesium burns much faster and fiercer in pure oxygen, compared to air.

Experiment 8.2: Investigate reactions of metals with oxygen (SB p. 130)

Resources

Student's Book, workbook, samples of lithium, potassium, sodium, calcium, magnesium, iron, copper, aluminium and zinc, nine gas jars, nine combustion spoons (deflagrating spoons), nine glass cover slips, petroleum jelly, Bunsen burner, oxygen, distilled water, universal indicator

Guidelines

Students should be able to observe the experiment and record their observations correctly in their workbook, using a table format such as the example on p. 131 of the SB.

Answers

Observations and Conclusions

Experiment 8.3: Investigate reactions of solid non-metals with oxygen (SB p.131)

Resources

Student's Book, workbook, Samples of carbon (charcoal powder), red phosphorus and flowers of sulphur, three gas jars, three combustion spoons, three glass cover slips, petroleum jelly, Bunsen burner, oxygen, distilled water, universal indicator.

Guidelines

Guide the students to follow the same procedure as for the reactions of metals in Experiment 8.2.

Students should be able to observe the experiment and record their observations correctly in their workbook, using a table format such as the example on p. 131 of the SB.

Answers

Observations and conclusions

Element	Flame colour	Products formed	Equation	Nature of solution
Carbon	Yellow	Colourless gas Dissolves in water	$C(s) + O_2(g) \rightarrow CO_2(g)$	Acidic
Phosphorus	Blue	Colourless gas Dissolves in water	$P_4(s) + 5O_2(g) \rightarrow 2P_2O_5(s)$	Acidic
Sulphur	Yellow	Dense white fumes Dissolves in water	$S(s) + O_2(g) \rightarrow SO_2(g)$	Acidic

Experiment 8.4: Investigate reactions of gaseous non-metals with oxygen

(SB p. 132)

Resources

Student's Book, workbook, Test tube, zinc, dilute hydrochloric acid, burning wooden splint, Bunsen burner

Guidelines

Demonstrate this experiment using the method on p. 132 of the SB.

Students should be able to observe the experiment and record their observations correctly in their workbook.

Answers

Observations and conclusions

The reaction between zinc and hydrochloric acid releases hydrogen gas:

$Zn(s) + 2HCl(aq) \rightarrow ZnCl_2(aq) + H_2(g)$

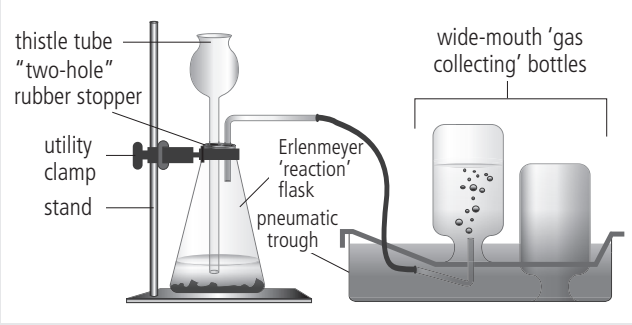
The hydrogen gas burns with a blue flame and a popping sound can be heard. We use this reaction to test for the presence of hydrogen gas.

We use liquefied petroleum gas (LPG) as town gas or cooking gas. The gas contains a mixture of compounds, such as propane and butane. This is the gas that produces the blue flame in a Bunsen burner.

Revision questions: Answers

(SB p. 135)

1. a) D b) B c) A d) C e) D
2. a) oxide
   b) electrolysis
   c) fractional distillation
   d) ozone
   e) respiration
3. a) Oxygen: O, 8, sulphur: S, 16, selenium: Se, 34, tellurium: Te, 52, polonium: Po, 84
   b) Oxygen, sulphur and selenium: non-metals, tellurium: metalloid, polonium: metal
   c) 6
   d) −2
   e) Selenium: semiconductor and is used in the manufacture of rectifiers, used to make photocells, used in laser printers and copiers, used in glassmaking, used in pigments
   Tellurium: semiconductor used in memory chips for computers, used in metallurgy, used in solar panels
   Polonium: radioactive and used in nuclear weapons
4. a) Oxygen: $1s^2\ 2s^2\ 2p^4$ or [He] $2s^2\ 2p^4$
   b) oxygen: O_2, ozone: O_3
   c) Ozone is formed when oxygen is radiated with ultraviolet rays and the oxygen molecules are broken down to oxygen atoms.
   $O_2(g) \rightarrow 2O\cdot(g)$
   Some of the energized oxygen atoms react with other oxygen molecules to form ozone.
   $O\cdot(g) + O_2(g) \rightarrow O_3(g)$
   i) O_2: 0
   ii) MgO_2: −1
   iii) H_2SO_4: −2
   iv) H_2O_2: −1
   v) KO_2: −1/2
   vi) F_2O: +2
5. Photochemical dissociation:
   $2H_2O(g) \rightarrow O_2(g) + 2H_2(g)$
   Photosynthesis:
   $6CO_2(g) + 6H_2O(l) \rightarrow C_6H_{12}O_6(s) + 6O_2(g)$
6. Decomposition of potassium chlorate:
   $2KClO_3(s) \rightarrow 2KCl(s) + 3O_2(g)$
   Decomposition of hydrogen peroxide:
   $2H_2O_2(aq) \rightarrow 2H_2O(l) + O_2(g)$
7. a) Any acid, alkali or ionic salt
   b) $2H_2O(l) \rightarrow 2H_2(g) + O_2(g)$
   c) Oxygen is produced at the positive anode.
8. a)



   b) A glowing wooden splint will catch alight.
9. a) Boiling point
   b) 1. The air is filtered to remove the dust particles.
      2. Carbon dioxide and water vapour are removed.
      3. The air is cooled under pressure to liquefy and passed through a fractional distillation column.
      4. Nitrogen has a lower boiling point and boils off first. Oxygen has a higher boiling point and collects closer to the bottom of the column.
      5. The cold gases are cooled and stored.
10. a) Non-metal: C (or any other example)
    Metal: Na or Fe (or any other example)

b) $C(s) + O_2(g) \rightarrow CO_2(g)$
$4Na(s) + O_2(g) \rightarrow 2Na_2O(s)$
$3Fe(s) + 2O_2(g) \rightarrow Fe_3O_4(s)$

c) CO_2 soluble – acidic solution
Na_2O soluble – basic solution
Fe_3O_4 insoluble

11. $CH_4(g) + 2O_2(g) \rightarrow CO_2(g) + 2H_2O(g)$

12. a) $4Al(s) + 3O_2(g) \rightarrow 2Al_2O_3(s)$
b) $2Mg(s) + O_2(g) \rightarrow 2MgO(s)$
c) $F_2(g) + O_2(g) \rightarrow F_2O_2(s)$
d) $2Cl_2(g) + 7O_2(g) \rightarrow 2Cl_2O_7(l)$
e) $3Fe(s) + 2O_2(g) \rightarrow Fe_3O_4(s)$
f) $4Li(s) + O_2(g) \rightarrow 2Li_2O(s)$
g) $2C(s) + O_2(g) \rightarrow 2CO(g)$
h) $2N_2(g) + O_2(g) \rightarrow 2N_2O(g)$

How are you doing?

Take this opportunity to ask learners if there is anything that they do not understand. You can check their understanding by asking them some questions about the information covered in the unit. Explain anything that learners do not understand.

Key words

chalcogens – the chemical elements in Group 16 of the Periodic Table; the group is also known as the oxygen family; it consists of the elements oxygen (O), sulphur (S), selenium (Se), tellurium (Te), and the radioactive element polonium (Po): 121

diatomic – a molecule made of two atoms of an element bonded together: 121

allotropes – forms of the same element with different structures, e.g. graphite and diamond: 121

semiconductors – elements that normally cannot conduct electricity, but that can have their conductivity improved by either raising the temperature or by adding certain impurities: 122

conductivity – the ability or power to conduct or transmit heat, electricity, or sound: 122

photoconductivity – electrical conductivity that is affected by exposure to electromagnetic radiation (as light): 122

rusting – the chemical reaction between iron and oxygen to form brown flakes of iron oxide that weaken the material: 133

photosynthesis – the process in which green plants synthesise glucose from CO_2 and H_2O in the presence of sunlight: 124

Checklist for Self-evaluation

Theme 2 Topic 8

EVALUATION GUIDE: Student should be able to:

	Criteria	4	3	2	1	Comments
8.1	write and draw the electron configuration of oxygen					
8.2	describe the laboratory preparation of oxygen					
8.3	describe the fractional distillation of liquid air					
8.4	state the physical and chemical properties of oxygen					
8.5	list the uses of oxygen					
8.6	write and balance equations of reactions of oxygen with other elements e.g. Mg, Fe, etc.					

Code for evaluation:			
4 – Very well	3 – Well	2 – Fairly well	1 – Not well at all

Topic 9: Halogens
Performance objectives
9.1 write the electronic configuration of halogens (Cl_2, Br_2, I_2) **9.2** state the physical properties of halogens and their gradation down the group **9.3** state the chemical properties of halogens and their gradation down the group **9.4** list some compounds of halogens **9.5** state the uses of halogens **9.6** describe the laboratory preparation of chlorine **9.7** demonstrate the bleaching action of chlorine and adduce the reason for the bleaching action.

Introduction

The elements in Group 17 of the Periodic Table of elements are known as the halogens. There are five members of the group: fluorine (F), chlorine (Cl), bromine (Br), iodine (I) and astatine (At). The name "halogen" means "salt-producing". When halogens react with metals, they produce a wide range of salts.

Of the halogens, chlorine has the largest distribution on Earth. You will not find it in its native form, but always as part of compounds, of which table salt (NaCl) is the most well-known. All of the halogens form acids when bonded to hydrogen.

Activity 9.1: Answer questions about halogens

(SB p. 139)

Resources

Periodic Table, Table showing physical properties of halogens and similarities among halogens, Student's Book and workbook

Guidelines

This activity can be used as a class test.

Explain the general properties of the halogen group in the periodic table.

Students should be able to answer questions about halogens.

Answers

1. Fluorine (F), chlorine (Cl), bromine (Br), iodine (I) and astatine (As)
2. Group 17
3. Fluorine and chlorine: gases, bromine: liquid, iodine: solid
4. Astatine is very rare and all its isotopes are radioactive and quickly decay into other elements.
5. Sodium chloride or common table salt
6. a) Electronegativity is the ability of an atom in a molecule to attract the bonding electron pairs.
   b) Fluorine
   c) F: 4.0, Cl: 3.0, Br: 2.8, I: 2.5, At: 2.1
7. a) Ionization energy is the amount of energy required to remove an electron from a neutral atom.
   Electron affinity is the amount of energy released when an atom accepts an electron to form an anion.
   b) The atoms of halogens have a high tendency to accept electrons from atoms of other elements. They also do not lose their own electrons easily. They form ions easily or bond with any other atom that can form a bonding pair with their unpaired p-electron.

Experiment 9.1: Test for halogens

(SB p. 140)

Resources

Student's Book, workbook, four test tubes in a test tube stand, potassium chloride (KCl), potassium bromide (KBr), potassium iodide (KI), sodium carbonate (Na_2CO_3), water, silver nitrate, medicine dropper, spatula, ammonia solution

Guidelines

Students should be able to observe the experiment and record their observations correctly in their workbook.

Answers

Observations and Conclusions

1. Silver nitrate forms precipitates with halides and carbonates. The reactions are:
   - $KCl(aq) + AgNO_3(aq) \rightarrow AgCl(s) + KNO_3(aq)$
   - $KBr(aq) + AgNO_3(aq) \rightarrow AgBr(s) + KNO_3(aq)$
   - $KI(aq) + AgNO_3(aq) \rightarrow AgI(s) +$
   - $KNO_3(aq)$
   - $Na_2CO_3(aq) + 2AgNO_3(aq) \rightarrow Ag_2CO_3(s) + 2NaNO3(aq)$
2. The colours of the precipitates vary:
   - $AgCl(s)$ is white
   - $AgBr(s)$ is cream
   - $AgI(s)$ is light yellow
   - $Ag_2CO_3(s)$ is brilliant white
3. The colours of silver bromide and silver iodide are very similar. To distinguish between them, we add ammonia solution. Silver bromide reacts with ammonia to form a soluble complex, whereas silver iodide does not react and remains a precipitate.

Experiment 9.2: Reaction of halides with metals (SB p. 142)

A Reactions with chlorine

Resources

Three gas cylinders, with greased cover slips, filled with chlorine gas, two combustion spoons, sodium, magnesium, iron sponge (steel wool), some flower petals, a small piece of coloured fabric. (You can find methods for the preparation of chlorine gas on page 147.)

B Reactions of iodine (teacher demonstration)

Resources

Asbestos or fire-resisting plate, tweezers or tongs, sharp knife, medicine dropper, pestle and mortar, sodium, magnesium powder, iodine

Guidelines

Guide students to follow the safety rules exactly.

Students should be able to do both experiments, and record their observations correctly in their workbook.

Answers

Observations and Conclusions

A

- Chlorine gas reacts violently with the metals.
- Sodium burns with a bright yellow flame and forms thick white fumes of sodium chloride. Much energy, in the form of heat and light, is released.
  $2Na(s) + Cl_2(g) \rightarrow 2NaCl(s)$
- Magnesium burns energetically with a bright white flame and white magnesium chloride forms. Energy is released as light and heat.
  $Mg(s) + Cl_2(g) \rightarrow MgCl_2(s)$
- The iron sponge glows yellow and dark red-brown vapours of iron(III) chloride form.
  $2Fe(s) + 3Cl_2(g) \rightarrow 2FeCl_3(s)$

B

1. The sodium-iodine mixture burns with a bright yellow flame and excessive purple iodine vapours are liberated. Yellow sparks shoot out in all directions.
   $2Na(s) + I_2(s) \rightarrow 2NaI(s)$
2. The reaction with magnesium is less violent and the colour of the flame is a dark reddish-purple.
   $Mg(s) + I_2(s) \rightarrow MgI_2(s)$
3. Heat is generated in both reactions, so both reactions are exothermic.
4. Water will catalyse the reactions.

Experiment 9.3: Test the reactivity of halogens (SB p. 144)

Resources

Student's Book, workbook, nine test tubes in a test tube stand, chlorine water, bromine water, iodine solution, chloroform ($CHCl_3$) or xylene (C_8H_{10}), potassium chloride (KCl), potassium bromide (KBr), potassium iodide (KI), water.

(Household bleach can be used instead of chlorine water, as it contains enough free chlorine in solution.)

Guidelines

Distinguish between colours of different halogens in a solvent:

- Displacement reactions of chlorine with halide solutions
- Displacement reactions of bromine with halide solutions

Students should be able to observe the experiments and record their observations correctly in their workbook.

Answers

Observations and Conclusions

1. The halogens display different colours when they dissolve in non-polar solvents. We use the colour formation to detect which halogen is in solution.
2. Chlorine can displace bromide and iodide ions from a solution.
3. Bromine can displace iodide ions from a solution.
4. The reactivity of halogens decreases from chlorine to iodine: $Cl_2 > Br_2 > I_2$

Case study: Chlorine (SB p. 146)

Answers

a) Chlorine is a bleaching agent and will bleach the colours in the costume, which will soon fade.
b) Water contains many pathogens that can make us ill and cause diseases, such as typhoid and gastro-enteritis. Chlorine oxidizes the bacteria and kills them.
c) Student's own research. Ancient methods were sufficient because the human population was so much smaller and contamination of water sources was not a factor.
d) Student's own ideas. No upliftment and development can be effective if people have to spend hours collecting water for domestic use. Water-borne diseases cause many people, especially very young and very old, to become ill. Other family members must car for them and they cannot go to work to earn wages.
e) Hydrogen peroxide can also be used as an oxidizing agent, but it is too unstable to use for commercial purposes. It will quickly decompose and there will be no residual H_2O_2 to kill newly introduced pathogens. (Refer back to previous topic on oxygen.)

Experiment 9.4: Preparation of chlorine (SB p. 148)

Resources

Student's Book, workbook, Bunsen burner, retort stand and clamp, round-bottomed flask, thistle funnel, delivery tube, conical flask (Erlenmeyer flask), manganese(IV) oxide (MnO_2) or potassium permanganate ($KMnO_4$), concentrated hydrochloric acid (HCl), concentrated brine solution (NaCl)

Guidelines

Set up the apparatus and demonstrate the laboratory preparation of chlorine.

Students should be able to observe the experiment and record their observations correctly in their workbook.

Revision questions: Answers

(SB p. 150)

1. a) C b) A c) D d) C e) A
2. a) electron configuration
   b) halogens
   c) electronegativity
   d) chlor-alkali process
   e) MnO_2 or $KMnO_4$
3. a) i) seven
      ii) one
   b) Chlorine: $1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^5$ or [Ne] $3s^2\ 3p^5$
   c) −1
4. The halogens have high ionization energies and high electron affinities. It means that halogen atoms do not easily lose electrons and that a large amount of energy is released when they accept an electron. These properties make them very reactive and they are not found in elemental form in nature.
5. Complete the table.

Name	Formula	State at STP	Appearance
Fluorine	F_2	Gas	Yellow-green gas
Chlorine	Cl_2	Gas	light green gas
Bromine	Br_2	Liquid	reddish-brown liquid
Iodine	I_2	Solid	dark purple crystals

6. a) The magnitude of reactivity decreases from fluorine to iodine.
   b) The colour intensity increases from fluorine to iodine.
7. a) $Br_2(l) + H_2(g) \rightarrow HBr(g)$
   b) $F_2(g) + 2Na(s) \rightarrow 2NaF(s)$
   c) $Cl_2(g) + Mg(s) \rightarrow MgCl_2(s)$
   d) $2B(s) + 3F_2(g) \rightarrow 2BF_3(g)$
   e) $3I_2(s) + N_2(g) \rightarrow 2NI_3(s)$
8. a) Exothermic, because energy in the form of heat and light is given off.
   b) $2Fe(s) + 3Cl_2(g) \rightarrow 2FeCl_3(s)$
9. a) Add a few drops of water, water is a catalyst that will start the reaction
   b) The powders react and a dark purple flame and purple vapours can be seen.
   c) $Mg(s) + I_2(s) \rightarrow MgI_2(s)$
10. a) Iodide
    b) $2I^-(aq) + Cl_2(aq) \rightarrow 2Cl^-(aq) + I_2(aq)$
    c) Halogens form diatomic molecules that are non-polar. The molecules of water are polar and the CCl_4 molecules are non-polar. Like dissolves like, and non-polar halogens will dissolve much better in non-polar solvents such as CCl_4.
    d) The reactivity of halogens decrease from chlorine to iodine: $Cl_2 > Br_2 > I_2$

How are you doing?

Take this opportunity to ask learners if there is anything that they do not understand. You can check their understanding by asking them some questions about the information covered in the unit. Explain anything that learners do not understand.

Key words

alchemists – a person who studies or practices alchemy; alchemy is a medieval chemical science which aimed to change base metals into gold, discover a universal cure for disease, and discover a means of indefinitely prolonging life: 138

carcinogenic – a substance or agent causing cancer: 140

Checklist for Self-evaluation

Theme 2 Topic 9

EVALUATION GUIDE: Student should be able to:

	Criteria	4	3	2	1	Comments
9.1	write and draw the electron structure of halogens					
9.2	describe the laboratory preparation of chlorine					
9.3	state the physical properties of halogens					
9.4	state the chemical properties of halogens					
9.5	explain the gradation of physical and chemical properties of halogen down the group					
9.6	name some natural compounds of halogens					
9.7	state the uses of halogens.					

Code for evaluation:			
4 – Very well	3 – Well	2 – Fairly well	1 – Not well at all

Topic 10:Nitrogen
Performance objectives
10.1 state the general properties of Group 15 elements **10.2** explain the laboratory preparation of nitrogen **10.3** explain the industrial preparation of nitrogen from liquid air **10.4** list the properties of nitrogen **10.5** outline the uses of nitrogen **10.6** list the oxides of nitrogen **10.7** explain the nitrogen cycle **10.8** explain the Haber process for the preparation of ammonia **10.9** state the uses of ammonia.

Introduction

The nitrogen family can be found in Group 15 of the Periodic Table of elements. It is also known as the pnictogens. This word is derived from the ancient Greek word meaning 'to choke', which refers to the choking smell of nitrogen gas. The family consists of the elements nitrogen (N), phosphorus (P), arsenic (As), antimony (Sb) and bismuth (Bi).

Activity 10.1: Test yourself (SB p. 154)

Resources

Periodic Table, Student's Book and workbook

Guidelines

This activity can be used as a class test.

Explain the general properties of the nitrogen family in the periodic table.

Students should be able to answer questions about members of the nitrogen family.

Answers

1. They are both allotropes of phosphorus and have the same electron configuration, red phosphorus is stable and not very reactive and has a polymeric structure, white phosphorus consists of P4 molecules, reacts very violently and is poisonous
2. $1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^6\ 4s^2\ 3d^{10}\ 4p^3$ or [Ar] $4s^2\ 3d^{10}\ 4p^3$
3. Bismuth has metallic properties.
4. Nitrogen and phosphorus are non-metals.
5. Bismuth has the largest atomic radius.
6. Nitrogen is a colourless gas.
7. Nitrogen has the highest first ionisation energy and does not easily lose an electron.
8. Common oxidation state for Group 15 elements is −3.

Experiment 10.1: Prepare nitrogen (SB p. 155)

Resources

Charts tabulating the properties of nitrogen group laboratory preparation of nitrogen, Student's Book, workbook, diagrams on p. 125 of SB, and laboratory apparatus and reagents as listed below for Method 1 and Method 2.

Method 1:
Ammonium hypochlorite ($CaOCl_2$, bleaching powder), water, concentrated ammonia solution (NH_3), balance, flask with stopper and side-arm, funnel with filter paper, Bunsen burner, tripod and wire gauze, delivery tube, gas cylinder, water trough

Method 2:
Same as for Method 1. The mixture must be heated very carefully.

Guidelines

Guide students to follow the safety rules exactly.

Students should be able to observe the experiments on the laboratory preparation of nitrogen, and record their observations correctly in their workbook.

Answers

Observations and Conclusions

Method 1

1. Nitrogen is prepared by the reaction between $CaOCl_2$ and NH_3:
   $2NH_3(aq) + 3CaOCl_2(aq) \rightarrow N_2(g) + 3H_2O(l) + 3CaCl_2(aq)$
2. By-products, such as chloramine and nitrogen trichloride, may form. These are poisonous and explosive.

Method 2

1. $NH_4NO_2(aq) \rightarrow N_2(g) + 2H_2O(l)$ or $NH_4Cl(s) + NaNO_2(s) \rightarrow N_2(g) + NaCl(s) + 2H_2O(g)$
2. Small amounts of impurities (NO and HNO_3) may also form. These can be removed by passing the gas through aqueous sulphuric acid, containing potassium dichromate.
3. Nitrogen is collected by downward displacement of air or water.

Experiment 10.2: Preparation of ammonia

(SB p. 159)

Resources

Student's Book, workbook, Bunsen burner, retort stand and clamp, small round-bottomed flasks or test tubes, delivery tube, ammonium chloride (NH_4Cl), calcium hydroxide ($Ca(OH)_2$), optional calcium oxide drying column

Guidelines

Demonstrate the test for ammonia gas using

- A damp red litmus paper
- Conc. HCL

Safety Note: Please note that this should only be carried out in a fume cupboard and only by a teacher.

Students should be able to observe the experiment and record their observations correctly in their workbook.

Answers

Observations and Conclusions

1. The reaction between calcium hydroxide and ammonium chloride is:
   $Ca(OH)2(s) + 2NH_4Cl(s) \rightarrow CaCl_2(s) + 2H_2O(l) + 2NH_3(g)$
2. Ammonia gas has a choking smell and must be prepared in a well-ventilated room or fume cupboard.
3. Check when the container is full by holding a wet red litmus paper at its mouth. It will turn blue when the container is full.
4. A calcium oxide drying column can be included in the apparatus if you want dry ammonia gas

Experiment 10.3: Fountain experiment

(SB p. 160)

Resources

Student's Book, workbook, Round-bottomed flask filled with ammonia, stopper with hole, glass tube with narrow point, glass beaker filled with water and a few drops of phenolphthalein

Guidelines

Guide the students to follow the same procedure as for the reactions of metals in Experiment 8.2.

Students should be able to observe the experiment and record their observations correctly in their workbook, using a table format such as the example on p. 118 of the SB.

Answers

Observations and Conclusions

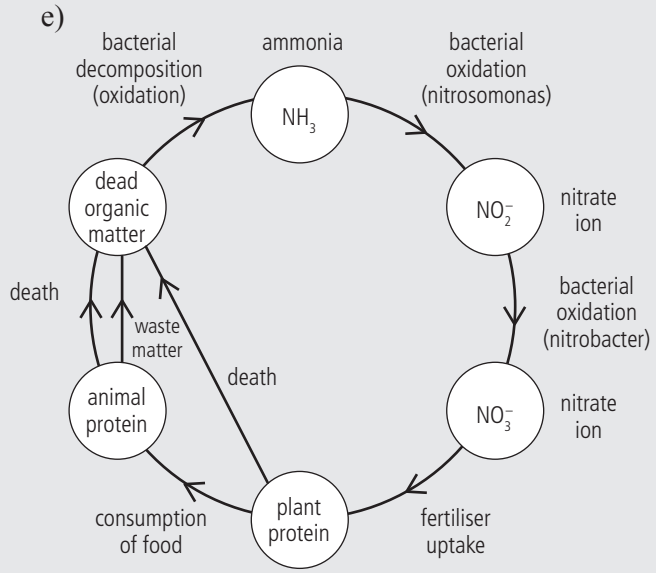
1. Ammonia is extremely soluble in water and one volume water will dissolve 1 300 volumes of ammonia at STP.
2. As the first drop of water reaches the top of the tube, some of the ammonia dissolves in it. This reduces the pressure in the upper flask and water rises up the tube, creating a fountain. Ammonia solution soon fills the round-bottomed flask.
   $NH_3(g) + H_2O(l) \rightarrow NH_4OH(aq)$
3. The presence of an indicator shows that the ammonia solution is alkaline. The colourless water with phenolphthalein turns pink in the fountain.

Revision questions: Answers

(SB p. 164)

1. a) D b) A c) C d) B e) A
2. a) Haber process
   b) fractional distillation
   c) nitrogen cycle
3. a) Group 15
   b) nitrogen (N), phosphorus (P), arsenic (As), antimony (Sb) and bismuth (Bi)
   c) −3
   d) Nitrogen, phosphorus: non-metals
      Arsenic, antimony: metalloids
      Bismuth: metal
4. Oxygen in the air reacts with copper to form copper oxide:
   $2Cu(s) + O_2(g) \rightarrow 2CuO(s)$

   An enclosed volume of air is allowed to flow over hot copper turnings which are then oxidized. When the oxygen has been removed from the air, only nitrogen and trace gases are left.
5. Physical properties: gaseous non-metal, no colour, odour, taste, very unreactive
   Chemical properties: diatomic molecules N_2 with strong covalent triple bonds between atoms, combines with oxygen at high temperatures, oxidation states of +5, +3, and −3
6. Used to make proteins in living organisms
   Liquid nitrogen is used for freezing and preserving organic samples
   Nitrogen gas is used as an inert replacement for air to fill packets of fresh food and light bulbs
7. a) The nitrogen cycle is a series of processes in which nitrogen compounds are interconverted in the environment and in living organisms.
   b) N_2 is too unreactive and must be changed into more reactive forms that plants can use.
   c) ammonia and nitrate ions
   d) Nitrogen cycle involves chemical changes, water cycle involves physical changes of state
   e)



8. a) Mix solid calcium hydroxide and ammonium chloride and heat gently.
   b) $Ca(OH)_2(s) + 2NH_4Cl(s) \rightarrow CaCl_2(s) + 2H_2O(l) + 2NH_3(g)$
   c) Downward displacement of air, NH_3 is less dense than air and very soluble in water.
9. a) $3H_2(g) + N_2(g) \rightleftharpoons 2NH_3(g)$
   b) Fractional distillation is used to produce nitrogen gas from liquid air
   c) The ammonia is separated from the reactants by liquefaction, as ammonia has a much higher boiling point than nitrogen and hydrogen. By releasing some pressure, the expansion causes cooling, and ammonia is condensed and removed.
10. a) Step 1: $4NH_3(g) + 5O_2 \rightarrow 4NO + 6H_2O(g)$
    Step 2: $2NO + O_2(g) \rightarrow 2NO_2$
    Step 3: $4NO_2 + 2H_2O(l) \rightarrow 4HNO_3(aq) + NO$
    b) Platinum/rhodium catalyst
    c) Oxidation states of N in:
    NH_3: −3
    NO: +2
    NO_2: +4
    HNO_3: +5

All the reactions are redox reactions because the oxidation number of N increases every time. N is oxidised in every reaction.

11. a) Haber process
    b) Ammonia
    c) Ostwald process
    d) Platinum/rhodium catalyst
    e) $4NH_3(g) + 5O_2(g) \rightarrow 4NO(g) + 6H_2O(g)$
    f) H_2O
    g) NH_4NO_3
    h) Fertiliser and explosive

How are you doing?

Take this opportunity to ask learners if there is anything that they do not understand. You can check their understanding by asking them some questions about the information covered in the unit. Explain anything that learners do not understand.

Key words

nitrifying) – the biological oxidation of ammonia or ammonium to nitrite, followed by the oxidation of the nitrite to nitrate: 157

denitrifying – to remove nitrogen or nitrogen groups from (a compound); to reduce (nitrates or nitrites) to nitrogen-containing gases, as by bacterial action on soil: 157

carnivorous – subsisting or feeding on animal tissues: 158

Haber Process – an industrial process for producing ammonia from nitrogen and hydrogen, using an iron catalyst at high temperature and pressure: 161

Ostwald Process – a chemical process for making nitric acid (HNO_3); developed by Wilhelm Ostwald and patented in 1902: 162

Checklist for Self-evaluation

Theme 2 Topic 10

EVALUATION GUIDE: Student should be able to:

	Criteria	4	3	2	1	Comments
10.1	explain the laboratory preparation of nitrogen					
10.2	list the properties of nitrogen					
10.3	list the oxides of nitrogen					
10.4	drawn and explain the nitrogen cycle					
10.5	explain the Huber process for the preparation of ammonia					
10.6	state the uses of nitrogen					
10.7	list the uses of ammonia.					

Code for evaluation:			
4 – Very well	3 – Well	2 – Fairly well	1 – Not well at all

Topic 11: Sulphur
Performance objectives
11.1 state the general properties of Group 16 elements **11.2** write and draw the electron configuration of sulphur **11.3** explain the meaning of allotropy **11.4** identify the allotropes of sulphur **11.5** state the uses of sulphur **11.6** name some common compounds of sulphur **11.7** determine the oxidation states of sulphur in its major compounds **11.8** describe the industrial preparation of H_2SO_4 by the contact process **11.9** state the uses of H_2SO_4.

Introduction

Sulphur is one of the few elements that are found in in elemental form in the Earth's crust, not combined with other elements. Sulphur is also found in many sulphide and sulphate minerals. Elemental sulphur is present in large amounts in volcanoes and it is released from volcanoes in the form of a gas. When the sulphur gas reaches the cold air, it changes back to a solid to form beautiful yellow deposits along the edge of the volcano.

Activity 11

Experiment 11.1: Prepare sulphur ointment

(SB p.170)

Resources

Powdered sulphur, glycerin, natural yoghurt, measuring spoon or spatula, glass beaker

Guidelines

Safety: Guide the students to be extremely careful when performing this experiment.

High concentrations of CS_2 are lethal, as it affects the nervous system. The smell of H_2S can only be detected at low concentrations. High concentrations of H_2S numb the senses and cause death quickly by respiratory paralysis. It is more toxic than cyanide.

Demonstrate to the class how to prepare sulphur ointment.

Students should be able to prepare sulphur ointment.

Revision questions: Answers

(SB p. 173)

1. a) C b) D c) A d) D e) B
2. • Group 16 is known as the oxygen group or chalcogens.
   • Members are: Oxygen (O); gaseous non-metal that exists as diatomic molecules, O_2
   • Sulphur (*s*); solid yellow non-metal that exists as S_8 rings
   • Selenium (Se); non-metal that exists as Se_8 molecules
   • Tellurium (Te); metalloid with a three-dimensional structure
   • Polonium (Po); radioactive metal
   • They have six valence electrons
   • The most common oxidation state is −2, but +2, +4 and +6 are possible
3. a) Sulphur: $1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^4$ or [Ne] $3s^2\ 3p^4$
   b) 6 valence electrons
   c) i) SO_2: +4
      ii) HSO_4^-: +6
      iii) Na_2SO_4: +6
      iv) $Na_2S_2O_3$: +2
      v) H_2S: −2
      vi) H_2SO_3: +4
      vii) S: 0
      viii) SO_3: +6
4. a) Sulphur is found in volcanic regions, as contaminant on crude oil and natural gas, and as part of minerals.
   b) Frasch process: superheated water is pumped down the pipes on the outside to melt the sulphur; compressed air is forced down the inside pipe; liquid sulphur/air mixture rises to the surface up the middle pipe.
5. • brittle, yellow, crystalline non-metal that exists as S_8 rings
   • insoluble in water, but soluble in organic solvents
   • does not conduct electricity, whether solid, molten or dissolved
   • relatively low melting and boiling point
6. a) Different forms of the same element in the same physical state
   b) Rhombic (α-form) and monoclinic (β-form)
      Both yellow solids that consists of S_8 rings
      Different crystalline structures: crystals of rhombic sulphur are denser and more stable than those of monoclinic sulphur at room temperature
   c) Plastic sulphur
7. a) $S(s) + O_2(g) \rightarrow SO_2(g)$
   b) Strong, choking smell and is toxic
   c) $2HCl(aq) + Na_2SO_3(aq) \rightarrow 2NaCl(aq) + H_2O(l) + SO_2(g)$ OR
      $Cu(s) + H_2SO_4(aq) \rightarrow CuSO_4(aq) + 2H_2O(l) + SO_2(g)$
   d) sulphur trioxide, SO_3
   e) $SO_3(g) + H_2O(l) \rightarrow H_2SO_4(aq)$

How are you doing?

Take this opportunity to ask learners if there is anything that they do not understand. You can check their understanding by asking them some questions about the information covered in the unit. Explain anything that learners do not understand.

Key words

Frasch Process – a method to extract sulphur from underground deposits; it is the only economic method of recovering sulphur from elemental deposits: 167

alchemical symbol – originally devised as part of alchemy, were used to denote some elements and some compounds until the 18th century: 167

vulcanisation – chemical process for converting natural rubber or related polymers into more durable materials via the addition of sulphur or other equivalent curatives or accelerators; the additives modify the polymer by forming cross-links (bridges) between individual polymer chains: 170

Checklist for Self-evaluation

Theme 2 Topic 11

EVALUATION GUIDE: Student should be able to:

	Criteria	4	3	2	1	Comments
11.1	write and draw the electron configuration of sulphur					
11.2	list and draw the allotropes of sulphur					
11.3	list the oxides of sulphur					
11.4	state the uses of sulphur					
11.5	explain the Contact Process for the manufacture of H_4SO_3					
11.6	list the uses of H_2SO_4					

Code for evaluation:			
4 – Very well	3 – Well	2 – Fairly well	1 – Not well at all

THEME 3

Chemistry and industry

Topic 12: Oxidation-reduction reactions

Performance objectives

12.1 define oxidation as: addition of oxygen, removal of hydrogen, process of electron loss, process of increase of oxidation number of a substance
12.2 define reduction as the reverse of any of the above processes
12.3 calculate oxidation numbers of elements using a set of arbitrary rules, viz:
- **12.3.1** oxidation number of free elements = 0
- **12.3.2** oxidation number of oxygen in any compound is –2, except in peroxides, where it is –1
- **12.3.3** oxidation number of H is +1, except in hydrides, where it is –1
- **12.3.4** oxidation number of a neutral molecule or compound is zero, for example H_2SO_4 = 0, etcetera

12.4 use oxidation numbers to name inorganic compounds, to include the number of oxygen atoms and water molecules (if hydrated)
12.5 determine the oxidation states of a number of common elements in their compounds
12.6 define oxidising and reducing agents in terms of:
- **12.6.1** addition and removal of oxygen and hydrogen respectively
- **12.6.2** loss and gain of electrons
- **12.6.3** change in oxidation numbers/states

12.7 write and balance redox equations.

Introduction

Oxidation-reduction (redox) reactions happen around us all the time. They range from the burning of fossil fuels to the action of household bleach. In industry, we obtain most metallic and non-metallic elements from their ores by the processes of oxidation or reduction. Many important redox reactions happen in water, but not all redox reactions occur in aqueous solution. In previous topics we saw that acid-base reactions are proton (H^+) transfer reactions. Redox reactions are electron transfer (e^-) reactions.

Activity 12.1: Test yourself (SB p. 176)

Resources

Student's Book and workbook, chalk board, flip-charts

Guidelines

This activity can be used as a class test.

Guide students to define oxidation and reduction.

Students should be able to know the various definitions of oxidation and reduction.

Answers

1. a) oxidation
   b) oxidation
   c) reduction
   d) reduction
   e) oxidation
2. a) reduced
   b) oxidized
   c) reduced
   d) reduced
   e) oxidized

Activity 12.2: Calculate oxidation numbers (SB p. 178)

Resources

Student's Book and workbook, chalk board, flip-charts

Guidelines

This activity can be used as a class test.

Guide students to calculate oxidation numbers, using the rules set out under the objectives, and give central atoms in compounds their IUPAC names.

Students should be able to know the set of rules that enables one to calculate oxidation numbers of central atoms in compounds.

Answers

1. The oxidation numbers of sulphur in these compounds:
   a) SO_2 +4
   b) Na_2SO_4 +6
   c) H_2S −2
   d) sulphur powder 0
   e) $H_SO_4^-$ +6
   f) $Na_2S_2O_3$ +2
   g) H_2SO_3 +4
   h) SO_3 +6
2. The oxidation numbers of nitrogen in these compounds:
   a) N_2O +1
   b) NO_2 +4
   c) NO +2
   d) NH_4OH −3
   e) N_2 0
   f) NH_3 −3
   g) $NaNO_3$ +5
   h) N_2O_4 +4
   i) NH_4^+ −3

Activity 12.3: Name the compounds (SB p. 180)

Resources

Student's Book and workbook, chalk board, flip-charts

Guidelines

This activity can be used as a class test.

Guide students to name the inorganic compounds.

Students should be able to write IUPAC names of compounds by first calculating the oxidation numbers of the central atom.

Answers

1. Write names for the following compounds.
   a) copper(II) tetraoxosulphate(VI) pentahydrate
   b) Iron(III) tetraoxosulphate(VI)
   c) Potassium manganate(VII)
   d) Lead(IV) oxide
   e) silver trioxonitrate(V)
2. Write formulae for the following compounds.
   a) $FeCl_2$
   b) PH_3
   c) $Pb(SO_3)_2$
   d) CO_2O_3
   e) $Al_2(CO_3)_3$

Activity 12.4: Identify the oxidising and reducing agent INDIVIDUAL (SB p. 182)

Resources

Student's Book and workbook, chalk board, flip-charts

Guidelines

This activity can be used as a class test.

Guide students to name the inorganic compounds.

Students should be able to write IUPAC names of compounds by first calculating the oxidation numbers of the central atom.

Answers

	Oxidising agent	Reducing agent
a)	Cu^{2+}	Zn
b)	Hg^{2+}	O_2^-
c)	Cr^{6+}	N in NH_4^+
d)	H atoms in H_2O	Ca
e)	Cl_2	Cl_2
f)	O_2	S^{2-}(ZnS)

Experiment 12.1: Investigating oxidising and reducing agents

(SB p. 183)

Resources

Method A:
Student's Book, workbook, Dilute hydrochloric acid (HCl(*aq*)), concentrated HCl, iron sulphide (FeS), dilute solutions of $FeCl_3$, $KMnO_4$ and $K_2Cr_2O_7$, test tube with a side arm, rubber tubing, glass tube, stopper, propette, three test tubes in test tube stand

Method B:
Potassium permanganate solution ($KMnO_4$), concentrated HCl, dilute HCl, sodium sulphite (Na_2SO_3), test tube with a side arm, rubber tubing, glass tube, stopper, propettes, test tubes in test tube stand

Guidelines

Guide the students to observe the experiment and record their findings in their workbook.
Safety notes

- Hydrogen sulphide is a poisonous gas that smells like rotten eggs. It irritates the eyes, nose and throat and higher concentrations will cause headaches, nausea and dizziness. Do the experiments in a well-ventilated room or in a fume cupboard. Prepare only enough H_2S to complete the experiment.
- $KMnO_4$ is a strong oxidiser and must be stored away from substances that can be oxidised. It stains skin, clothing and glassware. Remove stains with lemon juice, oxalic acid or a weak solution of H_2O_2.

Students should be able to identify oxidation and reduction half reactions in redox reactions.

Answers

Observations and Conclusions

A. Reducing action of hydrogen sulphide (H_2S)
Preparation of H_2S:
$2HCl(aq) + FeS(s) \rightarrow H_2S(g) + FeCl_2(aq)$
The reaction with iron(III) chloride:
$2FeCl_3(aq) + H_2S(g) + HCl(aq) \rightarrow 2FeCl_2(aq) + S(s) + 3HCl(aq)$
The greenish-brown Fe^{3+} ion is reduced to the green Fe^{2+} ion. The solution turns milky as the solid sulphur is formed. The reaction with potassium manganate(VII):
$2KMnO_4(aq) + 5H_2S(g) + 6HCl(aq) \rightarrow 2KCl(aq) + 2MnCl_2(aq) + 5S(s) + 8H_2O(l)$
The purple MnO_4^- ion is reduced to the colourless MN^{2+} ion. The solid sulphur gives the solution a milky effect:
$MnO_4^- + 8H^+ + 5e^- \rightarrow MN^{2+} + 4H_2O$
The reaction with potassium dichromate:
$K_2Cr_2O_7(aq) + 3H_2S(g) + 8HCl(aq) \rightarrow 2KCl(aq) + 2CrCl_3(aq) + 2S(s) + 7H_2O(l)$
The orange $Cr_2O_7^{2-}$ ion is reduced to the green Cr^{3+} ion.
$Cr_2O_7^{2-} + 14H^+ + 6e^- \rightarrow 2Cr^{3+} + 7H_2O$
The solution turns milky green.

B. Oxidising action of potassium manganate(VII) ($KMnO_4$)
The reaction between H_2S and $KMnO_4$ in the previous section illustrates the oxidising ability of $KMnO_4$. Here we will look at the reaction between $KMnO_4$ and SO_2.
Preparation of SO_2:
$Na_2SO_3(s) + 2HCl(aq) \rightarrow 2NaCl(aq) + H_2O(l) + SO_2(g)$
SO_2 can also be prepared using the reaction between copper metal and concentrated sulphuric acid. Place a small amount of copper filings in a test tube. Cover with concentrated H_2SO_4 and heat very carefully.
Move the test tube constantly through the flame to prevent the acid from boiling and splashing out of the test tube.
The reaction is:
$Cu(s) + 2H_2SO_4(l) \rightarrow CuSO_4(aq) + 2H_2O(l) + SO_2(g)$
You can either use a side-arm test tube and bubble the gas through the $KMnO_4$ solution, or dip a filter paper in $KMnO_4$ solution and hold it at the mouth of the test tube. The $KMnO_4$ will discolour.

The reaction between $KMnO_4$ and SO_2 is:
$2KMnO_4(aq) + 5SO_2(g) + 2H_2O(l) \rightarrow$
$K_2SO_4(aq) + 2MnSO_4(aq)$ 1+ $2H_2SO_4(aq)$
Note that no free sulphur forms and the resultant solution is colourless and clear.

Activity 12.5: Balancing reactions

(SB p. 186)

Resources

Student's Book and workbook, chalk board, flip-charts

Guidelines

This activity can be used as a class test.

Guide students to balance redox equations.

Students should be able to write and balance redox equations.

Answers

1. Oxidation reaction: $Mg \rightarrow Mg^{2+} + 2e^-$
   RA: Mg
   Reduction reaction: $2H^+ + 2e^- \rightarrow H_2$
   OA: H^+
   $Mg + 2H^+ + 2e^- \rightarrow Mg^{2+} + H_2 + 2e^-$
   add spectator ion Cl− and cancel e^-
   $Mg(s) + 2HCl(aq) \rightarrow MgCl_2(aq) + H_2(g)$
2. Oxidation reaction: $Fe \rightarrow Fe^{3+} + 3e^-$
   $\times 4$ RA: Fe
   Reduction reaction: $O_2 + 4e^- \rightarrow 2\ O_2^- \times$
   3 OA: O_2
   $4Fe + 3O_2 + 12e^- \rightarrow 4Fe^{3+} + 6O_2^- + 12e^-$
   cancel e−
   $4Fe(s) + 3O_2(g) \rightarrow 2Fe_2O_3(s)$
3. Oxidation reaction: $Na \rightarrow Na^+ + e^- \times 2$
   RA: Na
   Reduction reaction: $2H^+ + 2e^- \rightarrow H_2$
   OA: H^+
   $2Na + 2H^+ + 2e- \rightarrow 2Na^+ + H_2 + 2e^-$
   add spectator ion O_2^- and cancel e^-
   $2Na(s) + H_2O(l) \rightarrow NaOH(aq) + H_2(g)$
4. Oxidation reaction: $C^{2+} \rightarrow C^{4+} + 2e^-$
   $\times 3$
   Reduction reaction: $Fe^{3+} + 3e^- \rightarrow Fe$
   $\times 2$
   $2Fe^{3+} + 3C^{2+} + 6e^- \rightarrow 2Fe + 3C^{4+} + 6e^-$
   add spectator ion O_2^- and cancel e^-
   $Fe_2O_3(s) + 3CO(g) \rightarrow 2Fe(s) + 3CO_2(g)$

Revision questions: Answers

(SB p. 191)

1. a) B b) D
2. C
3. D
4. A
5. a) redox/oxidation-reduction reaction
   b) reduction
   c) oxidation
   d) reducing agent
   e) oxidation state
6. A reducing agent is oxidized and loses electrons during a reaction. An oxidizing agent is reduced and gains electrons during a reaction. Oxidation occurs when the atom loses electrons and when the oxidation number increases.
7. The oxidation number is an indication of an atom's relative richness in electrons and it signifies that number of charges the atom would have in a molecule if electrons were transferred completely.
8. a) iron in FeO: +2, $FeCl_3$: +3, Fe: 0, $FeBr_2$: +2
   b) copper in Cu_2O: +1, $CuCl_2$: +2: Cu: 0, CuO: +2, CuCl: +1
   c) carbon in C: 0, CO: +2, CO_2: +4, HCO_3^-: +4
   d) chlorine in Cl_2: 0, ClF: +1, Cl_2O: +1, ClO_3^-: +5, HCl: −1, HOCl: +1
9. a) 0 +2
   $Cu \rightarrow CuSO_4$ oxidation (increase in oxidation number of Cu)
   $H_2SO_4 \rightarrow SO_2$ reduction (decrease in oxidation number of S)
   H_2SO_4 is the oxidising agent
   b) $Cu \rightarrow Cu^{2+} + 2e-$
10. a) i) OA: Cl_2, RA: Br−
    ii) OA: Cu^{2+}, RA: Zn
    iii) OA: H^+: RA Mg
    b) i) K^+
    ii) SO_4^{2-}
    ii) Cl^-
    c) i) Colourless solution turn yellow-brown as a result of the Br_2 that forms
    ii) The blue colour of the solution disappears, the zinc powder dissolves, reddish brown copper precipitates.
    iii) Bubbles of H_2 gas form, the Mg dissolves

How are you doing?

Take this opportunity to ask learners if there is anything that they do not understand. You can check their understanding by asking them some questions about the information covered in the unit. Explain anything that learners do not understand.

Key words

oxidation – the reaction that takes place when an element or compound loses electrons; a reaction in which oxygen is added to the element or compound; when the oxidation number of an atom increases: 174

reduction – the reaction that takes place when a compound gains electrons; a reaction in which oxygen is removed from the element or compound: 174

water of crystallisation – water that is found in the crystalline framework of a metal complex or a salt, which is not directly bonded to the metal cation; or the water molecules forming an essential part of the crystal structure of some compounds: 179

redox reaction – an electron transfer reaction: 185

redox pair – the reducing agent and oxidising agent:185

spectator ions – ions that do not take part in the overall reaction: 176, 185

activity series – a list of metals ranked in order of decreasing reactivity to displace hydrogen gas from water and acid solutions; can also be used to predict which metals will displace other metals in aqueous solutions: 187

displacement – a reaction in which an ion or atom in a compound is replaced by an ion or atom of another element: 186

reducing agents – substances that can donate electrons to other substances: 182

Checklist for Self-evaluation

Theme 3 Topic 12

EVALUATION GUIDE: Student should be able to:

	Criteria	4	3	2	1	Comments
12.1	define oxidation and reduction					
12.2	explain redox reactions					
12.3	calculate oxidation numbers of central atoms in inorganic compounds and hence give their IUPAC names					
12.4	identify oxidation and reduction half reactions in redox reactions					
12.5	write formula of a compound given its IUPAC name					
12.6	write and balance redox equations.					

Code for evaluation:			
4 – Very well	3 – Well	2 – Fairly well	1 – Not well at all

Topic 13: The ionic theory

Performance objectives

13.1 explain the difference between:
- **13.1.1** electrovalent and covalent compounds
- **13.1.2** electrolytes and non-electrolytes

13.2 investigate the movement of ions in solution
13.3 distinguish between strong and weak electrolytes
13.4 rank and explain the position of ions in the electrochemical series
13.5 relate the order of ions in the electrochemical series to their rate of discharge from solution
13.6 state the factors affecting the preferential discharge of ions.

Introduction

The ionic theory was originally developed by a Swedish scientist, Svante Arrhenius, in 1887. According to this theory, soluble ionic salts dissolve in water to form an electrolyte that consists of positive and negative ions. This process is called ionisation. The total positive charge is equal to the total negative charge and the solution is electrically neutral. Metallic ions, ammonium ions and hydrogen ions are positive and non-metals and certain polyatomic ions form negative ions. The charge on an ion is equal to its valency. The ionic theory is used to explain the process of electrolysis. In an electrolyte, the ions are free to move in all directions. When an electric current is passed through the electrolyte, the negative ions are attracted to the positive electrode and they move in its direction. The positive ions are attracted to the negative electrode and move towards it. When ions reach the electrodes, they are discharged and they form atoms again.

Activity 13.1: Distinguish covalent from electrovalent compounds

(SB p. 194)

Resources

Samples of electrovalent and covalent compounds, Student's Book and workbook

Guidelines

This activity can be used as a class test.

Guide students to distinguish between electrovalent and covalent compounds.

Students should be able to list five electrovalent and covalent compounds respectively.

Answers

1. Covalent compounds: H_2S, NH_3, HOCl, N_2O_5
   Electrovalent compounds: NaBr, $FeCl_3$, PbO_2, $BeCl_2$
2. Covalent: H_2O, NO_2, P_2O_5, CO, CO_2, SO_3, CH_4 and all other organic molecules, H_2 and all other diatomic molecules, etc.
   Electrovalent compounds: All halide salts, e.g. NaCl, KBr, AgCl, MgI_2, all metal oxides, e.g. CuO, Na_2O, Fe_2O_3, all sulphates, e.g. $CuSO_4$, $FeSO_4$, all carbonates, e.g. $CaCO_3$, $MgCO_3$, all nitrates, e.g. KNO_3, $Mg(NO_3)_2$, etc.

Experiment 13.1: Investigating the conductivity of various liquids

(SB p. 196)

Resources

Student's Book, workbook, battery, connecting copper wires, two graphite rods, light bulb, ammeter, glass beaker, liquids to test: ethanol, kerosene (or benzene or chloroform), cooking oil, distilled water, ethanoic acid (vinegar), solutions to test: sodium chloride, sugar, sodium carbonate, sodium hydroxide and copper(II) chloride in water, dilute hydrochloric acid, oxalic acid, sulphuric acid and ammonia, naphthalene dissolved in benzene

Guidelines

Guide the students to follow the procedure on p. 196 of the SB, observe the experiment and record their observations correctly in their workbook.

Students should be able to compile a list of electrolytes and non-electrolytes, strong and weak electrolytes.

Answers

Observations and Conclusions
The beaker and graphite rods must be rinsed and dried after each reading, because particles from the liquid can cling to the beaker and rods and contaminate the next reading.

Non-electrolytes are: distilled water, ethanol, kerosene/benzene/chloroform, cooking oil, sugar solution, naphthalene in benzene.

Weak electrolytes are: ethanoic acid, ammonia solution, oxalic acid solution, sodium carbonate solution

Strong electrolytes are: sodium chloride solution, copper(II) chloride solution, dilute hydrochloric acid, dilute sulphuric acid, sodium hydroxide solution.

Activity 13.2: Classify solutions

(SB p. 183)

Resources

Student's Book, workbook

Guidelines

This activity can be used as a class test.

Guide students to distinguish between electrolytes and non-electrolytes.

Students should be able to classify the solutions into 3 types of electrolytes: strong, weak or non-electrolytes.

Answers

1. strong electrolytes: HNO_3, KOH, $CuSO_4$
   weak electrolytes: H_2CO_3, C_5H_5N, NH_3,
   non-electrolytes: H_2O, $C_6H_{12}O_6$, CH_3OH

Activity 13.3: Using standard electrode potential

(SB p. 202)

Resources

Student's Book, workbook

Guidelines

Guide the students in a discussion of the relationship between the nature of ions and their ranking in the electrochemical series.

Students should be able to give the order of discharge of a given list of ions.

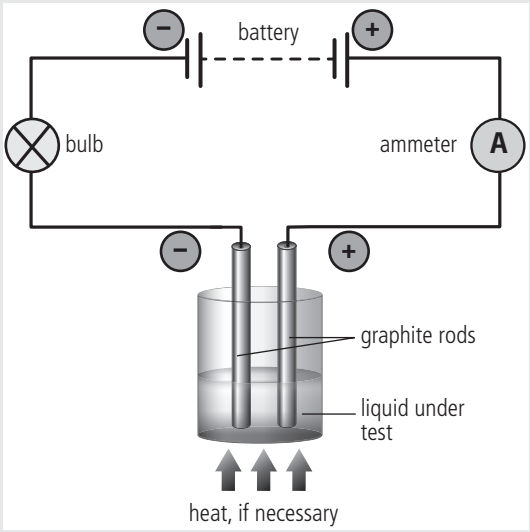
Answers

1. Cathode: Cu^{2+} and Na^+ migrate there. Cu^{2+} will discharge in preference to Na^+.
   Anode: Cl^- and I^- migrate there. I^- will discharge in preference to Cl^-.
   $Cu^{2+}(aq) + 2Cl^-(aq) + 2Na^+(aq) + 2I^-(aq) \rightarrow Cu(s) + I_2(s) + 2Na^+(aq) + 2Cl^-(aq)$
2. a) cathode: $Cu^{2+} + 2e- \rightarrow Cu$
   Anode: $4OH^- \rightarrow O_2 + 2H_2O + 4e^-$
   The removal of copper ions from the solution causes it to become colourless.
   b) cathode: $2H^+ + 2e^- \rightarrow H2$
3. Order of discharge: silver, nickel, magnesium, calcium
4. a) volts (V)
   b) Hydrogen electrode
   c) A platinum electrode is placed in an acidic solution that contains H^+ ions and hydrogen gas is bubbled through the solution.
   d) pressure of H_2 gas: 1 atm, temperature: 25 °C, concentration of H^+ ions: 1 mol·dm^{-3}
   $2H^+ + 2e^- \rightarrow H_2$
   $E° = 0{,}00$ V
5. The table contains all half-reactions written as reductions in order of increasing $E°$ values. Chemists use these tables to determine which substances will react spontaneously, and also to calculate the voltage supplied by a particular combination of half-cells.
6. a) No
   b) Yes
   c) Yes
   d) No

Revision questions: Answers

(SB p. xxx)

1. a) C, D b) C c) A, B d) D e) C
2. a) Electrolytes
   b) Anions
   c) Molecules
   d) Table of standard reduction potentials
   e) Standard hydrogen electrode
3. a) Strong electrolytes form when compounds ionize completely in water. Weak electrolytes form when compounds ionize partially in water Non-electrolytes are molecular substances that do not dissociate in water.
   b) Strong electrolytes: fused lead(II) bromide, copper(II) sulphate solution, molten table salt
   c) Non-electrolytes: Kerosene, sugar solution, cooking oil
   d) Weak electrolytes: Carbonic acid, ammonia solution
4. Tap water contains dissolved salts that are electrovalent compounds, such as sodium chloride. The ions can act as the charge carriers and conduct an electric current. Distilled water only contains water molecules that are non-conductors.
5. a) No, the ions are packed in a crystal lattice and cannot move to conduct a current.
   b) He should either dissolve the salt in water, or melt the salt to free the ions.
   c)



6. a) Yes, lithium is a strong reducing agent and can reduce water
   b) Yes, copper is a strong enough reducing agent to reduce silver ions
   c) No, mercury is not a strong enough reducing agent to reduce hydrogen ions
   d) No, iron is not a strong enough reducing agent to reduce zinc ions

How are you doing?

Take this opportunity to ask learners if there is anything that they do not understand. You can check their understanding by asking them some questions about the information covered in the unit. Explain anything that learners do not understand.

Key words

electrovalent compounds – the chemical bond between an anion and a cation; compounds formed by electrovalent bonds ionise when dissolved in water, with dissociation of the cations and anions; a type of chemical bond in which one atom loses an electron to form a positive ion and the other atom gains the electron to form a negative ion; the resulting ions are held together by electrostatic attraction; also called ionic bond: 193

covalent compounds – a molecule formed by covalent bonds, in which the atoms share one or more pairs of valence electrons: 193

interatomic – forces that hold atoms together in a molecule: 194

intermolecular – intermolecular forces attractive forces between molecules: 194

electrolyte – a liquid that contains ionic salts and can carry a current: 193

non-electrolyte – a substance that does not exist in an ionic form in aqueous solution: 195

fused – the liquid or melted state induced by heat: 193

Checklist for Self-evaluation

Theme 3 Topic 13

EVALUATION GUIDE: Student should be able to:

	Criteria	4	3	2	1	Comments
13.1	list five electrovalent and covalent compounds respectively					
13.2	list three each of weak and strong electrolytes					
13.3	give the order of discharge of a given list of ions					
13.4	state the factors affecting the discharge of ions in the solution.					

Code for evaluation:			
4 – Very well	3 – Well	2 – Fairly well	1 – Not well at all

Topic 14: Electrolysis

Performance objectives

14.1 explain the quantitative aspects of electrolysis
14.2 define electrolytes (strong, weak, fused/molten, non-electrolytes), electrolytic and electrochemical cells
14.3 differentiate between strong and weak electrolytes
14.1 illustrate the electrolysis of acidified water, copper(II) sulphates and brines
14.2 identify factors affecting the discharge of ions during electrolysis
14.3 construct the electrolytic and electrochemical cells
14.4 state Faraday's laws of electrolysis
14.5 calculate the amount of substances liberated or deposited at electrodes during electrolysis
14.6 explain the uses of electrolysis in extraction and purification of metals.

Introduction

Electrochemistry is the branch of chemistry that deals with the transformation between electrical energy and chemical energy. Some chemical reactions occur spontaneously and chemical energy is converted into electrical energy. We use these reactions to make portable electrical energy in the form of cells and batteries.

Other reactions require an input of electrical energy before the compounds will react. These reactions are called electrolytic reactions.

Electrochemical reactions have important practical applications in our everyday lives. In Topic 14 we are going to revise and discuss terminology used in electrochemical reactions.

Activity 14.1: Revise the basics of electrolysis

(SB p. 206)

Resources

Student's Book, workbook

Guidelines

This activity can be used as a class test.

Students should be able to define electrolysis

Answers

1.

Electrolytic cells	Voltaic cells
Electrical energy → chemical energy	Chemical energy → electrical energy
Reactions not spontaneous	Reactions spontaneous
Cathode negative, anode positive	Cathode positive, anode negative

2. Electrode: the solid substance that transfers electrons to and from the ions, atoms and molecules of the electrolyte
Electrolyte: a dissolved or molten ionic compound that can conduct electricity
Anode: the electrode where oxidation takes place, to which the anions go
Cathode: the electrode where reduction takes place, to which the cations go

Experiment 14.1: Investigating the electrolysis of copper(II) chloride

(SB p. 209)

Resources

Student's Book, workbook, battery, connecting copper wires, fitted with clips, two graphite rods, light bulb, glass beaker, copper(II) chloride, water, test tube

Guidelines

Guide the students to construct electrolytic and electrochemical cells by assembling a series circuit such as the one shown in Figure 14.3 in the SB.

Students should be able to set up the electrolytic and electrochemical cells, and observe and record their findings in their workbook.

Answers

Observations and Conclusions

- A reddish-brown deposit forms on the negative electrode (cathode). The positive copper ions move to the negative cathode and form a copper metal deposit on the electrode.
- A greenish gas forms bubbles on the positive electrode (anode). The smell of chlorine is distinctive.
- The electrolyte becomes lighter in colour. The copper ions have a blue colour. When the ions are removed from the solution to form copper metal, the colour will disappear in time.

Experiment 14.2: Investigating the electrolysis of acidified water

(SB p. 210)

Resources

Student's Book, workbook, Hofmann's voltammeter, battery, connecting copper wires, fitted with clips, glass beaker, dilute sulphuric acid, water, small funnel

Guidelines

Demonstrate the electrolysis of acidified water using Hoffman's voltammeter

Students should be able to observe the experiment and record their findings in their workbook.

Students should be able to describe the principle in the Hoffman's voltammeter in the electrolysis of acidified water.

Answers

Observations and Conclusions

In this experiment gas collects at the top of the arms of the voltammeter. The volume of the hydrogen gas is twice the amount of the oxygen gas.

Activity 14.2: Investigating electrolysis

(SB p. 210)

Resources

Student's Book, workbook, an electrochemical cell apparatus, copper sulphate solution

Guidelines

This activity can be used as a class test.

Demonstrate the electrolysis of copper sulphate.

Students should be able to explain the electrolysis of acidified water, copper II sulphates and brine.

Answers

1. a) Chlorine gas
   b) Oxidation: $2Cl^- \rightarrow Cl_2 + 2e^-$
   c) Sodium metal
   d) Reduction: $Na^+ + e^- \rightarrow Na$
   e) $2NaCl(l) \rightarrow 2Na(l) + Cl_2(g)$
2. a) Cu
   b) cathode
   c) $Cu^{2+} + 2e^- \rightarrow Cu$
   d) reduction
   e) O_2
   f) anode
   g) $2H_2O \rightarrow O_2 + 4H^+ + 4e^-$
   h) oxidation
   i) $2Cu^{2+}(aq) + 2H_2O(l) \rightarrow 2Cu(s) + O_2(g) + 4H^+(aq)$
3. a) Anode: $2H_2O \rightarrow O_2 + 4H^+ + 4e^-$
   b) Cathode: $2H_2O + 2e^- \rightarrow H_2 + 2OH^-$
   c) Net reaction: $6H_2O \rightarrow O_2 + 4H^+ + 4OH^- + 2H_2$
      Simplify: $2H_2O(l) \rightarrow O_2(g) + 2H_2(g)$
   d) The gases form in a ratio of 1 O_2 : 2 H_2.
   e) The electrode on the left is the anode, because less gas developed there.
   f) Hydrogen gas explodes with a popping sound when a burning splint is held at

the test tube mouth. Oxygen gas will set a glowing split alight.

g) In the left electrode (anode), H+ ions form. The solution will be acidic and the indicator will turn yellow. In the right electrode (cathode), OH– ions form. The solution will be basic and the indicator will turn blue.

Experiment 14.3: Investigating galvanic cells (SB p. 214)

Resources

Student's Book, workbook, Copper (Cu) and zinc (Zn) electrodes, 1 mol·dm^{-3} solutions of copper(II) sulphate ($CuSO_4$) and zinc sulphate ($ZnSO_4$), two 100 ml glass beakers, salt bridge consisting of glass U-tube, saturated potassium nitrate (KNO_3) and glass wool stoppers, ammeter, connecting wires

Guidelines

Demonstrate the electrolysis of brine

Students should be able carry out the electrolysis of brine, and observe and record their findings in their workbook.

Answers

Observations and Conclusions

- The reading on the ammeter is 1.10 V.
- This is the standard cell voltage for the Daniell cell at standard conditions of 25 °C and concentrations of 1 mol·dm^{-3}.
- After time, the mass of the zinc electrode will decrease as the zinc electrode dissolves in the solution due to the oxidation of zinc at the anode: $Zn \rightarrow ZN_2^+ + 2e^-$
- The Cu electrode will increase in mass, as Cu^{2+} ions are deposited on the electrode during the reduction of Cu^{2+}: $Cu^{2+} + 2e^- \rightarrow Cu$
- The ionic cell reaction is: $Zn(s) + Cu^{2+}(aq) \rightarrow ZN_2^+(aq) + Cu(s)$

Activity 14.3: Revise oxidation and reduction (SB p. 215)

Resources

Student's Book, workbook

Guidelines

This activity can be used as a class test.

Guide the students in a discussion on the properties of electrolysis.

Students should be able to answer questions on oxidation and reduction.

Answers

1. a) $Zn(s) + 2Ag^+(aq) \rightarrow ZN_2^+(aq) + 2Ag(s)$
   $E°_{cell} = E°_{OA} - E°_{RA} = 0{,}80 - (-0{,}76) = 1{,}56$ V
   b) $Cu(s) + 2Fe^{3+}(aq) \rightarrow Cu^{2+}(aq) + 2Fe^{2+}(aq)$
   $E°_{cell} = E°_{OA} - E°_{RA} = 0{,}77 - 0{,}34 = 0{,}43$ V
2. a) The platinum electrode is the cathode.
   b) $Fe \rightarrow Fe^{3+} + 3e^-$
   c) $O_2 + 2H^+ + 2e^- \rightarrow H_2O_2$
   d) $2Fe + 3O_2 + 6H+ \rightarrow 2Fe^{3+} + 3H_2O_2$
   e) K^+
   f) Platinum is an inert metal and does not react with the chemicals in the half-cell or their products.
3. a) $MnO_4^- + 8H^+ + 5e^- \rightarrow MN_2^+ + 4H_2O$
   b) Potassium permanganate is a strong oxidising agent and can easily oxidise oxalic acid to CO_2 and H^+ ions. Oxalic acid is a strong enough reducing agent to oxidise the permanganate ion to MN_2^+ and H_2O. The cell potential for the reaction is positive:
   $E°_{cell} = E°_{OA} - E°_{RA} = 1{,}51 - (-0{,}49) = 2{,}00$ V
   c) Oxidation: $H_2C_2O_4 \rightarrow 2CO_2 + 2H^+ + 2e^-$
   Reduction: $MnO_4^- + 8H^+ + 5e^- \rightarrow MN_2^+ + 4H_2O$
   Net ionic equation: $2MnO_4^- + 5H_2C_2O_4 + 6H^+ \rightarrow 2MN_2^+ + 10CO_2 + 8H_2O$

(Net equation: $2KMnO_4 + 5\ H_2C_2O_4 + 6HCl \rightarrow 2MnCl_2 + 10CO_2 + 8H_2O + 2KCl$)

4. a) The Cu will form the cathode and the Zn the anode.
   b) The cell sap contains salicylic acid and salts that will act as the electrolyte.
   c) The potatoes have less cell sap than the lemons.

Activity 14.4: Perform calculations related to electrolysis (SB p. 218)

Resources

Student's Book, workbook

Guidelines

This activity can be used as a class test.

Explain the Faraday's first and second laws of electrolysis.

Students should be able to state Faraday's first and second laws of electrolysis, and solve problems involving these laws.

Answers

1. $Ni^{2+} + 2e^- \rightarrow Ni$

$$m = \frac{Q}{F} \times \frac{M}{z}$$

$$0.5\text{ g} = \frac{Q}{96\ 500} \times \frac{58.7}{2}$$

$$Q = \frac{0.5}{29.35} \times 96\ 500$$

$$= 1\ 643.95\text{ C}$$

$$Q = It \Rightarrow t = \frac{Q}{I} = \frac{1\ 643.95}{3} = 548\text{ s}$$
$$= 9\text{ min } 8\text{ s}$$

2. $Cr^{3+} + 3e^- \rightarrow Cr$

$$m = \frac{Q}{F} \times \frac{M}{z}$$

$$0.5\text{ g} = \frac{Q}{96\ 500} \times \frac{52}{3}$$

$$Q = \frac{0.5}{17.33} \times 96\ 500$$

$$= 2\ 784.19\text{ C}$$

$$1\text{ h} = 60 \times 60\text{ s} = 3\ 600\text{ s}$$

$$Q = It \Rightarrow I = \frac{Q}{t} = \frac{2\ 784.19}{3\ 600} = 0.77\text{ A}$$

3. a) anode (oxidation): $4OH^- \rightarrow O_2 + 2H_2O + 4e^-$
   cathode (reduction): $2H^+ + 2e^- \rightarrow H_2$
   b) $Q = It = 1.60 \times 10 \times 60 = 960$ C
   anode: 4 mol e^- produce 1 mol O_2
   Charge on 1 mol $e^- = F = 96\ 500$ C
   $\Rightarrow 960\text{ C} = 9.95 \times 10^{-3}$ mol e^-

$$\text{mol } O_2 \text{ produced} = \frac{9.95 \times 10^{-3}}{4}$$
$$= 2.49 \times 10^{-3} \text{ mol } O_2$$

   cathode: 2 mol e^- produced 1 mol H_2

$$\text{mol } H_2 \text{ produced} = \frac{9.95 \times 10^{-3}}{2}$$
$$= 5.00 \times 10^{-3} \text{ mol } H_2$$

   c) 2.49×10^{-3} mol $\times$ 22,4
   $= 5.58 \times 10^{-2}$ dm^3 O_2
   5.00×10^{-3} mol $\times$ 22,4
   $= 1{,}12 \times 10^{-1}$ dm^3 H_2

Revision questions: Answers

(SB p. 221)

1. a) D b) A c) B d) C e) D
   f) B g) A
2. a) Electrolyte
   b) Galvanic cell
   c) Electrolytic cell
   d) Reduction
   e) Oxidation-reduction/ redox reactions
   f) Reducing agent
   g) Oxidation
   h) Salt-bridge
3.

Column A	Column B
a) Electrode where oxidation occurs	E anode
b) Chemical species that undergoes oxidation	J Reducing agent
c) The loss of electrons	G oxidation
d) Cell in which chemical energy is converted to electrical energy	K Galvanic cell
e) The volt reading when a battery is flat	A zero
f) The electrode to which the cations migrate in an electrochemical cell	F cathode
g) The volt reading for a spontaneous galvanic cell	C positive
h) The substance that causes another substance to be oxidised	I Oxidising agent
i) The type of half-reactions on the Table of Standard Potentials	H reduction
j) The type of cell that is used for the decomposition of chemical compounds	L Electrolytic cell
k) The cathode in an electrolytic cell	D negative
l) The cell potential when the galvanic cell is first assembled	B Maximum

4. a) The process in which electricity is used to bring about a chemical change and the electrical energy is converted to chemical energy.
   b) C, The chloride ions are negative and move towards the positive electrode.
   c) Fe
   d) $2Cl^- \rightarrow Cl_2 + 2e^-$
   e) The reduction reaction for Ca^{2+}/Ca has a lower $E°$ value than the reduction reaction for Na^+/Na. Na^+ is the stronger oxidising agent and will react in preference to Ca^{2+}.
   f) Na metal is less dense than the liquid NaCl and will float.
   g) The Na metal and Cl_2 gas will react violently if they come into contact.
   h) $Na^+ + e^- \rightarrow Na$

   $Q = It = 5 \times 2 \times 60 \times 60 = 3.6 \times 10^4$ C

   $m = \frac{Q}{F} \times \frac{M}{z}$

   $= \frac{3.6 \times 10^4}{96\ 500} \times \frac{23.0}{1}$

   $= \frac{0.5}{29.35} \times 96\ 500$

   $= 8.58$ g sodium
5. a) A liquid or solution that conducts electricity.
   b) Brine contains positive Na^+ and negative Cl− ions that are free to move.
   c) Chloride ions, chloride ions are oxidised to form Cl_2.
   d) $2H_2O + 2e^- \rightarrow H_2 + 2OH^-$
   e) Anode, chloride ions are oxidised to Cl_2 – oxidation at the anode
   f) Chlorine is poisonous and harmful to humans and the environment. Leaching of sodium hydroxide into groundwater can carry health risks for humans and the environment. Hydrogen gas can cause an explosion.
6. a) $H_2O_2 \rightarrow O_2 + 2H^+ + 2e^-$
   b) $H_2O_2 + 2H^+ + 2e^- \rightarrow 2H_2O$
   c) $2H_2O_2(aq) \rightarrow O_2(g) + 2H_2O(l)$
   d) MnO_2 is the catalyst.

e) The oxidising ability of hydrogen peroxide is responsible for its bleaching action.

7. Zinc is a stronger reducing agent than iron because it has a more negative $E°$ value. Zn will be oxidised in preference to Fe and will therefore protect the Fe from being oxidised.

8. a) Standard conditions: $O_2(g)$ pressure of 1 atm, $Pb^{2+}(aq)$ concentration of 1 mol·dm^{-3}, temperature of 25 °C (298 K)
   b) Anode: Pb^{2+}/Pb
   c) Oxidation: $Pb \rightarrow Pb^{2+} + 2e^-$
   d) Reduction: $O_2 + 4H^+ + 4e^- \rightarrow 2H_2O$
   e) Net reaction: $2Pb(s) + O_2(g) + 4H+(aq) \rightarrow 2Pb^{2+}(aq) + H_2O(l)$
   f) $E_{cell} = E°_{OA} - E°_{RA}$
   $= 1{,}23 - (-0{,}13)$
   $= 1{,}36$ V

9. a) The cadmium electrode is the anode.
   b) $Cd + 2OH^- \rightarrow Cd(OH)_2 + 2e^-$
   c) $Cd + 2NiO(OH) + 2H_2O \rightarrow 2Ni(OH)_2 + Cd(OH)_2$
   d) $E°_{cell} = E°_{cathode} - E°_{anode}$
   $= 0{,}48 - (-0{,}82) = 1{,}30$ V
   e) During recharge the electrons will flow from the cathode to the anode: NiO(OH) to Cd.

10. $Au^{3+} + 3e^- \rightarrow Au$

$$m = \frac{Q}{F} \times \frac{M}{z}$$

$$5.00 \text{ g} = \frac{Q}{96\,500} \times \frac{197.0}{3}$$

$$Q = \frac{5.00 \times 3}{197.0} \times 96\,500$$

$$= 7.35 \times 103 \text{ C}$$

$$20 \text{ min} = 20 \times 60 = 1\,200 \text{ s}$$

$$Q = It \Rightarrow I = \frac{Q}{t} = \frac{7.35 \times 10^3}{1\,200} = 6.12 \text{ A}$$

How are you doing?

Take this opportunity to ask learners if there is anything that they do not understand. You can check their understanding by asking them some questions about the information covered in the unit. Explain anything that learners do not understand.

Key words

electrolysis – when a liquid completes an electric circuit: 206

electrochemistry – studies chemical reactions which take place at the interface of an electrode, usually a solid metal or a semiconductor, and an ionic conductor, the electrolyte: 205

anode – the positive electrode: 205

cathode – the negative electrode: 205

electrolytic cell – the apparatus in which electrolysis is carried out: 206

galvanic cell – an electrochemical cell that transforms chemical energy to electric energy: 206

Faraday's Law of Electrolysis – whenever there is a change in the magnetic flux linkage with a conductor, an e.m.f. is induced; the induced e.m.f. in any closed circuit is equal to the rate of change of the magnetic flux through the circuit: 216

electroplating – the process whereby one metal is coated with another metal to give it more desirable or useful properties: 219

Checklist for Self-evaluation

Theme 3 Topic 14

EVALUATION GUIDE: Student should be able to:

	Criteria	4	3	2	1	Comments
14.1	define electrolysis					
14.2	set up the electrolytic and electrochemical cells					
14.3	carry out the electrolysis of acidified water					
14.4	explain Faraday's first and second laws of electrolysis					
14.5	solve problems involving Faraday's laws of electrolysis					
14.6	state the uses of electrolysis.					

Code for evaluation:			
4 – Very well	3 – Well	2 – Fairly well	1 – Not well at all

THEME 4

Chemistry of life

Topic 15: Hydrocarbons

Performance objectives

15.1 explain the structure of carbon and its valency
15.2 define hydrocarbons
15.3 give examples of hydrocarbons and their structure
15.4 explain isomerism and give examples
15.5 explain homologous series as it relates to the physical and chemical properties of hydrocarbons
15.6 distinguish between aliphatic and aromatic hydrocarbons.

Introduction

Organic chemistry is a specific field of study within the subject of chemistry. It involves the structure, properties, composition, reactions and preparations of carbon-based chemical compounds. These compounds may also contain other elements, such as hydrogen, nitrogen, oxygen, halogens, phosphorus and sulphur. We now know that the chemistry is the same in all molecules, irrespective of their source. Although the vitalistic theory is no longer accepted, we still study organic chemistry as a separate section of chemistry.

Activity 15.1: Test your knowledge of hydrocarbons

(SB p. 228)

Resources

Student's Book, workbook

Guidelines

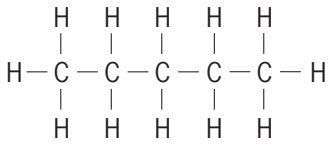
This activity can be used as a class test.

Discuss the structure and valency of carbons.

Students should be able to answer questions on carbons.

Answers

1. Carbon and hydrogen
2.
   - Carbon can make four covalent bonds
   - Carbon atoms can join each other to form long chains. Atoms of other elements can then attach to the chain.
   - The carbon atoms in a chain can be linked by single, double or triple covalent bonds.
   - Carbon atoms can also arrange themselves in rings.
3. a) Group IV, Period 2
   b) Valence electrons: 6
   c) Valency: 4
   d) $1s^2\ 2s^2\ 2p^2$
4. a) Pentane structural formula:



   b) Pentane condensed structural formula: $CH_3CH_2CH_2CH_2CH_3$ or $CH_3(CH_2)_3CH_3$
   c) Pentane skeletal formula:

Activity 15.2: Name organic molecules

(SB p. 214)

Resources

Student's Book, workbook

Guidelines

This activity can be used as a class test.

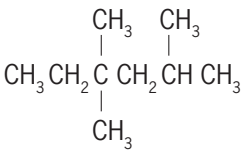
Discuss the worked examples on p. 235 of the SB.

Students should be able to use the IUPAC system to name organic molecules.

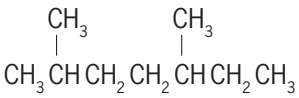
Answers

1. i) a) and c) are identical

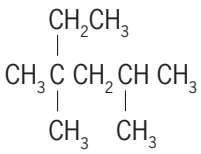
   ii)



2,4,4-trimethylhexane



2,5-dimethylheptane



2,4,4-trimethylhexane

Activity 15.3: Build molecular molecules

(SB p. 236)

Resources

Atomic model kits, or marbles and prestik, or jelly tots and toothpicks

Guidelines

This activity can be used as a class test.

Illustrate with models, the Stereo-chemistry of simple hydrocarbons.

Students should be able to draw and name simple hydrocarbons.

Answers

The results of the activity may differ from student to student.

Activity 15.4: Test yourself

(SB p. 236)

Resources

Student's Book, workbook

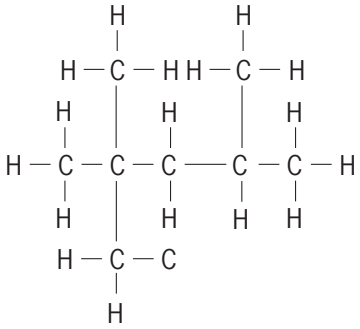
Guidelines

This activity can be used as a class test.

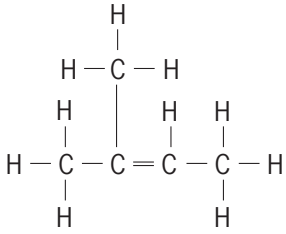
Students should be able to answer questions.

Answers

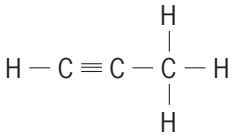
1. a) heptane
   b) 3-methyl-but-1-ene
   c) decane
   d) 2,3-dimethylpentane
   e) ethyne
   f) hex-3-yne
2. a)



   b)



   c)



Activity 15.5: Test your knowledge of hydrocarbons (SB p. 241)

Resources

Student's Book, workbook

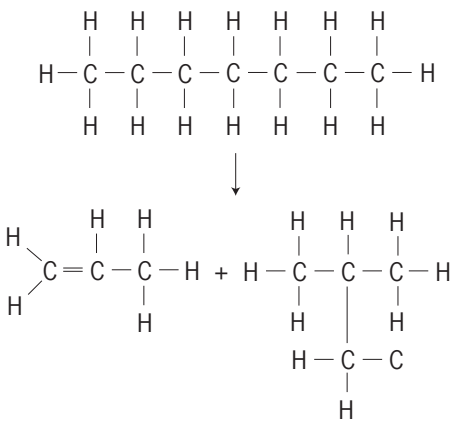
Guidelines

This activity can be used as a class test.

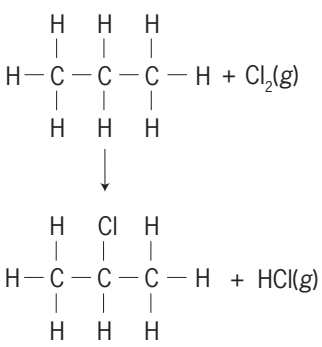
Students should be able to answer questions on hydrocarbons.

Answers

1. a) Oxygen
   b) $2C_2H_2 + 5O_2 \rightarrow 4CO_2 + 2H_2O$
   c) The reaction is very exothermic and the flame is very hot.
2. $2C_4H_{10} + 13O_2 \rightarrow 8CO_2 + 10H_2O$
3. a) and b)

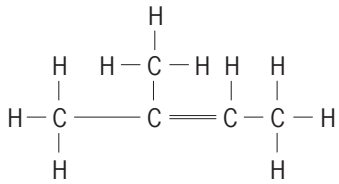


4. a)



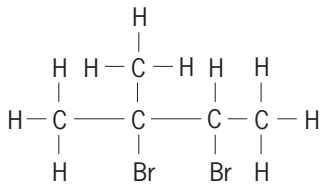
Chlorine can substitute any of the hydrogens in the propane molecule. Different products are possible.

   b) substitution
   c) nothing
5. a) alkenes
   b) C_nH_{2n}
   c)

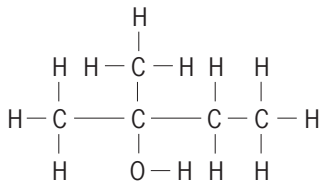


2-methylbut-2-ene

   d) i) The brown bromine solution discolours.
      ii) Halogenation (addition reaction)
      iii) 1,2-dibromo-2-methylbutane



   e) i) Water must be in excess, a small amount of strong acid (HCl, H_2SO_4 or H_3PO_4) to act as catalyst
      ii) 2-methylbutan-2-ol



Experiment 15.1: Investigating saturated and unsaturated molecules (SB p. 242)

Resources

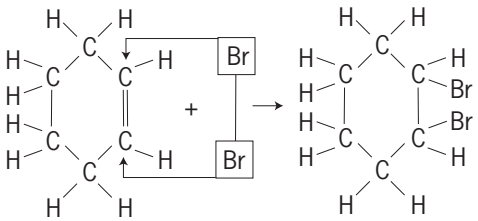
Student's Book, workbook, Cyclohexane, cyclohexene, test tubes with stoppers in test tube stand, bromine water, potassium manganate(VII) solution, calcium carbide (CaC_2) pellets, water, conical flask with stopper, dropping funnel, delivery tube

Guidelines

Guide the students to perform an experiment to distinguish between saturated and unsaturated hydrocarbons.

Students should be able to define homologous series, isomerism, saturated and unsaturated hydrocarbons.

Answers

1. Add a small amount of bromine solution to the unknown hydrocarbon. The alkene will discolour the bromine solution immediately from orange-brown to colourless.
2. $CaC_2(s) + 2H_2O(l) \rightarrow C_2H_2(g) + Ca(OH)_2(aq)$
3. There should be no difference in the reactivity of cyclohexene and ethyne. Both are unsaturated.
4. 

Activity 15.6: Answer questions about saturated and unsaturated compounds

(SB p. 243)

Resources

Student's Book, workbook

Guidelines

This activity can be used as a class test.

Guide a class discussion on the differences of saturated and unsaturated compounds.

Students should be able to define saturated and unsaturated hydrocarbons.

Answers

1. Saturated compounds: methane, octane, 2,3-dimethylnonane
   Unsaturated compounds: propene, butyne, buta-1,3-diene

Activity 15.7: Test yourself

(SB p. 245)

Resources

Student's Book, workbook

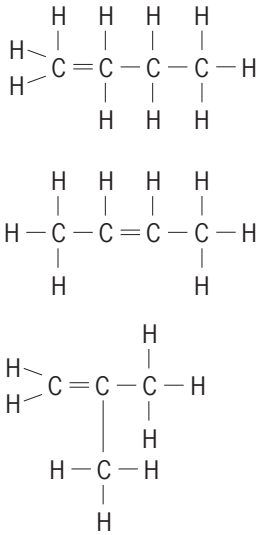
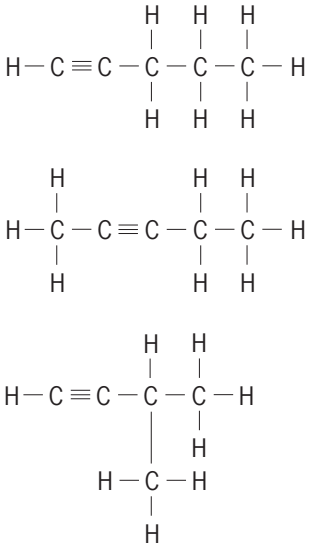
Guidelines

This activity can be used as a class test.

Guide a class discussion on the differences of saturated and unsaturated compounds.

Students should be able to draw and write the structure and name isomers of given hydrocarbons.

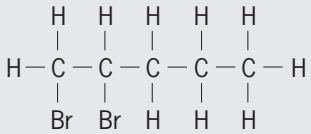
Answers

1. 
2. 

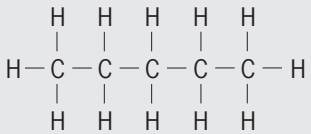
Revision questions: Answers

(SB p. 247)

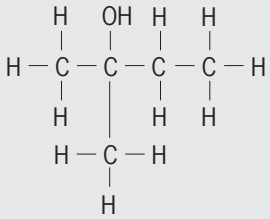
1. a) A b) B c) C d) D e) C
2. a) hydrocarbons
   b) functional group
   c) homologous series
   d) unsaturated
   e) alkynes
   f) C_nH_{2n+2}
   g) (structural) isomers
   h) IUPAC system
   i) alkanes
   j) hydration
3. a) General formula: All members of a homologous series have the same general formula. The general formula for the alkanes is C_nH_{2n+2} where *n* is the number of carbon atoms in the molecule.
   b) Homologous series: A family of organic molecules forms a homologous series. All the molecules in a homologous series have the same functional group, but different lengths of carbon chains. Alkanes, alkenes and alkynes are examples of organic homologous series.
   c) Functional group: The distinctive group of atoms attached to a carbon chain that all the members of the homologous series have in common is called the functional group. Functional groups are responsible for the characteristic chemical reactions of a homologous series. The functional group of the alkenes is a double bond >C=C<.
4. a) alkenes
   b) i)



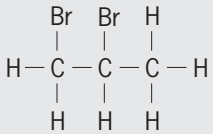
   ii) addition reaction: halogenation
   iii) the orange-brown colour of the bromine solution is discoloured
   c) i)



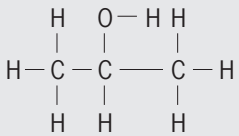
   ii) addition reaction: hydrogenation
   iii) Vegetable oil is hydrogenated to increase its saturation level and decrease the melting point to form soft solids (margarine) at room temperature.
   d) i) addition reaction: hydration
   ii)



   e) C: 2-methyl-but-1-ene
   D: 2-methyl-but-2-ene
   E: 3-methyl-but-1-ene
5. a) C_nH_{2n-2}
   b) i) addition reaction: halogenation
   ii) 1,2-dibromopropane



   c) i) propan-2-ol



   ii) Strong acid such as HCl, H_2SO_4 or H_3PO_4
6. a) Large hydrocarbons are cracked to form smaller alkenes that are more useful.
   b) C_8H_{18}
   c) $CH_3(CH_2)_6CH_3 \rightarrow CH_3CH_2CH_2CH_3 + CH_3CH{=}CHCH_3$

d) Butane and but-2-ene (or but-1-ene)
e) Thermal cracking happens at high temperature without a catalyst, catalytic cracking happens at lower temperature with a suitable catalyst.

How are you doing?

Take this opportunity to ask learners if there is anything that they do not understand. You can check their understanding by asking them some questions about the information covered in the unit. Explain anything that learners do not understand.

Key words

organic chemistry – a chemistry subdiscipline involving the scientific study of the structure, properties, and reactions of organic compounds and organic materials: 225

hydrocarbons – compounds that contain only carbon and hydrogen atoms: 226

homologous series – a family of organic molecules in which all the molecules have the same functional group, but different lengths of carbon chains: 228

saturated compound –organic molecule that contains only single bonds; no further atoms can be added to the molecule: 240

unsaturated compound – organic molecule that contains double or triple bonds; atoms can be added to the molecule in a chemical reaction: 238

isomers – molecules with the same molecular formula, but different structures: 245

Checklist for Self-evaluation

Theme 4 Topic 15

EVALUATION GUIDE: Student should be able to:

	Criteria	4	3	2	1	Comments
15.1	define: a) homologous series b) isomerism c) saturated and unsaturated hydrocarbons					
15.2	differentiate between aliphatic and aromatic hydrocarbons					
15.3	draw and name simple hydrocarbons					
15.4	write the structure and name isomers of given hydrocarbon.					

Code for evaluation:			
4 – Very well	3 – Well	2 – Fairly well	1 – Not well at all

Topic 16: Alkanols

Performance objectives

16.1 relate the structure of alkanes to that of alkanols
16.2 identify -OH as the functional group in alkanols
16.3 explain the increase in boiling point of alkanols compared with hydrocarbons in terms of hydrogen bonding
16.4 explain the polar nature of alcohols and its effect on the solubility of substances
16.5 determine solubility of common materials in water and alcohols
16.6 identify di-, tri- and polyhydroxy compounds by their structures and name them appropriately.

Introduction

Alkanols are organic molecules that contain carbon, hydrogen and oxygen atoms. The alkanols are a type of alcohol. Alkanols and alcohols contain the hydroxyl (-OH) functional group. In alkanols, the hydroxyl group is attached to a hydrocarbon chain.

Activity 16.1: Test yourself (SB p. 252)

Resources

Student's Book, workbook

Guidelines

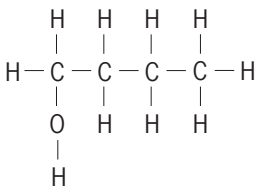
This activity can be used as a class test.

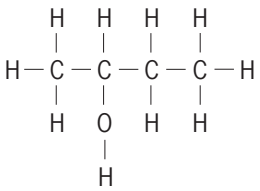
Guide a class discussion on the structures of alkanes and alkanols.

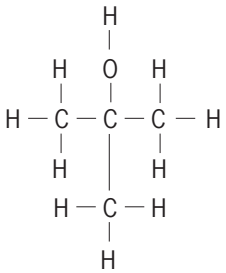
Students should be able to relate the structure of alkanes to that of alkanols.

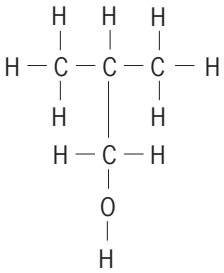
Answers

1. a)









b) butan-1-ol, butan-2-ol, 2-methyl-propan-2-ol, 2-methylpropan-1-ol

c) butan-1-ol: primary
butan-2-ol: secondary
2-methyl-propan-2-ol: tertiary
2-methylpropan-1-ol: primary

Activity 16.2: Test yourself (SB p. 253)

Resources

Student's Book, workbook

Guidelines

This activity can be used as a class test.

Guide a class discussion on how the increase in boiling point of alkanols compare with hydrocarbons in terms of hydrogen bonding.

Students should be able to explain this effect of boiling point on alkanols.

Answers

1. a) Intermolecular forces between alkane molecules: van der Waals forces, Intermolecular forces between alkanol molecules: hydrogen bonds.
   b) The boiling points of the alkanes increase with molecular mass. As the molecules become larger, greater parts of the molecules are closer to the adjacent molecules to form momentary dipoles. More energy is needed to free larger molecules from their neighbours, so the boiling point increases.
   c) Like dissolves like. This means that only molecules with intermolecular forces of similar strength will dissolve in each other. Both water and alkanols have polar molecules. The forces between water molecules are hydrogen bonds, and so are the forces between alkanol molecules. The hydrogen bonds between water molecules and between alkanol molecules can be replaced easily by hydrogen bonds between water and alkanol molecules. Alkanes are non-polar molecules with weak van der Waals forces between them. Van der Waals forces and hydrogen bonds have different order strengths and cannot replace each other. Alkanes do not mix with water.
   d) The O atom on one hydroxyl group forms a hydrogen bond with an H atom on an adjacent hydroxyl group. When there are more hydroxyl groups linked to a carbon chain, there are more places that can form hydrogen bonds. Hydrogen bonds have strong intermolecular forces and more hydrogen bonds need more energy to break the molecules apart to reach boiling point.

Experiment 16.1: Investigating some properties of alkanols (SB p. 256)

Resources

Student's Book, workbook, a selection of alkanols (for example, methanol, ethanol, ethan-1,2-diol, propan- 1,2,3-diol, phenol, and so on), test tubes with stoppers, test tube stand, potassium manganate(VII) solution, water, dilute sulphuric acid, sodium metal

Guidelines

Perform experiments to show oxidation of an alkanol with $KMnO_4$

Students should be able to observe experiments to observe the properties of alkanols, and observe and record their findings in their workbook.

Observations and conclusions

Part 1

Methanol and ethanol mix with water.

Higher alkanols mix slightly with water and the longer the alkyl chain, the less soluble the alkanol becomes.

Part 2

For primary and secondary alkanols you should see a brownish colour develop as the purple $KMnO_4$ colour disappears. Be careful not add too much $KMnO_4$. In an excess $KMnO_4$ the purple colour will persist even though you have a primary or secondary alkanol. With the tertiary alkanol you should see no colour change.

Part 3

Just as with water, the hydrogen atom of the hydroxyl group in alcohols and phenols can be displaced by sodium. The resulting alkoxides are strong bases.

Activity 16.3: Test yourself

(SB p. 258)

Resources

Student's Book, workbook

Guidelines

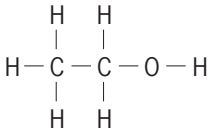
This activity can be used as a class test.

Show that methanol and ethanol are soluble in water.

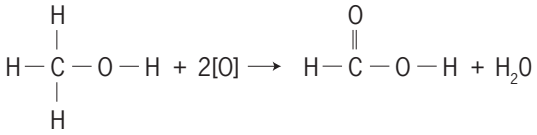
Students should be able to list the properties of alkanols.

Answers

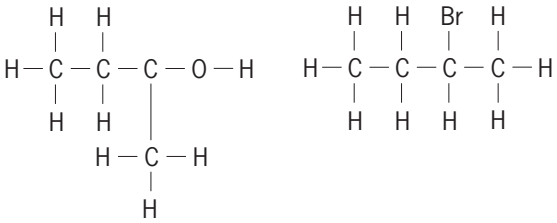
1. a) Alkanols
   b)



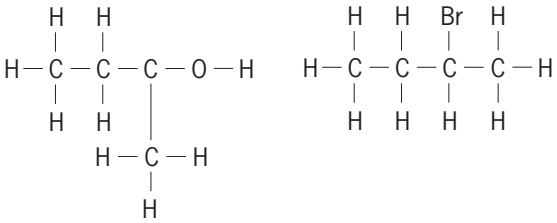
   c) $CH_3CH_2OH(l) + 3O_2(g) \rightarrow 2CO_2(g) + 3H_2O(g)$
   d) Sodium sinks to the bottom of the container and reacts with ethanol to liberate hydrogen gas. The ethanol turns pink, indicating that the solution becomes alkaline.
   e) Ethanol can undergo a dehydration reaction to form an alkene.
      i) Catalyst: heated ceramic surface
      ii) Heat ethanol with an excess of concentrated sulphuric acid or phosphoric acid
   f) i) Esterification reaction
      ii) Ethyl ethanoate + water
      iii) Concentrated sulphuric acid
2. a) Purple to colourless
   b)



   c) Methanoic acid + water
3. Alcohols and haloalkanes can be interconverted through substitution.
   a) Structure left is 1-methylpropan-1-ol

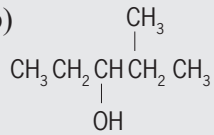
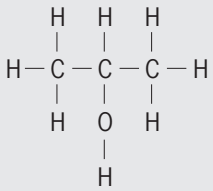
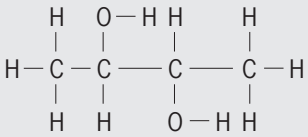
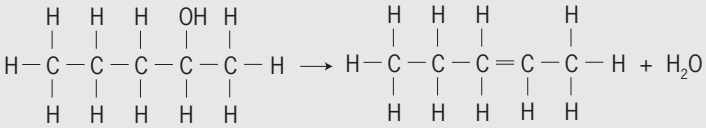
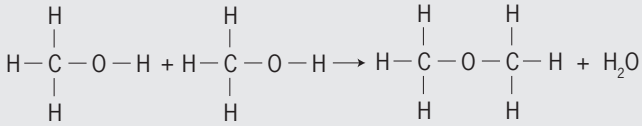


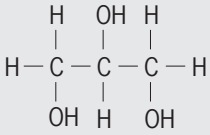
   b) $2NaBr + H_2SO_4 \rightarrow 2HBr$ (or HI) + Na_2SO_4
      1-methylpropan-1-ol + 2HBr (or HI) $\rightarrow$ product + H_2O
   c) structure right is 2-bromopropane (or 2-iodopropane)



Revision questions: Answers

(SB p. 261)

1. a) D b) A c) B d) B e) A
2. a) oxidation
   b) elimination
   c) elimination
   d) addition
   e) oxidation
   f) substitution
3. Ethanol is a polar molecule and will dissolve in water that also consists of polar molecules. (Like dissolves like.) Hexan-1-ol is a longer molecule with a hydroxyl group on the one end that will dissolve in water. The other side of the molecule consists of a hydrocarbon chain which is non-polar and that does not dissolve in water. So hexan-1-ol is only slightly soluble.
4. a) Methanol consists of polar molecules with strong hydrogen bonds between them. Ethane consists of non-polar molecules with weak van der Waals force. The boiling point of methanol is higher because more energy is needed to break the hydrogen bonds when the substance boils.
   b) Ethanol has a larger molecular mass and boiling points increase with molecular mass.
   c) Ethan-1,2-diol can form double the number of hydrogen bonds because it has 2 hydroxyl groups per molecule, which increases the boiling point.
5. a) $CH_3CH_2CH_2Br + NaOH \rightarrow CH_3CH_2CH_2OH + NaBr$
   b) Substitution reaction
   c) Propan-1-ol
6. a) ethanol
   b) butan-2-ol
   c) pentan-3-ol
   d) ethan-1,2-diol
   e) 2-methylpropan-2-ol
7. a) 
   b) 
8. a) 
   b) 
9. a) 
   
   b) pent-1-ene and pent-2-ene, and methoxymethane (dimethyl ether)
   c) i) Reactants: hydrogen bonding
      ii) Products: van der Waals forces
   d) Probably gaseous, because the intermolecular forces are weaker than in ethanol, though the molecular masses are somewhat higher
10. a) $C_3H_8O_3$

    b) propan-1,2,3-triol
    c) propane
    d) alkanols

e) i) The only intermolecular forces in propane are weak van der Waals forces whereas propanol has much stronger hydrogen bonds.

ii) As we move through the series, the number of hydroxyl groups, and hence hydrogen bonds per molecule, increases. Therefore the strength of the intermolecular forces increases.

f) The increase from propan-1-ol to propan-1,2-diol to glycerine: the number of hydroxyl groups and hydrogen bonds per molecule increases. The strength of the intermolecular forces also increases.

How are you doing?

Take this opportunity to ask learners if there is anything that they do not understand. You can check their understanding by asking them some questions about the information covered in the unit. Explain anything that learners do not understand.

Key words

positional isomers – one of two or more compounds with identical chemical composition, but having a different functional group location: 251

miscible – a liquid that mixes and dissolves in other liquids: 252

Checklist for Self-evaluation

Theme 4 Topic 16

EVALUATION GUIDE: Student should be able to:

	Criteria	4	3	2	1	Comments
16.1	explain the relationship between the structure of alkanes and alkanols					
16.2	list the properties of alkanols,					
16.3	describe the relationship between the -OH functional group and the properties of alkanols					
16.4	identify and name di-, tri- and polyhydroxy compound.					

Code for evaluation:			
4 – Very well	3 – Well	2 – Fairly well	1 – Not well at all

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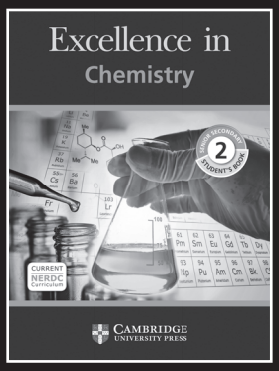

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